لجنة طب الأسنان

Dental
Committee

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#فريق_لجنة_طب_الاسنان_الاكاديمي

10.2 Pressure of a GAS

Because gas molecules are in constant motion, gas particles exert Pressure on any surface they encounter.

Atmospheric Pressure:

- Pressure exerted by column of air on area exposed to Earth's atmosphere.
- depends on location, temperature, and weather conditions.
- decreases as altitude increases because air becomes thinner.

About 760 mmHg at sea level barometer

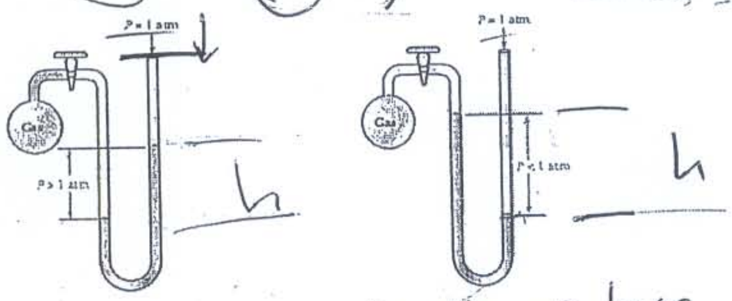
Instruments

- Barometer: $P_{gas} = P_{atm.}$
- Manometer
  - Closed-tube: $P_{gas} = h$
  - Open-tube:
    - $P_{gas} > P_{atm}$: $P_{gas} = P_{atm} + h$
    - $P_{gas} < P_{atm}$: $P_{gas} = P_{atm} - h$

$P_{atm} = 1$ atm

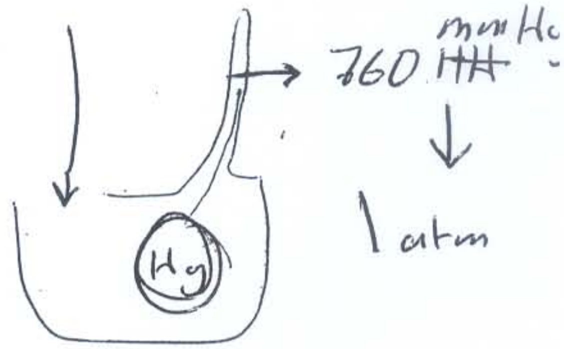
Note: 76 cm = 760 mmHg → torr

→ 1.00 atm = 760 mmHg



mmHg → torr

760 torr



1 atm → 760 mmHg

P_{atm}

Units of Pressure

Standard Atmospheric Pressure (1 atm): 760 mmHg at 0°C at sea level

1 atm ≡ 760 mmHg ≡ 760 torr = 101.325 kPa

Ex. 1: If the atmospheric pressure is measured to be 745 mmHg on a given day in Phoenix, express this atmospheric pressure in torr, atm, and kPa.

Answer:

Pressure in torr = pressure in mmHg = 745 torr

Pressure in atm = (? mmHg/760) = 0.98 atm

Pressure in kPa = (? mmHg x 101.325 kPa)/760 mmHg = 99.325 kPa

10.3 GAS LAWS

$$PV = nRT$$

P: pressure (atm)

V: volume (L)

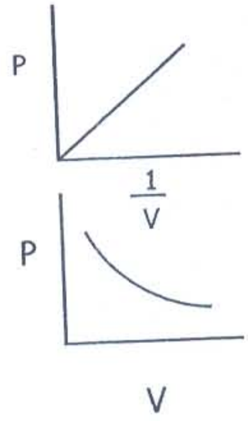
n: moles: (mol)

R: gas constant (0.0821 L.atm/mol.K)

T: temperature (K)

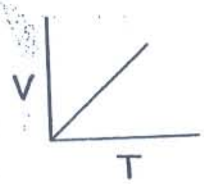
1- **Boyle's Law:** (P-V Relationship) [Constant T & n]

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



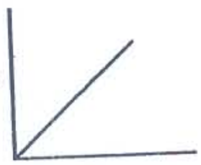
2- **Charles' Law:** (T-V Relationship)[Constant P & n]

$$T \propto V$$



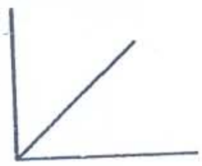
– Absolute Temperature ≡ zero Kelvin Temp., = -273.15 °C

3- **Gay Lussac's Law:** (P-T Relationship) [Constant V & n]



$$T \propto P$$

4- **Avogadro's Law:** (V - n Relationship) [Constant T & P]



$$n \propto V$$

For a given sample of gas under two different conditions we have:

Boyl $P_1V_1 = P_2V_2$

Charles $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

Gay-Lussac $\frac{P_1}{T_1} = \frac{P_2}{T_2}$

Avogadro $\frac{V_1}{n_1} = \frac{V_2}{n_2}$

Q.1: If 20 L of hydrogen gas are heated from 298.15K to 552.50°C, calculate the new volume?

$V_1 = 20$ L
$T_1 = 298.15$ K

$V_2 = ?$
$T_2 = 552.5 + 273 = 825.5$ K

$\frac{V_1}{T_1} = \frac{V_2}{T_2}$ >>> $V_2 = (V_1 x T_2)/T_1$
= (20 x 825.5)/298.15
= 55.37 L

Q.2: A 250 mL sample of helium at 722 mmHg is compressed until the new pressure is 3.60 atm. Calculate the new volume?

$P_1V_1 = P_2V_2$

$0.95 \times 250 = 3.60 \times V_2$

$V_2 = 0.95 \times 250 / 3.60$

$V_2 = 0.066$ L

STP: Standard Temperature & Pressure
0°C & 1 atm

Also, at STP 1mol = 22.4 L

$PV = nRT$
$1 \times V = 1 \times R \times 273$
$V = 22.4$

Q.:A sample of krypton gas at -10 °C. What is the pressure at STP?

Ans;

$T_1 = -10 + 273 = 263$ K
$P_1 = ?$

STP = 273 K, 1 atm
>>> $T_2 = 273$ K $P_2 = 1$ atm

$\frac{P_1}{T_1} = \frac{P_2}{T_2}$ >>> $P_1 = (T_1 x P_2)/T_2$
$= (263 \times 1)/273$
$= 0.96$ atm

$PV = nRT$
$P \times 22.4 = 1 \times 0.0821 \times 263.15$
$P = 0.96$ atm
#

10.4: The Ideal Gas Equation: معادلة الغاز المثالية

'' No attraction or repulsions between molecules ''
جاذبية

$$PV = nRT$$

P: pressure (atm)
V: volume (L)
n: moles: (mol)
R: gas constant (0.0821 L.atm/mol.K)
T: temperature (K)

$PV = nRT$
$PV = (m/MM)RT$
>> $MM = mRT/PV$

m : mass (g)
MM : molar mass (g/mol)

$d = m/V$
>>>> $MM = dRT/P$

d: density (g/ml)

n = m / MM
m = n × MM
d = m / V
d = n × MM / V
d = n × MM / V

Q.1: Calculate the volume for 1 mole of gas at STP?

PV =nRT
V = 1×0.0821×273 /1
V = 22.4 L

Q.2: How many moles of NO_2 gas occupy a volume of 5.00 L at 50.00°C and 735 torr?

PV =nRT
n = PV / RT
n = 0.97×5 /0.0821×323
n =0.18 mol

Q.3: An unknown gas having a mass of 6.15 g occupies a volume of 5 L at 874 torr and 23.50°C. Calculate the molar mass of the unknown gas?

MM = mRT/PV
= 6.15× 0.0821 × 296.5 /1.15 × 5
= 26 g/mol

Q.4: Which one of the following gases is lowest dense?

الرجوع إلى الجدول الدوري
العدد الكتلي (الذري)

16   12 + 1×4   34   79 + 1

a) NH_3 b) CH_4 c) Ar   d) HBr

Ans.: b) CH_4 (Since it has the lowest molar mass)

$PV = nRT \Rightarrow PV = \frac{m}{mm}RT = mm = \frac{dRT}{P}$

$mm \propto d$

Q.5: Calculate the density of NH_3 gas (17 g/mol) in 4.32 L container at 837 torr and 45 °C?

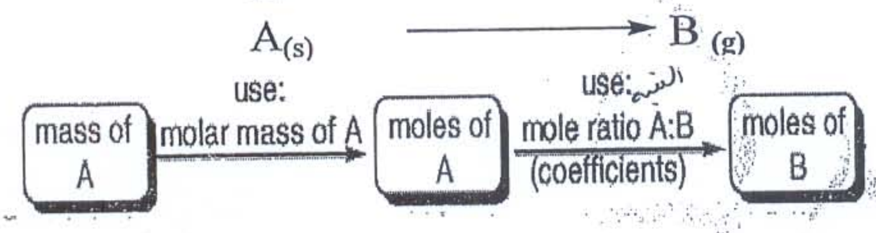
MM = dRT/ P

D = MM P/ RT

D= 17 × 1.1/0.0821 × 318

D = 0.719 g/L

10.5 Gas Stoichiometry



>>>> PV = nRT

Note: If the question need to find the limiting reactant you should find it

Ex.1: If 240 g of CH_4 are reacted with excess O_2, what is the volume of CO_2 (in liters) produced at 23°C and 0.985 atm?

$$CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(l)}$$

nCH_4 = 240 g / 16 g/mol = 15 mol

$$n_{CO_2} = \frac{1}{1} x \; n_{CH_4}$$

= 15 mol

Now, substitute n in the ideal gas law:

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

$$V_{CO_2} = \frac{(15.0\ \text{mol})\left(0.0821\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(23+273)\,\text{K}}{0.985\ \text{atm}} = 3.70 \times 10^2\ \text{L}$$

Ex.2: If 9.0 L of NO are reacted with excess O_2 at STP, what is the volume in liters of the NO_2 produced?

$$2NO + O_2 \rightarrow 2NO_2$$

Ans.:

The coefficients from a balanced equation can represent the volume ratio in which the gases in the equation react and are produced. Recall that Avogadro's Law states that $V \alpha\ n$.

$$9.0\ \text{L NO} \times \frac{2\ \text{volumes } NO_2}{2\ \text{volumes NO}} = 9.0\ \text{L } NO_2$$

Q. For the decomposition of $2NaN_3(s) \rightarrow 2Na(s) + 3N_2(g)$. How many grams of NaN_3 are required to provide 40 L of N_2 at 25 °C and 763 mmHg?

PV = nRT

n = PV / RT

n = 1× 40 / 0.0821 × 298

n = 1.63 mole

Mass = n × M.M

= 1.63 × 65

= 106 grams

$2NaN_3 \rightarrow 2Na + 3N_2$

$n = \frac{m}{Mm}$, $1 = \frac{m}{64}$ = 64 g

PV = nRT

1 × 40 = n × 0.0821 × 298.15

n = 1.6

$\frac{1.6}{3} = 0.5$

0.5 × 2 = 1

1 = n

10.6 Dalton's Law of Partial Pressures

partial pressure: pressures of individual gas components in a mixture

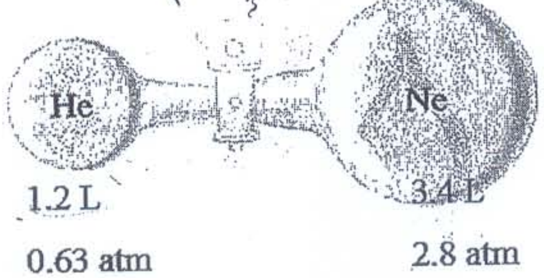
Dalton's Law of Partial Pressure:

- total pressure of a mixture of gases is the sum of the partial pressures of the gases present

$$P_{Total} = P_1 + P_2 + P_3 + \ldots$$

Also, $$P_{total} = (n_1 + n_2 + \ldots + n_n)\frac{RT}{V_{total}}$$

Q.: Consider the following apparatus. Calculate the partial pressures of helium and neon after the stopcock is open. The temperature remains constant at 16 °C?



Ans.:

$$n_{He} = \frac{PV}{RT} = \frac{(0.63\text{ atm})(1.2\text{ L})}{\left(0.0821\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(16+273)\text{K}} = 0.032\text{ mol He}$$

$$n_{Ne} = \frac{PV}{RT} = \frac{(2.8\text{ atm})(3.4\text{ L})}{\left(0.0821\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(16+273)\text{K}} = 0.40\text{ mol Ne}$$

The total pressure is:

$$P_{Total} = \frac{(n_{He} + n_{Ne})RT}{V_{Total}} = \frac{(0.032 + 0.40)\text{mol}\left(0.0821\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(16+273)\text{K}}{(1.2+3.4)\text{L}} = 2.2\text{ atm}$$

$P_i = X_i P_T$. The partial pressures of He and Ne are:

$$P_{He} = \frac{0.032 \text{ mol}}{(0.032 + 0.40)\text{mol}} \times 2.2 \text{ atm} = \mathbf{0.16 \text{ atm}}$$

$$P_{Ne} = \frac{0.40 \text{ mol}}{(0.032 + 0.40)\text{mol}} \times 2.2 \text{ atm} = \mathbf{2.0 \text{ atm}}$$

Q. What is the pressure in a 12.2 L vessel that contains 2.34 g of CO_2, 1.73 g SO_2 and 3.33 g of Ar at 42 °C?

$n_{CO_2} = 2.34 / 44 = 0.05$ $n_{SO_2} = 1.73 / 64 = 0.03$ $n_{Ar} = 3.33 / 40 = 0.08$

$$P_{total} = (n_1 + n_2 + \ldots + n_n)\frac{RT}{V_{total}}$$

$= (0.05 + 0.03 + 0.08) \times 0.0821 \times 315 / 12.2$

$= 0.34$

mole fraction (X_A): ratio of the number of moles of one component to sum total of all the moles of all components

$$X_A = \frac{\text{\# of moles of gas A}}{\text{Total \# of moles of all gases in mixture}}$$

Mole fraction to calculate the partial pressure of a gas in a mixture

- For system with more than many gases, the partial pressure of the n^{th} gas:

$$P_n = X_n P_{Total}$$

Ex. 1 A mixture of gases contains 4.465 mol of neon, 0.741 mol of argon, and 2.154 mol of xenon. Calculate the partial pressures of all the gases if the total pressure is 2.00 atm at a given temperature.

$P_{neon} = 1.2$ atm $P_{Argon} = 0.2$ atm $P_{xenon} = 0.60$ atm

Q1. In a gas mixture of He, Ne, and Ar of total pressure 8.40 atm, what is the Mole fraction of Ar if the respective partial pressures of He and Ne are 1.5 and 2.0 atm?

He + Ne + Ar
1,5 2
P_t = 8,4 atm
X_{Ar} = ??

8,4 = 1,5 + 2 + P_{Ar}
P_{Ar} = 4,4

$P_i = X_{Ar} \times P_t$

$\frac{4,4}{8,4} = \frac{X_A \times 8,4}{8,4}$

0,6 = X_{Ar}

$P_{total} = p_1 + p_2 + p_3$

$8.40 = 1.5 + 2 + p_{Ar}$

$p_{Ar} = 8.40 - (1.5 + 2)$

$= 4.9$ atm

$P_{Ar} = X_{Ar} \cdot p_{total}$

$4.9 = X_{Ar} \cdot 8.40$

$X_{Ar} = 0.583$

Q2. A gas mixture of total pressure 4.0 atm and 16.0 total moles contains gases X and Z. If partial pressure of Z is 2.75 atm, how many moles of X are in the mixture?

P = 4 atm
n_{tot} = 16
X + Z
?? n
2,75 atm

$P_i = \frac{n}{n_{tot}} \times P_t$

$P_{tot} = P_{i_X} + P_{i_Z}$

4 = P_{iX} + 2,75

P_{iX} = 1,25

$P_i = \frac{n}{n_{tot}} \times P_t$

1,25 = $\frac{n}{16} \times 4$

n = 5 mole.

$P_z = X_z \cdot P_{total}$

$2.75 = X_z \times 4$

$X_z = 0.686$

$X_z = n_z / n_{total}$

$0.686 = n_z / 16$

$n_z = 11$ mole

$n_{total} = n_x + n_z$

$16 = n_x + 11$

$n_x = 5$ mole

Gases is collected over water

$$P_{total.} = P_{H2O} + P_{gas}$$

Q. A mixture of helium (He) & neon(Ne) gases is collected over water at 28 °C & 745 mmHg. If the partial pressure of He is 368 mmHg, what is the partial pressure of Ne? Vapor pressure of H_2O at 28 °C = 28.3 mmHg

$$P_{total} = P_{Ne} + P_{He} + P_{H2O}$$

$$P_{Ne} = P_{total} - P_{He} - P_{H2O}$$

$$= 745\text{mmHg} - 368\text{mmHg} - 28.3\text{ mmHg}$$
$$= 349\text{ mmHg}$$

Molecular Speed (Root Mean Square Speed) [Urms]

Combining PV=nRT and KMT gives the ***root-mean-square speed, u_{rms}:***

$$U_{rms} = \sqrt{\frac{3RT}{Mwt.}}$$

U_{rms} : m/s R: 8.314 J/mol.k T: temperature (K) Mwt.: molar mass (Kg/mol)

Ex. Find Urms for $CO_{2(g)}$ (44g/mol) at 27°C?

Ans.: R=8.314 J/mol.k T= 300 K MM = 44×10^{-3} Kg/mol

>>>> Urms = 412 m/s

Q. Calculate the root-mean-square speed of a Rb (85.5 g/mol) at 1.7×10^{-7} K?

$$U_{rms} = \sqrt{\frac{3RT}{Mwt.}}$$

$U_{rms} = \sqrt{3 \times 8.314 \times 1.7\times10^{-7} / 85.5\times 10^{-3}}$

Ans.: 7.0×10^{-3} m/s

Q. at 200 K, the molecules or atoms of an unknown gas X, has an Urms of Ar atoms at 400 K. What is X? (Assume ideal behavior)

a) He   b) CO   c) HF   d) HBr

10.8 Gas Diffusion and Effusion

Diffusion: gradual mixing of molecules of one gas with molecules of another by virtue of their kinetic properties

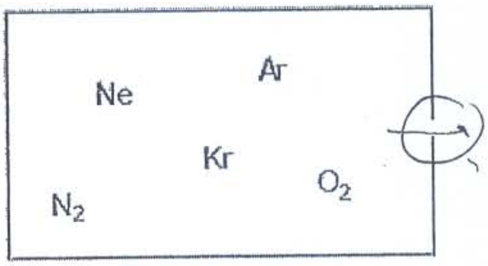
Which one of the following gases will diffuse the fastest at a given temperature?

a) NH_3 b) CH_4 c) Ar   d) HBr

Ans.: b) CH_4 (Since it has the lowest molar mass)

Effusion: process of a gas under pressure escaping from a container via a small opening.

Example: Consider a container sealed with a small opening. A mixture of several gases is added to the container. Assuming the gases do not react, indicate the order that the gases escape out of the container, starting with the gas that escapes the fastest. Explain why.



Ex 1. A sealed chamber containing ozone (O_3) gas also contains several other gases. Circle all the gases below that effuse faster than ozone at room temperature when a small opening is created in the chamber.

N_2O C_3H_8 Cl_2 CO_2 Ar

Graham's Law of Effusion

– **rate of effusion of molecules** $\propto \frac{1}{\sqrt{MM_{gas}}}$ (at same temperature and pressure)

For 2 gases: $$\frac{\text{rate of effusion for x}}{\text{rate of effusion for y}} = \frac{\sqrt{MM_y}}{\sqrt{MM_x}} = \frac{\text{time for y to effuse}}{\text{time for x to effuse}}$$

$$\frac{r_1}{r_2} = \frac{\sqrt{M.wt_{(2)}}}{\sqrt{M.wt_{(1)}}} = \frac{t_2}{t_1}$$

Ex.1: The rate of effusion for helium gas (4 g/mol) will how much faster Compared to the rate of effusion for carbon dioxide CO_2 (44 g/mol)?

Ans.:

$$\frac{r_1}{r_2}=\frac{\sqrt{Mwt._2}}{\sqrt{Mwt._1}}=\frac{\sqrt{44}}{\sqrt{4}}$$

$$\frac{r_1}{r_2}=3.3 \quad >>>> r_1=3.3\ r_2$$

$$\frac{r_2}{r_1}=\frac{\sqrt{M_{He}}}{\sqrt{M_{CO_2}}}=\frac{4}{44}=\frac{1}{11}$$

$r_1 = 11 r_2$

Questions

Q1. A sample of oxygen gas was found to effuse at a rate equal to three times that of unknown gas. What is the molecular weight of the unknown gas?

a) 288 b) 96 c) 55 d) 4

Q2. He gas effused through a hole in 53 seconds. An equivalent amount of an unknown gas, under the same conditions, effused through the hole in 248 seconds. What is the molecular weight of the unknown gas?

a) 0.18 g/mol b) 62 g/mol c) 88 g/mol d) 18.3g/mol

Q3. Consider the following reaction: $C_{12}H_{22}O_{11} + 12O_2 \longrightarrow 12CO_2 + 11H_2O$ If 11.7 g of $C_{12}H_{22}O_{11}$ is reacted with excess O2 , what volume of CO2 that Is formed at 1.0 atm and 37C ? (M.m $C_{12}H_{22}O_{11}$= 342.3 g/mole)?

a)22.4 L b)10.4 L c) 1.78 L d) 5.20 L

Q4. Consider the following reaction: $A_s + 2B_{(g)} \longrightarrow 4C_{(g)} + 2F_{(g)}$. If 20 L of B is reacted with an excess of A at STP , what is the total volume of gases at the end of reaction at STP?

a)20.0 L b)80.0 L c) 60.0 L d) 40.0 L

Answers

Questions	Answer
1	A
2	C
3	B
4	C

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