Kinetic Energy of Ideal Gas.

Molecules of ideal gases possess only translational motion. So they possess only translational kinetic energy.

Quantity of gas	Kinetic energy
Kinetic energy of a gas molecule (Emolecule)	$=\frac{1}{2}m v_{rms}^2 = \frac{1}{2}m\left(\frac{3kT}{m}\right) = \frac{3}{2}kT$ $\left[\text{As}\, v_{rms} \sqrt{\frac{3kT}{m}}\right]$
Kinetic energy of 1 mole (M gram) gas (Emole)	$=\frac{1}{2}M v_{rms}^2 = \frac{1}{2}M\frac{3RT}{M} = \frac{3}{2}RT$ $\left[\text{As}\, v_{rms} = \sqrt{\frac{3RT}{M}}\right]$
Kinetic energy of 1 gm gas (Egram)	$=\frac{3}{2}\frac{R}{M}T = \frac{3}{2}\frac{kN_A}{mN_A}T = \frac{3}{2}\frac{k}{m}T = \frac{3}{2}rT$

Here m = mass of each molecule, M = Molecular weight of gas and NA = Avogadro number = 6.023 × 1023

Important points

(1) Kinetic energy per molecule of gas does not depends upon the mass of the molecule but only depends upon the temperature of the gas.

As $E = \frac{3}{2}kT$ or $E \propto T$ i.e. molecules of different gases say He, H2 and O2 etc. at same temperature will have same translational kinetic energy though their rms speed are different. $\left[v_{rms} = \sqrt{\frac{3kT}{m}}\right]$

(2) Kinetic energy per mole of gas depends only upon the temperature of gas.

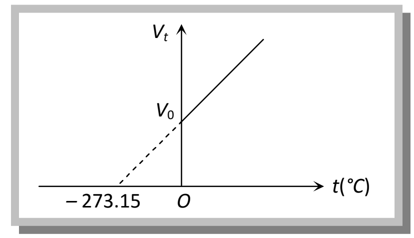
(3) Kinetic energy per gram of gas depend upon the temperature as well as molecular weight (or mass of one molecule) of the gas.

$$E_{gram} = \frac{3}{2}\frac{k}{m}T \qquad \therefore E_{gram} \propto \frac{T}{m}$$

From the above expressions it is clear that higher the temperature of the gas, more will be the average kinetic energy possessed by the gas molecules at T = 0, E = 0 i.e. at absolute zero the molecular motion stops.

(2) Charle's law

(i) If the pressure remains constant, the volume of the given mass of a gas increases or decreases by $\frac{1}{273.15}$ of its volume at 0°C for each 1°C rise or fall in temperature.



$$V_t = V_0\left(1 + \frac{1}{273.15}t\right)$$. This is Charle's law for centigrade scale.

(ii) If the pressure remaining constant, the volume of the given mass of a gas is directly proportional to its absolute temperature.

$V \propto T$ or $\frac{V}{T} = \text{constant}$ or $\frac{V_1}{T_1} = \frac{V_2}{T_2}$ [If m and P are constant]

(iii) $\frac{V}{T} = \frac{m}{\rho T} = \text{constant}$ [As volume $V = \frac{m}{\rho}$]

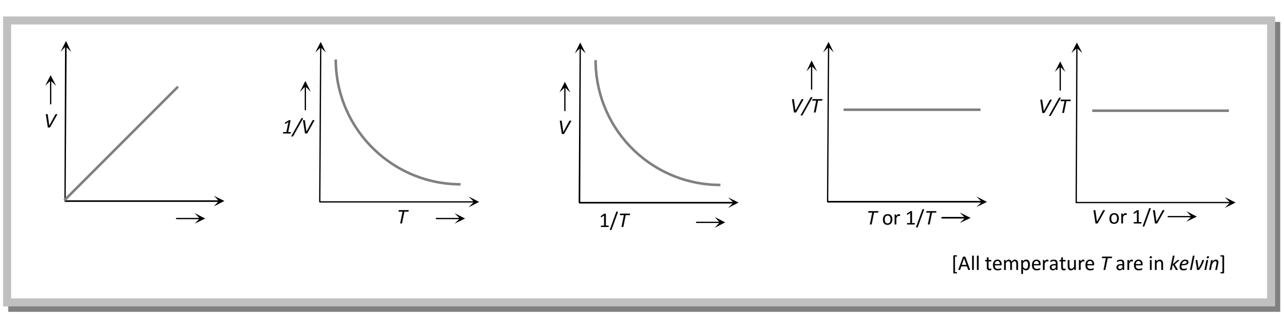
or $\rho T = \text{constant}$ or $\rho_1 T_1 = \rho_2 T_2$ [As m = constant]

(iv) According to kinetic theory of gases $P = \frac{1}{3}\frac{mN}{V}v_{rms}^2$

or $P \propto \frac{\text{Mass of gas}}{V}T$

If mass and pressure of the gas remains constant then $V \propto T$. This is in accordance with Charles law.

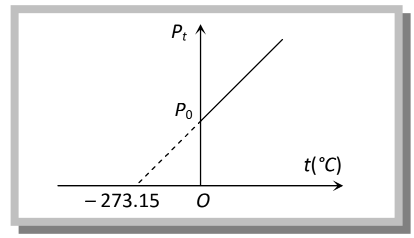
(v) Graphical representation: If m and P are constant



[All temperature *T* are in *kelvin*]

(3) Gay-Lussac's law or pressure law

(i) The volume remaining constant, the pressure of a given mass of a gas increases or decreases by $\frac{1}{273.15}$ of its pressure at 0°C for each 1°C rise or fall in temperature.



$$P_t = P_0\left[1 + \frac{1}{273.15}t\right]$$

This is pressure law for centigrade scale.

(ii) The volume remaining constant, the pressure of a given mass of a gas is directly proportional to its absolute temperature.

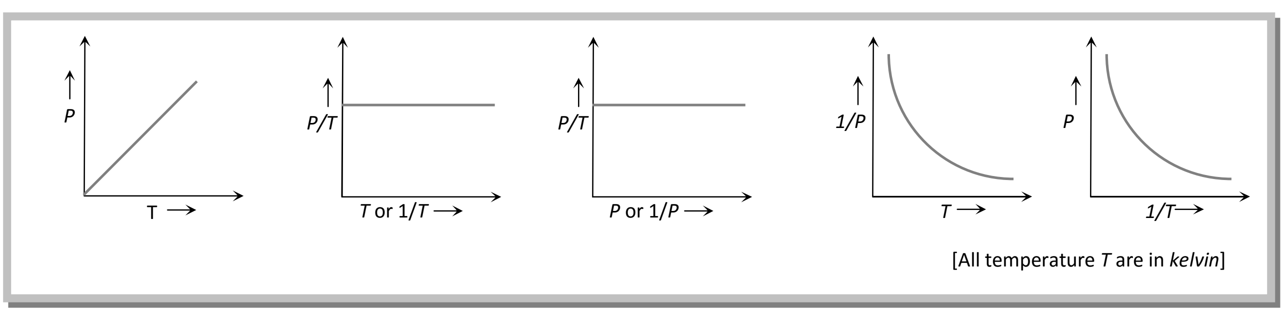
$P \propto T$ or $\frac{P}{T} = \text{constant}$ or $\frac{P_1}{T_1} = \frac{P_2}{T_2}$ [If m and V are constant]

(iii) According to kinetic theory of gases $P = \frac{1}{3}\frac{mN}{V}v_{rms}^2$ [As $v_{rms}^2 \propto T$]

or $P \propto \frac{\text{mass of gas}}{V}T$

If mass and volume of gas remains constant then $P \propto T$. This is in accordance with Gay Lussac's law.

(4) Graphical representation: If m and V are constants



(5) Avogadro's law: Equal volume of all the gases under similar conditions of temperature and pressure contain equal number of molecules.

According to kinetic theory of gases $PV = \frac{1}{3} m N v_{rms}^2$

For first gas, $PV = \frac{1}{3} m_1 N_1 v_{rms(1)}^2$ .....(i)

For second gas, $PV = \frac{1}{3} m_2 N_2 v_{rms(2)}^2$ .....(ii)

From (i) and (ii) $m_1 N_1 v_{rms1}^2 = m_2 N_2 v_{rms2}^2$ .....(iii)

As the two gases are at the same temperature $\frac{1}{2} m_1 v_{rms1}^2 = \frac{1}{2} m_2 v_{rms2}^2 = \frac{3}{2} kT$ $\Rightarrow$ $m_1 v_{rms1}^2 = m_2 v_{rms2}^2$ .....(iv)

So from equation (iii) we can say that $N_1 = N_2$. This is Avogadro's law.

(i) Avogadro's number (NA): The number of molecules present in 1 gm mole of a gas is defined as Avogadro number.

$N_A = 6.023 \times 10^{23}$ per gm mole $= 6.023 \times 10^{26}$ per kg mole.

(ii) At S.T.P. or N.T.P. (T = 273 K and P = 1 atm) 22.4 liter of each gas has 6.023×10^{23} molecule.

(iii) One mole of any gas at S.T.P. occupy 22.4 liter of volume

Example : 32 gm oxygen, 28 gm nitrogen and 2gm hydrogen occupy the same volume at S.T.P.

(iv) For any gas 1 mole = M gram = 22.4 liter = 6.023 × 1023 molecule.