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 ionized 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 behavior and properties of gases can be theoretically explained using the Kinetic Molecular Theory. This theory is a model that explains the behavior of gases using generalization 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 physicists 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 Clauses deauced 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 gas molecules under the given condition 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 c_2 and so on. Then,

$$c^{\bar{2}} = \frac{n_1c_1^2 + n_2c_2^2 + ...}{n_1 + n_2 + ...}$$

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

Root mean square velocity:

The square root of 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

1. Translational Motion

II. Rotational M

III. Vibrational Motion

A mono atomic molecule (0.9. 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 center 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 KE and PE is possessed by the gaseous molecules Thus,

Total KE = KE (Vib) + PE (Vib)

The vibrational motions in polyatomic molecules (like CO_2, NH_3etc) are quite complicated

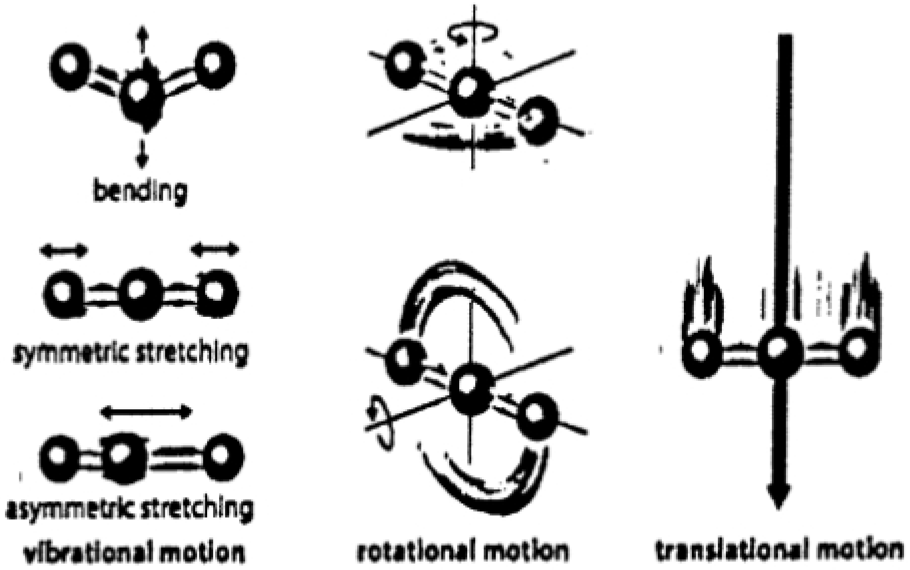


Figure 4.4 Different motions of molecules

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 = 98 °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: °C = TK - 273

(a) t° C = 340 TK - 273 67°C

(b)t °C = 200 TK - 273 = -73°C

(c) + C = 10 TK - 273 = -263°C

(d)t = 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) = $\frac{3}{2}$n RT

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

K.E = 3/2 × 2 × 8.314x10^{7} x 300 moles dm^{-3}× ergs mole^{-1} x K^{-1}x K

=7483.5 x 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:

SI Units:

SI unit for pressure is Pascal (Pa), in S.I. system, unit ot torce is Newton (N) and the unit of area is metre square (m). Thus Pascal is one Nm.

1 atm = 101325Pa = 101325 Nm^2

1 atm = 14.7 Psi (Pounds per square inch)

1 atm = 101325Pa

1 atm = 101.325 KPa

1 atm = 760 torr = 760mm of Hg

1J = 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 (1) Kilo Pascals (ii) atmosphere (iii) bars

Solution:

I. One Kilo Pascal = 1000 Pascals

or 1000 Pascals = one Kilo Pascal

10 Pa = $\frac{1}{1000}$ X 10 = 0.01 Kilo Pascal

II. One atmosphere = 101325 Pa

or 101325 Pa = one atm

10 Pa = $\frac{1}{101325}$ x 10 = 0.000098 atm

Activity for Students

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

Solution:

i. bar = 101325 pa

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

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

ii. 1.01325 bar = 101.325 kpa

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

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

iii. 1.01325 bar = 1 atm

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

35 bar = $\frac{1}{1.01325}$×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)

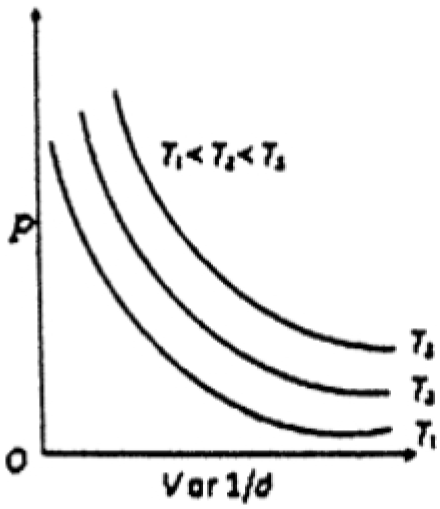
P_2 V2= K (For gas at pressure P_2)

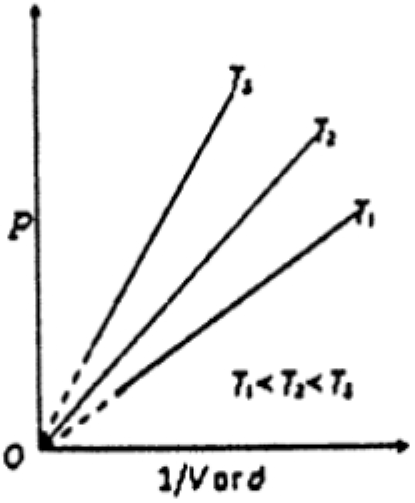
Therefore, $P_1 V_1 = P_2 V_2$ where $V_1$and P_1 are initial volume and Pressure while $P_2 V_2$ are their final Pressure and volume.

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.





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

Solution:

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

$P_1 = 1.80 \times 10^5\ Pa$ $P_2 = 115 \times 10^5\ Pa$

According to the Boyle's Law

$V_1\ P_1 = V_2\ P$

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

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

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

Q12. Example 6: At sea level, where the pressure was 775 torr, the gas in a balloon occupied 2m². 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 P_2 = 100 torr

$V_1 = 2m^3$ V_2= ?

According to the Boyle's Law

$V_1 P_1 = V_2 P_2$

$2m^3$ x 755 torr = V_2 × 100 torr

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

(a) P_1 = 755 torr P_2 = 10 torr

$V_1 = 2m^3$ V_2= ?

According to the Boyle's Law

$V_1 P_1 = V_2 P_2$

$2m^3$ x 755 torr = V_2 × 10 torr

Or
$$V_2 = \frac{2m^3 \times 755torr}{10torr}$$
$$= \frac{2 \times 755}{10} m^3 = 151m^3$$

Activity for Student

375cm^3 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 cm)

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

$$P_1V_1 = P_2V_2$$
$$375 \times 770 = V_2 \times 750$$
$$V_2 = \frac{375 \times 770}{750} = 385cm^3$$

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 various 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 observea 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 = $V_o + V_o\frac{t}{273}$

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

Where t °C + 273 = T (Absolute temperature)

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

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 α 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

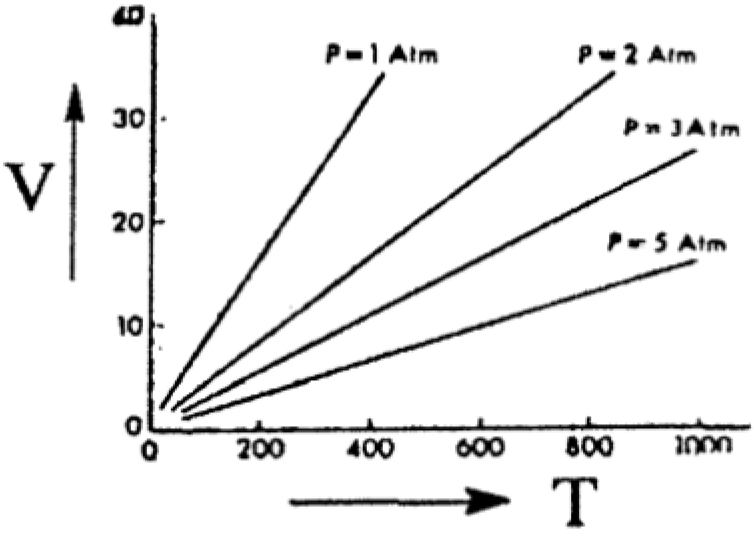
So, $\frac{V_1}{T_1} = K$ and $\frac{V_2}{T_2} = K$

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

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.

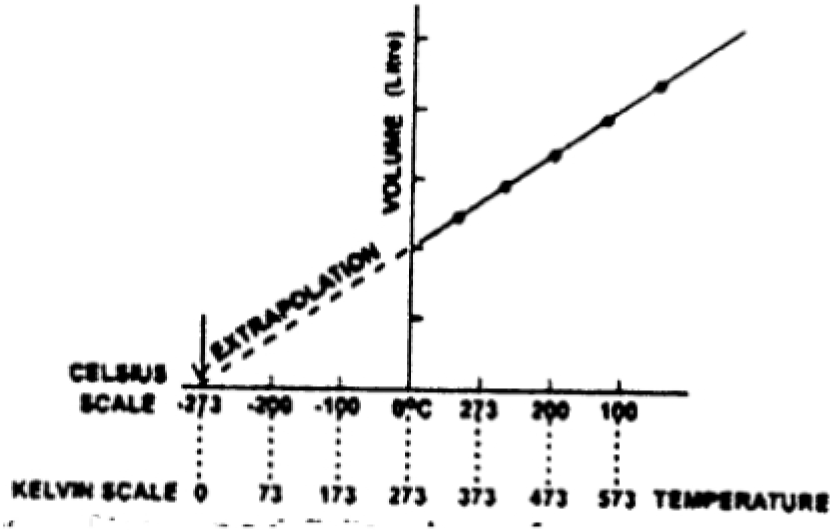
Ans: Absolute Zero:

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 liquefy 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 initialized 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 cm³ 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:

$V_1 = 50cm^3$ $V_2 = ?$

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

According to Charles's Law

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \text{ or } \frac{50cm^3}{288K} = \frac{V_2}{322} \text{ or } \frac{50 \times 323}{288} \times \frac{cm^3 \times K}{K}$$

$= 56cm^3$

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:

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

$T_1 = 17°C + 273$ $T_2 = 100°C + 273$

$= 290K$ $= 373K.$

According to Charles's Law

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \text{ or } \frac{125cm^3}{290K} = \frac{V_2}{373K} \text{ or } \frac{125 \times 373}{290} \times \frac{cm^3 \times K}{K}$$

Thus the new volume will be 161 cm^3

Activity for Student

A Volume-Temperature experiment which was performed on a sample of M 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 $V = 50\ dm^3$

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

$299.4\ K = T$

$t = 299.4 - 273$

$t = 26.40\ °C$

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 (28g) at S.T.P. = 22.4 dm^3 = 6.022×10^{23} molecules

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

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

1 mole of CO_2 (44g) at S.T.P. = 22.4 dm^3 = 6.022×10^{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^3at S.T.P. This volume of 22.4 dm^3 at STP is called Molar volume.

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

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

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

1 mole of CO_2 (44g) at S.T.P. = 22.4 dm^3 = 6.022 x 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' of 10g of ammonia at S.L.P.**

Solution:

i. **Number of molecules and atoms:**

According to Avogadro's Law, one mole of any gas at STP 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 x 10^{23} molecules

Now mass of NH_3 = 10g

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

= 0.588 moles

One mole of NH_3 at S.TP. = 6.022 x 10^{23} molecules

0.588 moles of NH_3 at STP = 0.588 x 6.022 10^{23} molecules

` = 3.54 x 10^{23} molecules

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

= 3.54 x 10^{23} x 4

= 14.16 x 10^{23}atoms

ii. **Volume in dm^3**

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 = 179

or 17g NH_3 = 22.4 dm^3

10g NH_3 = $\frac{22.4}{17}$ x 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 of Hydrogen (iii) 2.5 dm' of Nitrogen All gases are at S.T.P. (Ans: (i) 2.7 x 10^{22} (ii) 5.4 x 10^{22} (iii) 6.75 x 10^{22})

Solution:

i. 22.414 dm^3 of oxygen at STP = 1 mole of oxygen

= 6.023 X 10^{23}molecules

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

=2.7 x 10^{22} molecules

ii. 22.414 dm^3 of hydrogen at STP = 1 mole of hydrogen

= 6.023 x 10^{23} molecules

=6.023 x 10^{23}

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

5.4 x 10^{22} molecules

iii. 22.414 dm^3 of nitrogen at STP = 1 mole of nitrogen

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

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

$= 6.67 \times 10^{22}$

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

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

According to Charles's Law

$V \propto T$ (at constant... P and n)..............(2)

According to Avogadro's Law

$V \propto n$ (at constant. P and T) .... (3)

Where n is the number of moles of the gas.

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

$$V \propto \frac{1}{P} \times T \times n$$
$$V \propto \frac{Tn}{P}$$
$$V = R\frac{Tn}{P}$$
$$PV = nRT$$

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 $\frac{P_1V_1}{T_1} = R$ (For first condition of gas)

And $\frac{P_2V_2}{T_2} = R$ (For second condition of gas)

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

Q22. Example 10: A certain mass of gas occupies 1000 cm³ at 57°C and 723 mm of Hg. What will be the volume at S.T.P.?

Solution:

V_1 = 1000cm V_2 =?

P_1 = 723mm of Hg P_2= 760mm of Hg

T_1 = 57°C + 273 = 330TK T_2 =273TK

Now apply the general gas equation

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

$$V_2 = \frac{726 mmofhg \times 1000 cm^3}{330K} \times \frac{273K}{760 mmofhg}$$

$$V_2 = \frac{760 \times 100 \times 273}{330 \times 760} \times cm^3$$

V_2 =790.3 cm^3 So the volume of the gas at STP = 790.3cm^3

Activity for Student

Q.1 A given mass of gas occupies 850 cm^3 at 320K and 0.92×10^5 Nm^{-2} pressure. Calculate the new volume of the gas at STP (Ans: 660cm)

Solution:

V_1 = 850 cm 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.

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

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

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

$$V_2 = 658.410cm^3$$

Q.2 A container contains 2.5m^3 of Helium gas at 15°C and 98 Nm^{-1} pressures. What volume would this gas occupy at S.T.P? (Ans: 2.3cm')

Solution:

V_1 = 2.5 m^3 V_2=?

P_1 = 98 Nm^{-2} P_2 = 1 atm = 101325 Nm^{-2}

T = 15 °C + 273 K = 288 K T = 273 K

Now apply the general gas equation

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

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

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

$$V_2=0.0023m^{3}$$

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 (number of moles) $=\frac{W}{M}$, Where W = mass of the gas

M = molecular mass of the gas

Putting the value of n in

$$PV = \frac{W}{M}RT$$

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

or $$M = \frac{WRT}{PV} \qquad \text{............ (3)}$$

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 (dm^3):

According to the General Gas Equation PV = nRT .... (1)

But $n = \frac{W}{M}$, therefore, $PV = \frac{W}{M}RT$ or $PM = \frac{W}{V}RT$

As, $\frac{W}{M}$ = d (density).

So MP = d RT

Or $d = \frac{MP}{RT}$

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.

(1) 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. Now according to the general gas equation.

PV = nrt .... (1)

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

From the equation (1)

PV = Nrt

$$R = \frac{PV}{nT} = \frac{22.4dm^3 \times 1atm}{1mole \times 273K}$$
$$R = \frac{22.4 \times 1}{1 \times 273} dm^3 atm \times mole^{-1} K^{-1}$$
$$R = 0.0821 dm^3 atm \times mole^{-1} K^{-1}$$

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

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

=0.0821 dm^3 x 760mm of Hg mole^{-1} K^{-1}

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

= 62.4 dm^3torrmole^{-1} K^{-1} (one mm of Hg = one torr)

= 62.4 x1000 cm^3torrmole^{-1} K^{-1}

= 62.4 x 100cm torrmole^{-1} K^{-1}

=62400 cm^3torrmole^{-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 (1 dm^3 = 100 m^3)

P=101325 Nm^{-2}

n = 1 mole

T = 273 K

R = ?

$$R = \frac{PV}{nT} = \frac{0.0224 m^3 \times 101325 Nm^{-2}}{1 mole \times 273 K}$$
$$R = \frac{0.0224 \times 1}{1 \times 273} m^3 Nm^{-2} \times mole^{-1} K^{-1}$$
$$R = 0.0821 dmm \times mole^{-1} K^{-1}$$

Now we know that 1 Nm = 1J

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 = 8.31453 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)

$$M = \frac{1.29\ Kgm^{-3} \times 8.31453\ K^{-1}mole^{-1} \times 273.15K}{101325kgm^{-1}s^{-1}}$$

$$M = \frac{1.29\ \times 8.31453 \times 273.15}{101325} . \frac{Kgm^{-3} \times K^{-1}mole^{-1}K}{kgm^{-1}s^{-1}}$$

$$M = 0.0289Kgmole^{-1}$$

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 V_2= ? P_1 =720mm P P_2= 760mm

T_1 =25°C -273 = 298K T_2 = 273K

According to the General Gas Equation

or

$$P_1 = \frac{n_1RT}{V}$$

$$= \frac{6 \times 720 \times 298}{760 \times 273} \frac{dm^3 \times mmofHg \times K}{mmofHg \times K}$$

$$= 6.204dm^3$$

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

Solution:

No of moles of $CH_4 = \frac{mass}{molecular\ mass}$

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

T = 27°C + 273 = 300K

P = 2.5 atm

V = 2.46 dm^3

R =?

According to the General Gas Equation,

VP = nRT

$$R = \frac{VP}{nT}$$

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

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

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

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} = 17g $mole^{-1}$

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

Now according to the General Gas Equation

MP = dRT (Derived from General Gas Equation)

$$d = \frac{MP}{RT} = \frac{17gmole^{-1} \times 2.105atm}{0.0821dm^3atmmole^{-1}K^{-1} \times 273K}$$

$$d = \frac{17 \times 2.105}{0.0821 \times 273}$$

$$d = 1.598gdm^{-3}$$

Q32. Example 15: At 27°C a volume of 500 cm^3 of H_2measured 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:

$V_1 = 2dm^3$ $V_2 = 500cm^3 = \frac{500}{1000}$ $dm^3 = \frac{1}{2}dm^3$

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

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

For N_2:

$V_3 = 1000$ $cm^3 = 1$ dm^3

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

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

Total pressure = 0.1315atm+0.3945atm = 0.5260atm

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

Solution:

44 g of CO_2 = 1 mole

3g CO_2= $\frac{1}{44}$×3 = 0.068moles

Number of moles of CO_2 = 0.068 moles

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

Now according to the General Gas Equation

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

$$P_{CO_2} = \frac{0.068 mole \times 0.0821 dm^3 atm mole^{-1} K^{-1} \times 323K}{5dm^3}$$

$$P_{CO_2} = \frac{0.082 \times 0.068 \times 323}{5} = 0.678 atm$$

30g Ethane (C_2H_6) = 1 mole

2g Ethane (C_2H_6) = $\frac{1}{30}$x2 = 0.06 moles

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

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

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

P = 0.06 moles xR = $0.082dm^3atm\ K^{-1}\ mole^{-1}$ × 323K

= 0.06 × 0.082 x 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 (atm)	Volume (dm^3)	PV
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 behavior 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.

(1) Inter-molecular Forces of attraction:

These are no attractive or repulsive forces between the gas molecules

(2) Volume:

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

Explanation of deviations:

(1) 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.

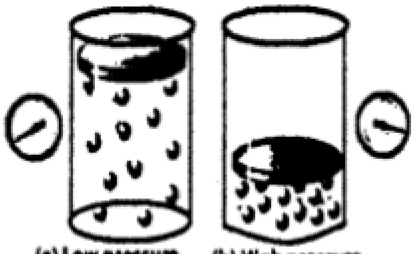
At a certain very low temperature the gases change into the liquid state. Therefore, attractive forces between the gas molecules become significant near liquefying temperature. That is why the ideal gases deviate from their original behavior at low temperature

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, gas shows non-ideal behavior at room temperature as compared to H_2 gas.

(2) 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

Suppose:

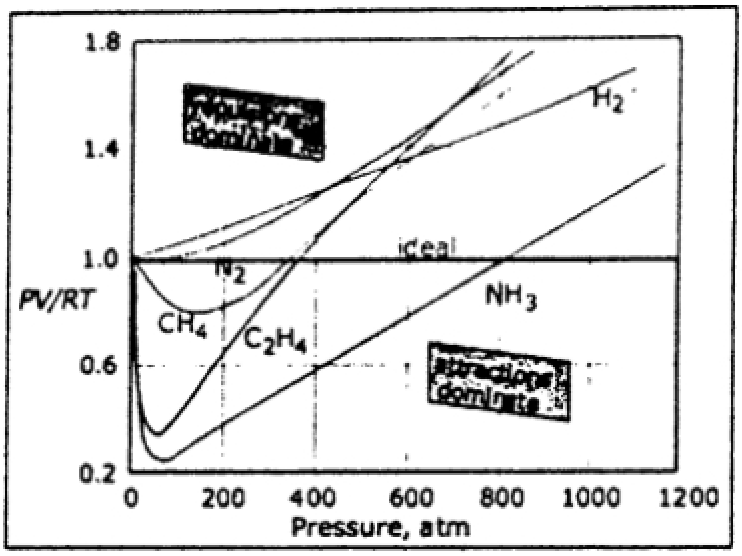
When pressure is one atm, individual volume of one mole of gas molecules say 32 cm is negligible as compared to the total volume of gas 22400 cm but if it is subjected to a pressure of 100 atm, the volume of the gas is reduced to 224 cm. Under this pressure, individual volume of gas molecules (32 cm')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

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

Ans: Graphical Representation:

For one mole of a gas, f a graph is plotted between $\frac{PV}{RT}$ and pressure thenfor an ideal gas, it must give a straight line the temperature remains at 0°C. But actually, it's seen that the gases do not give a straight line. They deviate from their original behavior 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 themolecules 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 bis individual volume of one mole of the gas molecules and n' is total number of moles of the gas molecules

Then, $V = V_{vessel} - nb$

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, a = proportionality constant

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

After the correction of volume and pressure. PV = nRT becomes

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

For 1 mole of gas, n = 1, then

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

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	0.0322
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
H2O	0.5536	0.0305
NH3	0.4225	0.0371
CH4	0.2283	0.0428
CC'2F2	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 mo1e $^{\cdot 1}$

b = 0.0428 dm^3 mo1e $^{\cdot 1}$

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=nRT/V

= 98.52atm

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

$$(P + n^2a/V^2)\ (V - nb) = nRT$$

Now dividing by (V-nb) on both sides Putting the following values in this equation

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

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

=$2.253dm^3$atmmole^{-1}, b=$0.0428dm^3$mole^{-1}

$$\text{IM x0.0821}dm^3 \text{mote}^{-1}\ K^{-1} \times 300K \qquad \text{IM x } 2.25dm^3 \text{atm}K^{-1}$$

$$0.25dm^3 - (\text{I M x } 0.0428dm^3 \text{mote}^{-1}) \qquad (0.25)^3dm^3$$

= 118.87 - 3.6048 = 115.2652 atm

Q40. Example18:EstimatethemolarvolumeofN2atSOOK and100atm by treating it as a Van der Waals' equation. (R = 0.0821 dm^3 atm K^{-1} mo1e $^{-1}$)

a = 1.39dm' atm Mo1e ^{2}

b = 3.913x10^{-2} dm3mo1e^{-1} **(Ans: Molar Volume = O.5 dm^3)**

Solution.

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

$$(P + n^2a/V)(V - b) = Rt$$

$$(p+\frac{a}{V^2})(v-b)=RT$$

$$(v-b)=\frac{RT}{P+\frac{a}{\frac{n^2r^2}{P^2}}}$$

Since PV=RT , V= $\frac{RT}{P}$, V^2= $\frac{n^2r^2}{P^2}$ putting in equation (i)

$$V=\frac{RT}{P+\frac{a}{\frac{n^2r^2}{P^2}}}+b$$

$$V=\frac{RT}{P+\frac{aP^2}{n^2r^2}}+b$$

$$V=\frac{0.0821x500}{100+\frac{1.39x100^2}{(0.0821)^2x(500)^2}}+3.913x10^{-2}$$

$$V=\frac{0.0821x500}{100+\frac{1.39x10000}{(0.0067x250000}}+3.913x10^{-2}$$

$$V=\frac{41.05}{100+\frac{13900}{1685.1025}}+3.913x10^{-2}$$

$$V=\frac{41.05}{100+8.23}+3.913x10^{-2}$$

V = 0. 4 + 3. 913 X 10^{-2}

V = 0.4 + 0.03913

V = 0.439

V=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 $(P_1)=P_A+P_B+P_c$

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 vapors 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 agas:

If and c are the number of moles of the gases and PA, P_8 and Pc be their partial pressure respectively.

Then according to the general gas equation

PV=nRT

For gas A:

$P_A=n_ART/V$

For gas B:

$P_B=n_BRT/V$

For gas C:

$P_C = n_C RT/V$

$P_t = n_t RT/V$

Where n1 (total number of moles of all the gases) = $n_a + n_b + n_c$ Dividing equation (1) by (4)

$P_A/P_t = n_A RT/V \ldots n_t RT/V$

$P_A/P_t = n_A RT/V \times V/ n_t RT$

$P_A/P_t = n_A/n_t = P_A = = n_A/n_t = P_A \alpha n_A$

Partial Pressure of any gas/total no of moles of all gases =

No of moles of that gas / total no of moles of all gases

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 agas:

We know that

Partial Pressure of any gas=No of moles of that gas

Pressure of all gas's Total no of moles of all gasses

Or $\frac{p_t}{p_t} = \frac{n_t}{n_t} p_t = \frac{n_t}{n_t} \times p_t$

$\frac{n_t}{n_t} = X$ where X, is called mole fraction of the component

So $p_t = p_t X$

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 1s insoluble in water. The total pressure consists of the pressure of the water vapors in addition to the pressure of the gas that is generated. The pressure due to water vapor 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 eg. Partial pressure of O in the outside air is higher(159g/cm^2 than in the lungs where the partial pressure of 02 is lower(116gcm $^{-2}$However at high altitudes, the pressure of oxygen decreases and the pilots _may have uncomfortable breathing in a non-pressurized cabin where the partial pressure of 0_2 is about 150g/cm

Respiration by deep sea divers:

Deep sea divers breathe air under increased pressure. At a depth of 40 meters, the pressure increases 5 times than the normal pressure. Therefore, regular air cannot be use in diver's tanks because the partial pressureof02 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.

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 compression chamber is slowly lowered over many hours. Nitrogen gas comes out of the blood and disposed of safely for normal breathing

Q46. Example 19: 1.00 mole of N2 and 3.00 moles of H2 are enclosed in a container of volume 10.0 dm3 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 N_2 (n_2)= 3.00 moles

Total number of moles (nro1a1) = 1 + 3 = 4 moles

V= 10.0 dm 3

T= 298K,

R = 0.082 dm3 atm KMole·1

According to the General Gas Equation,

PV=nRT

= or $P_{total} = \frac{n_{total}RT}{V}$

$$= \frac{4\times0.082\times298}{10}.\frac{dm^3 atmos K^{-1} Mole^{-1}\times K}{dm^3 Mole}$$

$$= \frac{4\times0.082\times298}{10} \text{atm}$$

=9.77atm

$$P_{N2} = \frac{n_{N2}}{n_{total}}\times P_{total} = \frac{1}{4}\times9.77 = 2.44atm$$

$$P_{N2} = \frac{n_{N2}}{n_{total}}\times P_{total} = \frac{3}{4}\times9.77 = 7.33atm$$

Q47.Example 20: The composition of dry air at sea level is approx. N_2=75.5, 02=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 $N_2(n_{N2})$ =75.5

No of moles of $0_{2}(n_{O2})$= 23.2

No of moles of Ar(n_{A2}) = 0.13

P_{total}= 1.00 atm.

$$P_{N2}=\frac{n_{N2}}{n_{total}}\times P_{total}$$

$$=\frac{75.52}{100.00}\times 1.00=0.752atm$$

$$P_{N2}=\frac{n_{O2}}{n_{total}}\times P_{total}= \ =\frac{23.2}{100.00}\times 1.00=0.23\angle u....$$

$$P_{N2}=\frac{n_{A2}}{n_{total}}\times P_{total}$$

$$=\frac{1.3}{100.00}\times 1.00=0.013atm$$

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 $35cm^3$

Calculate its dry volume at S.T.P. (V.P. of water at 6°C = 7mm.of Hg) Solution:

V1=$35cm^3$, V2=? P1 = 765-7 =758mm

T_1= 6 + 273=279K, T_2=273K P2 =760mm.

$$\frac{V_1P_1}{T_1}=\frac{V_2P_2}{T_2}\text{ or }\frac{V_1P_1T_2}{T_1P_2}=\frac{35cm^3\times 758mm\times 273K}{279K\times 760mmofHg}=34.2\ cm^3$$

Activity for students

A gas was collected over water at 10°C and 803mm of Hg. If the gas occupies73cm^3 calculate the volume of dry gas at S.T.P. (aqueous tension at10°c-9.2mm).

(Ans: = 73.55cm^3)

Solution:

Pmoist=P_{dry}+ Aqueous tension

Volume of gas at 10 °c = V_1 = 73 cm^3

T_1 = 10 °C + 273 = 283K

Values at STP:P_2 760

Tz= 273 K

Vz=?

According to the general gas equation

$$\frac{V_1P_1}{T_1}=\frac{V_2P_2}{T_2}$$

$$V_2=\frac{V_1P_1T_2}{T_1P_2}$$

$$V_2=\frac{706x73x273}{760x283}$$

V_2 = 0. 002m^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. IntermixingofNH$_3$andHCltogivesdensewhitefumesofNH$_4$Cl. NH$_3$+HCl NH$_4$Cl

iii. Reddish brown bromine (Br_2) gas taken in a test tube will combine with air taken in another test tube. Red coloration 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.Example22: Equal volumes of HCl and 502 are confined in a porous container. What would be the comparative rates of diffusion of these gas H1 through the porous walls? The molecular masses of HCl and 502 are 36.5 and 64 respectively.

Solution:

Molecular mass of HCl (M,) = 38

Molecular mass of S02 (M,) = 64

Rate of diffusion of HCl(r,)=?

Rate ofdiffusionofS0$_2$(r2)=?

According to the Graham's Law of Diffusion

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

r_1 (HCl)>r(SO_3)

1.33:1

Hence rate of diffusion of HCl will be 1.33 times the rate of diffusion of S0$_2$.

Q52. Example23: If 280 cm³of H_2 diffuse in 40 seconds, how long will it take 490 cm^3 of a gas (X), where vapor density is 25 to diffuse under the same conditions (molecular mass of H_2 is 2).

Solution:

For H2:

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

T, = 40 sec (time for the diffusion)

velocity 280cm^3

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

For Gas (X):

V = 490 cm³ Vapor density (V.D.) =25

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

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

Now according to Graham's Law of Diffusion

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

$$= \sqrt{\frac{50^{25} gmMole^{-1}}{2gmMole^{-1}}}$$

$\frac{7xt_2}{490} = \sqrt{\frac{25}{1}} = \frac{5}{1}$ or $7xt_2x1 = 490x5$=350 sec

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

ACTIVITY

If 465 cm³ of 502 can diffuse through porous partition in 30 seconds. How long will 620 cm of H2S take to diffuse through the same partition (H =1, S =32, 0=16) (Ans: 29.15sec)

Solution:

$v_{50\ 2}$=46Scm^3

$$r_{so_2} = \frac{velocity}{time} = \frac{465}{30} = 15.5cm^3 \sec^{-1}$$

M so_2 = 64 gmol

vHz s = 620 cm^3

vHs = 34 g m o l^{-1}

rHzs=

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

$$r_{H_2S} = \frac{15.5cm^3 \sec^{-1}}{0.72}$$

$$r_{H_2S} = 21.29cm^3 \sec^{-1}$$

$$r_{H_2S} = \frac{velocity}{time} = \frac{465}{21.29} = 29.15 \sec$$

$$T_{H_2S} = 29.15 \sec$$

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

Ans: Liquefaction of Gases:

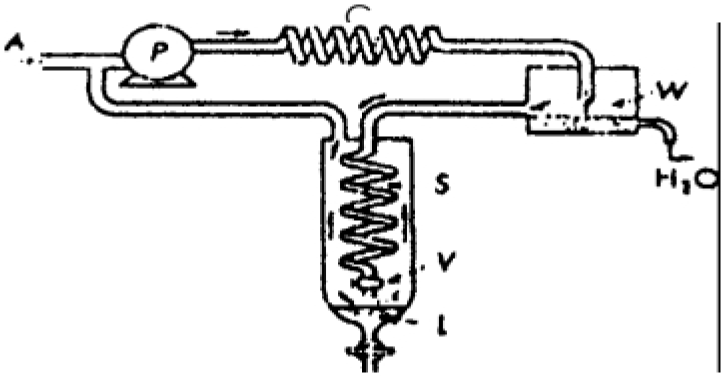
Principle:

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: Lynde'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.



Lynde's method for liquefaction of air

Working Steps:

a. Air is compressed to 200 at map prox.

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 1s 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 liquefied.

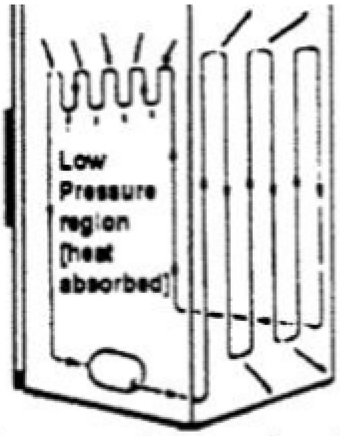
i. The liquid air collects in the chamber L and can be drawn off.

j. Any uncondensed air is re-circulated.

Q55. Gives the different application 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 compress 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 referegirator

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 vaporize rapidly. The heat of vaporization needed to do this is drawn from the kinetic energy of the Freon molecule 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.

Self-cooling pop can(container):

A container holding liquid CO_2 Is built right into the can. When the can is opened, the liquid CO_2 and escape out of the top of the can. The heat absorbed by the vaporizing 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: Velocity of molecules:

The equation for the root means square velocity deduced from kinetic equation

$$Crms = \sqrt{\frac{3RT}{M}}$$

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:

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

$$E_k = \frac{1}{2}mv^2 \Rightarrow \frac{2E_k}{m} = \sqrt{\frac{2E_k}{m}} \Rightarrow v = \sqrt{\frac{1}{m}}\sqrt{2E_k}$$

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

Therefore $v\alpha\sqrt{\frac{1}{m}}$

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:

i. 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:

$V_2 = V_1$ (Conversion factor due to temperature change) x (Conversion factor due to Pressure charge)

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

The ratio between T, and T2 is due to Charles law (V a T) and that of $P_1$and P2 is due to Boyle's law

$(V \alpha \frac{1}{P})$

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.

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 H2 gas collected over water at &•c and 755mm of Hg pressure occupied a volume of 35cm^3. calculate its dry volume at S.T.P.

Solution:

V,=35cm^3 V:=?

P,= 765 - 7=758mm P$_2$ = 760mm

T,= 6+273=279K T$_2$=273K

$\frac{V_1P_1}{T_1} = \frac{V_2P_2}{T_2}$ or V_2=34.2 cm^3

Q58. Example 25: A fire fighters air tank contains 12.0 dm^3 of air compressed to 1.40 x **10^4 KPa 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?**

Solution:

$V = 12 d.m^3$

$V_2 = ?$

$P = 1.40 X 10^{4}$ KPa

T_1 = 22°c + 273 = 295K

Apply the equation for conversion factors

P_2 = 102K.Pa

T_2 = 42°C + 273 = 315K

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

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

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

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

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

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

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=? R = 8.31KPa $dm^3 mole^{-1} K^{-1}$

n = 6.30 moles

T = 23°C + 273 = 296K

P = $550KP_a$

Now according to the Ideal gas law equation·

VP=nRT

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

V = $28.2dm^3$

Result:

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

Activity for students

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 100KP, and the person's body temperature is 31°C.

Solution:

n = 0.115 moles

V =?

P =100kPa= 0. 986atm

T= 37°C+273=310K

According to the• general gas equation

VP= nRT

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

Q60. Example 27: If a 0.75 dm^3cylinder 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 = 296K

R=8.31KPdm^3 K$^{\cdot 1}$ mole$^{\cdot 1}$

p=?

According to the Ideal Gas Law equation

VP=nRT

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

$$=\frac{6.5\times 8.31\times 296}{0.75}\cdot\frac{KP_a dm^3 K^{-1} Mole^{-1}\times K^{+1}}{dm^3}$$

=21318 KP_a Now 101.325 KP_a =1atm

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

=210.392atm

Activity for students

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: 683.28mm)

Solution:

V_1 = 130m^3

V2 = 150m^3

P_1 = 750 mm Hg

Pz=?

T_1=20°C+273K=293K

T_2 =35 °C + 273 K = 308K

$$\frac{V_1P_1}{T_1}=\frac{V_2P_2}{T_2} \text{ or } P_2=P_1\times\frac{T_2}{T_1}\times\frac{P_1V_1}{V_2}$$

$$P_2=\frac{750x130x308}{293x150}=68{\cdot}328$$

2. find the total pressure exerted by 2 grams of ethane and 3grams of CO_2 contained in a dm³ vessel at 50°C. (R = 0.082dm ^{3} atm K$^{\cdot 1}$ Mo1e$^{\cdot 1)}$

Solution:

Mass of Ethane given= 2g

Number of moles of ethane = 0. 066 mol

Mass of CO_2 given= 3g

Number of moles of C= 0. 068 mol

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

P_{ethane} V = nRT

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

$$P_{ethane}=\frac{o.66molx0.0821x323K}{5}=0.350atm$$

P_{CO_2} V = nRT

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

$$P_{CO_2}=\frac{o.68x0.0821x323}{5}=0.350atm$$

P_{total} = 0. 350 + 0.360

P_{total}= 0. 71 atm

Q61. Example28:(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_2molecules at S.T.P = 22.4dm^3

2.5 mole of Cl_2moleculeatS.T. P = 22.4 x2.5

= 56 dm^3

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

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

= 1.51 x10 ^{23} molecules

Q62. Example 29: A 2.60 dm^3container was evacuated and its mass was found to be 655.53 g. The container was then filled with CO2 gas at a pressure of 100 KP, and a temperature of 24°C, the container and gas were found to have a ma11 of 660,13 g, Calculate the standard molar volume of CO2.

Solution:

$Vco_{2=}$ 2.60dm^3

m_1 (mass of container) =655.539g

m_2{mass of container + gas) =.660.13g

Pco_2= I00KPa

T co_2= 24°C + 273 = 297 K

m co_2 (mass of CO2) = 660.13 - 655.53 =4.60g

mass of co_2 =660.13-655.53-4.60g

Convert the gas {CO2) to S.T.P by applying General Gas Equation

(V_1) V co_2 =2.60dm^3 v2=?

(P1)P co_2 =IOOKP

(T,)T co_2= 293K

P2=101.3KP|1] T_2 = 273K S.T.P

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

Standard Molar Volume {Vm) = Molar Volume x volume of CO2 at S.T.P.

$= 9.5652$ moles x $2.36dm^3$

$= 9.5652 \times 2.36$Moles.dm^3

$= 22.6\ dm^3$ Moles

The accepted value for Molar value= $22.4dm^3$

$$Vm = \frac{\text{Molar mass of}CO_2}{\text{Mass of }CO_2}$$

Thus, the conversion of volume of a gas to molar volume can be obtained by using the rates

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

Solution

$$p = 102KP_a = \frac{102000}{101325} = 1.006$$

T = 20 °C + 273 =293k

m = 66.184g

n=1.141mol

PV = nRT

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

V $=21.2dm^3$

Calculate the number of molecules atmospheric pressure and 10grams of ammonia at 1.5

(Ans: 3.60×10^{23} molecules)

Solution:

Mass of ammonia given =10 g

Molar mass of ammonia = 17 g/mol

Number of moles of ammonia=0.588moles

1 mole of ammonia contains= 6.02×10^{23}molecules

0. 588 moles of ammonia contains =$0.588 \times 6.02 \times 10^{23}$molecules

0.588 moles of ammonia contain= 3.60×10^{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.

Molecular gas→ atoms gas→ Ions.

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