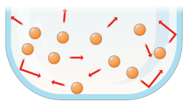
# GASES



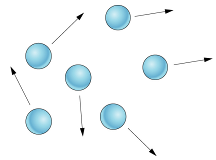






# Characteristics of Gases

1) Gases consist of hard, **SPHERICAL** particles (atoms or molecules)

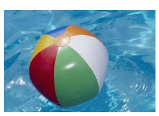
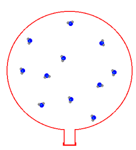


2) Particles are **SMALL** and do have ***MASS***

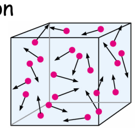
3) Large amounts of **EMPTY SPACE**… so they are easily **COMPRESSED**

# Characteristics of Gases

4) Easily **EXPAND** and uniformly fill any container (can mix with any other gas)

5) Move **RAPIDLY** and **CONSTANTLY**… exerting pressure (force / area) on their surroundings



# Kinetic Theory of Gases

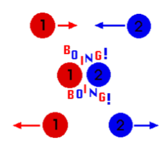
- Explains the properties of gases and why they behave the way they do…

1) All particles are in **constant**, random motion

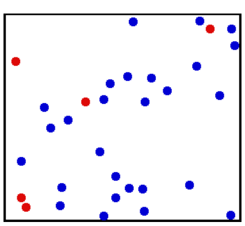
2) All collisions between particles are perfectly **elastic** and particles exert no forces on each other

3) Volume of the particles in a gas is **negligible**

4) Average kinetic energy of molecules is proportional to the **Kelvin** temperature



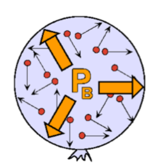

# Kinetic Theory of Gases



**CONSTANT MOTION WHICH IS RANDOM AND RAPID (ALLOWS FOR COMPRESSIBILITY AND PRESSURE) WITH COLLISIONS!!**

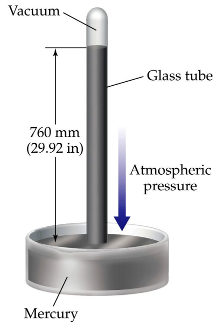
# Gas Pressure

*Force exerted by a gas on an object*



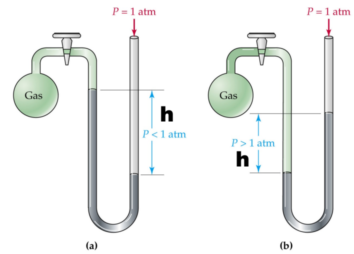
- Due to simultaneous collisions of billions of rapidly moving particles
- **BAROMETER**: device used to measure atmospheric pressure (invented by Evangelista Torricelli)
- Atmospheric pressure exists due to the force exerted by air molecules striking other objects

## Barometer




## Manometer

*Measures the pressure of a gas in a container*




**Gas pressure LESS than atmospheric**

**Gas = atmosphere - h**

**Gas pressure MORE than atmospheric**

**Gas = atmosphere + h**

## Units of Pressure

**1 atm = 760 mm Hg**

**1 atm = 760 torr**

**1 atm = 101.3 kPa = 101,325 Pa**

**STP = 0°C and 1 atm**

- **EXAMPLE:**

Express 3400 mm Hg in atm.

$$3400 \text{ mm Hg} \times \frac{1 \text{ atm}}{760 \text{ mm Hg}} = 4.5 \text{ atm}$$

## Gas Laws

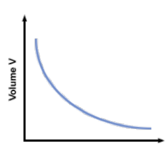
- Factors including PRESSURE, AMOUNT, TEMPERATURE, and VOLUME will affect a gas
- How a gas will respond to these changes can be predicted using certain gas laws

## Boyle's Law

**Pressure INCREASES, then volume DECREASES**

**(inverse relationship)**

Pressure times volume is a constant as long as temperature is constant: PV = k




## Boyle's Law

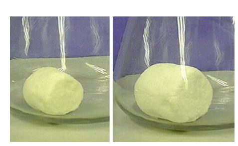
**Pressure INCREASES, then volume DECREASES**

**(inverse relationship)**

Temp is CONSTANT!!

$$P_1V_1 = P_2V_2$$

Only works at LOW pressures!



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- **REAL WORLD EXAMPLES:**




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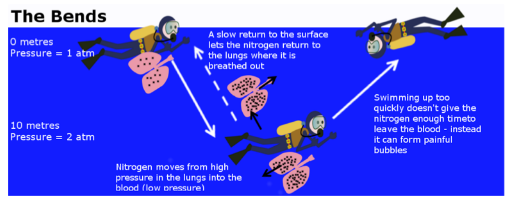


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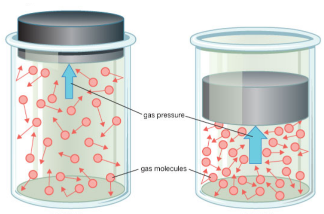



## Boyle's Law

**Pressure INCREASES, then volume DECREASES**

**(inverse relationship)**

- **EXAMPLE:**

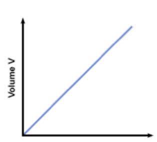
A balloon is filled with 25 L of air at 1.0 atm pressure. If the **pressure** is changed to 1140 mm Hg, what is the new **volume**?

## Charles' Law

**Volume INCREASES, as the temperature INCREASES**

**(proportional relationship)**

Volume is proportional to temperature if held at constant pressure: V = kT




## Charles' Law

**Volume INCREASES, as the temperature INCREASES**

**(proportional relationship)**

Pressure is CONSTANT!!

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Temperature is in KELVIN!!

(°C + 273 = K)

If °C used, could get zero on bottom which would not work!

## Charles' Law

**Volume INCREASES, as the temperature INCREASES**

**(proportional relationship)**

- **REAL WORLD EXAMPLES:**



## Charles' Law

**Volume INCREASES, as the temperature INCREASES**

**(proportional relationship)**


HEAT 'EM UP… SPEED 'EM UP! More KE in the boiling water so moving faster and colliding with walls with greater force!

## Charles' Law

**Volume INCREASES, as the temperature INCREASES**

**(proportional relationship)**

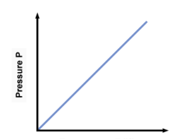
- **EXAMPLE:**

  What is the temperature of a gas expanded from 2.5 L at 25°C to 4.1 L at constant pressure?

## Gay-Lussac's Law

**Pressure INCREASES, as the temperature INCREASES**

**(proportional relationship)**

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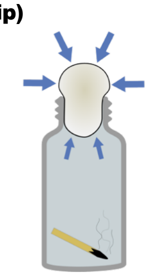


## Gay-Lussac's Law

**Pressure INCREASES, as the temperature INCREASES**

**(proportional relationship)**

- **REAL WORLD EXAMPLES:**


Do not heat aerosol cans!


## Gay-Lussac's Law

**Pressure INCREASES, as the temperature INCREASES**

**(proportional relationship)**

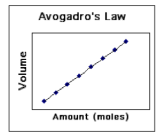
- **EXAMPLE:**

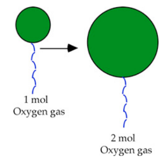
What is the **pressure** inside a 0.250 L can of deodorant that starts at 25°C and 1.2 atm if the **temperature** is raised to 100°C?

## Avogadro's Law

**Volume INCREASES, as the amount (# of moles) INCREASES**

**(proportional relationship)**




## Avogadro's Law

**Volume INCREASES, as the amount (# of moles) INCREASES**

**(proportional relationship)**

Pressure and temperature are CONSTANT!!

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

## Avogadro's Law

**Volume INCREASES, as the amount (# of moles) INCREASES**

**(proportional relationship)**

- **EXAMPLE:**

If you begin with 15.0 L of $H_2$, what **volume** of $N_2$ will be needed to complete this reaction and how many **liters** of $NH_3$ will form?

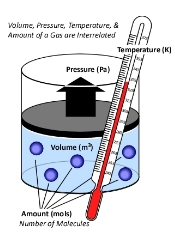
$N_2 + 3\ H_2 \rightarrow 2\ NH_3$

## Combined Gas Law

**Boyle's, Charles', and Gay-Lussac's Laws (and Avogadro's) COMBINED!**

If # of molecules is CONSTANT, then get rid of n!!

$$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}$$




Temperature is in KELVIN!!

## Combined Gas Law

**Boyle's, Charles', and Gay-Lussac's Laws COMBINED!**

- **EXAMPLE:**

A 15 L cylinder of gas at 486 kPa pressure and 25°C is heated to 75°C and compressed to 17 atm. What is the new **volume**?

## Practice

- **EXAMPLES:**

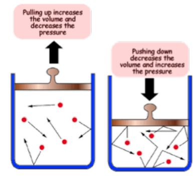
Bacteria produce methane gas in sewage-treatment plants. If a bacterial culture produces 60.0 mL of methane gas at 700.0 mm Hg, what **volume** (in L) would be produced at 760.0 mm Hg?

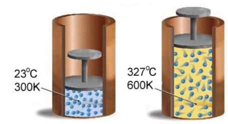
A pressure cooker raises the temperature of its contents by keeping the contents under pressure. The volume of the cooker is 4.0 L. Steam at 100.0°C and 1.00 atm usually cooks the food. If the cooker is placed at 1.15 atm, what will the **temperature** of the steam be in °C?

## KMT and Gas Laws

**BOYLE'S**

V is decreased, particles hit walls more often… increasing P





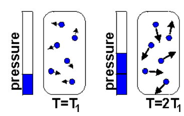

**CHARLES'**

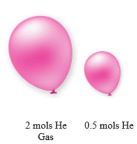
T is increased, increasing gas speeds and hitting walls with more force… more V

## KMT and Gas Laws

**GAY-LUSSAC'S**

T is increased, gas speeds increase and hit walls with more force… increasing P




**AVOGADRO'S**

n is increased, gas takes up more space… increasing V

## Ideal Gas Law

**Includes all FOUR factors that can affect a gas!**

$$PV = nRT$$

**Not restricted by STP conditions now!!**

n = # of moles OR grams / molar mass

R = ideal gas constant = 0.082 L · atm / K · mol

- *IDEAL GAS*: follows gas laws at **ALL** conditions of pressure and temperature… assumes particles have no volume and no attractive forces (These don't really exist!)
- *REAL GAS*: have volume and attract to each other and will only be ideal at very low pressures or high temperatures

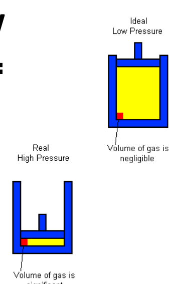
## Ideal Gas Law

- **Limitations of the Ideal Gas Law:**

1) Works well at LOW PRESSURES and HIGH TEMPERATURES (that's when real gases act ideally)

2) Most gases do not behave ideally above 1 atm pressure

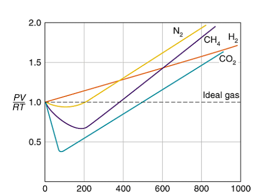
3) Does not work well near condensation conditions of a gas



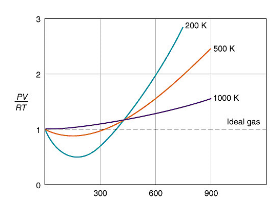

**Ideal gases don't exist because molecules do take up space and attractive forces exist otherwise liquids would not form… math easier and close approximation so we assume gases are ideal!**

## Ideal Gas Law

- **Limitations of the Ideal Gas Law:**




Gases are ideal at LOWER pressures!




Gases are ideal at HIGH temperatures!

## Van der Waals Equation

$$\left[P + \frac{an^2}{v^2}\right](V - nb) = nRT$$

$$\underbrace{\overset{\text{actual pressure}}{P} + a\left(\frac{n}{V}\right)^2}_{\text{ideal gas pressure}} \times \underbrace{(\overset{\text{actual volume}}{V} - nb)}_{\text{ideal gas volume}} = nRT$$

- Corrects for real gases:

Pressure is INCREASED… Intermolecular forces lower the real pressure!

Volume is DECREASED… Volume of the container available is not 100%

**REAL GASES HAVE ATTRACTIVE FORCES AND TAKE UP SPACE!!**

## Ideal Gas Law

- **EXAMPLES:**

Determine the volume occupied by 2.34 grams of carbon dioxide gas at STP.

If I have 2.4 moles of a gas held at a temperature of 97°C and in a container with a volume of 45,000 mL, what is the **pressure** of the gas?

What **volume** will one mole of $N_2$ occupy at STP?

## Ideal Gas Law

- Variations (by rearranging the equation):

MOLAR MASS

$$\mathbf{MM = \frac{gRT}{PV}} \qquad \mathbf{MM = \frac{dRT}{P}}$$

"Molar Mass Kitty Cat…"
Cats put dirt (dRT) over their pee (P)

DENSITY

$$\mathbf{d = \frac{MMP}{RT}}$$

**Knowing these can save you time!!**

## Ideal Gas Law

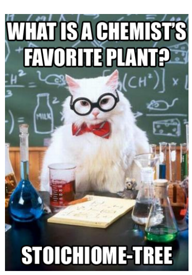
- **EXAMPLES:**

Calculate the density of nitrogen dioxide gas at 0°C and 1.00 atm.

If 2.06 g of a gas occupies 3.33 L at 17.0°C and 700. torr, what is the **molecular weight** of the gas?

## Gas Law Stoichiometry

- Use mole ratio from the equation to determine number of moles of an unknown substance (might have to do a limiting reagent problem)
- Use ideal gas law to find desired values
- Use total number of moles of gases formed from the equation (if needed)




**If at STP, use 1 mole = 22.4 L…**
**If not, do STOICH!**

## Gas Law Stoichiometry

- **EXAMPLE:**

What **volume** of gaseous products can be formed by the reaction of 8.0 g Zn with 8.0 g $NH_4NO_3$ at 0.998 atm and 22°C?

$Zn(s) + NH_4NO_3(s) \rightarrow N_2(g) + ZnO(s) + 2\ H_2O(g)$

**DO LIMITING REAGENT FIRST!**

**Asking for the volume of GAS products… there are TWO ($N_2$ and $H_2O$) in this reaction and there are 3 MOL TOTAL of gases formed (1 mol $N_2$ + 2 mol $H_2O$), so use that number of mol in the problem!**

## Gas Law Stoichiometry

- **EXAMPLE:**

What volume of gaseous products can be formed by the reaction of 8.0 g Zn with 8.0 g $NH_4NO_3$ at 0.998 atm and 22°C?

$Zn(s) + NH_4NO_3(s) \rightarrow N_2(g) + ZnO(s) + 2\ H_2O(g)$

$$8.0\text{ g Zn} \times \frac{1\text{ mol Zn}}{65.39\text{ g Zn}} \times \frac{3\text{ mol Gases}}{1\text{ mol Zn}} = 0.37\text{ mol}$$

**Limiting Reagent**

$$8.0\text{ g } NH_4NO_3 \times \frac{1\text{ mol } NH_4NO_3}{80.052\text{ g } NH_4NO_3} \times \frac{3\text{ mol Gases}}{1\text{ mol } NH_4NO_3} = 0.30\text{ mol}$$

## Gas Law Stoichiometry

- **EXAMPLE:**

What volume of gaseous products can be formed by the reaction of 8.0 g Zn with 8.0 g $NH_4NO_3$ at 0.998 atm and 22°C?

$Zn(s) + NH_4NO_3(s) \rightarrow N_2(g) + ZnO(s) + 2\ H_2O(g)$

**Now the number of mol of gaseous products is known… Use the Ideal Gas Law to find the volume of the gases at these conditions!**

## Gas Law Stoichiometry

- **EXAMPLE:**

What volume of gaseous products can be formed by the reaction of 8.0 g Zn with 8.0 g $NH_4NO_3$ at 0.998 atm and 22°C?

$Zn(s) + NH_4NO_3(s) \rightarrow N_2(g) + ZnO(s) + 2\ H_2O(g)$

$V = nRT / P$

$$V = \frac{(0.30\text{ mol})(0.082\text{ L atm mol}^{-1}\text{ K}^{-1})(295\text{ K})}{0.998\text{ atm}}$$

$V = 7.3$ L

## Gas Law Stoichiometry

- **EXAMPLE:**

Determine the volume of hydrogen gas that will form at 15.0°C and 780.0 mm Hg if 100.0 mL of 1.00 M HCl react completely with Mg metal.

## Gas Law Stoichiometry

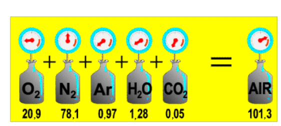
- **EXAMPLE:**

A sample of methane gas having a volume of 2.80 L at 25.0°C and 1.65 atm was mixed with a sample of oxygen gas having a volume of 35.0 L at 31.0°C and 1.25 atm. The mixture was then ignited to form carbon dioxide and water. Calculate the volume of $CO_2$ formed at a pressure of 2.50 atm and a temperature of 125°C.

## Dalton's Law

**Total pressure of a mixture of gases is equal to the sum of the PARTIAL PRESSURES of the component gases!**

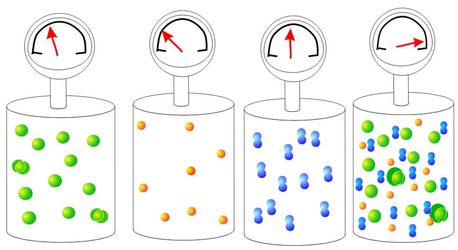
$$P_{Total} = P_1 + P_2 + P_3 + \dots P_n$$




Useful for Water Displacement problems!

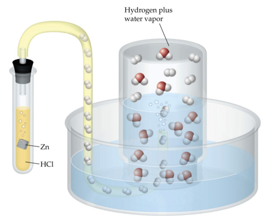
## Dalton’s Law

**2 atm + 1 atm + 3 atm = 6 atm**



## Water Displacement

- Common to collect gas in the lab by this method
- Some of the pressure is due to the **WATER VAPOR** collected as the gas was passing through the liquid… Need to correct for this!
- Look up the partial pressure due to water vapor at a given temperature and **SUBTRACT** that value from the total pressure to get actual $P_{gas}$:




$$\mathbf{P_{Total} = P_{Gas} + P_{H_2O}}$$

## Water Displacement

**Water Vapor Pressure Table**

| Temperature (°C) | Pressure (mmHg) | Temperature (°C) | Pressure (mmHg) | Temperature (°C) | Pressure (mmHg) |
|---|---|---|---|---|---|
| 0.0 | 4.6 | 19.5 | 17.0 | 27.0 | 26.7 |
| 5.0 | 6.5 | 20.0 | 17.5 | 28.0 | 28.3 |
| 10.0 | 9.2 | 20.5 | 18.1 | 29.0 | 30.0 |
| 12.5 | 10.9 | 21.0 | 18.6 | 30.0 | 31.8 |
| 15.0 | 12.8 | 21.5 | 19.2 | 35.0 | 42.2 |
| 15.5 | 13.2 | 22.0 | 19.8 | 40.0 | 55.3 |
| 16.0 | 13.6 | 22.5 | 20.4 | 50.0 | 92.5 |
| 16.5 | 14.1 | 23.0 | 21.1 | 60.0 | 149.4 |
| 17.0 | 14.5 | 23.5 | 21.7 | 70.0 | 233.7 |
| 17.5 | 15.0 | 24.0 | 22.4 | 80.0 | 355.1 |
| 18.0 | 15.5 | 24.5 | 23.1 | 90.0 | 525.8 |
| 18.5 | 16.0 | 25.0 | 23.8 | 95.0 | 633.9 |
| 19.9 | 16.5 | 26.0 | 25.2 | 100.0 | 760.0 |

## Water Displacement

- **EXAMPLE:**

A chemist collects a sample of $H_2S(g)$ over water at a temperature of 27.0°C. The total pressure of the gas that has displaced a volume of 15 mL of water is 207.33 kPa. What is the pressure of the $H_2S$ gas collected?

## Mole Fraction

- Total **number of moles** are important for Dalton’s Law, not the identity of the gas particles…

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

$$\mathbf{n_{Total} = n_1 + n_2 + n_3\ldots}$$

$$\frac{P_1}{P_{Total}} = \frac{n_1}{n_{Total}}$$

## Dalton’s Law

- **EXAMPLE:**

$N_2$, He, Ne, and Ar are placed into a container. The partial pressures of each are as follows: $N_2$ = 80 kPa, He = 73 kPa, and Ne = 49 kPa. The total pressure in the container is 310 kPa. What is the **partial pressure** of Argon?

Hydrogen gas, made from the reaction of aluminum and hydrochloric acid, was collected over water at 25.0°C. If the pressure was equal to 1.00 atm, what was the **pressure** of the hydrogen in torr?

# Dalton's Law

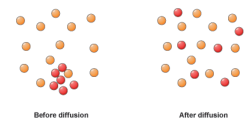
- **EXAMPLE:**

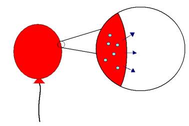
The total pressure of a gas mixture is 1.20 atm. The mixture contains 0.10 mol of $N_2$ and 0.20 mol of $O_2$. What is the **partial pressure** of $O_2$?

A sample of solid $KClO_3$ was heated and decomposed in a test tube. The oxygen produced was collected by water displacement at 22.0°C at a total pressure of 754 torr. The volume of gas collected was 0.650 L. Calculate the **partial pressure** of $O_2$ and the **mass** of $KClO_3$ that was decomposed.

# Diffusion vs. Effusion

**DIFFUSION**: gas molecules will move to fill an area until the molecules are evenly spread out



**EFFUSION**: when a gas is confined to a container that has a small opening, molecules will randomly encounter the opening and pass through it

# Diffusion vs. Effusion

- **Graham's Law of Diffusion / Effusion:** Rate at which a gas will travel / effuse is inversely proportional to the square root of the gas's molar mass

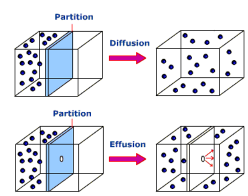
$$\frac{r_1}{r_2} = \sqrt{\left(\frac{MM_2}{MM_1}\right)}$$

**R** = rate (VELOCITY) of diffusion or effusion (m/s)

**MM** = molar mass (g/mol)

# Diffusion vs. Effusion

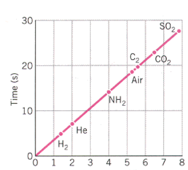
- **Graham's Law of Diffusion / Effusion:** Rate at which a gas will travel / effuse is inversely proportional to the square root of the gas's molar mass




# Diffusion vs. Effusion

- **Graham's Law of Diffusion / Effusion:** Rate at which a gas will travel / effuse is inversely proportional to the square root of the gas's molar mass

**When comparing effusion rates for gases… the LOWER the molar mass, the FASTER the gas will effuse!!!**




# Graham's Law

- **EXAMPLE:**

Helium is placed in scuba tanks so it effuses into the blood instead of $N_2$. How much **faster** is helium than $N_2$?

What is the **molar mass** of an unknown gas that effuses 2.12 times faster than pentane through a porous barrier?

## Kinetic Energy of Gases

- Directly proportional to the Kelvin temperature of the gas

**KE = ½ mv²**

m= mass
v= velocity

- **HIGHER** the temperature… **GREATER** the KE!
- Kinetic energy depends on temperature… Gases at the **SAME** temp have the **SAME** kinetic energy!!

**KE and Velocity are NOT the same!**

## Root Mean Square Velocity

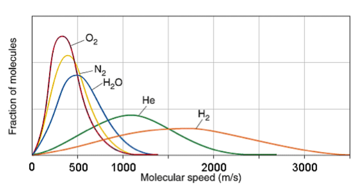
*Gas molecules have different speeds and in different directions… this tries to find a SINGLE value*

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

**Do not know or calculate just understand what it means…**

- At any given temperature, gas molecules have both high and low speeds
- As temperature **INCREASES**, distribution of speeds is across a wider range (flatter graph)
- Velocity of a gas dependent on mass and temperature (greater MM = slower / greater temp = faster)

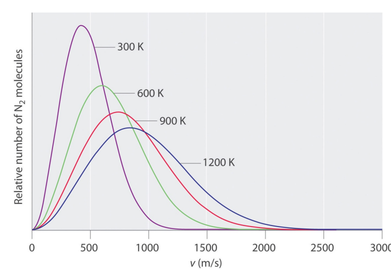
## Root Mean Square Velocity




**GREATER Molar Mass = SLOWER!**

Notice that some molecules are moving slower and some are moving faster for EACH!

## Root Mean Square Velocity




**GREATER Temp = FASTER!**
**(Flatter graphs too!!)**