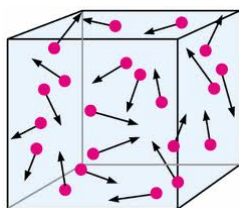
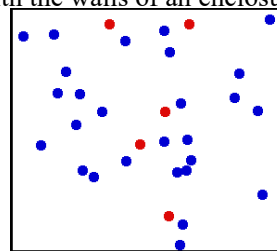


**5.6 - 5.8 Kinetic Molecular Theory (KMT)****Kinetic Theory Gas Model**

Gas particles can be considered point particles colliding with the walls of an enclosure randomly.



The higher the temperature, the faster molecules move, the more kinetic energy they have.

**Kinetic Theory Postulates**

For KMT to work, a few guiding assumptions are necessary:

1. The particles are so small as compared to the distances between them that the volume of individual particles can assumed to be negligible (zero).
2. The particles are in constant motion, and the collisions of the particles against the container walls are the cause of the pressure exerted by the gas.
3. Particles exert no forces on each other: they neither attract nor repel each other.
4. The average kinetic energy of gas particles is assumed to be directly proportional to the Kelvin temperature of the gas.

**KMT vs. Gas Laws**

How well, and how exactly, does the model apply to the gas laws we've seen?

1. Boyle's Law: Pressure vs. Volume.  
Pressure is inversely proportional to volume: as one goes up, the other goes down.  
KMT suggests that a decrease in volume causes particles to hit the walls of the container with greater frequency (thus raising the pressure).
2. Gay-Lussac's Law: Pressure vs. Temperature.  
KMT accounts for this because as the temperature of a gas increases, the speeds of its particles increases. As a result, particles collide with the container with more force, and therefore more pressure in a fixed volume.

**KMT vs. Gas Laws**

3. Charles's Law: Volume vs. Temperature.  
KMT shows a gas being heated as moving faster, thus imparting more force as it strikes the walls of the container. To prevent pressure from increasing, the density of the particles must decrease by an increase of volume.
4. Avogadro's Law: Particles vs. Volume - this ties into Dalton's Law of Partial Pressures.  
KMT explains: when increasing particles (regardless of size), to keep system pressure constant one needs a bigger corral.  
Yee-Haw!

Molecule corral!

**Kinetic Energy of Ideal Gas**

We have seen from KMT that the Kelvin temperature directly relates to average kinetic energy (also called internal energy) of the particles.

Note on R: thus far the Universal Gas Constant has been used to determine state parameters of gases.

We will look at energy too, but we need to use a different value for R (based on  $\text{m}^3$  and Pa).

$$KE_{Average} = \frac{3}{2}nRT$$

n = number of moles  
R = Universal gas constant  
(8.315 J/K mol)  
T = Kelvins

Note: for real gases, where vibrations factor in, the coefficient changes. That won't be explored here - assume all gases behave ideally.

**Energy Examples**

1. What's the average kinetic energy of 3.00 moles of oxygen gas at 25 degrees C?

$$KE_{avg} = \frac{3}{2} nRT$$

$$= \frac{3}{2} 3.00 \text{ moles} \cdot 8.315 \text{ J/mol} \cdot \text{K} \cdot 298 \text{ K}$$

$$= \boxed{1.12 \text{ E}4 \text{ J}}$$

2. What is the average kinetic energy per molecule?

$$KE_{avg} = \frac{3}{2} nRT$$

$$= \frac{3}{2} \cdot \frac{1 \text{ molecule}}{6.02 \text{ E}23 \text{ particles}} \cdot 8.315 \text{ J/mol} \cdot \text{K} \cdot 298 \text{ K}$$

$$= \boxed{6.17 \text{ E} - 21 \text{ J}}$$

**Relation to Kinetic Energy**

The kinetic energy of a particle of an ideal gas is:

$KE_{avg} = \frac{1}{2} m v_{rms}^2$	m = mass of one molecule (kg) $v_{rms}$ = root-mean-square speed (m/s) (to be explained)
<b>AP Equation</b>	

How do you determine the mass of one molecule?

$m = \frac{M}{N_A}$	m = mass (kg) M = Molar mass ( <b>kg/mol</b> ) $N_A$ = Avogadro's Number (6.02 E 23 particles/mol)
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**Root-Mean-Squared Speed**

Special kind of average – pertinent to large systems with positive and negative velocity magnitudes.

The square root of the average of the squared speed (m/s) of the gas molecules in a system.

$v_{rms} = \sqrt{v^2}$	$v_{rms} = \sqrt{\frac{v_1^2 + v_2^2 + \dots + v_n^2}{n}}$
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Note: for any gases at the same thermodynamic parameters (P, V, n, T), the amount of kinetic energy is the same. As gas molecules get larger, their speeds decrease, and it balances out such that they have the same average energy.

**Molecular Speed**

The average speed of particles depends on their mass: the smaller a particle, the faster it moves.

The proof for the following is in your book:

$v_{rms} = \sqrt{\frac{3RT}{M}}$	R = Universal gas constant (8.315 J/K mol) T = Kelvins M = molar mass (kg/mol)
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Note: in your book, velocity's symbol is  $u$ . In AP resources (and most other reference materials), it is denoted  $v$ . Thus it is  $v$  here as well.

**Molecular Speed Examples**

3. Comparing rms speeds of ammonia and methane, which would you expect to have the greater speed?

ammonia:  $\text{NH}_3$ : molar mass = 17.04 g/mol

methane:  $\text{CH}_4$ : molar mass = 16.05 g/mol

Methane should have greater speed.

4. At 35°C, what are the speeds?

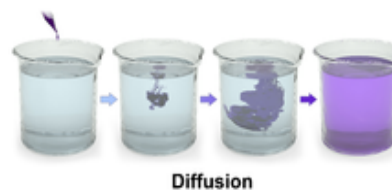
$$v_{rms \text{ NH}_3} = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3 \cdot 8.315 \frac{\text{J}}{\text{K} \cdot \text{mol}} \cdot 308 \text{ K}}{0.01704 \frac{\text{kg}}{\text{mol}}}} = \boxed{671 \text{ m/s}}$$

$$v_{rms \text{ CH}_4} = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3 \cdot 8.315 \frac{\text{J}}{\text{K} \cdot \text{mol}} \cdot 308 \text{ K}}{0.01605 \frac{\text{kg}}{\text{mol}}}} = \boxed{692 \text{ m/s}}$$

**Movement of Particles**

Diffusion – random movement of particles in fluids from area of high to area of low concentration.

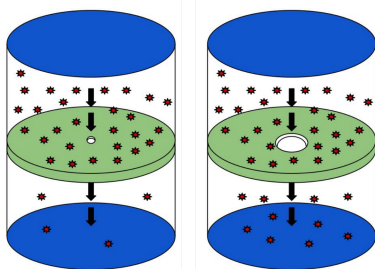
Rate depends on rms speed of particles, which depends on particle size – the smaller, the faster.



### Movement of Particles

Effusion – movement of a gas through a tiny hole into an evacuated chamber.

The smaller the gas particle, the faster it effuses - smaller particles move faster, and have a greater probability of passing through the hole as they strike the walls of the container more frequently.

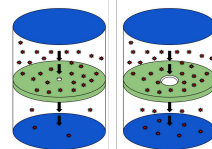


### Comparative Effusion Rate

In the mid 1800's, Scottish chemist Thomas Graham related effusion rate to molar mass.

This is called Graham's Law of Effusion, and compares molar masses directly, and indirectly compares average velocities of gases:

$$\frac{\text{Effusion Rate Gas 1}}{\text{Effusion Rate Gas 2}} = \sqrt{\frac{M_2}{M_1}} \quad M = \text{molar mass}$$



### 5. Effusion Problem

What is the effusion ratio of hydrogen gas vs. oxygen gas? In other words how much faster will hydrogen escape through a hole than oxygen?

Molar mass hydrogen: 2.02 g/mol.

Molar mass oxygen: 32.00 g/mol.

$$\frac{\text{Effusion Rate Hydrogen}}{\text{Effusion Rate Oxygen}} = \frac{\sqrt{O_2}}{\sqrt{H_2}} = \frac{\sqrt{32.00 \frac{\text{g}}{\text{mol}}}}{\sqrt{2.02 \frac{\text{g}}{\text{mol}}}} = 3.98$$

This means that hydrogen effuses about four times faster than oxygen!

### Real Gases

Ideal gases are a hypothetical concept - but real gases at low pressure or high temperature behave very similarly to ideal gases.

When measuring real gases, observed pressure is less than that calculated by the Ideal Gas Law.

The main factors contributing to these errors are:

A. volume effects - since gases have volume, the actual amount of volume available for particles to move through is somewhat less than that of the container. This would add to observed pressure.

B. attractive forces between particles - due to uneven electron distribution, all atoms and molecules will to some extent attract each other. This will lower observed pressure.

By making adjustments accounting for particle volume and attractive forces, correct gas pressures can be calculated.

### Homework

5.6 - 5.8 Booklet Problems.  
Due: next class.

Unit 5 Review Problems  
Due ???

### Molecular Race Demo

Which will move faster, HCl or NH<sub>3</sub>?

What are the molar masses of these two?

Postpone the race - more equipment needed.

