Gaseous State

Introduction

The gaseous state is one of the fundamental states of matter, characterized by weak intermolecular forces and high kinetic energy. Unlike solids and liquids, gases have no fixed shape or volume and expand to fill the container they are in. This lecture will cover the basic gas laws, kinetic molecular theory, real gas behavior, and important equations governing gases.

Properties of Gases

• Gases are highly compressible.
• They have low densities compared to solids and liquids.
• Gases exert uniform pressure in all directions.
• They expand to fill any container.

Ideal Gas Laws

The behavior of gases is described using several fundamental laws:

Boyle's Law (Pressure-Volume Relationship)

$PV = k$

At constant temperature, the volume of a gas is inversely proportional to its pressure.

$V \propto 1/P$

Charles's Law (Temperature-Volume Relationship)

$V \propto T$

At constant pressure, the volume of a gas is directly proportional to its absolute temperature.

Avogadro's Law (Volume-Mole Relationship)

$V \propto n$

At constant temperature and pressure, the volume of a gas is directly proportional to the number of moles of the gas.

Combined Gas Law

Combining Boyle's, Charles's, and Avogadro's laws:

$PV = nRT$

where:

- P = pressure (atm or Pa)
- V = volume (L or m^3)
- n = number of moles
- R = universal gas constant
- T = temperature (Kelvin)

Dalton's Law of Partial Pressures

The total pressure exerted by a mixture of gases is equal to the sum of the partial pressures of individual gases:

$P_{total} = P_1 + P_2 + P_3 + ...$

Kinetic Molecular Theory (KMT)

This theory explains gas behavior at the molecular level:

1. Gas molecules move randomly and continuously.
2. Collisions between gas molecules and container walls are elastic.

3. The volume of individual gas molecules is negligible compared to the container volume.
4. There are no intermolecular forces between gas molecules.
5. The average kinetic energy of gas molecules is proportional to absolute temperature:
$KE_{avg} = (3/2) k_B T$
where k_B is the Boltzmann constant.

Maxwell-Boltzmann Distribution

Describes the distribution of molecular speeds in a gas:
• Most molecules have a speed close to the most probable speed $v_{mp} = \sqrt{2RT/M}$
• The average speed $v_{avg} = \sqrt{8RT/\pi M}$
• The root mean square speed $v_{rms} = \sqrt{3RT/M}$

Real Gases and Deviations from Ideal Behavior

Van der Waals Equation
Ideal gases assume no intermolecular forces, but real gases deviate due to molecular interactions. The Van der Waals equation corrects for these:
$(P + a/V^2)(V - b) = RT$
where:
- a accounts for intermolecular attractions.
- b accounts for finite molecular volume.

Compressibility Factor (Z)

Defines how much a real gas deviates from ideal behavior: $Z = PV/nRT$

For an ideal gas, $Z = 1$; for real gases, $Z \neq 1$.

Diffusion and Effusion

- Graham's Law of Diffusion and Effusion: $r_1 / r_2 = \sqrt{(M_2 / M_1)}$

where r is the rate of diffusion and M is molar mass.

Applications of Gaseous State

- Industrial gas storage (LPG, CNG, oxygen tanks)
- Weather balloons and atmospheric pressure studies
- Respiratory physiology and anesthesia gases

Conclusion

Understanding the gaseous state is crucial for explaining thermodynamics, reaction kinetics, and real-world applications. The ideal gas laws, kinetic molecular theory, and real gas behavior form the foundation for further study in chemistry and physics.

Questions for Self-Assessment

1. How does temperature affect the kinetic energy of gas molecules?
2. Explain why real gases deviate from ideal behavior at high pressures.
3. Derive the root mean square velocity formula.
4. Compare and contrast Boyle's and Charles's laws.
5. How does Van der Waals' equation improve upon the ideal gas law?