

Conversions

Volume

1 Litre = 10^3 mL = 10^3 cm³ = 10^{-3} m³ = 1 dm³

Pressure

1 atm = 76 cm of Hg = 760 mm of Hg = 760 torr

1 atm = 1.01×10^5 Nm²

1 atm = 1.01×10^5 Pa

1 bar = 0.9862 atm $\approx$ 1 atm

mL = milli-litre

cm = centimeter

m = meter

dm = decimeter

atm = atmosphere

N = Newton

Pa = Pascals

Boyle's Law

Boyle's Law $PV = K$ $P_1V_1 = P_2V_2$

If the temperature is constant, as pressure of a gas increases the volume decreases.

According to Boyle's law,

$V \propto P^{-1}$

or, $PV = \text{constant}$ @particular temperature.

or, $P_1V_1 = P_2V_2$

Charles' Law

Charles's Law $\frac{V}{T} = K$ $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

If the pressure is constant, as temperature of a gas increases the volume increases.

Temperature must be in Kelvin for all gas laws

According to Charles' Law,

$V \propto T$ @constant Pressure

therefore, $V_1/T_1 = V_2/T_2$

or, $V_1T_2 = V_2T_1$

Gay-Lussac's Law

Gay-Lussac's Law $\frac{P}{T} = K$ $\frac{P_1}{T_1} = \frac{P_2}{T_2}$

As the temperature of an enclosed gas increases, the pressure increases at constant volume.

According to Lussac's law,

$P \propto T$

or, $PT^{-1} = \text{constant}$ @constant volume

$P_1T_2 = P_2T_1$

Ideal Gas Equation

$PV = nRT$

or, $PVT^{-1} = \text{constant} = nR$

R = Universal Gas Constant

Universal Gas Constant

[R] = Lit atm K⁻¹ mol⁻¹ = work done per Kelvin per mole

R = 0.0821 Lit atm K⁻¹ mol⁻¹

R = 8.314 J K⁻¹ mol⁻¹ = 8.314 Pa cm³ K⁻¹ mol⁻¹

R = 1.987 cal K⁻¹ mol⁻¹

R = 5.189×10^{19} eV K⁻¹ mol⁻¹

Dalton's Law of Partial Pressure

According to Dalton's Law of Partial Pressure,

$P = P_1 + P_2 + P_3 + \dots + P_n = \Sigma P_n$

Where P is the total pressure and P_1 , P_2 , P_3 ... are the partial pressures of gases 1, 2, 3 and so on.

$P_1 = P \cdot x_1$

where x_1 is the mole fraction of gas 1.

Graham's Law of Diffusion

$$r \propto \sqrt{\frac{1}{M}} \propto \sqrt{\frac{1}{d}} \\ \frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}} = \sqrt{\frac{d_2}{d_1}} = \frac{n_2}{n_1}$$

r = rate of diffusion

M = molar mass of given gas

d = density of given gas

n = number of moles of given gas

Velocity of Gas Particles

$$\text{Most Probable Velocity} \\ \alpha = \sqrt{\frac{2R}{M}} = \sqrt{\frac{2PV}{M}} \\ \text{Average Velocity} \\ v = \sqrt{\frac{8RT}{\pi M}} = \sqrt{\frac{8RPV}{\pi M}} \\ \text{Root Mean Square Velocity} \\ u = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3PV}{M}}$$

$\alpha : v : u = 1 : 1.128 : 1.224$

Kinetic Energy

$$PV = \frac{1}{2} m N u^2 \\ PV = \frac{1}{2} M u^2 \\ \text{Molar Kinetic Energy} \\ K.E. = \frac{1}{2} M u^2 = \frac{3}{2} PV = \frac{3}{2} RT \\ \text{Molecular Kinetic Energy} \\ K.E. = \frac{3}{2} \frac{RT}{N_A} = \frac{3}{2} k_B T$$

k = Boltzmann Constant = 1.380649×10^{-23} joule per kelvin

Van der waal's Equation

$$\text{Van der waal's Equation} \\ \left(P + \frac{a n^2}{V^2}\right)(V - nb) = nRT \\ \text{Dimensions of a \& b} \\ [a] = \text{atm} \cdot L^2 \cdot \text{mol}^{-2} = N \cdot m^4 \cdot \text{mol}^{-2} \\ [b] = L \cdot \text{mol}^{-1} = m^3 \cdot \text{mol}^{-1}$$

Molecular Collision

$$\text{Critical Temperature} \\ t_c = \frac{8a}{27bR} \\ \text{Critical Pressure} \\ P_c = \frac{a}{27b^2} \\ \text{Critical Volume} \\ V_c = 3b \\ \text{Relation between T, V \& P} \\ \frac{PV_c}{RT_c} = \frac{8}{3}$$

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