CHAPTER # 4
STATE OF MATTER
GASES

Q1. How many states of matter exist in nature?

Ans: Matter consists of three fundamental states, the gases, liquid and solid. Besides these three states, the fourth one is plasma state (a partially ionised gas at very high temperature).

Q2. Define and explain the term Brownian movement.

Ans: The word gas is derived from **chaos.** This is because the gaseous particles have chaotic (random) motion. The molecules travel in a straight line until they collide with each other or with the walls of the container. Such motion of the molecules is called Brownian Movement (Robert Brown 1827).

Q3. How can you explain the effect of pressure on gas molecules?

Ans: In a gas the average distance between the molecules is quite large as compared to their diameters. This shows that the intermolecular forces among the molecules are negligible at the lowest pressure. However these forces increase with the increase of pressure. For example the forces are negligible at one atmosphere (lowest) significant at about 10 atmospheres and very important at higher pressures (about 30 atmospheres). It is found by experiment that when the pressure is sufficiently low, all gases satisfy the perfect gas law $PV = nRT$

Q4. Earth's atmosphere is a mixture of gases how can you explain it by giving daily life examples?

Ans: The earth's atmosphere is a mixture of gases. Oxygen (21%) and Nitrogen (78%) are its two major components, both support life. In fact, the earth's atmosphere is an immense system of gases where many chemical reactions occur. For example photo-synthesis, combustion etc. Natural gas, the major component of which is methane, is used to heat homes in winter, to cook food and to drive vehicles etc.

Kinetic Molecular Theory of Gases:

The behaviour and properties of gases can be theoretically explained using the Kinetic Molecular Theory. This theory is a model that explains the behaviour of gases using generalisation about random moving molecules within a gas.

The Kinetic Molecular Theory was first postulated by **Daniel Bernoulli**, a Swiss mathematician. The Kinetic Theory of gases was developed by a number of well-known physicist such as **James Maxwell** (1859) and **Boltzmann** in (1870). The research was contained on the Kinetic Theory making it a cornerstone of our present view of molecular substances. In 1857, **Clausius** derived the Kinetic equation.

Physical theories are often given in terms of postulates, the basic statements from which all conclusions or prediction of theory is deduced.

Q5. Write and explain kinetic equation?

Ans: Kinetic equation:

R J Clausius deduced an equation for the pressure of an ideal gas from the postulates of Kinetic theory. This equation is called Kinetic equation.

Mathematically, $pv = \frac{1}{3} mnc^2$

Where
- p = Pressure
- v = Volume
- m = mass of one molecule of the gas
- n = number of moles of the gas
- $c^{\bar{2}}$ = mean square velocity of gas molecules

Mean square velocity:

Since the molecules of a gas under the given conditions do not have the same velocity, so mean square velocity is considered. If n_1 molecules have velocity c_1, n_2 molecules have velocity etc, then

$$c^{\bar{2}} = \frac{n_1 c_1^2 + n_1 c_2^2 + \cdots}{n_1 + n_2 + \cdots}$$

Where $c^{\bar{2}}$ is the average of all the possible velocities.

Root mean square velocity:

The square root of $c^{\bar{2}}$ is called root mean square velocity.

Q6. How many types of motion occur between gas molecule also elaborate these motions?

Ans: Motion of Particles of a Gas:

In gases, the molecular motion is of three types:

I. Translational Motion.
II. Rotational Motion
III. Vibrational Motion.

A mono atomic molecule (e.g. He) will show translational motion while a diatomic (H_2) and polyatomic molecules (CO_2, NH_3 etc.) will undergo, in addition to translational motion, the rotational and vibrational motions too.

I. Translational Motion:

The motion imparted to the gaseous molecules due to their motion in all possible directions is called translational motion and the energy as kinetic translation energy. In this case the entire molecules move from place to place.

II. Rotational Motion:

The motion imparted to the gaseous molecules as a result of net angular momentum about their centre of gravity is called rotational motion and the energy as kinetic rotational energy. In this case the molecule spins like a propeller.

III. Vibrational Motion:

The motion imparted to the gaseous molecules due to the oscillations is called vibrational motion and the energy as kinetic vibrational energy. In this case the molecules vibrate back and forth about the same fixed location.

In such an oscillating system, there is a continuous interchange between vibrational kinetic energies and potential energies. As a result both K.E and P.E is possessed by the gaseous molecules. Thus,

Total K.E = K.E (Vib.) + P.E (Vib.)

The vibrational motions in polyatomic molecules (like CO_2, NH_3 etc) are quite complicated.

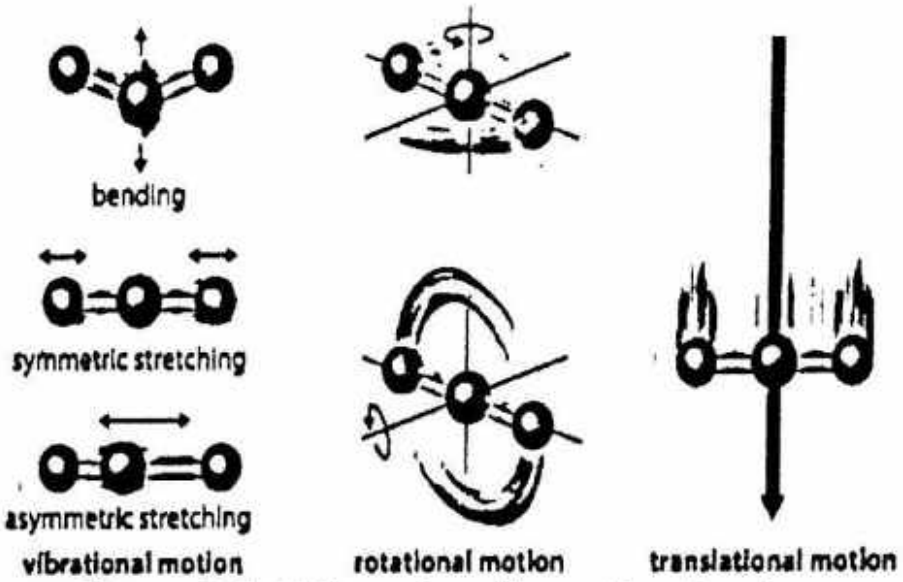


Figure 4.4 Different motions of molecules

Example 1:

Convert the following Celsius temperatures to Kelvin temperatures:

(a) -132°C **(b)** 96°C **(c)** 0°C **(d)** -12°C **(e)** 148°C

Solution:

Using the formula: K = °C + 273

(a) TK = -132 + 273 = 141 TK (b) TK = 96°C + 273 = 369 TK

(c) TK = 0°C + 273 = 273 TK (d) TK = -12°C + 273 = 261 TK

(e) TK = 148°C + 273 = 421 TK

Example 2:

Convert the following Kelvin temperatures to Celsius degree centigrade temperatures.

(a) 340 TK **(b)** 200 TK **(c)** 10TK **(d)** 405 TK

Solution:

Using the formula: t °C = TK - 273

(a) t °C = 340 TK - 273 = 67°C (b) t °C = 200 TK - 273 = -73°C

(c) t °C = 10 TK - 273 = - 263°C (d) t °C = 405 TK - 273 = 132°C

Q7. What is the total Kinetic Energy of translation in ergs of two moles of perfect gas at 27°C in calories?

Ans: Solution:

No of moles of perfect gas (n) = 2

$T = 27°C + 273 = 300K$

$R = 8.314 \times 10^7$ ergs $mole^{-1} K^{-1}$

Now, K.E of translation $(K.E) = \frac{3}{2} nRT$

or K.E = 3/2 × 2 moles $dm^{-3} \times 8.314 \times 10^7$ ergs $mole^{-1}K^{-1} \times 300K$

K.E = $3/2 \times 2 \times 8.314 \times 10^7 \times 300$ moles $dm^{-3} \times$ ergs $mole^{-1} \times K^{-1} \times K$

= 7483.5×10^7 ergs dm^3

Q8. Briefly explain gas pressure in the Light of Kinetic Molecular Theory.

Ans: Gas Pressure in the Light of Kinetic Molecular Theory:

According to the Kinetic molecular theory, gas molecules are in constant random motion. They move in straight line until they collide with another molecule or the walls of the container. The pressure of a gas in a container is due to the force exacted by its molecules on the walls of the container. The average kinetic energy of gas molecules is directly proportional to the Kelvin temperature. So the average kinetic energy of a collision when a gas molecule collides with the walls of a container will not charge at constant temperature.

Q9. Define gaseous pressure and its various units.

Ans: Pressure:

A force being exerted over a unit area is called Pressure.

So, $P = \frac{F}{A}$ where P is measured in Newton Force in joules and the area in m^2.

Pascal (Pa):

A Pascal (Pa) is defined as the force of one Newton (N) spread over an area of $1m^2$.

Units of Pressure:

i. SI Units:

SI unit for pressure is Pascal (Pa), in S.I. system, unit of force is Newton (N) and the unit of area is metre square (m^2). Thus Pascal is one Nm^{-2}.

1 atm = 101325Pa =101325 Nm^{-2}
1 atm = 14.7 Psi (Pounds per square inch)
1 atm = 101325Pa
1 atm = 101.325 KPa
1atm = 760 torr = 760mm of Hg
1 J = 1Nm =10^7 ergs = Kgm^2s^{-2}
1 Cal = 4.18J

Interconversion of Pressure in Pascals, Kilo Pascals, atmosphere and bar.

Example 4:

Convert 10 Pascals into (i) Kilo Pascals **(ii)** atmosphere **(iii)** bars

Solution:

(i) One Kilo Pascal = 1000 Pascals
or 1000 Pascals = one Kilo Pascal
10 Pa $= \frac{1}{1000} \times 10 = 0.01$ Kilo Pascal

(ii) One atmosphere = 101325 Pa
or 101325 Pa = one atm
10 Pa $= \frac{1}{101325} \times 10 = 0.000098$ atm

Activity for Students

Convert 35 bars into (i) Pascal (ii) Kilo Pascal (iii) Atmosphere

Solution:

(i) $1.01325\ bar = 101325\ pa$

$$1\ bar = \frac{101325}{1.01325}\ pa$$

$$35\ bar = \frac{101325}{1.01325} \times 35 = 35 \times 10^6\ pa$$

(ii) $1.01325\ bar = 101.325\ kpa$

$$1\ bar = \frac{101.325}{1.01325}\ kpa$$

$$35\ bar = \frac{101.325}{1.01325} \times 35 = 3.5 \times 10^3\ kpa$$

(iii) $1.01325\ bar = 1\ atm$

$$1\ bar = \frac{1}{1.01325}\ atm$$

$$35\ bar = \frac{1}{1.01325} \times 35 = 34.54\ atm$$

Q10. Briefly explain Boyle's Law. And why the Isotherm goes away from both the axis in the graphical explanation of Boyle's Law.

Ans: Boyle's Law:

It states that the pressure of a fixed amount of a gas varies inversely with its volume, when the temperature is maintained constant.

Mathematical form of Boyle's Law:

Mathematically, $P \propto \frac{1}{V}$ or $P = K \times \frac{1}{V}$ where K is proportionality constant The value of K is different for different amounts of the same gas.

Again $P = K \times \frac{1}{V}$ or $PV = K$, thus the product of volume and Pressure remains constant provided the temperature is kept constant.

$P_1 V_1 = K$ (For gas at pressure P_1)

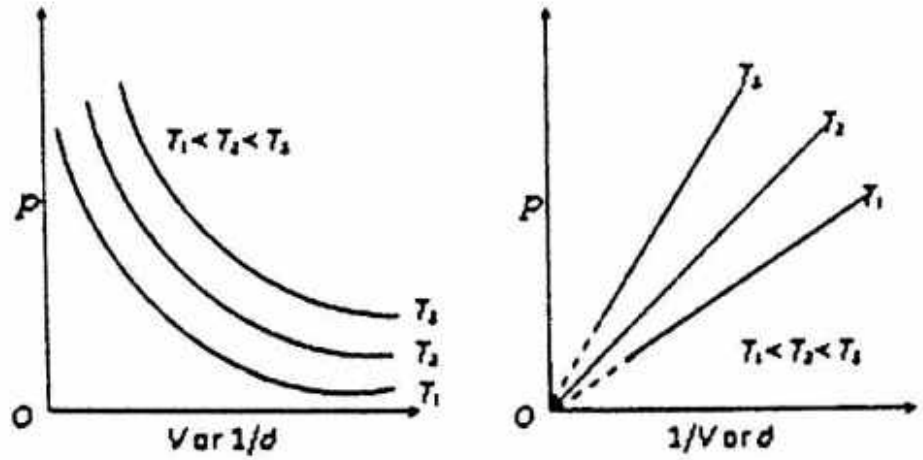
$P_2 V_2 = K$ (For gas at pressure P_2)

Therefore $P_1 V_1 = P_2 V_2$ where V_1 and P_1 are initial volume and Pressure while $P_2 V_2$ are their final Pressure and volume.

Graphical Representation of Boyle's Law:

Isotherm:

When the Pressure of a gas is plotted against volume at different temperatures, we get a family of curves as shown in the figure. Each curve is a hyperbola with different values of K. Each curve is known as **Isotherm** (constant temperature plot). As the temperature is increased, the Isotherm goes away from both the axis. This is because at **higher temperature, the volume** of the gas is increased.



Graphical representation of different forms of Boyles Law

Q11. Example 5: An ideal gas occupies a volume of 0.300 dm^3 at a pressure of 1.80×10^5 Pa. What is the volume of the gas maintained at the same temperature if the pressure is reduced to 1.15×10^5 Pa.

Solution:

$V_1 = 0.300\ dm^3$ $\quad V_2 = ?$

$P_1 = 1.80 \times 10^5\ Pa$ $\quad P_2 = 1.15 \times 10^5\ Pa$

According to the Boyle's Law

$$V_1 P_1 = V_2 P_2$$

or $$V_2 = \frac{V_1 P_1}{P_2}$$

$$= \frac{0.300\ dm^3 \times 1.80 \times 10^5\ Pa}{1.15 \times 10^5\ Pa}$$

$$= \frac{0.300 \times 1.80}{1.15}\ dm^3 = 0.469\ dm^3$$

Q12. Example 6: At sea level, where the pressure was 775 torr, the gas in a balloon occupied $2m^3$. What volume will the balloon expand when it has risen to an altitude where the pressure is

(a) 100 torr **(b)** 10 torr temperature remains constant

Solution:

(a) $P_1 = 755\ torr$ $\quad P_2 = 100\ torr$

$V_1 = 2m^3$ $\quad V_2 = ?$

According to the Boyle's Law

$$V_1 P_1 = V_2 P_2$$

$$2m^3 \times 755\ torr = V_2 \times 100\ torr$$

or $$V_2 = \frac{2m^3 \times 755\ torr}{100\ torr}$$

$$V_2 = \frac{2 \times 755}{100}\ m^3 = 15.10\ m^3$$

(b) $V_1 = 2m^3$ $V_2 = ?$

$P_1 = 755\ torr$ $P_2 = 10\ torr$

According to the Boyle's law

$V_1 P_1 = V_2 P_2$

$2m^3 \times 755\ torr = V_2 \times 10\ torr$

or $V2 = \frac{2m^3 \times 755\ torr}{10\ torr}$

$$= \frac{2 \times 755}{10}\ m^3 = 151 m^3$$

Activity For Student

375cm³ of a given gas has a pressure of 770mm of Hg. What will be the volume if the pressure is reduced to 750mm of Hg? (Ans: 385 cm³)

Solution:

$P_1 = 770\ mm\ of\ Hg$

$P_2 = 750\ mm\ of\ Hg$

$V_1 = 375\ cm^3$

$V_2 = ?$

According to the boyle's law

$P_1V_1 = P_2V_2$

$375 \times 770\ mm\ of\ Hg = V_2 \times 750\ mm\ of\ Hg$

$or\quad V_2 = \frac{375\ cm^3 \times 770\ mm\ of\ Hg}{750\ mm\ of\ Hg}$

$V_2 = \frac{375 \times 770}{750}\ cm^3 = 385\ cm^3$

Q13. Briefly explain Charle's Law.

Ans: Charle's Law:

It states that at constant pressure, the volume of a given mass of a gas varies linearly with the absolute temperature of the gas.

Critical definition:

Charles's law can also be stated as "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."

From his preliminary investigations, Charles's observed that the gases such as H_2, CO_2, O_2 etc. expanded equally on heating from **0** to **80°C** at constant pressure.

Derivation of critical form of Law:

Suppose the volume of a gas at 0°C = V_o

Volume at 1 °C = $V_o + V_o \frac{1}{273}$

Volume at 2 °C $= V_o + V_o \frac{2}{273}$

Volume at t °C $= V_o + V_o \frac{t}{273}$

$$V_t = V_o[1 + \frac{t}{273}] = V_o[\frac{273 + t}{273}]$$

$$= V_o[\frac{T}{273}] \quad \text{Where } t\,°C + 273 = T \text{ (Absolute temperature)}$$

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

Another definition:

Thus the Law states that the volume of a definite quantity of a gas at constant pressure is directly proportional to the absolute temperature.

Again $V \propto T$ or $V = KT$, where T is the temperature on absolute scale.

Thus the Law may also be defined as, "the ratio between volume and temperature is always a constant quantity."

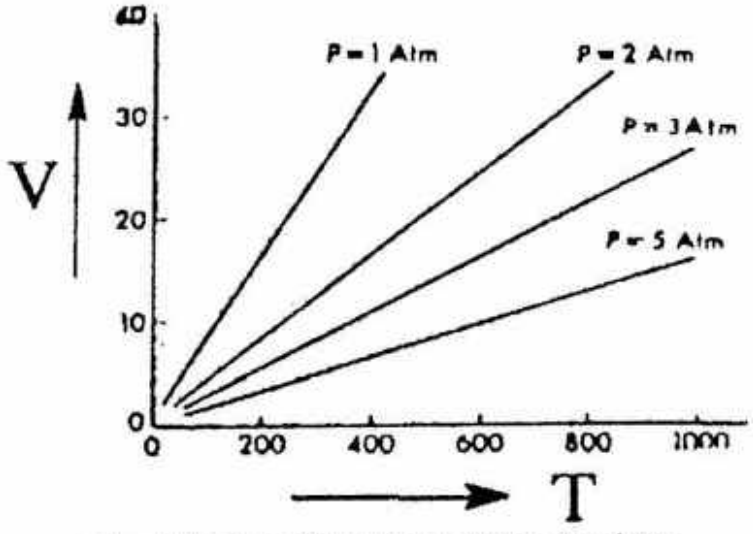
So $\frac{V_1}{T_1} = K$ and $\frac{V_2}{T_2} = K$

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

Where V_1 and T_1 are initial volume and absolute temperature. V_2 T_2 are final volume and absolute temperature.

Graphical Representation of Charles's Law:

The equation for Charles's Law gives a straight line proved by plotting a graph between volume and temperature. Different straight lines are obtained with different pressures, which are a constant quantity. Each constant pressure line is called an "Isobar".



Graphical verification of Charles law

Q14. How centigrade or Celsius scale can be converted to the absolute scale or Kelvin scale?

Ans: The temperature on degree centigrade or Celsius scale can be converted to the absolute scale or Kelvin scale (after the name lord Kelvin 1824-1907) by adding 273.15.

Thus $K = t\,°C + 273.15$

Now $V = KT$ or $\frac{V}{T} = K$

Q15. What is absolute zero? Explain it graphically also write its significance.

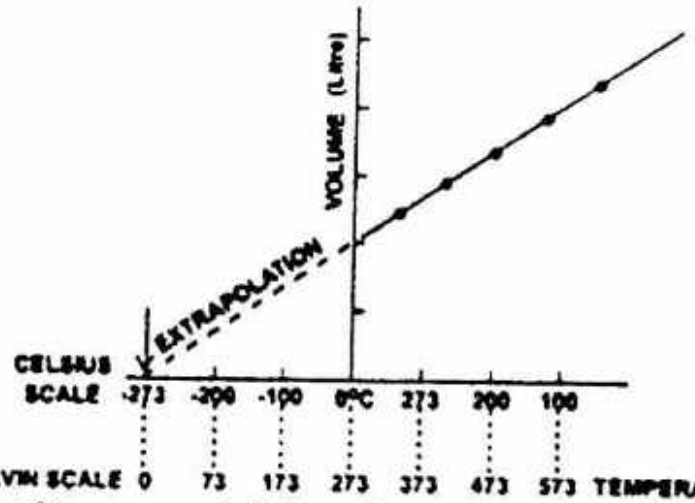
Ans: Absolute Zero:

According to critical definition of Charles's Law, at constant pressure, the volume of given mass of a gas increases or decreases by $\frac{1}{273}$ of its original volume at 0°C by 1°C rise or fall of temperature.

Let the volume of an ideal gas at 0°C by $V_o\ cm^3$

Volume at $-1°C = V_o - \frac{1}{273} V_o$

Volume at $-273°C = V_o - \frac{273}{273} V_o = 0$



Effect of decrease of temp on a definite volume of a gas at constant pressure

Thus at (exact −273°C), the volume of a given mass of a gas reduces to zero. A graph between volume and temperature gives a straight line and cuts the temperature axis at −273.15°C.

Thus the temperature at which the given volume of a gas reduces to zero is called Absolute zero. For this reason a new temperature scale has been developed. This new scale is called Absolute or Kelvin scale. It starts from −273.15°C which is marked as zero Kelvin.

−273.15°C = Zero Kelvin

This is equal for all other gases as well and is the lowest possible in the gaseous state but actually it does not happen because all the gases liquify or solidify before they reach this temperature (−273.15°C). This temperature is considered as the ***lowest possible temperature***.

Significance of Absolute Zero:

i. It is not possible to have a gas at Absolute zero because it **changes into liquid or solid states.**

ii. At –273.15°C, the gas particles do not move and **volume of gas becomes minimum.**

iii. Absolute zero may be initialised to get the **Kelvin scale of temperature.**

iv. Zero Kelvin (-273.16°C) is called as Absolute zero which is the **lowest possible temperature.**

Q16. Example 7: If 50 cm^3 of a gas in a syringe at 15°C is heated to 50°C and the piston of syringe is allowed to move outwards against constant atmospheric pressure, calculate the new volume of the hot gas.

Solution:

$V = 50cm^3$ $V_2 = ?$

$T_1 = 15°C + 273 = 288K$ $T_2 = 50°C + 273 = 323K$

According to Charles's Law

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \quad \text{or} \quad \frac{50\,cm^3}{288\,K} = \frac{V^2}{323} \quad \text{or} \quad \frac{50 \times 323}{288} \times \frac{cm^3 \times \cancel{K}}{\cancel{K}}$$

$$= 56\,cm^3$$

Therefore the new volume of the hot gas = $56cm^3$

Q17. Example 8: At 17°C, a sample of H_2 gas occupies 125 cm^3. What would be the volume at 100°C? (Pressure remains constant)

Solution:

$V_1 = 125\,cm^3$ $V_2 = ?$

$T_1 = 17°C + 273 = 290\,K$ $T_2 = 100°C + 273 = 373\,K$

According to Charles's Law

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

$$\frac{125\,cm^3}{290\,K} = \frac{V_2}{373\,K} \quad \text{or}$$

$$V_2 = \frac{125 \times 373}{290} \cdot \frac{cm^3 \times \cancel{K}}{\cancel{K}} = 161cm^3$$

Thus the new volume will be 161 cm^3

Activity For Student

A Volume-Temperature experiment which was performed on a sample of H_2 gas produced the Charles's Law relationship V= 0.167 T (V is in dm^3). At what temperature in degree Celsius would this sample of gas occupy 50 dm^3, if the pressure remains constant? (Ans: 26.40°C)

Solution: As given relation

$$V = 0.167\,T$$

$$Putting\ the\ V = 50\ dm^3$$

$$\frac{50}{0.167} = T$$

$$299.4\ K = T$$

$$t = 299.4 - 273$$

$$t = 26.40\ ^\circ C$$

Q18. Explain Avogadro's Law with the help of examples and what is its significance?

Ans: Avogadro's Law:

It states that equal volumes of all gases at the same temperature and pressure must contain equal number of molecules.

It means volume is directly proportional to the number of molecules at constant pressure and temperature.

$$V \propto n$$

Examples:

1 mole of N_2 (28 g) at S.T.P. $= 22.4\,dm^3 = 6.022 \times 10^{23}$ molecules

1 mole of Cl_2 (71 g) at S.T.P. $= 22.4\,dm^3 = 6.022 \times 10^{23}$ molecules

1 mole of O_2 (32 g) at S.T.P. $= 22.4\,dm^3 = 6.022 \times 10^{23}$ molecules

1 mole of CO_2 (44 g) at S.T.P. $= 22.4\,dm^3 = 6.022 \times 10^{23}$ molecules

Significance of Avogadro's Law:

Avogadro's law can be applied to determine:

i. Mass **ii.** Volume **iii.** Number of molecules of a gas

Q19. By using Avogadro's Law explain molar volume.

Ans: Molar volume:

1gram molecule or one mole of a gas will have 22.4 dm^3 at S.T.P. This volume of 22.4 dm^3 at S.T.P is called Molar volume.

1 mole of N_2 (28 g) at S.T.P. $= 22.4\,dm^3 = 6.022 \times 10^{23}$ molecules

1 mole of Cl_2 (71 g) at S.T.P. $= 22.4\,dm^3 = 6.022 \times 10^{23}$ molecules

1 mole of O_2 (32 g) at S.T.P. $= 22.4\,dm^3 = 6.022 \times 10^{23}$ molecules

1 mole of CO_2 (44 g) at S.T.P. $= 22.4\,dm^3 = 6.022 \times 10^{23}$ molecules

The standard temperature is 0°C and pressure as one atmosphere.

Q20. Example 9: Calculate the (I) number of molecules and atoms (ii) Number of mole (iii) Volume in dm^3 of 10g of ammonia at S.T.P.

Solution:

(i) Number of molecules and atoms:

According to Avogadro's Law, one mole of any gas at S.T.P contains 6.022×10^{23} molecules in case of di-atomic and poly-atomic molecules. For example, in the case of ammonia one mole of it contains 6.022×10^{23} molecules.

Now mass of NH_3 = 10g

Number of moles of NH_3 $= \frac{\text{mass}}{\text{molecular mass}} = \frac{10}{17}$

$= 0.588$ moles

One mole of NH_3 at S.T.P. $= 6.022 \times 10^{23}$ molecules.

0.588 moles of NH_3 at S.T.P. $= 0.588 \times 6.022 \times 10^{23}$ molecules

$= 3.54 \times 10^{23}$ molecules

Now number of atoms = no of molecules × atomicity (Number of atoms in one molecule of NH_3)

$= 3.54 \times 10^{23} \times 4$

$= 14.16 \times 10^{23}$ atoms

(ii) Volume in dm^3

We know that

$22.4 dm^3$ of any gas at S.T.P = 1 mole

$22.4 dm^3$ of NH_3 at S.T.P = 17g

or 17g $NH_3 = 22.4\ dm^3$

10g $NH_3 = \frac{22.4}{17} \times 10 = 13.176 dm^3$

Self Check Exercise 4.4

Calculate the number of molecules in (i) 1 dm^3 of oxygen (ii) 2 dm^3 of Hydrogen (iii) 2.5 dm^3 of Nitrogen All gases are at S.T.P.

(Ans: (i) 2.7×10^{22} (ii) 5.4×10^{22} (iii) 6.75×10^{22})

Solution:

(i) $22.414\ dm^3\ of\ oxygen\ at\ STP = 1\ mole\ of\ oxygen$

$= 6.023 \times 10^{23}\ molecules$

$1\ dm^3\ of\ oxygen\ at\ STP = \frac{6.023 \times 10^{23}}{22.414}$

$2.7 \times 10^{22}\ molecules$

(ii) $22.414\ dm^3\ of\ hydrogen\ at\ STP = 1\ mole\ of\ hydrogen$

$= 6.023 \times 10^{23}\ molecules$

$2\ dm^3\ of\ hydrogen\ at\ STP = 2 \times \frac{6.023 \times 10^{23}}{22.414}$

$5.4 \times 10^{22}\ molecules$

(iii)

$22.414\ dm^3\ of\ nitrogen\ at\ STP = 1\ mole\ of\ nitrogen$

$= 6.J23 \times 10^{23}\ molecules$

$2.5\ dm^3\ of\ nitrogen\ at\ STP = 2.5 \times \frac{6.023 \times 10^{23}}{22.414}$

$6.75 \times 10^{22}\ molecules$

Q21. Derive General Gas Equation or Ideal Gas Law Equation with the help of Boyle's Law, Charles's Law and Avogadro's Law.

Ans: Derivation of Ideal Gas Equation:

The ideal Gas Equation is a combination of three Laws:

1. Boyle's Law **2.** Charles's Law **3.** Avogadro's Law

Now according to Boyle's Law

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

According to Charles's Law

$$V \propto T \quad \text{(at constant.....P and n).....} \quad \textbf{(2)}$$

According to Avogadro's Law

$$V \propto n \quad \text{(at constant.....P and T).....} \quad \textbf{(3)}$$

Where n is the number of moles of the gas.

Combining equations (1), (2) and (3), we get

$$\boldsymbol{V \propto \frac{1}{P} \times T \times n}$$

$$\boldsymbol{V \propto \frac{Tn}{P}}$$

$$\boldsymbol{V = R\frac{Tn}{P}} \quad \text{or} \quad \boldsymbol{PV = nRT}$$

Where R is a constant called General Gas Constant. This equation is called General Gas Equation or Ideal Gas or Ideal Gas Law Equation.

If **n = 1** Then $PV = RT$ or $\frac{PV}{T} = R$

So $\frac{P_1 V_1}{T_1} = R$ (For first condition of gas)

and $\frac{P_2 V_2}{T_2} = R$ (For second condition of gas)

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

Q22. Example 10: A certain mass of gas occuples 1000 cm^3 at 57°C and 723 mm of Hg. What will be the volume at S.T.P?

Solution

$V_1 = 1000cm^3$ $V = ?$

$P_1 = 723mm$ of Hg $P = 760mm$ of Hg

$T_1 = 57°C + 273 = 330TK$ $T_2 = 273TK$

Now apply the general gas equation

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

$$\frac{726mm\text{ of Hg} \times 1000cm^3}{330K} = \frac{760mm\text{ of Hg} \times V_2}{273K}$$

$$V_2 = \frac{726\text{mm of Hg} \times 1000\text{cm}^3}{330\text{K}} \times \frac{273\text{K}}{760\text{ mm of Hg}}$$

$$V_2 = \frac{726 \times 1000 \times 273}{330 \times 760} \quad \frac{\text{mm of Hg} \times \text{cm}^3 \times \text{K}}{\text{K} \times \text{mm of Hg}}$$

$V_2 = 790.3\ \text{cm}^3$ So, the volume of the gas at S.T.P = 790.3cm^3

Activity For Student

Q.1 A given mass of gas occupies 850 cm^3 at 320K and $0.92 \times 10^{+5}$ Nm^{-2} pressure. Calculate the new volume of the gas at S.T.P.

(Ans: $660cm^3$)

Solution:

$V_1 = 850\ cm^3$

$V_2 = ?$

$P_1 = 0.92 \times 10^5\ Nm^{-2}$

$P_2 = 1\ atm = 101325\ Nm^{-2}$

$T_1 = 320\ K$

$T_2 = 273\ K$

Now apply the general gas equation.

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

$$\frac{0.92 \times 10^5\ Nm^{-2} \times 850\ cm^3}{320\ K} = \frac{101325\ Nm^{-2} \times V_2}{273\ K}$$

$$V_2 = \frac{0.92 \times 10^5\ Nm^{-2} \times 850\ cm^3 \times 273\ K}{320\ K \times 101325\ Nm^{-2}}$$

$V_2 = 658.410\ cm^3$

Q.2 A container contains $2.5m^3$ of Helium gas at 15°C and 98 Nm^{-2} pressures. What volume would this gas occupy at S.T.P?

(Ans: $2.3cm^3$)

Solution:

$V_1 = 2.5\ m^3$

$V_2 = ?$

$P_1 = 98\ Nm^{-2}$

$P_2 = 1\ atm = 101325\ Nm^{-2}$

$T_1 = 15\ ^\circ C + 273\ K = 288\ K$

$T_2 = 273\ K$

Now apply the general gas equation

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

$$\frac{98\ Nm^{-2} \times 2.5\ m^3}{288\ K} = \frac{101325\ Nm^{-2} \times V_2}{273\ K}$$

$$V_2 = \frac{98\ Nm^{-2} \times 2.5\ m^3 \times 273\ K}{288\ K \times 101325\ Nm^{-2}}$$

$V_2 = 0.0023\ m^3$

$$V_2 = 0.0023 \times 1000 = 2.3\ cm^3$$

Q23. List the Significance of General Gas Equation.

Ans: Significance of General Gas Equation:

With the help of the Ideal Gas equation, we can determine:

(a) Molecular mass of the gas. **(b)** Density of the gas.

Q24. Derive a relation for molecular mass of the gas with the help of general gas equation.

Ans: Molecular Mass of the Gas:

According to the general gas equation PV = nRT ..... **(1)**

But n (numbber of moles) = $\frac{W}{M}$ Where W = mass of the gas

M = molecular mass of the gas

Putting the value of n in

$$PV = \frac{W}{M} RT \quad \text{..... (1)}$$

$$MPV = WRT \quad \text{..... (2)}$$

$$\text{Or} \quad M = \frac{WRT}{PV} \quad \text{..... (3)}$$

So molecular mass (M) can be calculated if P, V W and T are known.

Q25. Derive a relation for density of the gas with the help of general gas equation.

Ans: Density of the gas (g dm^{-3}):

According to the General Gas Equation, $PV = nRT$ ..... (I)

but $n = \frac{W}{M}$, Therefore $PV = \frac{W}{M} RT$ or $MP = \frac{W}{V} RT$

as $\frac{W}{V} = d$ (density),

so $MP = d\ RT$

or $d = \frac{MP}{RT}$ (II)

So knowing the values of P, T and M, the density can be calculated.

Q26. Calculate the numerical value of R in different ways with the help of general gas equation.

Ans: Numerical Value of R:

The value of R can be calculated in different ways as shown below:

(I) If one mole of a gas is taken at S.T.P (273 K, 1 atm), then the volume occupied by it is $22.4 dm^3$. Now according to the general gas equation.

$$PV = nRT \quad (1)$$

Where $V = 22.4 dm^3$ P = 1 atm n = 1 mole T = 273 K Then R =?

From the equation (1)

$$PV = nRT$$

$$R = \frac{VP}{nT} = \frac{22.4\,dm^3 \times 1\,atm}{1\,mole \times 273\,K}$$

$$R = \frac{22.4 \times 1}{1 \times 273}\ dm^3 \times atm\ mole^{-1}\ K^{-1}$$

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

(ii) If pressure is measured in mm of Hg or torr and V in cm^3 then

$$R = 0.0821\,dm^3 \times atm\ mole^{-1}\ K^{-1}$$

$$= 0.0821\,dm^3 \times 760mm\ of\ Hg\ mole^{-1}\ K^{-1}$$

$$= 62.4\,dm^3 mm\ mole^{-1}\ K^{-1}$$

$$= 62.4\,dm^3 torr\ mole^{-1}\ K^{-1}\ (one\ mm\ of\ Hg\ =\ one\ torr)$$

$$= 62.4 \times 1000\ cm^3\ torr\ mole^{-1}K^{-1}$$

$$= 62.4 \times 100\,cm^3\ torr\ mole^{-1}\ K^{-1}$$

$$= 62400\,cm^3 torr\ mole^{-1}\ K^{-1}$$

Q27. Calculate the numerical value and SI unit of R with the help of general gas equation.

Ans: In SI Units, pressure is expressed in Nm^{-2} and volume in m^3, then

$V = 0.0224\,m^3$ $(1dm^3 = 10^{-3}\ m^3)$

$P = 101325\,Nm^{-2}$

n = 1 mole

T = 273 K

R = ?

$$R = \frac{VP}{nT} = \frac{0.0224\,m^3 \times 101325\,Nm^{-2}}{1\,mole \times 273\,K}$$

$$= \frac{0.0224 \times 101325}{1 \times 273}\ m^3\ Nm^{-2}\ mole^{-1}\ K^{-1}$$

$$R = \underline{8.3143\,Nm\ mole^{-1}\ K^{-1}}$$

Now we know that 1 Nm = 1 J

So, $R = 8.3143J\ mole^{-1}\ K^{-1}$

Q28. Example 11: Calculate the average molar mass of air at sea level and 0°C if the density of air is 1.29 Kgm^{-3}.

Solution:

At sea level the pressure may be taken equal to 1 atm or 101325 Pa. We know that $d = \frac{MP}{RT}$ or M (average molar mass) $= \frac{dRT}{P}$ .....(1)

where $d = 1.29\,Kgm^{-3}$

$R = 8.3145\,J\,K^{-1}\,mole^{-1}$

$T = 273.15\,K$

$P = 101325\,Nm^{-2} = Kgm^{-1}s^{-1}$

Now putting the values of these in equation .....(1)

$$M = \frac{1.29\,Kgm^{-3} \times 8.3145\,JK^{-1}mole^{-1} \times 273.15\,K}{101325\,Kgm^{-1}S^{-1}}$$

$$\therefore M = \frac{1.29 \times 8.3145 \times 273.15}{101325} \cdot \frac{Kgm^{-3}Kgm^{2}s^{-2}k^{-1}mole^{-1}}{Kgm^{-1}s^{-2}m^{-2}}$$

$M = 0.0289\,Kg\,mole^{-1}$

Q29. Example 12: A certain gas occupies a volume of 6 dm^3 under a pressure of 720 mm of Hg at 25°C. What volume will this gas occupy under standard conditions of Temperature and Pressure (S.T.P)?

Solution:

$V_1 = 6\,dm^3$ $\quad V_2 = ?$

$P_1 = 720\,mm$ $\quad P_2 = 760\,mm$

$T_1 = 25°C + 273$ $\quad T_2 = 273\,K$

$= 298\,K$

According to the General Gas Equation

or $\quad P_1 = \frac{n_1 RT}{V}$

$$= \frac{6 \times 720 \times 298}{760 \times 273} \frac{dm^3 \times \cancel{mm\ of\ Hg} \times K}{\cancel{mm\ of\ Hg} \times K} = 8.204\ dm^3$$

Q30. Example 13: Four grams of CH_4 at 27°C and a pressure of 2.5 atm occupy a volume of 2.46 dm^3. Calculate the value of R.

Solution:

$$\text{No of moles of } CH_4 = \frac{\text{Mass}}{\text{Molecular mass}}$$

$n = \frac{4}{16} = \frac{1}{4}$ moles

$T = 27°C + 273 = 300\,K$

$P = 2.5\,atm$

$V = 2.46\,dm^3$

$R = ?$

According to the General Gas Equation,

$VP = nRT$

or $\quad R = \frac{VP}{nT}$ ..... (1)

$$= \frac{2.46\,dm^3 \times 2.5\,atm}{\frac{1}{4}\,mole \times 300\,K}$$

$$= \frac{2.46 \times 2.5 \times 4}{300 \times 1}\ dm^3\ atm\ mole^{-1}\ K^{-1}$$

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

Q31. Example 14: Find the density of ammonia gas at 100°C when confined by a pressure of 1600 mm of Hg.

Solution: d = ?

T = 100°C + 273 = 373K

P = 1600 mm of Hg = $\frac{1600}{760}$ atm = 2.105 atm

$M_{NH_3} = 17\,g\ mole^{-1}$

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

Now according to the General Gas Equation,

MP = dRT (Derived from General Gas Equation)

or $$d = \frac{MP}{RT} = \frac{17g\ mole^{-1} \times 2.105\,atm}{0.0821\,dm^3 atm\ mole^{-1}\ K^{-1} \times 273\,K}$$

$$= \frac{17 \times 2.105}{0.082 \times 273}$$

$$= 1.598\,g\ dm^{-3}$$

Q32. Example 15: At 27°C a volume of 500 cm^3 of H_2 measured under a pressure of 400 mm of Hg and 1000 cm^3 of N_2 measured under a pressure of 600 mm of Hg are introduced into an evacuated 2 dm^3 flask. Calculate the resulting Pressure.

Solution:

We will apply Boyle's Law, as the temperature is constant.

For H_2:

$V_1 = 2dm^3$ $\quad V = 500cm^3 = \frac{500}{1000}\ dm^3 = \frac{1}{2}dm^3$

$P_1 = ?$ $\quad P_2 = 400\,mm = \frac{400}{760}\ atm = 0.526\,atm$

$V_1 P_1 = V_2 P_2$ or $2 \times P_2 = \frac{1}{2} \times 0.526$ $\quad P_2 = 0.1315\,atm$

For N_2:

V_3 = 1000 cm^3 = 1 dm^3

$P_3 = ?$ $\quad P_4 = \frac{400\,mm}{760\,mm} = 0.789\,atm$

$V_3 P_3 = V_4 P_4$ or $2 \times P_3 = 1 \times 0.789 = 0.3945\,atm$

Total pressure = 0 1315 atm + 0 3945 atm = 0 5260 atm

Q33. Example 16: Find the total pressure exerted by 2g of ethane and 3g of CO_2 contained in a 5dm³ vessel at 50°C.

Solution:

44 g of CO_2 = 1 mole

$$3\,g\,CO_2 = \frac{1}{44} \times 3 \quad 0.068\,moles$$

Number of moles of CO_2 = 0.068 moles

$V = 5\,dm^3$ $T = 50°C + 273 = 323K$ $R = 0.082\,dm^3\,atm\,K^{-1}\,mole^{-1}$

Now according to the General Gas Equation

$$VP = nRT \text{ or } P = \frac{nRT}{V}$$

$$P_{CO_2} = \frac{0.068\,mole \times 0.082 \times dm^3\,atm.K^{-1}\,mole^{-1} \times 323K}{5\,dm^3}$$

$$= \frac{0.068 \times 0.082 \times 323}{5} = 0.678\,atm$$

$30\,g$ Ethane $(C_2H_6) = 1$ mole

$$2\,g \text{ Ethane } (C_2H_6) = \frac{1}{30} \times 2 = 0.06\,moles$$

$V = 5\,dm^3$ $n = 0.6\,moles$ $T = 50°C + 273 = 323K$

$R = 0.082\,dm^3\,atm\,K^{-1}\,mole^{-1}$

$VP = n\,R\,T$ or

$$P = \frac{nRT}{V} = 0.06\,moles \times 0.082\,dm^3\,atm\,K^{-1}\,mole^{-1} \times 323K$$

$$= 0.06 \times 0.082 \times 323 = 6.317\,atm$$

Q34. Briefly explain why Real gases Deviate from the Gas Laws explain it with the help of example.

Ans: Deviation of Real gases from the Gas Laws:

At low pressures, the molecules are generally very far apart and the attractions between the molecules are insignificant. At temperatures above the condensation point, the Kinetic Energy of the gaseous molecules is so great that it overwhelms the attractive forces between them.

Example:

The example of Boyle's Law (PV = a (constant) be quoted as follows e.g. for H_2.

Pressure (In atm)	Volume (In dm^3)	P V
0.1	224.1	22.41
50	0.4634	23.71
1000	0.384	38.34

The result (PV) is quite different, temperature remains constant.

Q35. Briefly explain the causes of Deviation of Real Gases.

(OR)

Gases show non-ideal behaviour at low temperature and high pressure. Explain.

Ans: Causes of Deviations (Limitations of Boyle's Law and Charles's Law):

The ideal condition is that the ideal gas obeys Gas Laws at all temperature. But at low temperature and high pressure deviations caused by gases is due to the two faulty assumptions of kinetic molecular theory.

(I) Inter – molecular Forces of attraction:

These are no attractive or repulsive forces between the gas molecules.

(II) Volume:

The volume of the gas molecules is negligible as compared to the volume of the container.

Explanation of deviations:

(I) At lower temperature:

At high temperature the Kinetic energy of the gaseous molecules is very high. As a result the attractive forces between them are negligible. But when temperature is decreased, the Kinetic Energy of the molecules decreases. The intermolecular forces become significant. It means that the molecules come close to each other.

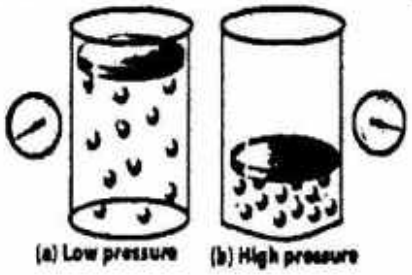
At a certain very low temperature the gases change into the liquid state. Therefore attractive forces between the gas molecules become significant near liquefying temperature. That is why the ideal gases deviate from their original behaviour at low temperature

Example:

SO_2 liquefies at –10°C while H_2 at –252.7°C. Therefore attractive forces between SO_2 molecules cannot be considered negligible at room temperature. For this reason SO_2 gas shows non-ideal behaviour at room temperature as compared to H_2 gas.

(II) At high pressure:

At low pressure actual volume of gas molecules is very small as compared to the volume of the container. However, this volume does not remain negligible at high pressure. This can be understood from the following figure.



Suppose:

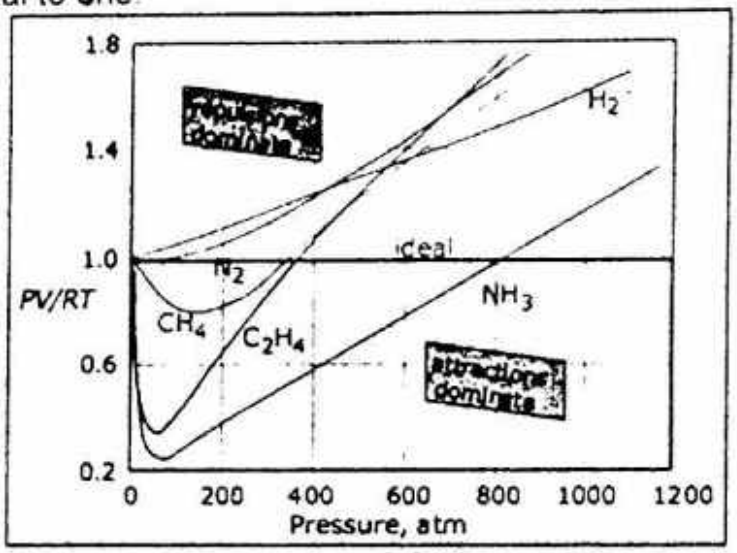
When pressure is one atm, individual volume of one mole of gas molecules say 32 cm^3 is negligible as compared to the total volume of gas 22400 cm^3. But if it is subjected to a pressure of 100atm, the volume of the gas is reduced to 224 cm^3. Under this pressure, individual volume of gas molecules (32 cm^3) is not negligible as compared a volume of the gas (244 cm^3).

Q36. Describe Graphical Representation of Deviation of Gases from ideal behaviour.

Ans: Graphical Representation:

For one mole of a gas, if a graph is plotted between $\frac{PV}{RT}$ and pressure, then for an ideal gas, it must give a straight line (the temperature remains at 0°C. But actually it is seen that the gases do not give a straight line. They deviate from their original behaviour as shown in the figure.

Actually, when the pressure is initially increased, it pushes the molecules closer and increases the intermolecular forces. Due to these forces, the volume of the gas shrinks more than that predicted from the gas laws. (i.e. Boyle's Law). This will decrease the value of $\frac{PV}{RT}$. That is why the gases do not show a straight line. $\frac{PV}{RT}$ is called compressibility factor. For ideal gases it should be one, whereas for real gases it is not equal to one.



The sudden decrease in the curve of C_2H_4 and NH_3 shows that these gases liquefy at the given pressure.

Q37. Explain and derive Van Der Waal's Equation for Real Gases.

Ans: Van Der Waal's Equation: (Correction Factors to Ideal Gas Equation):

The general gas equation is not applicable to real gases at all temperature and pressure. In order to make it applicable to real gases, a Dutch Scientist Van der Waals in 1873 made correction in volume and pressure of the gases and derived an equation known as Vander Waals equation.

Volume Correction:

Vander Waal thought that some of the volume of the vessel is occupied by the molecules of the gas and that volume is not available for the free movement of the molecules. Actually, we need the free volume of the gas and that is obtained when we subtract the individual volume of gas molecules from the volume of the vessel.

$\mathbf{V = V_{vessel} - V_{molecule}}$

V is free volume

$\mathbf{V_{molecule} = nb}$

where **b** is individual volume of one mole of the gas molecules and 'n' is total number of moles of the gas molecules.

Then, $\mathbf{V = V_{vessel} - n\,b}$ ..... **(i)**

Here 'b' is effective volume of gas molecules Keep it in mind that 'b' is not the actual volume of gas molecule, but is roughly equal to 4 times their molar mass volumes. If we have one mole of gas, then $b = 4V_m$ where, V_m is the actual volume of gas molecules for one mole of the gas.

Pressure correction:

The pressure which is exerted on the walls of the vessels is due to collisions of molecules of gas. Since there are some weak forces of attractions among the gas molecules, so the molecules cannot hit the walls of the vessels with that much force with which they should have been in the absence of these attractive forces.

It means that the pressure being observed on the walls of the vessels is a little bit lesser than the ideal pressure.

Therefore, $P_{observed} = P_{ideal} - P_{lessened}$

If the ideal pressure is denoted by P_i and the pressure lessened due to molecular attractions is denoted by P' then

$$P_{observed} = P_i - P'$$

and $$P_i = P_{observed} + P'$$

But the value of P' is inversely proportional to the density of the gas so,

$$P' \propto (\text{density})^2$$

$$P' = \frac{an^2}{V^2}$$

where **n** = number of moles, **V** = volume and **a** = proportionality constant

$$P_i = P + \frac{an^2}{V^2}$$

After the correction of volume and pressure, **PV = nRT** becomes

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

For 1 mole of gas, n = 1, then

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

Q38. List the Van der Waals' constants for some gases.

Ans:

Van der Waals' Constants for Some Gases		
Substance	**a (dm^3 atm $mole^{-1}$)**	**b (dm^3 $mole^{-1}$)**
H_2	0.0247	0.0266
He	0.0034	0.0237
N_2	0.1408	0.0391
O_2	0.1378	0.0318
Cl_2	0.6579	0.0562
Ar	0.1355	0.0 322
Kr	0.2349	0.0398
CO	0.1505	0.0399
NO	0.1358	0.0279
CO_2	0.3640	0.0427
HCl	0.3716	0.0408
SO_2	0.6803	0.0564

H_2O	0.5536	0.0305
NH_3	0.4225	0.0371
CH_4	0.2283	0.0428
CCl_2F_2	0.1066	0.0973

Q39. Example 17: One mole of methane gas is maintained at 300 K. Its volume is 250 cm^3. Calculate the pressure exerted by the gas under the following conditions.

(i) When the gas is ideal

(ii) When the gas is non-ideal, and at

$a = 2.253\ dm^3\ atm\ mole^{-1}$

$b = 0.0428\ dm^3\ mole^{-1}$

Solution:

(i) When the gas is ideal

General gas Equation is applied here.

(a) Gas = Methane:

n = 1 mole

T = 300K

V = $250cm^3$ = $0.25dm^3$

R = $0.0821\ dm^3\ atm\ Mole^{-1}$

P = ?

Now applying the General Gas Equation,

PV = nRT,

$$P = \frac{nRT}{V}$$

$$= \frac{1\text{mole} \times 0.0821\,\text{dm}^3\,\text{atm}\,\text{mole}^{-1}\,\text{K}^{-1} \times 300\,\text{K}}{0.25\,\text{dm}^3}$$

$$= \frac{1 \times 0.0821 \times 300}{0.25} \cdot \frac{\cancel{\text{mole}} \times \cancel{\text{dm}^3}\,\text{atm}\,\cancel{\text{mole}^{-1}}\,\cancel{\text{K}^{-1}} \times \cancel{\text{K}}}{\cancel{\text{dm}^3}}$$

= 98.52 atm

(ii) When the gas is non-ideal (Here Van der Waals' equation is applied)

$$(P + \frac{n^2a}{V^2})(V - nb) = nRT \quad \text{or}$$

$$PV - Pnb + \frac{n^2a}{V} - \frac{n^3ab}{V^2} = nRT$$

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

Now dividing by (V-nb) on both sides

$$P = \frac{nRT}{V - nb} - \frac{n^2a}{V\cancel{(V - nb)}}\left(\frac{\cancel{(V - nb)}}{V}\right)$$

$$P = \frac{nRT}{V - nb} - \frac{n^2 a}{V^2}$$

Putting the following values in this equation

n = 1 mole, R = 0.0821 atm mole^{-1} K^{-1}

V = 0.25dm^3, T = 300K,

a = 2.253dm^3 atm mole^{-1}, b = 0.0428dm^3 mole^{-1}

$$= \frac{1M \times 0.0821 dm^3\, mole^{-1} K^{-1} \times 300K}{0.25 dm^3 - (1M \times 0.0428 dm^3\, mole^{-1})} - \frac{1M \times 2.25 dm^3\, atmos\, K^{-1}}{(0.25)^3\, dm^3}$$

$$= \frac{1 \times 0.0821 \times 300}{0.25 - 0.0428} - \frac{2.253}{0.625} \cdot \frac{M \cancel{dm^3}\, atm.mole^{-1} \cancel{K^{-1}}\, \cancel{K^{-1}}}{\cancel{dm^3}} .(n = 1)$$

$$= \frac{24.63}{0.2072} - \frac{2.253}{0.625} .atm$$

$$= 118.87 - 3.6048 = 115.2652\, atmos$$

Q40. Example 18: Estimate the molar volume of N_2 at 500K and 100 atm by treating it as a Van der Waals' equation.
(R = 0.0821 dm^3 atm K^{-1} mole^{-1})
a = 1.39dm^6 atm Mole^{-2}
b = 3.913×10^{-2} dm^3 mole^{-1} (Ans: Molar Volume = 0.5dm^3)

Solution:

For 1 mole of a gas (n = 1) Van der Waals equation is given by

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

$$(V - b) = \frac{RT}{P + \frac{a}{V^2}} + b \text{ .....(i)}$$

Since $PV = RT \implies V = \frac{RT}{P} \implies V^2 = \frac{R^2T^2}{P^2}$ Putting in equation (i)

$$V = \frac{RT}{P + \frac{a}{\frac{R^2T^2}{P^2}}} + b$$

$$V = \frac{RT}{P + \frac{aP^2}{R^2T^2}} + b$$

$$V = \frac{0.0821 \times 500}{100 + \frac{1.39 \times 100^2}{(0.0821)^2 \times (500)^2}} + 3.913 \times 10^{-2}$$

$$V = \frac{41.05}{100 + \frac{1.39 \times 10000}{0.0067 \times 250000}} + 3.913 \times 10^{-2}$$

$$V = \frac{41.05}{100 + \frac{13900}{1685.1025}} + 3.913 \times 10^{-2}$$

$$V = \frac{41.05}{100 + 8.23} + 3.913 \times 10^{-2}$$

$$V = \frac{41.05}{108.23} + 3.913 \times 10^{-2}$$

$$V = 0.4 + 3.913 \times 10^{-2}$$

$$V = 0.4 + 0.03913$$

$$V = 0.439$$

$$V \approx 0.5\, dm^3$$

Q41. Define Dalton's Law of Partial Pressure and its applications.

Ans: Dalton's Law of Partial Pressure and its Applications:

An English Chemist John Dalton (1766-1844) showed in 1801 that "The total pressure exerted by a gaseous mixture is equal to the sum of partial pressures of each gas present in the mixture."

Conditions for Dalton's Law of Partial Pressure:

The law is only obeyed, if the component gases do not chemically react with each other. Moreover the gases must behave ideally. The partial pressure is the pressure exerted by one component of the gaseous mixture.

Mathematically:

Total Pressure $(P_t) = P_A + P_B + P_C$

Whereas P_A, P_B and P_C are the partial pressures of individual gases.

Example of Air in terms of Dalton's Law of Partial Pressure:

Air is a mixture of non-reacting gases. Percentage of each gas in the gaseous mixture is N_2 (78.08%), O_2 (20%) Ar (0.93%), CO_2 (0.03%). Traces of Ne, He, Kr, H_2, along with various amounts of water vapours and pollutant (which can pollute the atmosphere) gases such as oxides of sulphur and Nitrogen are also found in air. The total pressure exerted by the air will be the sum of partial pressure of each gas.

Q42. Derive Relationship between pressure and number of moles of a Gas with the help of Dalton's Law of Partial Pressure.

Ans: Relationship between pressure and number of moles of a gas:

If n_A, n_B and n_C are the number of moles of the gases and P_A, P_B and P_C be their partial pressure respectively.

Then, according to the general gas equation $PV = nRT$

For gas A:

$$P_A = \frac{n_A RT}{V} \qquad (1)$$

For gas B:

$$P_B = \frac{n_B RT}{V} \qquad (2)$$

For gas C:

$$P_C = \frac{n_C RT}{V} \qquad (3)$$

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

Where n_t (total number of moles of all the gases) = $n_a + n_b + n_c$

Dividing equation (1) by (4)

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

$$\frac{P_A}{P_t} = \frac{n_A \cancel{R} \cancel{T}}{\cancel{V}} \times \frac{\cancel{V}}{n_t \cancel{R} \cancel{T}}$$

$$\frac{P_A}{P_t} = \frac{n_A}{n_t} \Rightarrow P_A = \frac{n_A}{n_t} \times P_t \Rightarrow P_A \propto n_A$$

Similarly $P_B \propto n_B$ and $P_C \propto n_t$

In general

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

Q43. Derive a Relationship between pressure and mole fraction of a gas with the help Dalton's Law of Partial Pressure.

Ans: Relationship between pressure and mole fraction of a gas:

We know that

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

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

but $\frac{n_A}{n_t} = X_A$ where X_A is called mole fraction of the component

so, $P_A = P_t X_A$

Thus the partial pressure of any component is equal to the product of total pressure of all the components and mole fraction of that component.

Q44. Give at least three Applications of the Law of Partial Pressure.

Ans: Applications of the Law of Partial Pressure:

i. Collection of gases over water:

The Dalton's Law is particularly useful when a gas is generated and subsequently collected over water provided the gas is insoluble in water. The total pressure consists of the pressure of the water vapours in addition to the pressure of the gas that is generated. The pressure due to water vapour is called aqueous tension.

Thus $P_{Total} = P_{gas} + P_{Water}$ or $P_{gas} = P_{Total} - P_{Water}$

The gas contaminated with water may be called a moist gas

ii. Respiration at higher altitudes:

The respiration process in living things depends upon the differences in partial pressure e.g. partial pressure of O_2 in the outside air is higher ($159 g/cm^2$) than in the lungs where the partial pressure of O_2 is lower ($116 g\ cm^{-2}$). However at high altitudes, the pressure of oxygen decreases and the pilots may have uncomfortable breathing in a non-pressurised cabin where the partial pressure of O_2 is about $150 g/cm^2$

iii. Respiration by deep sea divers:

Deep sea divers breathe air under increased pressure. At a depth of 40 metres, the pressure **increases 5 times than** the normal pressure. Therefore regular air cannot be used in diver's tanks because the partial pressure of O_2 would be **795 mm of Hg (159×5 = 795 mm of Hg)** in that case. Therefore deep sea divers use to breathe a mixture of 96% N_2 and 4% O_2 in the respiration tank, the **scuba.**

A ***scuba*** (self-contained breathing apparatus having oxygen used by divers underwater) contains compressed air to breath. As diver returns to the surface, it becomes hazardous for him if not handled properly. As the diver comes up, the pressure of the surrounding water drops. Consequently, the compressed air in the lungs expands. Surfacing must be done very slowly so that the compressed air may escape out from the lungs without causing damage to them.

Q45. How deep sea divers respire on return from sea?

Ans: The percentage of N_2 in air is about 80%. If N_2 is present in diver's tank, the solubility of N_2 in blood can increase with increase in pressure in very deep sea. And blood of a diver may be saturated with N_2 during a dive. If the diver returns quickly and relatively at low pressure at the surface, he will face life threatening condition called the "bends" Deep sea divers must either use different solutions of gases such as ***He*** in O_2 or spend many hours in a decompression chamber after a dive. The pressure in a decompression chamber is slowly lowered over many hours. Nitrogen gas comes out of the blood and disposed off safely for normal breathing.

Q46. Example 19: 1.00 mole of N_2 and 3.00 moles of H_2 are enclosed in a container of volume 10.0 dm^3 at 298K. What are the partial pressures and the total pressure?

Solution:

No of moles of N_2 (n_1) = 1.00 mole

No of moles of H_2 (n_2) = 3.00 moles

Total number of moles (n_{Total}) = 1 + 3 = 4 moles

V = 10.0 dm^3

T = 298K, R = 0.082 dm^3 atm K^{-1} $Mole^{-1}$

According to the General Gas Equation, PV=nRT

$$\text{or} \quad P_{Total} = \frac{n_{Total} RT}{V}$$

$$= \frac{4 \times 0.082 \times 298}{10} \cdot \frac{\cancel{dm^3}\ \text{atmos}\ \cancel{K^{-1}}\ \cancel{Mole^{-1}} \times \cancel{K}}{\cancel{dm^3}\ \cancel{Mole}}$$

$$\frac{4 \times 0.082 \times 298}{10}\ \text{atm}$$

$$= 9.77\ \text{atm}$$

$$P_{N_2} = \frac{n_{N_2}}{n_{Total}} \times P_{Total} = \frac{1}{4} \times 9.77 = 2.44\,\text{atm}$$

$$P_{H_2} = \frac{n_{H_2}}{n_{Total}} \times P_{Total} = \frac{3}{4} \times 9.77 = 7.33\,atm$$

Q47. Example 20: The composition of dry air at sea level is approx. N_2 = 75.5, O_2=23.2. Ar =1.3 in percentage by weight. What is the partial pressure of each component when the total pressure is 1.000 atm.

Solution:

Let the total number of moles of air = 100 then

No of moles of $N_2\left(n_{N_2}\right)$ = 75.5

No of moles of $O_2\left(n_{O_2}\right)$ = 23.2

No of moles of Ar $\left(n_{A_r}\right)$ = 01 3

$$P_{Total} = 1.00\,atm.$$

$$P_{N2} = \frac{\text{No of moles of } N_2}{\text{Total no of moles}} \times P_{Total}$$

$$= \frac{75.52}{100.00} \times 1.00 = 0.755\,atm.$$

$$P_{O_2} = \frac{n_{O_2}}{n_{Total}} \times P_{Total} = \frac{23.2}{100} \times 1.00 = 0.232\,atm$$

$$P_{Ar} = \frac{n_{A_r}}{n_{Total}} \times P_{Total}$$

$$= \frac{1.3}{100} \times 1.00 = 0.013\,atm$$

Q48. Example 21: A certain mass of H_2 gas collected over water at 6°C and 765mm of Hg pressure occupied a volume of 35cm³. Calculate its dry volume at S.T.P. (V.P. of water at 6°C = 7mm of Hg)

Solution:

V_1 = 35cm³, V_2 = ? P_1 = 765-7 = 758mm P_2 = 760mm.
T_1 = 6 + 273 = 279K, T_2 = 273K

Applying the Gas Equation ,

$$\frac{V_1 P_1}{T_1} = \frac{V_2 P_2}{T_2} \quad \text{or} \quad V_2 = \frac{V_1 P_1 T_2}{T_1 P_2} = \frac{35\,cm^3 \times 758\,mm \times 273K}{279\,K \times 760\,mm\ of\ Hg} = 34.2\,cm^3$$

Activity For Student

A gas was collected over water at 10°C and 803mm of Hg. If the gas occupies 73cm³, calculate the volume of dry gas at S.T.P. (aqueous tension at 10°C = 9.2mm). (Ans: = 73.55 cm³)

Solution:

$$P_{moist} = P_{dry\ gas} + \text{Aqueous tension}$$

$$P_{dry\ gas} = P_{moist} - \text{Aqueous tension}$$
$$P_{dry\ gas} = 803\ \text{mm of Hg} - 97\ \text{mm of Hg}$$
$$P_{dry\ gas} = 706\ \text{mm of Hg}$$

Volume of gas at 10 °C = $V_1 = 73\ cm^3$

$T_1 = 10\ °C + 273 = 283\ K$

Values at STP : P_2 760

$T_2 = 273\ K$

$V_2 = ?$

According to the general gas equation

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$
$$V_2 = \frac{P_1V_1T_2}{P_2T_1}$$
$$V_2 = \frac{706 \times 73 \times 273}{760 \times 283}$$
$$V_2 = 0.0023\ m^3$$
$$V_2 = 65.41\ cm^3$$

Q49. Define diffusion with the help of example.

Ans: Diffusion:

The spontaneous intermixing of molecules of one gas with another at a given temperature and pressure is called Diffusion. **Diffusion** is the random movement of a gas from an area of higher concentration to an area of lower concentration.

Example:

(I) Fragrance of rose.

(II) Intermixing of NH_3 and HCl to gives dense white fumes of NH_4Cl.

$NH_3 + HCl \rightarrow NH_4Cl$

(III) Reddish brown bromine (Br_2) gas taken in a test tube will combine with air taken in another test tube. Red colouration of Br_2 gas will spread in the test tube containing air.

(Iv) If liquid ink is left in the bottom of test tube containing water, it will spread evenly throughout the water.

Q50. Define Effusion.

Ans: Effusion:

The movement of gaseous molecules through extremely small pores in a region of low pressure is called Effusion. This escaping of molecules is not due to collisions but due to their tendency to escape one by one.

Q51. Example 22: Equal volumes of HCl and SO_2 are confined in a porous container. What would be the comparative rates of diffusion of these gases through the porous walls? The molecular masses of HCl and SO_2 are 36.5 and 64 respectively.

Solution:

Molecular mass of HCl (M_1) = 36

Molecular mass of SO_2 (M_1) = 64

Rate of diffusion of HCl (r_1) = ?

Rate of diffusion of SO_2 (r_2) = ?

According to the Graham's Law of Diffusion

$$\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}} \qquad \sqrt{\frac{64}{36.5}}$$

$$r_1(HCl) \quad r_2(SO_2)$$

$$1.33 : 1$$

Hence rate of diffusion of HCl will be 1.33 times the rate of diffusion of SO_2.

Q52. Example 23: If 280 cm^3 of H_2 diffuse in 40 seconds, how long will it take 490 cm^3 of a gas (X), where vapour density is 25 to diffuse under the same conditions (molecular mass of H_2 is 2).

Solution:

For H_2:

$V = 280\ cm^3$, $M_1 = 2\ gm\ mole^{-1}$

$T_1 = 40$ sec (time for the diffusion)

r_1 = rate of diffusion of $H_2 = \frac{\text{velocity}}{\text{time}} = \frac{280\,cm^3}{40\,sec} = 7\,cm^3\,sec^{-1}$

For Gas (X):

$V = 490\ cm^3$, Vapour density (V.D.) = 25

$T_2 = ?$ M_2 (Molar mass) $= 2 \times$ V.D. $= 2 \times 25 = 50\ gm\ mole^{-1}$

(Rate of diffusion of gas X) $= \frac{\text{velocity}}{\text{time}} = \frac{490\,cm^3}{t_2\,sec} = \frac{490}{t_2}.cm^3\,sec^{-1}$

Now according to Graham's Law of Diffusion

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

$$= \sqrt{\frac{\cancel{50}^{25}\ \cancel{gm\ mole^{-1}}}{\cancel{2}\ \cancel{gm\ mole^{-1}}}}$$

and $\frac{7 \times t_2}{490} = \sqrt{\frac{25}{1}} = \frac{5}{1}$ or $7 \times t_2 \times 1 = 490 \times 5$

or $t_2 = \frac{\cancel{490}^{70} \times 5}{\cancel{7}} = 350\,sec$

Hence time for the diffusion of un-known gas = 350 sec

Activity

If 465 cm^3 of SO_2 can diffuse through porous partition in 30 seconds. How long will 620 cm^3 of H_2S take to diffuse through the same partition (H = 1, S = 32, O = 16) (Ans: 29.15 sec)

Solution:

$$v_{SO_2} = 465\ cm^3$$

$$r_{SO_2} = \frac{Velocity}{Time} = \frac{465}{30} = 15.5\ cm^3 s^{-1}$$

$$M_{SO_2} = 64\ gmol^{-1}$$

$$v_{H_2}s = 620\ cm^3$$

$$v_{H_2}s = 34\ gmol^{-1}$$

$$r_{H_2}s = ?$$

$$\frac{15.5\ cm^3 s^{-1}}{0.72} = \sqrt{\frac{34}{64}}$$

$$r_{H_2}s = \frac{15.5\ cm^3 s^{-1}}{0.72}$$

$$r_{H_2}s = 21.29\ cm^3 s^{-1}$$

$$r_{H_2}s = \frac{Velocity}{Time} = \frac{465}{21.29} = 29.15\ sec$$

$$T_{H_2}s = 29.15\ sec$$

Q53. Explain Joule-Thomson's effect and Liquefaction of Gases.

Ans: Liquefaction of Gases:

Principle:

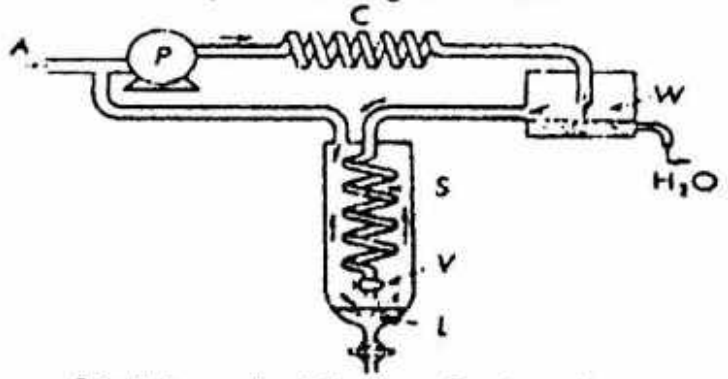
The gases can be liquefied by **Joule-Thomson's effect**. The liquefaction of a gas requires high pressure and low temperature. When a highly compressed gas is allowed to escape out through a throttle (small hole), the temperature falls to such an extent that it changes into the liquid form.

At high pressure, the gaseous molecules come close to each other with the result that the molecular attractions increase. When it is allowed to escape through a nozzle (small hole) into a region of low pressure, the molecules move apart. In doing so, the intermolecular attractions must be overcome. **This energy is taken from the molecules themselves.** Therefore the gas is cooled. This process is repeated for many times until the gas completely changes into the liquid form.

Q54. Describe principle and working of Linde's Method for liquefaction of gases.

Ans: Linde's Method (1895):

According to Joule-Thomson's effect "When a highly compressed gas is allowed to escape out through a throttle" the temperature of the gas falls to such an extent, that it changes into the liquid form". This is the basic principle of the adiabatic expansion which is thermally isolated from its environment, Joule Thomson expansion and subsequent cooling of the air.



Linde's method for liquefaction of air

Working Steps:

(a) Air is compressed to 200 atm approx.
(b) Most of the water in the air condenses and is removed.
(c) The heat generated as a result of compression is removed by passing the gas through coils C.
(d) The dry gas is then passed through a copper spiral coil S.
(e) It is then expanded to almost atmospheric pressure through a controlled valve V.
(f) The issuing gas, cooled now due to Joule-Thomson's effect is passed over the copper spiral and cool further the incoming compressed gas.
(g) The cycle is repeated several times.
(h) The temperature of the expanding gas finally drops and the remaining air is liquefied.
(I) The liquid air collects in the chamber L and can be drawn off.
(j) Any uncondensed air is re-circulated.

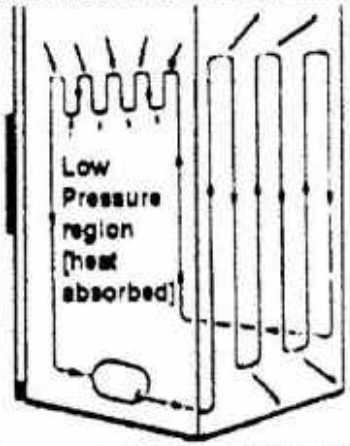
Q55. Gives the different applications of Linde's Method.

Ans: I. Construction of appliances:

The construction of appliances such as refrigerators, heat pumps and air conditioners all work on the principle of Joule-Thomson's effect.

e.g. In a refrigerator a compressor, a compresses a gas such as Freon.

Thus an increase in temperature takes place. The hot **Freon gas** moves to a condenser on the outside of the refrigerator. Air at room temperature cools the gas down and as a result condenses into liquid.



Cooling gas path in a refercgirator

The liquid Freon, now at room temperature, is then passed through a small hole in a restriction valve into the low pressure tubes which are inside the refrigerator. There, the Freon liquid vaporises rapidly. The heat of vaporisation needed to do this is drawn from the kinetic energy of the Freon molecules becoming very cold. This cold gas absorbs heat from the refrigerator and its contents, thereby cooling. Then the Freon gas is once again fed into the compressor and the cycle starts over again.



II. Self cooling pop can (container):

A small container holding liquid CO_2 is built right into the can. When the can is opened, the liquid CO_2 vaporises and escapes

out of the top of the can. The heat absorbed by the vaporising CO_2 can lower the temperature of the POP by about 16°C in a few seconds. Thus the temperature is lowered considerably.

Q56. Describe different Applications of Kinetic Molecular Theory.

Ans: i. Velocity of molecules:

The equation for the root mean square velocity deduced from kinetic equation is $Crms = \sqrt{\frac{3RT}{M}}$

Crms = root mean square velocity
M = molecular mass of the gas
T = absolute temperature.

This equation gives a quantitative relationship between the absolute temperature and the velocities of gas molecules. Higher the temperature of a gas, greater would be the velocities of the molecules.

ii. Graham's Law of diffusion of gases:

According to the postulates of Kinetic Molecular Theory, the Kinetic Energy of the particles is proportional to the Absolute Temperature.

$$\text{As, } \quad E_k = \frac{1}{2}mv^2 \quad \Rightarrow \quad \frac{2E_K}{m} = v^2 \Rightarrow v = \sqrt{\frac{2E_K}{m}} \Rightarrow v = \sqrt{\frac{1}{m}}\sqrt{2E_K}$$

as at constant temperature K.E will remain constant so $\sqrt{2E_k}$ = Constant

$$\text{Therefore} \quad v \propto \sqrt{\frac{1}{m}}$$

Thus the Kinetic Molecular theory predicts that the average volume depends upon the molecular mass (m). Further the average volume of the particles is inversely proportional to the square root of their molecular mass.

Graham's Law states that ***the rate of diffusion of a gas escaped is inversely proportional to the square root of the mass of its molecules***. This law is direct experimental evidence for the Kinetic Molecular Theory.

iii. Use of combined gas laws in calculations (Applying):

Combined Gas Laws:

The combination of Boyle's Law and Charles' Law is called "Combined Gas Law."

Applications:

1. It is applied to solve problems when volume of a fixed amount of gas is involved, when both the temperature and pressure change.
   Thus two conversions factors are involved.
   - **(a)** Conversion factor for pressure
     ___ Boyles' Law
   - **(b)** Conversion factor for temperature
     ___ Charles' Law

Calculations:

$$V_2 = V_1 \times \left(\begin{matrix}\text{Conversion factor due} \\ \text{to temperature change}\end{matrix}\right) \times \left(\begin{matrix}\text{Conversion factor due} \\ \text{to Pressure change}\end{matrix}\right)$$

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

The ratio between T_1 and T_2 is due to Charles' law $(V \propto T)$ and that of P_1 and P_2 is due to Boyle's law $(V \propto \frac{1}{P})$.

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

In combined gas law problems which involve a change in temperature and pressure conditions, the mass of the gas must remain constant.

2. The combined gas law is used when a fixed number of moles of a gas are maintained at different conditions of temperature and pressure.

iv. Ideal Gas Laws to calculate the Pressure or Volume of a gas:

Ideal Gas Laws i.e. Boyles' Law and Charles' Law when combined together give Idea gas law equation at S.T.P.

The equation is $PV = nRT$

If one of the variables is known, the other three can be calculated.

Q57. Example 124: A certain mass of H_2 gas collected over water at 6°C and 765mm of Hg pressure occupied a volume of 35 cm^3. Calculate its dry volume at S.T.P.

Solution:

$V_1 = 35cm^3$ $V_2 = ?$

$P_1 = 765 - 7 = 758mm$ $P_2 = 760mm$

$T_1 = 6 + 273 = 279K$ $T_2 = 273K$

$\frac{V_1 P_1}{T_1} = \frac{V_2 P_2}{T_2}$ or $V_2 = 34.2cm^3$

Q58. Example 25: A fire fighters air tank contains 12.0 dm^3 of air compressed to $1.40 \times 10^4 KP_a$ at $22°C$. What volume of air will this tank provide when it is used in a hot, smoke-filled building where the temperature is 42°C and pressure 102 KP_a.

Solution:

$V_1 = 12\,dm^3$ $V_2 = ?$

$P_1 = 1.40 \times 10^4\,KP_a$ $P_2 = 102\,KP_a$

$T_1 = 22°C + 273 = 295\,K$ $T_2 = 42°C + 273 = 315\,K$

Apply the equation for conversion factors

For comparison purposes, the Ideal Law equation can also be written as

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

$= 12dm^3 \times \frac{315}{295} \cdot \frac{\cancel{K}}{\cancel{K}} \times \frac{1.40 \times 10^4}{10^2} \cdot \frac{\cancel{KP_a}}{\cancel{KP_a}}$

$$= 12 \times \frac{315}{295} \times \frac{1.40 \times 10^4}{10^2}\ dm^3$$

$$= 12 \times \frac{315}{295} \times 1.40 \times 10^4 \times 10^{-2}\ dm^3$$

$$= 12 \times \frac{315}{295} \times 140\,dm^3 = 1793.89\,dm^3$$

$$= 1.79 \times 10^3\ dm^3$$

Result: The fire-fighter's air tank will provide $1.79 \times 10^3\ dm^3$ of air in the hot building.

Q59. Example 26: Calculate the volume that 6.30 moles of CO_2 gas occupy at 23°C and 550 KP_a pressure.

Solution:

$V = ?$ $\quad R = 8.31 KP_a\ dm^3\ mole^{-1} K^{-1}$

$n = 6.30\,moles$

$T = 23°C + 273 = 296\,K$

$P = 550\,KP_a$

Now according to the Ideal gas law equation

$VP = nRT$

$$= \frac{6.30 \times 8.31 \times 296}{550} \frac{mole^{+1}\ KP_a\ dm^3\ mole^{-1} K^{-1} \times K^{+1}}{KP_a}$$

$V = 28.2 dm^3$

Result:

6.30 moles of CO_2 occupy a volume of 28.2 dm^3 at 23°C and 550KPa pressure.

Activity For Student

A person can inhale a maximum of 0.115 moles of air per breath. Calculate the maximum volume of air, a person can inhale in one breath if the atmospheric pressure is 100 KP_a and the person's body temperature is 37°C.

Solution: $n = 0.115\,moles$

$V = ?$

$P = 100\,kPa = \frac{100000}{101325} = 0.986\,atm$

$T = 37\,°C + 273 = 310\,K$

According to the general gas equation

$$VP = nRT$$

$$V = \frac{nRT}{P} = \frac{0.115 \times 0.0821 \times 310}{0.986} = 2.96\,dm^3$$

Q60. Example 27: If a 0.75 dm^3 cylinder holds 6.5 moles of N_2 gas at 23°C, calculate the Pressure in the cylinder.

Solution:

$V = 0.75\ dm^3$

$N = 6.5$ moles

$T = 23°C + 273 = 296\,K$

$R = 8.31\,KP_a\ dm^3\ K^{-1}\ mole^{-1}$

$P = ?$

According to the Ideal Gas Law equation

$$VP = nRT$$

$$P = \frac{nRT}{V}$$

$$= \frac{6.5\,moles \times 8.31\,KP_a\ dm^3\ K^{-1}\ mole^{-1} \times 296\,K}{0.75\,dm^3}$$

$$P = \frac{6.5 \times 8.31 \times 296}{0.75} = \frac{\cancel{Moles^{-1}} \times KP_a\ \cancel{dm^3}\ \cancel{K^{-1}}\ \cancel{M^{-1}}\ .\cancel{K^{-1}}}{\cancel{dm^3}}$$

$= 21318\,KP_a$ Now $101.325\,KP_a = 1\,atm.$

$$21318\,KP_a = \frac{1}{101.325} \times 21318.0\,atm$$

$= 210.392$ atm

Activity For Student

1. 130 cm^3 of a gas exerts a pressure of 750 mm of Hg at 20°C. Calculate its pressure if its volume is increased to 150 cm^3 at 35°C. (Ans: 683.28 mm)

Solution:

$V_1 = 130\ m^3$

$V_2 = 150\ m^3$

$P_1 = 750\ mm\ Hg$

$P_2 = ?$

$T_1 = 20\ °C + 273\ K = 293\ K$

$T_2 = 35\ °C + 273\ K = 308\ K$

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \quad or \quad P_2 = \frac{P_1V_1 \times T_2}{T_1 \times V_2}$$

$$P_2 = \frac{750 \times 130 \times 308}{293 \times 150} = 683.28$$

2. Find the total pressure exerted by 2 grams of ethane and 3 grams of CO_2 contained in a 5dm^3 vessel at 50°C.

$(R = 0.082\,dm^3\,atm\,K^{-1}\,Mole^{-1})$

Solution:

$Mass\ of\ Ethane\ given = 2\ g$

$Number\ of\ moles\ of\ ethane = \frac{2}{30} = 0.066\ mol$

$Mass\ of\ CO_2\ given = 3\ g$

$$Number\ of\ moles\ of\ CO_2 = \frac{3}{44} = 0.068\ mol$$

$$P_{total} = P_{ethane} + P_{CO_2}$$

$$P_{ethane}\,V = nRT$$

$$P_{ethane} = \frac{nRT}{V}$$

$$P_{ethane} = \frac{0.066\ mol \times 0.0821 \times 323\ K}{5} = 0.350\ atm$$

$$P_{CO_2}\,V = nRT$$

$$P_{CO_2} = \frac{nRT}{V}$$

$$P_{CO_2} = \frac{0.068 \times 0.0821 \times 323}{5}$$

$$P_{CO_2} = 360\ atm$$

$$P_{total} = 0.350 + 0.360$$

$$P_{total} = 0.71\ atm$$

Q61. Example 28: (a) Determine the volume that 2.5 moles of Cl_2 gas occupies at S.T.P

(b) How many individual chlorine molecules are there in the 2.5 moles of Cl_2.

Solution:

(a) One mole of Cl_2 molecules at S.T.P = 22.4 dm^3

2.5 mole of Cl_2 molecule at S.T.P = 22.4 × 2.5

= 56 dm^3

(b) 1 mole of Cl_2 molecules at S.T.P = 6.022×10^{23} molecules

2.5 moles of Cl_2 molecule at S.T.P = $2.5 \times 6.022 \times 10^{23}$ molecules

= 1.51×10^{23} molecules

Q62. Example 29: A 2.60 dm^3 container was evacuated and its mass was found to be 655.53 g. The container was then filled with CO_2 gas at a pressure of 100 KP_a and a temperature of 24°C. The container and gas were found to have a mass of 660.13 g. Calculate the standard molar volume of CO_2.

Solution:

$V_{CO_2} = 2.60\ dm^3$

m_1 (mass of container) = 655.53g

m_2 (mass of container + gas) = 660.13g

$P_{CO_2} = 100\ KP_a$

$T_{Co_2} = 24°C + 273 = 297\ K$

m_{CO2} (mass of CO_2) = 660.13 – 655.53 = 4.60g

m_{CO_2} (mass of CO_2) = $660.13 - 655.53 - 4.60\ g$

Convert the gas (CO_2) to S.T.P by applying General Gas Equation

$(V_1)\ V_{CO_2} = 2.60\ dm^3$ $V_2 = ?$

$$\left.\begin{array}{ll}(P_1)P_{CO_2} = 100\,KP_a & P_2 = 101.3\,KP_a \\ (T_1)T_{CO_2} = 293\,K & T_2 = 273\,K\end{array}\right] \text{S.T.P}$$

$$\frac{V_1 P_1}{T_1} = \frac{V_2 P_2}{T_2} \quad \text{or} \quad V_2\,(dm^3) = \frac{V_1 P_1 T_2}{T_1 P_2} \cdot \frac{dm^3 \times \cancel{KP_a} \times \cancel{K}}{\cancel{KP_a} \times \cancel{K}}$$

$$= \frac{2.60 \times 100 \times 273}{293 \times 101.3} = 2.39\,dm^3$$

Conversion from volume (V_2) to the molar volume (V)

$$= \text{Molar volume} = \frac{\text{Molar mass}}{\text{Mass}} = \frac{44gm\,mole^{-1}}{4.6g} = 9.5652 \text{ mole}$$

Standard Molar Volume (Vm) = Molar Volume × volume of CO_2 at S.T.P.

$= 9.5652\,\text{moles} \times 2.36\,dm^3$

$= 9.5652 \times 2.36\,\text{Moles}.dm^3$

$= 22.6\,dm^3\,\text{Moles}$

(i) The accepted value for Molar value = 22.4 dm^3

(ii) $V_m = \frac{\text{Molar mass of } CO_2}{\text{Mass of } CO_2}$

Thus the conversion of volume of a gas to molar volume can be obtained by using the rates of Molar mass to mass.

Activity For Student

An empty container is filled with butane gas at 102KP_a and 20°C. It was then found to have a mass of 66.184gm. Determine the Molar mass of butane. (Ans: 272 dm^3)

Solution:

$P = 102\,kPa = \frac{102000}{101325} = 1.006$

$T = 20\,℃ + 273 = 293k$

$m = 66.184\,g$

$n = \frac{66.184}{58} = 1.141\,mol$

$PV = nRT$

$V = \frac{nRT}{P}$

$V = \frac{1.141 \times 0.0821 \times 293}{1.006}$

$V = 27.2\,dm^3$

Calculate the number of molecules in 10 grams of ammonia at 1.5 atmospheric pressure and 10°C. (Ans: 3.60×10^{23} molecules **)**

Solution:

$Mass\ of\ ammonia\ given = 10\,g$

$Mola\ mass\ of\ ammonia = 17\,gmol^{-1}$

$Number\ of\ moles\ of\ ammonia = \frac{10}{17} = 0.588\,moles$

$1\ mole\ of\ ammonia\ contains = 6.02 \times 10^{23}\ molecules$

$0.588\ moles\ of\ ammonia\ contains = 0.588 \times 6.02 \times 10^{23}\ molecules$

$0.588\ moles\ of\ ammonia\ contains = 3.60 \times 10^{23}\ molecules$

Q63. Describe the effect of heating gases to extremely high temperature.

Ans: Effect of Heating Gases to Extremely High Temperature:

A fourth type of matter i.e. Plasma State is formed when an extremely high temperature is applied to a molecular gas. The molecular gas is first converted to Atomic gas and ultimately to ions.

$$\text{Molecular gas} \rightarrow \text{Atomic gas} \rightarrow \text{Ions.}$$

Resulting Plasma consists of ions, positive or negative. The negatively charged ions i.e. electrons are mostly called as Plasma.

EXERCISE

MULTIPLE CHOICE QUESTIONS

1. Choose the correct answer (MCQs)

i. When compressed hydrogen is allowed to expand rapidly, it causes

(a) Cooling (b) Heating
(c) Liquefaction (d) Solidification

ii. 760 torr are equal to Pascal.

(a) 760 Pascal (b) 101325 Pascal
(c) 1.01325 Pascal (d) One Pascal.

iii. The number of molecules of 4 g H_2 are _____ number of molecules of 56 g of N_2.

(a) Equal to (b) less than (c) Greater than (d) None of these.

iv. According to the Kinetic theory of gases, the molecular collisions are elastic. Such collisions cause

(a) No energy change (b) A small energy change
(c) High energy change (d) None of these.

v. What volume of gas would one mole of hydrogen occupy at S.T.P?

(a) 11.2 dm^3 (b) 22.4dm^3 (c) 33.6dm^3 (d) 44dm^3.

vi. According to Graham's Law of diffusion, the rate of diffusion of H_2 and O_2 has the ratio

(a) 1:4 (b) $1:\sqrt{4}$ (c) 4:1 (d) 2:32.

vii. Deep sea diver's tank contains

(a) 96% N_2 + 4% O_2 (b) 4% N_2 + 96% O_2

$1\ mole\ of\ ammonia\ contains = 6.02 \times 10^{23}\ molecules$

$0.588\ moles\ of\ ammonia\ contains = 0.588 \times 6.02 \times 10^{23}\ molecules$

$0.588\ moles\ of\ ammonia\ contains = 3.60 \times 10^{23}\ molecules$

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vii. Deep sea diver's tank contains

(a) 96% N_2 + 4% O_2 (b) 4% N_2 + 96% O_2

(c) 50% N_2 + 50% O_2 (d) None of these

viii. In S.I units, the value of R is

(a) 8.3413 Nm K^{-1} $mole^{-1}$ (b) 8.3143 Nm K^{-1} $mole^{-1}$

(c) 0.0821 dm^3 atm K^{-1} $mole^{-1}$ (d) None of these

ix. According to Kinetic Molecular theory K.E of molecules increase when they

(a) Are mixed with other molecules at low temperature

(b) Are frozen into solid

(c) Are condensed into solid

(d) Are melted from solid to liquid state.

x. Which gas is more ideal at S.T.P.

(a) SO_2 (b) H_2S (c) NH_3 (d) H_2.

xi. On heating direct conversion of a solid in to gas is called

(a) evaporation (b) sublimation (c) diffusion (d) boiling

xii. In which of the following are the particles the most disordered.

(a) water (b) steam at 100°C

(c) impure water at 102°C (d) water at 0°C

xiii. When steam condenses, the particles

(a) shrink to a smaller size (b) lose energy to their surroundings

(c) move further apart (a) vibrate about fixed positions.

xiv. The particles of a gas can be described as

(a) only moving outwards in direction

(b) vibrating about defined position

(c) rising upwards

(d) moving randomly in all directions.

xv. Which of these changes would speed up the rate of diffusion the most.

No	Mass of Particle	Temperature of Surroundings
a	Decrease	Decrease
b	Decrease	Increase
c	Increase	Increase
d	Increase	Decrease

xvi. Which of the following gases has the lowest density under room conditions

(a) CO (b) N_2 (c) Ne (d) NH_3 (e) O_2

xvii. Which statement provides the best evidence that matter may exist as tiny particles moving at "random motion".

(a) Many elements conduct electricity.

(b) Air can be readily compressed.

(c) A small mass of water produces a much larger volume of steam.

(d) If a bottle of ether is opened, the smell is quickly detected in all parts of the room.

xviii. Which of the following is an example of diffusion?

(a) Bubbles rising in a beaker of boiling water.

(b) Steam condensing on a cold window.

(c) Water spreading out on the surface of a table.
(d) Hot air rising above a burning candle.
(e) The spreading of the smell of flowers in a garden.

xix. When a large beaker full of N_2 is inverted over a porous pot containing CO, the water level in the specified apparatus does not change. What is the reason for this

(a) The two gases have equal density.
(b) Nitrogen is an un-reactive gas.
(c) Both gases are diatomic.
(d) Both gases are soluble in H_2O.
(e) The gas particles are too large to pass through porous pot.

xx. Which statement explains why the compounds propane and CO_2 diffuse at the same rate?

(a) They are both gases.
(b) Their molecules contain carbon.
(c) They have the same relative molecular mass.
(d) Both are denser than air.
(e) Both molecules contain covalent bonds.

Answers

i. a	**ii. b**	**iii. a**	**iv. a**	**v. b**	**vi. c**	**vii. d**
viii. b	**ix. d**	**x. d**	**xi. b**	**xii. b**	**xiii. b**	**xiv. d**
xv. b	**xvi. d**	**xvii. d**	**xviii. e**	**xix. a**	**xx. c**	

2. Write brief answers to the following:

i. What is Absolute Zero and Absolute Scale of temperature?

Ans. Absolute Scale of temperature:

The temperature scale at which $-273.16\,°C$ is the starting point of the scale is called Absolute Scale of temperature.

Absolute scale of temperature = (Temp in °C + 273.16)K

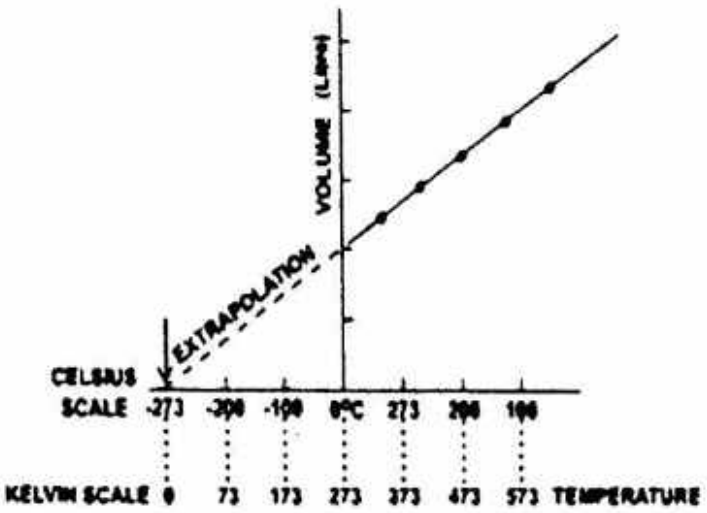
Absolute Zero.

According to critical definition of Charles's Law, at constant pressure, the volume of given mass of a gas increases or decreases by $\frac{1}{273}$ of its original volume at 0°C by 1°C rise or fall of temperature.

Let the volume of an ideal gas at 0°C by $V_o\ cm^3$

$$\text{Volume at } -1°C = V_o - \frac{1}{273} V_o$$

$$\text{Volume at } -273°C = V_o - \frac{273}{273} V_o = 0$$



Effect of decrease of temp on adefinite volume of a gas at constant pressure

Thus at (exact – 273°C), the volume of a given mass of a gas reduces to zero. A graph between volume and temperature gives a straight line and cuts the temperature axis at –273.15°C.

Thus the temperature at which the given volume of a gas reduces to zero is called Absolute zero. For this reason a new temperature scale has been developed. This new scale is called Absolute or Kelvin scale. It starts from –273.15°C which is marked as zero Kelvin.

– 273.15°C = Zero Kelvin

This is equal for all other gases as well and is the lowest possible in the gaseous state but actually it does not happen because all the gases liquify or solidify before they reach this temperature (– 273.15°C). This temperature is considered as the lowest possible temperature.

ii. Give any four applications of Plasma State.

Ans. Applications of Plasma:

Following are the applications of plasma.

i. They are used in the process of semiconductors, plasma, sterilization of some medical products, lamps and lasers.
ii. It is use in fluorescent bulbs
iii. It is use in Neon signs
iv. It is use to sterilize the operation theater instruments
v. It is use to control bacteria atmosphere pollution
vi. It is use in T.V screens
vii. It is use to pasteurize food

iii. What are the different units of gas constant (R.).

Ans. Following are the different units of general gas constant (R)

SI Unit:

$R = 8.3143\ Nm\ K^{-1}\ mole^{-1}$

$R = 8.3143\ J\ K^{-1}\ mole^{-1}\ (1Nm = 1J)$

In calories:

$As \quad 4.18J = Cal$

$$So,\ R = \frac{8.3143}{4.18} = 1.987\ Cal\ K^{-1}\ mole^{-1}$$

Standard temperature pressure:

$$R = 0.0824\ dm^3\ atm\ K^{-1}\ mole^{-1}$$

Pressure in ml and volume in dm^3:

$$62.4\ dm^3 mm\ of\ Hg\ K^{-1}\ mole^{-1}$$

Pressure in torr and volume in dm^3:

$$62.4\ dm^3\ torr\ K^{-1}\ mole^{-1}$$

When pressure is in torr and volume in cm^3:

$$62400\ cm^3\ torr\ K^{-1}\ mole^{-1}$$

iv. Define the Translational, Rotational and Vibration Energies.

Ans: i. Translational Energy:

The motion imparted to the gaseous molecules due to their motion in all possible directions is called translational motion and the energy as **kinetic translation energy**. In this case the entire molecules move from place to place.

ii. Rotational Energy:

The motion imparted to the gaseous molecules as a result of net angular momentum about their centre of gravity is called rotational motion and the energy as kinetic **rotational energy**. In this case the molecule spins like a propeller.

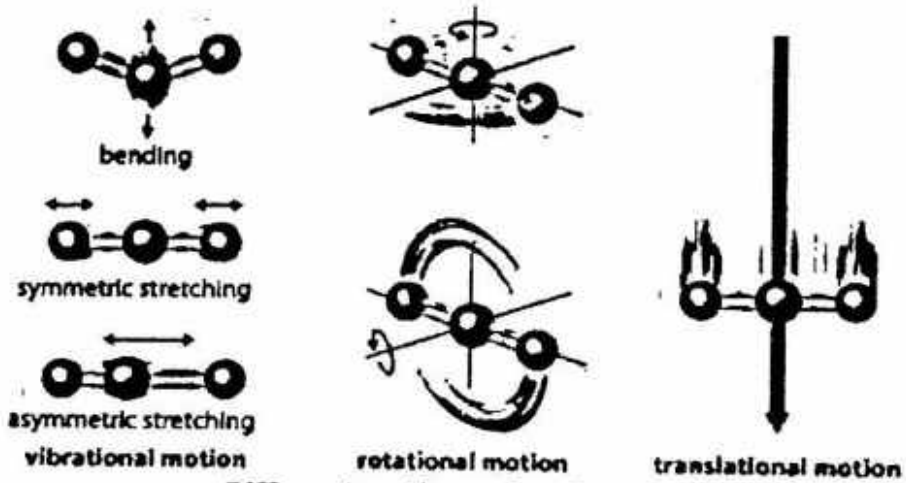
iii. Vibrational Energy:

The motion imparted to the gaseous molecules due to the oscillations is called vibrational motion and the energy as kinetic **vibrational energy.** In this case the molecules vibrate back and forth about the same fixed location.

In such an oscillating system, there is a continuous interchange between vibrational kinetic energies and potential energies. As a result both K.E and P.E is possessed by the gaseous molecules.

Thus, Total K.E = K.E (Vib.) + P.E (Vib.)

The vibrational motions in polyatomic molecules (like CO_2, NH_3 etc.) are quite complicated.



Different motions of molecules

v. Define Pressure. How will you explain it in different units?

Ans: Gaseous pressure and its various units:

A force being exerted over a unit area is called Pressure.

Explanation:

When a balloon is inflated, its walls expand. The balloon walls tend to collapse due to the squeezing of air trapped inside.

However this must be counterbalanced by a force exerted by the trapped gas. This force is called Pressure.

Mathematically:

So, $P = \frac{F}{A}$ where P is measured in Newton Force in joules and the area in m^2.

A Pascal (Pa) is defined as the force of one Newton (N) spread over an area of $1m^2$.

Units of Pressure:

i. SI Units:

SI unit for pressure is Pascal (Pa), in S.I. system, unit of force is Newton (N) and the unit of area is metre square (m^2). Thus Pascal is one Nm^{-2}.

1 atm	= 101325Pa =101325 Nm^{-2}
1 atm	= 14.7 Psi (Pounds per square inch)
1 atm	= 101325Pa
1 atm	= 101.325 KPa
1atm	= 760 torr = 760mm of Hg
1 J	= 1Nm =10^7 ergs = Kgm^2s^{-2}
1 Cal	= 4.18J

vi. Define Charles Law. How will you derive Absolute Zero with its help?

Ans. "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."

From his preliminary investigations, Charles's observed that the gases such as H_2, CO_2, O_2 etc. expanded equally on heating from 0 to 80°C at constant pressure.

Derivation of critical form of Law

Suppose the volume of a gas at 0°C = V_o

Volume at 1 °C $= V_o + V_o\frac{1}{273}$

Volume at 2 °C $= V_o + V_o\frac{2}{273}$

Volume at t °C $= V_o + V_o\frac{t}{273}$

$$V_t = V_o[1 + \frac{t}{273}]$$

$$= V_o[\frac{273 + t}{273}]$$

$$= V_o[\frac{T}{273}] \quad \text{Where } t\,°C + 273 = T \text{ (Absolute temperature)}$$

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

Thus the Law states that the volume of a definite quantity of a gas at constant pressure is directly proportional to the absolute temperature.

Again $V \propto T$ or $V = KT$, where T is the temperature on absolute scale. Thus the temperature on degree centigrade or Celsius scale can be converted to the absolute scale or Kelvin scale (after the name lord Kelvin 1824-1907) by adding 273.15.

Thus $K = t\,°C + 273.15$

Now $V = KT$ or $\frac{V}{T} = K$

Thus the Law may also be defined as, "the ratio between volume and temperature is always a constant quantity."

So $\frac{V_1}{T_1} = K$ and $\frac{V_2}{T_2} = K$

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

Where V_1 and T_1 are initial volume and absolute temperature. V_2 T_2 are final volume and absolute temperature.

Absolute Zero:

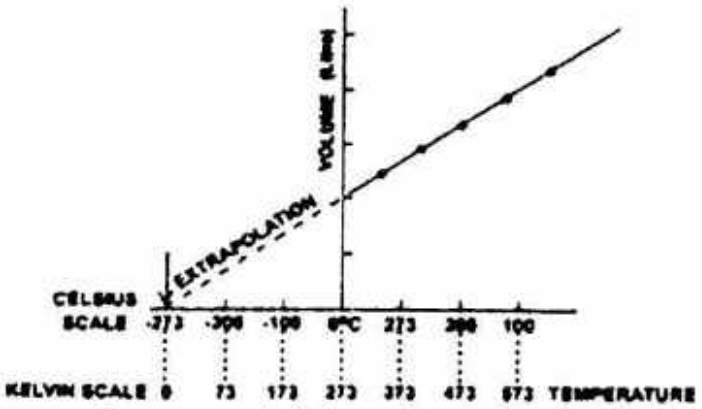
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Volume at $-273°C = V_o - \frac{273}{273} V_o = 0$

Thus at (exact $-273°C$), the volume of a given mass of a gas reduces to zero. A graph between volume and temperature gives a straight line and cuts the temperature axis at –273.15°C.



Effect of decrease of temp on a definite volume of a gas at constant pressure

Thus the temperature at which the given volume of a gas reduces to zero is called Absolute zero. For this reason a new temperature scale has been developed. This new scale is called Absolute or Kelvin scale. It starts from –273.15°C which is marked as zero Kelvin.

– 273.15°C = Zero Kelvin

This is equal for all other gases as well and is the lowest possible in the gaseous state but actually it does not happen because all the gases liquify or solidify before they reach this temperature (– 273.15°C). This temperature is considered as the lowest possible temperature.

vii. Define Avogadro's Law. What is its significance?

Ans. According to this Law, **"Equal volumes of all gases at same temperature and pressure must contain equal no of molecules".**

e.g. $1\ mole\ of\ H_2\ (2.016g) = 6.02 \times 10^{23}\ molecules$

$1\ mole\ of\ N_2\ (28g) = 6.02 \times 10^{23}\ molecules$

$1\ mole\ of\ O_2\ (32g) = 6.02 \times 10^{23}\ molecules$

Significances of Avogadro's Law:

Avogadro's law can be applied to determine:

1. Mass **2.** Volume **3.** Number of molecules of a gas.

(a) Exact No. of atoms or molecules as calculated by spectrometer are 6.02×10^{23}. This is called Avogadro's No. This help to calculate No. of atoms or molecules in any substance.

(b) 1 mole of any gas at S.T.P occupies volume = 22.4. Therefore if one have some gas at S.T.P. We can calculate its volume number of atom or number of molecules using Avogadro's Law.

viii. What is the difference between an Ideal gas and a real gas?

Ans: Difference between ideal and non ideal gases:

Ideal Gas	Non Ideal Gas
i. A gas which obeys the general gas equation ($PV = nRT$) and all the gas laws especially Boyle's Law and Charles's Law is called an Ideal gas.	**i.** The gases which do not obey general gas equation and all the gas laws strictly are called non-ideal gases.
ii. There are no attractive or repulsive forces among the gas molecules.	**ii.** There are attractive or repulsive forces among the gas molecules.
iii. Actual volume of gas molecules are negligible as compared to total volume of the gas.	**iii.** Actual volume of gas molecules are not negligible as compared to total volume of the gas.
iv. No such gas is exist in reality.	**iv.** All the gases are non ideal gases. E.g. O_2, H_2, N_2, CO_2 etc.

ix. How will you explain the relationship between Pressure and Mole of a gas with the help of Dalton's Law of Partial Pressures?

Ans. Relationship between pressure and no of moles of a gas:

If n_A, n_B and n_C are the number of moles of the gases and P_A, P_B and P_C be their partial pressure respectively.

Then, according to the general gas equation.

$PV = nRT$ or (4)

For gas A

$$P_A = \frac{n_A RT}{V} \quad (1)$$

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

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

Dividing equation (1) by (4)

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

Where n_t (total no. of moles of all the gases) = $n_a + n_b + n_c$

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

$$\frac{P_A}{P_t} = \frac{n_A \cancel{R} \cancel{T}}{\cancel{V}} \times \frac{\cancel{V}}{n_t \cancel{R} \cancel{T}}$$

$$\frac{P_A}{P_t} = \frac{n_A}{n_t} \Rightarrow P_A = \frac{n_A}{n_t} \times P_t \Rightarrow P_A \propto n_A$$

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

In general

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

x. How will you derive the General Gas Equation with the help of Boyle's Law and Charles's Law?

Ans: Derivation of Ideal Gas Equation:

The ideal Gas Equation is a combination of three Laws:

1. Boyle's Law **2.** Charles's Law **3.** Avogadro's Law

Now according to Boyle's Law

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

According to Charles's Law

$$V \propto T \quad \text{(at constant.....P and n).....} \quad \textbf{(2)}$$

According to Avogadro's Law

$$V \propto n \quad \text{(at constant.....P and T).....} \quad \textbf{(3)}$$

Where n is the number of moles of the gas.

Combining equations (1), (2) and (3), we get

$$V \propto \frac{1}{P} \times T \times n$$

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

$V = R\frac{Tn}{P}$ or $PV = nRT$

Where R is a constant called General Gas Constant. This equation is called General Gas Equation or Ideal Gas or Ideal Gas Law Equation.

If n = 1 Then $PV = RT$ or $\frac{PV}{T} = R$

So $\frac{P_1 V_1}{T_1} = R$ (For first condition of gas)

and $\frac{P_2 V_2}{T_2} = R$ (For second condition of gas)

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

xi. Ice on a car windscreen will disappear as you drive along, even without the heater on. Explain

Ans. Speed of air increases the rate of evaporation:

As the car moves the speed of air increases due to which the rate of evaporation of ice increases. Therefore the ice on the windscreen disappears due to evaporation.

xii. When salt is placed on ice, the ice melts. Explain

Ans. When salt is added the vapour pressure of ice decreases due to solute-solvent interactions. Therefore the freezing point of ice decreases and it melts.

OR

In Ice, water molecules are linked through hydrogen bonds. When salt like NaCl is placed on ice, the cations and anions of salt attract and attach water molecules with them. Therefore ice starts to melt and is converted into water.

xiii. When a jar of coffee is opened, people in all parts of the room soon notice the smell. Use the Kinetic Theory to explain how this happened.

Ans. It is because of diffusion effect of gases.

Explanation:

On opening the jar the vapours of coffee come out into the atmosphere. According to kinetic molecular theory, molecules of any gas are always in random motion so they spread in whole atmosphere of room because of diffusion. That is why every person soon feels the smell of coffee.

xiv. Why diffusion in gases is faster than in liquids.

Ans. According to kinetic molecular theory the molecules of gases have high kinetic energy and have large empty spaces between them therefore they can move freely in all directions. That is why the rate of diffusing of gases is high.

In liquids they are small empty spaces due to relatively strong intermolecular forces therefore the motion of molecules in liquids are limited so the rate of diffusion in liquids are slower than that of gases.

xv. A bubble of methane gas rises from the bottom of the North Sea. What will happen to the size of the bubble as it rises to the surface?

Ans. When a bubble of a gas rises from the bottom of Northern Sea, due to high pressure of water at the bottom, its size is very small at the bottom of the sea. But when it comes to the surface, its size enlarges due to the decrease in the pressure on the surface of sea.

At a depth of 40 meters the pressure increases 5 times then the normal pressure in a sea.

xvi. Use the Kinetic Theory to explain the following:

(a) When you take a block of butter out of the fridge, it is quite hard. However after 15 minutes it is soft enough to spread.

(b) When you come home from college and open the door, you can smell the tea which is being made.

(c) A football is blown up until it is hard on a hot summer. In the evening, the football feels softer.

(d) A windy day is a good drying day.

Solution:

(a) When you take a block of butter out of the fridge, it is quite hard. However after 15 minutes it is soft enough to spread.

Ans. At low temperature kinetic energy of molecule is less and therefore the force of attraction between the molecules is high due to which the block of butter is quite hard when it took out of fridge.

But when we take out a block of butter from the fridge into a region of relatively high temperature it absorbs heat which increases the kinetic energy of molecules therefore the force of attraction between the molecules also decreases due to which the block of butter become softer after some time.

(b) When you come home from college and open the door, you can smell the tea which is being made.

Ans. According to kinetic molecular theory, molecules of a gas move in all direction. They collide with each other and with the wall of container and also because of diffusion we can smell the tea which is been made in kitchen.

(c) A football is blown up until it is hard on a hot summer. In the evening, the football feels softer.

Ans. In hot summer due to high temperature the kinetic energy of gas molecules increases which cause the expansion of gas inside the football due to which we feel it hard.

In evening the temperature of the surrounding is comparatively low due to which the kinetic energy of the gas molecules also decreases which case the contraction of gas inside the football. That is why it feels softer.

(d) A windy day is a good drying day.

Ans. When the moving molecules of air collide with the water molecules they transfer the kinetic energy to the water molecules, giving them enough energy to break their bonds and and therefore 'escape' and evaporate. The wind would quickly carry away and remove any water molecules .That is why a windy day a good drying day.

xvii. Haemoglobin transports oxygen from the lungs to the rest of the body. One molecule of haemoglobia combines with 1.53 cm^3 of

oxygen at 37°C and 0.987KPa, what is the molar mass of haemoglobin.

Solution:

$V = 1.53\ cm^3 = \frac{1.53}{1000} = 0.00153\ dm^3$

$T = 37\ °C = 37 + 273 = 310\ K$

$P = 0.987\ KPa = 0.987 \times 1000 = 987\ Pa$

$R = 8.3145\ JK^{-1}mole^{-1}$

Molar mass of haemoglobin = M = ?

Molecular mass of 1 molecule of haemoglobin $C_{2952}H_{4664}O_{832}N_{812}S_8Fe_4$

$= 12 \times 2952 + 1 \times 4664 + 16 \times 832 + 14 \times 812 + 32 \times 8 + 56 \times 4$

$= 35424 + 4664 + 13312 + 11368 + 256 + 224 = 65248\ gmol^{-1}$

$= \frac{65248}{1000} = 65.248\ kgmol^{-1}$

$d = \frac{m}{V} = \frac{65.248}{0.00153} = 42645.75\ kgdm^{-3}$

Now, M (average molar mass) $= \frac{dRT}{P}$

$M = \frac{4264.5 \times 8.3145 \times 310}{987} = \frac{109919207}{987} = 1113666.98\ kgmol^{-1}$

xviii. James bought a huge balloon at the fair on a hot summer day. Before he went to bed, he noticed that the balloon was much smaller. Suggest two possible ways for the change in the shape of the balloon.

Ans.

i. In evening the temperature of surrounding is relatively low due to which the kinetic energy of the gas molecules filled in balloon is also less. Thus the gas contracts and compressed. That is why the size of the balloon becomes smaller and it feels softer.

ii. In hot summer the temperature is high due to which the kinetic energy of the gas molecules is also high. Thus the gas molecules escaped through the small opening on the top of balloon. The size of the balloon also decreases due to effusion process. Thus due to the decrease in the amount of gas balloon changes its shape and contracts.

3. What are the main postulates of the Kinetic Molecular theory of gases?

Ans: Kinetic Molecular Theory:

The kinetic molecular theory of gases is based upon the following postulates:

i. Gases are composed of molecules:

Gases are considered to be composed to minute discrete particles called molecules.

ii. Random motion:

The molecules move randomly in straight lines until they collide with one another or with the walls of container.

iii. Elastic Collision:

When molecules collide with one another, the collisions are perfectly elastic i.e. the total kinetic energy remains constant.

iv. Same mass and same size:

The molecules of a gas are thought to be of the same mass and size but are different from gas to gas.

v. Pressure due to collision:

The pressure is produced due to the collisions of the molecules.

vi. Kinetic energy directly proportional to the absolute temperature:

The average kinetic energy of a molecule is directly proportional to the absolute temperature. It means that the higher the temperature, greater will be the molecular kinetic energy

vii. Behaviour at low pressure:

At relatively low pressure, the average distances between molecules are large as compared with molecular diameters.

Faulty assumptions of kinetic molecular theory:

viii. No force of attraction between gas molecules:

There are no attractive or repulsive forces between the molecules.

ix. Volume of gas molecules is negligible:

The volume occupied by the molecules is negligible as compared to the total volume of the container.

4. Relate temperature to the average K.E of the particles in a substance.

Ans: The increase of temperature increases the average kinetic energy of the molecules and vice versa.

Derivation of relationship between Kinetic Energy and Temperature:

The average kinetic energy of the gaseous molecules is re-disturbed with rise or fall of temperature. It can be explained with the help of kinetic equation of gases.

According to kinetic equation of gases,

$$PV = \frac{1}{3} mn\overline{c^2} \quad \text{..... (1)}$$

and $$K.E = \frac{1}{2} m\overline{c^2} \quad \text{..... (2)}$$

Again we consider $PV = \frac{1}{3} mn\overline{c^2}$

$$= \frac{2}{3} n\left(\frac{1}{2} m\overline{c^2}\right)$$

$$= \frac{2}{3} n(K.E) \quad \text{..... (3)}$$

Now consider one mole of gas. It will posses Avogadro's number (N_A) of molecules,

Then $n = N_A$

Therefore $PV = 2/3\ N_A\ (K.E)$ ..... (4)

According to the General Gas Equation

$$PV = nRT \quad \text{..... (5)}$$

For one mole of a gas, $n = 1$ then $PV = RT$

Comparing equations (4) & (5) we get,

$$\frac{2}{3} N_A (K.E) = RT \qquad (6)$$

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

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

$$K.E = KT, \text{ where } \frac{3R}{2N_A} = K, \text{ a constant quantity.}$$

or $K.E \propto T$

Conclusion:

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

5. (a) What are the Standard Temperature and Pressure? (S.T.P).

(b) What is the density in grams per dm^3 of SO_2 at 25°C and 300 mm of Hg pressure?

Ans: (a) What are the Standard Temperature and Pressure? (S.T.P).

Standard temperature and pressure:

Ideal gas equation enables us to calculate the effect of a change in temperature and pressure on the volume of a gas. One cannot compare volumes unless they are stated at the same temperature and pressure. Gas volumes are usually compared at 0°C (273K) and 1 atm. These co9nditions are referred to as standard temperature and pressure (STP).

(b) What is the density in grams per dm^3 of SO_2 at 25°C and 300 mm of Hg pressure?

Solution: d = ?

T = 25°C + 273 = 298K

P = 300 mm of Hg = $\frac{300}{760}$ atm = 0.395 atm

$M_{SO_2} = 64\ gmol^{-1}$

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

Now according to the General Gas Equation,

MP = dRT (Derived from General Gas Equation)

or $d = \frac{PM}{RT} = \frac{0.395 \times 64}{0.0821 \times 298} = 1.03\ g\ dm^{-3}$

6. (a) Kinetically how will you interpret the effect of temperature on gaseous molecules?

(b) How will you explain the gas pressure with the help of Kinetic Theory?

Ans. (a) According to kinetic molecular theory of gases the kinetic energy of the gas molecules is directly proportional to temperature.

$$K.E \propto T$$

Thus the increase in temperature increases the average kinetic energy of the gas molecules and vice versa. The increase in temperature increase the collisions between the gas molecules due to which intermolecular forces decreases and molecules move apart from each other. Thus the volume of the gas increases.

(b) The gas molecules move randomly in straight lines until they collide with one another or with the walls of container. When molecules collide with one another, the collisions are perfectly elastic i.e. the total kinetic energy remains constant.

The pressure is produced due to the collisions of the molecules. The K.E. of molecules increases with increase in temperature and decreases with decrease in temperature that is why pressure of a gas increases at high temperature and decreases at low temperature.

7. (a) Define pressure. How will you express the unit of pressure in:
(i) SI units (ii) CGS system
(b) Explain the effect of change in pressure on the volume of a gas, temperature remains constant. — Boyle's Lav .
(c) How will you verify it graphically?
(d) A sample of air occupies 1 dm^3 at room temperature and pressure. What pressure is needed to compress it so that it occupies only 100 cm^3 at that temperature?

Ans. (a) Pressure is defined as the force per unit area.

(i) SI Units:

SI unit for pressure is Pascal (Pa), in S.I. system, unit of force is Newton (N) and the unit of area is metre square (m^2). Thus Pascal is one Nm^{-2}.

1 atm	= 101325Pa =101325 Nm^{-2}
1 atm	= 14.7 Psi (Pounds per square inch)
1 atm	= 101325Pa
1 atm	= 101.325 KPa
1atm	= 760 torr = 760mm of Hg
1 J	= 1Nm =10^7 ergs = Kgm^2s^{-2}
1 Cal	= 4.18J

(ii) CGS system:

The barye is the unit of pressure in CGS system. Its symbol is Ba. It is equal to 1 dyne per square centimeter (dynes cm^{-2}) or one decipascal, or one microbar.

So, In CGS system pressure is measured in dynes cm^{-2}.

(b) According to Boyles Law volume of a gas decreases with increase in pressure at constant temperature.

Effect of Change in Pressure on the Volume of a Gas (Boyle's Law)

In the middle of the 17th century, Robert Boyle (1627-1691) and his assistant Robert Hooke (1635-1702) made many investigations about the relationship between Pressure and Volume of a gas. In 1662, Robert Boyle reported that with increase in pressure on a gas, it's volume decreases. This was stated as Boyle's Law. It states that the pressure of a fixed amount of a gas varies inversely with its volume, the temperature is maintained constant.

Mathematically, $P \propto \frac{1}{V}$ or $P = K \times \frac{1}{V}$ where K is proportionality constant. The value of K is different for different amounts of the same gas.

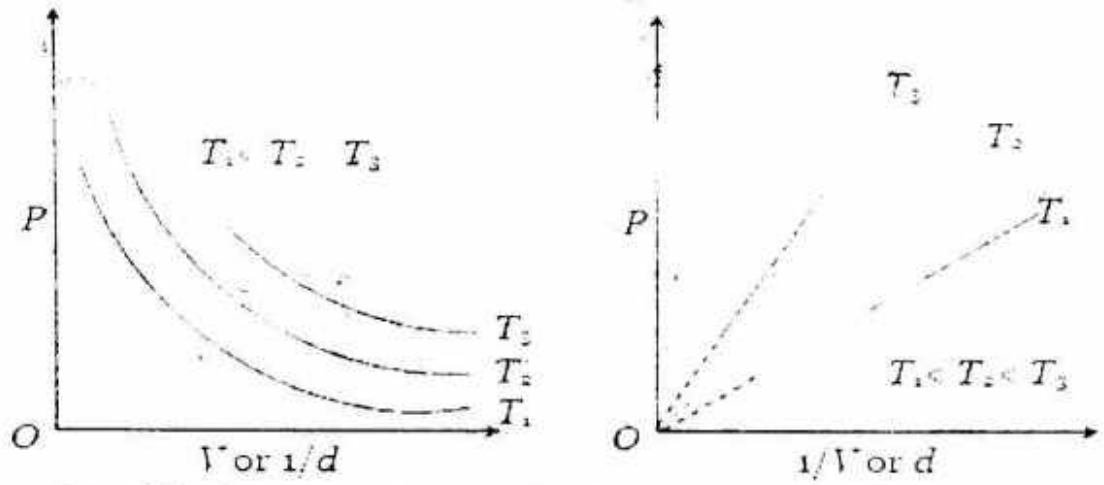
Again $P = K \times \frac{1}{V}$ or $PV = K$, thus the product of volume and Pressure remains constant provided the temperature is kept constant.

$P_1 V_1 = K$ (For gas at pressure P_1)
$P_2 V_2 = K$ (For gas at pressure P_2)
Therefore $V_1 P_1 = V_2 P_2$ where V_1 and P_1 are initial volume and Pressure while $P_2 V_2$ are their final Pressure and volume.

(c) Graphical Representation of Boyle's Law:

Isotherm:

When the Pressure of a gas is plotted against volume at different temperatures, we get a family of curves as shown in the figure. Each curve is a hyperbola with different values of K. Each curve is known as **Isotherm** (constant temperature plot). As the temperature is increased, the Isotherm goes away from both the axis. This is because at **higher temperature, the volume** of the gas is increased.



Graphical representation of different forms of Boyles Law

(d) Pressure = P_1 = 1 atm
Volume of gas 1 = V_1 = 100cm^3
Volume of gas 2 = V_2 = 100 cm^3
Required pressure = ?
Applying $P_1V_1 = P_2V_2$

$$1\ atm \times 1000\ cm^3 = P_2 \times 100\ cm^3$$

$$P_2 = \frac{1\ atm \times 1000\ cm^3}{100\ cm^3} = 10\ atm$$

8. How will you explain the effect of change in temperature on the volume of the gas.

(a) Give a graphic representation of the Charles's Law.

(b) To what temperature must a 1dm^3 sample of a perfect gas be cooled from room temperature in order to reduce its volume to 100 cm^3?

Ans: Charle's Law:

It states that at constant pressure, the volume of a given mass of a gas varies linearly with the absolute temperature of the gas.

Critical definition:

Charles's law can also be stated as "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."

From his preliminary investigations, Charles's observed that the gases such as H_2, CO_2, O_2 etc. expanded equally on heating from **0** to **80°C** at constant pressure.

Derivation of critical form of Law:

Suppose the volume of a gas at 0°C = V_o

Volume at 1 °C $= V_o + V_o\frac{1}{273}$

Volume at 2 °C $= V_o + V_o\frac{2}{273}$

Volume at t °C $= V_o + V_o\frac{t}{273}$

$$V_t = V_o[1 + \frac{t}{273}] \qquad = V_o[\frac{273 + t}{273}]$$

$= V_o[\frac{T}{273}]$ Where t °C $+ 273 = T$ (Absolute temperature)

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

Another definition:

Thus the Law states that the volume of a definite quantity of a gas at constant pressure is directly proportional to the absolute temperature.

Again $V \propto T$ or $V = KT$, where T is the temperature on absolute scale.

Thus the Law may also be defined as, "the ratio between volume and temperature is always a constant quantity."

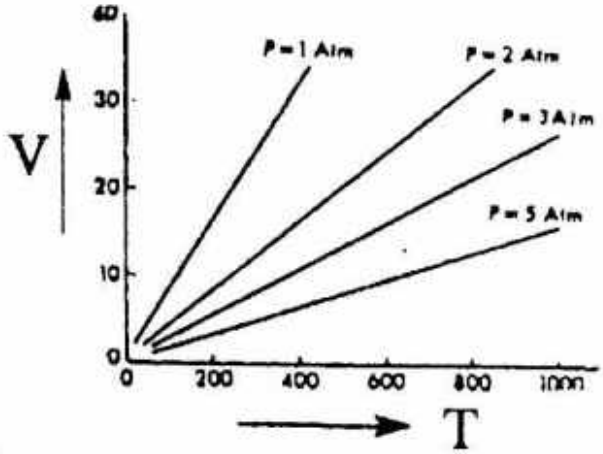
So $\frac{V_1}{T_1} = K$ and $\frac{V_2}{T_2} = K$

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

Where V_1 and T_1 are initial volume and absolute temperature. V_2 T_2 are final volume and absolute temperature.

(a) Graphical Representation of Charles's Law:

The equation for Charles's Law gives a straight line proved by plotting a graph between volume and temperature. Different straight lines are obtained with different pressures, which are a constant quantity. Each constant pressure line is called an "Isobar"



Graphical verification of Charles law

(b) Solution.

$$V_1 = 1\,dm^3 = 1000\,cm^3$$
$$T_1 = 25\,°C + 273 = 298\,K$$
$$V_2 = 100\,cm^3$$
$$T_2 = ?$$

According to Charles Law

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

OR

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

$$= 100 \times \frac{289}{1000} = 29.8K \qquad Ans$$

9. (a) How will you derive the Absolute Zero?

(b) What is the significance of Absolute Zero?

(c) How will you explain the Absolute Temperature scale on the bases of Charles Law?

(d) At 0°C and under a pressure of 1000 mm of Hg, a given weight of N_2 occupies a volume of one dm^3. At -100°C, the same weight of gas under the same pressure occupies a volume of 0.6313 dm^3. Calculate the Absolute Zero in °C.

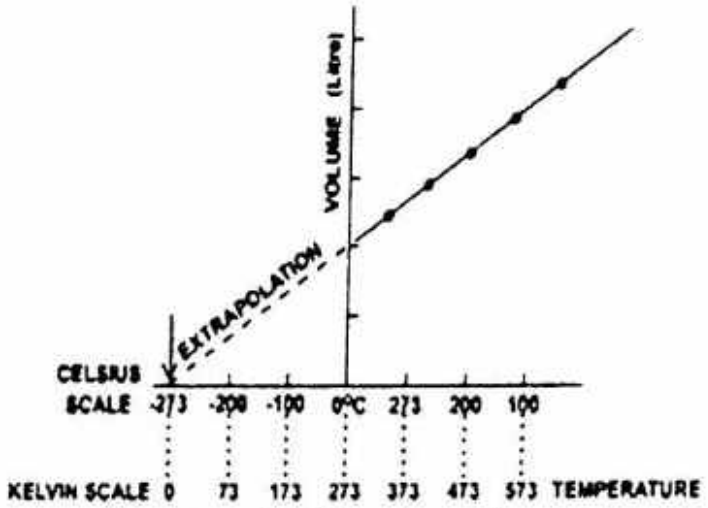
Ans. (a) Absolute Zero:

According to critical definition of Charles's Law, at constant pressure, the volume of given mass of a gas increases or decreases by $\frac{1}{273}$ of its original volume at 0°C by 1°C rise or fall of temperature.

Let the volume of an ideal gas at 0°C by $V_o\ cm^3$

Volume at $-1°C = V_o - \frac{1}{273} V_o$

Volume at $-273°C = V_o - \frac{273}{273} V_o = 0$



Effect of decrease of temp on adefinite volume of a gas at constant pressure

Thus at (exact – 273°C), the volume of a given mass of a gas reduces to zero. A graph between volume and temperature gives a straight line and cuts the temperature axis at –273.15°C.

Thus the temperature at which the given volume of a gas reduces to zero is called Absolute zero.

This is equal for all other gases as well and is the lowest possible in the gaseous state but actually it does not happen because all the gases liquify or solidify before they reach this temperature (– 273.15°C). This temperature is considered as the lowest possible temperature.

(b) Significance of Absolute Zero:

- It is not possible to have a gas at Absolute zero because it changes into liquid or solid states.
- At –273.15°C, the gas particles do not move and volume of gas becomes minimum.
- Absolute zero may be initialised to get the Kelvin scale of temperature.
- Zero Kelvin (-273.16°C) is called as Absolute zero which is the lowest possible temperature.

(c) According to critical definition of Charles's Law, at constant pressure, the volume of given mass of a gas increases or decreases by $\frac{1}{273}$ of its original volume at 0°C by 1°C rise or fall of temperature.

According to this statement the volume of the gas is zero at -273.15°C. Thus temperature at which the given volume of a gas reduces to zero is called Absolute zero and that temperature is -273.15°C or Zero Kelvin.

For this reason a new temperature scale has been developed. This new scale is called Absolute or Kelvin scale. It starts from –273.15°C which is marked as zero Kelvin. – 273.15°C = Zero Kelvin

(d) Solution:

Since at absolute zero, the volume of a gas is considered as zero, so we have following data.

$V_1 = 1\,dm^3$ $\quad t_1 = 0°C$

$V_2 = 0.6313\,dm^3$ $\quad t_2 = -100°C$

$V_o = 0$ $\quad t_o = ?$

The Charle's law can be written for Celsius scale as

$$V = kt + b$$

Since, according to Charle's law $V = kT$
where T = *temperature on Kelvin scale.*
Since $T = t + 273.15$,
where t = *temperature on Celsius scale.*
Put in above equation
$V = k(t + 273.15)$ $\quad or\ V = kt + k \times 273.15$ $\quad or\ V = kt + b$
Where 't' is the temperature in °C, k and b are constants. So, we can set up three equations for three conditions

i.e $\quad V_1 = kt_1 + b \qquad$ (i)
$\quad V_2 = kt_2 + b \qquad$ (ii)
$\quad V_o = kt_o + b \qquad$ (iii)

Put the values of V_1 and t_1 in Eq. (i), we have

$$1 = k(0) + b$$
$$or \quad b = 1$$

Put the values of V_2, t_2 and b in Eq. (ii), we have

$$0.6313 = k(-100) + 1$$
$$or\ k = \frac{0.6313 - 1}{-100} = 3.687 \times 10^{-3}$$

So, put the values of V_o, b and k in Eq. (iii), we have

$$0 = 3.687 \times 10^{-3} \times t_o + 1$$
$$t_o = \frac{0 - 1}{3.687 \times 10^{-3}} = -271.22\ °C$$

10. (a) Define and Explain Avogadro's Law, How does it help to determine
(i) Mass (ii) Volume (iii) Molecules of the gas
(b) Can we derive the Avogadro's Law from liquids and solids?

Ans. According to Avogadro's Law, "Equal volume of gases at same temperature and pressure have equal no of molecules. And at S.T.P, **1 mole of** a gas occupies a volume of 22.414 dm and contains 6.02×10^{23} molecules."

(i) Determination of Weight

Avogadro's Law may be used to calculate mass of a gas. Because according to Avogadro's Law

22.414 dm^3 of a gas at STP = 1 mole = $\frac{Mass\ in\ gram}{Molar\ mass}$

From this relation we can calculate the mass of a given substance

(ii) Determination of volume of a gas

According to Avogadro's Law
1 mole of any gas at STP = 22.414 dm^3
This relation may help to determine volume of any gas.

Example:

We know that
22.4dm^3 of any gas at S.T.P = 1 mole
22.4dm^3 of NH_3 at S.T.P = 17g
or 17g NH_3 = 22.4 dm^3

$$10g\ NH_3 = \frac{22.4}{17} \times 10 = 13.176\,dm^3$$

(iii) Determination of no. of molecules of a gas

According to Avogadro's Law

1 mole of any gas = 22.414 dm^3 at STP = 6.02×10^{23} moles

Example:

In the case of ammonia one mole of it contains 6.022×10^{23} molecules.

Now mass of NH_3 = 10g

Number of moles of NH_3 $= \frac{\text{mass}}{\text{molecular mass}} = \frac{10}{17}$

= 0.588 moles

One mole of NH_3 at S.T.P = 6.022×10^{23} molecules.

0.588 moles of NH_3 at S.T.P = $0.588 \times 6.022 \times 10^{23}$ molecules

= 3.54×10^{23} molecules

Now number of atoms = no of molecules × atomicity (Number of atoms in one molecule of NH_3)

= $3.54 \times 10^{23} \times 4$

= 14.16×10^{23} atoms

(b) Can we derive the Avogadro's Law from liquids and solids?

Ans. Avogadro's law is a law for gases. It is not applicable on liquids and solids because according to this law:

"Equal volumes of all gases at the same temperature and pressure must contain equal number of molecules."

11. (a) How will you derive the Ideal or General Gas Equation with the help of Boyle's Law, Charles Law and Avogadro's Law?

(b) How will you determine

(i) Molecular mass of the gas.

(ii) Density of the gas by Ideal Gas Equation.

(c) A gas that behaves ideally has a density of 1.92g dm^{-3} at 150 KPa and 298K. What is the molar mass of the sample? (Ans: = 31.73 g $mole^{-1}$)

Ans. Derivation of Ideal Gas Equation

The ideal Gas Equation is a combination of three Laws:

1. Boyle's Law 2. Charles's Law 3. Avogadro's Law

Now according to Boyle's Law

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

According to Charles's Law

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

According to Avogadro's Law

$$V \propto n \qquad \text{(at constant.....P and T)} \qquad (3)$$

Where n is the number of moles of the gas.

Combining equations (1), (2) and (3), we get

$$V \propto \frac{1}{P} \; Tn \; V \propto \frac{Tn}{P} \; V = R \frac{Tn}{P}$$

or $PV = nRT$

Where R is a constant called General Gas Constant. This equation is called General Gas Equation or Ideal Gas or Ideal Gas Law Equation.

(b) How will you determine

(i) Molecular mass of the gas.

Ans. (i) Molecular Mass of the Gas:

According to the general gas equation $PV = nRT$ ..... (1)

But n (no of moles) $= \frac{W}{M}$ Where W = mass of the gas

M = molecular mass of the gas

Putting the value of n in equation (1)

$$PV = \frac{W}{M} RT \qquad \text{..... (1)}$$

$$MPV = WRT \qquad \text{..... (2)}$$

$$\text{Or} \quad M = \frac{W R T}{P V} \qquad \text{..... (3)}$$

So molecular mass (M) can be calculated if P, V W and T are known.

(ii) Density of the gas by Ideal Gas Equation.

Ans. (ii) Density of the gas (g dm^{-3}):

According to the General Gas Equation, $PV = nRT$ ..... (1)

but $n = \frac{W}{M}$, Therefore $PV = \frac{W}{M} RT$ or $MP = \frac{W}{V} RT$

as $\frac{W}{V} = d$ (density),

so $MP = d\ RT$

or $d = \frac{MP}{RT}$ (2)

So knowing the values of P, T and M, the density can be calculated.

(c) A gas that behaves ideally has a density of 1.92g dm^{-3} at 150 KPa and 298K. What is the molar mass of the sample? (Ans: = 31.73 g $mole^{-1}$)

Ans. $Density\ of\ gas = d = 1.92\ g\ dm^{-3}$

$Pressure\ of\ gas = P = 150\ KPa = \frac{150000}{101325}$

$(\because 101325\ Pa = 1\ atm)$

$= 1.48\ atm$

$Temperature = T = 298K$

$$General\ gas\ constant\ =\ R\ =\ 0.0821\ dm^3\ atm\ K^{-1}\ mole^{-1}$$

$$Molar\ Mass\ of\ gas\ =\ M\ =?$$

$$Applying\ relation, d\ =\ \frac{PM}{RT}$$

$$M = \frac{dRT}{P} = \frac{1.92 \times 0.0821 \times 298}{1.48} = 31.73\ g\ mole^{-1}$$

12. (a) The value of Ideal Gas constant (R) can be calculated in three ways. Explain these three ways in detail.

(b) The density of air at 161.325 KPa and 298.15KPa is 1.59g dm^{-3}. Assuming that the air behaves as an ideal gas, calculate its molar mass. (Ans: = 24.447g $mole^{-1}$)

Ans. Numerical Value of R:

The value of R can be calculated in different ways as shown below:

(I) If one mole of a gas is taken at S.T.P (273 K, 1 atm), then the volume occupied by it is 22.4dm^3. Now according to the general gas equation

$$PV = nRT \qquad (1)$$

Where $V = 22.4\,dm^3$ $P = 1$ atm $n = 1$ mole $T = 273$ K Then R =?

From the equation (1)

$$PV = nRT$$

$$R = \frac{VP}{nT} = \frac{22.4\,dm^3 \times 1\,atm}{1\,mole \times 273\,K}$$

$$R = \frac{22.4 \times 1}{1 \times 273}\ dm^3 \times atm\ mole^{-1}\ K^{-1}$$

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

(II) If pressure is measured in mm of Hg or torr and V in cm^3 then

$$R = 0.0821\,dm^3 \times atm\ mole^{-1}\ K^{-1}$$

$$= 0.0821\,dm^3 \times 760mm\ of\ Hg\ mole^{-1}\ K^{-1}$$

$$= 62.4\,dm^3 mm\ mole^{-1}\ K^{-1}$$

$$= 62.4\,dm^3 torr\ mole^{-1}\ K^{-1}\ (one\ mm\ of\ Hg\ =\ one\ torr)$$

$$= 62.4 \times 1000\ cm^3\ torr\ mole^{-1} K^{-1}$$

$$= 62.4 \times 100\,cm^3\ torr\ mole^{-1}\ K^{-1}$$

$$= 62400\,cm^3 torr\ mole^{-1}\ K^{-1}$$

(III) In SI Units, pressure is expressed in Nm^{-2} and volume in m^3, then

$$V = 0.0224\,m^3\ (1dm^3 = 10^{-3}\ m^3)$$

$$P = 101325\,Nm^{-2}$$

$n = 1$ mole

$T = 273$ K

$R = ?$

$$R = \frac{VP}{nT} = \frac{0.0224\,m^3 \times 101325\,Nm^{-2}}{1\,mole \times 273\,K}$$

$$= \frac{0.0224 \times 101325}{1 \times 273}\ m^3\ Nm^{-2}\ mole^{-1}\ K^{-1}$$

$$R = \underline{8.3143\,Nm\ mole^{-1}\ K^{-1}}$$

Now we know that 1 Nm = 1 J

So, $R = 8.3143\,J\ mole^{-1}\ K^{-1}$

(b) The density of air at 161.325 KPa and 298.15KPa is 1.59g dm^{-3}. Assuming that the air behaves as an ideal gas, calculate its molar mass. (Ans: = 24.447g $mole^{-1}$)

Ans. $P = 161.325\ KPa = 161324\ Pa = \frac{161324}{101325} = 1.592\ atm$

$$T = 298.15\ K$$

$$d = 1.59\ dm^{-3}$$

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

Apply

$$M = \frac{dRT}{P} = \frac{1.59 \times 0.0821 \times 298.15}{1.592} = 24.48\ g\ mol^{-1}$$

13. (a) What is the difference between a real gas and an Ideal or perfect gas. What are the deviations of Ideal behaviour of Real Gases?

(b) What are the causes of deviations of real gases from their ideal behaviour. Explain these deviations at

i. Low temperature.

ii. High pressure.

Ans. (a) Difference between ideal and non ideal gases:

Ideal Gas	Non Ideal Gas
i. A gas which obeys the general gas equation ($PV = nRT$) and all the gas laws especially Boyle's Law and Charles's Law is called an Ideal gas.	**i.** The gases which do not obey general gas equation and all the gas laws strictly are called non-ideal gases.
ii. There are no attractive or repulsive forces among the gas molecules.	**ii.** There are attractive or repulsive forces among the gas molecules.
iii. Actual volume of gas molecules are negligible as compared to total volume of the gas.	**iii.** Actual volume of gas molecules are not negligible as compared to total volume of the gas.
iv. No such gas is exist in reality.	**iv.** All the gases are non ideal gases. E.g. O_2, H_2, N_2, CO_2 etc.

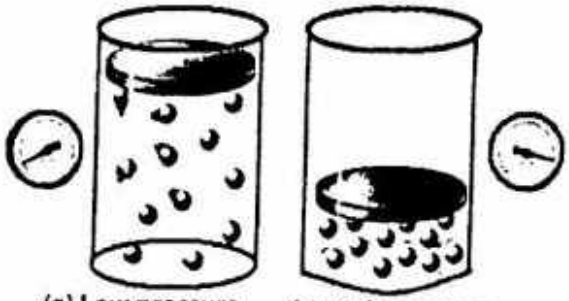
(b) (i) At lower temperature:

At high temperature the Kinetic energy of the gaseous molecules is very high. As a result the attractive forces between them are negligible. But when temperature is decreased, the Kinetic Energy of the molecules decreases. The intermolecular forces become significant. It means that the molecules come close to each other.

At a certain very low temperature the gases change into the liquid state. Therefore attractive forces between the gas molecules become significant near liquefying temperature. That is why the ideal gases deviate from their original behaviour at low temperature e.g. SO_2 liquefies at –10°C while H_2 at –252.7°C. Therefore attractive forces between SO_2 molecules cannot be considered negligible at room temperature. For this reason SO_2 gas shows non-ideal behaviour at room temperature as compared to H_2 gas.

(ii) At high pressure

At low pressure actual volume of gas molecules is very small as compared to the volume of the container. However, this volume does not remain negligible at high pressure. This can be understood from the following figure.



(a) Low pressure (b) High pressure

Explanation:

When pressure is one atm, individual volume of one mole of gas molecules say $32cm^3$ is negligible as compared to the total volume of gas $22400cm^3$. But if it is subjected to a pressure of 100atm, the volume of the gas is reduced to $224cm^3$. Under this pressure, individual volume of gas molecules ($32cm^3$)is not negligible as compared a volume of the gas ($244cm^3$).

14. How will you derive the Van der Waal's Equation keeping in view the General or Ideal Gas Equation (VP = nRT).

Ans. Van Der Waals Equation:

(CORRECTION FACTORS TO IDEAL GAS EQUATION)

The general gas equation is not applicable to real gases at all temperature and pressure. In order to make it applicable to real gases, a Dutch Scientist Van der Waals in 1873 made correction in volume and pressure of the gases and derived an equation known as Vander Waals equation.

Volume Correction:

Vander Waal thought that some of the volume of the vessel is occupied by the molecules of the gas and that volume is not available for the free movement of the molecules. Actually, we need the free volume of the gas and that is obtained when we subtract the individual volume of gas molecules from the volume of the vessel.

$V = V_{vessel} - V_{molecule}$

V is free volume

$V_{molecule} = nb$

where b is individual volume of one mole of the gas molecules and 'n' is total no. of moles of the gas molecules.

Then, $V = V_{vessel} - n\,b$ ..... (1)

Here 'b' is effective volume of gas molecules. Keep it in mind that 'b' is not the actual volume of gas molecule, but is roughly equal to 4 times their molar mass

volumes. If we have one mole of gas, then $b = 4V_m$ where, V_m is the actual volume of gas molecules for one mole of the gas.

Pressure correction:

The pressure which is exerted on the walls of the vessels is due to collisions of molecules of gas. Since there are some weak forces of attractions among the gas molecules, so the molecules cannot hit the walls of the vessels with that much force with which they should have been in the absence of these attractive forces.

It means that the pressure being observed on the walls of the vessels is a little bit lesser than the ideal pressure.

Therefore, $P_{observed} = P_{ideal} - P_{lessened}$

If the ideal pressure is denoted by P_i and the pressure lessened due to molecular attractions is denoted by P' then

$$P_{observed} = P_i - P'$$

and $$P_i = P_{observed} + P'$$

But the value of P' is inversely proportional to the density of the gas so,

$$P' \propto (\text{density})^2$$

$$p' = \frac{a n^2}{V^2}$$

where n = no of moles, V = volume and a = proportionality constant

$$P_i = p + \frac{a n^2}{V^2}$$

After the correction of volume and pressure, PV = nRT becomes

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

For 1 mole of gas, n = 1, then

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

15. (a) Define and explain the Dalton's Law of Partial Pressure. How does it explain the

(i) Relationship between Pressure and mole of a gas

(ii) Relationship between Pressure and Mole Fraction of component.

(iii) Give practical applications of the Dalton's Law of partial pressure.

Ans: Dalton's Law Of Partial Pressure And Its Applications:

An English Chemist John Dalton (1766-1844) showed in 1801 that "The total pressure exerted by a gaseous mixture is equal to the sum of partial pressures of each gas present in the mixture." The law is only obeyed, if the component gases do not chemically react with each other. Moreover the gases must behave ideally. The partial pressure is the pressure exerted by one component of the gaseous mixture.

Total Pressure $(P_t) = P_A + P_B + P_C$

Whereas P_A, P_B and P_C are the partial pressures of individual gases.

Particular Example:

Air is a mixture of non-reacting gases. Percentage of each gas in the gaseous mixture is N_2 (78.08%), O_2 (20%) Ar (0.93%), CO_2 (0.03%). Traces of Ne, He, Kr, H_2, along with various amounts of water vapours and pollutant (which can pollute the atmosphere) gases such as oxides of sulphur and Nitrogen are also found in air. The total pressure exerted by the air will be the sum of partial pressure of each gas.

(i) Relationship between Pressure and mole of a gas:

If n_A, n_B and n_C are the number of moles of the gases and P_A, P_B and P_C be their partial pressure respectively.

Then, according to the general gas equation.

$PV = nRT$ or (4)

For gas A

$$P_A = \frac{n_A RT}{V} \quad (1)$$

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

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

Dividing equation (1) by (4)

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

Where n_t (total no. of moles of all the gases) = $n_a + n_b + n_c$

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

$$\frac{P_A}{P_t} = \frac{n_A \cancel{R} \cancel{T}}{\cancel{V}} \times \frac{\cancel{V}}{n_t \cancel{R} \cancel{T}}$$

$$\frac{P_A}{P_t} = \frac{n_A}{n_t} \Rightarrow P_A = \frac{n_A}{n_t} \times P_t \Rightarrow P_A \propto n_A$$

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

In general

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

(ii) Relationship between pressure and mole fraction of a gas:

We know that

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

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

but $$\frac{n_i}{n_t} = X_i$$ where X_i is called mole fraction of the component

so, $P_i = P_t X_i$

Thus the partial pressure of any component is equal to the product of total pressure of all the components and mole fraction of that component.

(iii) Give practical applications of the Dalton's Law of partial pressure.

Applications of the Law of Partial Pressure

i. Collection of gases over water:

The Dalton's Law is particularly useful when a gas is generated and subsequently collected over water provided the gas is insoluble in water. The total pressure consists of the pressure of the water vapours in addition to the pressure of the gas that is generated. The pressure due to water vapour is called aqueous tension.

Thus $P_{Total} = P_{gas} + P_{Water}$ or $P_{gas} = P_{Total} - P_{Water}$

The gas contaminated with water may be called a moist gas.

ii. Respiration at higher altitudes:

The respiration process in living things depends upon the differences in partial pressure e.g. partial pressure of O_2 in the outside air is higher ($159g/cm^2$) than in the lungs where the partial pressure of O_2 is lower ($116g\ cm^{-2}$). However at high altitudes the pressure of oxygen decreases and the pilots may have uncomfortable breathing in a non-pressurised cabin where the partial pressure of O_2 is about $150g/cm^2$.

iii. Respiration by deep sea divers:

Deep sea divers breathe air under increased pressure. At a depth of 40 metres, the pressure increases 5 times than the normal pressure. Therefore regular air cannot be used in diver's tanks because the partial pressure of O_2 would be 795 mm of Hg (159x5 = 795 mm of Hg) in that case. Therefore deep sea divers use to breathe a mixture of 96% N_2 and 4% O_2 in the respiration tank, the scuba.

A scuba (self-contained breathing apparatus having oxygen used by divers underwater) contains compressed air to breath. As diver returns to the surface, it becomes hazardous for him if not handled properly. As the diver comes up the pressure of the surrounding water drops. Consequently, the compressed air in the lungs expands. Surfacing must be done very slowly so that the compressed air may escape out from the lungs without causing damage to them.

iv. How Deep sea divers respire on return from sea:

The percentage of N_2 in air is about 80%. If N_2 is present in diver's tank, the solubility of N_2 in blood can increase with increase in pressure in very deep sea. And blood of a diver may be saturated with N_2 during a dive. If the diver returns quickly and relatively at low pressure at the surface, he will face life threatening condition called the "bends". Deep sea divers must either use different solutions of gases such as He in O_2 or spend many hours in a decompression chamber after a dive. The pressure in a decompression chamber is slowly lowered over many hours. Nitrogen gas comes out of the blood and disposed off safely for normal breathing.

16. (a) By use of the Van der Waal's equation, find the temperature at which 3 moles of SO_2 will occupy a volume of 10 dm^3 at a pressure of 15 atm.

a= 2.253 atm dm^6 $mole^{-2}$, b= 0.0428 dm^3 $mole^{-1}$ (Ans:336°C)

Solution: $Here\ a = 2.253\ atm\ dm^6\ mole^{-2}$

$b = 0.0428\ dm^3\ mole^{-1}$

$n = 3$

$V = 10\ dm^3$

$P = 15\ atm$

$Putting\ their\ values\ in\ Vander - Waal's\ equation.$

$$\left(P + \frac{n^2 a}{v^2}\right)(v - nb) = nRT$$

$$\left(15 + \frac{(3)^2 \times 2.253}{(10)^2}\right)(10 - 3 \times 0.0428) = 3 \times 0.0821 \times T$$

$$(15.20277)(9.8716) = (0.2463)T$$

$$T = \frac{150.075664}{0.2463} = 609K$$

$$T = 609 - 273$$

$$T = 336^\circ C$$

(b) A certain gas occupies a volume of 6 dm^3 under a pressure of 720 mm of Hg at 25°C. What volume will this gas occupy under standard conditions? (Ans: = 5.2 dm^3)

Solution:

$V_2 = ?$

$V_1 = 6\ dm^3$

$P_1 = 720\ mm\ of\ Hg$

$P_2 = 760\ mm\ of\ Hg$

$T_1 = 25 + 273 = 298\ K$

$T_2 = 0^\circ C = 0^\circ C + 273 = 273K$

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

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

$$\frac{720 \times 6}{298} \times \frac{273}{760}$$

$$V_2 = 5.2\ dm^3$$

17. (a) Define and derive the Graham's Law of effusion and diffusion. What are its practical applications?

(b) The time required for a given volume of N_2 to diffuse through an orifice is 35 seconds. Calculate the molecular weight of a gas which requires 50 sec to diffuse through the same orifice under identical conditions.

(a) Define and derive the Graham's Law of effusion and diffusion. What are its practical applications?

Ans. Graham's Law of diffusion of gases:

Graham's Law of Diffusion and Effusion:

In 1831 Graham showed that "Rate of diffusion or effusion of a gas is inversely proportional to the square root of its density". Later in 1848 he himself showed that

"The rate of diffusion or effusion of a gas is inversely proportional to the square root of its molar mass."

Mathematical form of Graham's Law of Diffusion and Effusion:

Mathematically it can be shown as

$$r \propto \frac{1}{\sqrt{d}} \quad \text{(i) and} \quad r \propto \frac{1}{\sqrt{M}} \quad \text{.....(ii)}$$

where r = rate of diffusion,
d = density of the gas, and M = Molar mass of the gas

From equation (i) $r \propto \frac{1}{\sqrt{d}}$ or $r = k = \frac{1}{\sqrt{d}}$ .....(iii)

and from equation (ii) $r \propto \frac{1}{\sqrt{M}}$ or $r = k\frac{1}{\sqrt{M}}$ .....(iv)

where K = constant of proportionally.

Let us have two gases with r_1 and r_2 as their rates of diffusion and d_1, d_2, M_1, M_2 be their densities and molar masses respectively.

For two different gases, equation (iii) can be written as

$$r_1 = \frac{K}{\sqrt{d_1}} \quad \text{.....(v)} \qquad \text{and} \qquad r_2 = \frac{K}{\sqrt{d_2}} \quad \text{.....(vi)}$$

Dividing equation (v) by (vi),

$\frac{r_1}{r_2} = \sqrt{\frac{d_2}{d_1}}$ and may also be written as $\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}}$ because $d \propto M$

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

Applications of Graham's Law:

i. It can be used to determine the density of a gas.

ii. It can be used to determine the molar mass of a gas.

iii. It can be used to determine the rate of diffusion or effusion of a gas.

(b) The time required for a given volume of N_2 to diffuse through an orifice is 35 seconds. Calculate the molecular weight of a gas which requires 50 sec to diffuse through the same orifice under identical conditions.

Solution:

Let Time required for an unknown gas = t_x = 50 sec

Time required for N_2 gas = t_{N_2} = 35 sec

Since, all other conditions (e.g. Volume diffused, T etc.), are identical, therefore

Rate of diffusion of unknown gas $= r_x = \frac{1}{t_x} = \frac{1}{50} = 0.02\ sec^{-1}$

Rate of diffusion of N_2 gas $= r_{N_2} = \frac{1}{t_{N_2}} = \frac{1}{35} = 0.029\ sec^{-1}$

Molecular mass of nitrogen $= M_{N_2} = 28\ g\ mole^{-1}$

Molecular mass of unknown gas $= M_x = ?$

According to Graham's law of diffusion

$$\frac{r_{N_2}}{r_x} = \sqrt{\frac{M_x}{M_{N_2}}} \quad or \quad \frac{M_x}{M_{N_2}} = \left(\frac{r_{N_2}}{r_x}\right)^2$$

$$M_x = \left(\frac{r_{N_2}}{r_x}\right)^2 \times M_{N_2}$$

$$M_x = \left(\frac{0.029}{0.02}\right)^2 \times 28 = 2.1025 \times 28 = 58.87\ g\ mol^{-1}$$

18. Define Plasma and give its general characteristics.

Ans: Fourth State of Matter – THE PLASMA:

Plasma is composed of a mixture of neutral particles, positive ions and negative electrons. Plasma may be called as free electrons in metal.

Plasma is the fourth state of matter. It was identified by William Crooks in 1879. He obtained it by heating molecular gas changing into atomic form and then to ionic form at higher temperature.

(Molecular gas → Atomic gas → Ions)

At a higher temperature of **10^4 to 10^5 K**, electrons are removed from the atoms to form ions. Plasmas are studied in the atmosphere of the sun and stars and in the researches on nuclear fusion.

Plasma is composed of a mixture of un-ionized gas, free electrons and positively charged particles. Most of the universe contains matter in the Plasma state (about 99%).All shining stars are made up of Plasma. On our planet earth, Plasma does not exist in the **Free State**.

Charctristics of Plasma:

It is a substance in which many of the atoms or molecules are ionised effectively allowing charges to flow freely. It takes place at very high temperature. Plasma has the following properties:

a. Plasma consists of neutral particles, positive ions and negative electrons.
b. Plasma may also be called as free electrons in a metal.
c. Plasma is strongly influenced by both magnetic and electric force.
d. Plasma shows a characteristic glow depending upon the gas present in the discharge tube e.g. oxygen gives a red glow, hydrogen green and Nitrogen purple or pink glow.

19. (a) Define liquefaction of gases.

(b) How will you liquefy gases by Linde's method?

Ans. (a) Liquefaction of Gases:

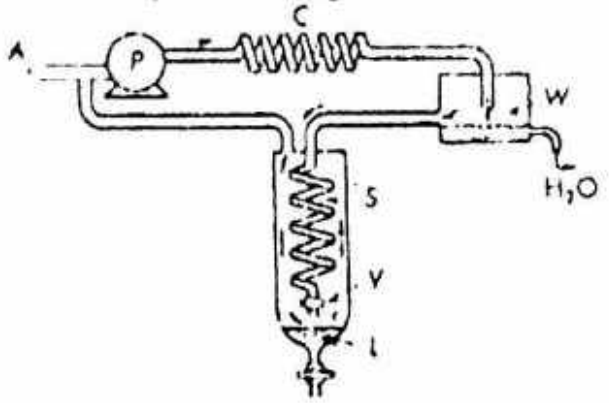
Principle:

The gases can be liquefied by Joule-Thomson's effect. The liquefaction of a gas requires high pressure and low temperature. When a highly compressed gas is allowed to escape out through a throttle (small hole), the temperature falls to such an extent that it changes into the liquid form.

At high pressure, the gaseous molecules come close to each other with the result that the molecular attractions increase. When it is allowed to escape through a nozzle (small hole) into a region of low pressure, the molecules move apart. In doing so, the intermolecular attractions must be overcome. This energy is taken from the molecules themselves. Therefore the gas is cooled. This process is repeated for many times until the gas completely changes into the liquid form.

(b) Linde's Method (1895):

According to Joule-Thomson's effect "When a highly compressed gas is allowed to escape out through a throttle" the temperature of the gas falls to such an extent, that it changes into the liquid form". This is the basic principle of the adiabatic expansion which is thermally isolated from its environment, Joule Thomson expansion and subsequent cooling of the air.



Working Steps:

- **(a)** Air is compressed to 200 atm approx.
- **(b)** Most of the water in the air condenses and is removed.
- **(c)** The heat generated as a result of compression is removed by passing the gas through coils C.
- **(d)** The dry gas is then passed through a copper spiral coil S.
- **(e)** It is then expanded to almost atmospheric pressure through a controlled valve V.
- **(f)** The issuing gas, cooled now due to Joule-Thomson's effect is passed over the copper spiral and cool further the incoming compressed gas.
- **(g)** The cycle is repeated several times.
- **(h)** The temperature of the expanding gas finally drops and the remaining air is liquefied.
- **(i)** The liquid air collects in the chamber L and can be drawn off.
- **(j)** Any uncondensed air is re-circulated.