1. Properties of gases.
4. Ideal gas law.
5. Dalton’s law of partial pressures.
1. Gases are mostly empty space; the volume of the particles is negligible.

1 mole of $\text{N}_2(\text{l})$ has a volume of 35 mL and a density of 0.81 g/mL

1 mole of $\text{N}_2(\text{g})$ has a volume of 22.4 L and a density of $1.2 \times 10^{-3}$ g/mL

2. Gases particles are in constant random motion. The collision of the particles with the walls of the container are the cause of the pressure exerted by the gas.
Kinetic Molecular Theory

3. Gas particles neither attract nor repel each other.

4. The average kinetic energy of a collection of gas particles is assumed to be directly proportional to the Kelvin temperature of the gas. $KE_{\text{AVE}} = \frac{3}{2}RT$, where $R = 8.3145 \text{ J/molK}$

How does kinetic molecular theory explain why gases behave according to the ideal gas law?
Kinetic energy – is the energy due to motion

\[
KE_{\text{AVE}} = \frac{1}{2}mv^2, \text{ where } m = \text{average mass, and } v = \text{average velocity}
\]

Why do we talk about AVERAGE kinetic energy?

A given gas sample has a distribution of kinetic energies, not just one value. The reason being there is a distribution of velocities for any gas sample at a given temperature.
Why is kinetic energy directly proportional to temperature?

As temperature increases, the velocity distribution shifts to higher and higher velocities. Since average velocity increases, so does the average kinetic energy. \( KE_{ave} = \frac{1}{2}mv^2 \)
What is the only variable we need in order to determine kinetic energy?

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 $H_2$ gas. How do each compare?

(a) Average kinetic energy?

(b) Velocities?
Kinetic Energy

You have two containers at the same temperature and pressure. Container #1 has one mole of O₂ gas, and container #2 has one mole of H₂ gas. How do each compare?

(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

**Effusion** – rate that a gas passes through a tiny pin hole.

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.

\[
\frac{\text{Rate of effusion for gas 1}}{\text{Rate of effusion for gas 2}} = \frac{\sqrt{M_2}}{\sqrt{M_1}}
\]
Graham’s Law of Effusion

We have already determined that the average velocity of $\text{H}_2$ is greater than that of $\text{O}_2$ at 298 K but by what factor?
Graham’s Law of Effusion

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\textsubscript{4}) gas is 47.8 mL/min. What is the molar mass of the unknown gas?
Graham’s Law of Effusion

What will happen if two cotton plugs, one dipped in \( \text{HCl} (aq) \) and one dipped in \( \text{NH}_3 (aq) \), are simultaneously inserted into the ends of a tube if:

a) Graham’s Law of Effusion holds true?

b) In reality?
Real Gas Behavior

At room temperature and one atmosphere pressure, most gases behave ideally.

At low temperatures and high pressures most gases do not follow the ideal gas law. WHY?
van der Waals Equation

\[ P_{\text{obs}} = \frac{nRT}{V-nb} - a(n/V)^2 \]

1. A factor \((a(n/V)^2)\) is added to the measured pressure in order to correct for the attractive forces of the gas molecules.

2. A term \((nb)\) is subtracted from the container volume in order to account for the small amount of space taken up by the gas molecules.
Real Gas Behavior

Why do gases behave most ideally at high temperature and low pressure?
van der Waals Equation

Without looking at tables of values, which of the following gases would you expect to have the largest value of the van der Waals constant b: H₂, N₂, CH₄, C₂H₆, or C₃H₈?