Chemistry 102 Summary June 19th

Kinetic Molecular Theory (KMT)
(a model that attempts to explain ideal gas behavior)

1. Gases are mostly empty space; the volume of the particles is negligible.
2. Gas particles are in constant random motion.
3. Gas particles neither attract nor repel each other.
4. Pressure is due to collisions of gas particles with container walls.
5. The average kinetic energy of a gas sample is proportional to the Kelvin temperature. \( KE_{\text{AVE}} = \frac{3}{2} RT \) where \( R = 8.3145 \text{ J/molK} \)

Discussion of KMT

- postulates 1 and 3 are important in order to explain why all gases follow ideal gas behavior. If gas molecule size and intermolecular forces were important, then each gas would have a different value of \( R \).

- **Kinetic Energy** is the energy due to motion \( (KE = \frac{1}{2}mv^2) \), where \( m \) is the average mass and \( v \) is the average velocity.

- We talk about average kinetic energy because a given gas sample has a distribution of kinetic energies, not just one value. See Figure 5.20. The reason being there is a distribution of velocities for any gas sample at a given temperature.

**Question:** Why is kinetic energy directly proportional to temperature?

**Answer:**

**Question:** What is the only variable we need in order to determine kinetic energy?

**Answer:**
**Question:** You have two containers at the same temperature and pressure. Container #1 has one mole of $O_2$ gas, and container #2 has one mole of hydrogen gas. How do each compare:

(a) Average kinetic energy?

(b) Velocities?

(c) Number of collisions per given area of container per second? How can the two containers be at the same pressure?

**Graham’s Law of Effusion:** found that the average velocity of a molecule is inversely proportional to the square root of the molar mass ($M$). In other words, the relative rates of effusion of two gases at the same temperature and pressure are given by the inverse ratio of the square roots of the masses of the gas particles.

**Diffusion:** rate that gases mix.

**Effusion:** rate that gases pass through a tiny hole (Fig 5.22).
- Both quantities are directly proportional to the average velocity of the gas.

\[ v = \text{average velocity} = \text{constant} \left( \frac{1}{\sqrt{M}} \right) \]

\[ \text{rate of effusion gas}_1/\text{rate of effusion gas}_2 = v_1/v_2 = \sqrt{M_2}/\sqrt{M_1} \]

- allows for the average velocity of one gas to be compared to another gas at constant temperature.
- we can’t really measure average velocities directly, therefore we use effusion as an experimentally determined value that can be related to average velocity.
Question: We have already determined that the average velocity of $H_2$ is greater than that of $O_2$ at 298 K but by what factor?
Answer:

Question: The rate of effusion of a particular gas was measured and found to be 24.0 mL/min. Under the same conditions, the rate of effusion of pure methane ($CH_4$) gas is 47.8 mL/min. What is the molecular mass of the unknown gas?
Answer:

Question: Figure 5.24 – When $NH_3$ gas molecules come in contact with $HCl$ gas molecules, a white cloud of $NH_4Cl$ (s) will form. On which side of the apparatus will the $NH_4Cl$ form?
Answer:

Real Gas Behavior
- At room temperature, 1 atmosphere pressure, most gases behave ideally.
- At low temperatures and very high pressures most gases do not follow the ideal gas law.
- At low temperatures and high pressures gas molecules have a volume and do exert attractive forces toward each other, i.e. certain postulates of KMT do not hold.
- The van der Waals equation attempts to correct for these assumptions in extreme conditions.

$$[P_{obs} + a(n/V)^2] X (V-nb) = nRT$$
- A factor is added to the measured pressure in order to correct for the attractive forces of the gas molecules.
- Because gas molecules attract one another, then they hit the container walls with a slightly less forceful collision then they would if the gas molecules were not attracted to one another. This results in a measured pressure which is less than the ideal pressure.
- For the volume term, a term is subtracted from the container volume in order to account for the small amount of space taken up by the gas molecules.

**Question:** Why do gases behave most ideally at high temperature and low pressure?