



Al-Mustaqbal University
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Second Stage



جامعة المستنقب
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Lecture (7 -8)

Molecular Energy: Kinetic vs. Potential

Second stage

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1. Molecular Energy: Kinetic vs. Potential

In gases, molecules possess **two main types of energy**, each playing a distinct role in gas behavior:

A. Kinetic Energy

- **Definition:** Energy resulting from the **motion of molecules** (translational, rotational, vibrational).
- **Characteristics:**
 - Directly related to **absolute temperature (T)**: higher T → faster molecules → higher kinetic energy.
 - Represents the majority of energy in an ideal gas.
 - Formula for a single molecule in 3D:

$$E_k = \frac{3}{2} k_B T$$

where k_B is Boltzmann's constant.

- **Effect on gas behavior:**
 - Gas pressure on container walls arises from collisions of fast-moving molecules.
 - Diffusion and thermal conductivity depend on kinetic energy.

B. Potential Energy

- **Definition:** Energy due to **intermolecular interactions**, such as:
 - Van der Waals forces (weak attraction between molecules)
 - Hydrogen bonds (in polar gases)
- **Characteristics:**
 - In **ideal gases**: nearly negligible because molecules are assumed non-interacting. → Most energy is kinetic.
 - In **real gases**:
 - Becomes significant at **high pressures** (molecules very close) or **low temperatures** (attractive forces dominate).
 - Influences thermal properties, e.g.:
 - Deviations from ideal gas law
 - Reduction in internal energy available for free motion
 - Boiling point and compressibility



❖ Detailed Comparison

Property	Kinetic Energy	Potential Energy
Source	Molecular motion	Intermolecular forces
Depends on	Temperature (T)	Gas density and interaction strength
Importance in ideal gas	Very high	Negligible
Importance in real gas	High	Significant at high P or low T
Effect on gas behavior	Pressure, diffusion, heat capacity	Deviations from ideal gas laws, phase changes

Illustrative Example:

- **Ideal gas (e.g., He at 300K):** Energy is almost entirely kinetic; molecules rarely interact → potential energy ≈ 0 .
- **Real gas (e.g., CO₂ at 200 atm and 250K):** Molecules are close; attractive forces are significant → potential energy contributes substantially to internal energy.

2. Evaporation from the Molecular Speed Perspective

A. Molecular Motion in Liquids

- Molecules in a liquid **do not move at the same speed**; their velocities follow a **statistical distribution** (similar to Maxwell–Boltzmann in gases).
- Most molecules have speeds near the average, but some move **much faster** than the average.

B. Mechanism of Evaporation

1. Escape of fast molecules:

- Only molecules with **sufficient kinetic energy** can overcome the **intermolecular attractive forces** at the liquid surface.
- These high-speed molecules escape into the vapor phase → **evaporation occurs**.

2. Reduction in average kinetic energy:

- Loss of the fastest molecules reduces the **average kinetic energy** of the remaining liquid molecules.
- Since **temperature is proportional to average kinetic energy**, the liquid **cools down**.



- This is the basis of **evaporative cooling**.

$$m \langle v^2 \rangle \frac{1}{2} = \langle KE \rangle \propto \text{Temperature}$$

C. Factors Affecting Evaporation

- **Temperature of the liquid:** higher $T \rightarrow$ more molecules exceed escape energy \rightarrow faster evaporation.
- **Surface area:** larger surface \rightarrow more molecules can escape per unit time.
- **Intermolecular forces:** weaker forces \rightarrow easier evaporation.
- **Vapor pressure of surroundings:** low surrounding vapor pressure \rightarrow faster evaporation (less opposition).

D. Practical Example: Sweating

- Sweat cools the skin because **water molecules with highest kinetic energy evaporate**, taking away energy from the skin \rightarrow **temperature drops locally**.
- This principle is widely used in **natural cooling processes, air conditioning, and industrial drying**.

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1. Not all molecules have the same speed; only the fastest escape.
2. Evaporation lowers the **average kinetic energy** \rightarrow lowers temperature.
3. This is a **molecular-level explanation of cooling by evaporation**.

3. Molecular Speeds and Diffusion

A. Relationship Between Molecular Speed and Diffusion

- **Diffusion** is the process by which molecules spread from a region of higher concentration to a region of lower concentration.
- The **rate of diffusion** depends on the **molecular speed**, which is related to the molecule's **kinetic energy**:

$$\sqrt{\frac{3k_B T}{m}} = \text{rms } v$$



- **Implication:** lighter molecules (smaller mass m) move faster \rightarrow diffuse more quickly than heavier molecules at the same temperature.

B. Graham's Law of Diffusion

Graham's Law quantitatively relates the diffusion rates of two gases to their molar masses:

$$\sqrt{\frac{r_2 M_1}{r_1 M_2}} = \frac{r_1}{r_2}$$

Where:

- r_1, r_2 = diffusion rates of gas 1 and 2
- M_1, M_2 = molar masses of gas 1 and 2

Key points:

- Lighter gases (H_2, He) diffuse faster than heavier gases (CO_2, O_2).
- This is because lighter molecules have higher **root-mean-square speeds** at the same temperature.

C. Molecular Explanation

1. **Kinetic energy distribution:** All molecules at a given temperature have approximately the same average kinetic energy:

$$T \propto \frac{1}{2} m v_{avg}^2$$

2. **Effect of mass:** For a lighter molecule (smaller m), the speed v must be larger to have the same kinetic energy.
3. **Result:** Lighter molecules travel further in the same time \rightarrow faster diffusion.

D. Practical Applications

1. **Oxygen diffusion in alveoli:**
 - Oxygen (O_2) is lighter than carbon dioxide (CO_2) \rightarrow diffuses rapidly into blood capillaries.
 - Understanding diffusion rates is critical for **efficient gas exchange in the lungs**.



2. Design of medical ventilation systems:

- Ventilation efficiency depends on **how fast gases spread and mix**.
- Lighter gases reach desired locations faster → important in anesthesia delivery and respiratory therapy.

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- **Higher molecular speed → faster diffusion.**
- **Graham's Law links diffusion rates to molar masses: lighter gases diffuse faster.**
- **Essential in respiratory physiology, medical gas delivery, and industrial gas transport.**

Molecular Collisions and Their Types

1. Molecular Collisions

- In gases, molecules move at high speeds in random paths, leading to continuous collisions with each other and the container walls.
- These collisions are the basis of gas behavior, including pressure, diffusion, and temperature.

2. Types of Collisions

A. Elastic Collisions

- Definition: Collisions in which total kinetic energy is conserved.
- Characteristics:
 - Molecules do not lose kinetic energy; they move with the same energy before and after collision.
 - Momentum is also conserved.
- Importance:
 - Main type assumed in the kinetic theory of gases.
 - Explains ideal gas behavior where collisions do not affect internal energy.

B. Inelastic Collisions

- Definition: Collisions in which part of the kinetic energy is converted into other forms of energy (vibrational, rotational, or internal excitations).
- Characteristics:
 - Molecules lose some translational kinetic energy.
 - Momentum is conserved.



- Importance:
 - Relevant in real gases, especially at high pressure or low temperature.
 - Explains phenomena such as energy dissipation and deviations from ideal gas laws.

3. Effects of Collisions

- Pressure: Each collision with the container wall contributes to pressure. More frequent collisions (higher molecular speeds or density) → higher pressure at constant volume.
- Temperature and Energy Distribution: Elastic collisions maintain temperature (average kinetic energy), while inelastic collisions redistribute energy into rotational or vibrational modes.
- Diffusion and Transport Properties: Collisions determine mean free path, affecting diffusion, viscosity, and thermal conductivity.

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Table	Type	Kinetic Energy	Momentum	Relevance
Elastic	Conserved	Conserved	Ideal gases, kinetic theory	
Inelastic	Partially lost → vibration/rotation	Conserved	Real gases, energy redistribution	