GASEOUS STATE SHORT NOTES

Gas has three main properties

a) Volume
b) Temperature
c) Pressure

Units of Volume

$1 m^3 = 10^3 L = 10^3 dm^3$
$= 10^6 cc = 10^6 ml$
$1 L = 1 dm^3$
$1 ml = 1 cc$

Units of pressure

1 atm = 76 cm of Hg
= 760 mm of Hg
= 760 torr
= 1.01×10^5 Pa
= 1.01 Bar

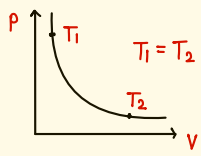
Units of Temperature

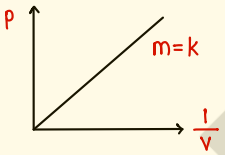
T(K) = T(°C) + 273
S.I unit = Kelvin

Boyle's law :

$P \propto \frac{1}{V}$ constant T and P

$PV = K$
$P_1V_1 = P_2V_2$



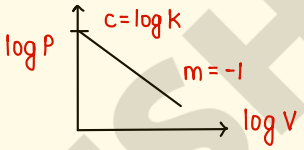


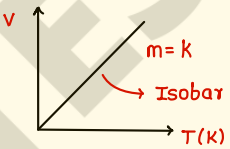
Charle's law

$V \propto T$ constant P and V

$\frac{V}{T} = K$

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





Absolute zero

T = 0 K = −273 °C

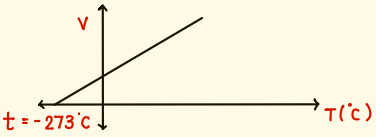


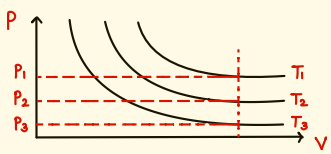
Gay Lussac's Law :

$P \propto T$ constant V and P

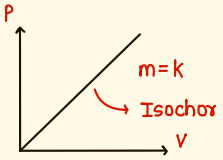
$\frac{P}{T} = K$

$\frac{P_1}{T_1} = \frac{P_2}{T_2}$





$P_1 > P_2 > P_3$ $T_1 > T_2 > T_3$



Avagadro's law

Under similar conditions of temperature and pressure; equal volume of all gases contain equal number of molecules.

Volume of gas ∝ moles

Equation of state :

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

Ideal gas equation

$PV = nRT$ → Temperature, Universal gas constant, moles, Volume, Pressure

R = 0.0821 L atm / mol k
= 0.0831 L Bar / mol k
= 8.314 J / mol k
= 2 Cal/mol K

Density of gas :

$d = \frac{PM}{RT}$ → Molar mass, g/L, atm, Kelvin

$0.0821 = \frac{1}{12}$

Dalton's law of partial pressure :

P_A = Partial pressure of A
P_B = Partial pressure of B
P_C = Partial pressure of C

$P_{Total} = P_A + P_B + P_C$

Expression of Dalton's rule :

$P_A = x_A P_T$
$P_B = x_B P_T$
$P_C = x_C P_T$

$x_A + x_B + x_C = 1$

Validation of dalton's law of partial pressure :

Valid only for non reacting gases at room temperature.

e.g. $H_2 + O_2 \rightarrow$ Valid.
$NH_3 + HCl \rightarrow$ Invalid.

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GASEOUS STATE SHORT NOTES

Graham's law of diffusion

$$r = \frac{PA}{\sqrt{MT}} \qquad \frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}} = \sqrt{\frac{d_2}{d_1}}$$

Calculation of determination of diffusion

a) Volume

$$\frac{r_1}{r_2} = \frac{V_1/t_1}{V_2/t_2} = \sqrt{\frac{M_2}{M_1}}$$

b) Distance

$$\frac{r_1}{r_2} = \frac{d_1/t_1}{d_2/t_2} = \sqrt{\frac{M_2}{M_1}}$$

c) moles

$$\frac{r_1}{r_2} = \frac{n_1/t_1}{n_2/t_2} = \sqrt{\frac{M_2}{M_1}}$$

Kinetic theory of gases

a) Gas is made up of large number of small tiny particles called molecules.
b) Effect of force of gravity on gas molecules is negligible due to their very small mass.
c) Particles of gas are in constant random motion.
d) Collisions are perfectly elastic in nature.
e) Average kinetic energy of a gas sample depends only on temperature.
f) Volume occupied by gas molecules is negligible in comparison to volume of container.
g) There is no force of attraction between gas molecules

Molecular speeds

$$V_{MPS} = \sqrt{\frac{2RT}{M}} = \sqrt{\frac{2P}{d}}$$

$$V_{avg} = \sqrt{\frac{8RT}{\pi M}} = \sqrt{\frac{8P}{\pi d}}$$

$$V_{RMS} = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3P}{d}}$$

$V_{MPS} < V_{avg} < V_{RMS}$

Average kinetic energy – depends only on temperature

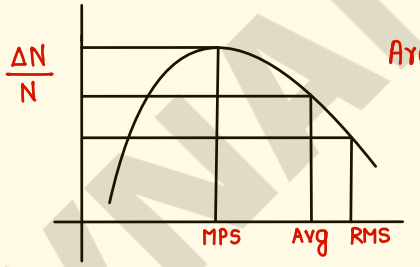
$(KE)_{avg} = \frac{3}{2} RT$ for 1 mole

$(KE)_{avg} = \frac{3}{2} nRT$ for n moles

$(KE)_{avg} = \frac{3}{2} kT$ for 1 molecule

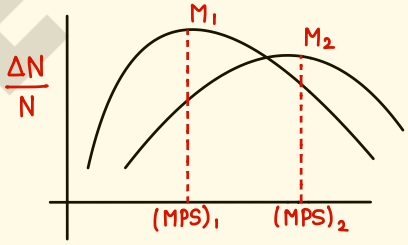
K = Boltzmann constant = 1.38×10^{-23}

Maxwell Boltzmann distribution



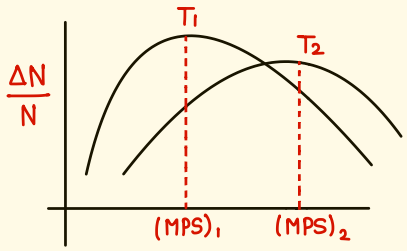
Area under curve = 1

Effect of molecular mass



$M_1 > M_2$

Effect of temperature :



$T_1 < T_2$

– On increasing temperature :

a) entire curve shift towards right.
b) Most probable speed increases.
c) Fraction of molecules having most probable speed decreases.
d) Area under curve remains same.

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GASEOUS STATE SHORT NOTES

Real gas

Gases which do not follow $PV = nRT$ at all temperature and pressure.

Note : A gas behaves as ideal gas only under two conditions

a) Low pressure

b) High temperature.

Vanderwaal's equation

a) Pressure correction = $P + \frac{an^2}{V^2}$ a = measure of force of attraction.

b) Volume correction = $V - nb$ b = size/volume of molecules.

Vanderwaal gas equation for real gas

$$\left(P + \frac{an^2}{V^2}\right)(V - nb) = nRT$$

Compressibility factor (Z)

$Z = \frac{PV}{nRT}$ $Z = \frac{V_{real\ gas}}{V_{ideal\ gas}}$

Case 1 : $Z = 1$

$PV = nRT$

Ideal gas

Case 2 : $Z > 1$ at High pressure +ve deviation

$V_{real} > V_{ideal}$

$P(V - nb) = nRT$ $Z = 1 + \frac{Pb}{RT}$

Case 3 : $Z < 1$ at low pressure −ve deviation

$V_{real} < V_{ideal}$

$\left(P + \frac{an^2}{V^2}\right)V = nRT$ $Z = 1 - \frac{an}{RTV}$

Note - To liquify gas

a) High pressure b) Low temperature.

Boyle's Law : Temp at which real gases obey ideal gas over a wide range of pressure.

$T_b = \frac{a}{bR}$

Critical temperature - Temperature above which gas cannot be liquified however high pressure is applied.

$T_c = \frac{8a}{27bR}$

Critical pressure - Minimum pressure that must be applied on a gas at T_c to liquify.

$P_c = \frac{a}{27b^2}$

Critical Volume - Volume occupied by 1 mole of gas at T_c and P_c

$V_c = 3b$

Inversion temperature - Temperature at which there is no heating or cooling of gas occur during adiabatic expansion.

$T_i = \frac{2a}{bR}$

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