A Gas

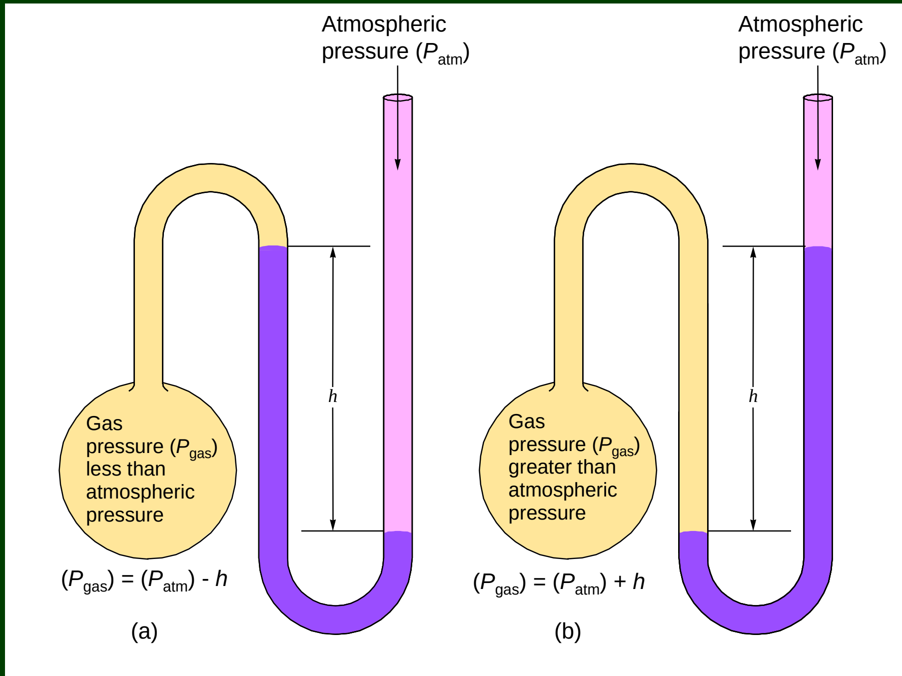
- Uniformly fills any container.
- Mixes completely with any other gas
- Exerts pressure on its surroundings.

Pressure

- is equal to force/unit area
- SI units = Newton/meter^2 = 1 Pascal (Pa)
- 1 standard atmosphere = 101,325 Pa
- 1 standard atmosphere = 1 atm = 760 mm Hg = 760 torr

A simple manometer – measures the pressure of a gas in a contained (in mm Hg = Torr)

a) Gas pressure = atmospheric pressure – h, b) Gas pressure = atmospheric pressure + h

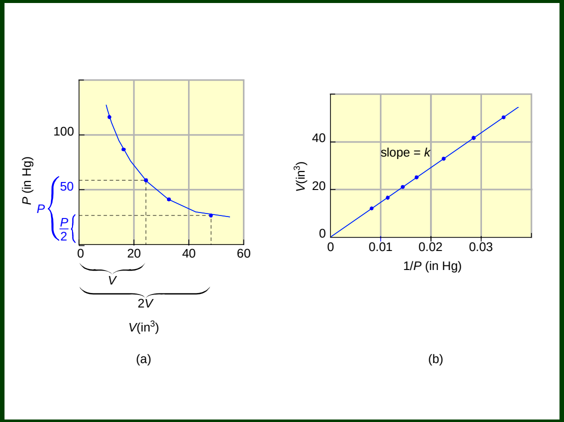


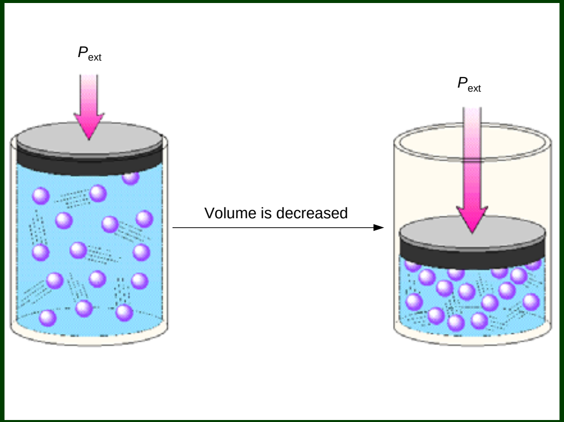
Boyle's Law*

Experiment

a) Volume doubles as the pressure is halved

b) A plot V versus 1/P gives a straight line and the slope gives the value of the contant k





Boyle's Law*

Generalization

Pressure × Volume = Constant   (T = constant)

$$P_1 V_1 = P_2 V_2 \quad (T = \text{constant})$$

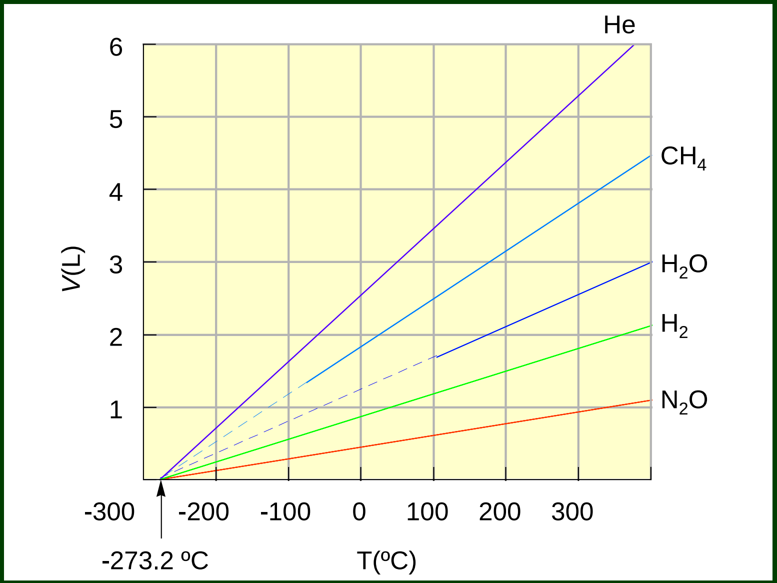
$$V \propto 1/P \quad (T = \text{constant})$$

(*Holds *precisely* only at very low pressures.)

A gas that strictly obeys Boyle's Law is called an ideal gas.

Charles's Law

A plot V versus T gives a straight line as shown for several gases. The solid lines represent experimental measurements. The dashed lines represent extrapolation of the data into regions where these gases would become liquids or solids



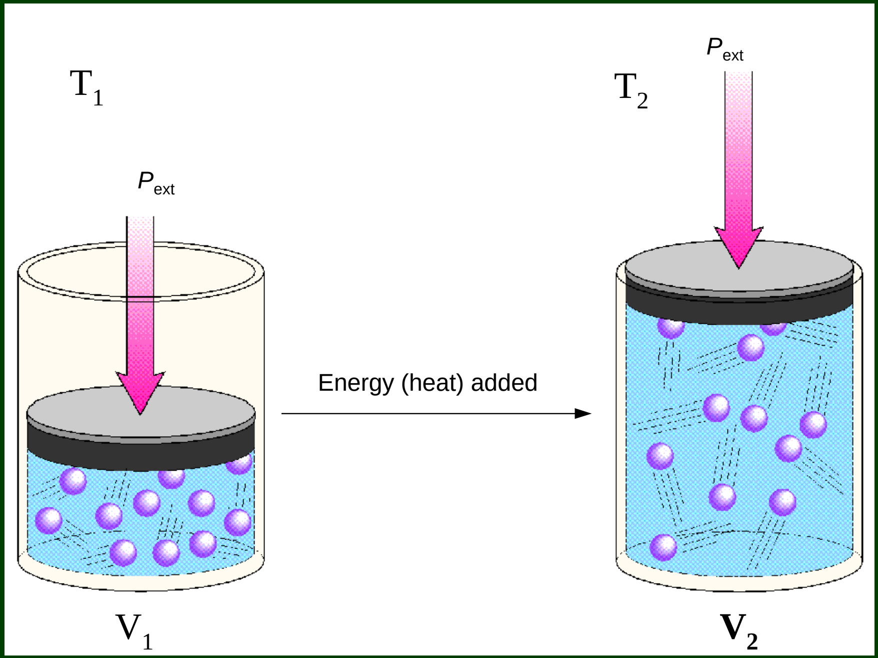
The volume of a gas is directly proportional to temperature, and extrapolates to zero at zero Kelvin.

$$V = bT \quad (P = \text{constant})$$

b = a proportionality constant

Charles's Law

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



Gas stoichiometry

Avogadro's Law

For a gas at constant temperature and pressure, the volume is directly proportional to the number of moles of gas (at low pressures).

$$V = an$$

a = proportionality constant

V = volume of the gas

n = number of moles of gas

Ideal Gas Law

- An equation of state for a gas.
- “state” is the condition of the gas at a given time.

$$PV = nRT$$

Ideal Gas Law

$$PV = nRT$$

R = proportionality constant

= 0.08206 L atm K^{-1} mol^{-1}

P = pressure in atm

V = volume in liters

n = moles

T = temperature in Kelvins

Holds closely at $P < 1$ atm

Standard Temperature and Pressure

"STP"

P = 1 atmosphere

T = 0°C

The molar volume of an ideal gas is 22.42 liters at STP

Dalton's Law of Partial Pressures

For a mixture of gases in a container,

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

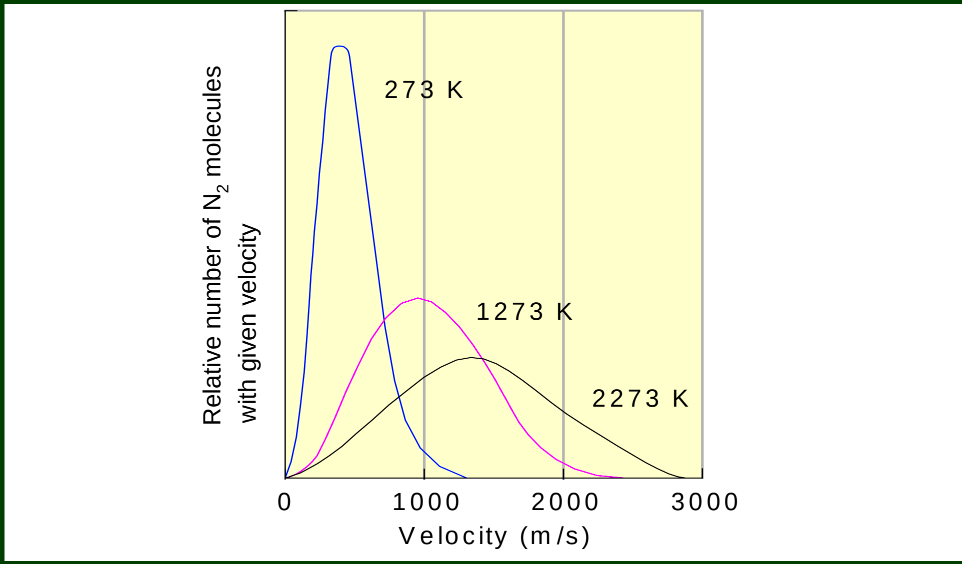
Kinetic Molecular Theory

1. Volume of individual particles is $\approx$ zero.
2. Collisions of particles with container walls cause pressure exerted by gas.
3. Particles exert no forces on each other.
4. Average kinetic energy $\propto$ Kelvin temperature of a gas.

The Meaning of Temperature

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

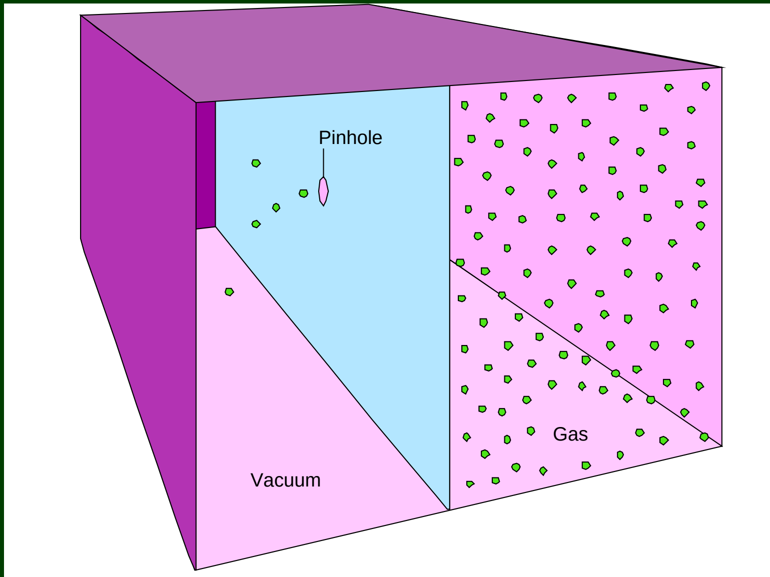
Kelvin temperature is an index of the random motions of gas particles (higher *T* means greater motion.)



A plot of the relative number of N_2 molecules that have a given velocity at three temperatures. Note that as the temperature increases both the average volocity (maximum on curves) and the spread of velocities increase

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

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



Effusion:

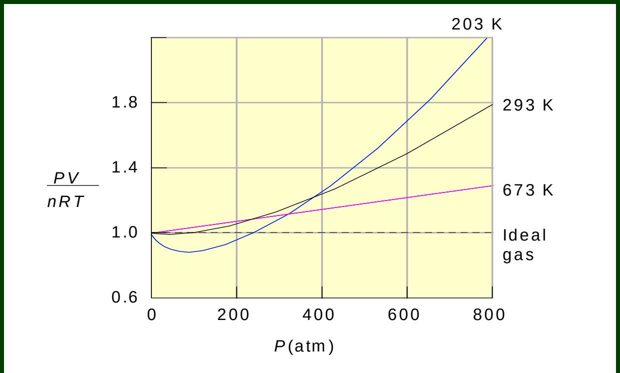
$$\frac{\text{Rate of effusion for gas 1}}{\text{Rate of effusion for gas 2}} = \frac{\sqrt{M_2}}{\sqrt{M_1}}$$

Diffusion:

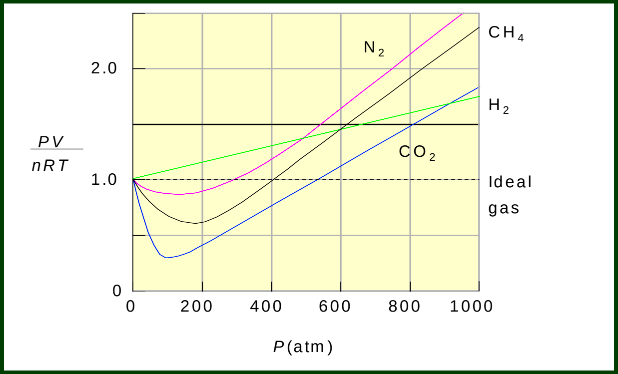
$$\frac{\text{Distance traveled by gas 1}}{\text{Distance traveled by gas 2}} = \frac{\sqrt{M_2}}{\sqrt{M_1}}$$

Real Gases

Plots PV/nRT versus P for nitrogen gas at three temperatures.



Plots PV/nRT versus P for several gases (at 200 K).



The ideal gas model works only for low pressures and high temperatures

Real Gases

Must correct ideal gas behavior when at high pressure (smaller volume) and low temperature (attractive forces become important).

Real Gases

Van der Waals' equation

$$[P_{\text{obs}} + a\,(n/V)^2] \times (V - nb) = nRT$$

↑ ↑

$\underbrace{\text{corrected pressure}}_{P_{\text{ideal}}}$ $\underbrace{\text{corrected volume}}_{V_{\text{ideal}}}$

Chemistry in atmosphere

$2NO_2(g) \rightarrow 2NO(g) + 2O(g)$
$2O\ (g) + 2O_2(g) \rightarrow 2O_3(g)$
$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$

$3O_2(g) \rightarrow 2O3(g)$