Kinetic-Molecular Theory

Unit: Thermodynamics

NGSS Standards/MA Curriculum Frameworks (2016): HS-PS2-8(MA)

AP® Physics 2 Learning Objectives/Essential Knowledge (2024): 9.1.A, 9.1.A.1, 9.1.A.1.i, 9.1.A.ii, 9.1.B, 9.1.B.1, 9.1.B.1.i, 9.1.B.1.ii

Mastery Objective(s): (Students will be able to…)

- Explain how each aspect of Kinetic-Molecular Theory applies to gases.

Success Criteria:

- Descriptions account for behavior at the molecular level.
- Descriptions account for measurable properties, *e.g.,* temperature, pressure, volume, *etc.*

Language Objectives:

- Explain how gas molecules behave and how their behavior relates to properties we can measure.

Tier 2 Vocabulary: kinetic, gas, ideal, real

Notes:

In chemistry you learned about matter, including its composition, structure, and changes that it can undergo. In physics, we are interested in matter to the extent that it can be used to bring objects or energy in contact with each other and transfer forces, energy or momentum from one object or collection of objects to another. This chapter is about gases and using properties of gases to convert between mechanical and thermal energy.

Properties of Different States of Matter

State	Description	Uses
solid	Particles rigidly bonded. Bonds difficult to break. (Definite shape & definite volume)	Construction materials where structure is important. Conduction of heat and/or electricity. Storage of heat as thermal mass.
liquid	Particles loosely bonded and have limited movement. Bonds continuously breaking & reforming. (Definite volume, but indefinite shape.)	Chemical reactions & heat transfer where continual mixing of materials is needed. Storage of heat as thermal mass.
gas	Particles not bonded and able to move freely. (Indefinite shape & volume.)	Heat and materials transfer in large spaces. ***Conversion of energy between heat and mechanical work.***

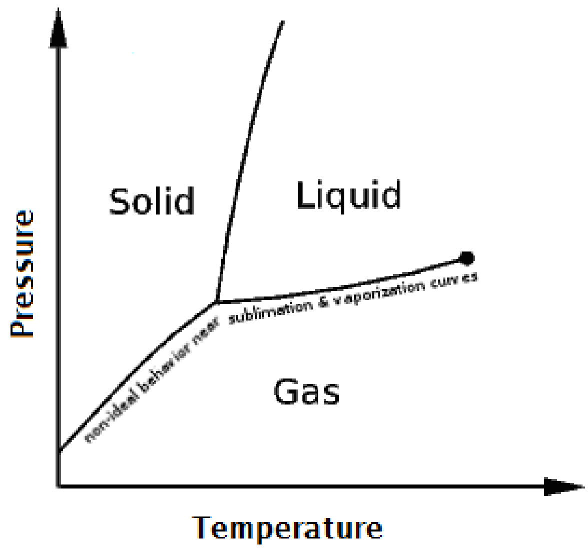
Kinetic-Molecular Theory (KMT)

Kinetic-Molecular Theory (KMT) is a theory, developed by James C. Maxwell and Ludwig Boltzmann, which predicts the behavior of gases by modeling them as moving molecules. (“kinetic” = “moving”.) The theory states that:

- Gases are made of very large numbers of molecules.*
- Molecules are constantly moving (obeying Newton’s laws of motion), and their speeds are constant.
- Molecules are very far apart compared with their diameter.
- Molecules collide with each other and walls of container in elastic collisions.
- Molecules behaving according to KMT are not reacting† or exerting any other forces (attractive or repulsive) on each other.

<u>ideal gas</u>: a gas whose molecules behave according to KMT. Most gases are ideal under *some* conditions (but not all). In general, gases behave ideally when they are not close to the solid or liquid regions of the phase diagram for the substance.

<u>real gas</u>: a gas whose molecules do ***<u>not</u>*** behave according to KMT. This can occur with all gases, most commonly at temperatures and pressures that are close to the solid or liquid regions of the phase diagram for the substance.



* “Particle” is more correct, but the theory is called Kinetic-***<u>Molecular</u>*** Theory. In this chapter we will use the terms “particle” and “molecule” interchangeably, with apologies to chemists.

† Of course, reactions can occur, but chemical reactions are part of collision theory, which is separate from KMT.

Measurable Properties of Gases

All gases have the following properties that can be measured:

Property	Variable	S.I. Unit	Description
amount	*N*	—	amount of gas (particles)
	n	mole (mol)	amount of gas (moles) (1 mol = 6.02×10^{23} particles)
volume	*V*	cubic meter (m^3)	space that the gas takes up
temperature	*T*	kelvin (K)	ability to transfer heat through collisions with other molecules (average kinetic energy of the particles)
pressure	*P*	pascal (Pa)	average force on the walls of the container due to collisions between the molecules and the walls

Notes about calculations:

- Moles are based on the definition that $1\,\text{mole} = 6.022\,140\,76 \times 10^{23}$ particles. 1 mole was originally the number of carbon atoms in exactly 12 grams of carbon-12, such that the molar mass of a substance is the same number of grams as the average atomic mass of one atom in atomic mass units. This definition persisted, despite the fact that the base mass unit of the MKS system is the kilogram.
- Temperature must be absolute, which means you *must* use Kelvin. A temperature of 0 in a gas laws calculation can only mean absolute zero.
- Pressures must be absolute. (For example, you can't use a tire gauge because it measures "gauge pressure," which is the difference between atmospheric pressure and the pressure inside the tire.) A pressure of 0 in a gas laws calculation can only mean that there are no molecules colliding with the walls.

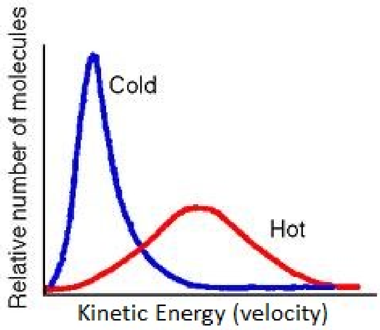
Other Common Units

- **Volume** can be measured in liters (L) or milliliters (mL).
  $1\ m^3 = 1\,000$ L and 1 L = 1 000 mL
- **Pressure** can be measured in many different units.
  1 atm = 101 325 Pa = 14.696 psi = 760 mm Hg = 29.92 in. Hg
  1 bar = 100 000 Pa = 100 kPa ≈ 1 atm

Temperature and Particle Motion

Particles of all substances, whether solids, liquids, or gases, are in constant motion. As described in *Phases & Phase Changes* starting on page 63, the particles of solids and liquids are bonded to one or more other particles. Gas particles, however, are not bonded to other particles. As stated above, these particles move at high speeds in straight lines until they collide with other particles or the walls of the container.

As described in *Heat & Temperature*, starting on page 39, temperature is related to the average kinetic energy of the particles. The kinetic energies follow the Maxwell-Boltzmann probability distribution, which is characterized by a longer tail on the right side. (The equation requires multivariable calculus and is beyond the scope of this course.)



The relationship between the kinetic energy and the average velocity of the molecules is characterized by the equation:

$$K_{avg} = \tfrac{3}{2} k_B T = \tfrac{1}{2} m v_{rms}^2$$

where:

K_{avg} = average kinetic energy of the particles (J)
k_B = Boltzmann constant $= 1.38 \times 10^{-23}\ \tfrac{\text{J}}{\text{K}}$
T = absolute temperature (K)
m = mass of a particle (kg)
v_{rms} = root mean square velocity ($\tfrac{\text{m}}{\text{s}}$)

The "root mean square velocity" simply means that v^2 is an average, which we call the "mean square velocity". *v* is the square root of this quantity, so we call it the "root mean square velocity".

Root mean square velocity is also discussed in *Thermodynamics*, starting on page 109.