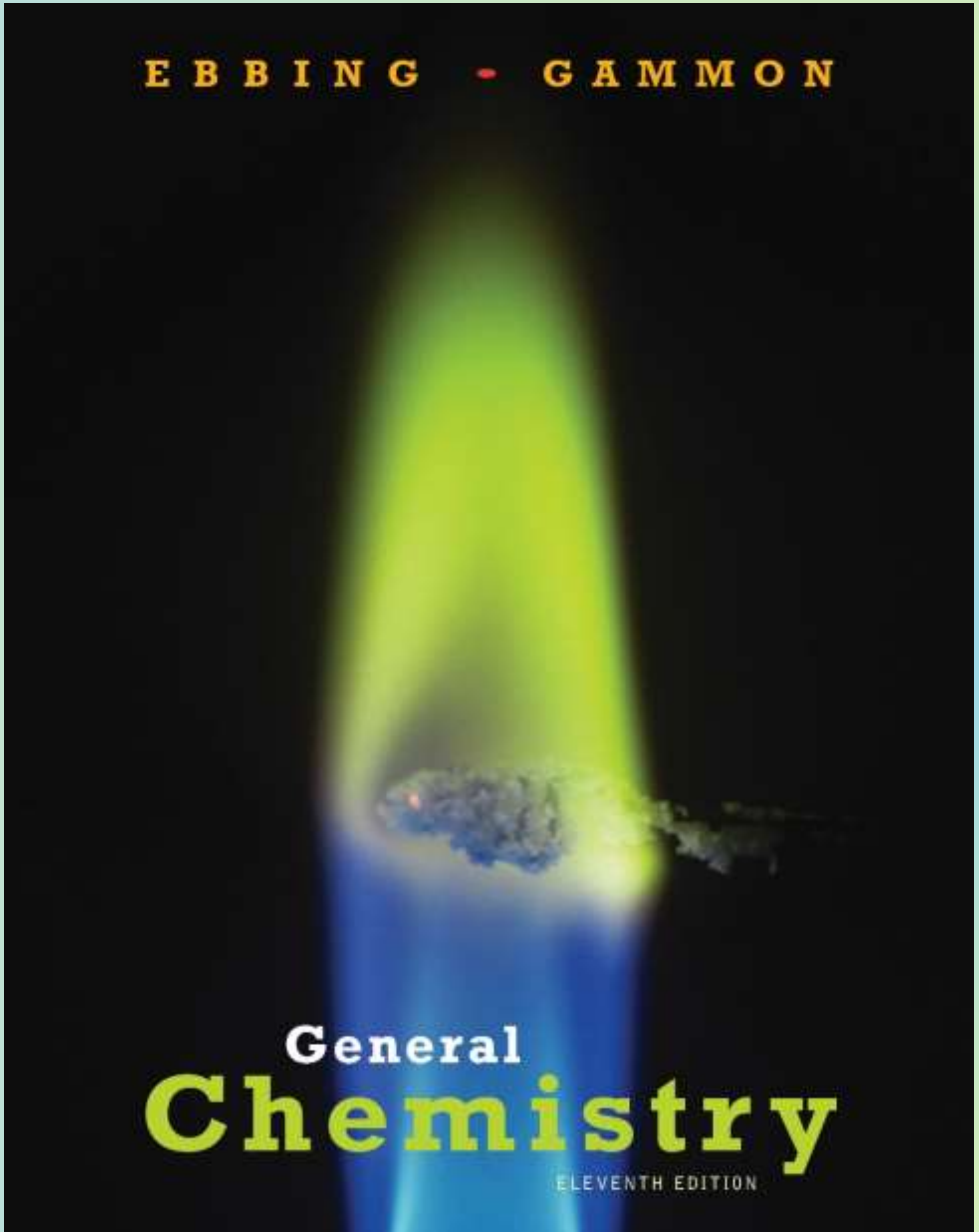


اللهم صلّ وسلّم على نبينا محمد وعلى آله وصحبه أجمعين

The Gaseous State

➢ Gas Laws

Low molecular weight

Most substances composed of (small molecules) 1. are gases under normal conditions or else are 2. easily vaporized liquids

(Liquid → gas)

Table 5.1 Properties of Selected Gases

Name	Formula	Color	Odor	Toxicity
Ammonia	NH_3	Colorless	Penetrating	Toxic
Carbon dioxide	CO_2	Colorless	Odorless	Nontoxic
Carbon monoxide	CO	Colorless	Odorless	Very toxic
Chlorine	Cl_2	Pale green	Irritating	Very toxic
Hydrogen	H_2	Colorless	Odorless	Nontoxic
Hydrogen sulfide	H_2S	Colorless	Foul	Very toxic
Methane	CH_4	Colorless	Odorless	Nontoxic
Nitrogen dioxide	NO_2	Red-brown	Irritating	Very toxic

5.1 Gas Pressure and Its Measurement

Pressure is defined as *the force exerted per unit area of surface*

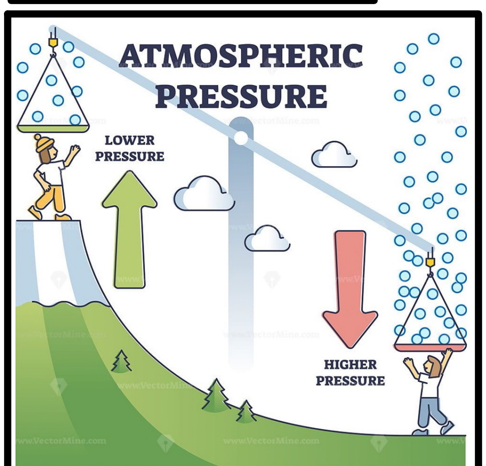
✓ Force = mass × constant acceleration of gravity

✓ *SI unit of pressure*, *kg/(m·s²)*, is given the name **pascal (Pa)**

① ✓ A **barometer** is *a device for measuring the pressure of the atmosphere*

atmospheric pressure

② ✓ A **manometer**, *a device that measures the pressure of a gas or liquid in a vessel*

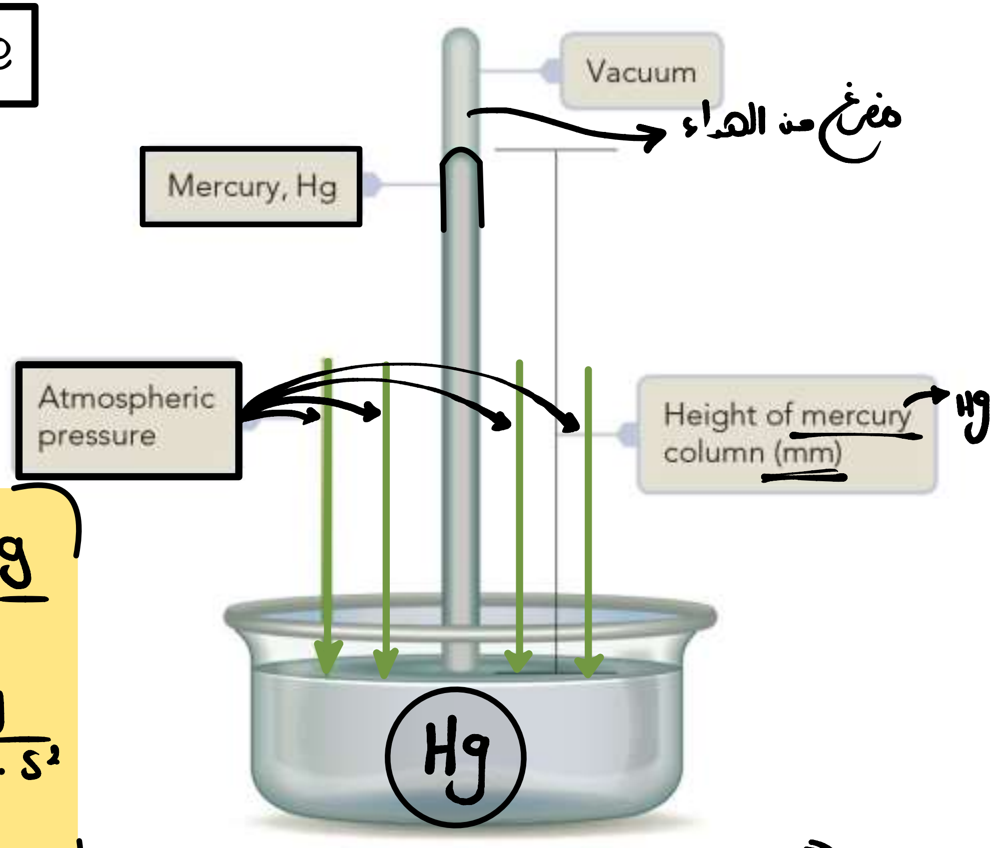


SI Unit:

$$P = \frac{F}{A} \Rightarrow \frac{m \times g}{a}$$

$$\frac{kg \cdot m}{m^2 \cdot s^2} = \frac{kg}{m \cdot s^2}$$

$$\rightarrow (kg/m.s^2)$$



mercury barometer

e.x: Pressure of a coin (9.3 mm in radius and 2.5 g) → r

Force = mass x g = (2.5×10^{-3} kg) × (9.81 m/s^2)

Area = π x (radius)^2 = 3.14 x (9.3×10^{-3} m)^2

For the Circle

2.5g x (9.8 ≈ 10 = g)

$$\text{Pressure} = \frac{\text{force}}{\text{area}} = \frac{(2.5 \times 10^{-2})\text{kg}\cdot\text{m/s}^2}{2.7 \times 10^{-4}\ \text{m}^2} = 93\ \text{kg/(m}\cdot\text{s}^2) = 93\ \text{Pa}$$

✓ The general relationship between the pressure P and the height h of a liquid column in a barometer or manometer is:

$$P = gdh$$

Gravitational acceleration (g), Density (d), Height (h)

Constant for the Substance

Table 5.2 Important Units of Pressure

Unit	Relationship or Definition
Pascal (Pa)	kg/(m·s^2)
Atmosphere (atm)	1 atm = 1.01325×10^5 Pa ≃ 101 kPa
mmHg, or torr	760 mmHg = 1 atm / 760 torr = 1 atm
Bar	1.01325 bar = 1 atm

Example 5.1 Converting Units of Pressure

(Q) The pressure of a gas in a flask is measured to be 797.7 mmHg. What is this pressure in pascals and atmospheres?

Solution Conversion to pascals:

given quantity

$$797.7 \text{ mmHg} \times \frac{1.01325 \times 10^5 \text{ Pa}}{760 \text{ mmHg}} = \mathbf{1.064 \times 10^5\ Pa}$$

desire quantity

Conversion to atmospheres:

$$797.7 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}} = \mathbf{1.050\ atm}$$

➢ 5.2 Empirical Gas Laws

1.) Bowles law 2.) Charles law 3.) Combined law 4.) Avogadro's law

1. ➢ Boyle's Law: Relating Volume and Pressure

Reminder!

Boyle's law:
$PV = \text{constant}$ (for a * given amount of gas at * fixed temperature)

the volume of a sample of gas at a given temperature varies inversely with the applied pressure. That is, ***V* α 1/*P***, where *V* is the volume, *P* is the pressure,

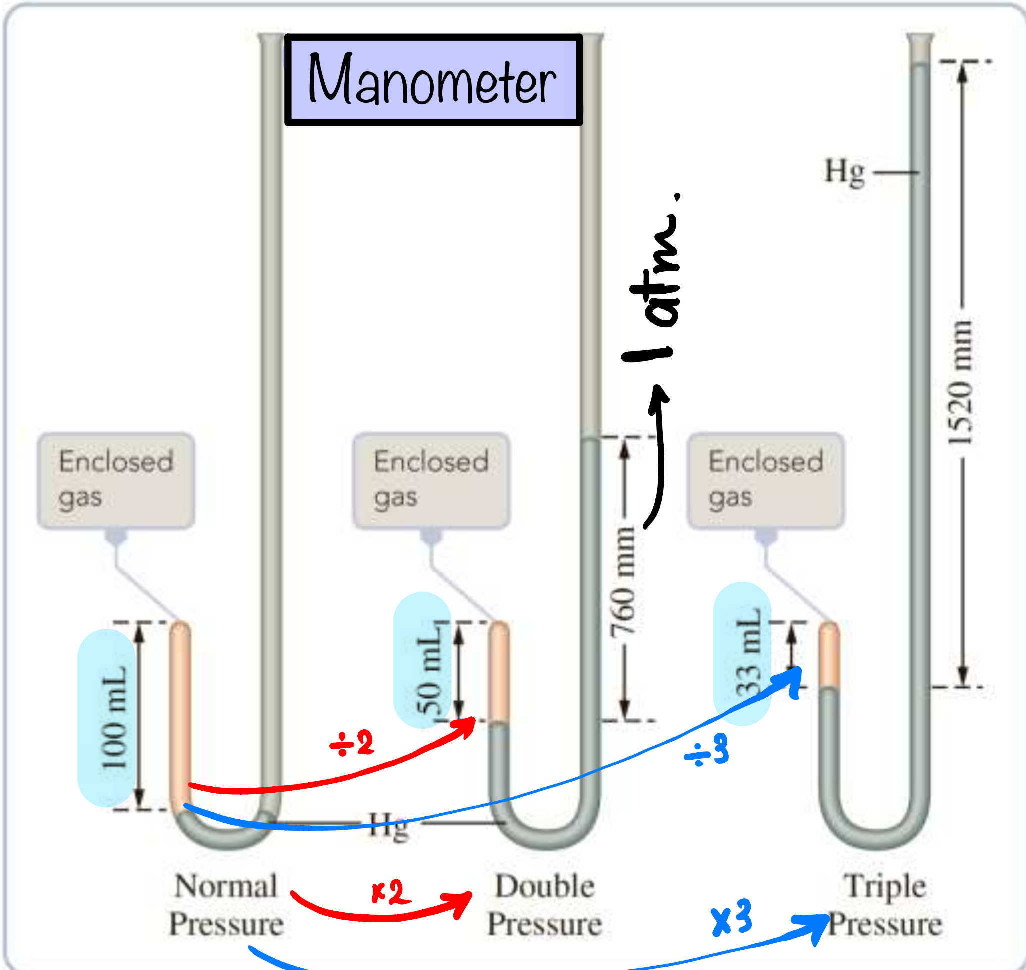


Table 5.3 Pressure–Volume Data for 1.000 g O_2 at 0°C

P (atm)	V (L)	PV
0.2500	2.801	0.7002
0.5000	1.400	0.7000
0.7500	0.9333	0.7000
1.000	0.6998	0.6998
2.000	0.3495	0.6990
3.000	0.2328	0.6984
4.000	0.1744	0.6976
5.000	0.1394	0.6970

Increasing pressure ↓ (P column); Increasing volume ↑ (V column). P: 0.2500 → 0.5000 ×2; V: 2.801 → 1.400 ÷2. PV: Constant

* $V \propto \frac{1}{P}$

→ $V = \text{Constant} \times \frac{1}{P}$

« $P \times V = \text{Constant}$ »

$P_1 V_1 = P_2 V_2 = P_3 V_3 = \text{Constant}$

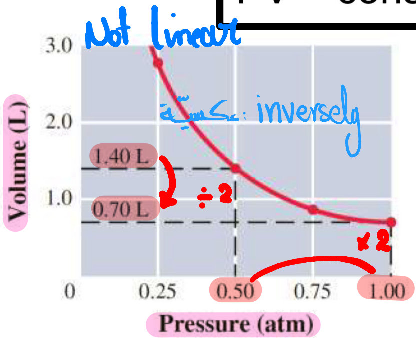
« $P_i V_i = P_f V_f$ »

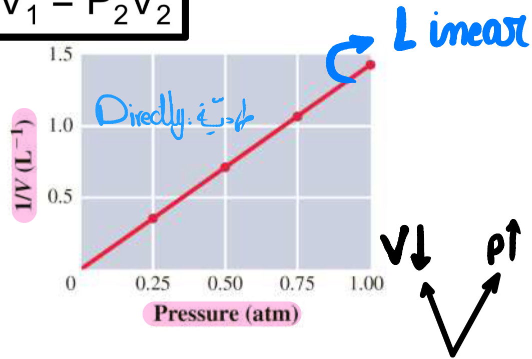
➢ Boyle's experiment:

The volume of the gas at normal atmospheric pressure (760 mmHg) is 100 mL. When the pressure is doubled by adding 760 mm of mercury, the volume is halved (to 50 mL).

Tripling the pressure decreases the volume to one-third of the original (to 33 mL).

$$PV = \text{constant} \rightarrow P_1V_1 = P_2V_2$$





(Q) A volume of air occupying 12.0 dm^3 at 98.9 kPa is compressed to a pressure of 119.0 kPa. The temperature remains constant. What is the new volume?

Example 5.2

$$P_iV_i = P_fV_f$$

Reminder!

$$V_f = V_i \times \frac{P_i}{P_f} = 12.0\ dm^3 \times \frac{98.9\ \cancel{kPa}}{119.0\ \cancel{kPa}} = \mathbf{9.97\ dm^3}$$

dm^3 ⟷ 1L

Exercise 5.1 A gas in a container had a measured pressure of 57 kPa. Calculate the pressure in units of atm and mmHg.

See Problems 5.37 and 5.38.

* 57 Kpa → atm → mmHg

* 1 atm = 1.0325 × 10^5 p ≃ 101 k pa

$$57\ Kpa \times \frac{1\ atm}{101\ k\ pa} = [.564\ atm]$$

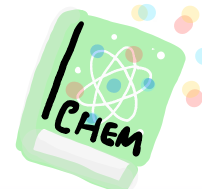
* 760 mmHg = 1 atm

$$.564\ atm \times \frac{760\ mm\ Hg}{1\ atm} = 428.64\ mmHg \rightarrow [4.3 \times 10^3\ mm\ Hg]$$

Exercise 5.2 A volume of carbon dioxide gas, CO_2, equal to 20.0 L was collected at 23°C and 1.00 atm pressure. What would be the volume of carbon dioxide collected at 23°C and 0.830 atm?

See Problems 5.39, 5.40, 5.41, and 5.42.

Non-toxic



$$\langle\langle p_i \times V_i = p_f \times V_f \rangle\rangle$$

$$20.0L \times 1.0\ atm = V_f \times .830\ atm$$

$$V_f = 24.09\ L = \boxed{24.1L}$$

} 23C° _ Co2
At given Amount of gas _ and fixed Temprature ...

Charles's Law: Relating Volume and Temperature

It must be in Kelvin

2.

$$\frac{V}{T} = \text{constant}$$

1\. (for a given amount of gas at a 2. fixed pressure)

$$\frac{V_f}{T_f} = \frac{V_i}{T_i}$$

Reminder!

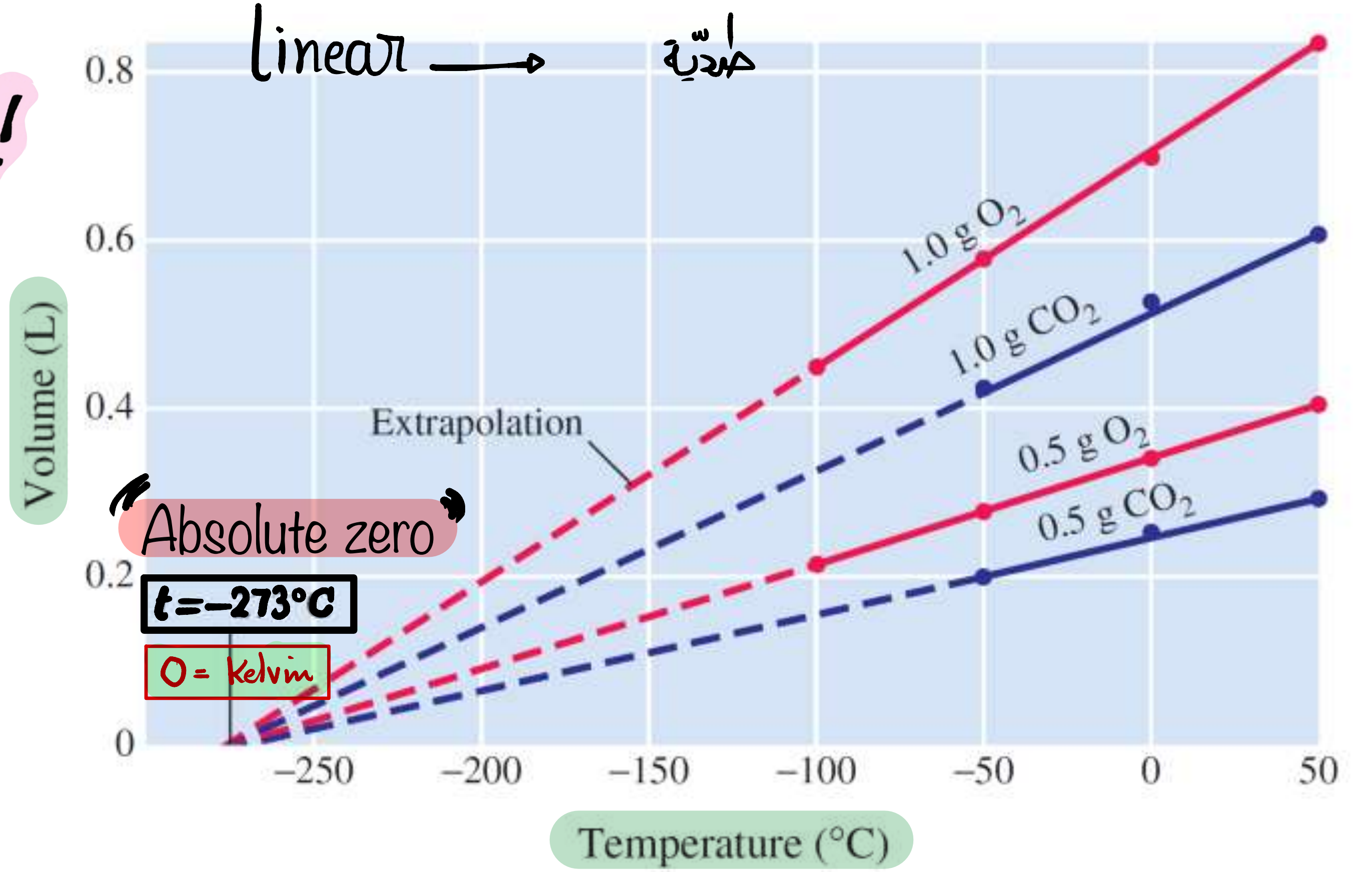
* $V \propto T$

$(V = \text{Constant } T)$

$\rightarrow \frac{V}{T} = \text{Constant}$

$\frac{V_1}{T_1} = \frac{V_2}{T_2} = \frac{V_3}{T_3} = \text{Constant}$

$\left(\frac{V_i}{T_i} = \frac{V_f}{T_f}\right)$



Exercise 5.3 If you expect a chemical reaction to produce 4.38 dm³ of oxygen, O_2, at 19°C and 101 kPa, what will be the volume at 25°C and 101 kPa?

First, convert the temperatures to the Kelvin.

$$T_i = (19 + 273) = 292\ \text{K}$$
$$T_f = (25 + 273) = 298\ \text{K}$$

Apply Charles's law

$$V_f = V_i \times \frac{T_f}{T_i} = 4.38\ \text{dm}^3 \times \frac{298\ \text{K}}{292\ \text{K}} = 4.4\underline{7}0 = 4.47\ \text{dm}^3$$

Example 5.3

Earlier we found that the total volume of oxygen that can be obtained from a particular tank at 1.00 atm and 21°C is 785 L (including the volume remaining in the tank). What would be this volume of oxygen if the temperature had been 28°C?

T_i = 21C° + 273 = 294K

T_f = 28C° + 273 = 301K

$$\left(\frac{V_i}{T_i} = \frac{V_f}{T_f}\right) \rightarrow \frac{785\ \text{L}}{294\ \text{K}} = \frac{V_2}{301\ \text{K}}$$

→ 804 L

➢ Combined Gas Law: Relating Volume, Temperature, and Pressure

✓ Boyle's law ($V \alpha 1/P$) and Charles's law ($V \alpha T$) can be combined to:

$$V \alpha T/P$$

$$V = \text{constant} \times \frac{T}{P} \quad \text{or} \quad \frac{PV}{T} = \text{constant} \quad \text{(for a given amount of gas)}$$

$$\frac{P_f V_f}{T_f} = \frac{P_i V_i}{T_i}$$

nf → It must be in Kelvin ← ni

Reminder!

Given amount of gas

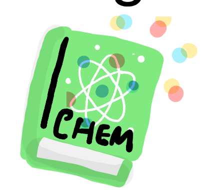
Example 5.4

(Q) A 39.8 mg sample of caffeine gives 10.1 cm³ of N_2 gas at 23°C and 746 mmHg. What is the volume of N_2 at 0°C and 760 mmHg?

$T_i = (23 + 273)\ \text{K} = 296\ \text{K}$

$T_f = (0 + 273)\ \text{K} = 273\ \text{K}$

$$V_f = V_i \times \frac{P_i}{P_f} \times \frac{T_f}{T_i} = 10.1\ \text{cm}^3 \times \frac{746\ \text{mmHg}}{760\ \text{mmHg}} \times \frac{273\ \text{K}}{296\ \text{K}} = \mathbf{9.14\ cm^3}$$



(Q) What will be the final pressure of a sample of nitrogen gas with a volume of 950. m^3 at 745 torr and 25.0 °C if it is heated to 60.0 °C and given a final volume of 1150 m^3?

$$\frac{P_i \times V_i}{T_i} = \frac{P_f \times V_f}{T_f} \rightarrow \left(\frac{745 \times 950\ m^3}{298\ K} = \frac{P_f \times 1150\ m^3}{333\ K} \right)$$

* $T_i = 25.0\ C^\circ + 273 = 298\ K$
* $T_f = 60.0\ C^\circ + 273 = 333\ K$

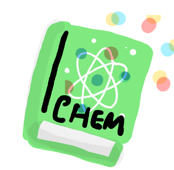
* $P_f = 687.68$ torr *

Exercise 5.4

A balloon contains 5.41 dm^3 of helium, He, at 24°C and 101.5 kPa. Suppose the gas in the balloon is heated to 35°C. If the helium pressure is now 102.8 kPa, what is the volume of the gas?

$T_i = (24 + 273) = 297$ K

$T_f = (35 + 273) = 308$ K



$$V_f = V_i \times \frac{P_i}{P_f} \times \frac{T_f}{T_i} = 5.41\ dm^3 \times \frac{101.5\ kPa}{102.8\ kPa} \times \frac{308\ K}{297\ K} = 5.5\underline{3}9 = \boxed{5.54\ dm^3}$$

1. $V \propto T$ Charles law

$V = \text{Constant}\ T$

$\frac{V}{T} = \text{Constant}$

$\frac{V_1}{T_1} = \frac{V_2}{T_2} = \frac{V_3}{T_3} = \text{Constant}$

" $\frac{V_i}{T_i} = \frac{V_f}{T_f}$ "

2. $V \propto \frac{1}{P}$ Boyles law

$V = \text{Constant} \times \frac{1}{P}$

$P \times V = \text{Constant}$

$P_1 V_1 = P_2 V_2 = P_3 V_3 = \text{Constant}$

" $P_i V_i = P_f V_f$ "

* $V \propto \frac{1}{P}$
* $V \propto T$
* $V \propto \frac{T}{P}$

$V = \text{Constant}\ \frac{T}{P}$

$\frac{PV}{T} = \text{Constant}$

$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} = \frac{P_3 V_3}{T_3}$

$$\frac{P_i V_i}{T_i} = \frac{P_f V_f}{T_f}$$

3. Combined Law

3. Avogadro's Law: Relating Volume and Amount

- French chemist Joseph Louis Gay-Lussac concluded from experiments on gas reactions that: *

* the volumes of reactant gases at the same pressure and temperature are in ratios of small whole numbers (*the law of combining volumes*).

Question (53-54).

$$2H_2(g) + 1O_2(g) \longrightarrow 2H_2O(g)$$

2 volumes   1 volume

when the question say at STP:
1. T = 0C°
2. P = 1 atm
3. 1 mol of gas → 22.4 L

- **Avogadro's law**:

equal volumes of any two gases at the same temperature and pressure contain the same number of molecules. / Moles

- *volume of **one mole** of gas* is called the molar gas volume, V_m.

$V_1 = V_2$
$T_1 = T_2$
$P_1 = P_2$
⟹ $n_1 = n_2$ for sure

Avogadro's law: → 1 mol → 22.4L STP ←
V_m = specific constant (=22.4 L/mol at STP)
(depending on T and P but independent of the gas)

Reminder!

STP = Standard Temperature and Pressure (*0°C and 1 atm*)

* $V \propto \frac{1}{P}$
* $V \propto T$
* $V \propto n$

⇓

$V \propto \frac{nT}{P}$ → $V = \text{Constant} \frac{nT}{P}$

$\frac{PV}{nT} = \text{Constant} = R$

P (atm), V (L), n (mol), T (K)

* $R = 0.0821 \text{ atm.L/mol.K}$ *

"$PV = nRT$"

"Ideal gas Law"

((Explanation))

➢ 5.3 The Ideal Gas Law

$$PV = nRT$$

Reminder!

Example 5.6

(Q) How many grams of oxygen are there in a 50.0-L gas cylinder at 21°C when the oxygen pressure is 15.7 atm?

$$PV = \frac{m}{mm} RT \Rightarrow 15.7 \times 50.0 = \frac{m}{31.998} \times 294 \times 0.0821$$

(atm, L, g/mol, K, R = 0.0821)

$(m = 1040.64\ g = 1.04 \times 10^3 g)$

Exercise 5.6

What is the pressure in a 50.0-L gas cylinder that contains 3.03 kg of oxygen, O_2, at 23°C?

$$PV = nRT \Rightarrow P \times 50.0 = \frac{3.03 \times 10^3}{31.998} \times 0.0821 \times 296 =$$

(L, g, R =, K)

$(P = 46.02\ atm)$

(Q) Calculate the volume (in L) occupied by 7.40 g of NH_3 at STP

STP: 0C° , 1atm

① $V = 7.40\ g\ NH_3 \times \frac{1\ mol\ NH_3}{17.03\ g\ NH_3} \times \frac{22.41\ L}{1\ mol\ NH_3}$

$= 9.74\ L$

② IGL: $PV = nRT$

Example 5.5 Deriving Empirical Gas Laws from the Ideal Gas Law

Gaining Mastery Toolbox

Critical Concept 5.5
The ideal gas law can be used to derive relationships between n, P, T, and V. Interpreting the equation $PV = nRT$ can lead you to the relationships between each of the variables. For example, for a sample of gas the pressure (P) is directly proportional to the number of moles (n) of gas and the temperature (T) of the gas.

Solution Essentials:
- Ideal gas law ($PV = nRT$)
- Units of pressure (Table 5.2)
- Kelvin temperature

Prove the following statement: the pressure of a given amount of gas at a fixed volume is proportional to the absolute temperature. This is sometimes called *Amontons's law*. In 1702, Guillaume Amontons constructed a thermometer based on measurement of the pressure of a fixed volume of air. The principle he employed is now used in special gas thermometers to establish the Kelvin scale.

Problem Strategy Because we need a relationship between pressure and temperature, the ideal gas law is a logical starting place.

Solution From the ideal gas law,

$$PV = nRT$$

Solving for P, you get

$$P = \left(\frac{nR}{V}\right)T$$

Note that everything in parentheses in this equation is constant. Therefore, you can write

$$P = \text{constant} \times T$$

Or, expressing this as a proportion,

$$\boldsymbol{P \propto T}$$

Boyle's law and Charles's law follow from the ideal gas law by a similar derivation.

(continued)

Exercise 5.5 Show that the moles of gas are proportional to the pressure for constant volume and temperature.

See Problems 5.55 and 5.56.

$PV^* = nRT^*$ → $P = n\underline{R}$ Constant

$P = n \times \text{Constant}$ → «$P \propto n$»

$$* \quad PV = nRT \longrightarrow \frac{P}{RT} = \frac{n}{V} \longrightarrow \frac{P}{RT} = M \longrightarrow$$

($\frac{n}{V}$ → Molarity)

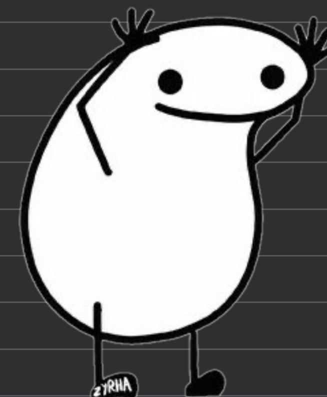
$$\longrightarrow \boxed{P = MRT}$$

$$* \quad PV = nRT \longrightarrow \frac{P}{V} = \frac{m}{mm} \times RT \longrightarrow$$

$$PM_m = \frac{m}{V} \times RT \longrightarrow PM_m = dRT$$

(d → Density)

$$\longrightarrow \boxed{PM_m = dRT}$$



Gas Density; Molecular-Weight Determination

$$PM_m = dRT$$

d is the density of the gas in g/L

Reminder!

Example 5.7

(Q) What is the density of oxygen, O_2, in grams per liter at 25°C and 0.850 atm?

$d = PM_m/RT = (0.85 \times 32)/(0.082 \times 298) = 1.11$ g/L

Exercise 5.8 A sample of a gaseous substance at 25°C and 0.862 atm. has a density of 2.26 g/L. What is the molecular weight of the substance?

$M_m = dRT/P = (2.26 \times 0.082 \times 298) / 0.862 = 64.1$ g/mol

$1\ mL = 1\ cm^3$

$1\ dm^3 = 1\ l$

$1\ m^3 = 1000\ L$

Exercise 5.7 Calculate the density of helium, He, in grams per liter at 21°C and 752 mmHg. The density of air under these conditions is 1.188 g/L. What is the difference in mass between 1 liter of air and 1 liter of helium? (This mass difference is equivalent to the buoyant, or lifting, force of helium per liter.)

See Problems 5.63, 5.64, 5.65, and 5.66.

$* \; PM_m = dRT \; *$

$.989 \times 4.0026 = d \times 0.0821 \times 294$

$\rightarrow d = 0.164 \text{ g/L}$

$T = 21 + 273 \rightarrow 294\text{K}$

$P = \frac{752}{760} = .989 \text{ atm}$

Mass O_2 − Mass He =

* Difference in Mass between 1L of air and 1L of He = 1.024 g

Example 5.8 Determining the Molecular Weight of a Vapor

Gaining Mastery Toolbox

Critical Concept 5.8
The molar mass of a substance is the grams of substance per mol of substance (g/mol). The ideal gas law can be used to determine the moles of a gas sample when the pressure (P), temperature (T), and volume (V) are known.

Using the apparatus described in Figure 5.17, a 200.0-mL flask at 99°C and a pressure of 733 mmHg is filled with the vapor of a volatile (easily vaporized) liquid. The mass of the substance in the flask is 0.970 g. What is the molecular weight of the liquid?

Problem Strategy You can calculate the moles of vapor from the ideal gas law and then calculate the molar mass by dividing mass by moles. Note that the temperature of the vapor is the same as the temperature of the boiling water and that the pressure of the vapor equals the barometric pressure.

(continued)

Volume: 200 mL → .2 L

$T = 99°C + 273 \rightarrow 372\text{K}$

$\frac{733 \text{ mmHg}}{760} \rightarrow .964 \text{ atm}$

mass = .970 g

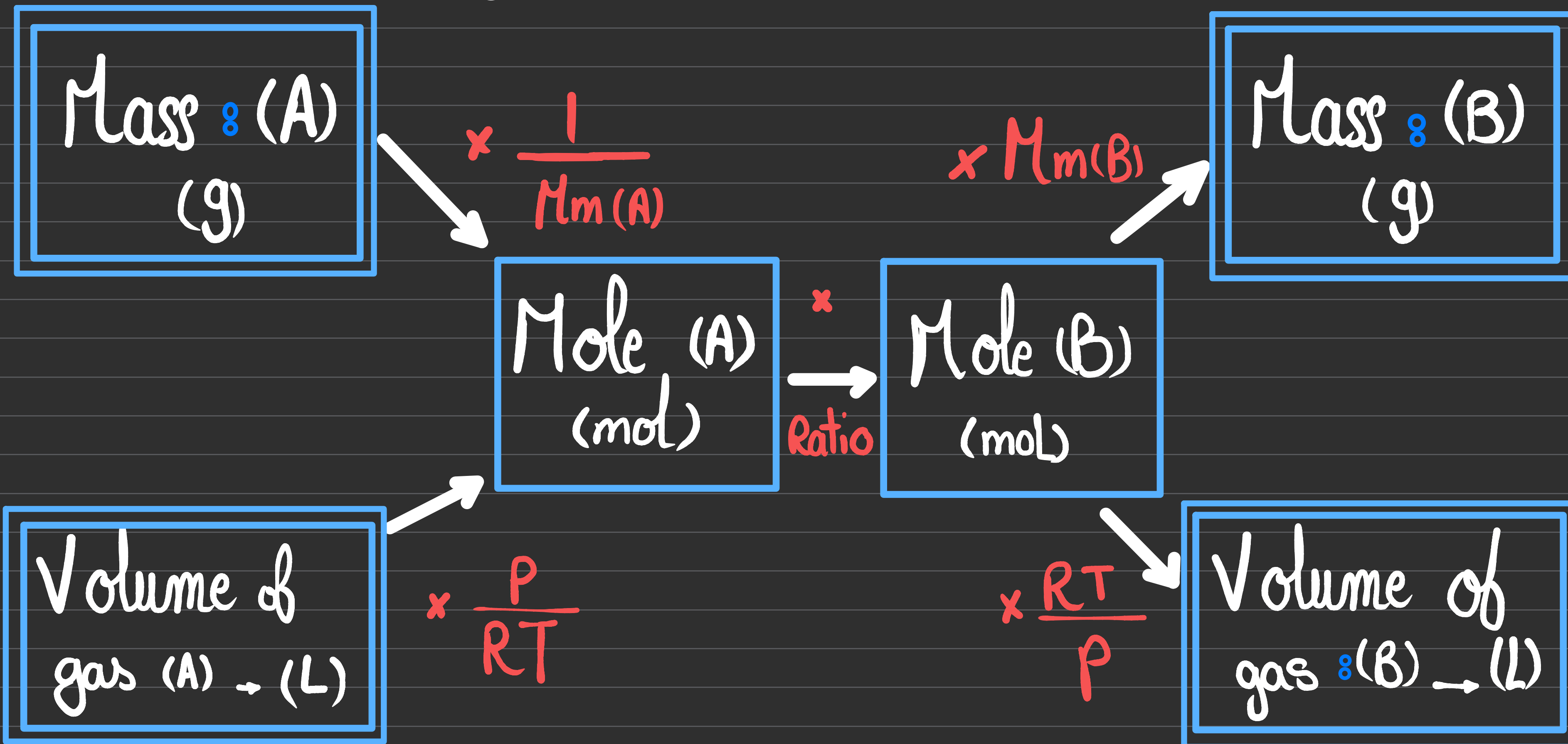
$PM_m = dRT$

$PM_m^? = \frac{m}{V} RT$

$M_m = 153.656$

→ 154 amu.

Stoichiometry Problems Involving Gas Volumes



$$PV = nRT$$

5.4 Stoichiometry Problems Involving Gas Volumes

$6NaN_3(s) + Fe_2O_3(s) \rightarrow 3Na_2O(s) + 2Fe(s) + 9N_2(g)$

Calculate the volume of N_2 generated at 80°C and 823 mmHg by the decomposition of 60.0 g of NaN_3

Exercise 5.9 How many liters of chlorine gas, Cl_2, can be obtained at 40°C and 787 mmHg from 9.41 g of hydrogen chloride, HCl, according to the following equation?

$2KMnO_4(s) + 16HCl(aq) \rightarrow 8H_2O(l) + 2KCl(aq) + 2MnCl_2(aq) + 5Cl_2(g)$

$$9.41\,g\,HCl \times \frac{1\,mol\,HCl}{36.46\,g} \times \frac{5\,Cl_2}{16\,HCl} \times \frac{0.0821 \times 313}{1.035} = (20.L)$$

R, T, P

n → Cl_2

Example 5.9:

$$6NaN_{3(s)} + Fe_2O_{3(s)} \longrightarrow 3Na_2O_{(s)} + 2Fe_{(s)} + 9N_{2(g)}$$

Volume N_2 = ?

T = 80C° → 353 K

P = 828 mmHg → 1.08 atm

NaN_3 = 60.0g → Mass

Mass NaN_3 → Mole NaN_3 → Mole N_2 $\xrightarrow{\frac{RT}{P}}$ Volume N_2

$$60.0\ g \times \frac{1\ mole\ NaN_3}{64.921\ g} \times \frac{9\ mol\ N_2}{6\ mol\ NaN_3} \times \frac{0.0821 \times 353.15\ K}{1.08\ atm} = 38.2\ L$$

(9 mol N_2: Desire; 6 mol NaN_3: given; 0.0821: R; 353.15 K: T; 1.08 atm: P)

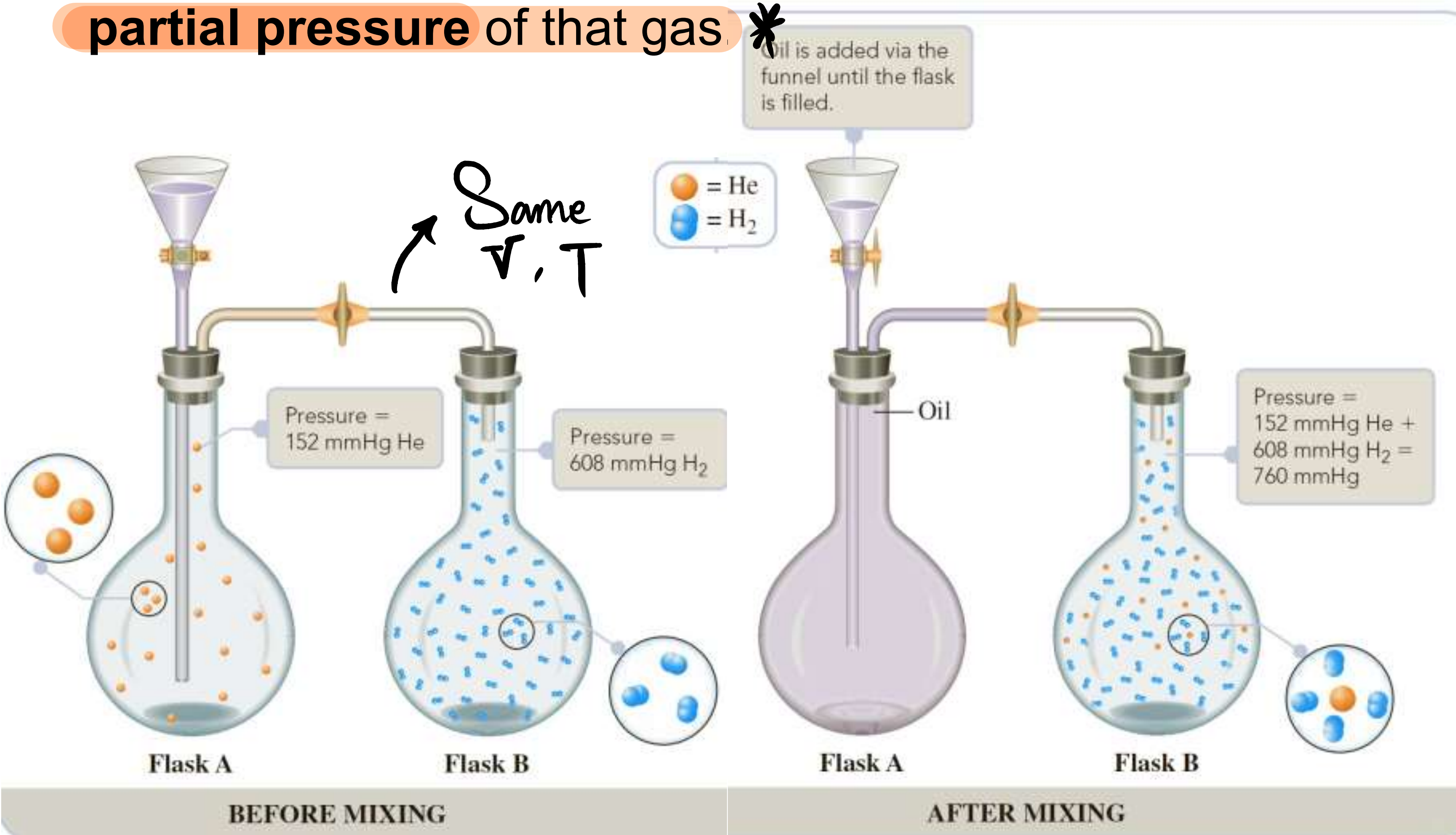
$n = N_{2(g)}$

5.5 Gas Mixtures; Law of Partial Pressures

➢ Partial Pressures and Mole Fractions

✓ **Dalton's law of partial pressures:**

* *The pressure exerted by a particular gas in a mixture* is the **partial pressure** of that gas *



Explanation

1. $P_T = P_A + P_B + P_C + \dots$ (T → Total) * Constant T, V

2. $\frac{n_T RT}{V} = \frac{n_A RT}{V} + \frac{n_B RT}{V} + \frac{n_C RT}{V}$

3. $\frac{n_T RT}{V} = \frac{RT}{V}(n_a + n_B + n_C + \dots)$

4. $n_T = (n_a + n_B + n_C + \dots)$ (T → total)

5. $\left[\frac{n_a}{n_T} = \frac{P_A}{P_T}\right] = X_A$ → mole Fraction.

Dalton's law of partial pressures:

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

✓ The individual (partial pressures) follow the ideal gas law. For component *A*,

$$P_A V = n_A RT$$

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

Reminder!

Example 5.10

(Q) A 1.00-L sample of dry air at 25°C and 786 mmHg contains 0.925 g N_2, plus other gases including oxygen, argon, and carbon dioxide.

a. What is the partial pressure (in mmHg) of N_2 in the air sample?

b. What is the mole fraction and mole percent of N_2 in the mixture?

$$0.925 \text{ g } N_2 \times \frac{1 \text{ mol } N_2}{28.0 \text{ g } N_2} = 0.0330 \text{ mol } N_2$$

$$P_{N_2} = \frac{n_{N_2}RT}{V} = \frac{0.0330 \text{ mol} \times 0.0821 \text{ L·atm/(K·mol)} \times 298 \text{ K}}{1.00 \text{ L}} = 0.807 \text{ atm } (= \mathbf{613 \text{ mmHg}})$$

Mixture

Mixture / Sample

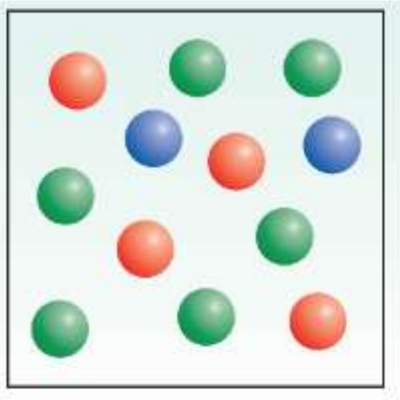
$$\text{Mole fraction of } N_2 = \frac{P_{N_2}}{P} = \frac{613 \text{ mmHg}}{786 \text{ mmHg}} = \mathbf{0.780}$$

mmHg

Air contains **78.0 mole percent** of N_2.

(Q) Each of the color spheres represents a different gas molecule. Calculate the partial pressures of the gases if the total pressure is 2.6 atm.

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



* $X_r = \frac{4}{12} = .333$ mol ($4 \to n_r$, $12 \to n_t$)

* $X_b = \frac{2}{12} = .166$ mol ($2 \to n_b$, $12 \to n_t$)

* $X_g = \frac{6}{12} = .5$ mol ($6 \to n_g$, $12 \to n_t$)

→ $X_r = \frac{n_A}{n_T} = \frac{P_r}{P_T}$

$.333 = \frac{P_r}{2.6} = .8658$ atm

→ $X_B = \frac{n_B}{n_T} = \frac{P_B}{P_T} =$

$.166 = \frac{P_B}{2.6} = .4316$ atm

→ $X_g = \frac{n_g}{n_T} = \frac{P_g}{P_T}$

$.5 = \frac{P_g}{2.6} = 1.3$ atm

(Q) A mixture consists of 122 moles of N_2, 137 moles of C_3H_8, and 212 moles of CO_2 at 200 K in a 75.0 L container. What is the total pressure of the gas and the partial pressure of CO_2?

A. 46.4 atm, 20.9 atm

B. 103 atm, 26.7 atm

C. 103 atm, 46.4 atm

D. 103 atm, 29.9 atm

E. 46.4 atm, 46.4 atm

1. $P_{total} = \frac{(471 \text{ moles})(0.0821 \text{ L atm mol}^{-1} \text{ K}^{-1})(200 \text{ K})}{75.0 \text{ L}}$ (n_T, R, T; V = SAMPLE)

$P_{total} = 103 \text{ atm}$

2. mole fraction CO_2: $\frac{212 \text{ moles } CO_2}{122 + 137 + 212 \text{ total}} = 0.450$ (n_{CO_2} / n_T)

$$P_{CO_2} = (\chi_{CO_2})(P_{total}) = (0.450)(103 \text{ atm})$$

$P_{CO_2} = 46.4 \text{ atm}$

Partial presure Co_2

$P = \frac{nRT}{V} = 46.4$

$$\text{Mole fraction of } A = \left(\frac{n_A}{n} = \frac{P_A}{P}\right)$$

(Q) A mixture of 250 mL of methane, CH_4, at 35° C and 0.55 atm and 750 mL of propane, C_3H_8, at 35° C and 1.5 atm was introduced into a 10.0 L container. What is the mole fraction of methane in the mixture?

A. 0.50

B. 0.11

C. 0.89

D. 0.25

E. 0.33

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

* الـ Partial pressure
الـ gas ← يجب أن
يكون مأخوذ من الـ
Volume » الكلي
V Total

1. $P_1V_1 = P_2V_2$ (P)

$$P_{CH_4} = \frac{0.55 \text{ atm} \times 0.250 \text{ L}}{10.0 \text{ L}} = 0.0138 \text{ atm}$$

$$P_{C_3H_8} = \frac{1.5 \text{ atm} \times 0.750 \text{ L}}{10.0 \text{ L}} = 0.112 \text{ atm}$$

$$\chi_{CH_4} = \frac{0.0138 \text{ atm}}{0.0138 \text{ atm} + 0.112 \text{ atm}} = 0.110$$

(P_{CH_4} ← 0.0138 atm; $P_{P\,Total}$ ← 0.0138 atm + 0.112 atm)

Boyle'S law

* ممكن
يلعب بـ Temprature
وماتكون ثابتة ←
يستخدم الـ Compined law

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

2. $PV = nRT$ (n)

Volume CH_4; n → CH_4

$n = \frac{PV}{RT}$ ← $PV = nRT = 5.4 \times 10^{-3}$ mol

Volume C_3H_8; n → C_3H_8

$n = \frac{PV}{RT}$ ← $PV = nRT = 0.044$ mol

> .108

$$\frac{n_{CH_4}}{n_{Total}}$$

Exercise 5.10 A 10.0-L flask contains 1.031 g O_2 and 0.572 g CO_2 at 18°C. What are the partial pressures of oxygen and carbon dioxide? What is the total pressure? What is the mole fraction of oxygen in the mixture?

See Problems 5.81, 5.82, 5.83, and 5.84.

Volume = 10.0L

T = 291 K°

O_2 = 1.031 g — 0.572 g = CO_2 — T = 18C°

1) partial presure O_2 = pO_2 ⟹ $\frac{nRT}{V}$

$n = \frac{1.031}{32.\ g/mol} = 0.0322$ mol

$n = \frac{m}{mm}$

nO_2 R

$pO_2 = \frac{0.0321 \times 00821 \times 291}{10.0}$ V_T T_t

→ 0.0769 atm

2) partial presure CO_2 = pCO_2 ⟹ $\frac{nRT}{V}$

$n = \frac{0.572\ g}{43.999} = .013$ mol nCO_2

$\frac{m}{mm}$ R

$pCO_2 = \frac{0.013 \times 0.0821 \times 291}{10}$ V_t T_t

0.031 atm

→

3.) Total presure = $pCO_2 + PO_2$

$0.0769 + 0.031 =$

$P_T \rightarrow ,1079$ atm

4.) The mole fraction O_2 ??

$$X_{O_2} = \frac{n_{O_2}}{n_T} = \frac{P_{O_2}}{P_T}$$

$$* \frac{0,0769}{,1079} = ,712$$

$$* \frac{0.0321}{0.0451} = .712$$

Done by: Joud Jaber