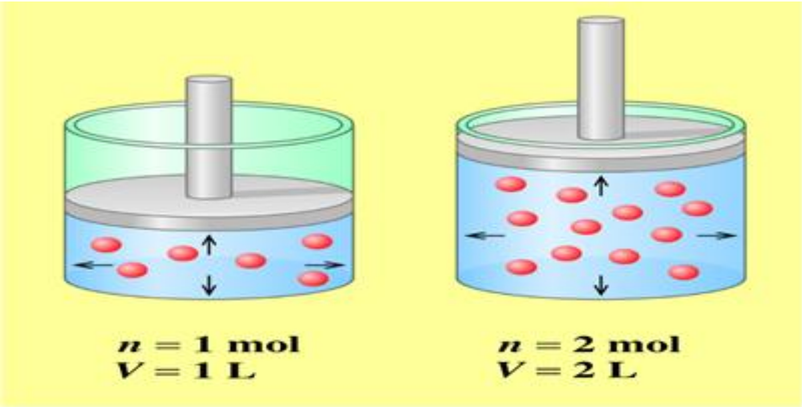
# Avogadro's law

- ***Avogadro's law*** States that the volume of a sample of gas is directly proportional to the number of moles in the sample at constant temperature and pressure.

$$V \alpha n \qquad \frac{V1}{n1} = \frac{V2}{n2}$$

- If the amount of gas in a container is **increased** the volume is **increased**.
- If the amount of gas in a container is **decreased** the volume is **decreased**.




***<u>Examples1</u>*** **a sample of gas with a volume of 9.20L is known to contain 1.225 mole if the amount of gas is increased to 2,85 mol what new volume will result if the pressure and temperature remain constant?** **Given** $V_1$ = 9.20 L $V_2$ = ???? $n_1$=1.225 mol $n_2$= *2.85 mol*

***<u>Solution</u>*** $$\frac{V1}{n1} = \frac{V2}{n2}$$

$$\frac{9.20\ \mathrm{L}}{1.225\ \mathrm{mol}} = \frac{V2}{2.85\ mol}$$

$$\mathbf{V_2 = 21.4\ L}$$

***Examples 2*** **What happens to the density of a gas as :**

**A) The gas is heated in a constant-volume container**

**b) The gas is compressed at constant temperature**

**c) Additional gas is added to a constant –volume container?**

## Ideal Gas Law

The law that describes the pressure, volume , temperature, and number of moles of a gas.

- So far we've seen that

$$V \propto 1/P \quad \text{(Boyle's law)}$$
$$V \propto T \quad \text{(Charles' law)}$$
$$V \propto n \quad \text{(Avogadro's law)}$$

- Combining these, we get

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

then becomes

$$V = R\frac{nT}{P} \quad \text{or} \quad PV = nRT$$

Which is the ideal –gas equation **{** also called the ideal gas law **}**

- ***P*** **pressure. (1 atm)**
- ***V*** **volume.(1 L = 1000 ml), (1 L = 1000 $cm^3$)**
- ***n*** **moles ( mole = gram/M.Wt). =** $\frac{\text{grams}}{\frac{\text{grams}}{\text{mole}}}$
- ***T*** **temperature (K =( 0C + 273) where K is the temperature expressed in Kelvin**

***R*** **= the Gas Constant** = $0.0821\frac{\text{L .atm}}{\text{mol. K}}$ **or** $\mathbf{8.3145}\frac{\text{J}}{\text{mol. K}}$

# Gas density

Ideal Gas Law → $PV = n \quad RT$

$$PV = \frac{mass}{m.wt} RT$$

$$P\,V\,m.wt = mass \quad RT$$

**Density** → $\frac{mass}{V} = \frac{P\,m.wt}{RT}$

***Examples*1** **What is the density of carrbon tetrafluouride ($CF_4$) at 1.00 atm and 50 $^0$C ?**

Pressure= 1 atm

m.wt ( $CF_4$) = 12+ (18.9 x4) = 88 g/mole

R = 0.08206 L. atm /mole. K

K= 50 +2733.15 = 323.15K

$$Density = \frac{P\,m.wt}{RT}$$

$$\frac{(1\,\cancel{atm})\ (88\frac{g}{\cancel{mol}})}{(0.08206\ L.\cancel{atm}\ /\cancel{mole}.\cancel{K})\ (323.15\cancel{K})}$$

Density = 3.32 g/L

**Examples(2):** **A sample of $H_2$ gas has a volume of 8.56 L at a temperature of 0°C and pressure of 1.5 atm. Calculate the moles of gas present.**

***PV = nRT*** 1.5 atm x 8.56 L = n 0.082 L.atm/mole.K x (o+273)

**n=0.57mole**

**Examples(3):** **What volume does 40.0 g of N2 gas occupy at 10°C and 750 mmHg?**

$$\frac{1atm}{x} = \frac{760mmHg}{750mmHg} \qquad x= 0.98 \text{ atm}$$

P v=nRT 0.98 atm x v = $\frac{40}{2x\ 14}$ $x$ (0.082 L.atm/mol.k) x (10+273)k **v =33.82L**

**Examples (4):** **A 23.8L cylinder of contains oxygen at 20.0 °C and 732 mmHg. How many moles of oxygen does it contain?**

$$\frac{1atm}{x} = \frac{760mmHg}{732mmHg} \qquad x= 0.98 \text{ atm}$$

**P v=nRT 0.98 atm x 23,8 L = n (0.082 L.atm/mol.k) x 293k** **n =0.95 mol**

**Examples(5):** **The density of a gas was measured at 1.50 atm and 27 0C and found to be 1.95 g/L calculate the molar mass of the gas ?**

**P v=nRT** P v= $\frac{wt}{m.wt}$ RT $m.wt = \frac{wt}{v} \ \frac{RT}{P}$

$$m.wt = \frac{1.95 \text{ x } 0.082 \text{ x } (27+273)}{1.50}$$

$m.wt$ = **31 .98 g/mole**

**Examples (6):** **A 2.10 vessel contain 4.65gof gas at 1.00 atm and 27.0 0C what is the molar mass of the gas**

**P v=nRT** P v= $\frac{wt}{m.wt}$ RT $m.wt = \frac{wt}{v} \ \frac{RT}{P}$

$$m.wt = \frac{4.66g \text{ x } 0.082 \text{ x } 327}{2.1 \text{ x1 atm}}$$

$m.wt$ = **59.37 g/mole**

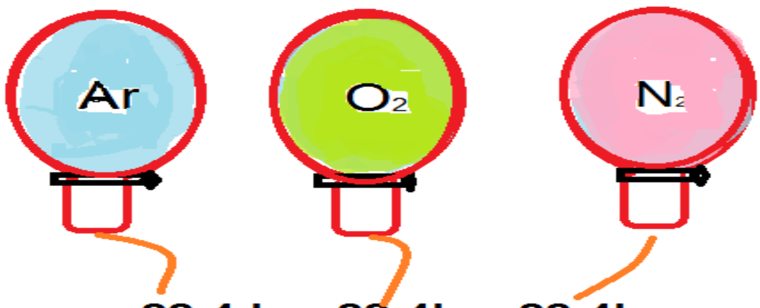
# STP: Standard Temperature & Pressure

Suppose we have 1.000 mol of an ideal gas at 1.000 atm and 273.15 k according to the ideal-gas equation the volum of the gas is:-

$$V(\text{Molar volume}) = \frac{nRT}{P} = \frac{(1.000\ mol)0.08206\frac{L.atm}{mol.k}}{1.000\ atm} = 22.41\ \ L$$

**Molar volume**: At **1atm** and **273.15 k** the volume of **1mole** of gas is **22.41 L**

**STP: Standard Temperature & Pressure =** 0 $^0$C (**273.15 k**) and **1.0 atm**




| | Ar | $O_2$ | $N_2$ |
|---|---|---|---|
| volume: | 22.4 L | 22.4L | 22.4L |
| Mass: | 40g | 32g | 28g |
| Quantity: | 1 mol | 1mol | 1mol |
| Presure:- | 1 atm | 1 atm | 1atm |
| Temperture:- | 273 K | 273 K | 273K |

**Example(1):** **What mass (in grams) of $Na_2O_2$ is necessary to consume 1.00 L of $CO_2$ at STP?**

$$2Na_2O_2(s) + 2CO_2(g) \rightarrow 2Na_2CO_3(s) + O_2(g)$$

**P v=nRT**

(1 atm ) (1.00L) = n (0.08206 ) 273.15 K    n= 0.04 mole

$n = \frac{wt}{m.wt}$    $0.04 \text{ mole} = \frac{wt}{78}$    **wt = 3.12**

**Example(2):** **What is the volume (in liters) occupied by 49.8 g of HCl at STP?** **Atomic weight of ( H=1 , Cl=35.5) ?**

PV = nRT

PV = M/MwRT

V = MRT/PM.w    (M.w = molecular weight of HCl =36.5 g/mole )

= 49.8x.0082x273/36.5x1

**= 30.54 litres**

# Dalton's Law of Partial Pressure

* Many gas samples are mixture of gases. For example, the air we breathe is a mixture of mostly oxygen and nitrogen gases.

*Since gas particles have no attractions towards one another, each gas in a mixture behaves as if it is present by itself, and is not affected by the other gases present in the mixture.

*In a mixture, each gas exerts a pressure as if it was the only gas present in the container. This pressure is called ***partial pressure*** of the gas.

* In a mixture, the sum of all the partial pressures of gases in the mixture is equal to the total pressure of the gas mixture. This is called ***Dalton's law*** of partial pressures.

$$P_{total} = \sum P_{partial}$$

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

**Total pressure of a gas mixture = Sum of the partial pressure of the gases in the mixture**

$$PV = nRT \quad \text{or} \quad P = \frac{nRT}{V}$$

**Therefore**

$$P_1 = \frac{n1\ R\ T1}{V1} \qquad P_2 = \frac{n2\ R\ T2}{V2} \qquad P_3 = \frac{n3\ R\ T3}{V3}$$

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

$$V_1 = V_2 = V_3 = V$$

$$T_1 = T_2 = T_3 = T$$

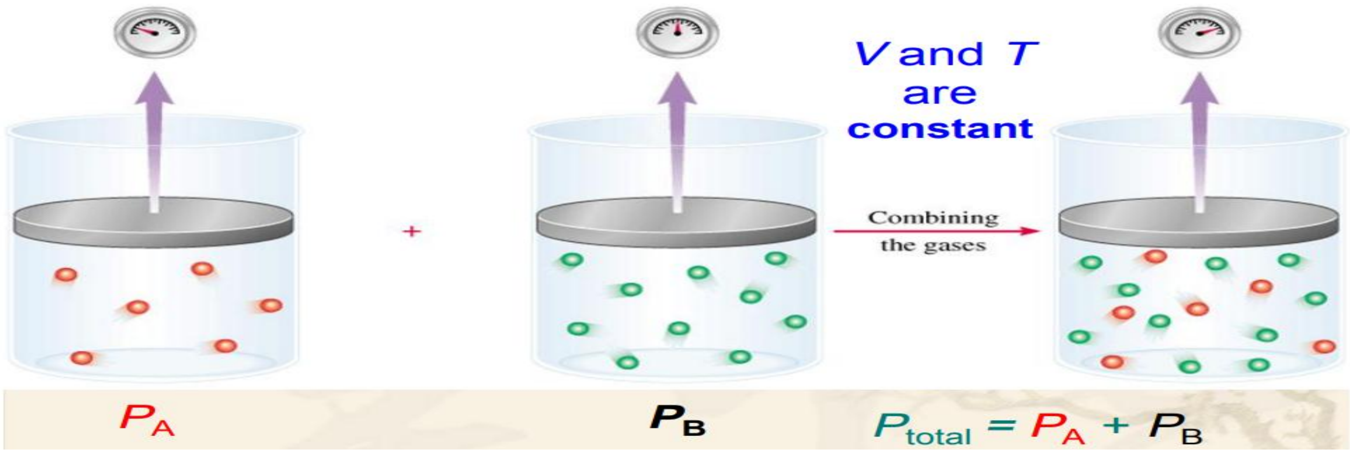
$$P_{total} = \frac{n1\ R\ T1}{V1} + \frac{n2\ R\ T2}{V2} + \frac{n3\ R\ T3}{V3}$$

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

In other wards, the total pressure of a mixture of gases is equal to the sum of the moles of all gases in the mixture times RT/V .the general form of this equation is

$$P_{total} = \left(\sum n_{each\ gas}\right)\frac{RT}{V}$$

The partial pressure of each gas in a mixture is proportional to the amount (mole) of gas present in the mixture. For example, in a mixture of gases consisting of 1 mole of nitrogen and 1 mole of hydrogen gas, the partial pressure of each gas is one-half of the total pressure in the container.




**Examples (1):** **Two 10L tanks, one containing propane gas at 300 torr and the other containing methane at 500 torr, are combined in a 10L tank at the same temperature. What is the total pressure of the gas mixture?** $P_{total} = P_1 + P_2 = 300\text{ torr} + 500\text{ torr} = \underline{\mathbf{800\ torr}}$

**Examples (2):** **A scuba tank contains a mixture of oxygen and helium gases with total pressure of 7.00 atm. If the partial pressure of oxygen in the tank is 1140 mmHg, what is the partial pressure of helium in the tank?** $\frac{1\ atm}{x} = \frac{760 mmg}{1140} = 1.5\text{ atm}$

$P_{total} = P_{oxygen} + P_{helium}$ 7atm= 1.5 atm + $P_{helium}$ $\underline{P_{helium} = \mathbf{5.5}\text{ atm}}$

**Examples (3):** **Determine the partial pressures and the total pressure in a 2.50-L vessel containing the following mixture of gases at 15.8°C: 0.0194 mol He, 0.0411 mol H2, and 0.169 mol Ne.**

**Solution:** **Step 1** Since each gas behaves independently, calculate the partial pressure of each using the ideal gas equation:

$$P_{He} = \frac{(0.0194\ mol\ He)\left(0.08206\ L.\frac{atm}{K.MOL}\right)(288.95\ k)}{2.50L} = 0.184 \text{ atm}$$

$$P_{H2} = \frac{(0.0411\ mol\ H2)\left(0.08206\ L.\frac{atm}{K.MOL}\right)(288.95\ k)}{5.00L} = 0.390 \text{ atm}$$

$$P_{Ne} = \frac{(0.169\ mol\ Ne)\left(0.08206\ L.\frac{atm}{K.MOL}\right)(288.95\ k)}{2.50\ L} = 1.60 \text{ atm}$$

$P_{total} = \sum P_{partial}$ $P_{total}$ = 0.184 atm + 0.390 atm + 1.60 atm = **2.17 atm**

**Examples (4):** **A 1.00-L vessel contains 0.215 mole of N2 gas and 0.0118 mole of H2 gas at 25.5° C. Determine the partial pressure of each component and the total pressure in the vessel?**

**Think About It** The total pressure in the vessel can also be determined by summing the number of moles of mixture components (0.215 + 0.0118 = 0.227 mol) and solving the ideal gas equation for $P_{total}$:

$$P_{total} = \left(\sum n_{each\ gas}\right)\frac{RT}{V} \qquad P_{total} = \frac{(0.227\text{ mol})(0.08206\text{L·atm/K·mol})(298.65\text{ K})}{1.00atm} = \mathbf{5.56\ atm}$$

## Kinetic Molecular Theory of Gases

**The** kinetic- molecular theory summarized by the following statements

1. Gases consist of small particles (atoms or molecules) that move randomly with rapid velocities.
2. Gas particles have little attraction for one another. Therefore, attractive forces between gas molecules can be ignored.
3. Gas particles collide with each other and with walls of the container. The force of collisions of the gas particles with the walls of the container causes pressure
4. The average kinetic energy of gas molecules is directly proportional to the absolute temperature (Kelvin).

# Graham's law Describes Diffusion & Effusion

The dependence of molecular speed on mass has two interesting consequences.

***Diffusion*:** Mixing of different gases by random molecular motion with frequent collisions is called diffusion.

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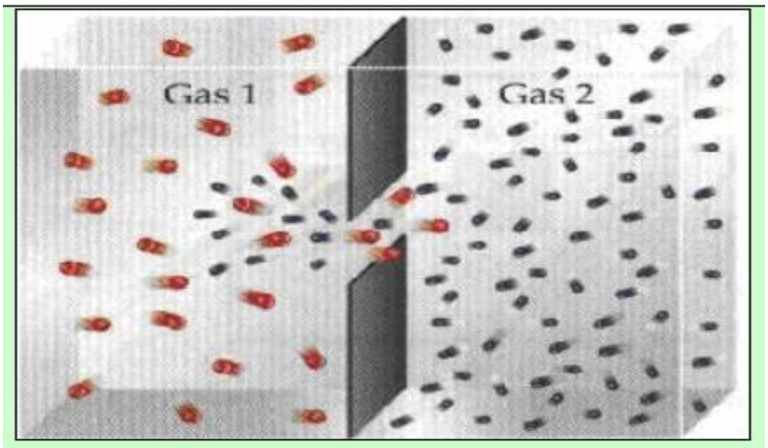
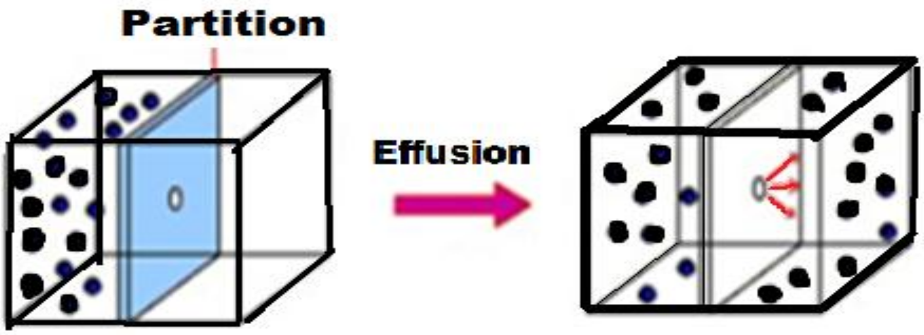



**Effusion:** Escape of gas molecules through a tiny hole into an evacuated space



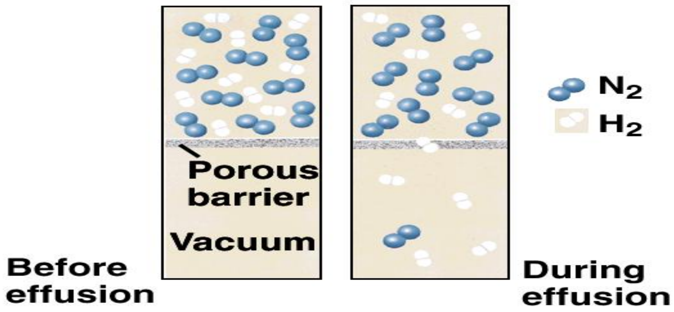
● **Graham 's Law** *studied relationship between effusion rates and molecular masses for series of gases.*

● *The Rate of effusion of a gas is* **inversely** *proportional to the square root of the molecular mass for series of gases(at constant p and T)*

● *Heavier gases effuse more slowly*

● *lighter gases effuse more rapidly*

$$\frac{Rate\ of\ effusion for\ gas1}{Rate\ of\ effusion\ for\ gas2} = \sqrt{\frac{MW2}{MW1}}$$




$$\frac{r(H2)}{r(N2)} = \sqrt{\frac{MW(N2)}{MW(H2)}}$$

Example : Calculate the ratio of the effusion rates of $N_2$ and $O_2$, $\frac{r_{N_2}}{r_{O_2}}$

$M_{N_2}$= 28 g/mol

$M_{O_2}$= 32 g/mol

$$\frac{r_{N_2}}{r_{O_2}} = \sqrt{\frac{M_{O_2}}{M_{N_2}}}$$

$$= \sqrt{\frac{32}{28}}$$

$$\frac{r_{N_2}}{r_{O_2}} \boxed{= 1.07}$$