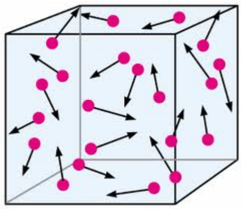
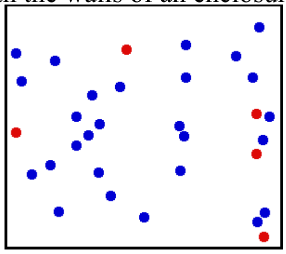
2.A.5 - Kinetic Theory of Gases



Kinetic Theory

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.

Internal Energy of a Gas (Symbol = U)

Total Internal Energy (U) (in Joules) equals the kinetic energy of a gas sample.

This is a measure of energy that could be recovered from the gas to do something, whether it's work, or changing something's temperature.

Far reaching repercussions: in thermodynamics, as a gas' parameters (V, P, T, n) change, the energy of the system changes also.

Monatomic vs. Diatomic Molecules

Monatomic Gases (He, Ne, Ar, Kr, Xe, Rn) have one atom and only translational motion.

Diatomic gases: (H_2, N_2, O_2, F_2, Cl_2) vibrate and rotate also, so have more energy.
Note: these are even less ideal than monatomic gases, due to their additional energy storage modes.

Internal Energy of Monatomic Gas

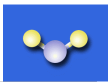
Monatomic gases resemble an ideal gas, as far as internal (kinetic) energy goes (they don't rotate or vibrate).

$K = U = \frac{3}{2}Nk_BT$	$= \frac{3}{2}nRT$
Microscopic N = Atoms of gas k_B = Boltzmann's constant (1.38 E - 23 J/K) T = Kelvins	Macroscopic n = moles of gas R = Universal gas constant (8.31 J/mol x K) T = Kelvins

AP Equation 9 - Resources 3 (microscopic only).

Internal Energy - Diatomic Molecules

Diatomic gases have more modes of movement than monatomic gases (rotation, vibration), they have more internal (kinetic) energy at some temp.



Same basic equation:

$K = U = \frac{5}{2}Nk_BT$	$= \frac{5}{2}nRT$
Microscopic: N = molecules of gas k_B = Boltzmann's Constant (1.38 E - 23 J/K) T = Kelvins	Macroscopic: n = moles of gas R = Universal gas constant (8.31 J/mol • K) T = Kelvins

1. Energy Example

What's the internal energy in 3.00 moles of oxygen gas at 25 degrees C?

$$U = \frac{5}{2} nRT$$

$$= \frac{5}{2} 3.00\, moles \bullet 8.31 J / mol \cdot K \bullet 298 K$$

$$= 1.86 E4 J$$

Relation to Kinetic Energy

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

$$K = U = \frac{1}{2} m v_{rms}^2$$

m = mass of one molecule (kg)
v_{rms} = root-mean-square speed (m/s) (to be explained)

How do you determine the mass of one molecule?

$$m = \frac{M}{N_A}$$

m = mass (kg)
M = Molar mass (**kg**/mol)
N_A = Avogadro's Number (6.02 E 23 particles/mol)

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.

$$v_{rms} = \sqrt{\overline{v^2}} \qquad v_{rms} = \sqrt{\frac{v_1^2 + v_2^2 + \ldots v_n^2}{n}}$$

Note: for any gases at the same thermodynamic parameters (P, V, n, T), the amount of internal energy is the same. As gas molecules get larger, their speeds decrease, and it balances out such that they have the same average energy.

2. Root-Mean-Squared Example

What is the average, and rms average of these speeds: -1.6 m/s, 2.3 m/s, 0.8 m/s?

Average: $\frac{-1.6 \text{ m/s} + 2.3 \text{ m/s} + 0.8 \text{ m/s}}{3} = 0.5$ m/s

rms average: $v_{rms} = \sqrt{\overline{v^2}}$

$$v_{rms} = \sqrt{\frac{(-1.6 \text{ m/s})^2 + (2.3 \text{ m/s})^2 + (0.8 \text{ m/s})^2}{3}}$$

$= 1.7$ m/s

Movement of Particles

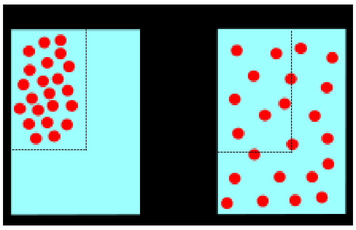
Diffusion – random movement of particles in fluids from area of high to area of low concentration.

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


Diffusion

Movement of Particles

Osmosis – diffusion of a liquid or gas across a permeable membrane.



Homework

2.A.5 Booklet Problems.
Due: next class.

Unit 2.A Review Problems
Due Thursday, 9/27

Molecular Race Demo

Which will move faster, HCl or NH_3?

What are the molar masses of these two?

Postpone the race - more equipment needed.