The Kinetic-Molecular Theory of Gases

based on the idea that particles are always in motion

Five assumptions:

1. Most of the volume occupied by a gas is empty space

2. Collisions between gas particles and container walls are elastic collisions; no net loss of total kinetic energy

3. Gas particles are in continuous, rapid, random motion.

4. There are no forces of attraction between gas particles.

5. The average kinetic energy of the gas particles depends on the temperature of the gas.
   - At the same temperature, lighter gas particles, have higher average speeds than do heavier gas particles

- The KMT applies only to ideal gases
- An **ideal gas** is a hypothetical gas that perfectly fits all the assumptions of the kinetic-molecular theory

Explaining the Nature of Gases:

- Expansion (gases don't have a definite shape or volume - #3 and 4)
- Fluidity (gas particles glide past one another - #4)
- Low Density (particles are far apart in the gaseous state - #1)
- Compressibility (gas are mostly empty space space - #1)
- Diffusion and Effusion (gases are in constant motion with no forces of attraction - # 3 and 4)
- Exert Pressure (gases undergo elastic collisions - #2)

Conditions where Real Gases Deviate from ideal

- Low temperatures
- High Pressure

Particles are "forced" to interact to become liquids or solids.

Diffusion and Effusion

- Gases spread out and mix with one another, even without being stirred.
- **Diffusion**: spontaneous mixing of two substances caused by their random motion; from a higher concentration to a lower concentration
- **Effusion**: process where the molecules of a gas confined in a container randomly pass through a tiny opening in the container

Graham's Law

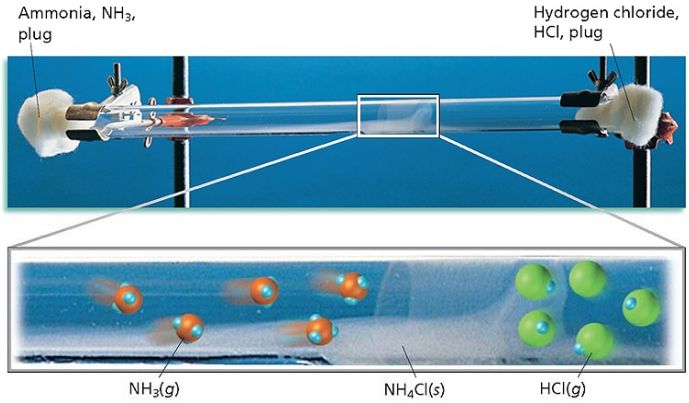
- Rate of effusion and diffusion of gases at the same temperature and pressure are inversely proportional to the square roots of their molar masses

$$\frac{\text{rate of effusion of } A}{\text{rate of effusion of } B} = \frac{\sqrt{M_B}}{\sqrt{M_A}}$$

- Molecules of low mass effuse faster than molecules of high mass.

Graham's Law

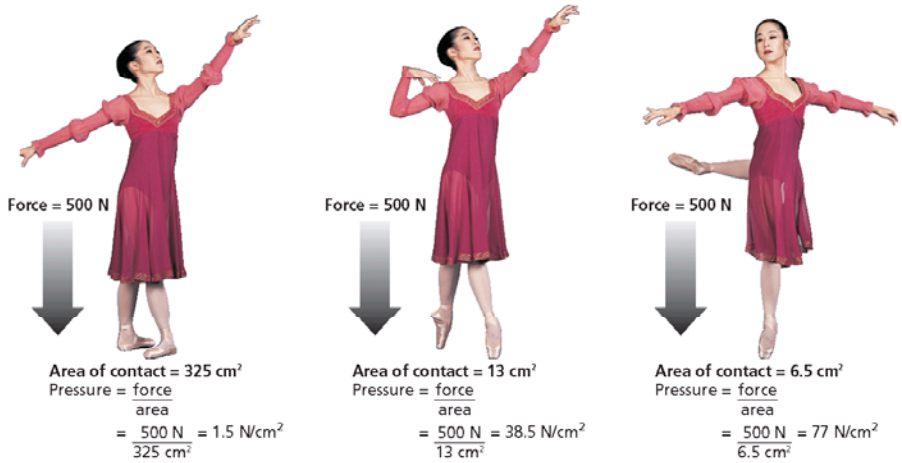
Cotton plugs moistened with ammonia and hydrogen chloride were placed at opposite ends of the glass tube several minutes before this photograph was taken. Why does the white ring of ammonium chloride form closer to the right end than the left end?



Pressure and Force

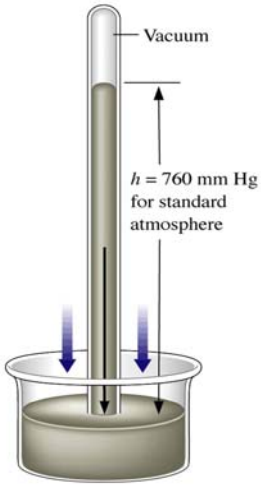
- **Pressure** (P) is defined as the force per unit area on a surface
- Caused by collisions of the gas molecules with each other and with surfaces they come in contact with
  - The greater the force, the greater the pressure
  - The smaller the area, the greater the pressure

Relationship Between Pressure, Force, and Area



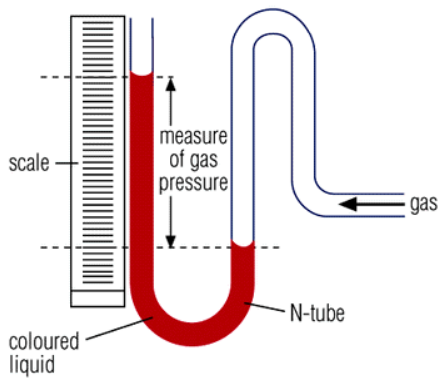
Measuring Pressure

- A **barometer** is a device used to measure atmospheric pressure.
- Introduced by Evangelista Torricelli in the early 1600s.



✓ Consists of Hg in a glass tube
✓ One end is sealed and the other immersed in a reservoir of Hg
✓ Any space at the top of the tube is a vacuum
✓ When the two forces (upward due to atmosphere or downward due to gravity) equalize, the atmospheric pressure can be read

A **manometer** is a device used to measure pressure in a laboratory experiment.



Measuring Pressure (Continued)

- The common unit of pressure is **millimeters of mercury**, symbolized mm Hg.
  - Standard air pressure is 760 mm Hg (29.9 in Hg)
    - A pressure of 1 mm Hg is also called 1 torr in honor of Torricelli for his invention of the barometer.
- **Torr** 1 torr = 1 mm Hg
- **Atmosphere** (atm) 1 atm = 760 mm Hg
- **Pascal** (Pa) 1Pa = 1 N/m^2
  - One pascal is a very small unit. It is more convenient to express pressure in kilopascals (kPa).
  - 1 atm is equal to 101.325 kPa
- **Pounds per square inch** **(psi)**

1 atm = 760 mmHg = 760 torr = 101.325 kPa

Dalton's Law of Partial Pressures

- The pressure of each gas in a mixture is called the **partial pressure** of that gas.
- John Dalton discovered that the pressure exerted by each gas in a mixture is independent of that exerted by other gases present.
- **Dalton's law of partial pressures** states that the total pressure of a gas mixture is the sum of the partial pressures of the component gases.

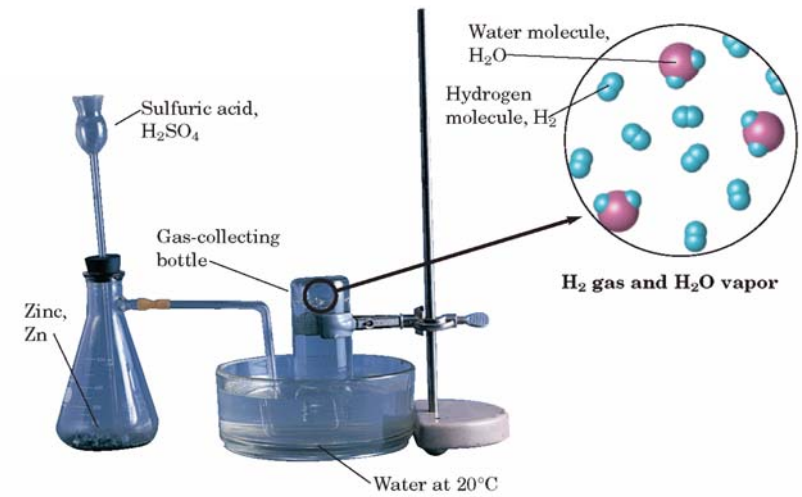
$$P_{total} = P_1 + P_2 + P_3 \ldots$$

Dalton's Law of Partial Pressures
Gases Collected by Water Displacement

- Gases produced in the laboratory are often collected over water.
- To determine the pressure of a gas inside a collection bottle, you would use the following equation:

$$P_{atm} = P_{gas} + P_{H_2O}$$

Model for a Gas Collected Over Water



Sample Problem

Oxygen gas from the decomposition of potassium chlorate, $KClO_3$, was collected by water displacement. The barometric pressure and the temperature during the experiment were 731.0 torr and 20.0°C. respectively. What was the partial pressure of the oxygen collected?

Given: $P_T = P_{atm} = 731.0$ torr

$P_{H_2O} = 17.5$ torr (vapor pressure of water at 20.0°C)

$P_{atm} = P_{O_2} + P_{H_2O}$

Solution: $P_{O_2} = P_{atm} - P_{H_2O}$

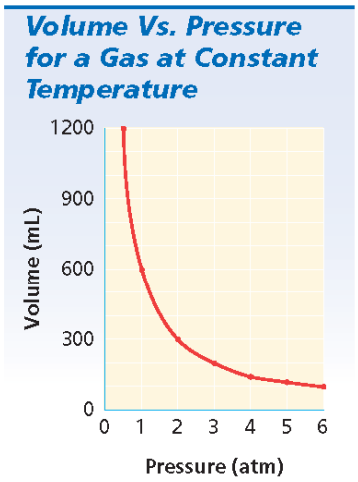
$P_{O_2} = 731.0 \text{ torr} - 17.5 \text{ torr} = 713.5 \text{ torr}$

Gas Laws

1. **Boyles' Law** (pressure-volume relationship)

- At constant temperature, pressure and volume are inversely proportional

$$P_1V_1 = P_2V_2$$

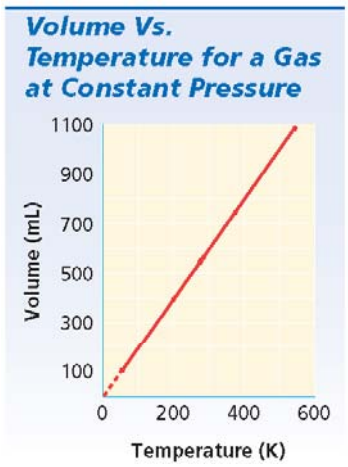


2. **Charles' Law** (Volume- temperature relationship)

- discovered by the French scientist Jacques Charles in 1787

- At constant pressure, the volume of a gas is directly proportional to the temperature (use Kelvin).

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$



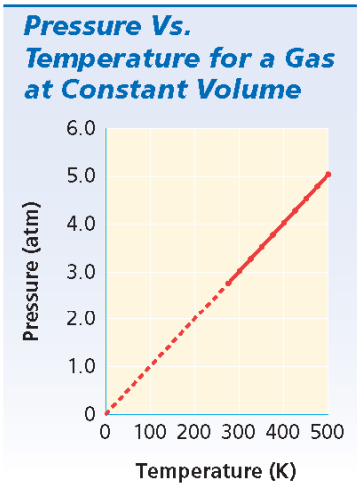
The temperature –273°C is referred to as **absolute zero,** and is given a value of zero in the Kelvin temperature scale. **K = 273 + °C.**

3. **Gay-Lussac's Law** (Pressure-Temperature relationship)

Named after Joseph Gay-Lussac, who discovered it in 1802

- At constant volume, pressure and temperature (in Kelvin) of a contained gas are directly proportional

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$



Combined Gas Law

- Expresses the relationship between pressure, volume, and temperature of a fixed amount of gas.

$$\frac{V_1P_1}{T_1} = \frac{V_2P_2}{T_2}$$

Avogadro's Law

- In 1811, Amedeo Avogadro stated that equal volumes of gases at the same temperature and pressure contain equal numbers of molecules.
- The volume occupied by one mole of gas at STP is known as the **standard molar volume of a gas**, which is 22.41410 L (rounded to 22.4 L).

Ideal Gas Law

$$PV = nRT$$

Where: *P = pressure*
V = volume (in L)
T = temperature (in Kelvin)
n = moles
R = constant (known as the ideal gas constant)
Its value depends on the units chosen for pressure

Numerical Values of the Gas Constant

Units of *R*	Numerical value of *R*	Units of *P*	Units of *V*	Units of *T*	Units of *n*
$\frac{L \cdot mm\ Hg}{mol \cdot K}$	62.4	mm Hg	L	K	mol
$\frac{L \cdot atm}{mol \cdot K}$	0.0821	atm	L	K	mol
$\frac{J}{mol \cdot K}$ *	8.314	atm	L	K	mol
$\frac{L \cdot kPa}{mol \cdot K}$	8.314	kPa	L	K	mol

Note: 1 L•atm = 101.325 J; 1 J = 1 Pa•m^3
* SI units

Finding Molar Mass from the Ideal Gas Law

- moles (n) = mass (*m*) / Molar Mass (M)
- Substitute into the ideal gas law and rearrange to determine the molar mass of a gas

$$M = \frac{m\,R\,T}{PV}$$

Finding Density from the Ideal Gas Law

- Density (D) = mass (*m*) / volume (V)
- Substitute into the ideal gas law and rearrange to determine the density of a gas

$$D = \frac{M\,P}{R\,T}$$