Announcements – 10/6/00

- ***Quiz Today***
- ***Demo today!***

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Gas Mixtures

- Dalton's Law of Partial Pressures:

  *The **TOTAL PRESSURE** of a gas mixture is the **SUM of the pressures** exerted by each of the gases in the mixture.*

  $$P_{total} = P_1 + P_2 + P_3 + \ldots\ldots$$

At constant Temp and Volume:

$P_1 = n_1 (RT/V) \rightarrow P_1 \propto n_1$

- partial pressure of a gas in a mixture is proportional to the number of mol of that gas in the mixture

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Partial Pressures: Example

- If dry air is 21.0% O_2, 78.0% N_2 and 1.0% Ar (wt%), what is the *partial pressure* of O_2 in a dry air sample at **1.00 atm** total pressure?

$$P_{O_2} = X_{O_2} P_{Total}$$

We need X_{O_2}:

$$X_{O_2} = \frac{n_{O_2}}{n_{O_2} + n_{N_2} + n_{Ar}} = \textit{mol fraction of } O_2$$

We need mol of each:

$$n_{O_2} = 21.0 \text{ g } O_2 \times \frac{1 \text{ mol } O_2}{31.998 \text{ g } O_2} = 0.6563 \text{ mol } O_2 \text{ (in 100 g Air)}$$

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Partial Press Example: Cont'd

- Putting it all together:

$$X_{O_2} = \frac{n_{O_2}}{n_{O_2} + n_{N_2} + n_{Ar}} = \frac{0.6563 \text{ mol}}{0.6563 + 2.784 + 0.02503 \text{ mol}}$$

$$= \textbf{\textit{0.18939}}$$

Finally: $P_{O_2} = X_{O_2} P_{Total}$

$= (0.18939)(1.00 \text{ atm}) = \textbf{\textit{0.18939 atm}}$

In *torr*: $0.18939 \text{ atm} \times \frac{760 \text{ torr}}{1 \text{ atm}} = 143.9 \text{ torr} = \textbf{\textit{144. Torr}}$

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Kinetic Theory of Gases

- Developed in the 1800's by physicists Clausius, Maxwell and Boltzmann
- **Model of Gas Behavior:**
  - Pure gas is ***dilute***
    - atoms/molecules separated by distances >> size
  - Molecular motion is:
    - continuous, random
    - in straight lines between *collisions*
    - defined by a *distribution* of velocities
  - Collisions:
    - are elastic (no energy loss)
    - are the ***only*** interactions
    - change the *direction* of motion

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More Gas Kinetic Theory

- **Implications:**

1. **Temperature**
   - The *average kinetic energy* of the molecules does not change over time (at constant temperature)
   - The *average kinetic energy* of the molecules is proportional to the **absolute temperature**
2. **Pressure**
   - The pressure exerted by a gas is caused by collisions of the gas molecules with the walls of the container
   - The *magnitude* of the pressure depends on collision frequency and force

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Gas Kinetic Theory: Qualitatively

- Can explain empirical observations:
  - Boyle's Law (PV = constant)
    - *Increase volume*: fewer molecules per unit volume, so there are fewer collisions with the container walls (*decreased pressure).*
  - Charles' Law (P/T = constant)
    - *Increase Temp*: molecules move faster, so there are more collisions with the container walls (*increased pressure*).

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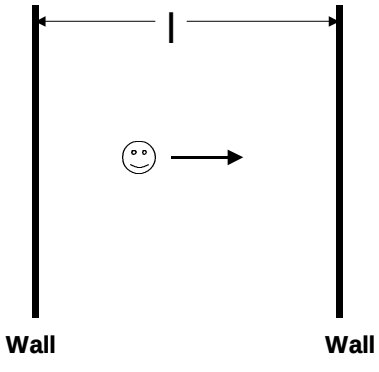
Gas Kinetic Theory: Quantitatively

- Consider a molecule in a cube moving in one direction:



l = *distance between walls*

v = l/t = velocity of molecule

l/v = time to travel across cube

$\Delta t = 2l/v$ = *time to travel from wall to wall round trip*

Momentum of the molecule:

$$p = mv_x$$

Due to collision:

$$\Delta p = 2mv_x$$

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Quantitative Gas Kinetic Theory: Cont'd

- What is the *force* exerted by the molecule on the wall due to the collision?

Force (F) = $\Delta p/\Delta t = 2mv/(2l/v)$

$$\mathbf{F = mv^2/l}$$

- Relating this to *pressure:*

Pressure = force/unit area

$P = F/l^2 = (mv^2/l)/l^2 = mv^2/l^3$

So: $\boxed{\mathbf{P = \frac{mv^2}{V}}}$ ***Boyle's Law!***

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Now, apply to an *aggregate* of gas particles

- For N_0 molecules moving in 3-dimensional space, it can be shown that:

$$\mathbf{PV = 1/3\ N_0 m\langle v^2\rangle}$$

where $\langle v^2\rangle$ is the *mean squared speed* of the aggregate of molecules

-Again, we've got **Boyle's Law** (PV = constant) at constant temperature and for a fixed number of gas molecules

But, what about ***temperature***?

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Kinetic Energy

- Recall that kinetic energy is: $1/2\ mv^2$
- *Rearranging* the previous equation *and substituting* in the kinetic energy:

$$PV = 2/3\ N_0\langle\varepsilon_k\rangle$$

Where $\langle\varepsilon_k\rangle$ is the ***average*** *kinetic energy* ***per particle***

- We also know that $N_0\langle e_k\rangle$ is the <u>*total kinetic energy*</u> (E_k) for a mole of gas molecules, so:

$$E_k = 3/2\ PV$$

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Kinetic Energy and Temperature

- Recalling the IGL, we can substitute *nRT* for *PV:*

$$E_k = 3/2\ nRT$$

-for a fixed amount of gas, *temperature* is determined ***solely*** by the <u>total kinetic energy</u> of the gas molecules

-Temp is independent of the *mass* or *density* or of the ***composition*** of the gas

-the *velocities* of gas molecules must then be dependent upon their masses

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Temperature and Velocity

- Bringing back our initial result:

$$PV = 1/3\ N_0 m\langle v^2 \rangle$$

- And recognizing that: $\mathbf{N_0 m = M}$ (molar mass)
- And that for a mole of gas: **PV = RT**
- We obtain: $\mathbf{1/3\ M\langle v^2 \rangle = RT}$
- Rearranging:

$$\langle v^2 \rangle = 3RT/M$$

- And taking the square root of both sides:

$$\boxed{v_{rms} = (3RT/M)^{1/2}}$$

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rms Velocity and Molar Mass

- Knowing that:

$$v_{rms} = (3RT/M)^{1/2}$$

Thus, the rms velocity of a gas at a constant temperature must vary with gas composition.

For example: N_2 at 25.0 °C

$$v_{rms} = (3(8.3145\ J/mol\text{-}K)(298.15\ K)/28.014 \times 10^{-3}\ kg/mol)^{1/2}$$

$$= \mathbf{515\ m/s}$$

Other examples: **Helium** (4.003 g/mol) *@ 1360 m/s*

Sulfur Hexafluoride (146 g/mol) *@ 226 m/s*

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