

# Ideal and Real Gases

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## Introduction

To master the foundational principles of Physical Chemistry, a clear understanding of the gaseous state is essential. This module focuses on:

- Establishing the critical distinction between the theoretical behavior of the Ideal Gas and the actual, non-ideal behavior of Real Gases.
- Utilizing the Ideal Gas model as a simplified tool for straightforward calculations.
- Applying necessary corrections via the Real Gas model to accurately predict gas properties under all conditions, especially those encountered in industrial and advanced research settings.

## Learning Objectives

By the end of this module, you will be able to:

- Apply the general gas equation(  $PV=nRT$ ) to solve for the state variables of an ideal gas in various conditions.
- Apply Dalton's law to determine the partial pressures and composition of gas mixtures.
- Explain the behavior of real gases, specifically the sources of their deviation from the ideal model, using the Compressibility Factor ( $Z$ ).
- State and apply the Van der Waals equation to model the behavior of real gases.

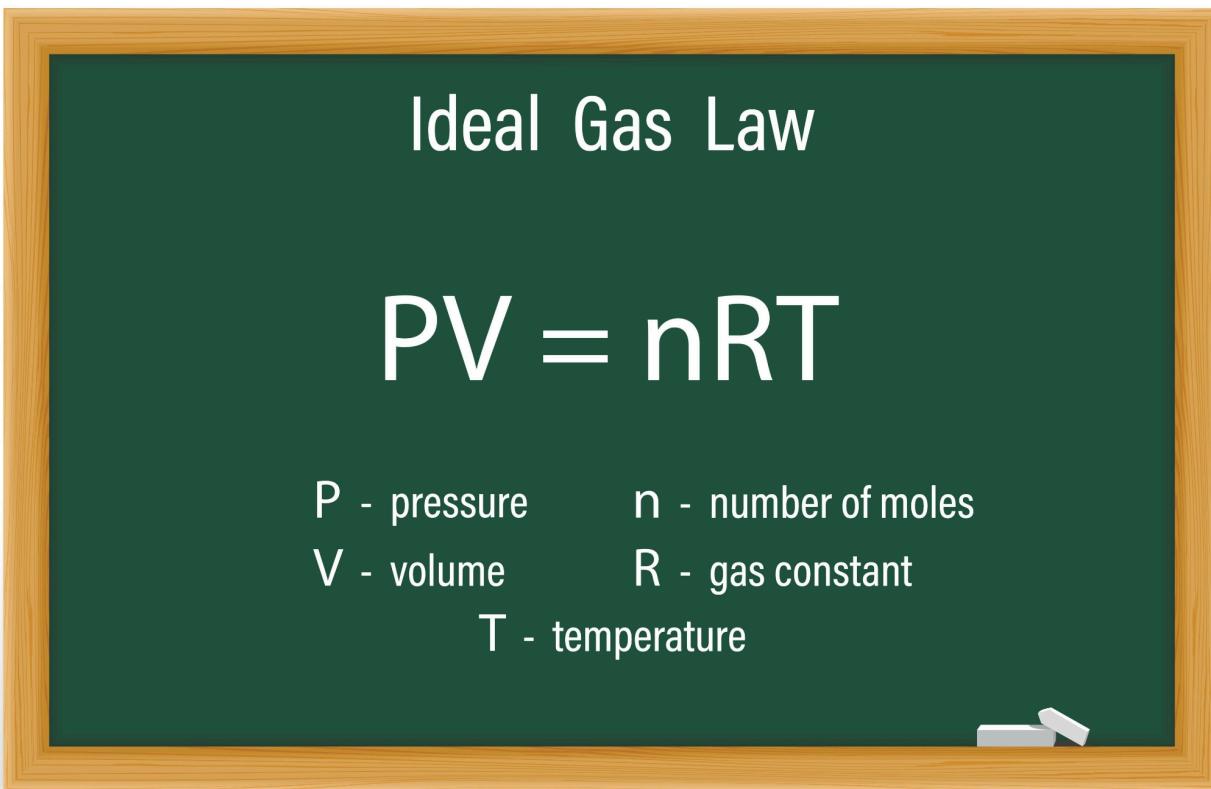
## Key Concepts and Definitions

Term	Definition
Ideal Gas	A hypothetical gas whose molecules exhibit no intermolecular forces and occupy a negligible volume. It perfectly obeys the Ideal Gas Law.
Real Gas	An actual gas whose molecules occupy finite volume and experience attractive/repulsive intermolecular forces, leading to deviation from ideal behavior.
Partial Pressure	The pressure exerted by an individual component gas in a mixture of non-reacting gases.
Compressibility Factor (Z)	A factor used to quantify the deviation from ideal behavior, calculated as $Z = (P * V) / (n * R * T)$ . For an ideal gas, $Z = 1$ .
Boyle Temperature	The specific temperature at which a real gas behaves ideally over a wide range of pressures because the effects of attractive and repulsive forces effectively cancel each other out. $T_B = a / (R * b)$ .

## Detailed Discussion

### 1. The Ideal Gas Law and Derived Laws

The **Ideal Gas Law** is the combined expression of Boyle's, Charles', and Avogadro's Laws, serving as the fundamental equation of state for ideal gases. It relates the macroscopic properties of pressure (**P**), volume (**V**), temperature (**T**), and the amount of gas (**n**) through the Ideal Gas Constant (**R**):



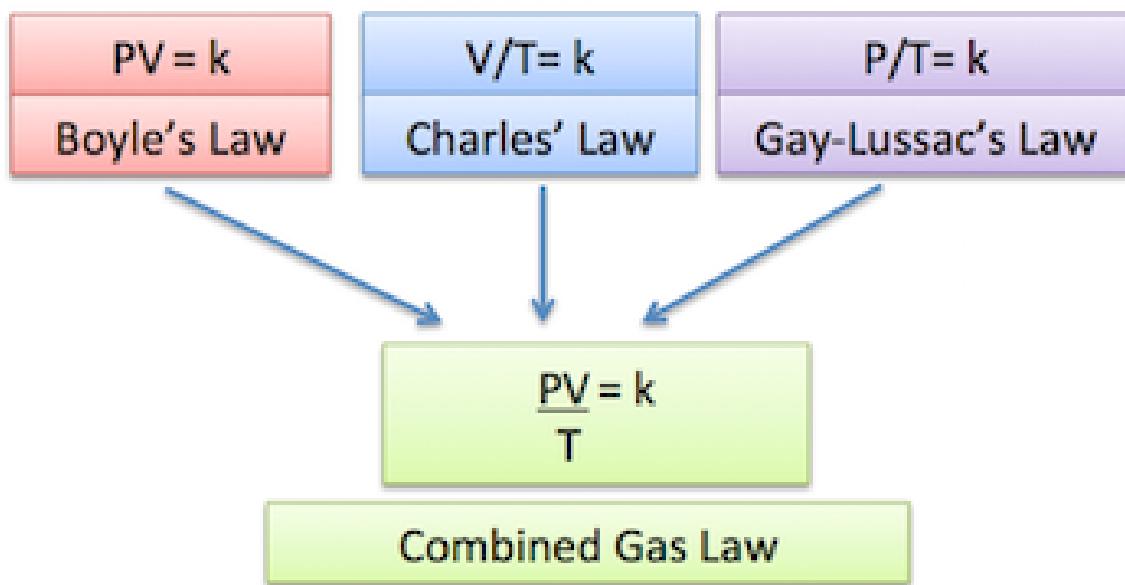
**Ideal Gas Law**

$$PV = nRT$$

P - pressure      n - number of moles  
V - volume      R - gas constant  
T - temperature

The Ideal Gas Constant (**R**) is commonly used in values such as **0.08206 L \* atm / (mol \* K)** or **8.314 J / (mol \* K)**.

The **Combined Gas Law** is used when a fixed amount of gas (constant  $n$ ) undergoes a change from an initial state to a final state. This formula combines *Boyle's*, *Charles'*, and *Gay-Lussac's laws*:



## 2. Mixtures of Gases

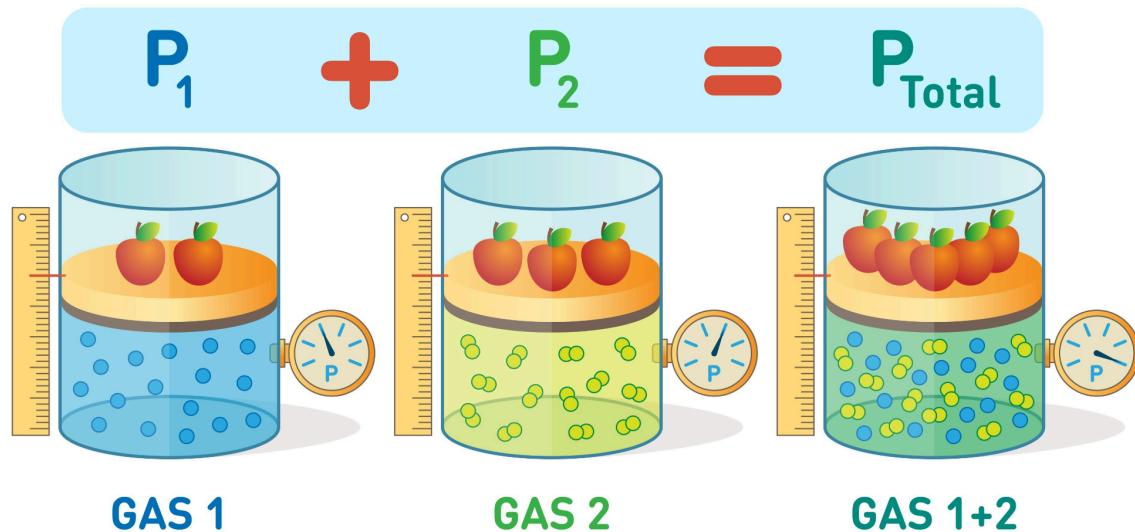
**Dalton's Law of Partial Pressures** states that the total pressure ( $P_{\text{total}}$ ) exerted by a mixture of non-reacting gases is simply the sum of the partial pressures ( $P_i$ ) that each individual gas would exert if it were alone in the container. This is a critical concept for atmospheric and industrial gas processes:

**Dalton's Law:**  $P_{\text{total}} = P_1 + P_2 + \dots + P_i$

The partial pressure of any component gas ( $P_i$ ) can also be calculated using its **mole fraction** ( $x_i$ ) within the mixture, which defines its contribution to the total pressure:

**Partial Pressure (Mole Fraction):**  $P_i = x_i * P_{\text{total}}$

## DALTON'S LAW OF PARTIAL PRESSURE



### 3. Kinetic Molecular Theory (KMT)

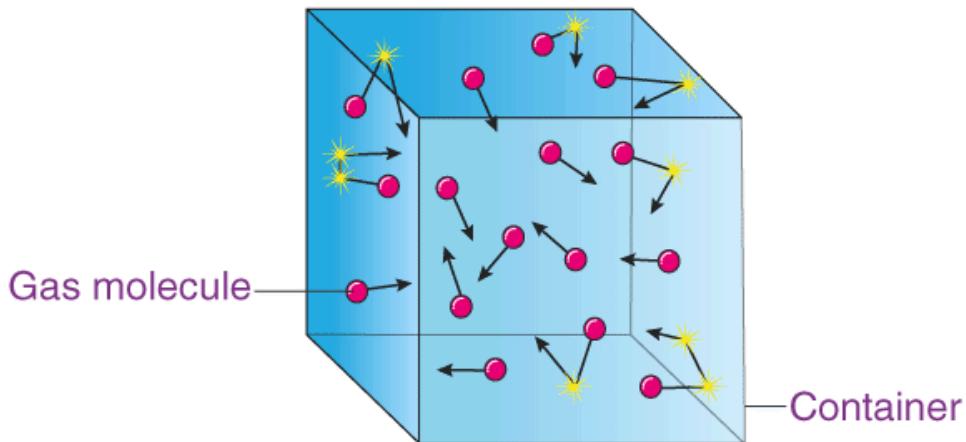
KMT provides the theoretical foundation for the Ideal Gas Law based on several postulates, primarily that gas particles are in continuous, random motion and collisions are perfectly elastic. A key result of KMT is the relationship between temperature and energy: the **average kinetic energy** ( $KE_{avg}$ ) of a mole of gas is directly proportional to the absolute temperature ( $T$ ).

**Average Kinetic Energy:**  $KE_{avg} = \frac{3}{2} * R * T$  per mole

The speeds of gas molecules are described by the Maxwell-Boltzmann distribution. This distribution shows that gas molecules move at a wide range of speeds that changes with temperature.

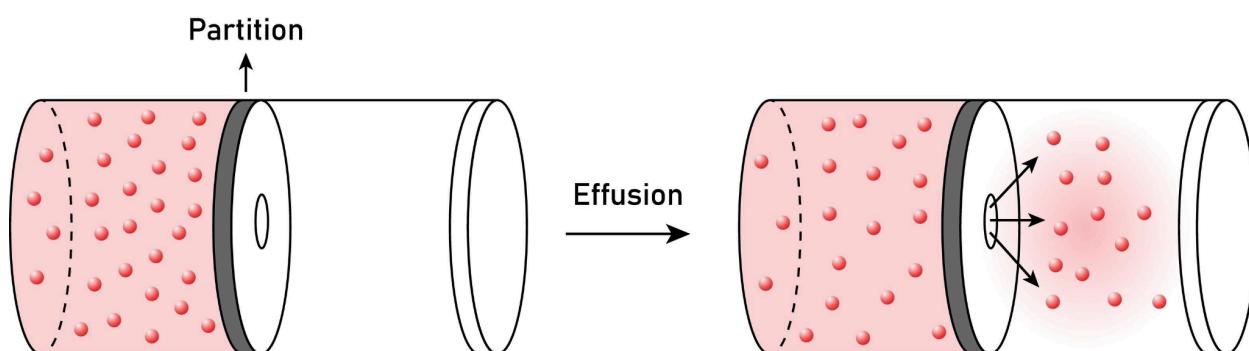
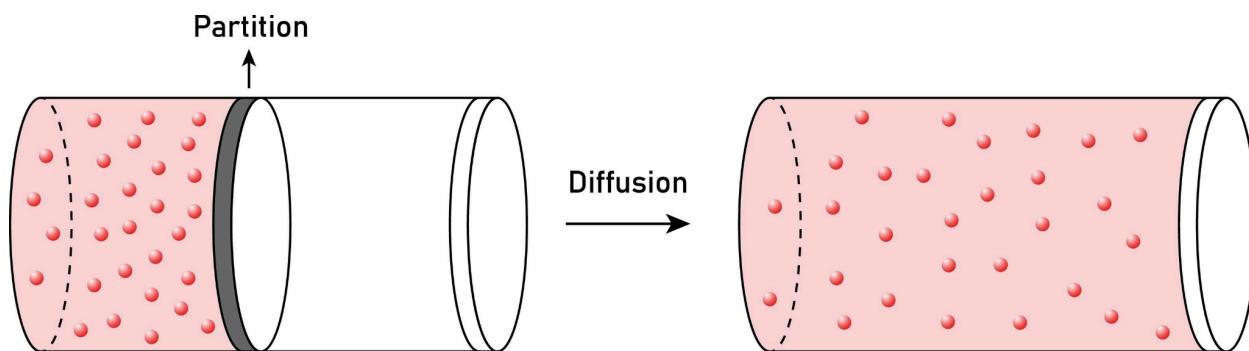
The **Root Mean Square (rms) Speed** ( $u_{rms}$ ) is a common measure of molecular velocity, where  $M$  is the molar mass of the gas.

**Root Mean Square Speed:**  $u_{rms} = \sqrt{3 * R * T / M}$



#### 4. Gas Movement: Graham's Law

**Graham's Law** governs the rate of **effusion** (escape through a small hole) and **diffusion** (mixing). It states that the rate of movement is inversely proportional to the square root of the molar mass (M). This means lighter gases move and mix faster than heavier gases.

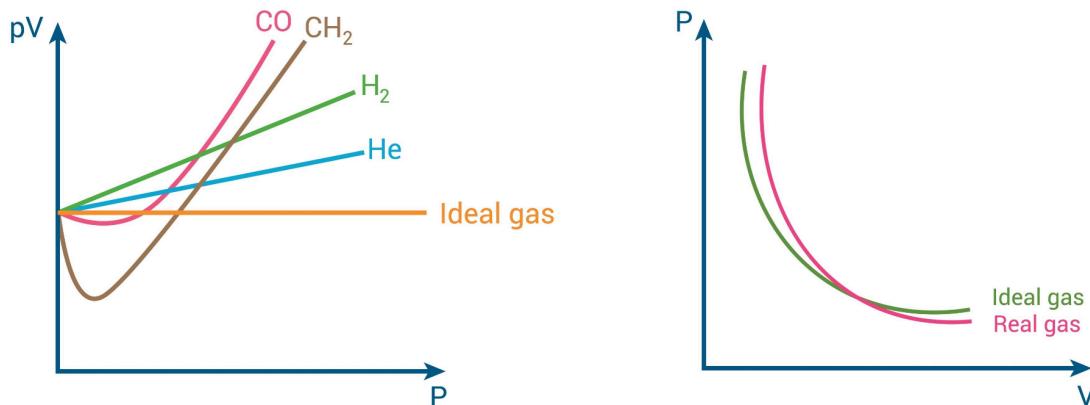


## 5. Real Gas Deviation and the Compressibility Factor

Real gases deviate from ideal behavior because the core assumptions of KMT fail under conditions of high pressure or low temperature.

- Finite Molecular Volume (Repulsion):** At high pressure, the volume occupied by the gas molecules themselves is no longer negligible compared to the container volume. This physical repulsion leads to a larger measured PV product than predicted. (Repulsive forces dominate,  $Z > 1$ ).
- Intermolecular Forces (Attraction):** At low temperature or moderate pressure, attractive forces between molecules reduce the force of their collisions with the container walls, leading to a smaller measured pressure. (Attractive forces dominate,  $Z < 1$ ).

The **Compressibility Factor** ( $Z = PV / nRT$ ) quantifies this deviation. Ideal behavior is  $Z=1$ .



## 6. The Van der Waals Equation

The **Van der Waals equation** is the most significant empirical correction to the Ideal Gas Law for real gases, introducing two constants characteristic of each gas to account for the physical limitations of real molecules:

- **Correction for Attraction (a):** The term ( $a * n^2 / V^2$ ) is added to the measured pressure (P). The constant  $a$  accounts for the intermolecular attractive forces that effectively reduce the observed pressure.
- **Correction for Volume (b):** The term ( $n * b$ ) is subtracted from the measured volume (V). The constant  $b$  accounts for the finite, non-zero volume occupied by the molecules themselves.

### Van der Waals Equation

$$(P + \frac{an^2}{V^2})(V - nb) = nRT$$

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