GASES

PROPERTIES:

The effects of attractive forces are minimal; we usually assume that the particles move independently**.**

- compressible
- indefinite shape and volume
- expand without limits
- diffuse into each other completely
- low density (usually measured in g/L)
- Measurements in **T** = temperature (K), **P** = pressure, **V** = volume, **n** = number of moles.

Ideal Gases:

- Exhibit specific mathematical relationships among the number of particles present, the temperature, the pressure and the volume.
- have negligible volume
- no attractive forces between molecules
- elastic collisions
- Real gases can behave as ideal gases if they avoid extremely low temperatures, very small volumes, and very high pressures (ie: under typical lab conditions).
- All real gases are observed to deviate from ideal behavior, particularly under conditions that are close to those resulting in condensation. Except at extremely high pressures that are not typically seen in the laboratory, deviations from ideal behavior are the result of intermolecular attractions among gas molecules. These forces are strongly distance dependent, so they are most significant during collisions.
- Observed deviations from ideal gas behavior can be explained through an understanding of the structure of atoms and molecules and their intermolecular interactions (EK.2.A.2).

Pressure:

- Pressure = force/area (F/A)
- Measured in Pascals (1 Pa = 1 N/m^2) or kilopascals (kPa)
- Common units of pressure used are atm, torr, mm Hg

1 atm = 760 torr = 760 mm Hg = 101,325 Pa = 1.01325 bar = 14.696 psi

- **Barometers** measure atmospheric pressure.
- **1 atm** is the atmospheric pressure which will support a column of mercury 760 mm high.
- Atmospheric pressure decreases with an increase in altitude (as does the density of air).
- **Manometers -** measure gas pressures. A glass tube is partially filled with mercury and has one arm open to the atmosphere and the other attached to the container of gas.

$\mathbf{P_{(gas)} = P_{(bar)} + \Delta h}$ P(bar) = barometric (atmospheric) pressure Δh = difference in height

$\mathbf{P = g\,d\,h}$ g = 9.8 m/s d = density h = height of column

Standard Temperature and Pressure (STP) = 0.00 ºC (273K) and 1.000 atm

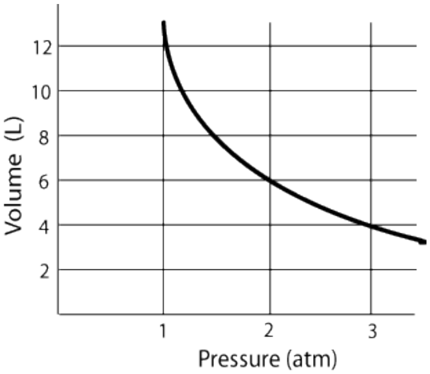
CALCULATIONS AND MATHEMATICAL RELATIONSHIPS

Combined Gas Law:

- **$P_1V_1 / T_1 = P_2V_2 / T_2$**

Boyle's Law (constant temperature):

- **$P_1V_1 = P_2V_2$**
- For a given amount of gas at a constant temperature, the volume of the gas varies inversely with its pressure.



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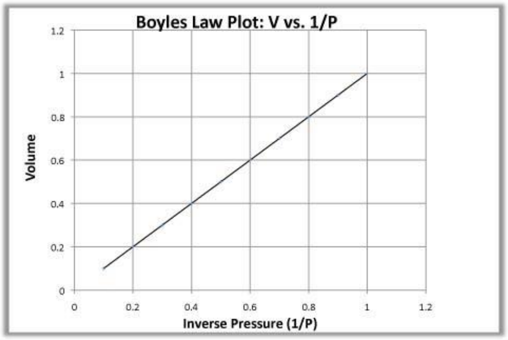

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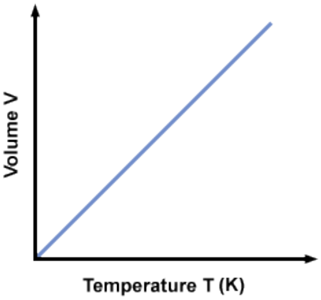
Charles' Law (constant pressure):

- **$V_1/T_1 = V_2/T_2$**
- The volume of a fixed amount of gas at a constant pressure is directly proportional to its Kelvin temperature.



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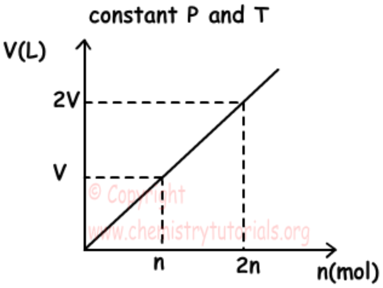
Avogadro's Law (mole-volume):

- At a fixed temperature and pressure, the volume of a gas is directly proportional to the number of moles of that gas.
- At STP **1 mol gas = 22.4 L**



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Ideal Gas Law:

PV = nRT R= universal gas constant = $0.08206\ \text{L atm mol}^{-1}\ \text{K}^{-1}$
$= 8.314\ \text{J mol}^{-1}\ \text{K}^{-1}$
$= 62.36\ \text{L torr mol}^{-1}\text{K}^{-1}$

(A) Molecular Mass:

- **n=m/M** **n** = number of moles **m** = mass (g) **M** = molar mass (g/mol)

(B) Density:

- **D = m/V**

Gases in Reaction Stoichiometry:

(A) Gay-Lussac's Law (combining volumes):

- At constant temperature and pressure, the volumes of gaseous reactants and products are in small whole number ratios.
- This relates the gaseous species in any reaction, even if some of the reactants or products are solids or liquids.
- The reaction does not have to take place at a particular temperature or pressure.

(B) Mass-Volume Relationships:

Dalton's Law of Partial Pressures:

- $P_{total} = P_A + P_B + P_C + \ldots\ldots.$
- The total pressure exerted by a mixture of gases is equal to the sum of the partial pressures exerted by the separate gases; the pressure exerted by each component is independent of the other components.

- $P_A = P_{total}\ (X_A)$ **where** X_A **= moles A/ total moles**

Collection of gas over water:

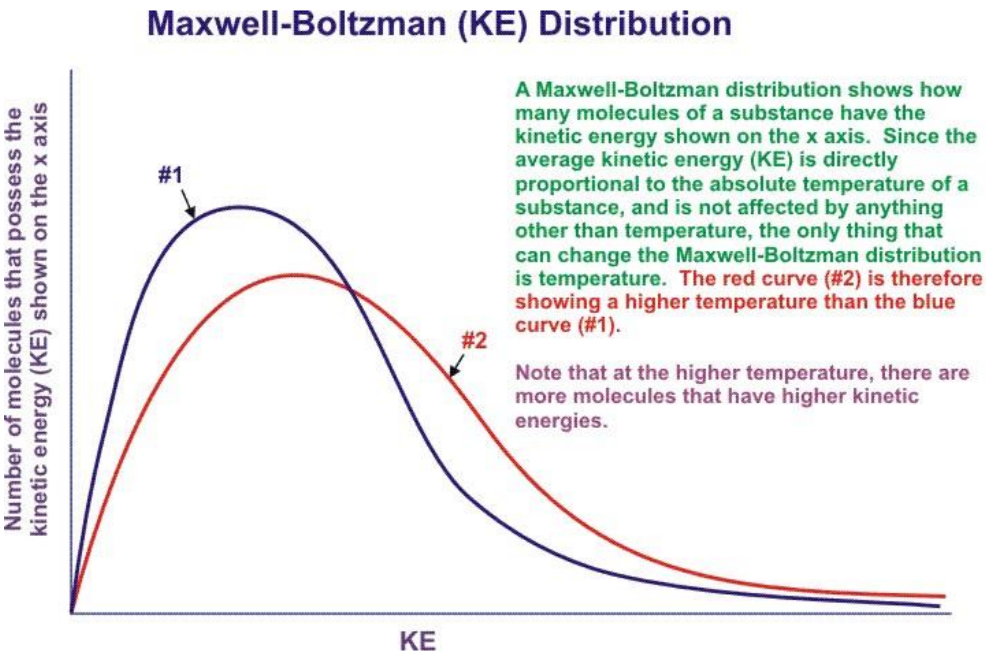
- Gases that are insoluble and that don't react with water can be collected over water.
- A gas collected over water is saturated with water vapor.
- $P_{gas} = P_{bar} - P_{H2O}$

Kinetic Molecular Theory:

(A) Kinetic Energy of a Gas Sample:

- $KE_{\text{per mole}} = 3/2\ RT$
- $KE_{\text{per molecule}} = 1/2\ mv^2$

(B) Maxwell-Boltzmann Distribution



Maxwell-Boltzman (KE) Distribution

A Maxwell-Boltzman distribution shows how many molecules of a substance have the kinetic energy shown on the x axis. Since the average kinetic energy (KE) is directly proportional to the absolute temperature of a substance, and is not affected by anything other than temperature, the only thing that can change the Maxwell-Boltzman distribution is temperature. The red curve (#2) is therefore showing a higher temperature than the blue curve (#1).

Note that at the higher temperature, there are more molecules that have higher kinetic energies.

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Graham's Law (Diffusion and Effusion):

- **Diffusion -** the intermingling of gases. Difficult to calculate since the molecules collide frequently.
- **Effusion -** the leaking out of a gas through a small hole (orifice).
- $\frac{\text{rate 1}}{\text{rate 2}} = \sqrt{\frac{M2}{M1}}$
- At a given temperature, the rates of effusion of gas molecules are inversely proportional to the square roots of their molar masses.