Kinetic Molecular Theory (KMT)

Video Workbook with Dr. B

More guides at
www.breslyn.org

Key ideas about KMT:

- Molecules colliding with walls of a container cause pressure.
- Temperature increases movement and collisions.

KMT is the basis for the Combined and Ideal Gas Laws (and much more).

KMT Assumptions & Limitations

Gas laws Playlist

Try to visualize what the molecules are doing AND what you can observe in real life.

Five Assumptions of KMT

Assumption One: Negligible Particle Volume

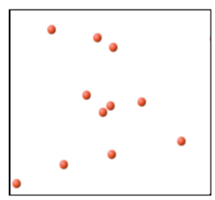
A gas is composed of tiny particles visualized as spheres.

The **volume** occupied by the **molecules** of the gas is **negligible** compared to the volume of the gas**.**

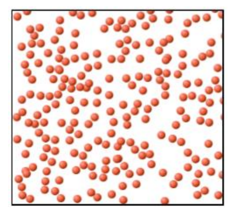
The orange spheres can be atoms (e.g. He, Ar) or molecules (e.g. O_2, CO_2).

Limitations (Very Important)

Under high pressure there are more molecules, so the volume occupied by the molecules matters.



Low pressure, molecule size doesn't matter as much.



Molecules occupy more of the volume.

Assumption Two: Constant, Random Motion

Molecules are always moving in random directions with varying speeds.

Important: the speed of the molecules is also random.

When molecules hit the walls of the container this causes pressure to be exerted.

Limitations

In the real world, collisions between real gases aren't always perfectly elastic.

This is because:

- Intermolecular attractions can slow the movement of molecules and reduce their kinetic energy.
- Kinetic energy can be converted to other forms of energy (vibrational, rotational, or electronic excitation).

Assumption Three: No Attractive Forces

Molecules do not attract or repel each other.

In an ideal gas the molecules just bounce off each other and the walls of the container.

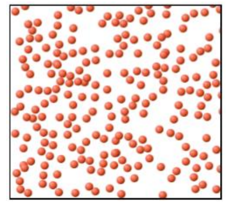
Intermolecular forces like dipole-dipole, dispersion forces, or hydrogen bonding are ignored.

Possible Intermolecular Forces

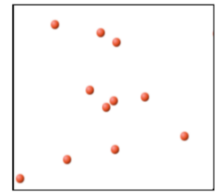
- *London Dispersion Forces*
  (Induced Dipole - Induced Dipole Forces)
- *Dipole-Dipole*
- *Hydrogen Bonding*

Limitations (Very Important)

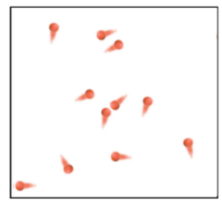
Under high pressures and low temperatures attractions between molecules becomes more likely.



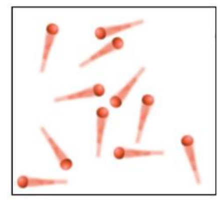
High pressure, molecules are closer and can attract each other.



Low pressure, less opportunities to interact.



Lower temperatures, moving slowly so they have more time to interact.



High temperatures, moving fast so there is less time to interact.

When don't gases follow Assumptions One and Two?

Assumption Four: Perfectly Elastic Collisions

When molecules collide, no kinetic energy is lost.

If the molecules have different speeds, their energy can be transferred.

The collisions with the walls of the container are also considered to be perfectly elastic.

Limitations

In the real world, collisions between real gases aren't always perfectly elastic.

This is because:

- Intermolecular attractions can slow the movement of molecules and reduce their kinetic energy.
- Kinetic energy can be converted to other forms of energy (vibrational, rotational, or electronic excitation).

Assumption Five: Kinetic Energy is Proportional to Temperature	*Limitations*
The average kinetic energy of gas molecules is directly proportional to the temperature of the gas. As the temperature of a gas increases, the average kinetic energy of its molecules also increases. The higher the temperature, the faster molecules move. For temperature we're talking about the **average** kinetic energy of **all** the molecules in a container. They individual speeds will vary.	This only becomes a problem at very high pressures and at very low temperatures. The assumption works well for gases that closely resemble ideal gases, but real gases may exhibit deviations.

Guides

KMT and the Gas Laws (this guide)

Combined Gas Law

Ideal Gas Law

Report errors and suggestions to DrB@breslyn.org

www.Breslyn.org