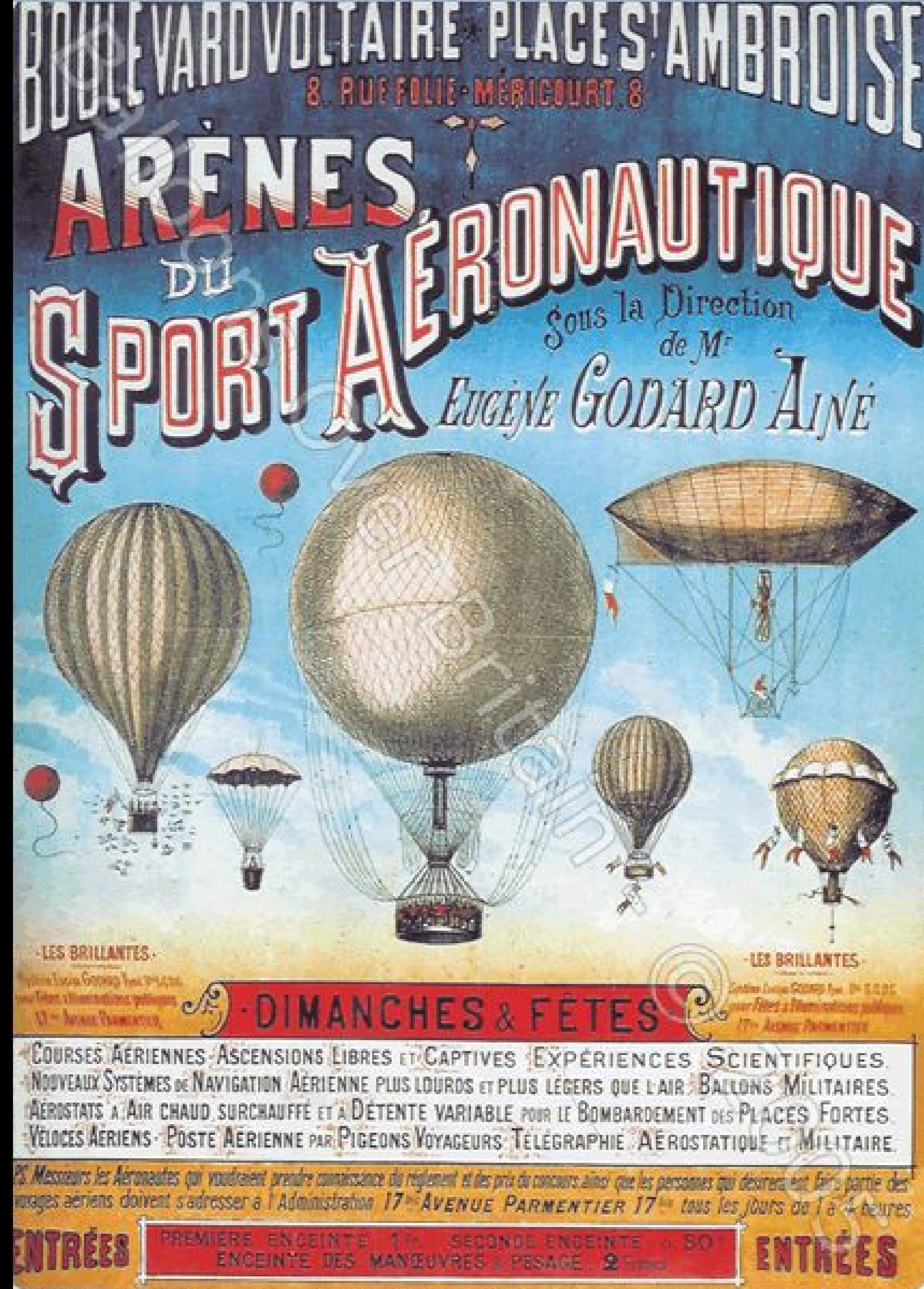
Unit 5: Gases and Gas Laws



Kinetic Molecular Theory

- Particles of matter are ALWAYS in motion
- Volume of individual particles is ≈ zero.
- Collisions of particles with container walls cause pressure exerted by gas.
- Particles exert no forces on each other.
- Average kinetic energy ∝ Kelvin temperature of a gas.

The Meaning of Temperature

$$(\mathrm{KE})_{\mathrm{avg}} = \frac{3}{2} RT$$

- Kelvin temperature is an index of the random motions of gas particles (higher *T* means greater motion.)

Kinetic Energy of Gas Particles

At the same conditions of temperature, all gases have the same average kinetic energy.

$$\boxed{KE = \frac{1}{2}mv^2}$$

m = *mass*

v = **velocity**

Measuring Pressure

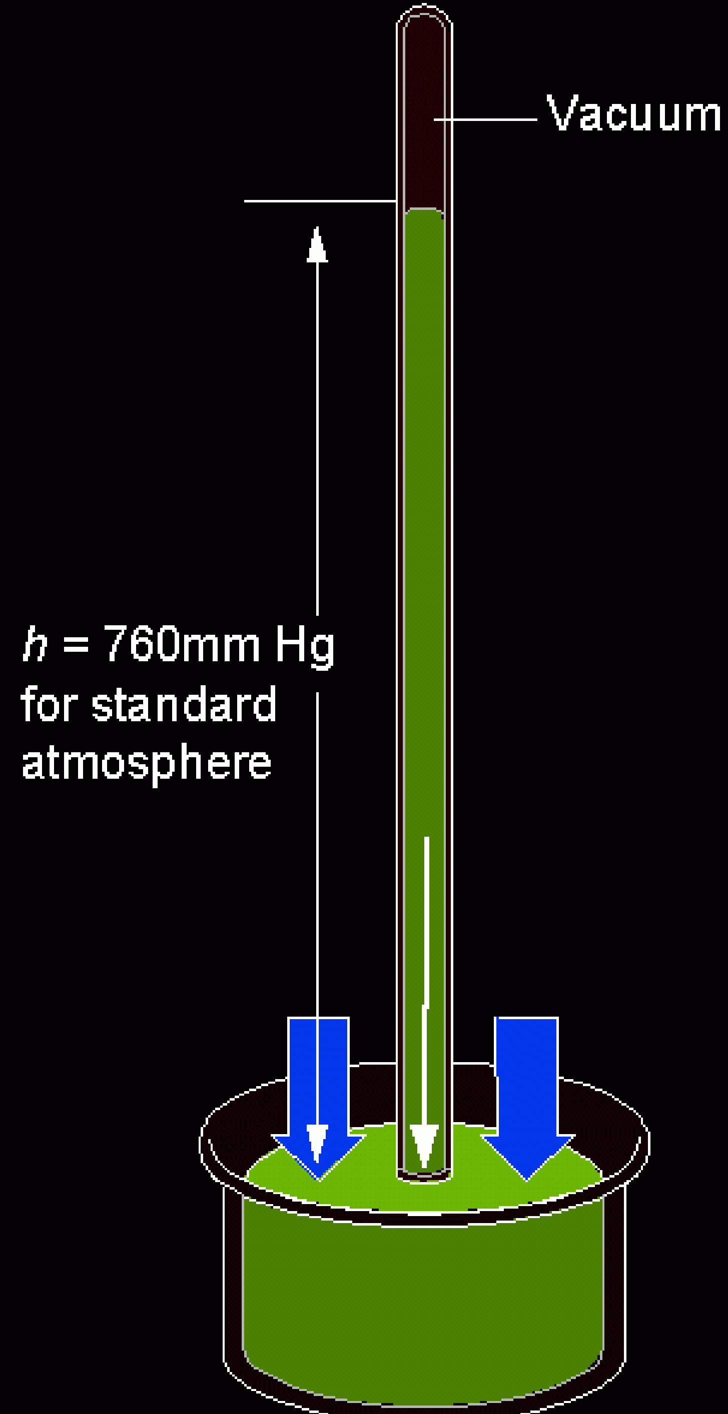
The first device for measuring atmospheric pressure was developed by Evangelista Torricelli during the 17th century.

The device was called a "barometer"

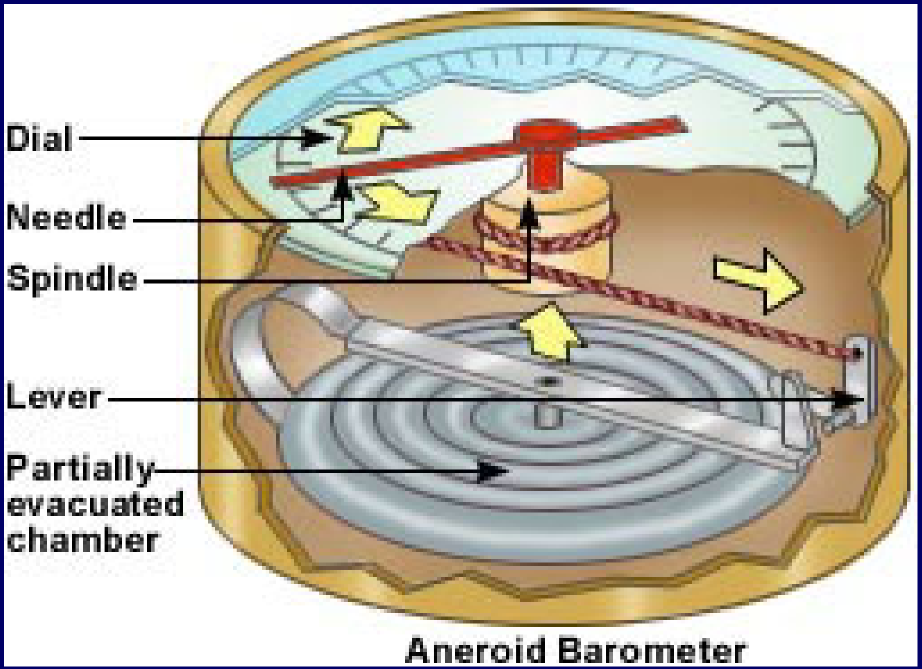
- Baro = weight
- Meter = measure

An Early Barometer

The normal pressure due to the atmosphere at sea level can support a column of mercury that is 760 mm high.



The Aneroid Barometer



Pressure

- Is caused by the collisions of molecules with the walls of a container
- is equal to force/unit area
- SI units = Newton/meter^2 = 1 Pascal (Pa)
- 1 standard atmosphere = 101.3 kPa
- 1 standard atmosphere = 1 atm =

  760 mm Hg = 760 torr

Units of Pressure

Unit	*Symbol*	*Definition/Relationship*
Pascal	Pa	SI pressure unit 1 Pa = 1 newton/meter^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

The Nature of Gases

- Gases expand to fill their containers
- Gases are fluid – they flow
- Gases have low density
  - 1/1000 the density of the equivalent liquid or solid
- Gases are compressible
- Gases effuse and diffuse

Standard Temperature and Pressure "STP"

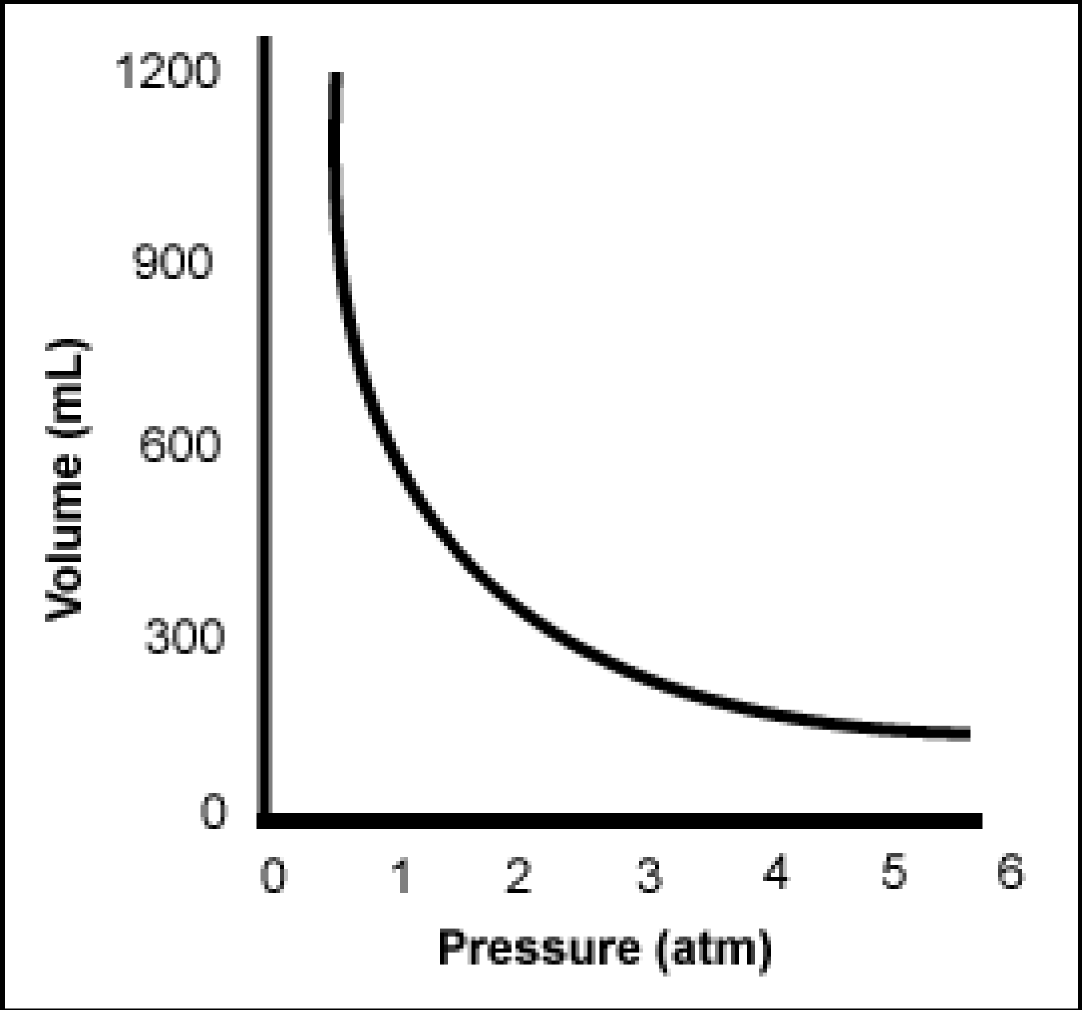
- P = 1 atmosphere, 760 torr, 101.3 kPa
- T = 0°C, 273 Kelvins
- The molar volume of an ideal gas is 22.42 liters at STP

Boyle's Law

Pressure is inversely proportional to volume when temperature is held constant.

$$P_1V_1 = P_2V_2$$

A Graph of Boyle's Law



Converting Celsius to Kelvin

Gas law problems involving temperature require that the temperature be in KELVINS!

Kelvins = °C + 273

°C = Kelvins - 273

Charles's Law

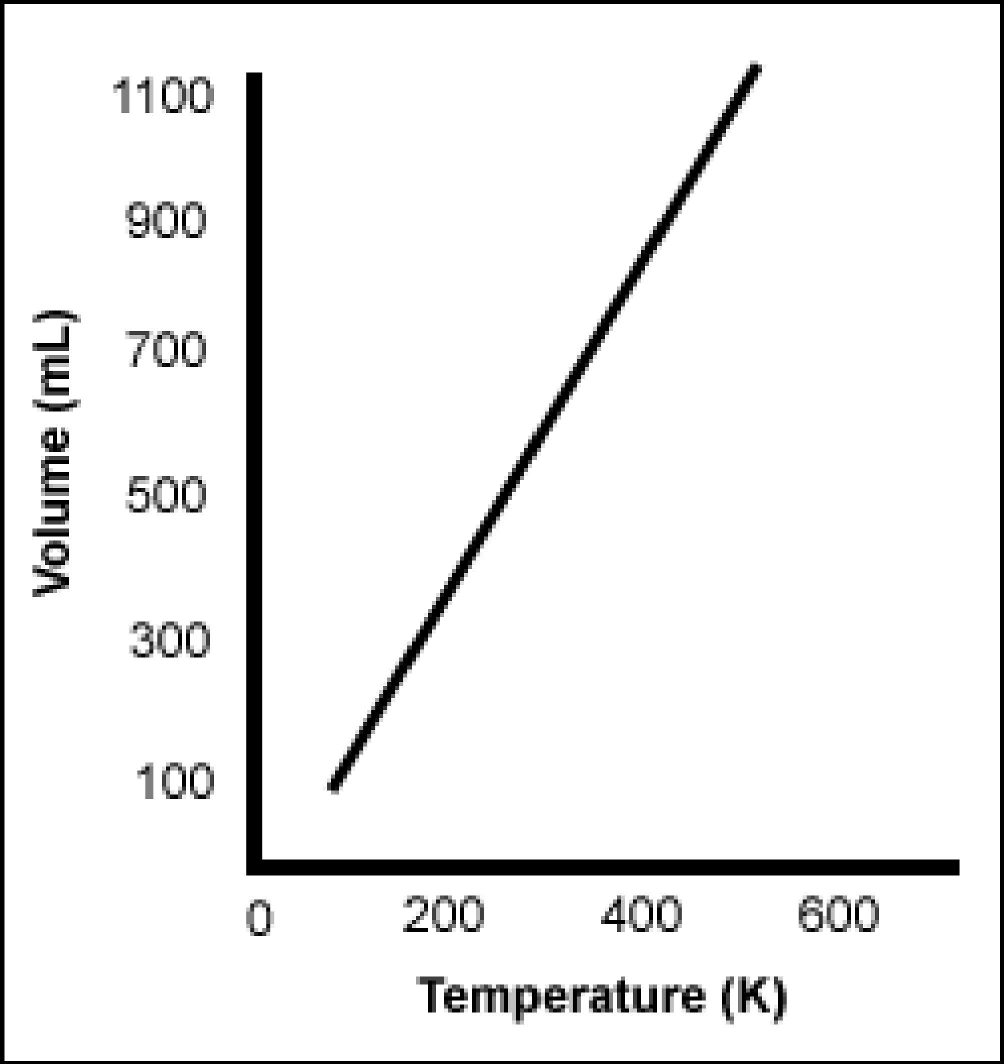
- The volume of a gas is directly proportional to temperature, and extrapolates to zero at zero Kelvin.

(*P* = constant)

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \qquad (P = \text{constant})$$

Temperature MUST be in KELVINS!

A Graph of Charles' Law



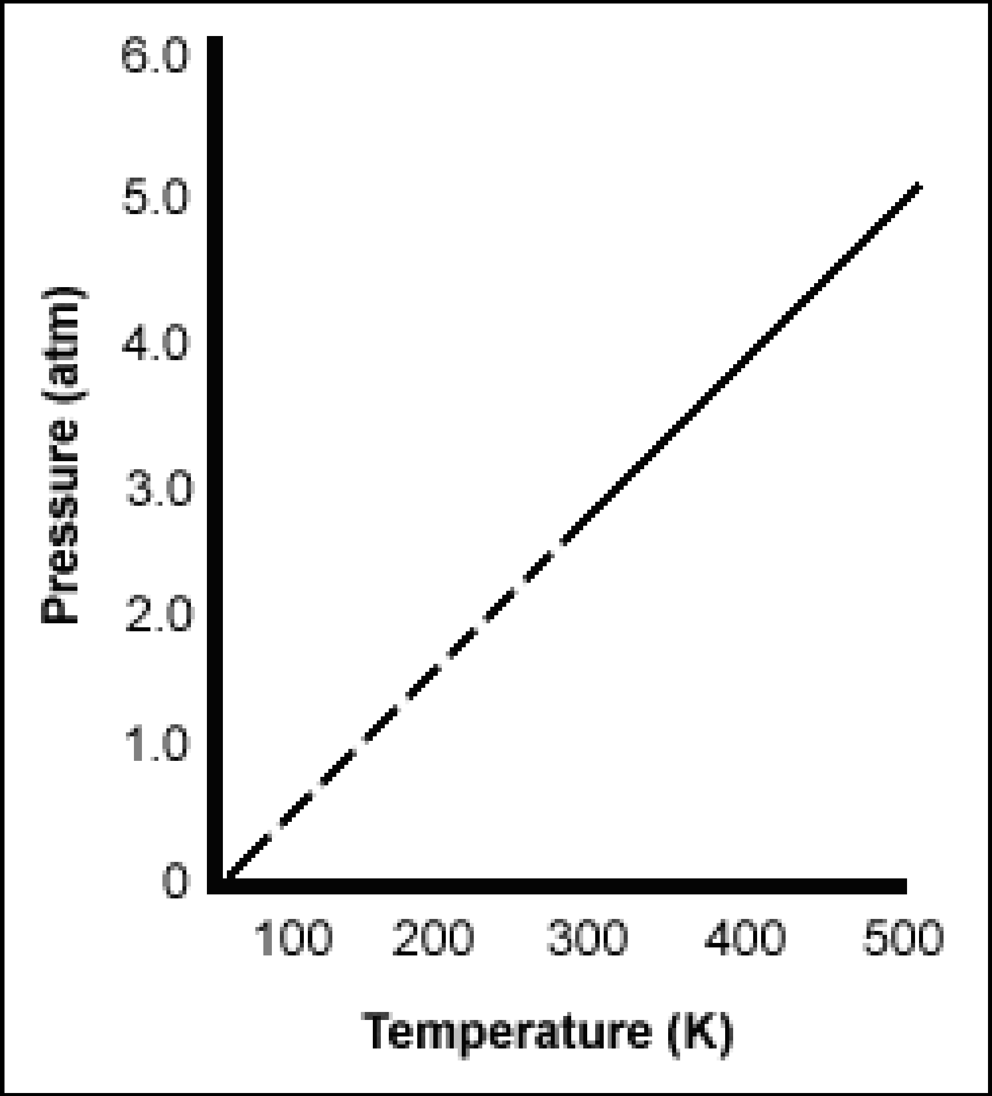
Gay Lussac's Law

The pressure and temperature of a gas are directly related, provided that the volume remains constant.

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

Temperature MUST be in KELVINS!

A Graph of Gay-Lussac's Law

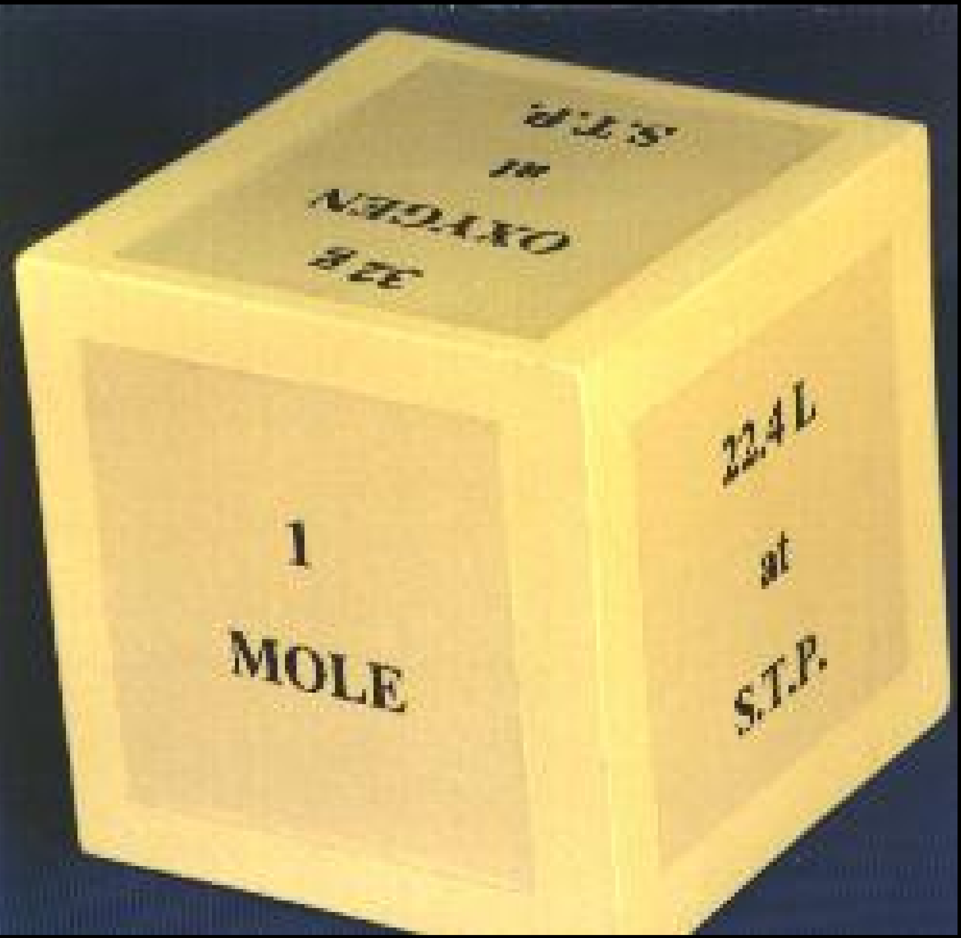


The Combined Gas Law

The combined gas law expresses the relationship between pressure, volume and temperature of a fixed amount of gas.

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

Boyle's law, Gay-Lussac's law, and Charles' law are all derived from this by holding a variable constant.



Standard Molar Volume

Equal volumes of all gases at the same temperature and pressure contain the same number of molecules.

– **Amedeo Avogadro**

Ideal Gas Law

$$PV = nRT$$

- P = pressure in atm
- V = volume in liters
- n = moles
- R = proportionality constant
  = 0.08206 L atm/ mol·K
- T = temperature in Kelvins

Holds closely at $P < 1$ atm

Gas Density

$$Density = \frac{mass}{volume} = \frac{\text{molar mass}}{\text{molar volume}}$$

… so at STP…

$$Density = \frac{\text{molar mass}}{22.4\text{ L}}$$

Density and the Ideal Gas Law

Combining the formula for density with the Ideal Gas law, substituting and rearranging algebraically:

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

M = Molar Mass

P = Pressure

R = Gas Constant

T = Temperature in Kelvins

Ideal Gases

Ideal gases are imaginary gases that perfectly fit all of the assumptions of the kinetic molecular theory.

- Gases consist of tiny particles that are far apart relative to their size.
- Collisions between gas particles and between particles and the walls of the container are elastic collisions
- No kinetic energy is lost in elastic collisions

Ideal Gases (continued)

- Gas particles are in constant, rapid motion. They therefore possess kinetic energy, the energy of motion
- There are no forces of attraction between gas particles
- The average kinetic energy of gas particles depends on temperature, not on the identity of the particle.

Real Gases Do Not Behave Ideally

Real gases **DO** experience inter-molecular attractions

Real gases **DO** have volume

Real gases **DO NOT** have elastic collisions

Deviations from Ideal Behavior

Likely to behave nearly ideally	*Likely not to behave ideally*
Gases at high temperature and low pressure	Gases at low temperature and high pressure
Small non-polar gas molecules	Large, polar gas molecules

Dalton's Law of Partial Pressures

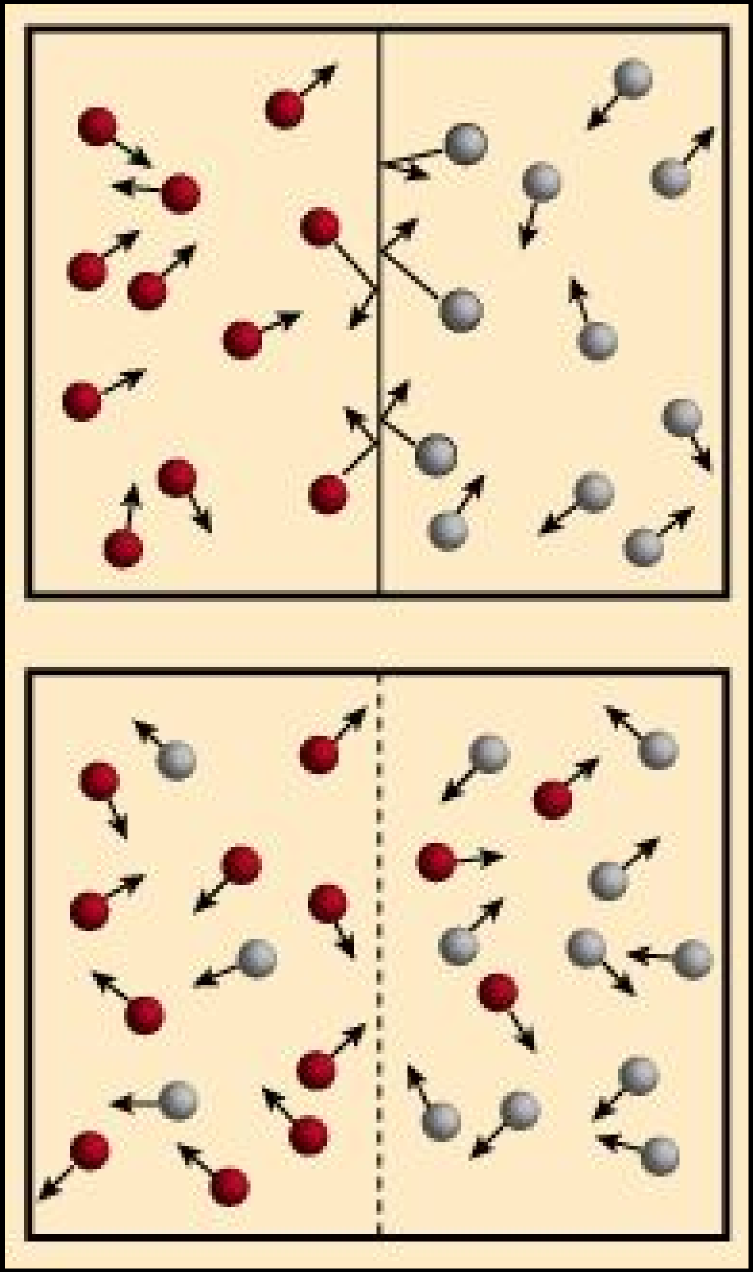
For a mixture of gases in a container,

$$P_{\text{Total}} = P_1 + P_2 + P_3 + \ldots$$

This is particularly useful in calculating the pressure of gases collected over water.

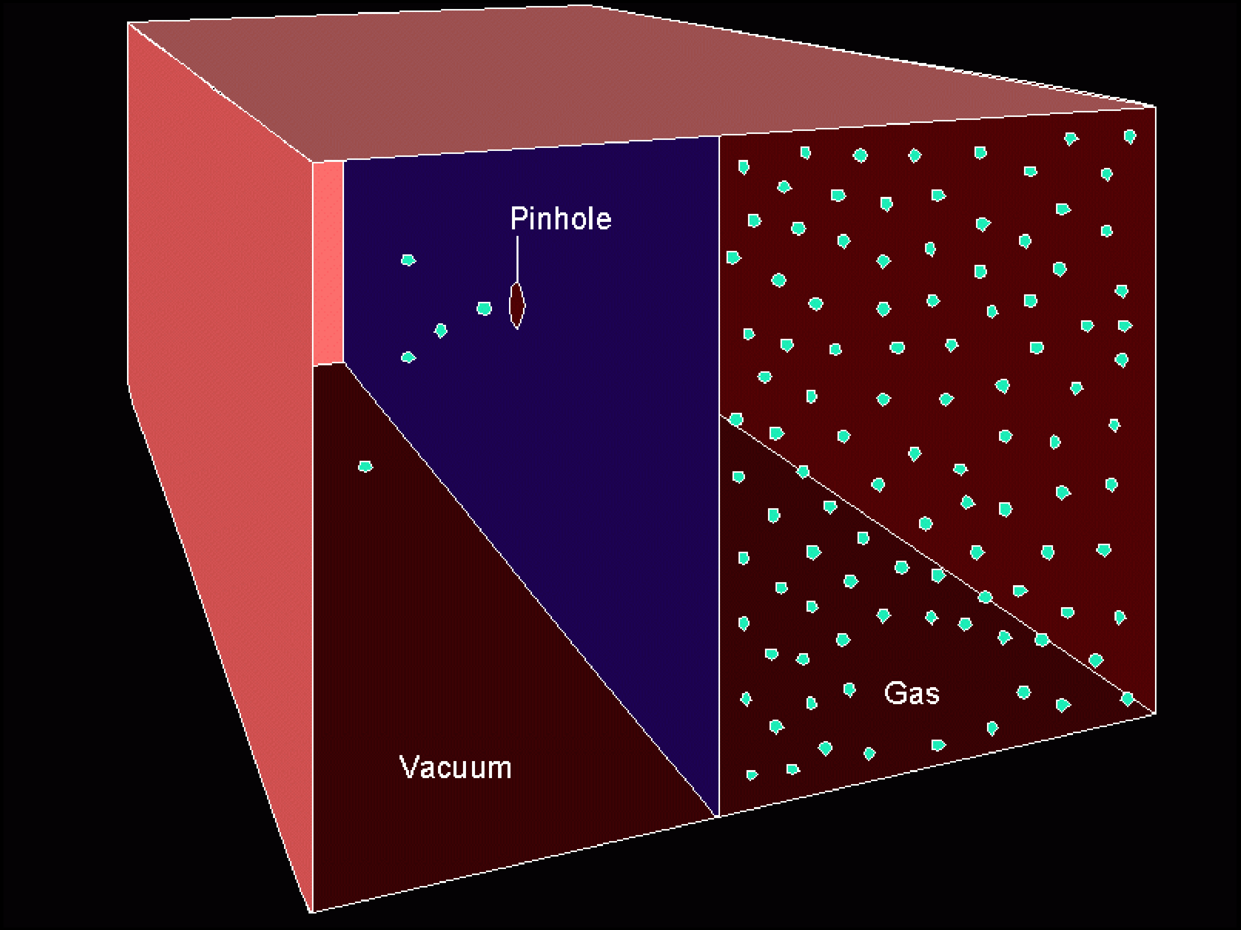
Diffusion

Diffusion: describes the mixing of gases. The rate of diffusion is the rate of gas mixing.



Effusion

Effusion: describes the passage of gas into an **evacuated** chamber.



Graham's Law
Rates of Effusion and Diffusion

Effusion:

$$\frac{\text{Rate of effusion for gas 1}}{\text{Rate of effusion for gas 2}} = \frac{\sqrt{M_2}}{\sqrt{M_1}}$$

Diffusion:

$$\frac{\text{Distance traveled by gas 1}}{\text{Distance traveled by gas 2}} = \frac{\sqrt{M_2}}{\sqrt{M_1}}$$