

Name \_\_\_\_\_

# Matter and Phase Changes

## The Kinetic-Molecular Theory

The Kinetic-Molecular Theory (KMT) explains the behavior of gases based on the idea that gas particles are in constant motion. The key assumptions are:

1. Gases consist of large numbers of tiny particles that are far apart relative to their size.
2. Collisions between gas particles and between particles and container walls are elastic collisions.
3. Gas particles are in continuous, rapid, random motion.
4. There are no forces of attraction or repulsion between gas particles.
5. The average kinetic energy of gas particles depends on the temperature of the gas.

## Dalton's Law of Partial Pressure

Dalton's Law states that the total pressure of a mixture of gases is equal to the sum of the partial pressures of the individual gases. This law is also explained by the KMT: each gas in the mixture exerts its own pressure independently, and the total pressure is simply the sum of these individual pressures. Mathematically:  $P_{\text{total}} = P_1 + P_2 + P_3 + \dots + P_n$ , where  $P_{\text{total}}$  are the partial pressures of each gas in the mixture.

## Graham's Law of Diffusion Rates

Graham's Law states that the rate of diffusion of a gas is inversely proportional to the square root of its molar mass. According to Graham's law, lighter gases effuse faster than heavier gases. This can be explained by the KMT: lighter molecules have higher average speeds and therefore collide with the hole more frequently, leading to a faster rate of effusion.

This can be expressed as:

$$\frac{\text{Rate}_1}{\text{Rate}_2} = \sqrt{\frac{M_2}{M_1}}$$

where Rate<sub>1</sub> and Rate<sub>2</sub> are the diffusion rates of gases 1 and 2, and M<sub>1</sub> and M<sub>2</sub> are their molar masses.



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