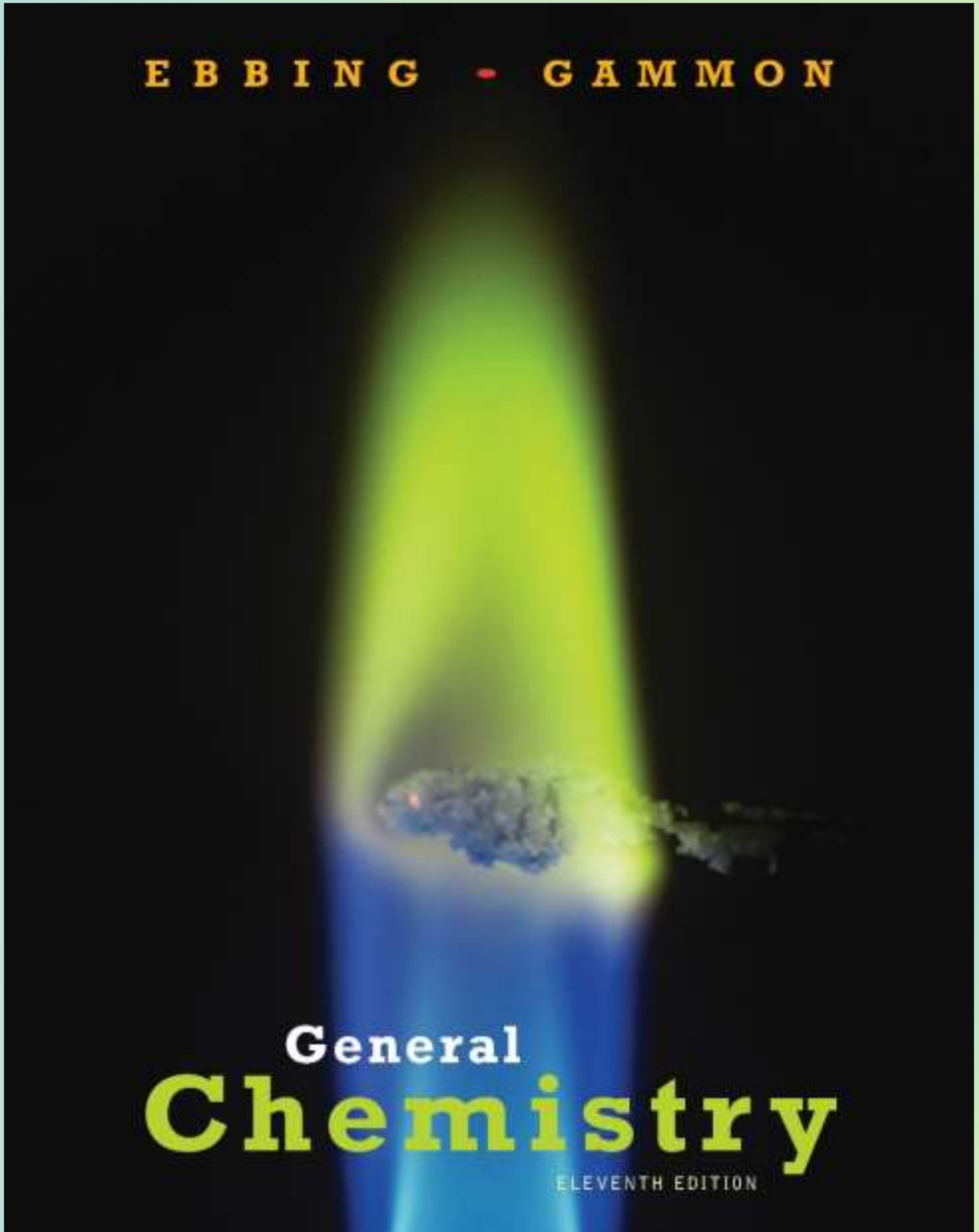
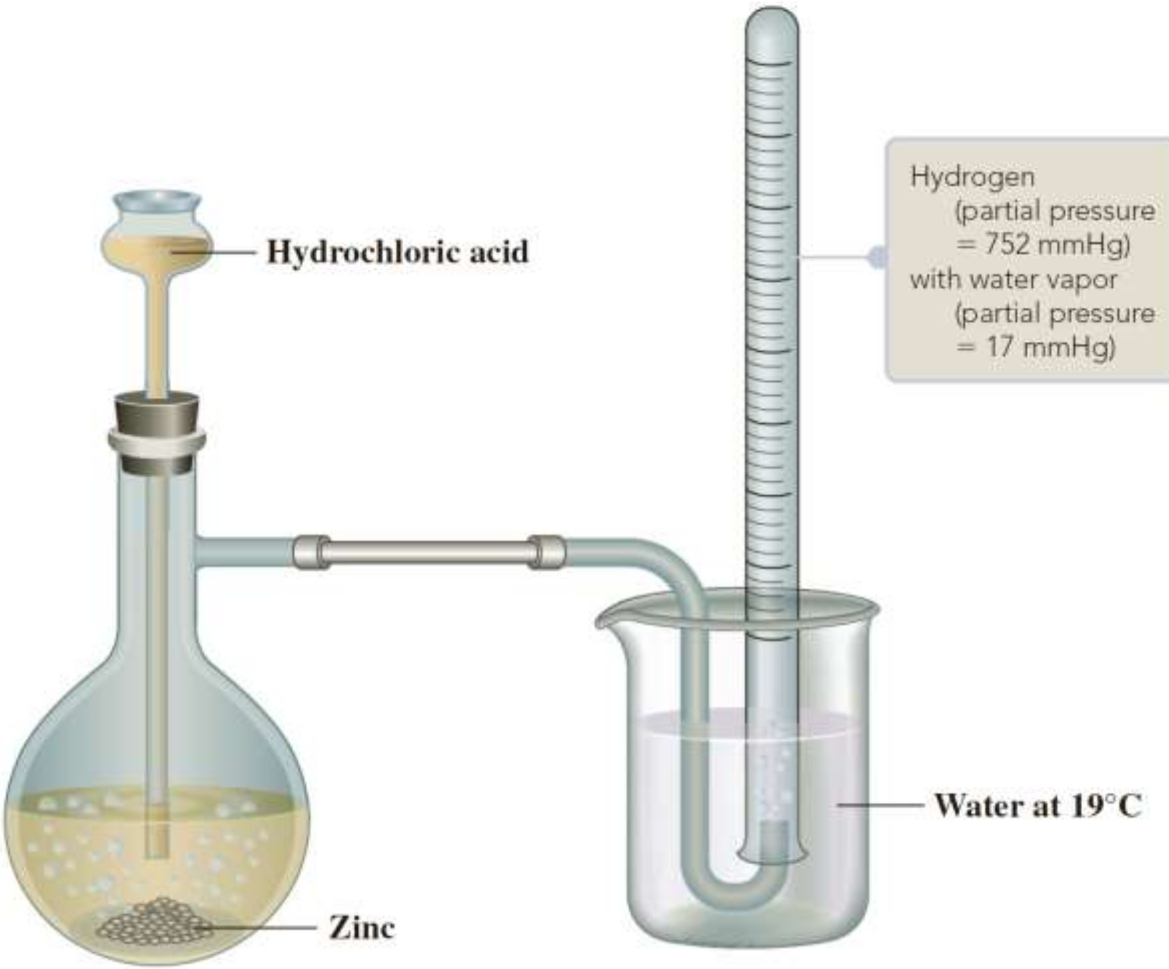


اللهم صلّ وسلّم على نبينا محمد وعلى آله وصحبه أجمعين

The Gaseous State

➢ Collecting Gases over Water



Vapor Pressure of Water at Various Temperatures*

Temperature (°C)	Pressure (mmHg)
0	4.6
10	9.2
15	12.8
17	14.5
19	16.5
21	18.7
23	21.1
25	23.8
27	26.7
30	31.8
40	55.3
60	149.4
80	355.1
100	760.0

Example 5.11

Hydrogen gas is produced according to the following reaction:

$$2HCl(aq) + Zn(s) \longrightarrow ZnCl_2(aq) + H_2(g)$$

The gas is collected over water. If 156 mL of gas is collected at 19°C and 769 mmHg total pressure, what is the mass of hydrogen collected? The vapor pressure of water at 19°C is 16.5 mmHg من الجدول:

$$P_{H_2}V = n_{H_2}RT$$

$$n_{H_2} = \frac{P_{H_2}V}{RT}$$

$$n_{H_2} = \frac{0.99 \times 156 \times 10^{-3}}{0.0821 \times 292.15} = 0.00645 \, n\, H_2$$

$$N_{H_2} = \frac{m_{H_2}}{Mm\, H_2}$$

$$0.00645 \times 2.016 =$$

$$H_2 = 0.0129\, g$$

$$19\,C° + 273 \rightarrow 292.15$$

$$P_{H_2} + P_{H_2O} = P_T$$

$$16.5 + P_{H_2} = 769$$

$$752.5\ mmHg = P_{H_2}$$

$$\frac{752.5}{760} = 0.99\ atm$$

Example 5.11 Calculating the Amount of Gas Collected over Water

Gaining Mastery Toolbox

Critical Concept 5.11
When gases are collected over water, there will always be water molecules (water vapor) present in the gas mixture. The partial pressure of water vapor depends *only* on temperature.

Solution Essentials:
- Vapor pressure
- Dalton's law of partial pressures
- Ideal gas law
- Units of pressure (Table 5.2)
- Kelvin temperature

Hydrogen gas is produced by the reaction of hydrochloric acid, HCl, on zinc metal.

$$2HCl(aq) + Zn(s) \longrightarrow ZnCl_2(aq) + H_2(g)$$

The gas is collected over water. If 156 mL of gas is collected at 19°C (two significant figures) and 769 mmHg total pressure, what is the mass of hydrogen collected?

Problem Strategy The gas collected is hydrogen mixed with water vapor. To obtain the amount of hydrogen, you must first find its partial pressure in the mixture, using Dalton's law (Step 1). Then you can calculate the moles of hydrogen from the ideal gas law (Step 2). Finally, you can obtain the mass of hydrogen from the moles of hydrogen (Step 3).

Solution

Step 1: The vapor pressure of water at 19°C is 16.5 mmHg. From Dalton's law of partial pressures, you know that the total gas pressure equals the partial pressure of hydrogen, P_{H_2}, plus the partial pressure of water, P_{H_2O}.

$$P = P_{H_2} + P_{H_2O}$$

Substituting and solving for the partial pressure of hydrogen, you get

$$P_{H_2} = P - P_{H_2O} = (769 - 16.5)\text{ mmHg} = 752\text{ mmHg}$$

Step 2: Now you can use the ideal gas law to find the moles of hydrogen collected. The data are

Variable	***Value***
P	$752\text{ mmHg} \times \dfrac{1\text{ atm}}{760\text{ mmHg}} = 0.989\text{ atm}$
V	$156\text{ mL} = 0.156\text{ L}$
T	$(19 + 273)\text{ K} = 292\text{ K}$
n	?

From the ideal gas law, $PV = nRT$, you have

$$n = \frac{PV}{RT} = \frac{0.989\text{ atm} \times 0.156\text{ L}}{0.0821\text{ L}\cdot\text{atm/(K}\cdot\text{mol)} \times 293\text{ K}} = 0.00641\text{ mol}$$

Step 3: You convert moles of H_2 to grams of H_2.

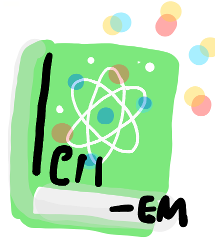
$$0.00664\text{ mol }H_2 \times \frac{2.02\text{ g }H_2}{1\text{ mol }H_2} = \mathbf{0.0130\text{ g }H_2}$$

Exercise 5.11 Oxygen can be prepared by heating potassium chlorate, $KClO_3$, with manganese dioxide as a catalyst. The reaction is

$$2KClO_3(s) \xrightarrow[MnO_2]{\Delta} 2KCl(s) + 3O_2(g)$$

How many moles of O_2 would be obtained from 1.300 g $KClO_3$? If this amount of O_2 were collected over water at 23°C and at a total pressure of 745 mmHg, what volume would it occupy?

See Problems 5.87 and 5.88.



* $PV = nRt \rightarrow O_2$ Volume

$$0.0159 = \frac{[0.9525] \times V}{0.0821 \times (23 + 273.15)} = 0.4058\ L$$

* $P_{total} = PO_2 + P_{water}$

745 mmHg = PO_2 + 21.1 mmHg – من الجدول

* $1.300\,g\,KClO_3 \times \frac{1}{122.55} \times \frac{3}{2} = 0.0159$ mol O_2

(Q)An unknown gas was collected by water displacement. The following data was recorded: T = 27.0 °C; P = 750 torr; V = 37.5 mL; Gas mass = 0.0873 g;

$P_{H_2O(vap)}$ = 26.98 torr → من الجدول

Determine the molecular weight of the gas.

A. 5.42 g/mol

B. 30.2 g/mol

C. 60.3 g/mol

D. 58.1 g/mol

E. 5.81 g/mol

➢ Kinetic-Molecular Theory

According to this theory, *a gas consists of molecules in constant random motion*.

Kinetic energy, E_k, is the energy associated with the motion of an object of mass *m*.

$$E_k = \frac{1}{2}m \times (\text{speed})^2$$

Kinetic-theory model of gas pressure According to kinetic theory, gas pressure is the result of the bombardment of the container walls by constantly moving molecules.

5.6 Kinetic Theory of an Ideal Gas

➢ Postulates of Kinetic Theory

Postulate 1: *Gases are composed of molecules whose size is negligible compared with the average distance between them*. Most of the volume occupied by a gas is (empty space). This means that you can usually ignore the volume occupied by the molecules.

تشغله..

Postulate 2: *Molecules move randomly in straight lines in all directions and at various speeds.*

This means that properties of a gas that depend on the motion of molecules, such as pressure, will be the same in all directions.

Postulate 3: *The forces of attraction or repulsion between two molecules (intermolecular forces) in a gas are very weak or negligible, except when they collide.*

This means that a molecule will continue moving in a straight line with undiminished speed until it collides with another gas molecule or with the walls of the container.

Postulate 4: *When molecules collide with one another, the collisions are elastic.* In an elastic collision, the total kinetic energy remains constant; (no kinetic energy is lost)

Postulate 5: *The average kinetic energy of a molecule is proportional to the absolute temperature*

CONCEPT CHECK 5.5

Consider a sealed glass bottle of helium gas at room temperature. If you immerse the bottle in an ice water bath, how will this immersion affect the pressure of the gas?

I. Pressure increase

II. Pressure decrease

III. No pressure change

The best explanation for your answer is?

ينخفض تواتر وقوة
اصطدامات ذرات الهيليوم
بجدران الحاوية مع
انخفاض درجة الحرارة.

a The force of the collisions of the helium atoms with the container walls increases with decreasing temperature.

b The helium atoms have significantly greater volumes at lower temperatures.

c The frequency and force of the collisions of the helium atoms with the container walls decrease with decreasing temperature.

d The average kinetic energy of the helium atoms remains constant when the temperature decreases.

e The concentration of helium decreases, resulting in a decrease in the frequency of the collisions with the container walls.

The Ideal Gas Law from Kinetic Theory

$$P \propto \text{frequency of collisions} \times \text{average force}$$

$$P \propto \left(u \times \frac{1}{V} \times N\right) \times mu$$

$$PV \propto Nmu^2$$

$$PV \propto nT$$

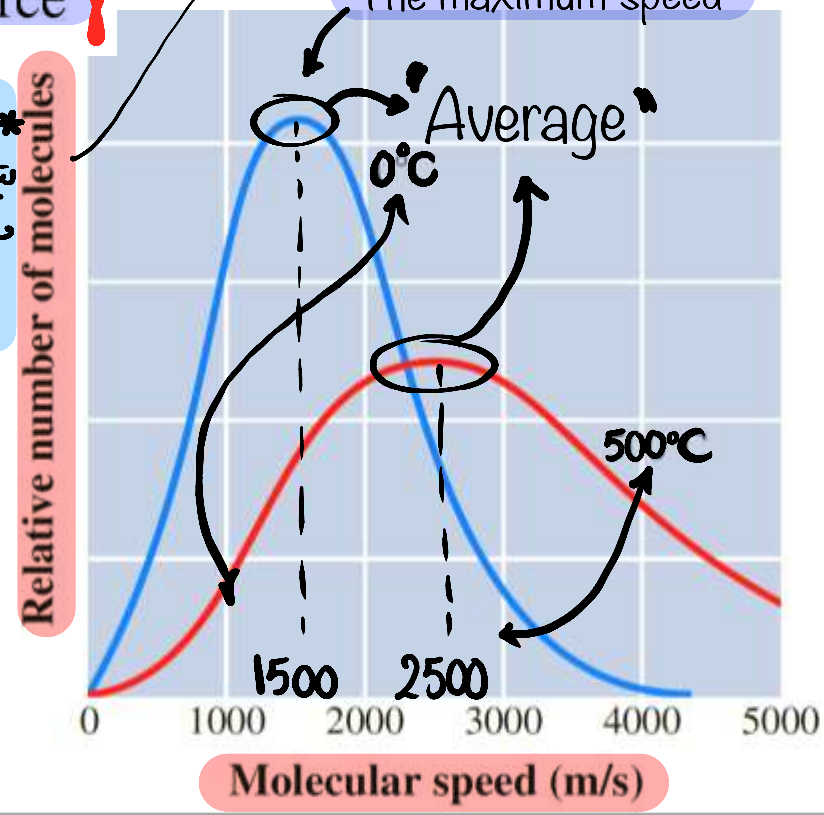
$$\boxed{PV = nRT}$$

5.7 Molecular Speeds; Diffusion and Effusion

- **Molecular Speeds**

root-mean-square (rms) molecular speed (u)

$$u = \sqrt{\frac{3RT}{M_m}} = \left(\frac{3RT}{M_m}\right)^{\frac{1}{2}}$$



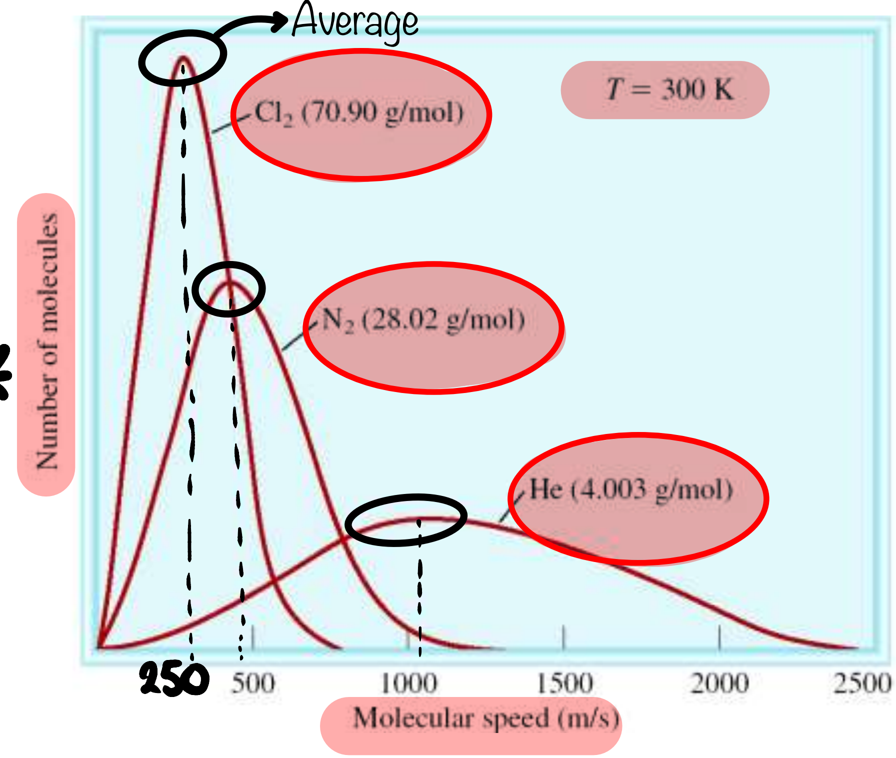
Maxwell's distribution of molecular speeds The distributions of speeds of H_2 molecules are shown for 0°C and 500°C. Note that the speed corresponding to the maximum in the curve (the most probable speed) increases with temperature.

$$\uparrow u = \sqrt{\frac{3RT^{*}}{M_m \downarrow}} = \left(\frac{3RT}{M_m}\right)^{\frac{1}{2}}$$

Constant

* R (= 8.314 kg·m²/(s²·K·mol)), *
* T (K), and M_m (kg/mol), *

(g) → kg



Example 5.12

(Q) Calculate the rms speed of O_2 molecules in a cylinder at 21°C and 15.7 atm

$$u = \left(\frac{3 \times 8.31\ \text{kg} \cdot \text{m}^2/(\text{s}^2 \cdot \text{K} \cdot \text{mol}) \times 294\ \text{K}}{32.0 \times 10^{-3}\ \text{kg/mol}}\right)^{\frac{1}{2}} = \mathbf{479\ m/s}$$

Exercise 5.13 At what temperature do hydrogen molecules, H_2, have the same rms speed as nitrogen molecules, N_2, at 455°C? At what temperature do hydrogen molecules have the same average kinetic energy?

Determine the rms molecular speed for N_2 at 455°C (728 K):

(Speed)

1. $$u = \left(\frac{3RT}{M}\right)^{\frac{1}{2}} = \left(\frac{3 \times 8.31\ \text{kg} \bullet \text{m}^2/(\text{s}^2 \bullet \text{K} \bullet \text{mol}) \times 728\ \text{K}}{28.02 \times 10^{-3}\ \text{kg/mol}}\right)^{\frac{1}{2}} = 804.81\ \text{m/s}$$

2. $$T = \frac{u^2 M}{3R} = \frac{(804.81\ \text{m/s})^2 (2.016 \times 10^{-3}\ \text{kg/mol})}{(3)(8.31\ \text{kg} \bullet \text{m}^2/\text{s}^2 \bullet \text{K} \bullet \text{mol})} = 52.4\ \text{K}$$

Reminder!

Any two gases at the same temperature will have the same average kinetic energy

Because the average kinetic energy of a molecule is proportional to only T

➢ Diffusion and Effusion

✓ **Diffusion** is *the process whereby a gas spreads out through another gas to occupy the space uniformly*.

✓ **Effusion** is *the process in which a gas flows through a small hole in a container*.

Average speed :

$$u = \sqrt{\frac{3RT}{M_m}}$$

✓ **Graham's law of effusion**

Graham's law of effusion:

$$\text{Rate of effusion of molecules} \propto \frac{1}{\sqrt{M_m}}$$

1. (for the same container at 2. constant T and P)

Speed

Time

$$\frac{\text{Rate }(1)}{\text{Rate }(2)} = \frac{T_{(2)}}{T_{(1)}} = \sqrt{\frac{M_{m\,(2)}}{M_{m\,(1)}}}$$

(Q) Calculate the ratio of effusion rates of molecules of carbon dioxide, CO_2, and sulfur dioxide, SO_2, from 1. the same container and 2. at the same temperature and 3. pressure. Example 5.13

$$\frac{\text{Rate of effusion of } CO_2}{\text{Rate of effusion of } SO_2} = \frac{\dfrac{1}{\sqrt{M_m(CO_2)}}}{\dfrac{1}{\sqrt{M_m(SO_2)}}} \rightarrow$$

$M_m\ CO_2 < M_m\ SO_2$

Effusion > Effusion

$$\frac{\text{Rate of effusion of } CO_2}{\text{Rate of effusion of } SO_2} = \sqrt{\frac{M_m(SO_2)}{M_m(CO_2)}} = \sqrt{\frac{64.1\ \text{g/mol}}{44.0\ \text{g/mol}}} = \mathbf{1.21}$$

- Sulfur dioxide effusses 0.82644 times slower than Carbon dioxide
- carbon dioxide effuses 1.21 times faster than sulfur dioxide

Exercise 5.12 What is the rms speed (in m/s) of a carbon tetrachloride molecule at 22°C? → CCl_4

See Problems 5.89, 5.90, 5.91, and 5.92.

$$u = \sqrt{\frac{3RT}{M_m}} = \sqrt{\frac{3 \times 8.31 \times 295}{153.82 \times 10^{-3}}} = 218.65\ \text{m/s}$$

T → K; M_m → kg

g → kg

Exercise 5.14 If it takes 3.52 s for 10.0 mL of He to effuse through a hole in a container at a particular temperature and pressure, how long would it take for 10.0 mL of O_2 to effuse from the same container at the same temperature and pressure? (Note that the **rate of effusion can be given in terms of volume of gas effused per second**.)

Mm He < Mm O_2

$$\frac{\text{Rate of effusion of } O_2}{\text{Rate of effusion of He}} = \sqrt{\frac{M_m(\text{He})}{M_m(O_2)}} = \sqrt{\frac{4.00 \text{ g/mol}}{32.00 \text{ g/mol}}} = 0.35$$

بالتالي O_2 أبطأ من الـ He ⇐ بمقدار

→ Rate of effusion of O_2 = 0.35 × rate of effusion of He.

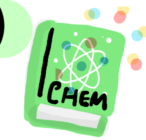
$$\frac{\text{Volume of } O_2}{\text{Time for } O_2} = 0.35 \times \frac{\text{Volume of He}}{\text{Time for He}}$$

* Speed He > Speed O_2
* Time He < Time O_2
* 3.52 s < 9.96 s

$$\frac{10.0 \text{ mL}}{\text{Time for } O_2} = 0.35 \times \frac{10.0 \text{ mL}}{3.52 \text{ s}}$$

$$\text{Time for } O_2 = \frac{3.52 \text{ s}}{0.35355} = 9.96 \text{ s}$$

Exercise 5.15 If it takes 4.67 times as long for a particular gas to effuse as it takes hydrogen under the same conditions, what is the molecular weight of the gas? (Note that the rate of effusion is inversely proportional to the time it takes for a gas to effuse.)

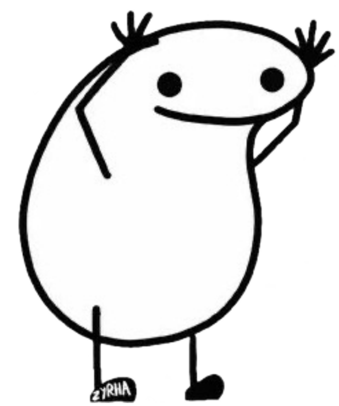


$$\frac{\text{Rate of effusion of } H_2}{\text{Rate of effusion of gas}} = \frac{\text{time for gas}}{\text{time for } H_2} = \sqrt{\frac{M_m(\text{gas})}{M_m(H_2)}} = 4.67$$

$$M_m(\text{gas}) = (4.67)^2 \times M_m(H_2) = (4.67)^2 \times 2.016 \text{ g/mol} = 43.96 \text{g/mol}$$

(Q) For the series of gases He, Ne, Ar, H_2, and O_2 what is the order of increasing rate of effusion?

Substance	He	Ne	Ar	H_2	O_2
MM	4	20	40	2	32



Lightest are fastest: $H_2 > He > Ne > O_2 > Ar$

Same as: $Ar < O_2 < Ne < He < H_2$

Small ⟶ high

Learning Objectives

Important Terms

5.1 Gas Pressure and Its Measurement

- Define *pressure* and its units.
- Convert units of pressure. **Example 5.1**

kg · m / s² ; $\frac{F}{A} = \frac{mg}{A}$; Pressure gas - liquid

Important Terms:
- pressure
- pascal (Pa)
- barometer — *Atmospheric pressure*
- manometer
- millimeters of mercury (mmHg or torr)
- atmosphere (atm) — *1 atm → 1.01 × 10⁵ pas; 1.01 KPas*
- bar

5.2 Empirical Gas Laws

- Express *Boyle's law* in words and as an equation.
- Use Boyle's law. **Example 5.2**
- Express *Charles's law* in words and as an equation.
- Use Charles's law. **Example 5.3**
- Express the *combined gas law* as an equation.
- Use the combined gas law. **Example 5.4**
- State *Avogadro's law*.
- Define *standard temperature and pressure* (*STP*).

$P_iV_i = P_fV_f$ (T, N) ; $\frac{V_i}{T_i} = \frac{V_f}{T_f}$ (P, N*) ; $\frac{VP}{T} = \frac{PV}{T}$; AT given (T, P) STP T=0°C, P = 1 atm ; 1 mole = 22.4 L ; the Volume of one mole of the gas (= 22.4 L)*

Important Terms:
- Boyle's law
- Charles's law
- Avogadro's law
- molar gas volume (V_m)
- standard temperature and pressure (STP)

5.3 The Ideal Gas Law

- State what makes a gas an *ideal gas*.
- Learn the *ideal gas law* equation.
- Derive the empirical gas laws from the ideal gas law. **Example 5.5**
- Use the ideal gas law. **Example 5.6**
- Calculate gas density. **Example 5.7**
- Determine the molecular weight of a vapor. **Example 5.8**
- Use an equation to calculate gas density.

PV = nRT ; PMm = dRT

Important Terms:
- molar gas constant (R) — *0.0821 (atm.L / mol . K) ; 8.134 (kg . . . .)*
- ideal gas law

You
What makes the gas to be an ideal gas?

ChatGPT
An ideal gas is a theoretical concept used in physics and chemistry to simplify the behavior of real gases under certain conditions. For a gas to be considered ideal, it should adhere to the following assumptions:

1. **Negligible Volume:** The volume occupied by the gas particles is considered negligible compared to the overall volume of the gas.
2. **No Intermolecular Forces:** There are no attractive or repulsive forces between gas particles. This means no intermolecular forces like van der Waals forces or hydrogen bonding are present.
3. **Elastic Collisions:** Collisions between gas particles are perfectly elastic, meaning kinetic energy is conserved during collisions.
4. **Continuous, Random Motion:** Gas particles move in constant, random motion. Their velocity and direction change due to collisions.
5. **No Energy Loss:** There is no energy loss due to friction or other dissipative forces.

It's important to note that real gases deviate from ideal behavior under certain conditions, especially at high pressures and low temperatures. In such cases, corrections are applied using equations of state or other models to better describe the gas behavior.

5.4 Stoichiometry Problems Involving Gas Volumes

- Solve stoichiometry problems involving gas volumes. **Example 5.9**

gas → mole → mole → Volume $\frac{RT}{P}$

5.5 Gas Mixture; Law of Partial Pressures

= Dalton's law

- Learn the equation for *Dalton's law of partial pressures*.
- Define the *mole fraction* of a gas.
- Calculate the partial pressure and mole fractions of a gas in a mixture. **Example 5.10**
- Describe how gases are collected over water and how to determine the *vapor pressure* of water.
- Calculate the amount of gas collected over water. **Example 5.11**

TABLE ; $P_{Total} = P_{water} + P_{gas}$; $P_T = P_A + P_B + \ldots$ Mixture ; $X_A = \frac{n_A}{n_T} = \frac{P_A}{P_T}$ (T, V) constant

Important Terms:
- partial pressure
- Dalton's law of partial pressures
- mole fraction

5.6 Kinetic Theory of an Ideal Gas

- List the five postulates of the *kinetic theory*.
- Provide a qualitative description of the gas laws based on the kinetic theory.

kinetic-molecular theory of gases (kinetic theory)

5.7 Molecular Speeds; Diffusion and Effusion

- Describe how the *root-mean-square (rms) molecular speed* of gas molecules varies with temperature.
- Describe the molecular-speed distribution of gas molecules at different temperatures.
- Calculate the rms speed of gas molecules. **Example 5.12**
- Define *effusion* and *diffusion*.
- Describe how individual gas molecules move undergoing diffusion.
- Calculate the ratio of effusion rates of gases. **Example 5.13**

root-mean-square (rms) molecular speed
diffusion
effusion
Graham's law of effusion

5.8 Real Gases

- Explain how and why a *real gas* is different from an ideal gas.
- Use the van der Waals equation. **Example 5.14**

van der Waals equation

Key Equations

$PV = \text{constant} \quad (\text{constant } n, T)$

$\dfrac{V}{T} = \text{constant} \quad (\text{constant } n, P)$

$V_m = \text{specific constant}$ (depending on T, P; independent of gas)

$PV = nRT$

$PM_m = dRT$

$P = P_A + P_B + P_C + \cdots$

$\text{Mole fraction of } A = \dfrac{n_A}{n} = \dfrac{P_A}{P}$

$u = \sqrt{\dfrac{3RT}{M_m}} = \left(\dfrac{3RT}{M_m}\right)^{\frac{1}{2}}$

$\text{Rate of effusion} \propto \dfrac{1}{\sqrt{M_m}}$ (same container at constant T, P)

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

Done by: Joud Jaber