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CHAPTER: 4

GASES

MATTER :- Matter is anything which has mass and occupies space.

SOLIDS :-

- have definite volume and definite shape
- They are close together and fixed due to greatest interaction forces

LIQUIDS :-

- have definite volume but not definite shape
- Adopt the shape of container in which they are placed

GASES :-

- Neither have definite shape nor volume
- word "gas" is derived from Chaos because particles are in random motion

BROWN MOVEMENT :- The molecules travel in straight line until they collide with each other and with the walls of container
"found by Robert Brown (1827)"

INTIMATE MIXING :- When diffusion takes place between a liquid & a gas.

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PLASMA → • 4th State of Matter • Stars are made of plasma
• Ionized gas mixture • indefinite volume & shape
• E.g.: - Fire

→ Out of 118 elements only 11 elements are gases under normal conditions which are

- | | | | |
|------------|------------|------------|---------|
| • Hydrogen | • Oxygen | • Chlorine | • Xenon |
| • Helium | • Fluorine | • Argon | • Radon |
| • Nitrogen | • Neon | • Krypton | |
-
-

KMT G

Kinetic Molecular Theory of Gases

→ Was proposed by Swiss Mathematician Bernoulli¹ in 1738 he suggested that gas molecules are in constant motion

→ In 1857, "Rudolf Clausius" (1822-1888) postulated KMT

→ further developed by "James Maxwell" (1831-1879) in 1859.

→ and "Ludwig Boltzmann" (1844-1906) in 1870 and Van der Waal

→ This theory is sometimes also called "Billiard ball Model"

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POSTULATES OF KMT

Only Headings, Refer book for detail

- 1) MINUTE DISCRETE PARTICLES [Molecules]
- 2) STRAIGHT LINE OR BROWNIAN MOTION
- 3) PERFECTLY ELASTIC COLLISIONS
- 4) MASS AND SIZE OF MOLECULES
- 5) K.E & TEMPERATURE
- 6) PRESSURE BY COLLISIONS
- 7) LOW PRESSURE CONDITION
- 8) NO ATTRACTIVE OR REPULSIVE FORCES
- 9) NEGLIGIBLE VOLUME OF GASES

→ Kinetic Equation:- $\frac{1}{3} m N \bar{c}^2$

P = Pressure, V = Volume, m = mass of one molecule

\bar{c}^2 = mean square velocity of gas molecules

MOTION OF PARTICLES

- Translational Motion:- Straight line or linear motion of molecules & has kinetic translation energy.
- Rotational Motion:- Spins or rotates like propeller due to net angular momentum
- Vibrational Motion:- Oscillatory or back & forth motion due to kinetic vibrational energy.
- Continuous interchange between K.E & P.E

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ABSOLUTE TEMPERATURE

⇒ CHARLE'S LAW:⁶⁶ Charles's law states that the volume of a gas is directly proportional to the temperature considering pressure as constant.

OR

ACC TO BOOK:⁶⁶ It states that the volume of a given mass of a gas increases or decreases by $\frac{1}{273}$ times of original volume at 0°C and reduces to zero at -273.15°C .

⇒ IMPORTANT POINTS

- At this temperature (-273.15°C) gas is not in gaseous state but changes into liquid or solid.
- Lord Kelvin (1827-1907) used another scale called Absolute Kelvin temperature scale.
- This scale starts with -273.15°C or 0 Kelvin.
- This is the lowest temperature recorded in history.

⇒ Relation:⁻ $K = ^{\circ}\text{C} + 273$

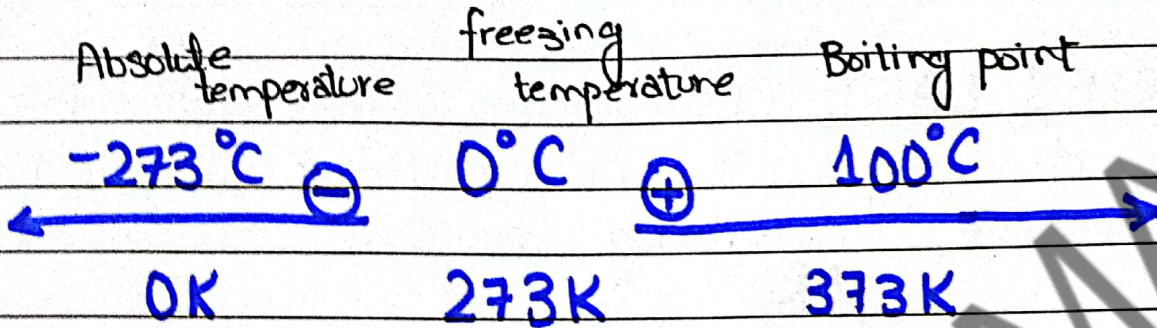
⇒ EXPLANATION

It is an important/fundamental principle in field of thermodynamics.

For every increase or decrease in temperature from 0°C , the volume of gas will change by $\frac{1}{273}$ of its original volume at 0°C .

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At Absolute temperature the volume will become zero



With every 1°C increase $\rightarrow \frac{1}{273}$ times of Volume will increase

With every 1°C decrease $\rightarrow \frac{1}{273}$ times of volume will decrease

E, Same from 2°C , 3°C & 4°C etc

RELATION BTW T & K.E

MAIN RELATION

The temperature & Kinetic Energy are both interdependent. Increasing the temperature, increases the K.E as well & vice versa

DERIVATION OF RELATION

The average K.E of molecules is redistributed with rise or fall of temperature which is explained through relation

KINETIC EQUATION OF GASES

According to Kinetic Equation of gases,

$$PV = \frac{1}{3} m N \bar{c}^2 \rightarrow \text{①}$$

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KINETIC ENERGY

The Equation for K.E is:- $K.E = \frac{1}{2} mc^2 \rightarrow \textcircled{2}$

MATHEMATICAL STEPS

Multiply the Eq $\textcircled{1}$ with $\frac{2}{3}$ so

$$PV = \frac{1}{3} m N C^2 \times \frac{2}{2}$$

$$PV = \frac{2N}{3} \left(\frac{1}{2} mc^2 \right)$$

As $K.E = \frac{1}{2} mc^2$ so replacing with eq $\textcircled{2}$

$$PV = \frac{2N}{3} (K.E)$$

AVAGADRO'S NUMBER

Considering 1 mole which contains 6.023×10^{23} or Avagadro's number of molecules so,

$N \rightarrow N_A$ so,

$$PV = \frac{2}{3} N_A (K.E)$$

GENERAL GAS EQUATION

As $PV = nRT$

As there is 1 mole so ($n=1$) so,

$$PV = (1)RT$$

$$PV = RT$$

MATHEMATICAL STEPS

Comparing R.H.S of Eq $\textcircled{4}$ & $\textcircled{5}$

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$$\frac{2}{3} N_A (K.E) = RT$$

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

As $\frac{3R}{2 N_A}$ are constants & equal to K so,

$$K.E = K T$$

Removing constant we get

$$K.E \propto T$$

CONCLUSION

• This concludes that K.E is directly proportional to temperature.

• Kinetic Energy & Temperature

The Kelvin temperature of a gas is actually the measure of average translational K.E of its molecules.

- In Solids temperature becomes a vibrational K.E
- In Gases & liquids, temperature explains the average translational energy of molecules.

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$$\text{Total } E \text{ (Vib)} = \text{K.E (Vib)} + \text{P.E (Vib)}$$

PRESSURE

DEFINITION :- "The force exerted by gas molecules per unit Area per second is called gas pressure or simply."

FORMULA :- $\text{Pressure} = \frac{\text{Force}}{\text{Area}}$

Atmospheric Pressure :- Force exerted by atmosphere on Unit Area

Barometer :- The most familiar instrument for measuring Atmospheric Pressure with the help of liquid Mercury.

Torr :- mmHg is also called Torr in honour of Evangelista Torricelli who invented Barometer in 1643.

Pascal :- A pascal (Pa) is defined as the force of one newton (N) spread over an Area of 1m^2 thus its also called (Nm^{-2}) .

UNITS

→ Commonly used units are

- Pascal
- cm Hg
- Atmosphere (Atm)
- mmHg
- torr
- Psi (pounds per square inch)

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Pascal origin:- Pascal (Pa) is named after famous mathematician & physicist **Blaise Pascal** who first proposed that atmospheric pressure **decreases** with increasing altitude.

COMMON CONVERSIONS

$$1 \text{ atm} = 760 \text{ mmHg} = 760 \text{ torr} = 76 \text{ cmHg}$$

$$1 \text{ atm} = 101325 \text{ Pa} = 101325 \text{ Nm}^{-2}$$

$$1 \text{ atm} = 14.7 \text{ Psi}$$

$$1 \text{ atm} = 101.325 \text{ kPa}$$

$$1 \text{ Joule} = 1 \text{ Nm} = 10^7 \text{ ergs} = \text{kgm}^2\text{s}^{-2}$$

$$1 \text{ Calorie} = 4.18 \text{ Joules}$$

$$1 \text{ atm} = 1.01325 \text{ bar} = 1013.25 \text{ millibar (mbar)}$$

$$1 \text{ atm} = 29.92 \text{ inches of Hg}$$

BOYLE'S LAW

BACKGROUND

In 1662, **Robert Boyle** & his assistant **Boyle Hooke** used a J shaped glass tube closed at one end to measure the volume of gas at different pressures. They kept temperature constant & observed inverse relation btw volume & pressure.

STATEMENT

The volume of a fixed amount of gas is inversely proportional to pressure at constant temperature.

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MATHEMATICAL RELATION

$$V \propto \frac{1}{P}$$

$$V = K \frac{1}{P} \quad [K \text{ is proportionality constant}]$$

$$\underline{PV = K}$$

66 MATHEMATICAL DEFINITION

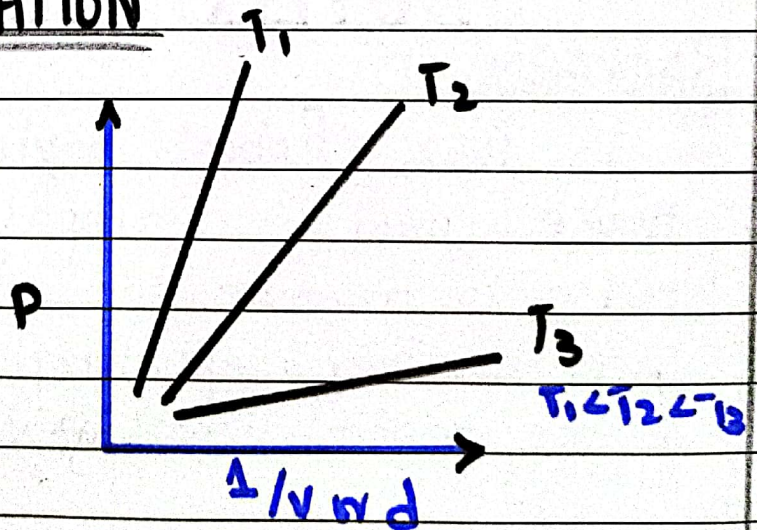
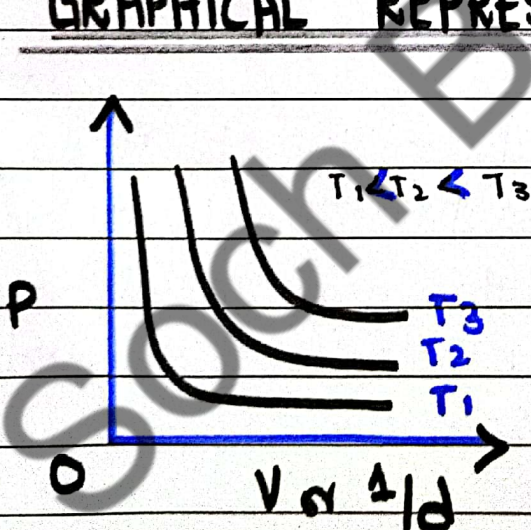
At constant temperature, the product of pressure & volume is constant for the given mass of gas.

$$P_1 V_1 = K$$

$$P_2 V_2 = K$$

$$P_1 V_1 = P_2 V_2$$

GRAPHICAL REPRESENTATION



The curve is called Hyperbola or Isotherm. [Iso → same] [therm → Temp]

The process in which temperature is constant → Isothermal process

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CHARLES LAW

BACKGROUND

In 1787, a french physicist "Jacques Charles" who was the first person to fill a balloon with Hydrogen & who made the first "Solo balloon flight" found the relation between temperature & volume.

STATEMENT

Charles law states that "Volume of fixed amount of gas is directly proportional to absolute temperature at constant pressure."

Boyle / 2nd Statement "At constant pressure, the volume of given mass of gas increases or decreases by $\frac{1}{273}$ times of its original volume at

0°C for every 1°C rise or fall in temperature respectively"

MATHEMATICAL RELATION

$$V \propto T$$

$$V = KT$$

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

If we have temperatures " T_1, T_2 " & Volume " V_1, V_2 " then

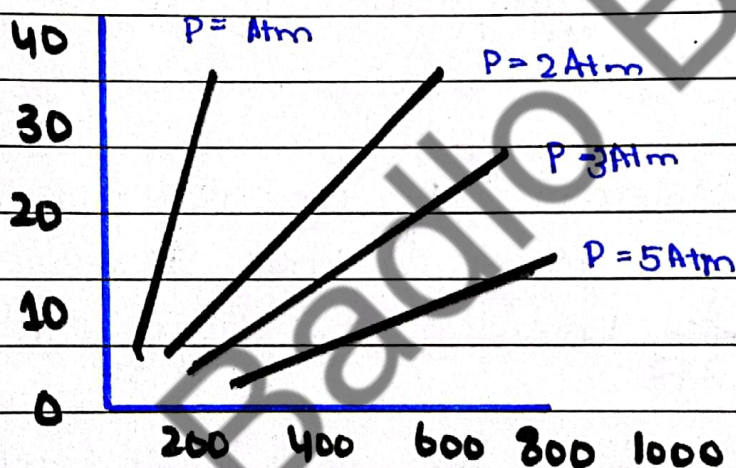
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$$\frac{V_1}{T_1} = k \rightarrow \frac{V_2}{T_2} = k$$

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

for the given mass of a gas, ratio of volume & temperature remains constant at constant pressure.

GRAPHICAL REPRESENTATION



The straight line drawn between temperature and volume shows **direct relation** keeping **pressure constant**. Each line is called **isobar**.

CRITICAL LAW DERIVATION

Suppose volume of gas V_0 at $0^\circ\text{C} = V_0$

$$\text{Volume at } 1^\circ\text{C} = V_0 + \frac{V_0}{273}$$

$$\text{Volume at } 2^\circ\text{C} = V_0 + V_0 \frac{2}{273}$$

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$$\text{Volume at } t^{\circ}\text{C} = V_0 + V_0 \frac{t}{273}$$

$$V_t = V_0 \left[1 + \frac{t}{273} \right]$$

$$= V_0 \left[\frac{273+t}{273} \right] \quad \rightarrow \text{(taking LCM)}$$

$$\text{As } t^{\circ} + 273 = T \quad [\text{Absolute temperature}]$$

$$= V_0 \left[\frac{T}{273} \right] \quad \rightarrow \text{(rearranging)}$$

$$V_t = \frac{V_0 T}{273}$$

$$\text{RESULT :- } V_t = \frac{V_0 T}{273}$$

Thus this indicates the volume of a definite quantity of a gas at constant pressure is directly proportional to Temperature

ABSOLUTE ZERO

DEFINITION

“The temperature at which the molecular motion ceases (stops) and a substance would have zero kinetic energy.” OR

“The hypothetical temperature at which the given volume of a gas reduces to Absolute zero”

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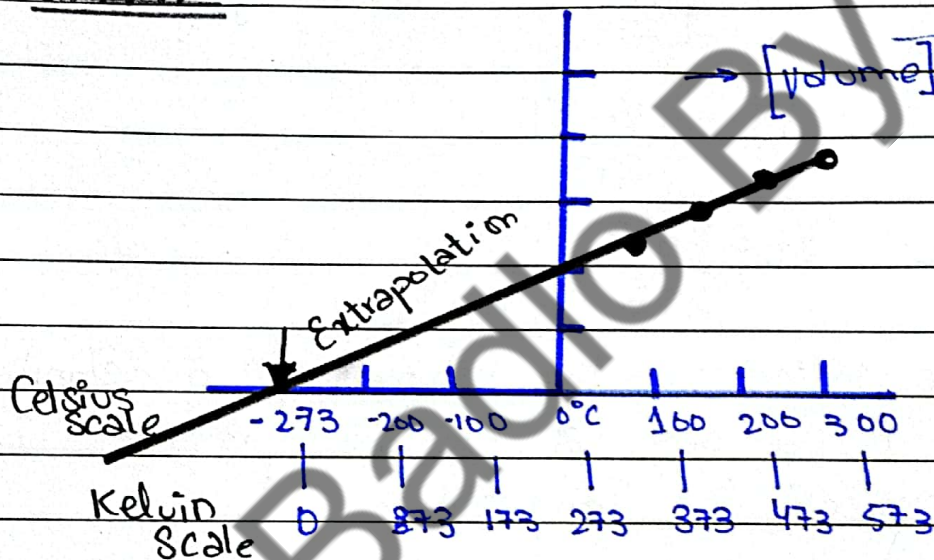
MATHEMATICALLY

$$\text{Volume at } -1^\circ\text{C} = V_0 - \frac{1}{273} V_0$$

$$\text{Volume at } -273^\circ\text{C} = V_0 - \frac{273}{273} V_0 = 0$$

thus at -273°C , the volume of a gas reduces to zero, is called Absolute zero

GRAPH



KELVIN SCALE

A new scale called Kelvin scale or Absolute scale was invented which starts from -273.15°C or Zero Kelvin

0 Kelvin
 -273°Celsius

273 Kelvin
 0°Celsius

373 Kelvin
 100°Celsius

→ -273°C or 0K is very hard to obtain & most gases liquify or solidify before they reach this temperature.

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AVAGADRO'S LAW

BACKGROUND

The Italian scientist "Amedeo Avogadro" (1766-1856), a professor of higher physics at the University of Turin for many years, observed Amount of moles - Volume relationship at constant temperature & pressure.

EXPLANATION

This shows that by taking equal volume of 2 gases at a fixed temperature & gas contain equal number of molecules.

STATEMENT

The law can be stated in ^{three} ways but the two ways are:-

Avogadro's law states that, "Equal volumes of all gases at the same temperature and pressure must contain equal number of molecules"

OR

"The volume of gas is directly proportional to number of moles at constant pressure and temperature"

MATHEMATICALLY

$$V \propto n \quad \text{or} \quad V = Kn \quad \text{or} \quad V = k \times n$$

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IMPORTANCE / AVADRO'S NO

The exact number of (atoms, molecules) has been calculated in one mole of a substance with the help of Spectrometer and may be called as Avagadro's Number, whose numerical value is 6.022×10^{23} particles per mole. It was confirmed by Joseph Loschmidt

IDEAL GAS EQUATION

NON-IDEAL / ACTUAL GAS:-

The gases who donot obey gas laws strictly are known as non-ideal or real gas

IDEAL GAS:- which obey gas laws strictly at all conditions of temperatures & pressures are called ideal or perfect gas.

BACKGROUND:-

The equation that shows the simultaneous changes in pressure and temperature on the volume of given amount of gas is called ideal gas equation

DERIVATION

The ideal gas Equation is a combination of 3 laws
Boyle's law Charles's law Avagadro's law

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According to Boyle's law,

$$V \propto \frac{1}{P} \quad (\text{at constant } T \text{ and } n) \rightarrow (1)$$

According to Charles's Law,

$$V \propto T \quad (\text{at constant } P \text{ \& } n) \rightarrow (2)$$

According to Avogadro's law,

$$V \propto n \quad (\text{at constant } P \text{ \& } T) \rightarrow (3)$$

Combining eq 1, 2, 3 we get

$$V \propto \frac{1}{P} n T, \quad V \propto \frac{n T}{P}, \quad V = \frac{R n T}{P}$$

$PV = RnT$, where R is general gas equation
if $n=1$ then,

$$PV = RT$$

FINAL VALUE :- $PV = RT$

FOR TWO CONDITIONS:-

$$\frac{P_1 V_1}{T_1} = R$$

$$\frac{P_2 V_2}{T_2} = R$$

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

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UNIT TABLE

V	P	n	T
cm ³	mmHg or torr	mole	K

V	P	n	T
m ³	Nm ⁻²	mole	K

V	P	n	T
dm ³	atm	mole	K

VALUES OF 'R'

* $R = 0.0821 \text{ dm}^3 \text{ atm mole}^{-1} \text{ K}^{-1}$

* $R = 62400 \text{ cm}^3 \text{ torr mole}^{-1} \text{ K}^{-1}$

* $R = 8.3143 \text{ Nm mole}^{-1} \text{ K}^{-1}$ [In SI units]

FORMULAS

⇒ FOR MOLECULAR MASS OF GAS - $M = \frac{WRT}{PV}$

⇒ FOR DENSITY OF GAS - $d = \frac{MP}{RT}$

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RELATION B/W TEMPERATURE AND K.E

RELATION

The temperature of a gas depends on Kinetic Energy of molecules. The increase in temperature increases the average Kinetic Energy & vice versa.

DERIVATION

The average Kinetic Energy of molecules is redistributed with rise and fall of temperature.

Kinetic Equation

It can be explained with the help of Kinetic Equation which states that

$$PV = \frac{1}{3} mNc^2 \quad \text{---(i)}$$

Kinetic Energy

Kinetic Energy is stated as :-

$$\frac{1}{2} mv^2, \text{ for gas molecules in this case}$$
$$K.E = \frac{1}{2} mc^2 \quad \text{---(ii)}$$

Mathematical Steps

Multiplying Equation (i) by $\frac{2}{2}$ so,

$$PV = \frac{1}{3} mNc^2 \times \frac{2}{2}$$

$$= \frac{2}{6} mNc^2$$

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$$PV = \frac{2}{3} N \left(\frac{1}{2} mc^2 \right)$$

As we know $K.E = \frac{1}{2} mc^2$ so,

$$PV = \frac{2}{3} N (K.E)$$

Avagadro's Number

As we know 1 mole of gas contains 6.023×10^{23} particles which is called Avagadro's Number represented by "N_A" so.

$$PV = \frac{2}{3} N_A (K.E) \quad \text{--- (iii)}$$

Ideal gas Equation

$$PV = nRT$$

for 1 mole of gas $n=1$ hence,

$$PV = RT \quad \text{--- (iv)}$$

Mathematical Steps

Comparing Equation (iii) & (iv),

$$\frac{2}{3} N_A (K.E) = RT$$

$$2 N_A (K.E) = RT \cdot 3$$

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

$$K.E = K T$$

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where $k = \frac{3R}{2N_A}$ is the constant value

So,

$K.E \propto T$

RESULT

$K.E \propto T$

Kinetic Energy is directly proportional to temperature.
Increasing K.E will increase Temperature & vice versa.

DEVIATION FROM IDEAL GAS BEHAVIOUR

IDEAL GAS :- Gas which obeys $PV = nRT$ and all the gas laws. No such gas is known.

NON-IDEAL GAS :- Gases which do not obey gas laws strictly are called non-ideal and actual gases.

FAULTY ASSUMPTIONS

→ Van der Waals attributed deviation of real gases from ideal behaviour due to faulty assumptions in KMT

→ No attractive or repulsive forces between gas molecules

→ Volume of gas is negligible compared to container

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AT LOW TEMPERATURE

At high temperature, kinetic energy of molecules is high thus attractive forces are negligible. When temperature is decreased, kinetic energy of molecules decreases thus intermolecular forces become more significant as molecules come close.

At a certain very low temperature, the gas changes into liquid thus attractive forces become significant near ~~liquidifying~~ liquefying temperature.

DEPENDENCE OF DEVIATION

Stronger the intermolecular forces, more polar will deviate more from ideal gas.

Weaker the intermolecular forces, less polar it is & will deviate less.

In case of same molecular forces, molecular weight will be considered. One with more will have more volume & will deviate more.

HIGH PRESSURE

At low pressure, volume is negligible whereas at high pressure, volume is not negligible due to intermolecular forces.

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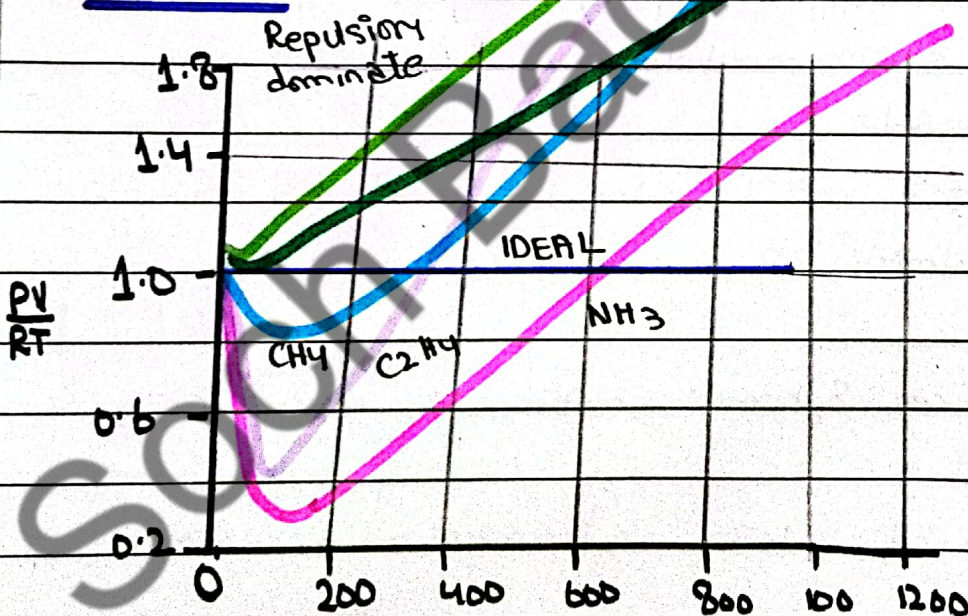
GRAPHICAL REPRESENTATION

→ IDEAL GAS:- for 1 mole of gas, graph between $\frac{PV}{RT}$ and pressure must be straight [temperature remains at 0°]

→ NON-IDEAL GAS:- Due to increase in pressure, volume will decrease more than its predicted due to intermolecular forces. This decreases value of $\frac{PV}{RT}$

→ COMPRESSIBILITY FACTOR:- $\frac{PV}{RT}$ is called compressibility factor which is 1 for ideal gas but not 1 for other gases

→ DIAGRAMS:-



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VANDER WAAI'S EQUATION

BACKGROUND

The Dutch physicist Johannes Van der Waals, in 1873, at the university of Amsterdam modified the general gas equation in order to calculate the volume and pressure of a real gas. He recognized that the ideal-gas equation can be corrected by performing the **Volume correction** and **pressure correction**. He got a NOBEL PRIZE in 1910 for his work.

VOLUME CORRECTION

MAIN POINTS

When a gas is compressed to a point where it is not possible to compress it anymore, At this stage they produce force of repulsion because gas molecules have definite volume which is very small but not negligible.

- Effective volume of gas molecules per mole = b
- Available free space or volume to gas molecules = Volume of vessel minus molecules

DERIVATION

$$V = V_{(\text{vessel})} - V_{(\text{molecules})}$$

$V \rightarrow$ free volume

$V_{\text{molecules}} \rightarrow nb$

$b \rightarrow$ individual volume of one mole of the gas.

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$n \rightarrow$ Total no of moles

" b is also the effective or excluded volume of gas"

EFFECTIVE VOLUME

The volume occupied by one mole of a gas at highly compressed state is called effective volume. It is four times of the actual volume of the gas molecules

$$"b = 4V_m"$$

Where $V_m \rightarrow$ Actual volume of one mole of gas

$$\text{So, } V = V_{\text{vessel}} - nb \rightarrow (1)$$

PRESSURE CORRECTION

MAIN POINTS

A molecule about to strike the walls of container will experience an attractive force from other molecules & thus pressure with which it should have been hitting is decreased

$$\text{therefore, } P_{\text{observed}} = P_{\text{ideal}} - P_{\text{lessened}}$$

DERIVATION

$$\text{ideal pressure} = P_i$$

$$\text{Pressure lessened} = P^s$$

$$P_{\text{observed}} = P_i - P^s$$

$$P_i = P_{\text{observed}} + P^s$$

$$\text{Ideal Pressure} = \text{Pressure observed} + \text{Pressure lessened}$$

P^0 [lessened] depends on forces of attraction between molecules

The net force of attraction is proportional to concentrations of type A & type B molecules

$$P^0 \propto C_A \cdot C_B$$

As $C = \frac{\text{mole}}{\text{dm}^3} = \frac{n}{V}$ so,

$$P^0 \propto \frac{n}{V} \cdot \frac{n}{V}$$

$$P^0 \propto \frac{n^2}{V^2}$$

$$P^0 = \frac{a n^2}{V^2} \quad \text{so} \quad P_i = P + \frac{a n^2}{V^2} \quad \rightarrow (2)$$

'a' = co-efficient of attraction per unit volume

$$PV = nRT$$

[putting eq (1) & (2)]

$$\left(P + \frac{a n^2}{V^2} \right) (V - nb) = nRT \quad , \quad n=1$$

$$\left(P + \frac{a}{V^2} \right) (V - b) = RT$$

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DALTON'S LAW OF PARTIAL PRESSURE

BACKGROUND

The English scientist John Dalton (1766-1844) in 1803, made an important contribution to the study of gaseous mixtures.

STATEMENT

Dalton's law of partial pressures states that the "total pressure" exerted by a mixture of non-reacting gases is equal to the sum of the partial pressures of all the gases present in the mixture.

EXAMPLE

Suppose in a system three gases A, B, C are present. The partial pressures of these gases are:-

- P_A :- Partial pressure of gas A
- P_B :- Partial pressure of gas B
- P_C :- Partial pressure of gas C

then Dalton's law can be stated as mathematically:-

$$P_T = P_A + P_B + P_C + \dots$$

Where P_T is total pressure of mixture of gases.

AIR

The air we breathe is a mixture of gases, containing 78.084% nitrogen, 20.948% oxygen, 0.934% argon, 0.035% CO_2 & other gases. Total pressure is sum of partial pressure of each gas.

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RELATION BTW PRESSURE & MOLES

If n_A , n_B and n_C are the numbers of moles of gases and P_A , P_B and P_C are their partial pressures respectively,

According to $PV = nRT$

$$\text{For gas A, } P_A = \frac{n_A RT}{V} \rightarrow (1)$$

$$\text{For gas B, } P_B = \frac{n_B RT}{V} \rightarrow (2)$$

$$\text{For gas C, } P_C = \frac{n_C RT}{V} \rightarrow (3)$$

$$P_t = \frac{n_t RT}{V} \rightarrow (4)$$

n_t (total no of moles of all gases) = $n_A + n_B + n_C$

Dividing eq (1) by (4)

$$\frac{P_A}{P_t} = \frac{\frac{n_A RT}{V}}{\frac{n_t RT}{V}}$$

$$\frac{P_A}{P_t} = \frac{n_A RT}{V} \times \frac{V}{n_t RT}$$

$$\frac{P_A}{P_t} = \frac{n_A}{n_t}, \quad P_A = \frac{n_A}{n_t} \times P_t$$

Similarly $P_B \propto n_B$ & $P_C \propto n_C$

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Partial Pressures of any gas = $\frac{\text{No of moles of gas}}{\text{Total no of moles}} \times \text{Total pressure}$

RELATION BTW PRESSURE AND MOLE FRACTION

Partial Pressure of any gas = $\frac{\text{No of moles of that gas}}{\text{Total no of moles}} \times \text{Total pressure of all gases}$

$$\frac{P_i}{P_t} = \frac{n_i}{n_t} \quad \text{or} \quad P_i = \frac{n_i}{n_t} \times P_t$$

$\frac{n_i}{n_t} = \chi_i$, χ_i is called mole fraction

$$P_i = P_t \chi_i$$

Partial pressure of any component is equal to the product of total pressure of all the components & mole fraction of that component.

APPLICATIONS

COLLECTION OF GASES OVER WATER

The sum of water vapours and insoluble gas collected over water causes total pressure on the water. Pressure due to vapours is called aqueous tension.

$$P_{\text{total}} = P_{\text{gas}} + P_{\text{water vap}}$$

$$\text{OR } P_{\text{gas}} = P_{\text{total}} - P_{\text{water vap}}$$

RESPIRATION AT HIGHER ALTITUDES

→ NORMAL CONDITIONS:- The partial pressure of O_2 in the outside air is higher (159 mm Hg) than in the lungs where partial pressure is lower (116 mm Hg). It is easy to respire because O_2 moves from higher to lower pressure easily.

→ AT ALTITUDES:- At altitudes, the partial pressure of O_2 is about 150 mm Hg that's why pilots may have uncomfortable breathing in a non-pressurized cabin.

RESPIRATION BY DEEP SEA DIVERS

→ At a depth of 40 metres, the pressure increases 5 times than the normal pressure, thus it becomes 795 mm Hg.

→ Therefore they use mixture of 96% He and

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4% O_2 [Heliox] in the scuba

→ Diver needs to come slowly on surface because compressed air in lungs expands on sudden arriving at surface

RETURN FROM SEA

→ If Diver takes N_2 in tank, the blood of diver gets saturated with N_2 and quickly arriving on land can cause life threatening condition Bends

→ They use different mixtures of gases or spent hours in decompression chamber

GRAHAM'S LAW

BACKGROUND

English Scientist, Thomas Graham in 1831 (1805 - 1869) measured rates at which equal volumes of various gases escaped from an apparatus at the same rate & pressure.

STATEMENT

"Rate of diffusion or effusion of a gas is inversely proportional to the square root of its density & molar weight."

DERIVATION

$$\sigma \propto \frac{1}{\sqrt{d}}$$

$$\sigma = k \times \frac{1}{\sqrt{d}}$$

$$\sigma = \frac{k}{\sqrt{d}}$$

$$\sigma \propto \frac{1}{\sqrt{M}}$$

$$\sigma = k \times \frac{1}{\sqrt{M}}$$

$$\sigma = \frac{k}{\sqrt{M}}$$

$$\sigma_1 = \frac{k}{\sqrt{d_1}}$$

$$\sigma_2 = \frac{k}{\sqrt{d_2}}$$

$$\frac{\sigma_1}{\sigma_2} = \sqrt{\frac{d_2}{d_1}}$$

$$\frac{\sigma_1}{\sigma_2} = \sqrt{\frac{M_2}{M_1}}$$

$$\frac{\sigma_1}{\sigma_2} = \sqrt{\frac{d_2}{d_1}} = \sqrt{\frac{M_2}{M_1}}$$

Date _____

THE FOURTH STATE OF MATTER

DEFINITION

"The fourth very high-temperature physical state of matter which has ionized mixture of ions, electrons and neutral atoms, is called plasma."

BACKGROUND

- It was first identified in a Crookes tube by British English Chemist & physicist, William Crookes in 1879 & he called it "radiant matter".
- Word PLASMA was given by "Dr. Irving Langmuir".

PROPERTIES

- It is known since Big Bang & occupies 99% of Universe
- Overall Neutral → equal no. of electrons & positive ions
- Good conductors & affected by magnetic field
- Neither definite shape nor volume
- Easily compressible
- Free electrons in metals can be considered plasma
- Show characteristic glow

APPLICATIONS

- Light up our offices & homes, make our computers, & electronic equipment.
- Drive lasers, help clean environment, make tools corrosion-resistant
- found in fluorescent lamps, used for cleaning & sterilization of food.

FORMATION: Electrons leave the atom due to high energy in gas, causing to convert into plasma.