N27 – Gases

Gas Density and More

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:

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

$$M = \frac{DRT}{P}$$

D = Density

M = Molar Mass

P = Pressure

R = Gas Constant

T = Temperature in Kelvins

"Molar Mass Kitty puts Dirt Over its Pee" - Ha!

Kinetic Energy of Gas Particles

At the same conditions of temperature, all gases have the same <u>average</u> kinetic energy.

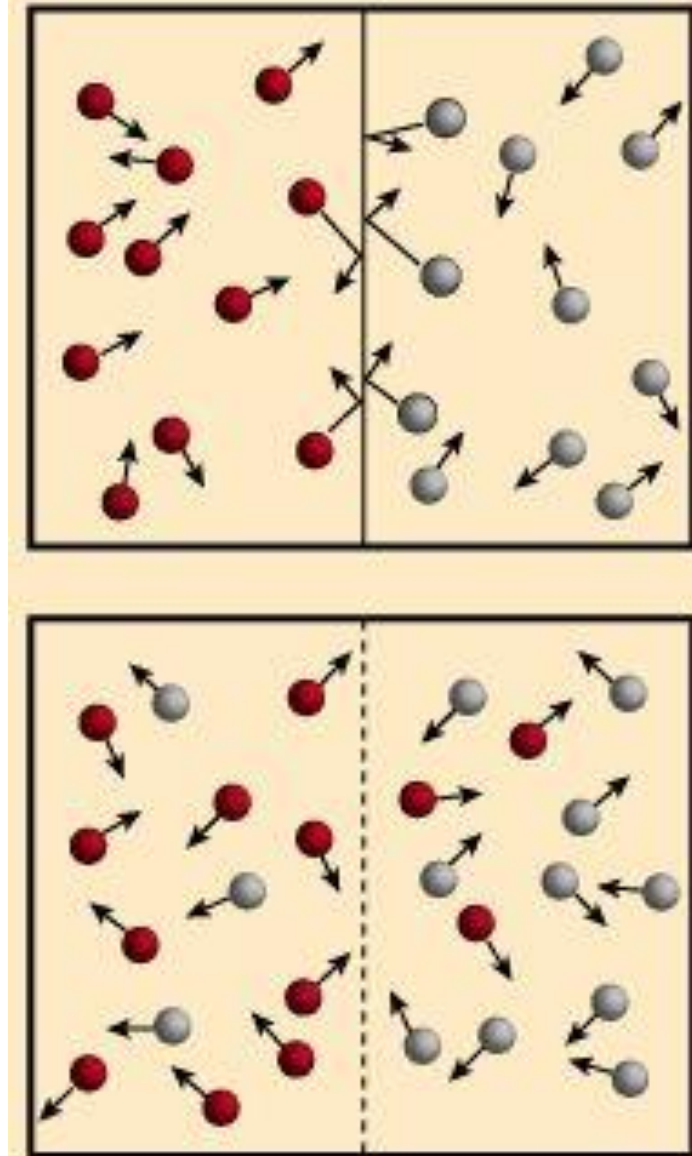$$KE = \frac{1}{2}mv^2$$

The Meaning of Temperature

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

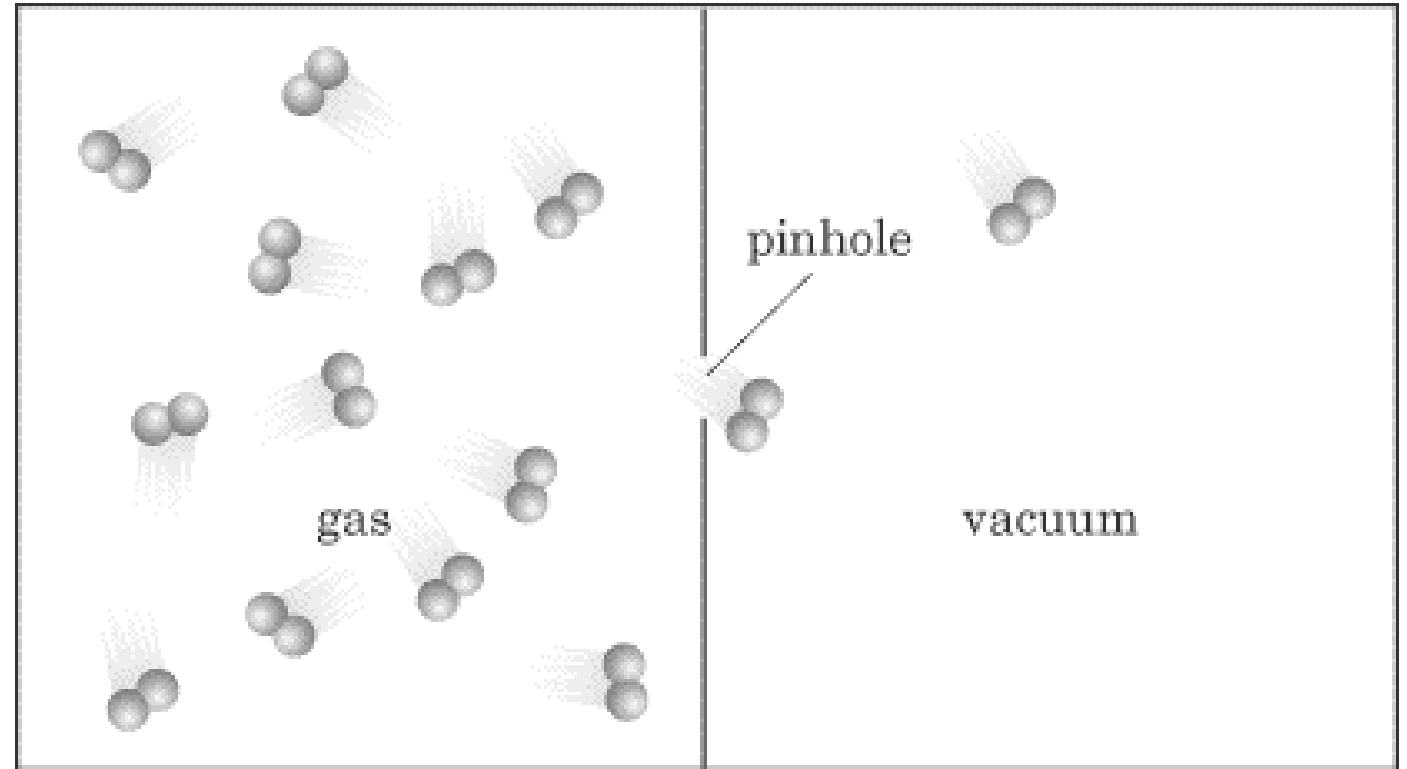$$(KE)avg = \frac{3}{2}RT$$

Diffusion

Diffusion describes the mixing of gases.

The rate of diffusion is the rate of gas mixing.

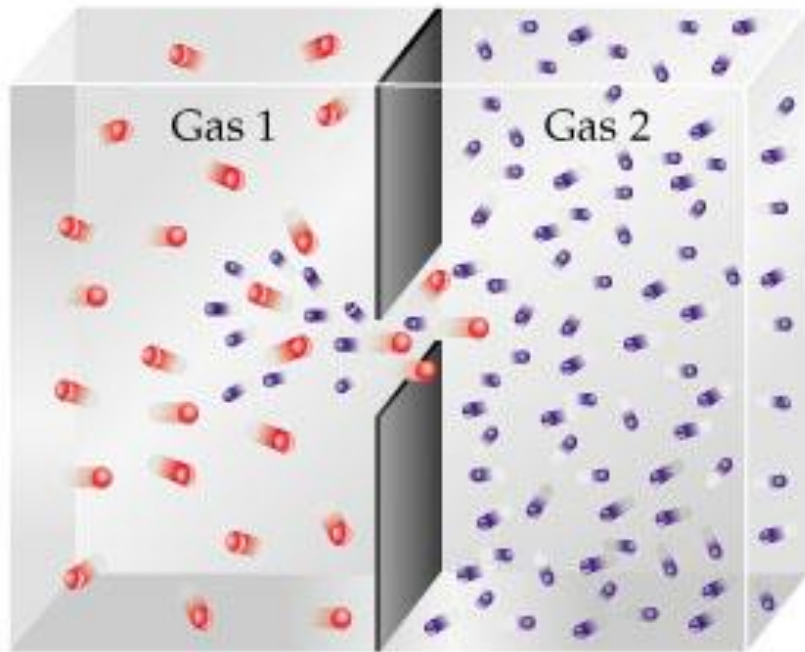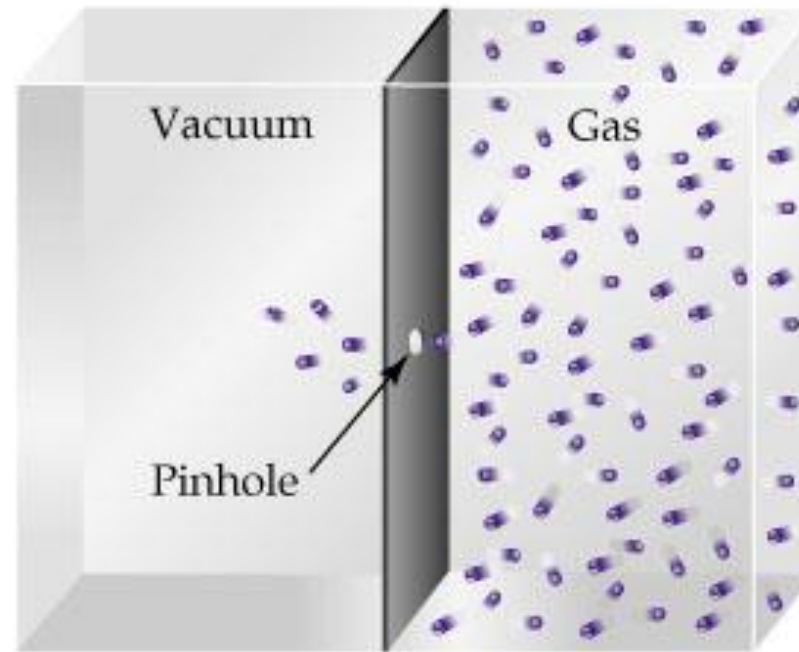# Effusion

Effusion describes the passage of gas into an evacuated chamber.

Diffusion versus Effusion

Diffusion

Effusion

Graham's Law

Rate of Effusion:

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

Rate of Diffusion:

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

M = molar mass

Often given seconds it takes the gases to travel versus distance. No big deal! Use that data.

If they tell you gas 1 travels 4 times faster, that is this part of the equation:

$$\frac{Rate\ gas\ 1}{Rate\ gas\ 2}$$

Careful to notice it isn't gas 1 on the top of both parts of the equation!

Practice

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

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

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

Real Gases

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

$$\left(P_{observed} + a \left(\frac{n}{V} \right)^2 \right) \text{x} \left(V - nb \right) = nRT$$

corrected pressure

P_{ideal}

Attractive forces in a real gas = less collisions

corrected volume

V_{ideal}

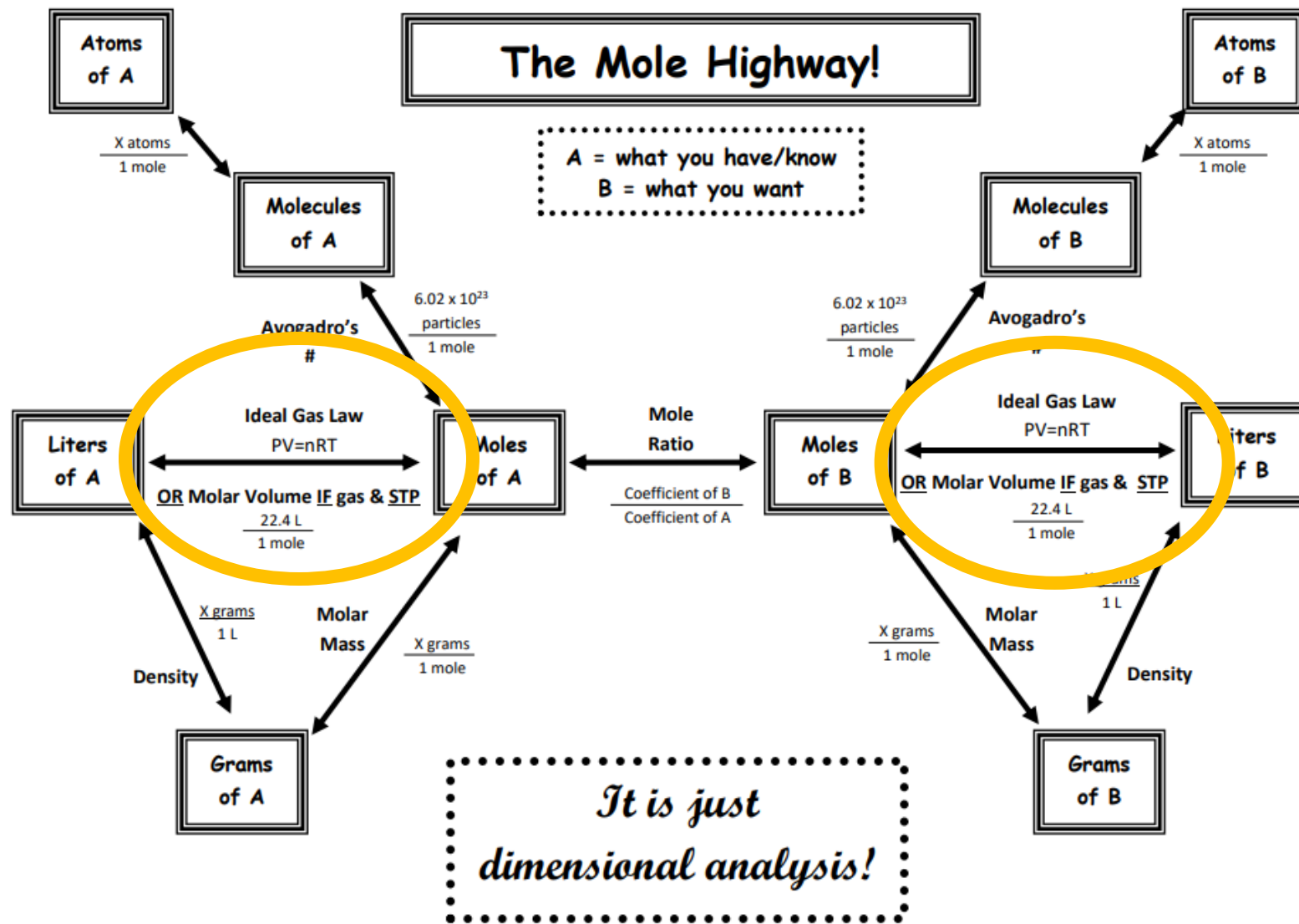*Real gases take up some volume since they are not "point particles"*

Root Mean Square Velosity

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

**R** = universal gas constant (the energy one, 8.314)

T = Kelvin Temperature

M = molar mass in _KILOGRAMS_ (b/c of the Joule in "R")

Gas Stoichiometry

The Mole Highway!

A = what you have/know
B = what you want

Atoms of A

$\dfrac{\text{X atoms}}{\text{1 mole}}$

Molecules of A

$\dfrac{6.02 \times 10^{23} \text{ particles}}{\text{1 mole}}$

Avogadro's #

Ideal Gas Law
PV=nRT

OR Molar Volume IF gas & STP

$\dfrac{22.4 \text{ L}}{\text{1 mole}}$

Liters of A

Moles of A

$\dfrac{\text{X grams}}{\text{1 L}}$

Density

Molar Mass

$\dfrac{\text{X grams}}{\text{1 mole}}$

Grams of A

Mole Ratio

$\dfrac{\text{Coefficient of B}}{\text{Coefficient of A}}$

Atoms of B

$\dfrac{\text{X atoms}}{\text{1 mole}}$

Molecules of B

$\dfrac{6.02 \times 10^{23} \text{ particles}}{\text{1 mole}}$

Avogadro's

Ideal Gas Law
PV=nRT

OR Molar Volume IF gas & STP

$\dfrac{22.4 \text{ L}}{\text{1 mole}}$

Moles of B

Liters of B

$\dfrac{\text{X grams}}{\text{1 mole}}$

Molar Mass

$\dfrac{1 \text{ L}}{}$

Density

Grams of B

It is just dimensional analysis!

Gas Stoichiometry #1

If reactants and products are at the same conditions of temperature and pressure, then mole ratios of _gases_ are also **volume ratios.**
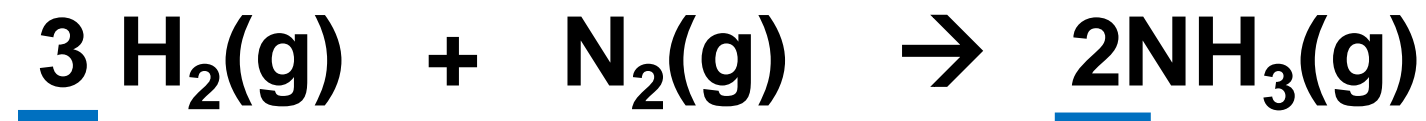
$$3\ H_2(g)\quad +\quad N_2(g)\quad \rightarrow\quad 2NH_3(g)$$

3 moles H_2 + 1 mole N_2 $\rightarrow$ 2 moles NH_3
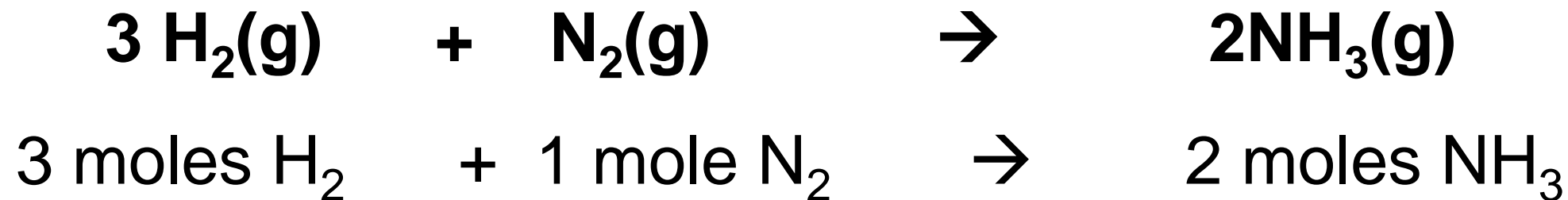
So how many liters of each?

3 liters H_2 + 1 liter N_2 $\rightarrow$ 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?

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

$$\frac{12\ L\ H_2}{} \cdot \frac{2\ L\ NH_3}{3\ L\ H_2} = 8.0\ \ L\ NH_3$$

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 KClO₃	1 mol KClO₃	3 mol O₂	22.4 L O₂
	122.55 g KClO₃	2 mol KClO₃	1 mol O₂

$$= 13.7 \quad L \ O_2$$

Gas Stoichiometry #4

How many liters of O_2, at 37.0°C and 0.930 atm, can be collected from the decomposition of 50.0 g of $KClO_3$?

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

50.0 g KClO₃	1 mol KClO₃	3 mol O₂	= "n" mol O₂
	122.55 g KClO₃	2 mol KClO₃	0.612 mol O₂

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

Kinetic Molecular Theory

- **Particles of matter are ALWAYS in motion**
- **Individual particles have a volume ≈ zero.**
- **Pressure is caused by collisions of particles with the container walls**
- **No attraction or repulsive forces between particles**
- **Average kinetic energy ∝ Kelvin temperature of a gas.**