Gases Density & More

Gas Density

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

… so at STP…

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

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

Kinetic Energy of Gas Particles

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

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

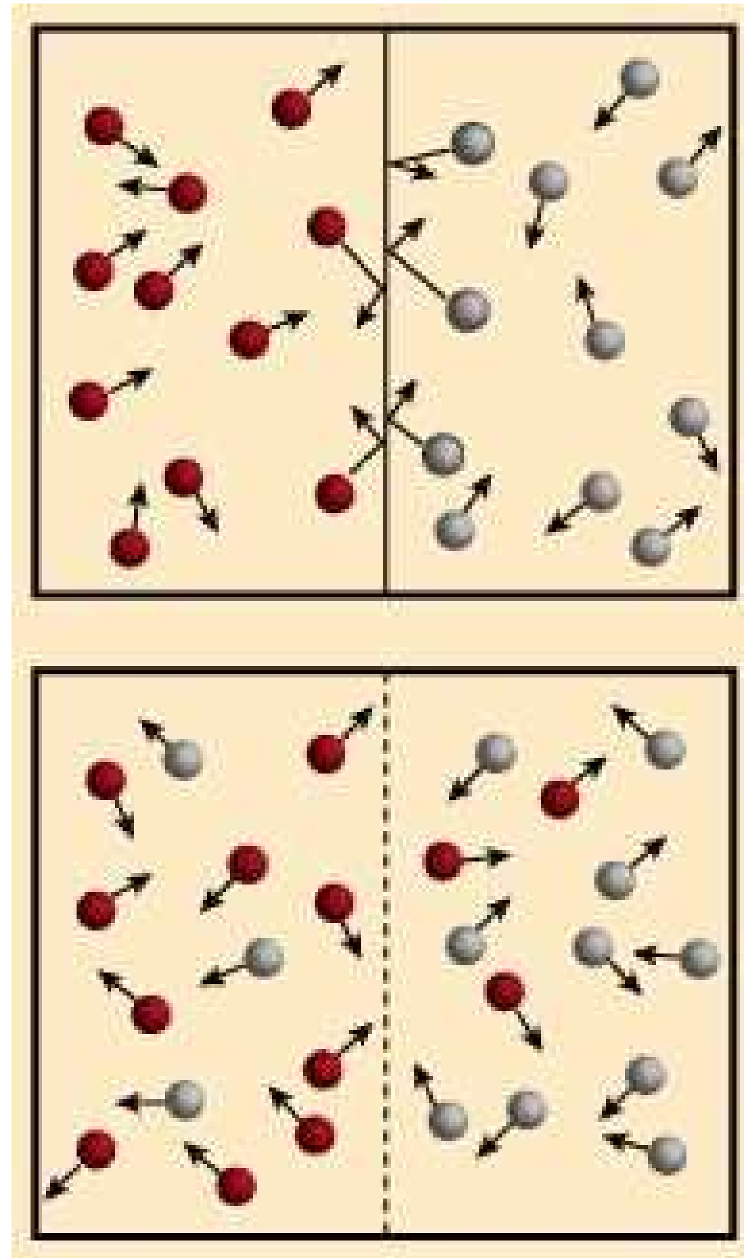
The Meaning of Temperature

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

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

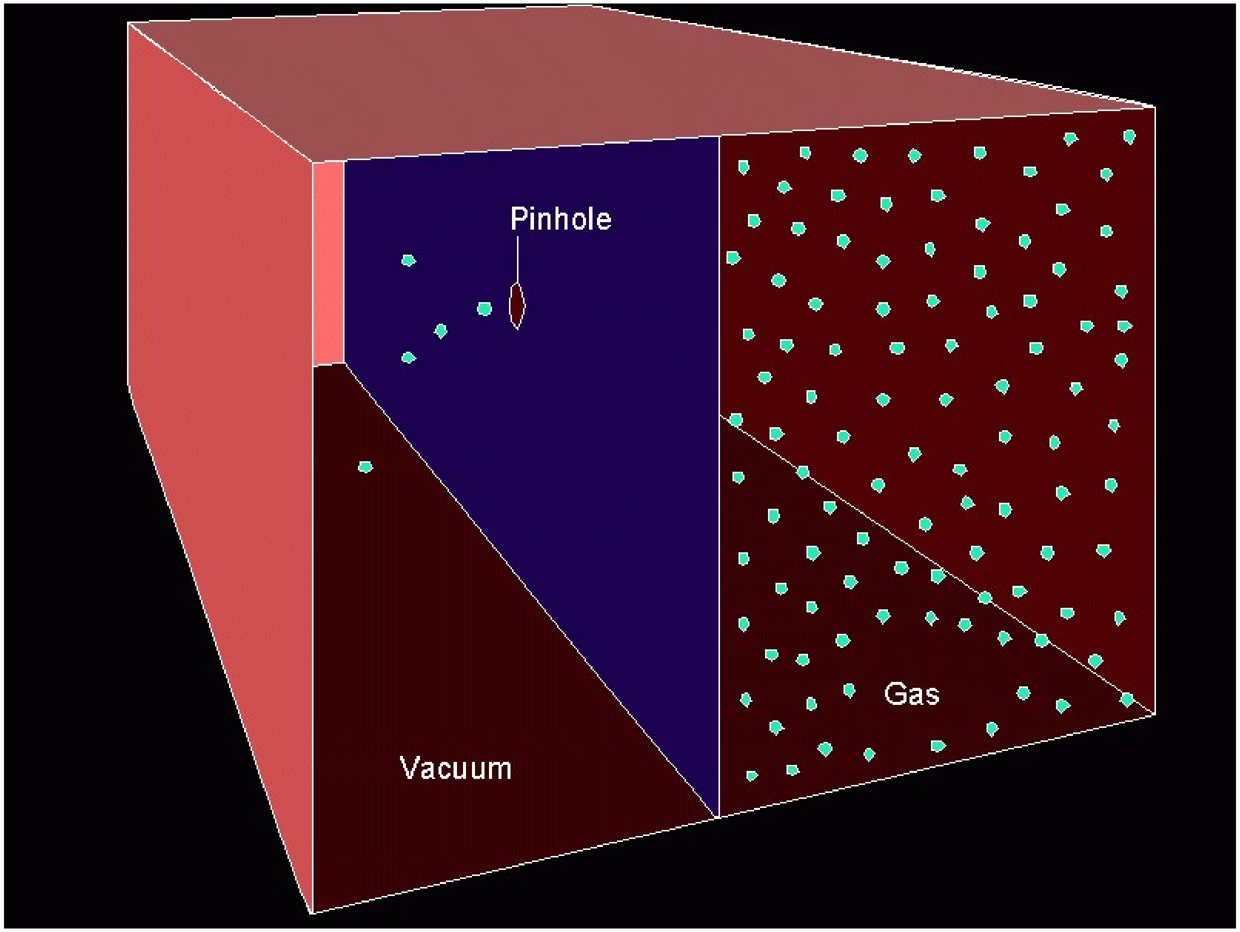
Diffusion

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:

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

Practice

- Under the same conditions of temperature and pressure, does hydrogen iodide or ammonia effuse faster? Calculate the relative rates at which they effuse.

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

- $\frac{\boldsymbol{Rate\ of\ NH_3}}{\boldsymbol{Rate\ of\ HI}} = \frac{\sqrt{\boldsymbol{HI}}}{\sqrt{\boldsymbol{NH_3}}} = \frac{\sqrt{\mathbf{127}}}{\sqrt{\mathbf{17}}} = \mathbf{2.74}$

Real Gases

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

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

corrected pressure	corrected volume
P_{ideal}	V_{ideal}
Attractive force = less collisions	**Actual volume of particles**

Root Mean Square Velocity

$$u_{rms} = \sqrt{\frac{3RT}{M}}$$

R = universal gas constant (the energy one)

T = Kelvin Temperature

M = molar mass in KILOGRAMS (b/c of the Joule in “R”)

Gas Stoichiometry #1

If reactants and products are at the same conditions of temperature and pressure, then mole ratios of <u>*gases*</u> are also volume ratios.

$3\ H_2(g) \quad + \quad N_2(g) \quad \rightarrow \quad 2NH_3(g)$

3 moles H_2 + 1 mole N_2 → 2 moles NH_3

3 liters H_2 + 1 liter N_2 → 2 liters NH_3

Gas Stoichiometry #2

How many liters of ammonia can be produced when 12 liters of hydrogen react with an excess of nitrogen?

$$3\ H_2(g) \quad + \quad N_2(g) \quad \rightarrow \quad 2NH_3(g)$$

$$\frac{12\ \text{L}\ H_2}{} \left| \frac{2\ \text{L}\ NH_3}{3\ \text{L}\ H_2} \right. = 8.0\ \text{L}\ NH_3$$

Gas Stoichiometry #3

How many liters of oxygen gas, at STP, can be collected from the complete decomposition of 50.0 grams of potassium chlorate?

$$2\ KClO_3(s) \rightarrow 2\ KCl(s) + 3\ O_2(g)$$

50.0 g $KClO_3$	1 mol $KClO_3$	3 mol O_2	22.4 L O_2
	122.55 g $KClO_3$	2 mol $KClO_3$	1 mol O_2

= 13.7 L O_2

Gas Stoichiometry #4

How many liters of oxygen gas, at 37.0°C and 0.930 atmospheres, can be collected from the complete decomposition of 50.0 grams of potassium chlorate?

$$2\ KClO_3(s) \rightarrow 2\ KCl(s) + 3\ O_2(g)$$

50.0 g $KClO_3$	1 mol $KClO_3$	3 mol O_2
	122.55 g $KClO_3$	2 mol $KClO_3$

= "n" mol O_2

0.612 mol O_2

$$V = \frac{nRT}{P} = \frac{(0.612\,\text{mol})(0.0821\,\frac{L \cdot atm}{mol \cdot K})(310\,K)}{0.930\,atm} = 16.7\ L$$

Overview: 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.