Kinetic Molecular Theory of GASES
Kinetic Theory is based on two ideas:

1) Particles of matter are always in motion
2) This motion has consequences (Properties)
PROPERTIES OF GASES
1) Expansion (Expands)

No definite shape or volume

2) Fluidity (Fluid)

Gas particles glide easily past one another
3) **Low Density**

The density of a substance in the gaseous state is about $\frac{1}{1000}$ the density of the same substance in the liquid or solid state.
4) Compressibility (Highly Compressible)

During compression of a gas, gas particles are crowded close together.

With sufficient compression, the volume of a gas can be decreased thousands of times.
5) **Diffusion/Effusion**

**Diffusion** = spontaneous mixing of the particles of two substances without being stirred

Rate of diffusion depends on:
- Speed of gas particles  (Higher speed = Faster diffusion)
- Diameter of gas particles (Smaller = Faster Diffusion)
- Attraction Force between gas particles  
  (Small attraction = Faster diffusion)
5) Diffusion/Effusion

**Effusion** = gas particles under pressure pass through a very small opening from one container to another

Rates of effusion is directly proportional to velocity of particles

(Faster particles = faster effusion)
Kinetic Molecular Theory (KMT) explains WHY gases have these properties.

Kinetic Molecular Theory states five assumptions...
Assumption # 1

Gases consist of large numbers of tiny particles that are far apart relative to their size

- Most of the volume occupied by a gas is empty space
Assumption # 2

Particles of a gas are in constant, rapid & random motion therefore possessing Thermal Energy

Thermal energy = energy of random motion
• Assumption # 3

Collisions between particles of a gas and between particles & container walls are **elastic collisions**

- Elastic collisions = no net loss or gain of thermal energy
Assumption # 4

There are no forces of attraction or repulsion between the particles of a gas.

- They do not stick together but immediately bounce off of each other like billiard balls.
• Assumption # 5

The average thermal energy of the particles of a gas depends on the temperature
- If temperature goes up, $E_{th}$ goes up (direct proportion)
  \[ E_{th} = \frac{1}{2} mv^2 \]
  $m =$ mass  \hspace{1cm} $v =$ velocity
- If same gas, mass is the same therefore $E_{th}$ depends on velocity
- With different gases, low mass means higher average speeds
Using Kinetic Molecular Theory to EXPLAIN properties
1) Gases Expand because...
- #2 Particles are in constant, random motion
- #4 There are no attractive forces between particles

2) Gases act like Fluids because...
- #4 There are no attractive forces between particles
3) **Gases have a low density because...**

- #1 Gases consist of tiny particles that are far apart
- #4 There are no attractive forces between particles
4) **Gases are highly compressible because...**

- #1 Gases consist of tiny particles that are far apart

*Fig. 15.1 Compression of a gas by applying pressure*
5) Gases diffuse & effuse because ...

- #2 Particles are in constant, random motion
Ideal Gas = imaginary gas that conforms perfectly to all assumptions of the Kinetic Molecular Theory

Real Gas = a gas that does not completely obey all the assumptions of the Kinetic Molecular Theory (KMT)
Ideal Versus Real Gases

Real gases deviate from ideal gases because...
1) Particles of real gases occupy space
2) Particles of real gases exert attractive forces on each other

Real gases behave like ideal gases when...
1) Particles are very far apart
2) Particles have high thermal energies
3) Particles have a weak attraction to each other
• Measureable quantities of gases:
  1) Volume
  2) Temperature
  3) Pressure
  4) Quantity or number of moles

Describe the volume, temperature & pressure needed for a gas to
a) Act most like an ideal gas
b) Deviate the most from an ideal gas

Ideal Versus Real Gases