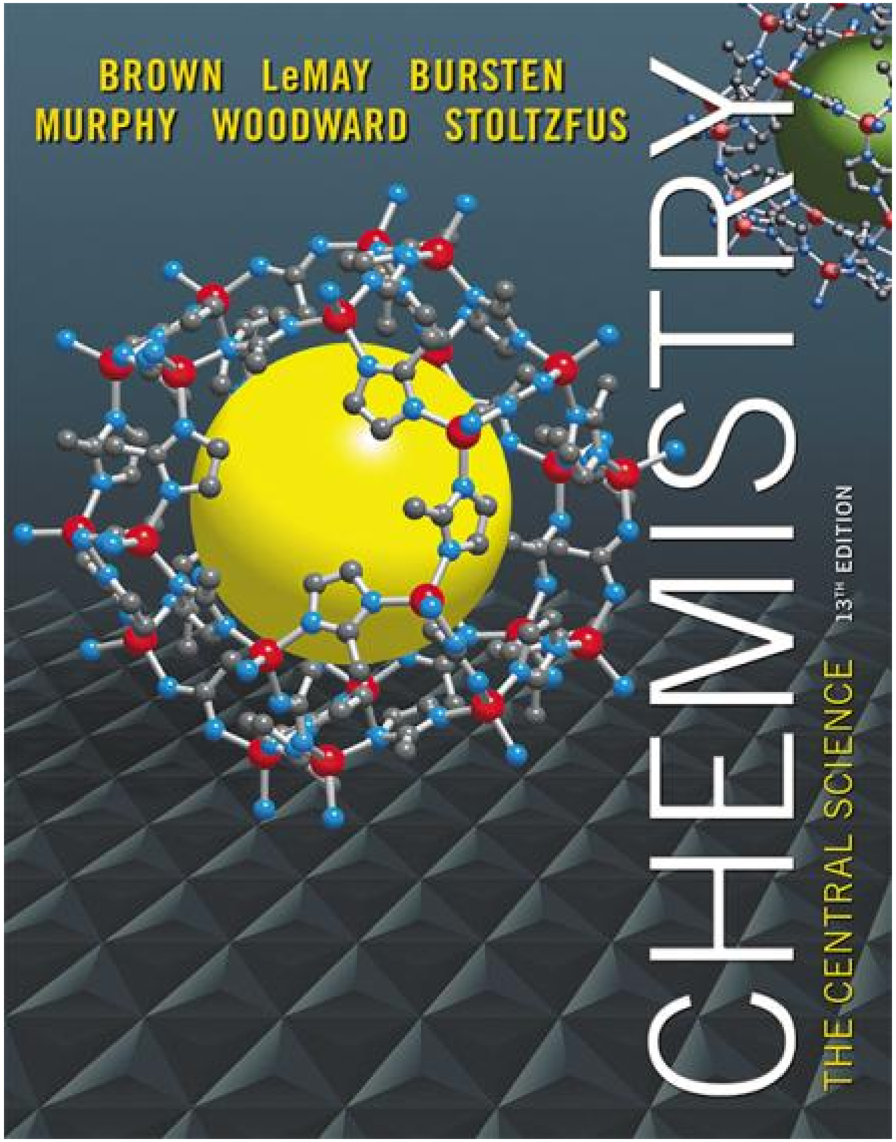


Lecture Presentation

Chapter 10

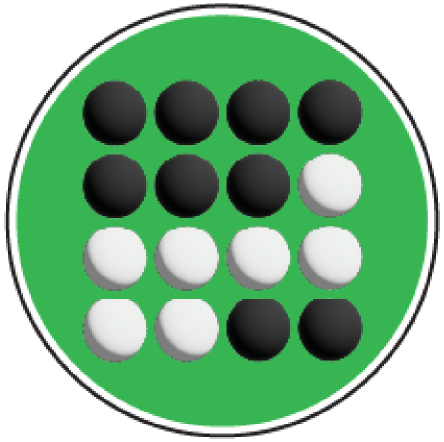
Gases

Characteristics of Gases

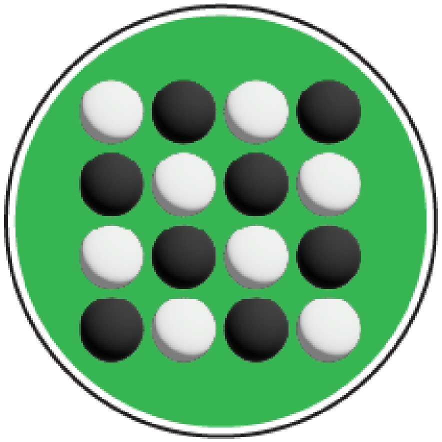
- Physical properties of gases are all similar.
- Composed mainly of nonmetallic elements with simple formulas and low molar masses.
- Unlike liquids and solids, gases
  - expand to fill their containers.
  - are highly compressible.
  - have extremely low densities.
- Two or more gases form a homogeneous mixture.

Mixtures

Heterogeneous Mixture



Homogeneous Mixture



Gases

Properties Which Define the State of a Gas Sample

1) Temperature
2) Pressure
3) Volume
4) Amount of gas, usually expressed as number of moles

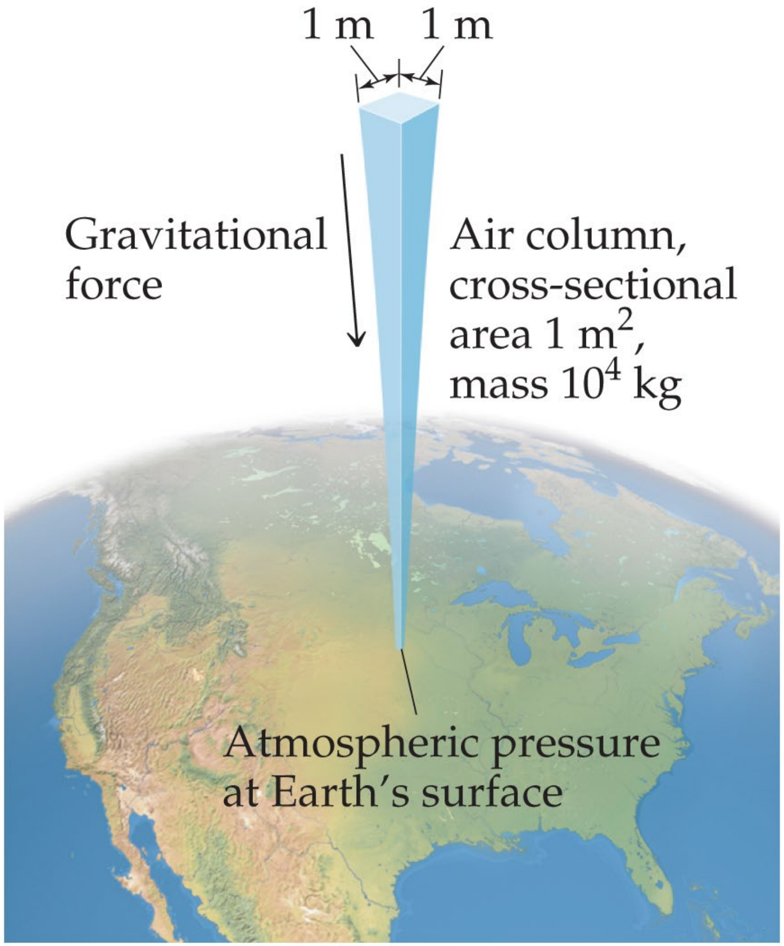
- Having already discussed three of these, we need to define pressure.

Pressure

- **Pressure** is the amount of force applied to an area:

$$P = \frac{F}{A}$$

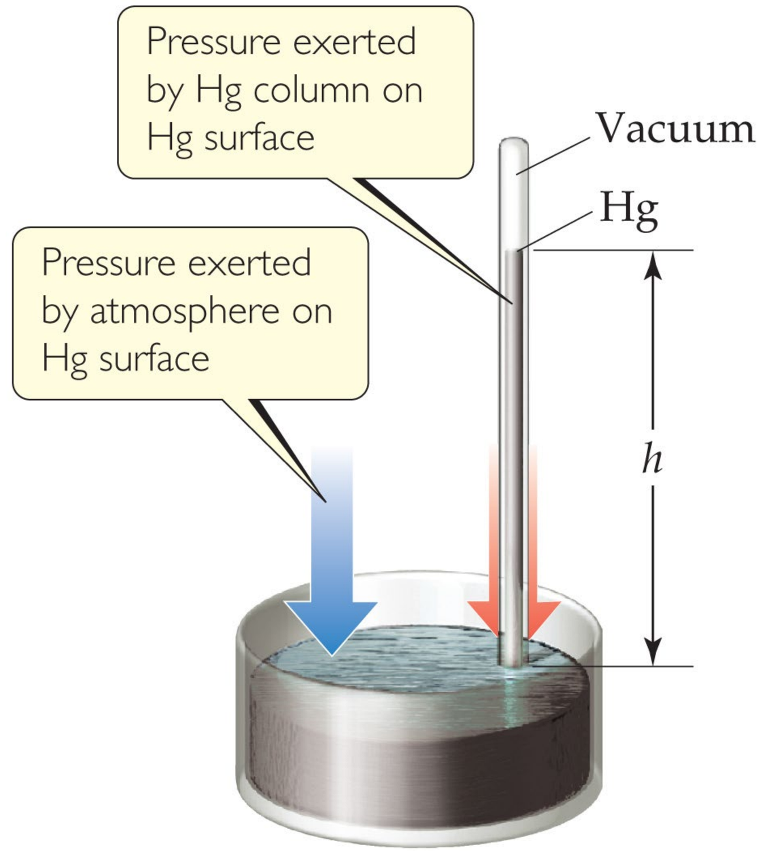
- **Atmospheric pressure** is the weight of air per unit of area.



Units of Pressure

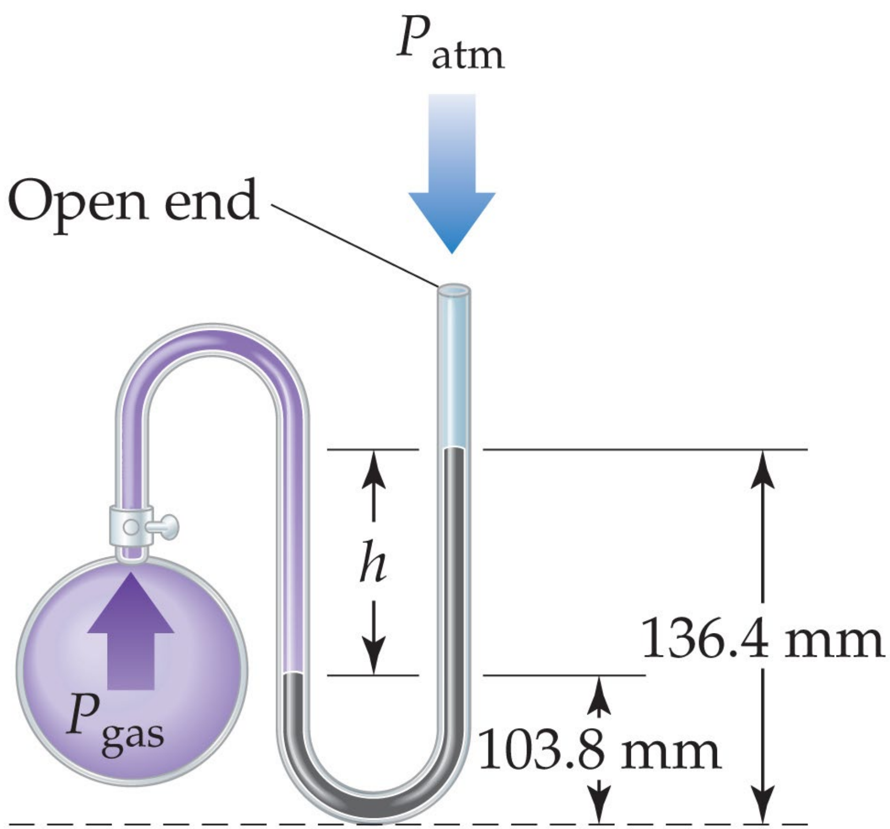
- **Pascals**: 1 Pa = 1 N/m^2 (SI unit of pressure)
- **Bar**: 1 bar = 10^5 Pa = 100 kPa
- **mm Hg or torr**: These units are literally the difference in the heights measured in mm of two connected columns of mercury, as in the **barometer** in the figure.
- **Atmosphere**:

1.00 atm = 760 torr = 760 mm Hg = 101.325 kPa



Gases

Manometer



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

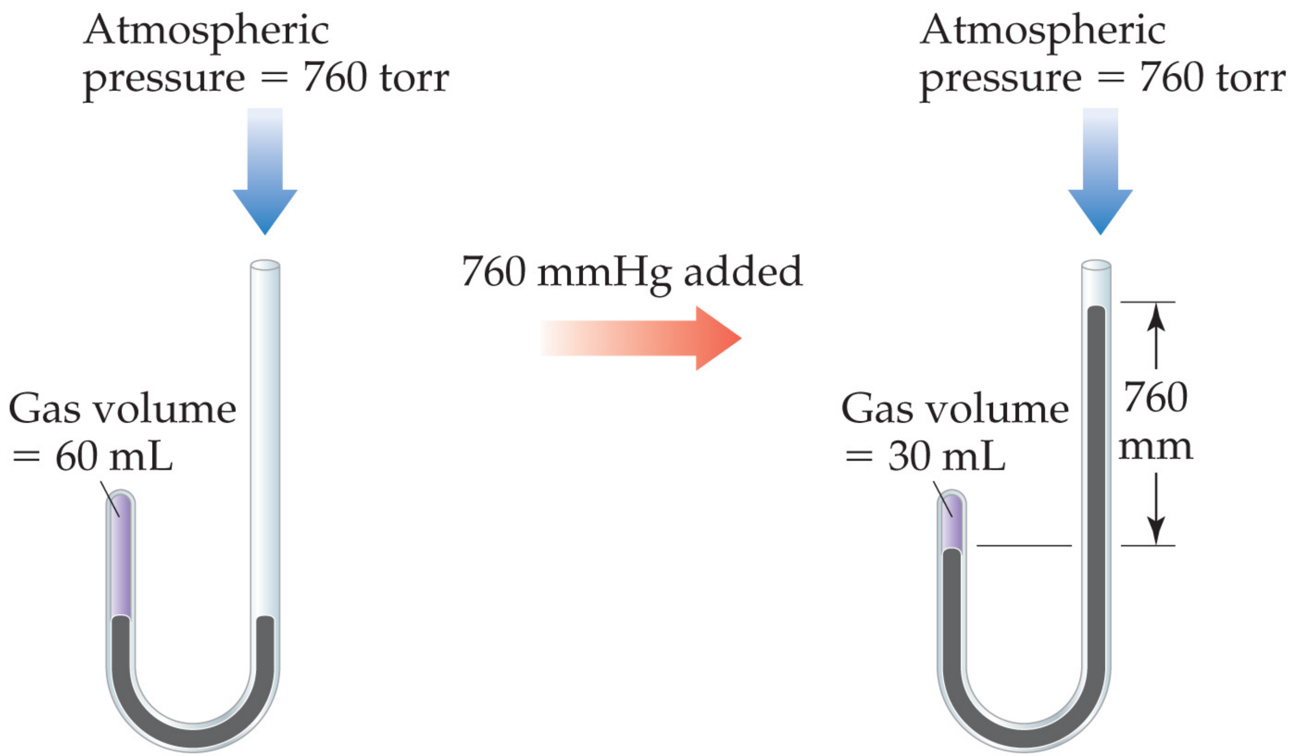
The **manometer** is used to measure the difference in pressure between atmospheric pressure and that of a gas in a vessel. (The barometer seen on the last slide is used to measure the pressure in the atmosphere at any given time.)

Standard Pressure

- Normal atmospheric pressure at sea level is referred to as **standard atmospheric pressure**.
- It is equal to
  - 1.00 atm.
  - 760 torr (760 mmHg).
  - 101.325 kPa.

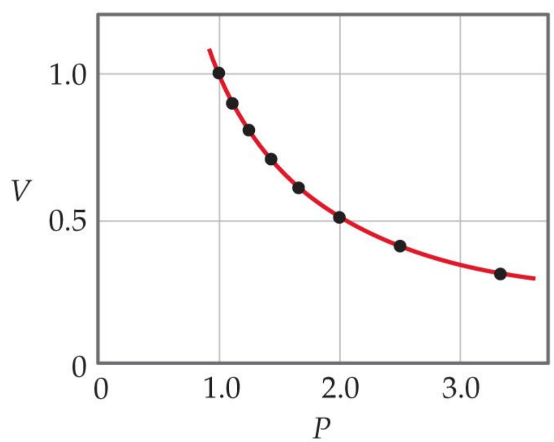
Boyle's Law

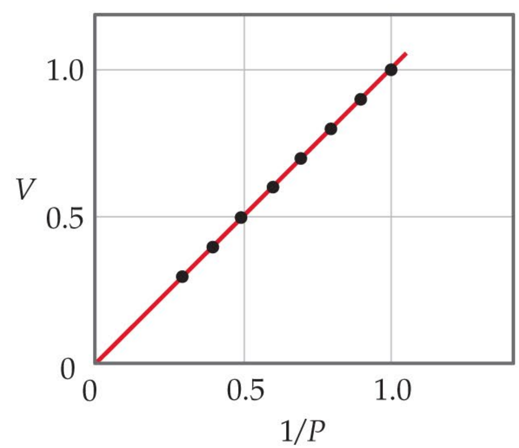
The volume of a fixed quantity of gas at constant temperature is inversely proportional to the pressure.



Mathematical Relationships of Boyle's Law

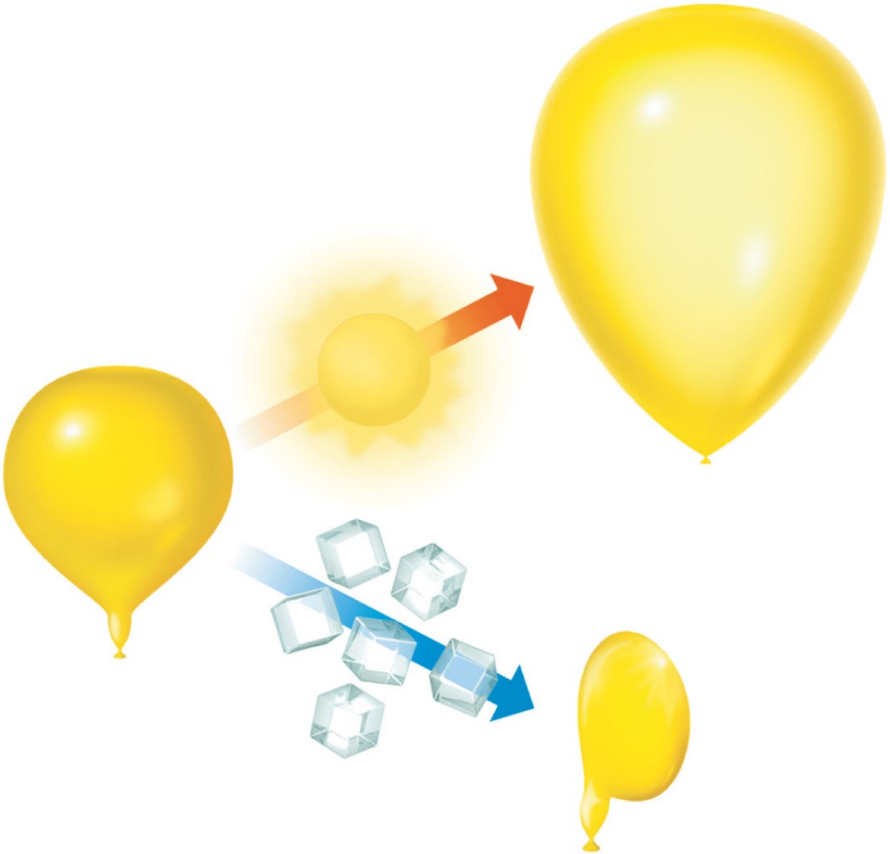
- *PV* = a constant
- This means, if we compare two conditions: $P_1V_1 = P_2V_2$.
- Also, if we make a graph of *V* vs. *P*, it will *not* be linear. However, a graph of *V* vs. 1/*P* (reciprocal of P) *will* result in a linear relationship!





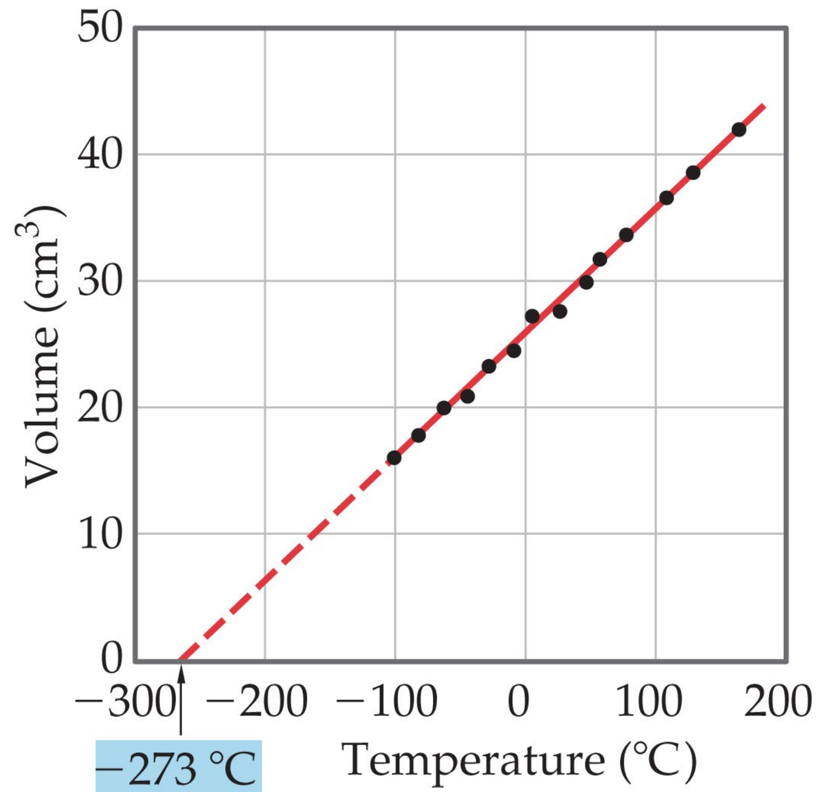
Charles's Law

- The volume of a fixed amount of gas at constant pressure is directly proportional to its absolute temperature.



Mathematical Relationships of Charles's Law

- V = constant × T
- This means, if we compare two conditions: $V_1/T_1 = V_2/T_2$.
- Also, if we make a graph of V vs. T, it will be linear.



Gases

Avogadro's Law

- The volume of a gas at constant temperature and pressure is directly proportional to the number of moles of the gas.
- Also, at STP, one mole of gas occupies 22.4 L.
- Mathematically: V = constant × n, or $V_1/n_1 = V_2/n_2$

Volume	22.4 L	22.4 L	22.4 L
Pressure	1 atm	1 atm	1 atm
Temperature	0 °C	0 °C	0 °C
Mass of gas	4.00 g	28.0 g	16.0 g
Number of gas molecules	6.02×10^{23}	6.02×10^{23}	6.02×10^{23}

Ideal-Gas Equation

- So far we've seen that

$V \propto 1/P$ (Boyle's law).

$V \propto T$ (Charles's law).

$V \propto n$ (Avogadro's law).

- Combining these, we get

$$V \propto \frac{nT}{P}$$

- Finally, to make it an equality, we use a constant of proportionality (R) and reorganize; this gives the Ideal-Gas Equation: $PV = nRT$.

Density of Gases

If we divide both sides of the ideal-gas equation by V and by RT, we get

$$n/V = P/RT.$$

Also: moles × molecular mass = mass

$$n \times \mathrm{M} = m.$$

If we multiply both sides by M, we get

$$m/V = MP/RT$$

and m/V is density, d; the result is:

$$d = MP/RT.$$

Density & Molar Mass of a Gas

- To recap:
- One needs to know only the molecular mass, the pressure, and the temperature to calculate the density of a gas.
- $d = MP/RT$
- Also, if we know the mass, volume, and temperature of a gas, we can find its molar mass.
- $M = mRT/PV$

Volume and Chemical Reactions

- The balanced equation tells us relative amounts of moles in a reaction, whether the compared materials are products or reactants.
- $PV = nRT$
- So, we can relate volume for gases, as well.
- For example: use ($PV = nRT$) for substance A to get moles A; use the mole ratio from the balanced equation to get moles B; and ($PV = nRT$) for substance B to get volume of B.

Dalton's Law of Partial Pressures

- If two gases that *don't* react are combined in a container, they act as if they are alone in the container.
- The total pressure of a mixture of gases equals the sum of the pressures that each would exert if it were present alone.
- In other words,

$$P_{\text{total}} = p_1 + p_2 + p_3 + \ldots$$

Mole Fraction

- Because each gas in a mixture acts as if it is alone, we can relate amount in a mixture to partial pressures:

$$\frac{P_1}{P_t} = \frac{n_1 RT/V}{n_t RT/V} = \frac{n_1}{n_t}$$

- That ratio of moles of a substance to total moles is called the **mole fraction**, χ.

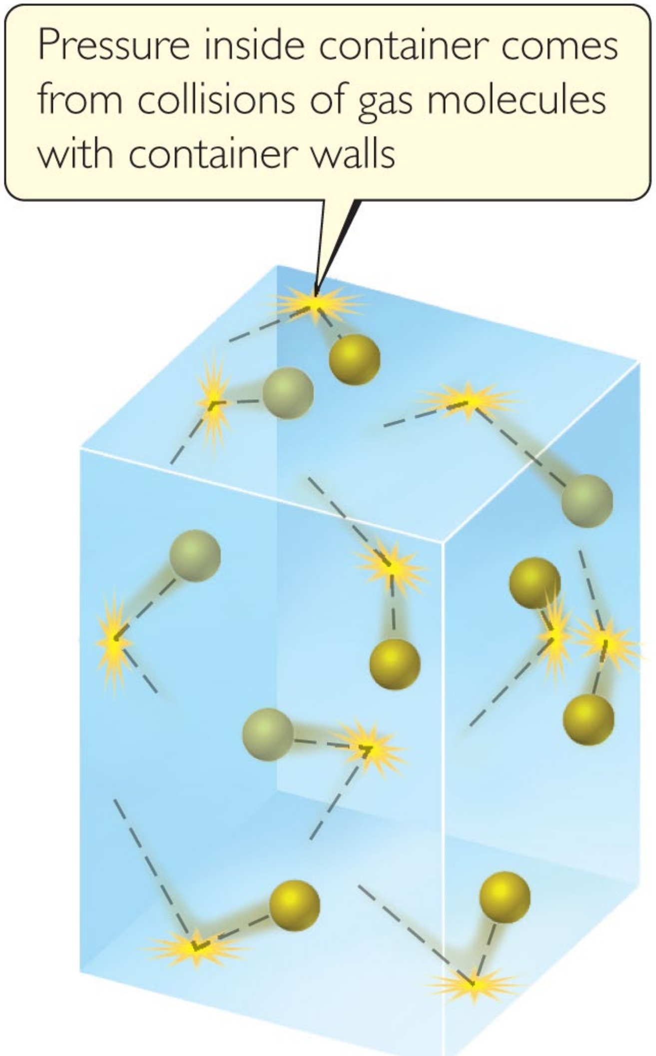
$$X_1 = \frac{\text{Moles of compound 1}}{\text{Total moles}} = \frac{n_1}{n_t}$$

Pressure and Mole Fraction

- The end result is

$$P_1 = \left(\frac{n_1}{n_t}\right)P_t = X_1P_t$$

Kinetic-Molecular Theory

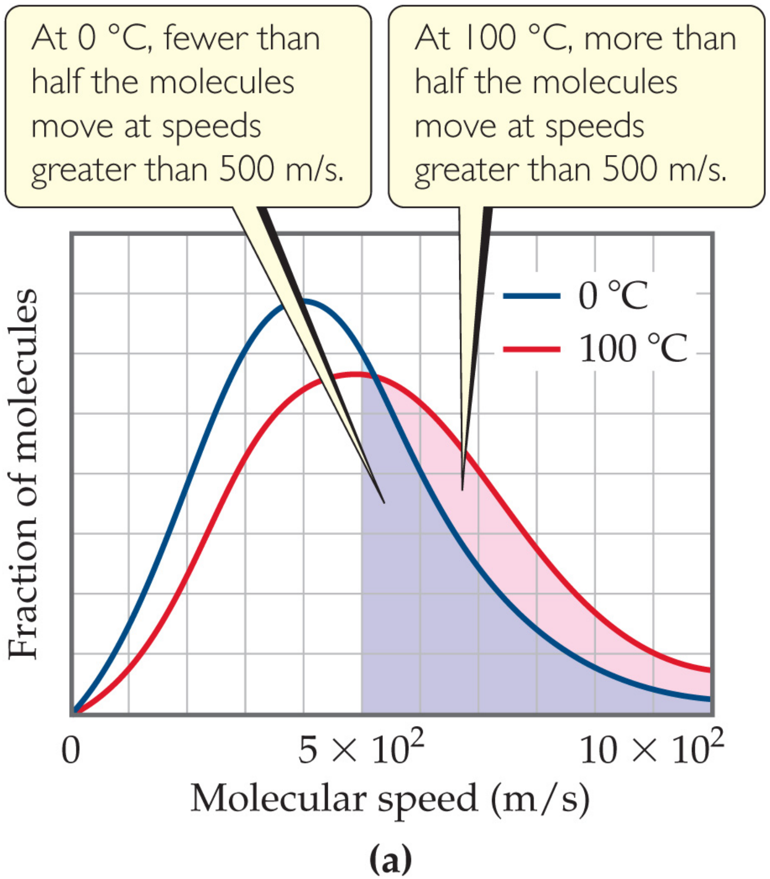


- *Laws* tell us *what* happens in nature. Each of the gas laws we have discussed tell us what is observed under certain conditions.
- *Why* are these laws observed? We will discuss a *theory* to explain our observations.

Main Tenets of Kinetic-Molecular Theory

1) Gases consist of large numbers of molecules that are in continuous, random motion.
2) The combined volume of all the molecules of the gas is negligible relative to the total volume in which the gas is contained.
3) Attractive and repulsive forces between gas molecules are negligible.

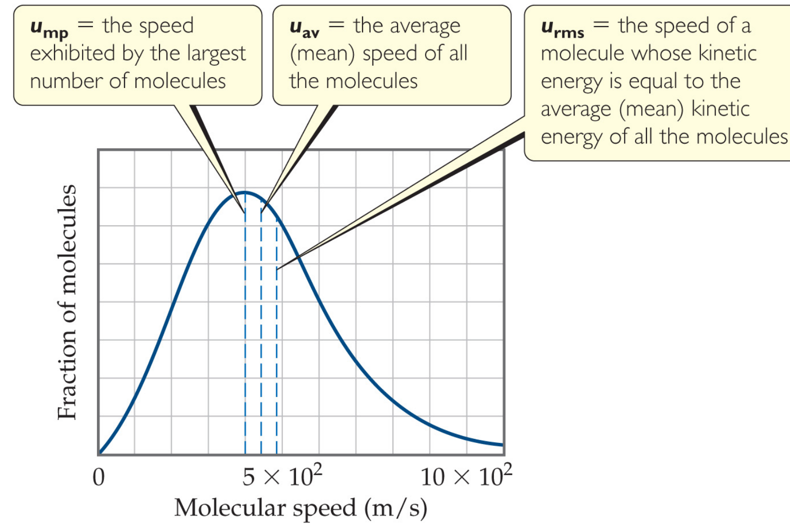
Main Tenets of Kinetic-Molecular Theory



4) Energy can be transferred between molecules during collisions, but the *average* kinetic energy of the molecules does not change with time, as long as the temperature of the gas remains constant.
5) The average kinetic energy of the molecules is proportional to the absolute temperature.

How Fast Do Gas Molecules Move?

- Temperature is related to their *average* kinetic.
- Individual molecules can have different speeds of motion.
- The figure shows three different speeds:



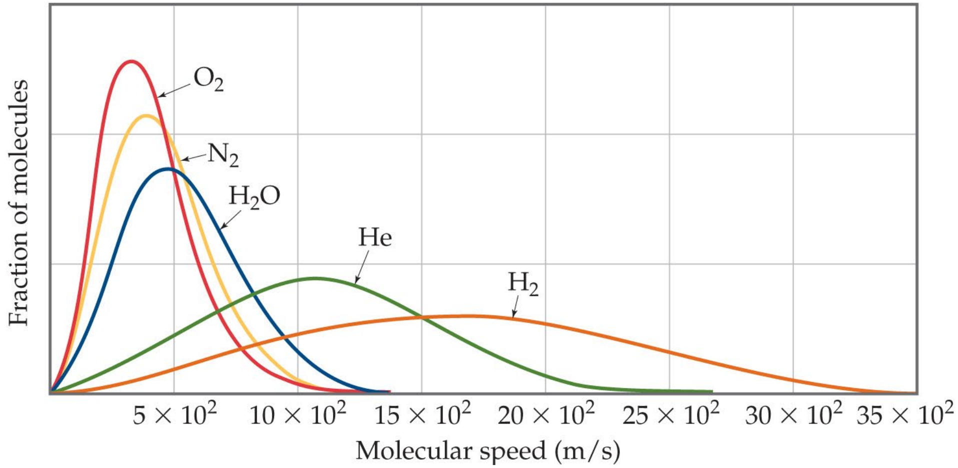
(b)

- ➢ u_{mp} is the most probable speed (most molecules are this fast).
- ➢ u_{av} is the average speed of the molecules.
- ➢ u_{rms}, the root-mean-square speed, is the one associated with their average kinetic energy.

u_{rms} and Molecular Mass

$$u_{rms} = \sqrt{\frac{3RT}{\mathcal{M}}}$$

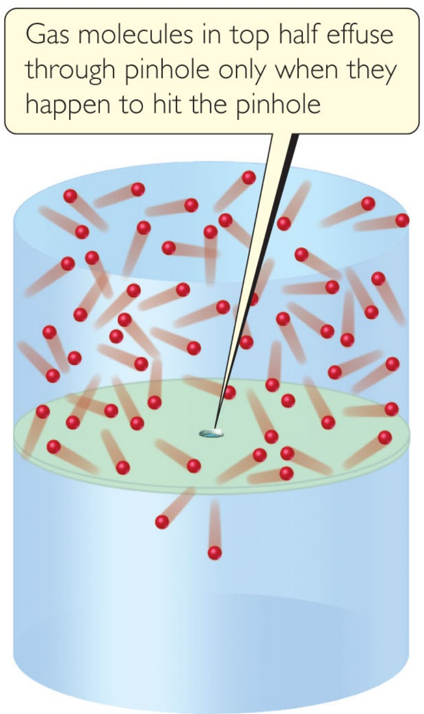
- At any given temperature, the average kinetic energy of molecules is the same.
- So, ½ m $(u_{rms})^2$ is the same for two gases at the same temperature.
- If a gas has a low mass, its speed will be greater than for a heavier molecule.



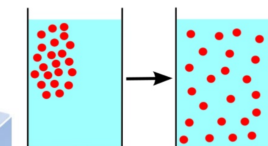
Gases

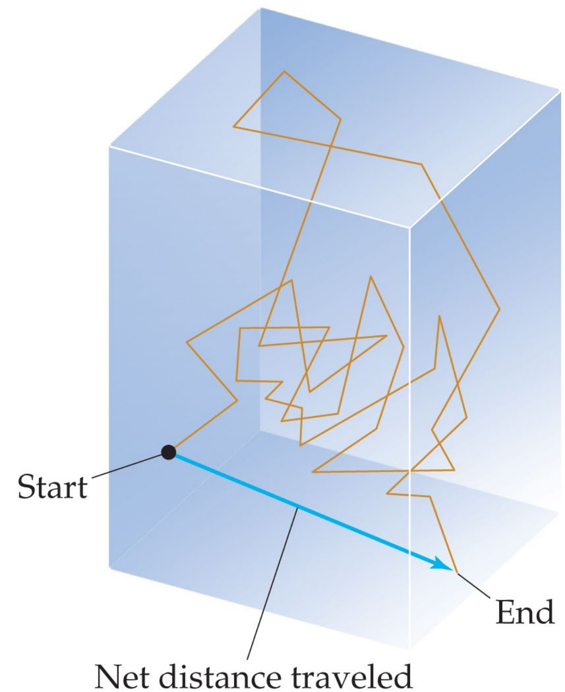
Effusion & Diffusion

Effusion（渗透）is the escape of gas molecules through a tiny hole into an evacuated space.



Diffusion（扩散）is the spread of one substance throughout a space or a second substance.





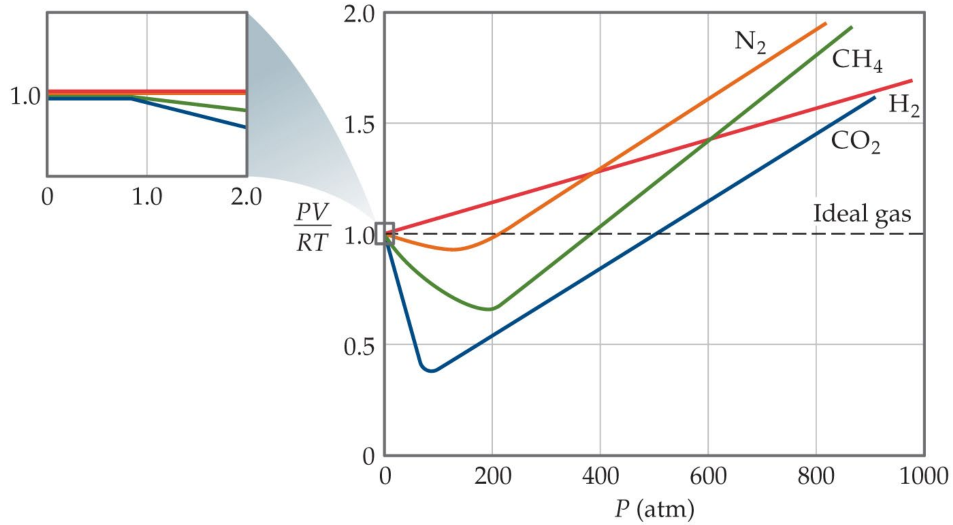
Graham's Law Describes Diffusion & Effusion

- Graham's Law relates the molar mass of two gases to their rate of speed of travel.
- The "lighter" gas always has a faster rate of speed.

rate $\frac{r_1}{r_2} = \sqrt{\frac{\mathcal{M}_2}{\mathcal{M}_1}}$ Molar mass

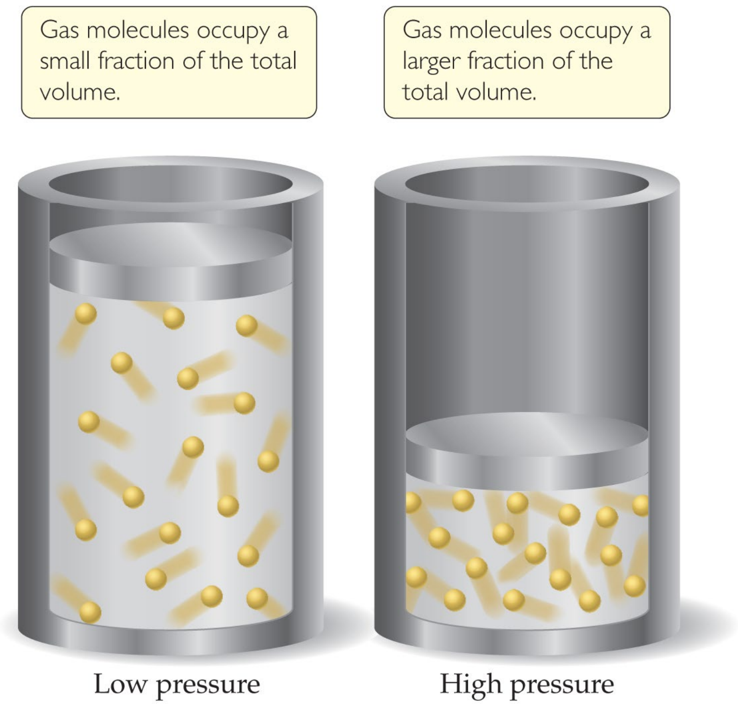
Real Gases

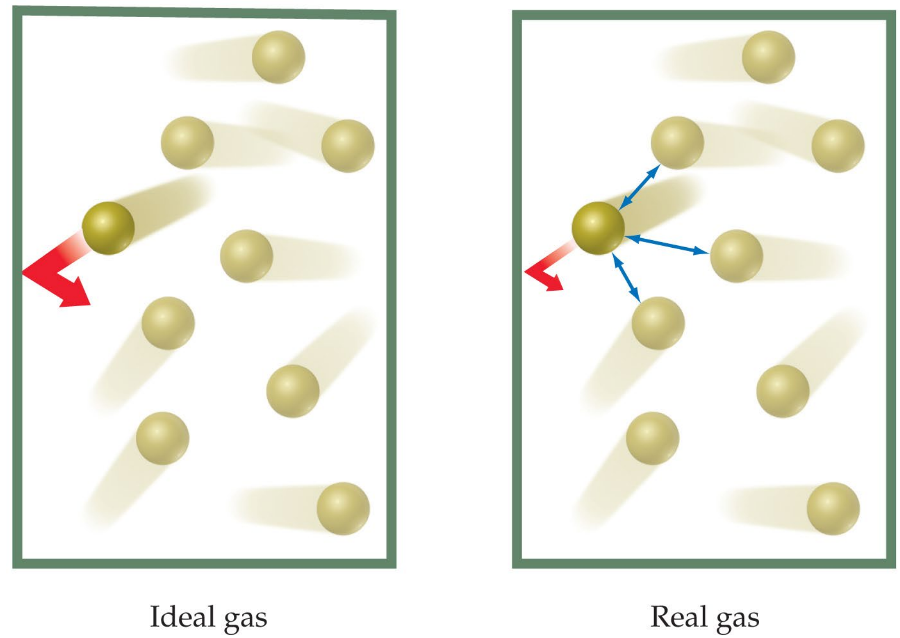
- In the real world, the behavior of gases only conforms to the ideal-gas equation at relatively high temperature and low pressure.
- Even the same gas will show wildly different behavior under high pressure at different temperatures.



Gases

Deviations from Ideal Behavior





The assumptions made in the kinetic-molecular model (negligible volume of gas molecules themselves, no attractive forces between gas molecules, etc.) break down at high pressure and/or low temperature.

Corrections for Nonideal Behavior

- The ideal-gas equation can be adjusted to take these deviations from ideal behavior into account.
- The corrected ideal-gas equation is known as the **van der Waals equation**.
- The pressure adjustment is due to the fact that molecules attract and repel each other.
- The volume adjustment is due to the fact that molecules occupy some space on their own.

The van der Waals Equation

$$\left(P + \frac{n^2a}{V^2}\right)(V - nb) = nRT$$

Table 10.3 Van der Waals Constants for Gas Molecules

Substance	a (L^2-atm/mol^2)	b (L/mol)
He	0.0341	0.02370
Ne	0.211	0.0171
Ar	1.34	0.0322
Kr	2.32	0.0398
Xe	4.19	0.0510
H_2	0.244	0.0266
N_2	1.39	0.0391
O_2	1.36	0.0318
F_2	1.06	0.0290
Cl_2	6.49	0.0562
H_2O	5.46	0.0305
NH_3	4.17	0.0371
CH_4	2.25	0.0428
CO_2	3.59	0.0427
CCl_4	20.4	0.1383