Gaseous State Cheat Sheet

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Conversions

Volume

1 Litre= 10^3 mL= 10^3 cm^3= 10^{-3} m^3= 1 dm^3

Pressure

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

1 atm= $1.01*10^5$ Nm^2

1 atm= $1.01*10^5$ Pa

1 bar= 0.9862 atm≈ 1 atm

mL= milli-litre

cm= centimeter

m= meter

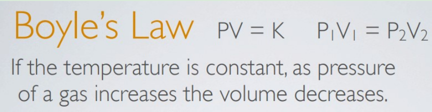
dm= decimeter

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.

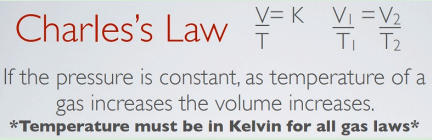
According to Boyle's law,

$V \propto P^{-1}$

or, PV = 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

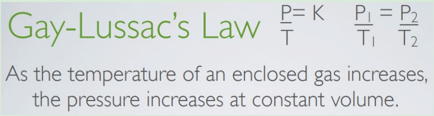
According to Charles' Law,

$V \propto T$ @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} = constant @constant volume

$P_1T_2 = P_2T_1$

Ideal Gas Equation

PV=nRT

or, PVT^{-1} = 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 cm^3 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,

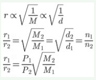
$P = P_1+P_2+P_3+.....+P_n = \Sigma P_n$

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= $P*x_1$

where x_1 is the mole fraction of gas 1.

Graham's Law of Diffusion



$$r \propto \sqrt{\frac{1}{M}} \propto \sqrt{\frac{1}{d}}$$

$$\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}} = \sqrt{\frac{d_2}{d_1}} = \frac{n_1}{n_2}$$

$$\frac{r_1}{r_2} = \frac{P_1}{P_2}\sqrt{\frac{M_2}{M_1}}$$

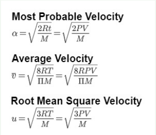
r= rate of diffusion

M= molar mass of given gas

d= density of given gas

n= number of moles of given gas

Velocity of Gas Particles



Most Probable Velocity

$$\alpha = \sqrt{\frac{2Rt}{M}} = \sqrt{\frac{2PV}{M}}$$

Average Velocity

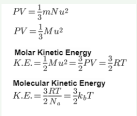
$$\bar{v} = \sqrt{\frac{8RT}{\Pi M}} = \sqrt{\frac{8RPV}{\Pi M}}$$

Root Mean Square Velocity

$$u = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3PV}{M}}$$

α:v:u = 1 : 1.128 : 1.224

Kinetic Energy



$$PV = \frac{1}{3}mNu^2$$

$$PV = \frac{1}{3}Mu^2$$

Molar Kinetic Energy

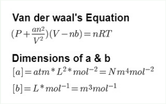
$$K.E. = \frac{1}{2}Mu^2 = \frac{3}{2}PV = \frac{3}{2}RT$$

Molecular Kinetic Energy

$$K.E. = \frac{3}{2}\frac{RT}{N_a} = \frac{3}{2}k_bT$$

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

Van der waal's Equation



Van der waal's Equation

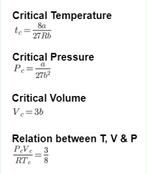
$$(P + \frac{an^2}{V^2})(V - nb) = nRT$$

Dimensions of a & b

$$[a] = atm * L^2 * mol^{-2} = Nm^4mol^{-2}$$

$$[b] = L * mol^{-1} = m^3mol^{-1}$$

Molecular Collision



Critical Temperature

$$t_c = \frac{8a}{27Rb}$$

Critical Pressure

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

Critical Volume

$$V_c = 3b$$

Relation between T, V & P

$$\frac{P_cV_c}{RT_c} = \frac{3}{8}$$

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