GASES

1. Common Quantities of Gases

A. Pressure

Pressure is the amount of force exerted per unit area of surface

$$\text{Pressure} = \frac{\text{Force}}{\text{Area}} \text{ or } P = \frac{F}{A}$$

- Units of pressure are, g/cm^2, kg/cm^2, N/m^2, Pa or kPa
- Pressure of a liquid depends on the density and height of the liquid. ***P = h. d***
- Atmospheric pressure is due to the mass of gases in the atmosphere. It has a mass of 5.10^{15} tons.
- Atmospheric pressure can be measured by barometers.
- Units for atmospheric pressure are atm, cmHg, mmHg or torr.
- Evangelista Torricelli discovered the atmospheric pressure.
- 1 atm = 76 cmHg and is called standard atmospheric pressure.

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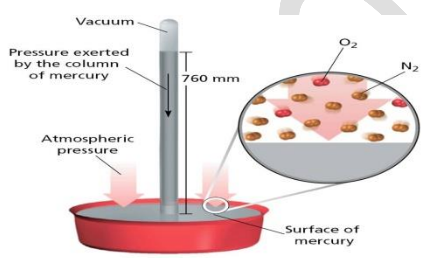
P = h. d

= $76 cm \times 13.6 g/cm^3$

= $1033.6 g/cm^2$

= 1 atm is standard atmospheric pressure measured at 0^oC and at sea level.



Units of Pressure

Unit	Abbreviation	Equivalent number of pascals
Atmosphere	atm	1 atm = 101 325 Pa
Bar	bar	1 bar = 100 025 Pa
Millimeter of mercury	mm Hg	1 mm Hg = 133.322 Pa
Pascal	Pa	1
Pounds per square inch	psi	1 psi = $6.892\ 86 \times 10^3$ Pa
Torr	torr	1 torr = 133.322 Pa

Example 1

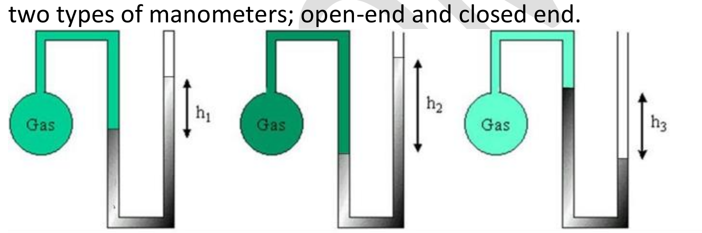
What is the minimum height of glass used to measure the atmospheric pressure at STP if water is used instead of mercury?

Solution

Atmospheric pressure is 76 cmHg at STP, therefore

$P = h.d = 76 \times 13.6 = h_{water} \times 1 \Rightarrow h_{water} = 1033\ cm = 10.33\ m$

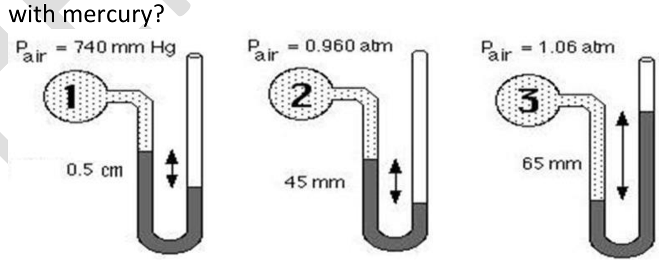
Gas pressures are measured by manometers, there are two types of manometers; open-end and closed end.



$P_{gas} = h_1$ $P_{gas} = P_{atm} + h_2$ $P_{gas} = P_{atm} - h_3$

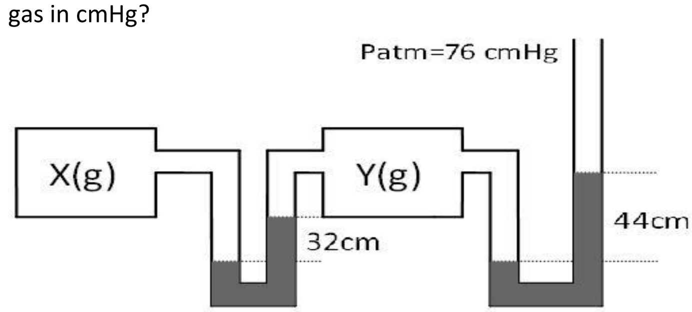
Example 2

Find the pressures of gases in each of the manometers filled with mercury?



Example 3

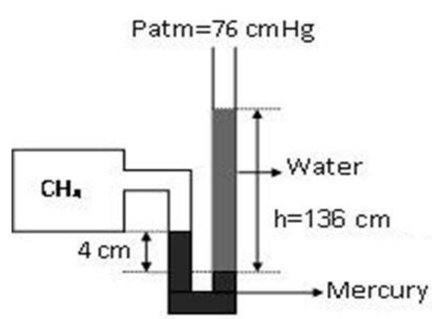
In the system given in figure below, what is the pressure of X gas in cmHg?



Example 4

Calculate the pressure of CH_4 gas in cmHg in the manometer given in the figure?

($d_{Hg}=13.6\ g/cm^3$, $d_{H2O}=1\ g/cm^3$)



B. Volume

- Volume of gases are measured as the volume of their containers.
- Gaseous substances do not have constant volume, they can expand or shrink to fit their containers.
- 1 mole of any gas occupies 22.4 L volume at STP.
- Units to express about the volume of gases are L, mL, dm^3 and cm^3.

C. Temperature

- Temperature is another important factor which affects all substances not only solids or liquids but also gases.
- Gaseous substances expand when temperature is increased.
- The unit is usually used to measure the temperature of gases is Kelvin (K).
- $T(K) = t(^oC) + 273$

Example 5

Convert the units of temperature in each of the following.

A. 27^oC = …….K
B. 546 K = …….^oC
C. 127^oC = …….K
D. -73^oC = ……..K

D. Amount

- Generally amount of gases are expressed in moles. But sometimes it can be given in volume or mass. Following equation is used to calculate the moles of gases in moles;

$$n = \frac{N}{N_A} = \frac{m}{MM} = \frac{V}{22.4}$$

Example 6

Convert the following amounts in mole.

A. 4.4 g CO_2 gas = …….mol
B. 11.2 L of He gas at STP = …….mol
C. $36.12x10^{22}$ NO_2 molecules = …….mol
D. 4.48 dm^3 SO_2 gas at STP = ……..mol

2. Kinetic Theory of Gases

- The kinetic-molecular theory is a model that is used to predict gas behavior. Ludwig Boltzmann, Austrian physicist, stated that about gases;
- Volume of a molecule of a gas is extremely small in comparison to the total volume of the gas. Therefore the volumes of gas molecules are neglected, but their masses are not.
- The attractive and repulsive forces between gas molecules are negligible.
- The motions of gas molecules are constant, rapid, random, straight line and three-dimensional.
- Gas molecules collide with each other and with the walls of their container. These collisions are totally elastic collisions. Therefore total kinetic energy remains constant.
- The average kinetic energy of the molecules of any gas is the same at a given temperature and it is directly proportional to the absolute (Kelvin) temperature.

3. Graham's Law of Diffusion

- The distribution of gases in air or in other gases is called the diffusion of gases. (Thomas Graham first studied on, in 1826)
- Graham found out that small gas molecules move more rapid than larger gas molecules.
- Movement of gas molecules through on a very small hole into a region of lower pressure is known as effusion.
- According to kinetic theory of gases, at the same temperature, kinetic energy of gas molecules are the same.

$$KE_X = KE_Y$$

$$\frac{1}{2} m_x V_x^2 = \frac{1}{2} m_Y V_Y^2$$

- The ratio of the molecular masses is the same as the ratio of molar masses, then;

$$\frac{V_x}{V_y} = \sqrt{\frac{m_y}{m_x}} = \sqrt{\frac{MM_y}{MM_x}}$$

- If densities or time needed to move is considered the formula will be ;

$$\frac{V_x}{V_y} = \sqrt{\frac{MM_y}{MM_x}} = \sqrt{\frac{d_y}{d_x}} = \frac{t_y}{t_x}$$

Example 7

Compare the diffusion rates He and CH_4 at the same temperature (He:4, CH_4: 16).

Solution

$$\frac{V_{He}}{V_{CH_4}} = \sqrt{\frac{MM_{CH_4}}{MM_{He}}} \quad V_{He} = 2 \times V_{CH_4}$$

Example 8

Under the same conditions, gas Y has a diffusion rate 4 times greater than that of SO_2 gas. What is the molecular mass of the gas Y? (S:32, O: 16).

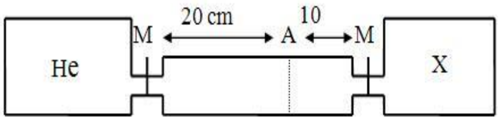
Solution

$$\Rightarrow \frac{V_Y}{V_{SO_2}} = \sqrt{\frac{MM_{SO_2}}{MM_Y}} = 4$$

$$V_Y = 4\times V_{SO_2}$$

Example 9

When the valves of M opened simultaneously, the gases of He and X meet at the point of A. What is the molecular mass of X?(He: 4)



4. Ideal Gas Equation

- There are four fundamental gas laws summarized as follows;
  Boyle's Law-Pressure & Volume Relation
  Charle's Law-Temperature & Volume Relation
  Gay-Lussac's Law-Temperature & Pressure Relation
  Avagadro's Law- Amount & Volume Relation
- These laws combine and form an equation called ideal gas equation.
  PV = nRT

Where P is pressure in atm, V is volume in L, n is mole number in mol, R is constant and T is temperature in K.

Example 10

What is the value of ideal gas constant at STP?

Solution

At STP 1 mol gas has a volume of 22.4 L.

$$PV = nRT$$

$$1atm\times 22.4L = 1mol\times R\times(0^oC+273)$$

$$R = 0.082\,\frac{atmL}{molK}$$

$$\left(\sqrt{\frac{64}{MM_Y}}\right)^2 = 4^2 \Rightarrow \frac{64}{MM_Y} = 16 \Rightarrow MM_Y = \frac{64}{16}$$

Example 11

What would be the pressure in atm exerted by 16 g He in 16.4 L container at -73^{o}C? (R: 0.082)

Solution

$$n_{He} = \frac{m}{MM} = \frac{16}{4} = 4mol$$

$$PV = nRT$$

$$P\times 16.4 = 4\times 0.082\times(-73^oC+273)$$

$$P = 4atm$$

Example 12

What is the volume of 33 g CO_2 in L under a pressure of 4 atm and at a temperature of 273^{o}C? (R: 0.082)

Solution

$$n_{CO_2} = \frac{m}{MM} = \frac{33}{44} = 0.75mol$$

$$PV = nRT$$

$$4\times V = 0.75\times 0.082\times(273^oC+273)$$

$$V = 8.39L$$

Example 13

What is the temperature in ^{o}C of 0.2 mol of a gas that exerts a pressure of 2280 mmHg in a 400 mL closed container?
(R: 0.082)

Solution

$$P = \frac{2280}{760} = 3atm,\quad V = \frac{400}{1000} = 0.4L$$

$$PV = nRT$$

$$3\times 0.4 = 0.2\times 0.082\times(t^oC+273)$$

$$t = -200^oC$$

Example 14

The density of CH_4 gas under a pressure of 4.1 atm and at a temperature of t ^{o}C is 2.5 g/L. What is the value of t? (CH_4: 16)
(R: 0.082)

Solution

$$PV = nRT\text{, then}$$

$$n = \frac{m}{MM} \text{ and } d = \frac{m}{V} \Rightarrow P\times\frac{m}{d} = \frac{m}{MM}\times R\times T$$

$$P = \frac{d}{MM}\times R\times T,\quad 4.1 = \frac{2.5}{16}\times 0.082\times(t^oC+273)$$

$$t^oC = 47^oC$$

Example 15

What is the density of argon in g/L gas under a pressure of 4 atm and at a temperature of 273^{0}C? (Ar: 40)
(R: 0.082)

Solution

$$P = \frac{d}{MM}\times R\times T,\quad 4 = \frac{d}{40}\times 0.082\times(273+273)$$

$$d = 3.57g/L$$

5. Gas Mixtures and Partial Pressure

- Individual pressure of a gas in a mixture of gases is called <u>*partial pressure*</u>.
- Partial pressure of a gas is directly proportional to its mole fraction in the mixture

In a mixture of two gases;

$$P_{Total} = P_A + P_B \text{ and } n_{Total} = n_A + n_B$$

Partial Pressures;

$$P_A = \frac{n_A}{n_{Total}}\times P_{Total} \text{ and } P_B = \frac{n_B}{n_{Total}}\times P_{Total}$$

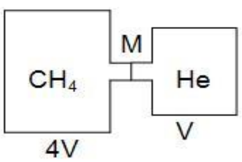
Example 16

There are 4.4 g of CO_2 and 2.8 g of N_2 gases in a 11.2 L container at 27^{o}C. Calculate the partial pressures of CO_2 and N_2 gases in the mixture. (R: 0.082, C:12, O:16, N:14)

Solution

$$n_{CO_2} = \frac{4.4}{44} = 0.1 \text{ mol and } n_{N_2} = \frac{2.8}{28} = 0.1 \text{ mol}$$
$$n_{total} = 0.1 + 0.1 = 0.2 \text{ mol and } P_{total}V = n_{total}RT$$

Example 17



The balloons contain CH_4 and He gases with equal masses at the same temperature. If the pressure of CH_4 gas is equal 1 atm, what will be the final pressure in atm after the valve, M, is opened? (C:12, H:1, He:4)

Solution

$$P_{total} \times 11.2 = 0.2 \times 0.082 \times (27 + 273)$$
$$P_{total} = 0.44 \text{ atm}$$
$$P_{CO_2} = \frac{0.1}{0.2} \times 0.44 = 0.22 \text{ atm and } P_{N_2} = \frac{0.1}{0.2} \times 0.44 = 0.22 \text{ atm}$$

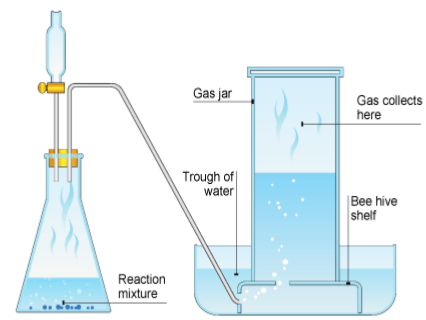
$$n_{CH_4} = \frac{m}{16} \text{mol and } n_{He} = \frac{m}{4} \text{mol} \Rightarrow n_{total} = \frac{m}{16} + \frac{m}{4} = \frac{5m}{16} \text{mol}$$
$$PV = nRT \text{ for } CH_4 \Rightarrow 1 \times 4V = \frac{m}{16}\text{RT} \Rightarrow \frac{m}{V}RT = 64$$
$$P_{total} \times V_{total} = n_{total}RT \Rightarrow$$
$$P_{total} \times \cancel{5}V = \frac{\cancel{5}m}{16}RT \Rightarrow P_{total} = \frac{m}{16V}RT$$
$$P_{total} = \frac{64}{16} = 4 \text{ atm}$$

6. Vaporization, Boiling and Boiling Point

- The liquid molecules which have enough kinetic energy to overcome the surface tension and the attraction force between molecules escape from the liquid and become vapor. The process in which liquid molecules turn into vapor is called **vaporization**.
- Vaporization may take place at any temperature. If not, we would have some difficulty in our daily life. For instance, wet clothes could not dry at room temperature.
- When the vapor pressure of a liquid is equal to the atmospheric pressure, the liquid starts to boil.
- Boiling point is the temperature at which the vapor pressure of a liquid is equal to atmospheric pressure. The temperature at which the vapor pressure of a liquid is equal to the standard atmospheric pressure is called the normal boiling point.
- In general, the higher vapor pressures the lower boiling point at constant temperature.
- Liquids boil at lower temperatures than the standard boiling points at high altitude.

7. Collecting Gases over Water

When a gas is collected over water in a bottle vapor pressure of water must be considered as well as the pressure of the gas. Therefore total pressure in the bottle is equal to;



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

Example 18

Some N_2 gas is collected over water at 18.5^{o}C and a total pressure of 756 mmHg. Calculate the partial pressure of N_2 in mmHg. (The vapour pressure of water is equal to 16 mmHg at 18.5^{o}C)

Solution

$$P_{total} = P_{N_2} + P_{water\ vapor}$$
$$756 = P_{N_2} + 16$$
$$P_{N_2} = 740 \text{ mmHg}$$