Intro to Gases

Pressure (_P_): the _force_ per unit area on a surface.

➔ The pressure of a gas is the force that the gas exerts on the _walls_ of its container.

1. A _barometer_ is used to measure atmospheric pressure. When you see the word "barometric" or "barometer", it's telling you the pressure.
2. Units for pressure: on F.C. **1 atm = 760 mmHg = 760 torr = 101.3 kPa**

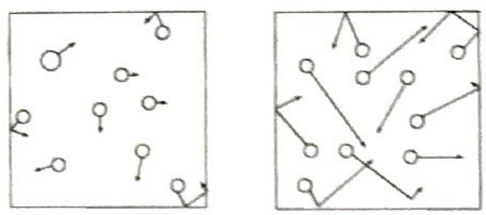
   *note: mmHg = torr! (same exact unit, two names)

Temperature (_T_): a measure of the average _Kinetic_ energy of the particles in a sample of matter.

- Kinetic energy (_KE_) is the energy of _motion_
- Kinetic energy is given by the following equation: $KE = \frac{1}{2}mv^2$

Temperature ∝ Kinetic Energy
∴ Same Temp → Same (average) KE (!!)

Particle diagrams with _vectors_ can be used to illustrate differing kinetic energies at differing temperatures: the _longer_ the vector, the _greater_ the kinetic energy.



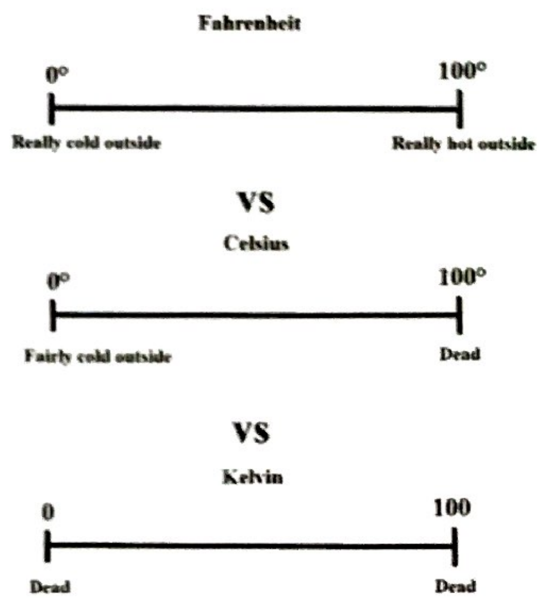
Low Temperature   High Temperature

There are two temperature scales to be familiar with:

1. **Celsius** (°C) scale: the temperature based on water
   a. H_2O: freezes at _ø_ °C and boils at _100_ °C
2. **Kelvin** (K) scale: the absolute temperature scale used in the scientific community.
   a. There are no negative values on the Kelvin scale!!
   b. H_2O: freezes at _273_ K and boils at _373_ K.
   c. _Absolute zero_ is the theoretical lowest temperature possible at which all molecular motion stops.
   d. To convert between °C and K use the formula:

$$K = °C + 273$$

on F.C

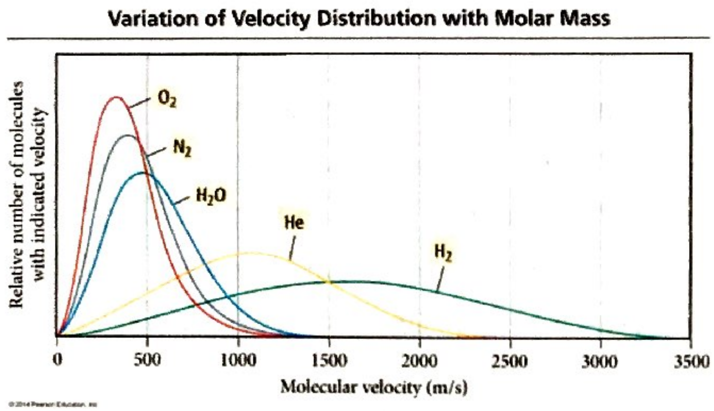


Standard Temperature and Pressure: (_STP_) standard conditions of **1.0 atm, 273.15 K**

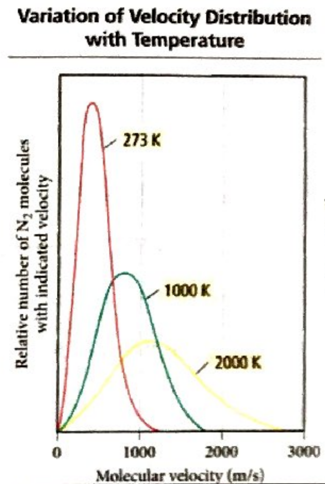
$$\textbf{STP} = \textbf{1.0 atm (pressure), 273.15 K (temperature)}$$

on F.C.

Two factors affect the speed (velocity) of gas particles: **molar mass** and **temperature**!



1) (Assuming same temperature)



2) (Assuming same particle)

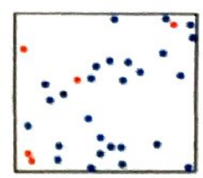
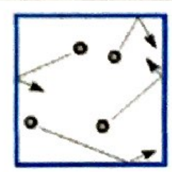
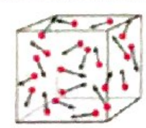
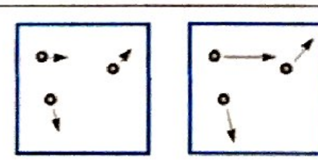
Heat 'em up, Speed 'em up!

In summary,

1. At the same temperature (i.e. same kinetic energy), heavier gas particles are slower than light particles.
2. The same gas particle will move Slower at low temperatures and faster at high temperatures.

Kinetic Molecular Theory (KMT): In Five Postulates

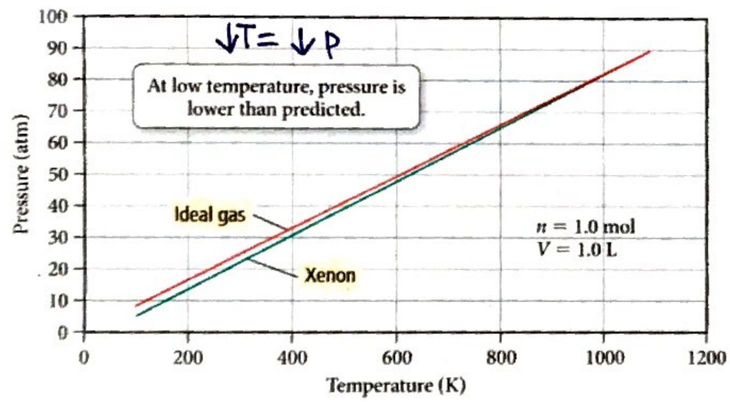
Ideal gases are assumed to behave according to these postulates!

1. Gas particles Small, space between them large. • Gas particles = negligible volume • Gas are compressible because of large spaces between particles	
2. Gas particles bounce off walls (and each other) with no loss of energy → " elastic " collisions Force of particles hitting wall = gas pressure !	
3. Gases particles are in constant, rapid, random motion.	
4. Gas particles do NoT attract or repel each other! • Gas particles can easily glide past each other because attractive forces (IMFs) are insignificant	¯_(ツ)_/¯
5. AVERAGE kinetic energy of a gas is proportional to the velocity of its particles. Heat 'em up, Speed 'em up!	

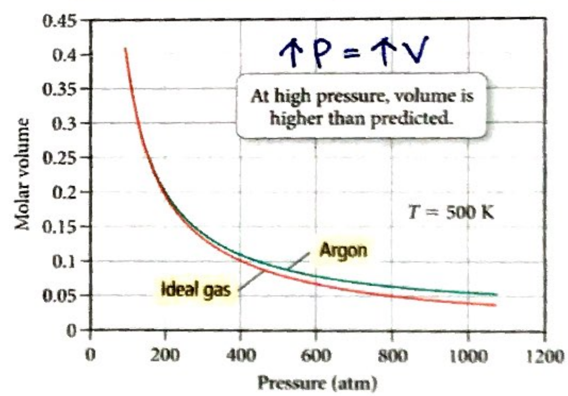
Ideal vs. Real Gases

1. **Ideal Gas**: an imaginary gas that perfectly fits all the assumptions of the kinetic-molecular theory.
2. **Real Gas**: a gas that does NOT behave completely according to the assumptions of the kinetic-molecular theory.
    - The more polar (more IMFs) the gas molecules = greater the deviation from ideal behavior.
        - Pressure will be lower than predicted because of intermolecular attractions.
    - At low temperatures, gas particles have insufficient kinetic energy to overcome attractions (IMFs).
        - Pressure will be lower than predicted because of intermolecular attractions.
    - At high pressures, distance between particles in likely to be small relative to the size of the particles.
        - Volume will be higher than predicted because of non-negligible particle size.

Nonideal Behavior: The effect of intermolecular forces



Nonideal Behavior: The effect of particle volume



Gases behave most like ideal gases when they have:

1. **High Temperature** – because KE overcomes IMFs
2. **Low Pressure** – because gas particles are relatively far apart (molecular volume is low relative to container volume)
3. **Small Intermolecular Forces** (IMFs):
    a. Small particles (low London dispersion forces)
    b. Non-polar (no dipole-dipole IMFs)

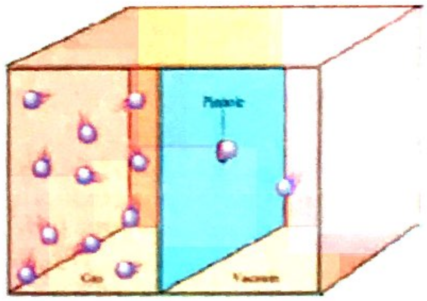
Interpreting Data from Non-Ideal Gases

Data	What happened?	Because Ideal Gas doesn't account for:
↓P (than predicted by KMT)	↓ T or ↑ IMFs (or both!)	Attractive forces between particles
↓V (than predicted by KMT)	↓ T or ↑ IMFs (or both!)	Attractive forces between particles
↑V (than predicted by KMT)	↑ P or tiny container	Size of particles

Graham's Law of Diffusion and Effusion

Diffusion: the mixing of gases. The rate of diffusion is the rate of the mixing.

Effusion: the passage of gas through a tiny hole into an evacuated chamber (as shown below). The rate of effusion measures the speed at which the gas is transferred into the chamber.



- The rates of diffusion and effusion of a gas are related to the average Velocity of its particles.
- For gases at the same temperature, this means the rate of gas movement is inversely proportional to the square root of its molar mass.

$$rate\ of\ gas\ movement \propto \frac{1}{\sqrt{MM}}$$

→ Lighter gases (lower molar mass) diffuse and effuse more rapidly.

Note: "heavier" ≠ "bigger". Be very careful about word choice when answering FR problems!

Let's Practice!

1. Consider two 1.0 L balloons at STP: one is filled with helium gas and the other with carbon dioxide gas.
   a. Do the atoms in the helium sample have the same average kinetic energy as the atoms in the carbon dioxide sample? Justify your answer.

   Yes, b/c average KE is the same at the same temp!

   b. Do the atoms in the helium sample have the same average velocity as the atoms in the carbon dioxide sample? Justify.

   Nope. He atoms have a lower molar mass than CO_2 molec., so at the same temp, He atoms will have a greater average velocity than CO_2, b/c $KE = \frac{1}{2}mv^2$.

   c. Which balloon will deflate faster? Explain.

   The He balloon will deflate first - the lighter He atoms will effuse out of the balloon more quickly than the heavier CO_2 molecules at the same temperature, b/c $KE = \frac{1}{2}mv^2$.