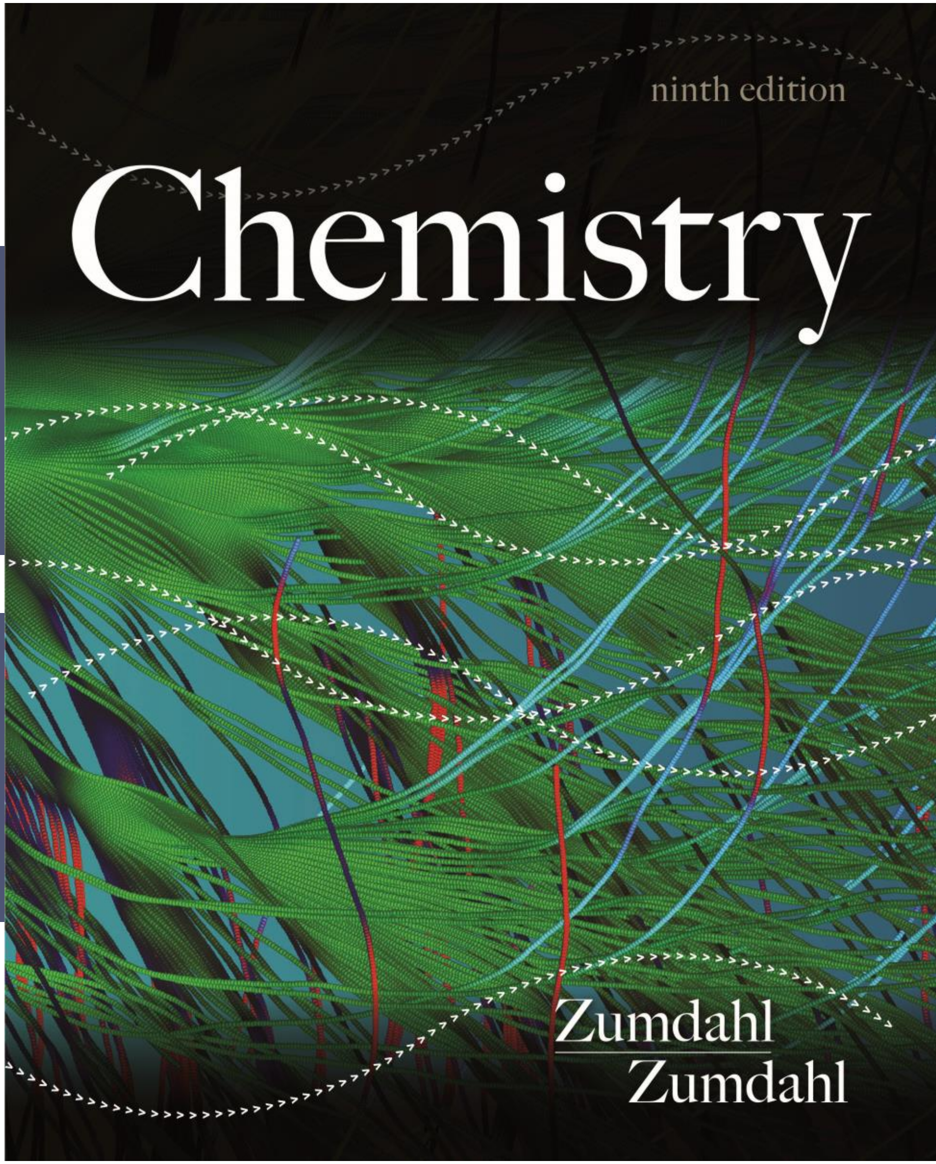


Chapter 5

Gases

Answer in your notes, compare with partner

- The pressure of a gas is measured as 49 torr
  - Represent this pressure in both atmospheres and pascals

Interactive Example 5.1 - Solution

$$49\ \cancel{\text{torr}} \times \frac{1\ \text{atm}}{760\ \cancel{\text{torr}}} = 6.4 \times 10^{-2}\ \text{atm}$$

$$6.4 \times 10^{-2}\ \cancel{\text{atm}} \times \frac{101{,}325\ \text{Pa}}{1\ \cancel{\text{atm}}} = 6.5 \times 10^{3}\ \text{Pa}$$

Ideal Gas Law (Continued)

- Can be rearranged to:

$$PV = nRT$$

- Equation of state for a gas, where the state of the gas is its condition at a given time
  - Any gas that obeys this law is said to be behaving ideally

Molar Volume of an Ideal Gas

- **Molar volume**: For 1 mole of an ideal gas at 0°C and 1 atm, the volume of the gas is 22.42 L
  - **Standard temperature and pressure (STP)**: Conditions 0°C and 1 atm

Answer in your notes, compare with partner

- What if STP was defined as normal room temperature (22° C) and 1 atm?
  - How would this affect the molar volume of an ideal gas?
    - Include an explanation and a number

Molar Mass of a Gas

- Ideal gas law is essential for the calculation of molar mass of a gas from its measured density

$$n = \frac{\text{grams of gas}}{\text{molar mass}} = \frac{\text{mass}}{\text{molar mass}} = \frac{m}{\text{molar mass}}$$

  - Substituting into the ideal gas equation gives:

$$P = \frac{nRT}{V} = \frac{(m/\text{molar mass})RT}{V} = \frac{m(RT)}{V(\text{molar mass})}$$

Answer in your notes, compare with a partner

- The density of a gas was measured at 1.50 atm and 27° C and found to be 1.95 g/L
  - Calculate the molar mass of the gas

Interactive Example 5.14 - Solution

- Where are we going?
  - To determine the molar mass of the gas
- What do we know?
  - P = 1.50 atm
  - T = 27°C + 273 = 300 K
  - d = 1.95 g/L

Interactive Example 5.14 - Solution (Continued)

- What information do we need?
  - $\text{Molar mass} = \dfrac{dRT}{P}$
  - R = 0.08206 L · atm/K · mol
- How do we get there?

$$\text{Molar mass} = \frac{dRT}{P} = \frac{\left(1.95\,\frac{\text{g}}{\cancel{\text{L}}}\right)\left(0.08206\,\frac{\cancel{\text{L}}\cdot\cancel{\text{atm}}}{\cancel{\text{K}}\cdot\text{mol}}\right)\left(300\ \cancel{\text{K}}\right)}{1.50\ \cancel{\text{atm}}} = 32.0\ \text{g/mol}$$

- Reality check - These are the units expected for molar mass

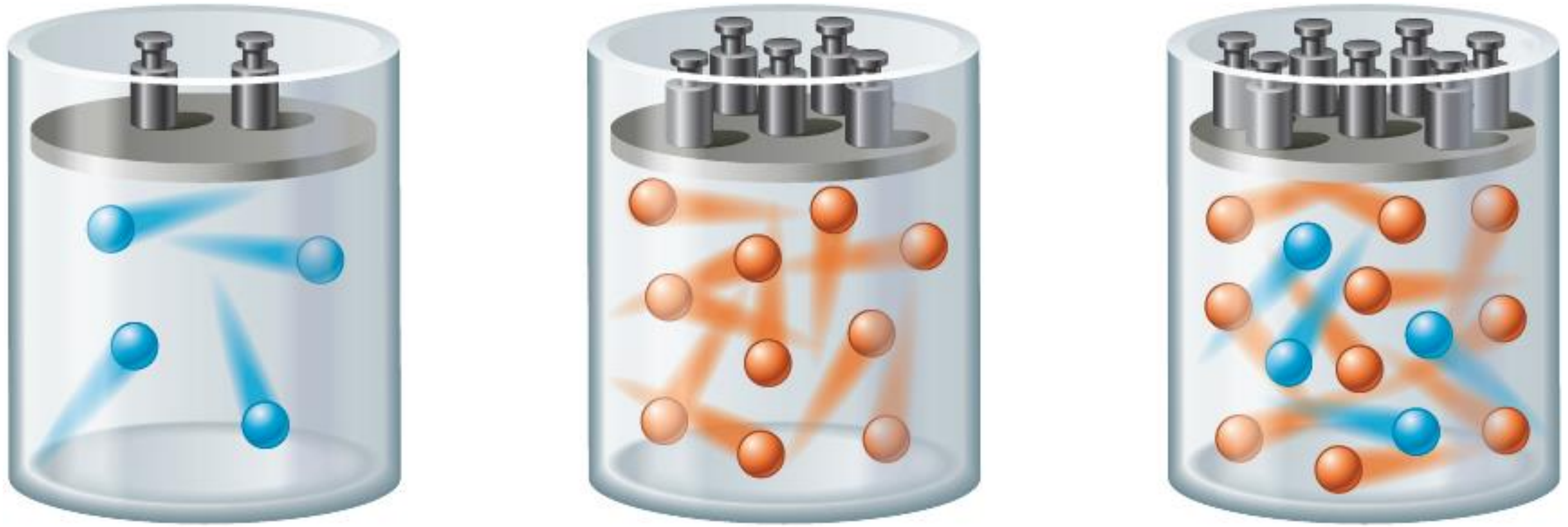
Law of Partial Pressures - John Dalton

- For a mixture of gases in a container, the total pressure exerted is the sum of the partial pressures

$$P_{\text{TOTAL}} = P_1 + P_2 + P_3 + \ldots$$

  - **Partial pressure**: Pressure that a gas would exert if it were alone in a container

Figure 5.12 - Schematic Diagram of Dalton's Law of Partial Pressures



Write and answer in your notes, compare with partner

- Mixtures of helium and oxygen can be used in scuba diving tanks to help prevent "the bends"
  - For a particular dive, 46 L He at 25° C and 1.0 atm and 12 L O_2 at 25° C and 1.0 atm were pumped into a tank with a volume of 5.0 L
    - Calculate the partial pressure of each gas and the total pressure in the tank at 25° C

Interactive Example 5.15 - Solution

- Where are we going?
  - To determine the partial pressure of each gas
  - To determine the total pressure in the tank at 25° C
- What do we know?

	He	O_2	Tank
P	1.00 atm	1.00 atm	? atm
V	46 L	12 L	5.0 L
T	25°C + 273 = 298 K	25°C + 273 = 298 K	25°C + 273 = 298 K

Interactive Example 5.15 - Solution (Continued 1)

- What information do we need?
  - Ideal gas law

$$PV = nRT$$

  - $R = 0.08206\ \mathrm{L \cdot atm/K \cdot mol}$
- How do we get there?
  - How many moles are present for each gas?

$$n = \frac{PV}{RT}$$

Interactive Example 5.15 - Solution (Continued 2)

$$n_{\text{He}} = \frac{(1.0\ \cancel{\text{atm}})(46\ \cancel{\text{L}})}{(0.08206\ \cancel{\text{L}} \cdot \cancel{\text{atm}}/\cancel{\text{K}} \cdot \text{mol})(298\ \cancel{\text{K}})} = 1.9\ \text{mol}$$

$$n_{\text{O}_2} = \frac{(1.0\ \cancel{\text{atm}})(12\ \cancel{\text{L}})}{(0.08206\ \cancel{\text{L}} \cdot \cancel{\text{atm}}/\cancel{\text{K}} \cdot \text{mol})(298\ \cancel{\text{K}})} = 0.49\ \text{mol}$$

Interactive Example 5.15 - Solution (Continued 3)

- What is the partial pressure for each gas in the tank?
  - The tank containing the mixture has a volume of 5.0 L, and the temperature is 25° C
  - We can use these data and the ideal gas law to calculate the partial pressure of each gas

$$P = \frac{nRT}{V}$$

$$P_{He} = \frac{(1.9\ \cancel{\text{mol}})(0.08206\ \cancel{\text{L}} \cdot \text{atm}/\cancel{\text{K}} \cdot \cancel{\text{mol}})(298\ \cancel{\text{K}})}{5.0\ \cancel{\text{L}}} = 9.3\ \text{atm}$$

Interactive Example 5.15 - Solution (Continued 4)

$$P_{O_2} = \frac{(0.49\ \cancel{\text{mol}})(0.08206\ \cancel{\text{L}} \cdot \text{atm}/\cancel{\text{K}} \cdot \cancel{\text{mol}})(298\ \cancel{\text{K}})}{5.0\ \cancel{\text{L}}} = 2.4\ \text{atm}$$

- What is the total pressure of the mixture of gases in the tank?
  - The total pressure is the sum of the partial pressures

$$P_{\text{TOTAL}} = P_{\text{He}} + P_{O_2} = 9.3\ \text{atm} + 2.4\ \text{atm} = 11.7\ \text{atm}$$

Mole Fraction (χ)

- Ratio of number of moles of a given component in a mixture to the total number of moles
- Example
  - For a given component in a mixture, χ_1 is calculated as follows:

$$\chi_1 = \frac{n_1}{n_{\text{TOTAL}}} = \frac{n_1}{n_1 + n_2 + n_3 + \ldots}$$

Answer in your notes, compare with partner

- The partial pressure of oxygen was observed to be 156 torr in air with a total atmospheric pressure of 743 torr
  - Calculate the mole fraction of O_2 present

Interactive Example 5.16 - Solution

- Where are we going?
  - To determine the mole fraction of O_2
- What do we know?
  - P_{O_2} = 156 torr
  - P_{TOTAL} = 743 torr

Interactive Example 5.16 - Solution (Continued)

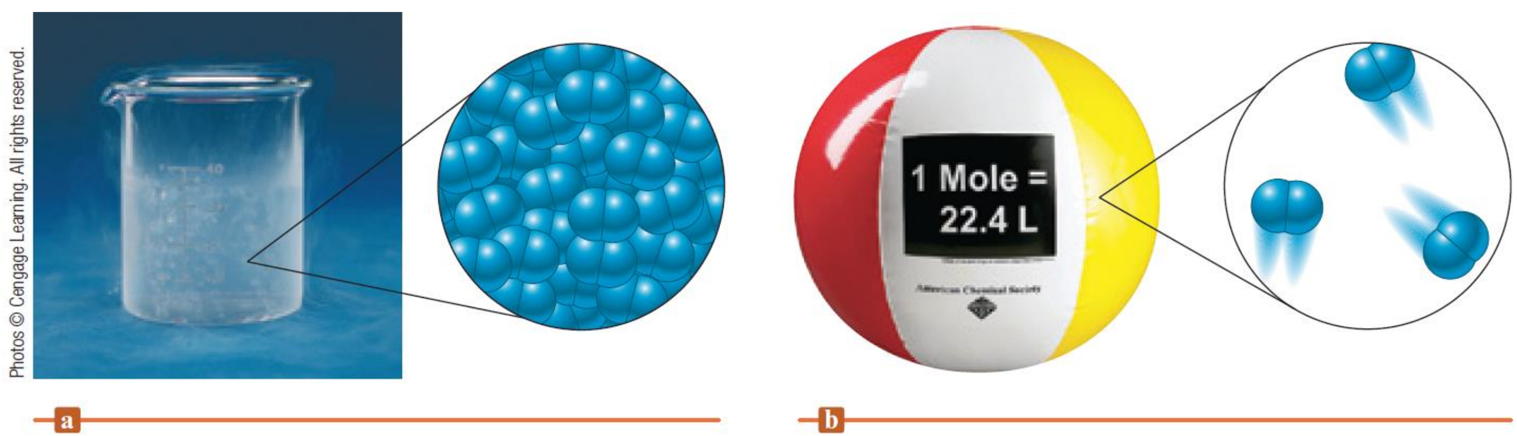
- How do we get there?
  - The mole fraction of O_2 can be calculated from the equation

$$\chi_{O_2} = \frac{P_{O_2}}{P_{TOTAL}} = \frac{156 \ \cancel{\text{torr}}}{743 \ \cancel{\text{torr}}} = 0.210$$

    - Note that the mole fraction has no units

Postulates of the Kinetic Molecular Theory

1. Particles are so small compared with the distances between them that the volume of the individual particles can be assumed to be negligible (zero)



Postulates of the Kinetic Molecular Theory (Continued 1)

2. Particles are in constant motion
   - Collisions of the particles with the walls of the container are the cause of the pressure exerted by the gas
3. Particles are assumed to exert no forces on each other
   - Particles are assumed neither to attract nor to repel each other

Postulates of the Kinetic Molecular Theory (Continued 2)

4. Average kinetic energy of a collection of gas particles is assumed to be directly proportional to the Kelvin temperature of the gas

Figure 5.16 - Effects of Decreasing the Volume of a Sample of Gas at Constant Temperature

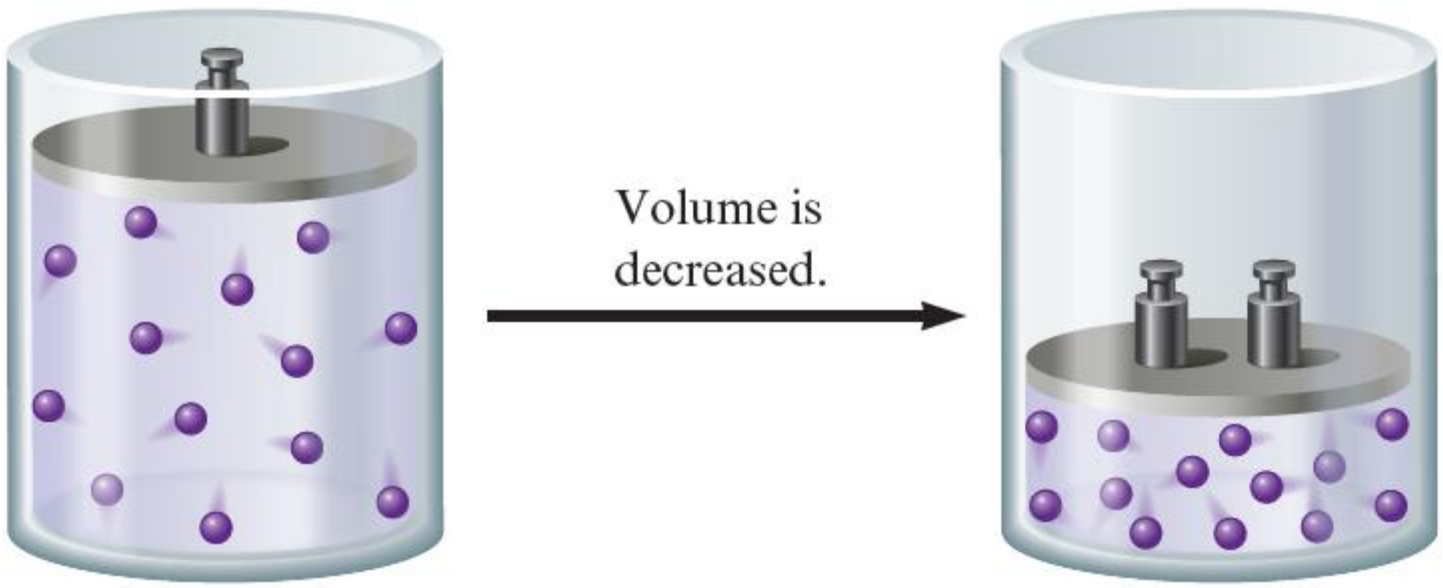
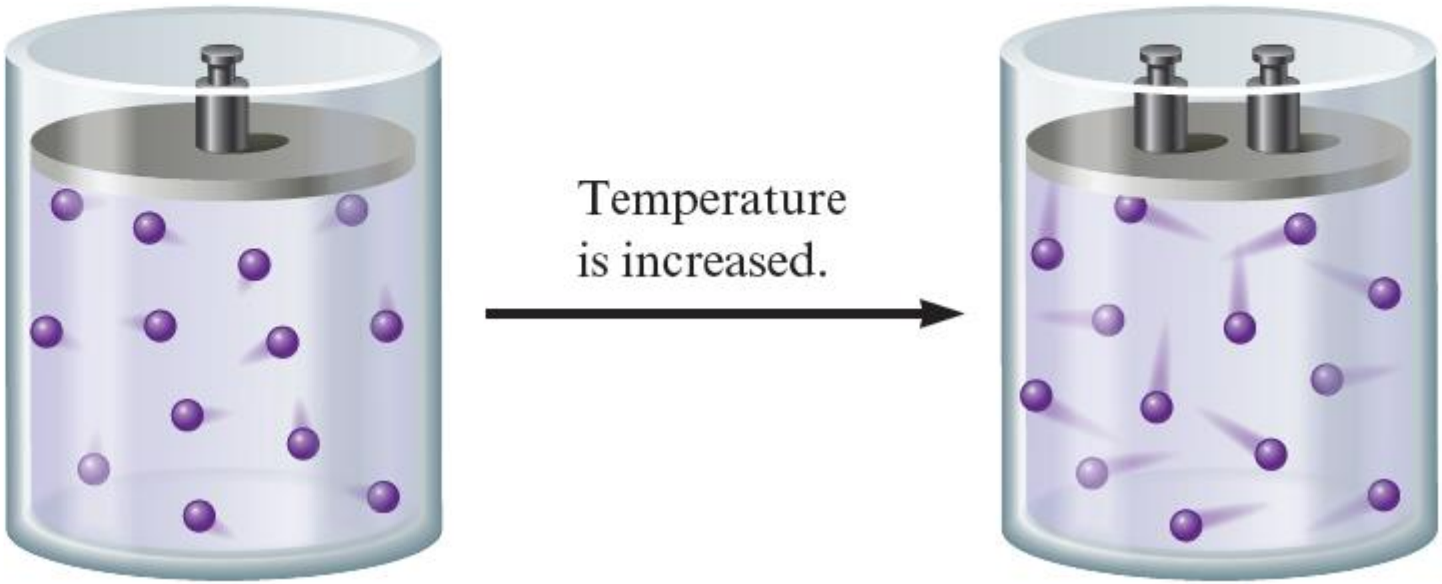
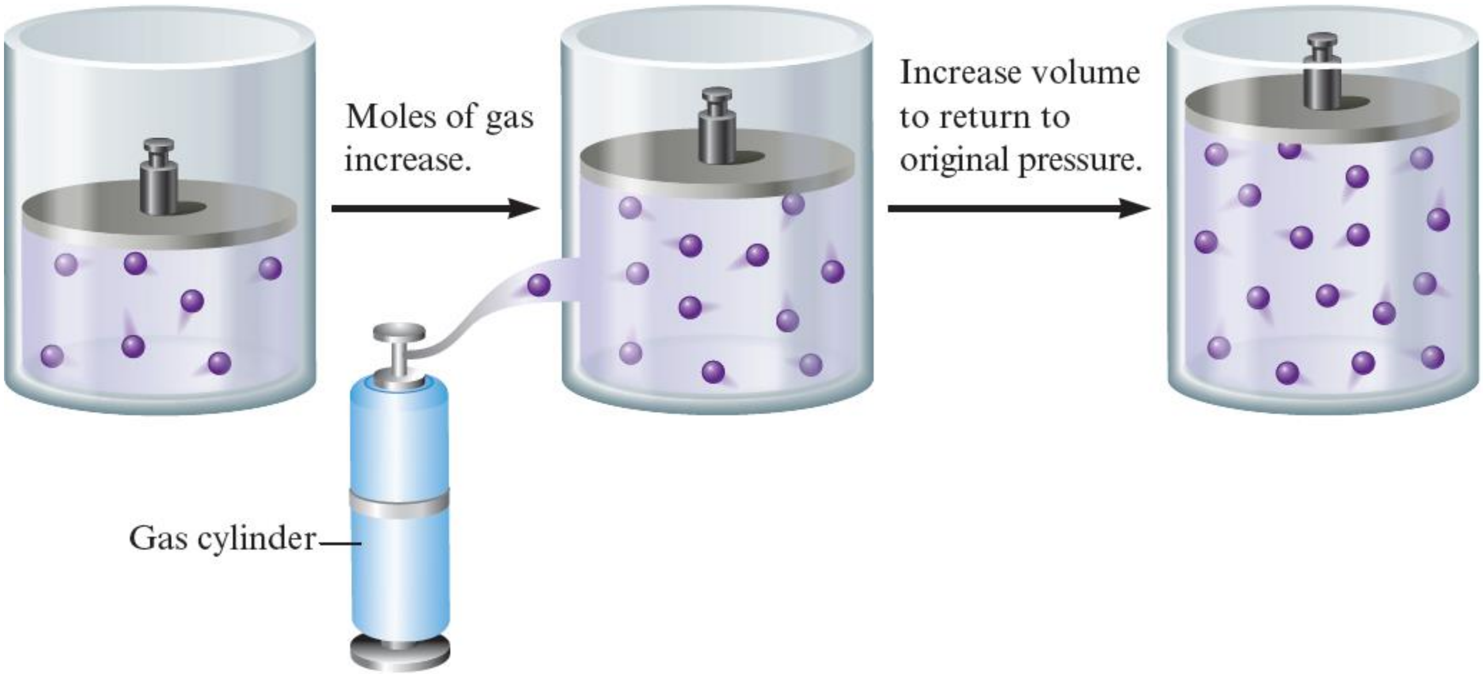


Figure 5.16 - Effects of Increasing the Temperature of a Sample of Gas at Constant Volume



Volume and Number of Moles (Avogadro's Law)
(Continued)

- Pressure can return to its original value if volume is increased

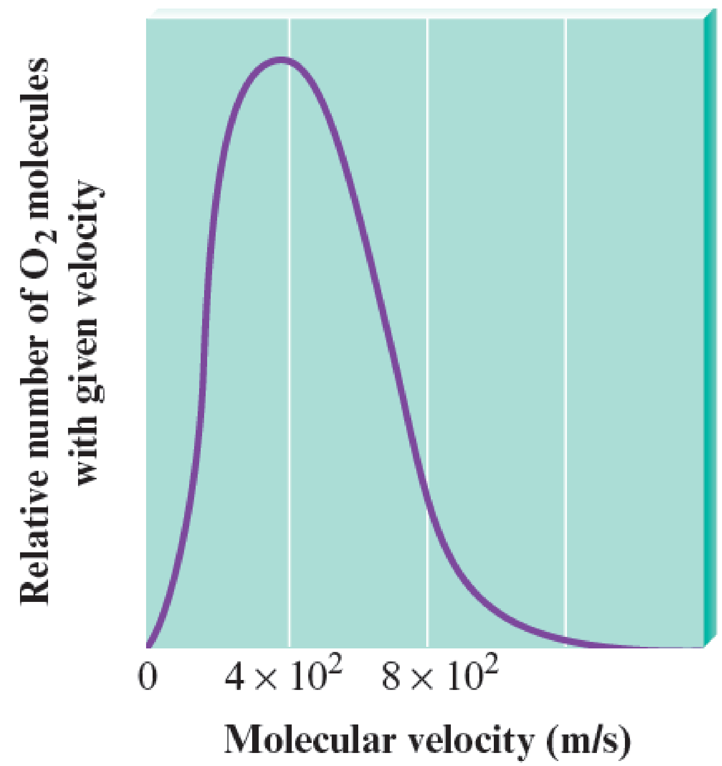


Mixture of Gases (Dalton's Law)

- Total pressure exerted by a mixture of gases is the sum of the pressures of the individual gases
- The KMT assumes that:
  - All gas particles are independent of one another
  - Volume of individual particles are not important

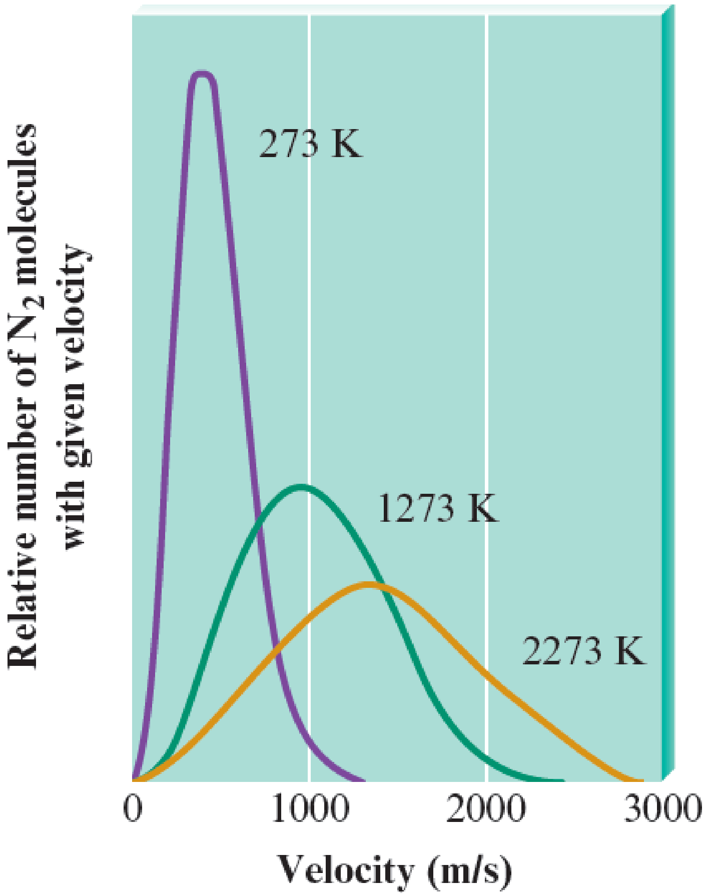
Effect of Collisions among Gas Particles

- When particles collide and exchange kinetic energy, a large range of velocities is produced
  - This plot depicts the relative number of O_2 molecules that have a given velocity at STP



Effect of Temperature on Velocity Distribution

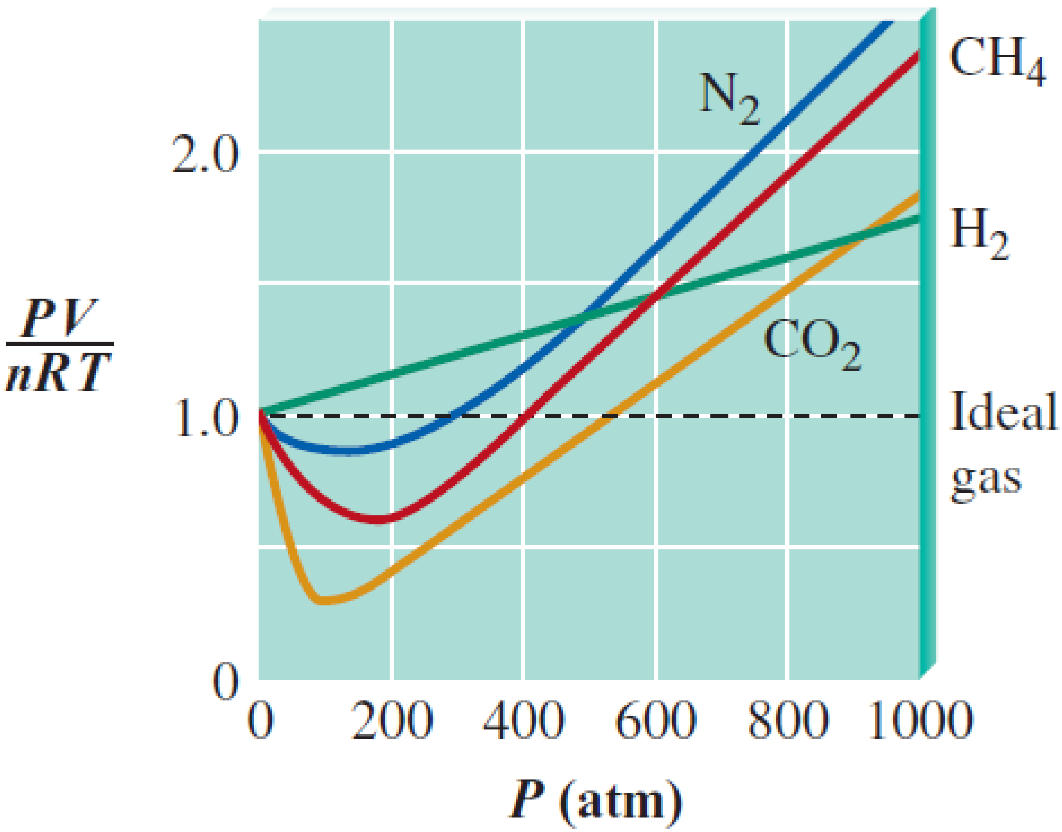
- As the temperature increases, the range of velocities becomes larger
  - Peak of the curve reflects the most probable velocity



Ideal Gas Behavior

- Exhibited by real gases under certain conditions of:
  - Low pressure
  - High temperature

Figure 5.25 - Plots of *PV/nRT* versus *P* for Several Gases (200 K)



Answer with your partner, then compare with another group

- You have learned that no gases behave perfectly ideally, but under conditions of high temperature and low pressure (high volume), gases behave more ideally
  - What if all gases always behaved perfectly ideally?
    - How would the world be different?