States of Matter
Lesson 4.5

CHEMISTRY 2
HONORS

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Further Applications of the Ideal-Gas Equation

Gas Densities and Molar Mass

- Density has units of mass over volume.
- Rearranging the ideal-gas equation with $\mathcal{M}$ as molar mass we get

$$PV = nRT$$

$$\frac{n}{V} = \frac{P}{RT}$$

$$\frac{n\mathcal{M}}{V} = d = \frac{P\mathcal{M}}{RT}$$

Examples

1. What is the density of nitrogen gas at 35 C and 1.50 atm?

2. What is the density of chlorine gas at 100.5 and 135.2 kPa?

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   1.67 g/L

2. What is the density of chlorine gas at 100.5 C and 135.2 kPa?

   3.09 g/L

- The molar mass of a gas can be determined as follows:

$$\mathcal{M} = \frac{dRT}{P}$$

OR

$$M = \frac{mRT}{PV}$$

Volumes of Gases in Chemical Reactions

- The ideal-gas equation relates *P*, *V*, and *T* to number of moles of gas.
- The *n* can then be used in stoichiometric calculations.

Examples

1. 4.83 g of a gas occupy a 1.50 L flask at 25 C and 112.4 kPa. What is the molar mass of the gas?

   71.0 g/mol

2. An evacuated 2.00-L flask has a mass of 225.34 g. A gas, at 75 C and 1.25 atm pressure, causes the mass of the flask to increase to 232.68 g. What is the molar mass of the gas?

   83.8 g/mol

Gas Mixtures and Partial Pressures

- Since gas molecules are so far apart, we can assume they behave independently.
- Dalton's Law: in a gas mixture the total pressure is given by the sum of partial pressures of each component:

$$P_{\text{total}} = P_1 + P_2 + P_3 + \cdots$$

- Each gas obeys the ideal gas equation:

$$P_i = n_i \left(\frac{RT}{V} \right)$$

- Combining the equations

$$P_{\text{total}} = (n_1 + n_2 + n_3 + \cdots)\left(\frac{RT}{V}\right)$$

Partial Pressures and Mole Fractions

- Let n_i be the number of moles of gas i exerting a partial pressure P_i, then

$$P_i = X_i P_{\text{total}}$$

where X_i is the **mole fraction** (n_i/n_t).

Examples:

1. Hydrogen gas is added to a 2.00-L flask at a pressure of 5.6 atm. Oxygen gas is added until the total pressure in the flask measures 8.4 atm. What is the mole fraction of hydrogen in the flask?

2. 1.35 moles of argon and 2.75 moles of neon are placed in a 15.0-L tank at 35 C. What is the total pressure in the flask? What is the pressure exerted by neon?

Examples:

1. Hydrogen gas is added to a 2.00-L flask at a pressure of 5.6 atm. Oxygen gas is added until the total pressure in the flask measures 8.4 atm. What is the mole fraction of hydrogen in the flask?

   0.67

2. 1.35 moles of argon and 2.75 moles of neon are placed in a 15.0-L tank at 35 C. What is the total pressure in the flask? What is the pressure exerted by neon? P_{tot} = 6.91 atm

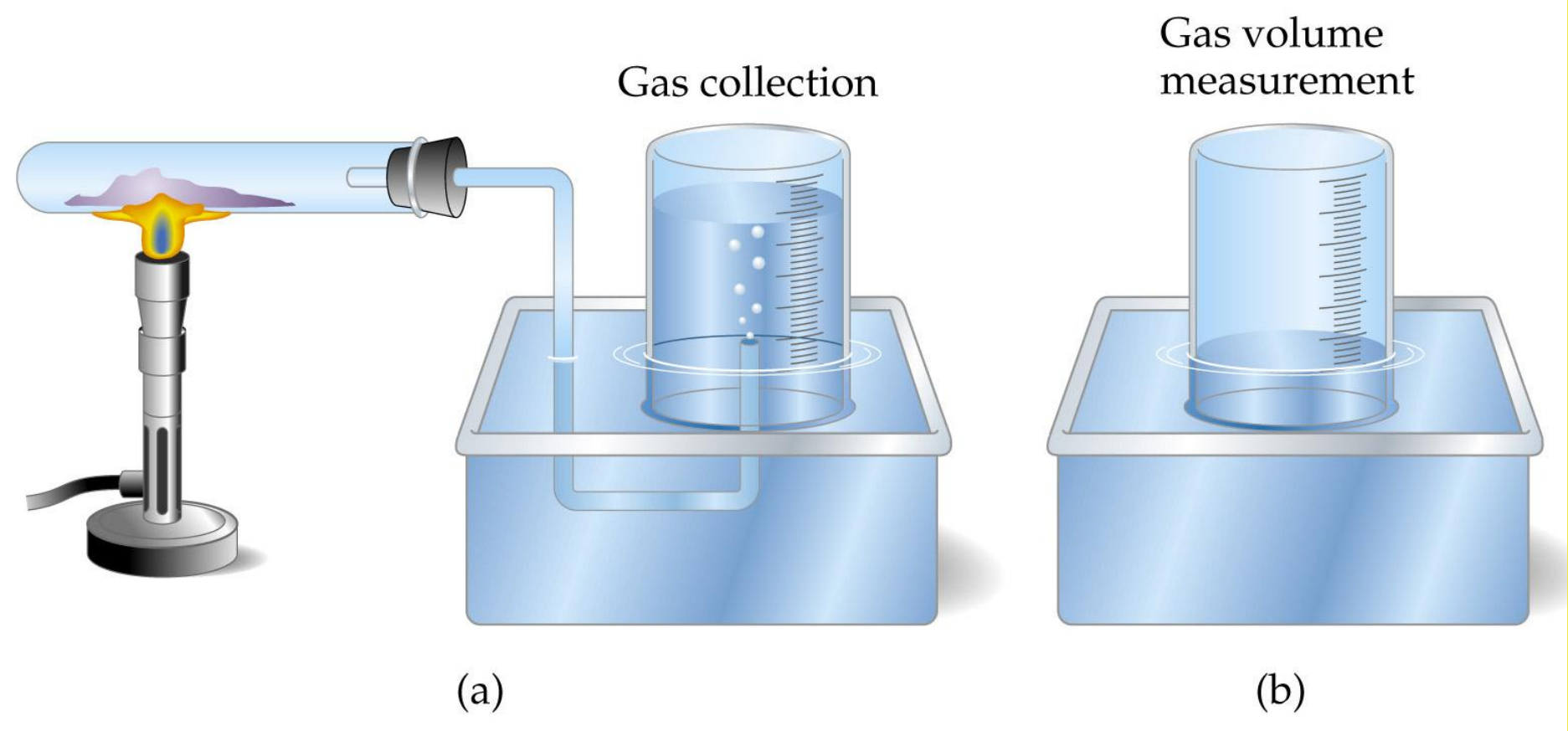
   P_{Ne} = 4.63 atm

Collecting Gases over Water

- It is common to synthesize gases and collect them by displacing a volume of water.
- To calculate the amount of gas produced, we need to correct for the partial pressure of the water:

$$P_{\text{total}} = P_{\text{gas}} + P_{\text{water}}$$

Collecting Gases over Water



Example - 150.82 mL of an unknown gas is collected over water at 27 C and 1.032 atm. The mass of the gas is 0.1644 g. What is the molar mass of the gas? The vapor pressure of water at 27 C is 26.74 torr

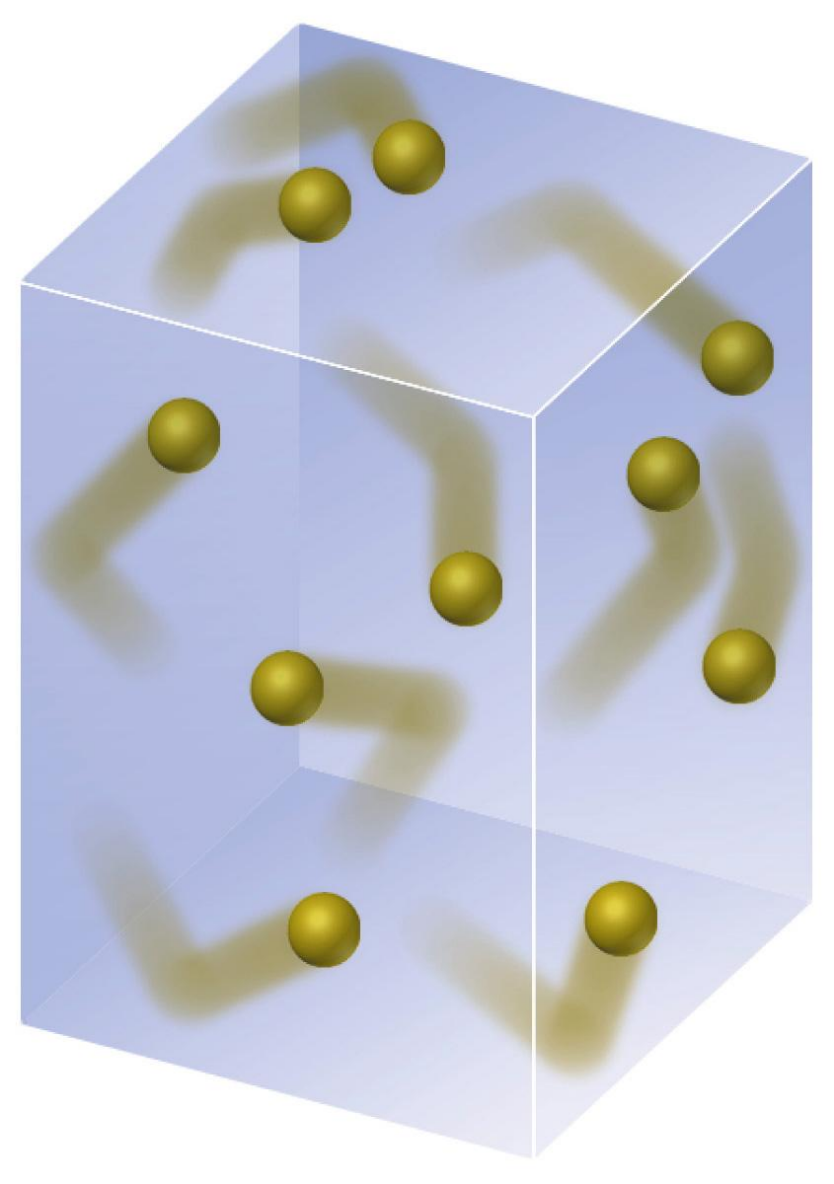
26.9 g/mol

Kinetic Molecular Theory

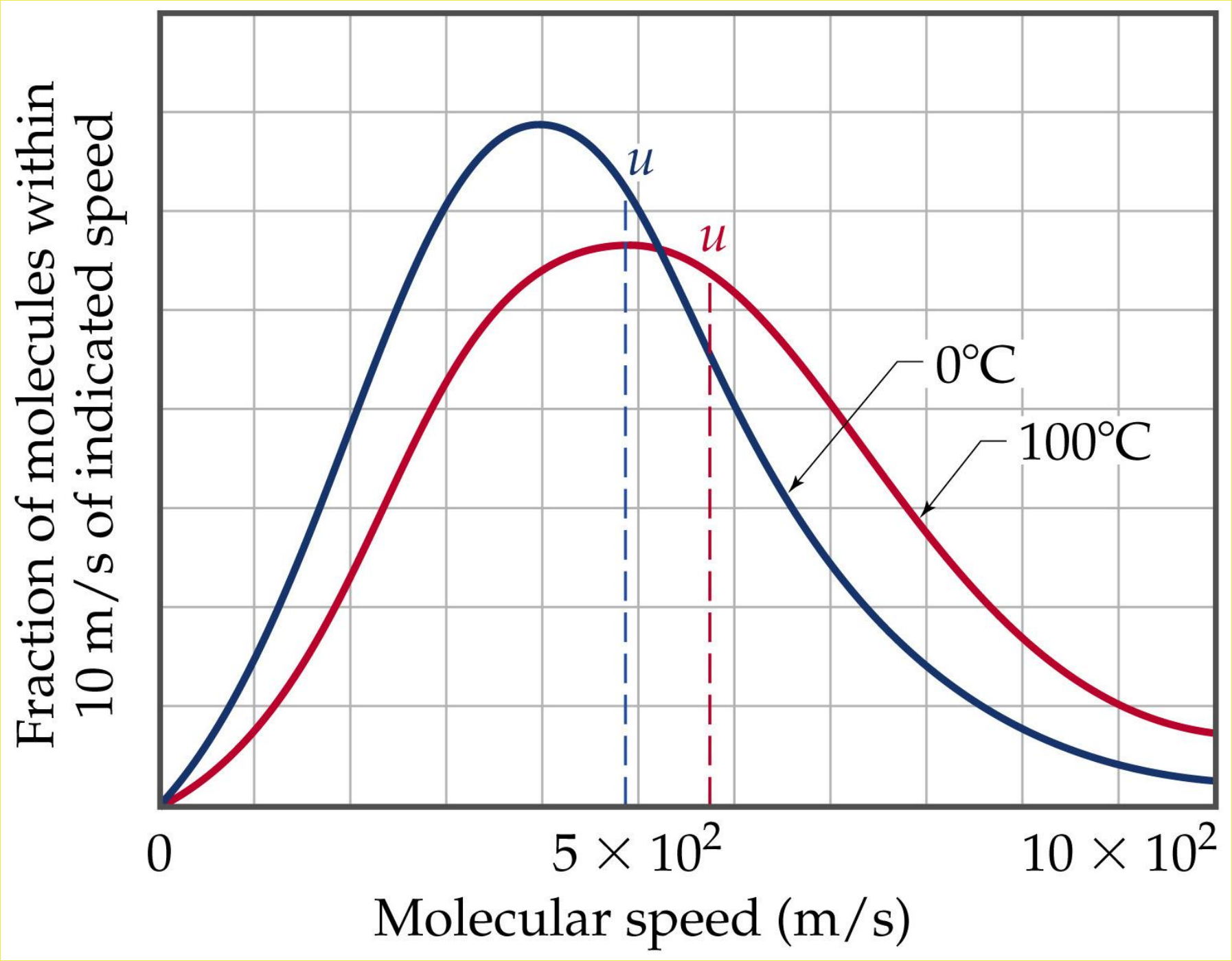
- Theory developed to explain gas behavior.
- Theory of moving molecules.
- Assumptions:
  - Gases consist of a large number of molecules in constant random motion.
  - Volume of individual molecules negligible compared to volume of container.
  - Intermolecular forces (forces between gas molecules) negligible.

- Assumptions:
  - Energy can be transferred between molecules, but total kinetic energy is constant at constant temperature.
  - Average kinetic energy of molecules is proportional to temperature.
- Kinetic molecular theory gives us an understanding of pressure and temperature on the molecular level.
- Pressure of a gas results from the number of collisions per unit time on the walls of container.

- Magnitude of pressure given by how often and how hard the molecules strike.
- Gas molecules have an average kinetic energy.
- Each molecule has a different energy.



- There is a spread of individual energies of gas molecules in any sample of gas.
- As the temperature increases, the average kinetic energy of the gas molecules increases.


Fraction of molecules within 10 m/s of indicated speed
u
u
0°C
100°C
0
5 × 10²
10 × 10²
Molecular speed (m/s)

- As kinetic energy increases, the velocity of the gas molecules increases.
- Root mean square speed, *u*, is the speed of a gas molecule having average kinetic energy.
- Average kinetic energy, ε, is related to root mean square speed:

$$\varepsilon = \frac{1}{2}mu^2$$

Application to Gas Laws

- As volume increases at constant temperature, the average kinetic energy of the gas remains constant. Therefore, *u* is constant. However, volume increases so the gas molecules have to travel further to hit the walls of the container. Therefore, pressure decreases.
- If temperature increases at constant volume, the average kinetic energy of the gas molecules increases. Therefore, there are more collisions with the container walls and the pressure increases.