

## GASEOUS STATE SOLUTIONS

### LEVEL – 1

#### Gas Law's

1. If 20 cm<sup>3</sup> gas at 1 atm. Is expanded to 50 cm<sup>3</sup> at constant T, then what is the final pressure  
 (1)  $20 \times \frac{1}{50}$                       (2)  $50 \times \frac{1}{20}$                       (3)  $1 \times \frac{1}{20} \times 50$                       (4) None of these

(1)

At constant T,  $P_1V_1 = P_2V_2$

$$1 \times 20 = P_2 \times 50; P_2 = \frac{20}{50} \times 1$$

2. Which of the following statement is false

- (1) The product of pressure and volume of fixed amount of a gas is independent of temperature  
 (2) Molecules of different gases have the same K.E. at a given temperature  
 (3) The gas equation is not valid at high pressure and low temperature  
 (4) The gas constant per molecule is known as Boltzmann constant

(1)

$P \cdot V = \text{Constant}$  at constant temperature. As temperature changes, the value of constant also changes.

3. 10 g of a gas at atmospheric pressure is cooled from 273°C to 0°C keeping the volume constant, its pressure would become

- (1)  $\frac{1}{2}$  atm                      (2)  $1/273$  atm                      (3) 2 atm                      (4) 273 atm

$$T_1 = 273^\circ\text{C} = 273 + 273^\circ\text{K} = 546^\circ\text{K}$$

$$T_2 = 0^\circ\text{C} = 273 + 0^\circ\text{C} = 273^\circ\text{K}$$

$$P_1 = 1; P_2 = ?$$

According to Gay – Lussac's law

$$\frac{P_1}{T_1} = \frac{P_2}{T_2} \therefore P_2 = \frac{P_1 T_2}{T_1} = \frac{1 \times 273}{546} \text{ atm} = \frac{1}{2} \text{ atm.}$$

4. 400 cm<sup>3</sup> of oxygen at 27°C were cooled to -3°C without change in pressure. The contraction in volume will be

- (1) 40 cm<sup>3</sup>                      (2) 30 cm<sup>3</sup>                      (3) 44.4 cm<sup>3</sup>                      (4) 360 cm<sup>3</sup>

(1)

$$V_2 = \frac{T_2}{T_1} \cdot V_1 = \frac{270}{300} \cdot 400 \text{ cm}^3 = 360 \text{ cm}^3$$

$$\text{Contraction} = V_1 - V_2 = 400 - 360 = 40 \text{ cm}^3$$

5. In a closed flask of 5 litres, 1.0 g of H<sub>2</sub> is heated from 300 to 600 K. which statement is not correct
- (1) Pressure of the gas increases
  - (2) The rate of collision increases
  - (3) The number of moles of gas increases**
  - (4) The energy of gaseous molecules increases

(3)  
At constant V of a definite mass

$$\frac{P_1}{T_1} = \frac{P_2}{T_2} \therefore \frac{P_1}{P_2} = \frac{300}{600} = \frac{1}{2} \text{ i.e. Pressure increases and on increasing temperature energy of molecules}$$

increases so the rate of collisions also increases and number of moles remains constant because there is neither addition nor removal of gas in the occurring.

6. For an ideal gas number of moles per litre in terms of its pressure P, gas constant R and temperature T is
- (1) PT/R
  - (2) PRT
  - (3) P/RT**
  - (4) RT/P

$$PV = nRT \therefore \frac{n}{V} = \frac{P}{RT}$$

7. If two moles of an ideal gas at 546 K occupy a volume of 44.8 litres, the pressure must be
- (1) 2 atm**
  - (2) 3 atm
  - (3) 4 atm
  - (4) 1 atm

$$P = \frac{nRT}{V} = \frac{2 \times 0.0821 \times 546}{44.8} = 2 \text{ atm.}$$

8. Pure hydrogen sulphide is stored in a tank of 100 litre capacity at 20°C and 2 atm pressure. The mass of the gas will be
- (1) 34 g
  - (2) 340 g
  - (3) 282.4 g**
  - (4) 28.24 g

(3)  
$$n = \frac{PV}{RT} = \frac{m}{M}$$

$$m = \frac{MPV}{RT} = \frac{34 \times 2 \times 100}{0.082 \times 293} = 282.4 \text{ gm}$$

9. One litre of a gas weighs 2 g at 300 K and 1 atm pressure. If the pressure is made 0.75 atm, at which of the following temperatures will one litre of the same gas weight one gram
- (1) 450 K**
  - (2) 600 K
  - (3) 800 K
  - (4) 900 K

(1)

$$\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2} \therefore T_2 = \frac{P_2 V_2}{P_1 V_1} T_1 \frac{m_1}{m_2}$$

$$= \frac{0.75}{1} \times \frac{1}{1} \times \frac{2}{1} \times 300 = 450$$

10. The density of a gas at 27°C and 1 atm is d. Pressure remaining constant at which of the following temperatures will its density become 0.75 d

- (1) 20°C                      (2) 30°C                      **(3) 400 K**                      (4) 300 K  
(3)

At constant pressure

$$V \propto nT \propto \frac{m}{M} T$$

$$\frac{V_1}{V_2} = \frac{m_1 T_1}{m_2 T_2} \therefore \frac{T_1}{T_2} = \frac{V_1}{V_2} \times \frac{m_2}{m_1} = \frac{d_2}{d_1}$$

$$\Rightarrow \frac{300}{T_2} = \frac{0.75d}{d}$$

$$\Rightarrow T_2 = \frac{300}{0.75} = 400K$$

11. A vessel contains 0.1 mole of He, 0.1 mole of O<sub>2</sub>, and 0.3 moles of N<sub>2</sub>. The total pressure is 1 atmosphere. The pressure exerted by O<sub>2</sub> is

- (1) 380 mm of Hg      (2) 456 mm of Hg      (3) 304 mm of Hg      **(4) 152 mm of Hg**  
(1)

Total moles = 0.1(He) + 0.1 (O<sub>2</sub>) + 0.3 (N<sub>2</sub>) = 0.5 moles

Pressure exerted by O<sub>2</sub> = mole fraction of O<sub>2</sub> × total pressure

$$= \frac{0.1}{0.5} \times 1 = \frac{1}{5} \times 760 = 152\text{mm}$$

12. A cylinder containing LPG can withstand a pressure of 14.9 atm. The pressure gauge of the cylinder indicates 12 atm at 27°C. Due to a sudden fire, the cylinder explodes. What should have been its temperature?

- (1) 124 K                      (2) 372 K                      (3) 346 K                      (4) 432 K  
(2)

$$\frac{P_1}{T_1} = \frac{P_2}{T_2} \text{ or } \frac{12}{300} = \frac{14.9}{T_2} \text{ or } T_2 = 372 \text{ K. We have considered pressure as 14.9 since the cylinder}$$

explodes at this pressure.

13. A 10 litre gas is inserted into a car tyre at 4 atm and 27°C. The temperature of tyre increase to 57°C during driving. What would be the pressure of gas during driving?

- (1) 3.1 atm                      **(2) 4.4 atm**                      (3) 300 mm Hg                      (4) 2270 mm Hg

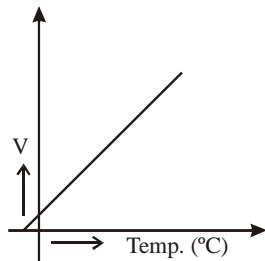
14. What fraction of air is expelled out if a flask containing V litre gas is heated from 27°C to 327°C?

- (1) 25%                      **(2) 50%**                      (3) 75%                      (4) 10%

15. A balloon of volume 200 litre having ideal gas at 1 atm pressure and at 27°C, when rises to a height where atmospheric pressure is 380 mm Hg and temperature is -3°C, balloon will

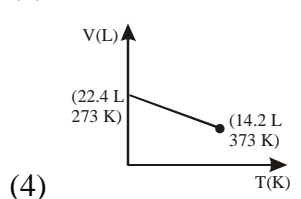
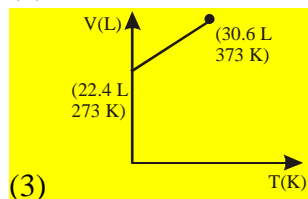
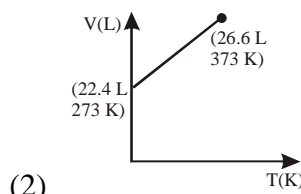
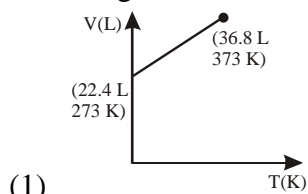
- (1) Contract
- (2) Expand
- (3) No change in volume of balloon
- (4) Initially expand and then contract

16. The following graph illustrates



- (1) Dalton's law
- (2) Charles's law
- (3) Boyle's law
- (4) Gay-Lussac's law

17. Which of the following volume (V) – temperature (T) plots represents the behaviour of one mole of an ideal gas at one atmospheric pressure?



18. A gas will approach ideal behaviour at

- (1) Low temperature and low pressure
- (2) Low temperature and high pressure
- (3) High temperature and low pressure
- (4) High temperature and high pressure

19. An ideal gas will have maximum density when

- (1) P = 0.5 atm, T = 600 K
- (2) P = 2 atm, T = 150 K
- (3) P = 1 atm, T = 300 K
- (4) P = 1 atm, T = 500 K

20. 120 g of an ideal gas of molecular weight 40 are confined to a volume of 20 L at 400 K. Using R = 0.0821 L atm K<sup>-1</sup> mole<sup>-1</sup>, the pressure of the gas is

- (1) 4.90 atm
- (2) 4.92 atm
- (3) 5.02 atm
- (4) 4.96 atm

21. Which of the following plots for a fixed amount of an ideal gas at constant temperature will be linear with the slope equal to zero  
 (1) P against V      (2) P against 1/V      (3) PV against P      (4) 1/P against V

**Dalton's law of partial pressure**

22. A whether balloon with hydrogen at 1 atm and 27°C has volume equal to 12000 litres. On ascending it reaches a place where the temperature is -23°C and pressure is 0.5 atm. The volume of the balloon is  
 (1) 24000 litres      (2) 20000 litres      (3) 10000 litres      (4) 12000 litres

$$V_2 = \frac{P_1 T_2}{P_2 T_1} \cdot V_1 = \frac{1}{0.5} \times \frac{250}{300} \times 12000 \text{ lit.}$$

$$= 20000 \text{ lit}$$

23. Dalton's law of partial pressure will not apply to which of the following mixture of gases  
 (1) H<sub>2</sub> and SO<sub>2</sub>      (2) H<sub>2</sub> and Cl<sub>2</sub>      (3) H<sub>2</sub> and O<sub>2</sub>      (4) O<sub>2</sub> and Cl<sub>2</sub>

(2)  
 Because H<sub>2</sub> and Cl<sub>2</sub> gases may react with each other to produce HCl gas hence Dalton's law is not applicable.

24. Equal masses of two gases of molecular weight 4 and 40 are mixed. The pressure of the mixture is 1.1 atm. The partial pressure of the light gas in this mixture is  
 (1) 0.55 atm      (2) 0.11 atm      (3) 1 atm      (4) 0.12 atm

$$\text{No. of moles of lighter gas} = \frac{m}{4}$$

$$\text{No. of moles of heavier gas} = \frac{m}{40}$$

$$\text{Total no. of moles} = \frac{m}{4} + \frac{m}{40} = \frac{11m}{40}$$

$$\text{Mole fraction of lighter gas} = \frac{\frac{m}{4}}{\frac{11m}{40}} = \frac{10}{11}$$

$$\text{Partial pressure due to lighter gas} = P_0 \times \frac{10}{11}$$

$$= 1.1 \times \frac{10}{11} = 1 \text{ atm}$$

25. There are  $6.02 \times 10^{22}$  molecules each of  $N_2$ ,  $O_2$  and  $H_2$  which are mixed together at 760 mm and 273 K. The mass of the mixture in grams is

(1) 6.2 (2) 4.12 (3) 3.09 (4) 7

(1)

$6.02 \times 10^{22}$  molecules of each  $N_2$ ,  $O_2$  and  $H_2$

$$= \frac{6.02 \times 10^{22}}{6.02 \times 10^{23}} \text{ moles of each}$$

Weight of mixture = weight of 0.1 mole  $N_2$  + weight

Of 0.1 mole  $H_2$  + weight of 0.1 mole of  $O_2$

$$= (28 \times 0.1) + (2 \times 0.1) + (32 \times 0.1) = 6.2 \text{ gm}$$

26. A pre-weighted vessel was filled with oxygen at S.T.P. and weighted. It was then evacuated, filled with  $SO_2$  at the same temperature and pressure, and again weighted, the weight of oxygen will be

(1) The same as that of  $SO_2$  (2)  $\frac{1}{2}$  that of  $SO_2$   
(3) Twice that of  $SO_2$  (4) One fourth that of  $SO_2$

(2)

Since, P, V, T are same so  $N_{O_2} = N_{SO_2}$

$$\therefore \frac{M_{O_2}}{M_{SO_2}} = \frac{M_{O_2}}{M_{SO_2}}$$

$$\frac{\text{M.wt. of } O_2}{\text{M.wt. of } SO_2} \Rightarrow \frac{M_1}{M_2} \Rightarrow \frac{32}{64} = \frac{1}{2}$$

The weight of oxygen will be  $\frac{1}{2}$  that of  $SO_2$

27. A closed vessel contains equal number of nitrogen and oxygen molecules at a pressure of P mm. if nitrogen is removed from the system then the pressure will be

(1) P (2) 2P (3) P/2 (4) P<sup>2</sup>

(3)

Since no. of molecules of gas is halved so pressure should also be halved.

28. Which of the following gas mixture is not applicable for Dalton's law of partial pressure

(1)  $SO_2$  and  $Cl_2$  (2)  $O_2$  and  $N_2$  (3)  $H_2$  and  $O_2$  (4)  $Cl_2$  and  $N_2$

(1)

Mixture of  $SO_2$  and  $Cl_2$  are reacted chemically and forms  $SO_2Cl_2$ . That is why Dalton's Law is not applicable to mixture of these gases.

29. What is the molecular weight of a gas whose density at 40°C and 785 mm of Hg pressure is 1.3 g L<sup>-1</sup>?

(1) 32.00 (2) 40.00 (3) 15.00 (4) 98.00

(1)

We know that,  $PV = nRT$

$$PV = \frac{\omega}{M}RT \text{ or } P = \frac{\omega}{V} \times \frac{1}{M}RT$$

$$\text{Or, } P = \frac{d}{M}RT$$

$$\text{Now, } \frac{785}{760} = \frac{1.3}{M} \times 0.0821 \times 313 \text{ or } M = 32.4 \text{ or } 32$$

30. A gas is found to have the formula  $(CO)_x$ . the vapour density of the gas is 70. The value of  $x$  is  
 (1) 7 (2) 4 (3) 5 (4) 6  
 (4)  
 $(CO)_x = (6+8)x = 70 \text{ or } x = 5$

31. Which of the following mixtures of gases does not obey Daltons law of partial pressure  
 (1)  $O_2$  and  $CO_2$  (2)  $N_2$  and  $O_2$  (3)  $Cl_2$  and  $O_2$  (4)  $NH_3$  and  $HCl$

32. If the four tubes of a car are filled to the same pressure with  $N_2$ ,  $O_2$ ,  $H_2$  and  $Ne$  separately, then which one will be filled first?  
 (1)  $N_2$  (2)  $O_2$  (3)  $H_2$  (4)  $Ne$

**Graham's law of diffusion**

33. The ratio of the rate of diffusion of a given element to that of helium is 1.4. the molecular weight of the element is  
 (1) 2 (2) 4 (3) 8 (4) 16

$$\frac{r_g}{r_{He}} = \sqrt{\frac{M_{He}}{M_g}} \therefore M_g = M_{He} \cdot \frac{r_{He}^2}{r_g^2}$$

$$= \frac{4}{(1.4)^2} \approx \frac{4}{2} = 2 \text{ [Note : } 1.4 \approx \sqrt{2} \text{]}$$

34. The molecular weight of a gas which diffuses through a porous plug at  $1/6^{\text{th}}$  of the speed of hydrogen under identical conditions is  
 (1) 27 (2) 72 (3) 36 (4) 48

$$r_g = \frac{1}{6} r_{H_2}; M_g = M_{H_2} \cdot \left[ \frac{r_{H_2}}{r_g} \right]^2$$

$$= 2 \times 6^2 = 2 \times 36 = 72$$

35. If rate of diffusion of A is 5 times that of B, what will be the density ration of A and B  
 (1) 1/25 (2) 1/5 (3) 25 (4) 4

$$r_a = 5r_b; \frac{d_a}{d_b} = \left[ \frac{r_b}{r_a} \right]^2 = \left( \frac{1}{5} \right)^2 = \frac{1}{25}$$

36. At constant volume and temperature conditions, the rate of diffusion  $D_A$  and  $D_B$  of gases A and B having densities  $\rho_A$  and  $\rho_B$  are related by the expression

$$(1) D_A = \left[ D_B \cdot \frac{\rho_A}{\rho_B} \right]^{1/2} \quad (2) D_A = \left[ D_B \cdot \frac{\rho_B}{\rho_A} \right]^{1/2} \quad (3) D_A = D_B \left( \frac{\rho_A}{\rho_B} \right)^{1/2} \quad (4) D_A = D_B \left( \frac{\rho_B}{\rho_A} \right)^{1/2}$$

$$\frac{D_A}{D_B} = \sqrt{\frac{\rho_B}{\rho_A}} = \left[ \frac{\rho_B}{\rho_A} \right]^{1/2} \therefore D_A = D_B \left( \frac{\rho_B}{\rho_A} \right)^{1/2}$$

37. A bottle of ammonia and a bottle of dry hydrogen chloride connected through a long tube are opened simultaneously at both ends. The white ammonium chloride first formed will be

- (1) At the centre of the tube  
 (2) Nearer to the hydrogen chloride bottle  
 (3) Nearer to the ammonia bottle  
 (4) Throughout the length of the tube

$\text{NH}_4\text{Cl}$  ring will first formed near the HCl bottle because rate of diffusion of  $\text{NH}_3$  is more than that of HCl because  $M_{\text{NH}_3} : M_{\text{HCl}} = 17 : 36.5$ . So  $\text{NH}_3$  will reach farther towards the HCl bottle and will react there with HCl to form  $\text{NH}_4\text{Cl}$ .

38. A gas diffuses at a rate which is twice that of another gas B. the ratio of molecular weights of A to B is

- (1) 1.0                      (2) 0.75                      (3) 0.50                      (4) 0.25

$$\Rightarrow \frac{M_A}{M_B} = \frac{1}{(2)^2} = \frac{1}{4} = 0.25$$

39. Two grams of hydrogen diffuse from a container in 10 minutes. How many grams of oxygen would diffuse through the same container in the same time under similar conditions

- (1) 0.5 g                      (2) 4 g                      (3) 6 g                      (4) 8 g

$$\frac{r_{\text{H}_2}}{r_{\text{O}_2}} = \sqrt{\frac{32}{2}} = 4$$

$$\therefore \frac{r_{\text{H}_2}}{r_{\text{O}_2}} = \sqrt{\frac{32}{2}} = 4$$

$$\therefore \frac{n_{\text{H}_2}}{n_{\text{O}_2}} = 4 (\text{as time is same})$$

$$n_{\text{H}_2} \text{ diffused} = \frac{2}{2} = 1$$

$$\therefore n_{\text{O}_2} \text{ diffused} = \frac{1}{4}$$



$$\therefore \text{mass of O}_2 \text{ diffused} = \frac{32}{4} = 8\text{g}$$

40. X ml of  $\text{H}_2$  gas effuses through a hole in a container in 5 seconds. The time taken for the effusion of the same volume of the gas specified below under identical condition is

- (1) 10 seconds : He     **(2) 20 seconds :  $\text{O}_2$**      (3) 25 seconds : CO     (4) 55 seconds :  $\text{O}_2$

$$r \propto \frac{1}{\sqrt{M}} \therefore r = \frac{\text{Volume effused}}{\text{time taken}} = \frac{V}{t}$$

$$\frac{V}{t} \propto \frac{1}{\sqrt{M}} \therefore \text{for same volumes (V constant)}$$

$$t \propto \sqrt{M} \therefore \frac{t_1}{t_2} = \sqrt{\frac{M_1}{M_2}}$$

$$t_{\text{He}} = t_{\text{H}_2} \sqrt{\frac{M_{\text{He}}}{M_{\text{H}_2}}} = 5\sqrt{\frac{4}{2}} = 5\sqrt{2}\text{s.}$$

$$t_{\text{O}_2} = t = 5\sqrt{\frac{32}{2}} = 20\text{s}$$

$$t_{\text{CO}} = 5\sqrt{\frac{28}{2}} = 5\sqrt{14}\text{s};$$

$$t_{\text{CO}_2} = 5\sqrt{\frac{44}{2}} = 5\sqrt{22}\text{s}$$

41. At constant temperature and pressure which gas will diffuse faster  $\text{H}_2$  or  $\text{O}_2$ ?

- (1) Hydrogen**    (2) Oxygen  
(3) Both will diffuse in same rate                    (4) None of the above

(1) Rate of diffusion  $\propto \frac{1}{\sqrt{\text{Molecular mass}}}$  that is why  $\text{H}_2$  gas diffuse first.

42. What is the ratio of diffusion rate of oxygen and hydrogen

- (1) 1 : 4**    (2) 4 : 1    (3) 1 : 8    (4) 8 : 1

$M_1 = 32$  for  $\text{O}_2$ ,  $M_2 = 2$  for  $\text{H}_2$

$$\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}}; \frac{r_1}{r_2} = \sqrt{\frac{2}{32}} = \sqrt{\frac{1}{16}} = \frac{1}{4}$$

43. According the Graham's law at a given temperature, the ratio of the rates of diffusion  $r_A / r_B$  of gases A and B is given by

- (1)  $(P_A / P_B) (M_A / M_B)^{1/2}$     (2)  $(M_A / M_B) (P_A / P_B)^{1/2}$   
**(3)  $(P_A / P_B) (M_B / M_A)^{1/2}$**     (4)  $(M_A / M_B) (P_B / P_A)^{1/2}$

**Kinetic theory of gases Different kinds of velocities**

44. According to kinetic theory of gases, for a diatomic molecule
- (1) The pressure exerted by the gas is proportional to the average velocity of the molecules
  - (2) The pressure exerted by the gas is proportional to the root mean square velocity of the molecules
  - (3) The root mean square velocity is inversely proportional to the temperature
  - (4) The mean translational kinetic energy of the molecules is proportional to the absolute temperature.

According to K.T.G

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

$$\therefore \text{K.E} \propto T$$

45. The average K.E. of an ideal gas in calories per mole is approximately equal to
- (1) Three times the absolute temperature
  - (2) Absolute temperature
  - (3) Two times the absolute temperature
  - (4) 1.5 times the absolute temperature

$$\text{K.E.} = \frac{3}{2} .RT = \frac{3}{2} .2.T \quad \because R \approx 2\text{calK}^{-1}\text{mol}^{-1}$$

$$\text{K.E.} = 3T.$$

46. Kinetic energy and pressure of a gas per unit volume are related as

$$(1) P = \frac{2}{3} E \quad (2) P = \frac{3}{2} E \quad (3) P = \frac{1}{2} E \quad (4) P = 2E$$

$$\text{T.K.E.} = \frac{3}{2} nRT = \frac{3}{2} PV$$

$$\therefore P = \frac{2}{3} \frac{E}{V} \text{ for unit volume (V = 1)}$$

$$\Rightarrow P = \frac{2}{3} E$$

47. Indicate the incorrect statement for a 1-L sample of N<sub>2</sub>(g) and O<sub>2</sub>(g) at 298 K and 1 atm pressure
- (1) The average translational KE per molecule is the same for N<sub>2</sub> and O<sub>2</sub>
  - (2) The RMS speed is same for both N<sub>2</sub> and O<sub>2</sub>
  - (3) The density of N<sub>2</sub> is less than that of O<sub>2</sub>
  - (4) The total translational KE of both N<sub>2</sub> and O<sub>2</sub> is the same.

Kinetic energies per molecule will be same because it is proportional to absolute temperature only.

$$\frac{d_{N_2}}{d_{CO_2}} = \frac{M_{N_2}}{M_{CO_2}} = \frac{28}{44} \text{ i.e. } d_{N_2} < d_{CO_2}$$

Total translational kinetic energy will also be same because at same temperature & pressure number of molecules present in same volume would be same But  $u_{\text{RMS}} \propto \frac{1}{\sqrt{M}}$

48. The average kinetic energy of an ideal gas per molecule in SI units at 25°C will be  
 (1)  $6.7 \times 10^{-21} \text{ kJ}$       (2)  $6.7 \times 10^{-21} \text{ J}$       (3)  $6.7 \times 10^{-20} \text{ kJ}$       (4)  $7.16 \times 10^{-20} \text{ kJ}$

Average kinetic energy per molecule

$$= \frac{3}{2} kT = \frac{3}{2} \times 1.38 \times 10^{-23} \times 300 \text{ J} = 6.17 \times 10^{-21} \text{ J}$$

49. The ratio between the root mean square velocity of  $\text{H}_2$  at 50 K and that of  $\text{O}_2$  at 800 K is  
 (1) 4      (2) 2      (3) 1      (4) 1/4

$$\frac{U_{\text{H}_2}}{U_{\text{O}_2}} = \sqrt{\frac{T_{\text{H}_2}}{M_{\text{H}_2}} \cdot \frac{M_{\text{O}_2}}{T_{\text{O}_2}}} = \sqrt{\frac{50}{2} \cdot \frac{32}{800}} = 1$$

50. The root mean square velocity of an ideal gas in a closed container of fixed volume is increased from  $5 \times 10^4 \text{ cm s}^{-1}$  to  $10 \times 10^4 \text{ cm s}^{-1}$ . Which of the following statement correctly explains how the change is accomplished

- (1) By heating the gas, the temperature is doubled  
 (2) By heating the gas, the temperature is quadrupled (i.e. made four times)  
 (3) By heating the gas, the volume is quadrupled  
 (4) By heating the gas, the pressure is doubled

$$\frac{U_1}{U_2} = \sqrt{\frac{T_1}{T_2}} \quad \therefore \frac{T_1}{T_2} = \left( \frac{5 \times 10^4}{10 \times 10^4} \right)^2 = \frac{1}{4}$$

51. The rms speed of  $\text{N}_2$  molecules in a gas is  $u$ . If the temperature is doubled and the nitrogen molecules dissociate into nitrogen atoms, the rms speed becomes

- (1)  $u/2$       (2)  $2u$       (3)  $4u$       (4)  $14u$

Temperature is doubled and due to dissociation M.W. is halved. So  $u_{\text{RMS}}$  is doubled.

52. The ratio of most probable velocity to that of average velocity is

- (1)  $\pi/2$       (2)  $2/\pi$       (3)  $\sqrt{\pi}/2$       (4)  $2/\sqrt{\pi}$

$$\frac{V_{\text{mp}}}{V_{\text{av}}} = \frac{\sqrt{\frac{2RT}{M}}}{\frac{\sqrt{8RT}}{\sqrt{\pi M}}} = \frac{\sqrt{\pi}}{2}$$

53. If the  $V_{\text{rms}}$  is  $30R^{1/2} \text{ cm/s}$  at 27°C then calculate the molar mass of gas in kilogram

- (1) 1      (2) 2      (3) 4      (4) 0.0001

$$V_{\text{rms}} = \sqrt{\frac{3RT}{M}}$$

$$\sqrt{30^2 R} = \sqrt{\frac{3RT}{M}} \Rightarrow 30 \times 30R = \frac{3R \times 300}{M}$$

$$\Rightarrow M = \frac{3 \times 300}{30 \times 30} = 1 \text{ gm} = 0.001 \text{ kg.}$$

54. The average kinetic energy of an ideal gas per molecule in SI units at 25°C will be

- (1)  $6.17 \times 10^{-21} \text{ J}$  (2)  $6.17 \times 10^{-22} \text{ J}$   
 (3)  $6.17 \times 10^{-20} \text{ J}$  (4)  $7.16 \times 10^{-20} \text{ J}$

The average kinetic energy of an ideal gas per molecule in SI units at 25°C will be  $6.17 \times 10^{-21} \text{ J}$

The average kinetic energy of an ideal gas per molecule is given by the expression

$$\text{Average kinetic energy} = \frac{3}{2} kT$$

Here, k is boltzmann constant and T is absolute temperature.

$$k = 1.36 \times 10^{-23} \text{ J/K}$$

$$T = 25 + 273 = 298 \text{ K}$$

$$\text{Average kinetic energy} = \frac{3}{2} \times 1.36 \times 10^{-23} \text{ J/K} \times 298 \text{ K} = 6.17 \times 10^{-21} \text{ J}$$

55. The r.m.s. velocity of hydrogen is  $\sqrt{7}$  times the r.m.s velocity of nitrogen. If T is the temperature of the gas, then

- (1)  $T(\text{H}_2) = T(\text{N}_2)$  (2)  $T(\text{H}_2) > T(\text{N}_2)$   
 (3)  $T(\text{H}_2) < T(\text{N}_2)$  (4)  $T(\text{H}_2) = \sqrt{7} T(\text{N}_2)$

The rms velocity of hydrogen is  $\sqrt{7}$  times the rms velocity nitrogen.

$$u_{\text{H}_2} = \sqrt{7} u_{\text{N}_2}$$

The rms velocity of a gas is directly proportional to square root of temperature.

$$u \propto \sqrt{T}$$

$$\sqrt{T_{\text{H}_2}} = \sqrt{7} \sqrt{T_{\text{N}_2}}$$

$$T_{\text{H}_2} = 7T_{\text{N}_2}$$

$$T_{\text{H}_2} > T_{\text{N}_2}$$

56. The r.m.s velocity of CO<sub>2</sub> at a temperature T (in Kelvin) is  $x \text{ cm sec}^{-1}$ . At what temperature (in Kelvin), the r.m.s. velocity of nitrous oxide would be  $4x \text{ cm sec}^{-1}$ ?

- (1) 16 T (2) 2 T (3) 4 T (4) 32 T

57. Which of the following has maximum root mean square velocity at the same temperature?

- (1) SO<sub>2</sub>                      (2) CO<sub>2</sub>                      (3) O<sub>2</sub>                      (4) H<sub>2</sub>

Solution: At same temperature, Root Mean Square velocity is greatest for the gas having lowest Molecular Weight.

58. The temperature of an ideal gas is reduced from 927°C to 27°C, the r.m.s. velocity of the molecules becomes

- (1) double the initial value                      (2) half of the initial value  
(3) four times the initial value                      (4) ten times the initial value

$$v_{rms} \text{ speed} = \sqrt{\frac{3RT}{M}}$$

when other terms are constant,  $v_{rms} \propto \sqrt{T}$ .

$$\begin{aligned} (v_{rms})_1 &= \sqrt{(927 + 273)} = \sqrt{1200} \\ (v_{rms})_2 &= \sqrt{(27 + 273)} = \sqrt{300} \end{aligned} \quad \left. \begin{array}{l} \text{Temperatures are} \\ \text{converted into} \\ \text{Kelvin} \end{array} \right\}$$

$$\therefore \frac{(v_{rms})_1}{(v_{rms})_2} = \frac{\sqrt{1200}}{\sqrt{300}} = 2$$

$$\therefore \boxed{(v_{rms})_2 = \frac{1}{2} \cdot (v_{rms})_1}$$

Correct choice - option B

59. The ratio of the root mean square velocity of H<sub>2</sub> at 50 K and that of O<sub>2</sub> at 800 K is:

- (1) 4                      (2) 2                      (3) 1                      (4) 1/4

As we know,

$$v_{rms} = \sqrt{\frac{3RT}{M}}$$

For H<sub>2</sub> at 50 K,

$$v_1 = \sqrt{\frac{3R \times 50}{2}}$$

For O<sub>2</sub> at 800 K,

$$v_2 = \sqrt{\frac{3R \times 800}{32}}$$

$$\therefore \frac{v_1}{v_2} = \frac{\sqrt{\frac{3R \times 50}{2}}}{\sqrt{\frac{3R \times 800}{32}}} = 1$$

**Real Gases**

60. The compressibility factor of a gas is less than unity at STP. Therefore  
 (1)  $V_m > 22.4$ litres      (2)  $V_m < 2.44$ litres      (3)  $V_m = 2.44$ litres      (4)  $V_m = 44.8$ litres

$$Z = \frac{PV}{nRT} < 1 \text{ at STP (as given)}$$

$$\left. \begin{array}{l} nRT > PV \\ n \times 0.0821 \times 273 > 1V_m \\ 22.41 \text{ litres} > V_m \end{array} \right\} \begin{array}{l} R = 0.0821 \\ T = 273K \\ P = 1 \\ n = 1 \end{array}$$

61. What is the pressure of 2 mole of  $NH_3$  at  $27^\circ C$  at  $27^\circ C$  when its volume is 5 litre in vander Waal's equation ( $a = 4.17$ ,  $b = 0.03711$ )  
 (1) 10.33 atm      (2) 9.33 atm      (3) 9.74 atm      (4) 9.2 atm

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

$$\left( P - \frac{(2)^2 \times 4.17}{(5)^2} \right) (5 - 2 \times 0.03711) = 2 \times 0.0821 \times 300$$

$$P = \frac{2 \times 0.0821 \times 300}{5 - 2 \times 0.03711} - \frac{4.7 \times 2^2}{5^2}$$

$$\Rightarrow 10 - 0.66 = 9.33 \text{ atm.}$$

62. The compressibility factor of a gas is define as  $Z = PV/RT$ . The compressibility factor of ideal gas is  
 (1) 0      (2) Infinity      (3) 1      (4) -1

$$Z = \frac{PV}{RT}; \text{ for ideal gas } PV = RT; \text{ so } Z = 1$$

63. Gases deviate from the ideal gas behaviour because their molecules  
 (1) Possess negligible volume      (2) Have forces of attraction between them  
 (3) Are polyatomic      (4) Are not attracted to one another

Because molecules of real gases have intermolecular forces of attraction so the effective impact on the wall of container is diminished. Pressure of real gas is reduced by  $\frac{a}{V^2}$  factor hence behaviour of real gas deviate from ideal behaviour.

64. Vanderwaal's equation of state is obeyed by real gases. For n moles of a real gas, the expression will be

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

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

(3)  $\left(P + \frac{na}{V^2}\right)(nV-b) = nRT$

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

By definition, Vanderwaals, equation of state is :

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

65. The Vander Waal's equation reduces itself to the ideal gas equation at

(1) High pressure and low temperature

(2) Low pressure and low temperature

(3) Low pressure and high temperature

(4) High pressure and high temperature

Real gas behaves as an ideal gas at very low pressure and high temperature.

66. The temperature at which real gases obey the ideal gas laws over a wide range of pressure is called

(1) Critical temperature

(2) Boyle temperature

(3) Inversion temperature

(4) Reduced temperature

At Boyle temperature real gas follows ideal gas most closely for a wide range of temperature.

67. At high temperature and low pressure, the Vander Waal's equation is reduced to

(1)  $\left(P + \frac{a}{V_m^2}\right)(V_m) = RT$

(2)  $PV_m = RT$

(3)  $P(V_m - b) = RT$

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

At high temperature and low pressure, Vander Waal's equation is reduced to ideal gas equation.

$$PV = nRT$$

$$PV = RT \text{ (For 1 mole of gas)}$$

68. At low pressure, the Vander Waal's equation is reduced to

(1)  $Z = \frac{PV_m}{RT} = 1 - \frac{a}{V_m RT}$

(2)  $Z = \frac{PV_m}{RT} = 1 + \frac{b}{RT} P$

(3)  $PV_m = RT$

(4)  $Z = \frac{PV_m}{RT} = 1 - \frac{aV_m}{RT}$

When pressure is low

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

$$\text{or, } PV = RT - \frac{a}{V}$$

$$\text{or, } \frac{PV}{RT} = 1 - \frac{a}{VRT}$$

$$Z = 1 - \frac{a}{VRT}$$

69. Pressure exerted by 1 mole of methane in a 0.25 litre container at 300K using vander Waal's equation (given  $a = 2.253 \text{ atm}^2 \text{ mol}^{-2}$ ,  $b = 0.0428 \text{ lit mol}^{-1}$ ) is

- (1) 82.82 atm                      (2) 152.51 atm                      (3) 190.52 atm                      (4) 70.52 atm

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

$$\left(P + \frac{2.253}{0.25 \times 0.25}\right)(0.25 - 0.0428)$$

$$= 0.0821 \times 300$$

$$\text{or } (P + 36.048)(0.2072) = 24.63$$

$$\Rightarrow P + 36.048 = 118.87 \Rightarrow P = 82.82 \text{ atm.}$$

70. The value of Vander Waal's constant  $a$  for the gases  $\text{O}_2$ ,  $\text{N}_2$ ,  $\text{NH}_3$  and  $\text{CH}_4$  are 1.360, 1.390, 4.170 and  $2.253 \text{ litre}^2 \text{ atm mole}^{-2}$ . The gas which can most easily be liquefied is

- (1)  $\text{O}_2$                       (2)  $\text{N}_2$                       (3)  $\text{NH}_3$                       (4)  $\text{CH}_4$

The values of vander waals' constant 'a' for the gases  $\text{O}_2, \text{N}_2, \text{NH}_3$  and  $\text{CH}_4$  are 1.360, 1.390, 4.170 and  $2.253 \text{ L atm mol}^2$  respectively. The gas which can most easily be liquefied is  $\text{NH}_3$ . Ammonia has high value of the van der waals' constant 'a' which is the measure of the attractive forces present between the gas molecules. Higher is the value of 'a', stronger are the attractive forces present in the gas molecules and more easily the gas can be liquefied.

71. The compressibility factor for  $\text{H}_2$  and He is usually:

- (1)  $>1$                       (2)  $=1$                       (3)  $<1$                       (4) Either of these

The compressibility factor for  $\text{H}_2$  and He is usually greater than 1.

72. At high pressures, the vander Waal's equation reduces to

- (1)  $PV = RT - \frac{a}{V}$                       (2)  $PV = RT$                       (3)  $PV = RT + Pb$                       (4)  $PV = \frac{aRT}{V^2}$



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

$\underline{P} \uparrow \quad P \gg \frac{a}{V^2}$

then

$$(P)(V - b) = nRT$$

divided by  $nRT$

$$PV - Pb = nRT$$

$PV = nRT + Pb$

73. At low pressure vander Waal's equation for 3 moles of a real gas will have its simplified form

- (1)  $\frac{PV}{RT - \frac{3a}{V}} = 3$       (2)  $\frac{PV}{RT + Rb} = 3$       (3)  $\frac{PV}{RT + 3Pb} = 1$       (4)  $\frac{PV}{RT - \frac{9}{V}} = 3$

At low Pressure:

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

$\downarrow$   
neglected

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

$$PV + \frac{an^2}{V} = nRT$$

$$PV = n \left( RT - \frac{an}{V} \right)$$

$$\frac{PV}{RT - \frac{3a}{V}} = 3 \quad \leftarrow \quad \left( \frac{PV}{RT - \frac{an}{V}} = n \right)$$

**Liquefaction of gases and critical constants and Virial Equation**

74. The temperature at which the second virial coefficient of real gas is zero is called

- (1) Critical temperature      (2) Eutetic point  
 (3) Boiling point      (4) Boyle's temperature

Second Virial coefficient  $B = b - \frac{a}{RT}$

$$\therefore T = \frac{a}{Rb}$$

This temperature is called Boyle's temperature.

75. An ideal gas can't be liquefied because  
 (1) Its critical temperature is always above 0°C  
 (2) Its molecules are relatively smaller in size  
 (3) It solidifies before becoming a liquid  
 (4) Forces operative between its molecules are absent  
 An ideal gas can't be liquefied because molecules of ideal gas have not force of attraction between them.

76. However great the pressure, a gas cannot be liquefied above its  
 (1) Boyle temperature (2) Inversion temperature  
 (3) Critical temperature (4) Room temperature  
 At above critical temperature, substances are existing in gaseous state, since gas cannot be liquefied above it.

77. Which set of conditions represents easiest way to liquefy a gas  
 (1) Low temperature and high pressure (2) High temperature and low pressure  
 (3) Low temperature and low pressure (4) High temperature and high pressure  
 Gases can be liquefied most easily at a low temperature and high pressure.

78. Adiabatic demagnetisation is a technique used for  
 (1) Adiabatic expansion of a gas (2) Production of low temperature  
 (3) Production of high temperature (4) None  
 Adiabatic demagnetisation is a technique of liquefaction of gases in which temperature is reduced.

79. The Vander Waal's constant 'a' for the gases O<sub>2</sub>, N<sub>2</sub>, NH<sub>3</sub> and CH<sub>4</sub> are 1.3, 1.390, 4.170 and 2.253 L<sup>2</sup> atm mol<sup>-2</sup> respectively. The gas which can be most easily liquefied is  
 (1) O<sub>2</sub> (2) N<sub>2</sub> (3) NH<sub>3</sub> (4) CH<sub>4</sub>  
 Value of constant 'a' is highest than other for NH<sub>3</sub> that's why NH<sub>3</sub> can be most easily liquefied.

80. A gas can be liquefied  
 (1) Above its critical temperature (2) At its critical temperature  
 (3) Below its critical temperature (4) At any temperature  
 The temperature below which the gas can be liquefied by the application of pressure alone is called critical temperature.

81. The Vander Waal's parameters for gases W, X, Y and Z are

Gas	A (atm L <sup>2</sup> mol <sup>-2</sup> )	b (L mol <sup>-1</sup> )
W	4.0	0.027
X	8.0	0.030
Y	6.0	0.032
Z	12.0	0.027

Which one of these gases has the highest critical temperature?

- (1) W                      (2) X                      (3) Y                      (4) Z

For Z critical temperature is highest as it has the highest value of a/b.

$$T_c = \frac{8a}{27Rb} \Rightarrow T_c = \frac{8 \times 12}{27 \times .0821 \times .027} = 1603.98K$$

82. Which of the following is incorrect for critical temperature

- (1) It is the highest temperature at which liquid and vapour can coexist  
 (2) Beyond the critical temperature, a gas cannot be liquefied by compression  
 (3) At critical temperature ( $T_c$ ) the surface tension of the system is zero  
 (4) At critical temperature the gas and the liquid phases have different critical densities

All differences between liquid and gas phases vanish at the critical temperature including surface tension.

83. Weight of 112 ml of oxygen at STP on liquefaction would be

- (1) 0.32 g                      (2) 0.64 g                      (3) 0.16 g                      (4) 0.96 g

Vol. of  $O_2$  at STP = 112mL

$$\therefore \text{moles of } O_2 \text{ at STP} = \frac{112}{22400} = \frac{1}{200}$$

$$\therefore \text{mass of } O_2 = \frac{1}{200} \times 32g = 0.16g$$

84. The ratio of Boyle's temperature and critical temperature for a gas is

- (1)  $\frac{8}{27}$                       (2)  $\frac{27}{8}$                       (3)  $\frac{1}{2}$                       (4)  $\frac{2}{1}$

$$T_B = \frac{a}{bR}$$

$$T_C = \frac{8}{27} \frac{a}{bR}$$

$$T_C = \frac{8}{27} T_B$$

$$\frac{T_B}{T_C} = \frac{27}{8}$$

The ratio of Boyle's temperature and critical temperature for gas is  $\frac{27}{8}$ .

**LEVEL – 2**

**Measurable properties of gases and Gas Laws**

- Which one of the following statements is not correct about the three states of matter i.e. solid, liquid and gas?
  - Molecules of a solid possess least energy whereas those of a gas possess highest energy
  - The density of solid is highest whereas that of gases is lowest
  - Gases like liquids possess definite volumes**
  - Molecules of a solid possess vibratory motion

Gases do not have definite shape and volume. Their volume is equal to the volume of the container.
- A volume of  $1 \text{ m}^3$  is equal to
  - $1000 \text{ cm}^3$
  - $100 \text{ cm}^3$
  - $10 \text{ cm}^3$
  - $10^6 \text{ cm}^3$**

$1 \text{ m}^3 = 10^3 \text{ dm}^3 = 10^3 \text{ L} = 10^6 \text{ cm}^3 = 10^6 \text{ m}^2$
- Which one of the following is not a unit of pressure
  - Newton**
  - Torr
  - Pascal
  - Bar

Newton is unit of force.
- $1^\circ\text{C}$  rise in temperature is equal to a rise of
  - 274 K
  - 1 K**
  - $5/9^\circ \text{ F}$
  - $33^\circ \text{ F}$

If  $\Delta T = 1^\circ\text{C}$  then  $\Delta T = 1\text{K}$ .
- Which one of the following statements is wrong for gases
  - Gases do not have a definite shape and volume
  - Volume of the gas is equal to the volume of the container confining the gas
  - Confined gas exerts uniform pressure on the walls of its container in all directions
  - Mass of the gas cannot be determined by weighing a container in which it is enclosed**

The mass of gas can be determined by weighing the container, filled with gas and again weighing the container after removing the gas. The difference between the two masses gives the mass of the gas.
- In a closed flask of 5 litres, 1.0g of  $\text{H}_2$  is heated from 300 to 600 K. Which statement is not correct?
  - Pressure of the gas increases
  - The rate of collision increases
  - The number of moles of gas increases**
  - The energy of gaseous molecules increases

In a closed flask, on heating the no. of moles does not change, as long as the gas doesn't dissociate.
- Which is not true in case of an ideal gas?
  - It cannot be converted into a liquid
  - There is no interaction between the molecules
  - All molecules of the gas move with same speed**
  - At a given temperature, PV is proportional to the amount of the gas

Molecules of a gas move with different speeds which keep changing after they collide with other molecules.

8. In the ideal gas equation, the gas constant R may have a unit of  
 (1) Mole-atm/degree K (2) litre mole  
 (3) litre-atm/degree K/mole (4) erg/degree K

$$R = 0.0821 \text{ L-atm mol}^{-1} \text{ K}^{-1}$$

9. In the equation  $PV = nRT$ , which one cannot be the numerical value of R?

- (1)  $8.31 \times 10^7 \text{ erg K}^{-1} \text{ mol}^{-1}$  (2)  $1.98 \text{ cal mol}^{-1} \text{ K}^{-1}$   
 (3)  $8.31 \text{ JK}^{-1} \text{ mol}^{-1}$  (4)  $8.31 \text{ L-atm.mol}^{-1} \text{ K}^{-1}$

$$R = 0.0821 \text{ L-atm mol}^{-1} \text{ K}^{-1}$$

10. The correct value of the gas constant 'R' is close to:

- (1) 0.082 litre-atmosphere K (2)  $0.082 \text{ litre-atmosphere K}^{-1} \text{ mol}^{-1}$   
 (3)  $0.082 \text{ litre-atmosphere}^{-1} \text{ K mole}^{-1}$  (4)  $0.082 \text{ litre}^{-1} \text{ atmosphere}^{-1} \text{ K mol}$

$$R = 0.082 \text{ litre atm K}^{-1} \text{ mol}^{-1}$$

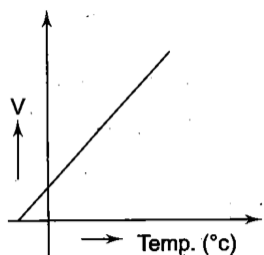
11. Under what conditions will a pure sample of an ideal gas not only exhibit a pressure of 1 atm but also a concentration of 1 mole litre<sup>-1</sup>? (R = 1/12 litre atm mol<sup>-1</sup> deg<sup>-1</sup>)

- (1) At STP (2) When V = 22.4 litres  
 (3) When T = 12K (4) Impossible under any conditions

$$PV = nRT \text{ or } P = \frac{n}{V} RT = CRT.$$

$$\text{Hence } 1 = 1 \times 0.082 \times T \text{ is } T = \frac{1}{0.082} = 12\text{K}.$$

12. The following graph illustrates



- (1) Dalton's law (2) Charles' law (3) Boyle's law (4) Gay-Lussac's law  
 Charles' law deals with variation of volume with temperature at constant pressure.

13. For an ideal gas, number of moles per litre in terms of its pressure P, gas constant R and temperature T is

- (1)  $PT/R$  (2)  $PRT$  (3)  $P/RT$  (4)  $RT/P$

$$PV = nRT \text{ or } \frac{n}{V} = \frac{P}{RT} \text{ i.e. no of moles per litre} = P / RT.$$

14. If pressure becomes double at the same absolute temperature on 2L CO<sub>2</sub>, then the volume of CO<sub>2</sub> becomes

- (1) 2L                      (2) 4L                      (3) 5L                      **(4) 1L**

Pressure is double, volume is halved (at const. temp.)

15. The volume of a gas is 100 mL at 100°C. If pressure remains constant then at what temperature it will be 200 mL?

- (1) 200°C                      **(2) 473°C**                      (3) 746°C                      (4) 50°C

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}, \frac{100}{373} = \frac{200}{T_2}, T_2 = 746\text{K} = 473^\circ\text{C}$$

16. The density of methane at 2.0 atm pressure and 27°C is:

- (1) 0.13 g L<sup>-1</sup>                      (2) 0.26 g L<sup>-1</sup>                      **(3) 1.30 g L<sup>-1</sup>**                      (4) 26.0 g L<sup>-1</sup>

$$d = \frac{PM}{RT} = \frac{2 \times 16}{0.082 \times 300} = 1.30\text{gL}^{-1}$$

17. The volume of 2.8 g of carbon monoxide at 27°C and 0.821 atm pressure is (R=0.0821 lit atm K<sup>-1</sup> mol<sup>-1</sup>)

- (1) 0.3 litre                      (2) 1.5 litre                      **(3) 3 litre**                      (4) 30 litre

$$2.8\text{g CO} = 2.8 / 28\text{mol} = 0.1\text{mol},$$

$$PV = nRT \quad \text{or}$$

$$V = \frac{nRT}{P} = \frac{0.1 \times 0.0821 \times 300}{0.821} = 3\text{litre}.$$

18. A gas is found to have a density of 1.80 g L<sup>-1</sup> at 760 mm pressure and 27°C. The gas may be

- (1) O<sub>2</sub>                      **(2) CO<sub>2</sub>**                      (3) NH<sub>3</sub>                      (4) SO<sub>2</sub>

$$M = d \frac{RT}{P} = \frac{1.8 \times 0.082 \times 300}{760 / 760} \approx 44$$

Hence the gas is CO<sub>2</sub>.

19. 4.0g of argon (at mass = 40) in a bulb at a temperature of TK has a pressure P atm. When the bulb was placed in hotter bath at a temperature 50° more than the first one, 0.8g of gas had to be removed to get the original pressure. T is equal to

- (1) 510 K                      **(2) 200 K**                      (3) 100 K                      (4) 73 K

$$n_i = \frac{4}{40} = 0.1 \qquad n_f = \frac{3.2}{40} = 0.08$$

$$T_i = TK \qquad T_f = (T + 50)K$$

$$\therefore 0.1T = 0.08(T + 50)$$

$$\text{or } 10T = 8T + 400$$

$$\therefore T = 200K.$$

20. An evacuated bulb of  $0.1 \text{ m}^3$  size is filled with  $\text{UF}_6$  at  $300 \text{ K}$  and  $2.46 \times 10^{-2} \text{ atm}$  pressure. The mass of  $\text{UF}_6$  in the bulb is (mol mass of  $\text{UF}_6$  is  $350$  and  $R$  is  $0.082 \text{ litre atm deg}^{-1} \text{ mol}^{-1}$ )

- (1)  $17.5\text{g}$                       (2)  $35\text{g}$                       (3)  $1.75\text{g}$                       (4)  $3.5\text{g}$

$$PV = nRT = \frac{w}{M}RT \text{ or } w = \frac{PMV}{RT}$$

$$= \frac{2.46 \times 10^{-2} \times 350 \times 0.1 \times 10^3}{0.082 \times 300}$$

$$= 35\text{g} (1\text{m}^3 = 10^3\text{L})$$

### Dalton's Law of partial pressure

21. If the density of air at  $298 \text{ K}$  and  $101.325 \text{ kPa}$  is  $1.161 \text{ kg m}^{-3}$ , then assuming air to behave like an ideal gas, the average molar mass of air ( $\text{g mol}^{-1}$ ) is

(1)  $\frac{1.161 \times 0.821 \times 298}{101.325}$                       (2)  $\frac{1.161 \times 8.314 \times 298}{101.325}$

(3)  $\frac{1.161 \times 8.314 \times 298}{101.325 \times 10^3}$                       (4)  $\frac{1.61 \times 0.0821 \times 298}{101.325 \times 10^3}$

$$m = d \frac{RT}{P} = \frac{1.61 \times 8.314 \times 298}{101.325} \text{ g/mol.}$$

22. Equal volumes of two gases which do not react together are enclosed in separate vessels. Their pressures at  $100 \text{ mm}$  and  $400 \text{ mm}$  respectively. If the two vessels are joined together, then what will be the pressure of the resulting mixture (temperature remaining constant)?

- (1)  $125 \text{ mm}$                       (2)  $500 \text{ mm}$                       (3)  $1000 \text{ mm}$                       (4)  $250 \text{ mm}$

After mixing, volume is doubled, so pressures are halved. Hence total pressure after mixing =  $50 + 200 = 250 \text{ mm}$ .

23. Equal weights of ethane and hydrogen are mixed in an empty container at  $25^\circ\text{C}$ . The fraction of the total pressure exerted by hydrogen is

- (1)  $1 : 2$                       (2)  $1 : 1$                       (3)  $1 : 16$                       (4)  $15 : 16$

$$n(\text{C}_2\text{H}_6) = \frac{w}{30}; n(\text{H}_2) = \frac{w}{2}$$

$$(\text{H}_2) = \frac{w/2}{w/2 + w/20} = \frac{1}{1 + \frac{1}{15}} = \frac{15}{16} = \frac{P_{\text{H}_2}}{P_{\text{tot}}}$$

24. Pressure of a mixture of 4g of O<sub>2</sub> and 2g of H<sub>2</sub> confined in a bulb of 1 litre at 0°C is  
 (1) 25.215 atm      (2) 31.205 atm      (3) 45.215 atm      (4) 15.210 atm

$$n_{\text{O}_2} = \frac{4}{32}, n_{\text{H}_2} = \frac{2}{2} \text{ Total } n = \frac{9}{8}$$

$$PV = nRT, P = \frac{9}{8} \times 0.0821 \times \frac{273}{1} = 25.215 \text{ atm}$$

25. If three unreactive gases having partial pressures P<sub>A</sub>, P<sub>B</sub> and P<sub>C</sub> and their moles are 1, 2, and 3 respectively then their total pressure will be:

(1)  $P = P_A + P_B + P_C$       (2)  $P = \frac{P_A + P_B + P_C}{6}$

(3)  $P = \frac{\sqrt{P_A + P_B + P_C}}{3}$       (4) none

The total pressure is sum of partial pressures irrespective of the number of their moles.

26. What is the ratio of masses of oxygen and helium taken together in a vessel with a total pressure of 10 atm given the partial pressure of He in the vessel is 2 atm?  
 (1) 32      (2) 8      (3) 4      (4) 16

$$P_{\text{tot}} = 10\text{atm and } P_{\text{He}} = 2\text{atm}$$

$$\therefore X_{\text{He}} = \frac{2}{10} = 0.2$$

$$\therefore X_{\text{O}_2} = 1 - 0.2 = 0.8$$

$$\text{Now, } \frac{X_{\text{He}}}{X_{\text{O}_2}} = \frac{N_{\text{He}}}{N_{\text{O}_2}} = \frac{M_{\text{He}}}{4} \times \frac{32}{M_{\text{O}_2}}$$

$$\therefore \frac{0.2}{0.8} = \frac{M_{\text{He}}}{M_{\text{O}_2}} \times 8$$

$$\therefore \frac{M_{\text{O}_2}}{M_{\text{He}}} = 32$$

27. 3.2 g of oxygen (At. wt. = 16) and 0.2 g of hydrogen (At. wt. = 1) are placed in a 1.12 litre flask at 0°C. The total pressure of the gas mixture will be  
 (1) 1 atm      (2) 4 atm      (3) 3 atm      (4) 2 atm



$$3.2\text{g O}_2 = 0.1\text{mol}, 0.2\text{g H}_2 = 0.1\text{mol},$$

$$\text{Total } n = 0.2\text{mol},$$

$$P = \frac{nRT}{V} = \frac{0.2 \times 0.082 \times 273}{1.12} = 4\text{atm}$$

28. A sample of  $\text{O}_2$  gas is collected over water at  $23^\circ\text{C}$  at a barometric pressure of 751 mm Hg (vapour pressure of water at  $23^\circ\text{C}$  is 21 mm Hg). The partial pressure of  $\text{O}_2$  gas in the sample collected is  
 (1) 21 mm Hg      (2) 751 mm Hg      (3) 0.96 atm      (4) 1.02 atm  
 $(751 - 21)\text{mm} = 730\text{mmHg}$

$$\therefore P_{\text{O}_2} = \frac{730}{760} = 0.96\text{atm}$$

29. Two glass bulbs A and B are connected by very small tube having a stop cock. Bulb A has a volume of 100 ml and contains a gas while bulb B was empty and had a volume of 150 ml. on opening the stop-cock, the pressure of the gas in bulb A will fall down to  
 (1) 80%      (2) 60%      (3) 40%      (4) 20%

Suppose initially, the pressure of the gas in bulb A = P mm, Initial volume = 100 ml. After opening the stop - cock, final volume =  $100 + 150 = 250$  ml.

$$\text{Now } P_1V_1 = P_2V_2 \text{ i.e., } P \times 100 = P_2 \times 250$$

$$\text{Or } P_2 = \frac{100}{250}P = 0.40P = 40\% \text{ of } P.$$

30. Air contains 79%  $\text{N}_2$  and 21%  $\text{O}_2$  by volume. If the barometric pressure is 750 mm Hg, the partial pressure of oxygen is  
 (1) 157.5 mm of Hg      (2) 175.5 mm of Hg      (3) 3125.0 mm of Hg      (4) None of these

Under similar conditions of temp. and pressure, equal volumes contain equal no. of moles. Hence mole fraction of  $\text{O}_2 = 0.21$

$$\therefore \text{Partial pressure of } \text{O}_2 = 0.21 \times 750 = 157.5\text{mm of Hg.}$$

### GRAHAM LAW OF DIFFUSION

31. Which of the following pairs will diffuse at the same rate through a porous plug?  
 (1)  $\text{CO}, \text{NO}_2$       (2)  $\text{NO}_2, \text{CO}_2$       (3)  $\text{NH}_3, \text{PH}_3$       (4)  $\text{NO}, \text{C}_2\text{H}_6$

Gases with same M.W. diffuse at the same rate under similar conditions.

32. At constant volume and temperature conditions, the rate of diffusion  $D_A$  and  $D_B$  of gases A and B having density  $\rho_A$  and  $\rho_B$  are related by the expression,

$$(1) D_A = \left[ D_B \cdot \frac{\rho_A}{\rho_B} \right]^{1/2} \qquad (2) D_A = \left[ D_B \cdot \frac{\rho_B}{\rho_A} \right]^{1/2}$$

$$(3) D_A = D_B \left( \frac{\rho_A}{\rho_B} \right)^{1/2}$$

$$(4) D_A = D_B \left( \frac{\rho_B}{\rho_A} \right)^{1/2}$$

$$\frac{D_A}{D_B} = \left( \frac{\rho_B}{\rho_A} \right)^{1/2} \text{ by Graham's law of diffusion.}$$

33. According to Graham's law at a given temperature, the ratio of the rates of diffusion  $r_A / r_B$  of gases A and B having unequal partial pressure is given by

(where P and M are the pressure and molecular weights of gases A and B respectively)

$$(1) (P_A / P_B)(M_A / M_B)^{1/2}$$

$$(2) (M_A / M_B)(P_A / P_B)^{1/2}$$

$$(3) (P_A / P_B)(M_B / M_A)^{1/2}$$

$$(4) (M_A / M_B)(P_B / P_A)^{1/2}$$

At constant temperature, rate of diffusion  $\propto \sqrt{\frac{1}{d}}$  or  $\sqrt{\frac{1}{M}}$  and  $\propto P$ . Hence

$$r_A / r_B = (P_A / P_B)(M_B / M_A)^{1/2}$$

34. If some moles of  $O_2$  diffuse in 18 sec and same moles of other gas diffuse in 45 sec then what is the molecular weight of the unknown gas

$$(1) \frac{45^2}{18^2} \times 32$$

$$(2) \frac{18^2}{45^2} \times 32$$

$$(3) \frac{18^2}{45^2 \times 32}$$

$$(4) \frac{45^2}{18^2 \times 32}$$

$$\frac{r_{O_2}}{r_x} = \sqrt{\frac{M_x}{M_{O_2}}}$$

As volumes are in the same ratio as the no. of moles

$$\frac{n/18}{n/45} = \sqrt{\frac{M_x}{32}}$$

$$\frac{45}{18} \sqrt{\frac{M_x}{32}}$$

$$M_x = \frac{45^2}{18^2} \times 32$$

35. 50 ml of gas A diffusion through a membrane in the same time as 40ml of a gas B under identical pressure-temperature conditions. If the molecular weight of A is 64, M.W. of B would be

$$(1) 100$$

$$(2) 250$$

$$(3) 200$$

$$(4) 80$$

$$\frac{r_A}{r_B} = \sqrt{\frac{M_B}{M_A}}, \frac{50/t}{40/t} = \sqrt{\frac{M_B}{64}}$$

$$\frac{50}{40} = \sqrt{\frac{M_B}{64}} \text{ or } M_B = 100$$

36. The rate of diffusion of methane at a given temperature is twice than of gas X. The molecular weight of X is

- (1) 64.0                      (2) 32.0                      (3) 4.0                      (4) 8.0

$$\frac{r_{CH_4}}{r_x} = \sqrt{\frac{M_X}{M_{CH_4}}}$$

$$2 = \sqrt{\frac{M_X}{16}},$$

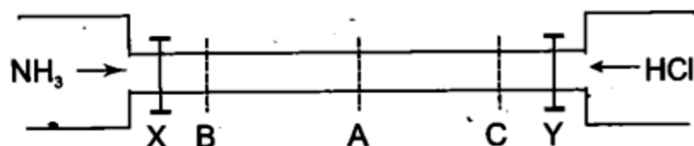
$$\Rightarrow M_X = 64.$$

37. The density of methane at 2.0 atmosphere pressure and 27°C is

- (1) 0.13 gL<sup>-1</sup>              (2) 0.26 gL<sup>-1</sup>              (3) 1.30 gL<sup>-1</sup>              (4) 26.0 gL<sup>-1</sup>

$$d = \frac{PM}{RT} = \frac{2 \times 16}{0.082 \times 300} = 1.30 \text{ gL}^{-1}$$

38.



The valves of X and Y are opened simultaneously. The white fumes of NH<sub>4</sub>Cl will first form at:

- (1) A                      (2) B  
 (3) C                      (4) A, B and C simultaneously

$$r \propto \frac{1}{\sqrt{M}} \text{ So, NH}_3 \text{ diffuses with faster rate.}$$

39. For 10 minutes each, at 0°C from two identical holes nitrogen and a unknown gas are leaked into a common vessel of 3L capacity under identical conditions. The resulting pressure is 5.6 atm and the mixture contains 0.5 mole of nitrogen. Then the molar mass of the unknown gas is:

- (1) 224 g mol<sup>-1</sup>              (2) 112 g mol<sup>-1</sup>              (3) 400 g mol<sup>-1</sup>              (4) 540 g mol<sup>-1</sup>

$$n_T = \frac{PV}{RT} = \frac{5.6 \times 3}{0.0821 \times 273}$$

$$n_T = 0.75$$

$$\text{Moles of unknown gas} = 0.75 - 0.5 = 0.25$$

$$\frac{n_g}{n_{N_2}} \cdot \frac{t_{N_2}}{t_g} = \sqrt{\frac{M_{N_2}}{M_g}} \Rightarrow \frac{1}{2} \sqrt{\frac{28}{M_g}}$$

$$\Rightarrow M_g = 112 \text{ g mol}^{-1}$$

40. 20 L so  $\text{SO}_2$  diffuses through a porous partition in 60 seconds. Volume of  $\text{O}_2$  that will diffuse under similar conditions in 30 seconds will be:

- (1) 12.14 L                      (2) 14.14                      (3) 18.14 L                      (4) 28.14 L

$$\frac{20}{60} \times \frac{30}{V} = \sqrt{\frac{32}{64}} \Rightarrow V = 10\sqrt{2}$$

**KINETIC THEORY OF GASES**

41. The r.m.s. velocity of gases can be calculated from the expression

- (1)  $\sqrt{\frac{3P}{\rho}}$                       (2)  $\sqrt{\frac{3PV}{mass}}$                       (3)  $\sqrt{\frac{3RT}{M.W}}$                       (4) All the above

$$u_{RMS} = \sqrt{\frac{3RT}{M.W.}}$$

Now,  $PV = nRT$

$$\therefore RT = \frac{PV}{n}$$

$$\therefore u_{RMS} = \sqrt{\frac{3PV}{n.(MW)}} = \sqrt{\frac{3PV}{\frac{mass}{MW}.MW}}$$

$$\therefore u_{RMS} = \sqrt{\frac{3PV}{mass}}$$

$$\text{Also, } u_{RMS} = \sqrt{\frac{3P}{\frac{mass}{Vol}}} = \sqrt{\frac{3P}{\zeta}}$$

42. Root mean square velocity of a gas molecule is proportional to (m: mass of gas taken)

- (1)  $m^{1/2}$                       (2)  $m^0$                       (3)  $m^{-1/2}$                       (4) m

$$u_{RMS} = \sqrt{\frac{3RT}{M}}$$

where M is the M.W. of the gas. Thus,  $u_{RMS}$  is independent of mass of the gas taken.

Hence,  $u_{RMS} \propto m^0$  where m is mass of gas taken.

43. Helium atom is two times heavier than a hydrogen molecule at 298K, the average kinetic energy of helium is

- (1) two times that of a hydrogen molecule                      (2) same as that of a hydrogen molecule  
 (3) four times that of a hydrogen molecule                      (4) half that of a hydrogen molecule

Average kinetic energy depends only on temperature and does not depend upon the nature of the gas.

44. By how many folds the temperature of a gas would have to increase when the r.m.s. velocity of gas molecules in a closed container of fixed volume is to be increased from  $5 \times 10^4 \text{ cms}^{-1}$  to  $10 \times 10^4 \text{ cms}^{-1}$

(1) 0.5 times                      (2) 2 times                      **(3) 4 times**                      (4) 16 times

$$\frac{u_1}{u_2} = \sqrt{\frac{T_1}{T_2}}$$

$$\frac{5 \times 10^4}{10 \times 10^4} = \sqrt{\frac{T_1}{T_2}} \text{ or } \frac{T_1}{T_2} = \frac{1}{4}$$

Or  $T_2 = 4 \times T_1$

45. If  $C_1, C_2, C_3, \dots$  represent the speed of  $n_1, n_2, n_3, \dots$  molecules, then the root mean square speed is

(1)  $\left( \frac{n_1 C_1^2 + n_2 C_2^2 + n_3 C_3^2 + \dots}{n_1 + n_2 + n_3 + \dots} \right)^{1/2}$

(2)  $\frac{(n_1 C_1^2 + n_2 C_2^2 + n_3 C_3^2 + \dots)^{1/2}}{n_1 + n_2 + n_3 + \dots}$

(3)  $\frac{(n_1 C_1^2)^{1/2}}{n_1} + \frac{(n_2 C_2^2)^{1/2}}{n_2} + \frac{(n_3 C_3^2)^{1/2}}{n_3} + \dots$

(4)  $\frac{(n_1 C_1^2 + n_2 C_2^2 + n_3 C_3^2 + \dots)^{1/2}}{(n_1 + n_2 + n_3 + \dots)}$

By definition,  $u = \frac{n_1 C_1^2 + n_2 C_2^2 + \dots}{n_1 + n_2 + \dots}$

46. The ratio among most probable velocity, average velocity and root mean square velocity is given by

(1) 1 : 2 : 3                      (2)  $1 : \sqrt{2} : \sqrt{3}$                       (3)  $\sqrt{2} : \sqrt{3} : \sqrt{8/\pi}$                       **(4)  $\sqrt{2} : \sqrt{8/\pi} : \sqrt{3}$**

$\alpha : v : u = \sqrt{2} : \sqrt{8/\pi} : 3$

47. The root mean square speed at STP for the gases  $H_2, N_2, O_2$  and HBr are in the order

(1)  $H_2 < N_2 < O_2 < HBr$                       **(2)  $HBr < O_2 < N_2 < H_2$**

(3)  $H_2 < N_2 = O_2 < HBr$                       (4)  $HBr < O_2 < H_2 < N_2$

$$u_{RMS} \propto \sqrt{\frac{1}{M}}$$

Molecular masses of  $H_2, N_2, O_2$  and HBr are 2, 28, 32, and 81.

48. The temperature of the gas is raised from  $27^\circ C$  to  $927^\circ C$ , the root mean square velocity is

(1)  $\sqrt{927/27}$  times the earlier value                      (2) same as before

(3) halved                      **(4) doubled**

$u \propto \sqrt{T}$  where T is in Kelvin

$$u_1 / u_2 = \sqrt{T_1 / T_2} = \sqrt{\frac{27 + 273}{927 + 273}}$$

$$\frac{u_1}{u_2} = \sqrt{\frac{300}{1200}} = \frac{1}{2}, u_2 = 2u_1$$

49. Which of the following expression correctly represents the relationship between the average molar kinetic energy,  $\overline{K.E.}$ , of CO and  $N_2$  molecules at the same temperature?

(1)  $\overline{KE}_{CO} = \overline{KE}_{N_2}$

(2)  $\overline{KE}_{CO} > \overline{KE}_{N_2}$

(3)  $\overline{KE}_{CO} < \overline{KE}_{N_2}$

(4) cannot be predicted unless the volume of the gases are given

$\overline{K.E.} \propto T$

Hence at the same temperature,  $\overline{KE}_{CO} = \overline{KE}_{N_2}$

50. The r.m.s velocity of hydrogen is  $\sqrt{7}$  times the r.m.s velocity of nitrogen. If T is the temperature of the gas,

(1)  $T(H_2) = T(N_2)$

(2)  $T(H_2) > T(N_2)$

(3)  $T(H_2) < T(N_2)$

(4)  $T(H_2) = \sqrt{7}T(N_2)$

$$u = \sqrt{\frac{3RT}{M}} \therefore \frac{u(H_2)}{u(N_2)} = \sqrt{\frac{T(H_2)}{M(H_2)} \times \frac{M(N_2)}{T(N_2)}}$$

$$\text{Or } \sqrt{7} = \sqrt{\frac{T(H_2)}{T(N_2)} \times \frac{28}{2}}$$

$$\text{Or } 7 = \frac{T(H_2)}{T(N_2)} \times 14 \text{ or } \frac{T(H_2)}{T(N_2)} = \frac{1}{2}$$

$$\text{Or } T(N_2) = 2 \times T(H_2) \text{ i.e. } T(N_2) > T(H_2)$$

51. A gas container observes Maxwell distribution of speeds. If the number of molecules between the speed 5 and 5.1 km sec<sup>-1</sup> 25<sup>o</sup> C be 'n' what would be the number of molecules between this range of speed if the total number of molecules in the vessel are doubled?

(1)  $n^2$

(2) n

(3) n/2

(4) 2n

At a particular temp. the fraction of molecules possessing a particular speed remains constant. As the total no. of molecules is doubled, to keep the fraction same, no. of molecules possessing speed in the same range is also doubled.

52. The ratio of root mean square velocity to average velocity of gas molecules at a particular temperature is

(1) 1.086:1

(2) 1:1.086

(3) 2:1.086

(4) 1.086:2

$$u_{AVG} = \sqrt{\frac{8RT}{\pi M}} = \sqrt{3} : \sqrt{8/\pi}$$

$$= \sqrt{3\pi/8} : 1$$

$$= 1.086 : 1$$

53. If a gas is expanded at constant temperature
- (1) the pressure increases
  - (2) the kinetic energy of the molecules remains the same
  - (3) the kinetic energy of the molecules decreases
  - (4) the number of molecules of the gas increases
- At constant temp., the K.E. of the molecules remains the same.

**REAL GASES**

54. Gases deviate from the ideal gas behavior because their molecules
- (1) possess non-zero volume
  - (2) have force of attraction between them
  - (3) experience force of repulsion from each other
  - (4) all of the above
- Gases deviate from ideal behaviour because of intermolecular forces of attraction and repulsion and also the volume of these gas molecules are non-zero.

55. In van der Waal's equation of state for a non-ideal gas, the term that accounts for intermolecular forces of attraction is :
- (1)  $(V - b)$
  - (2)  $(RT)^{-1}$
  - (3)  $\left(P + \frac{a}{V^2}\right)$
  - (4)  $RT$
- Intermolecular forces of attraction are given by the term  $\left(P + \frac{a}{V^2}\right)$  which is also called the pressure correction factor.

56. The Vanderwaal's constant 'a' for the gases  $O_2, N_2, NH_3$  and  $CH_4$  are 1.3, 1.390, 4.170 and 2.253  $L^2 atm mol^{-2}$  respectively. The gas which can be most easily liquefied is
- (1)  $O_2$
  - (2)  $N_2$
  - (3)  $NH_3$
  - (4)  $CH_4$
- 'a' is directly related to forces of attraction. Hence greater the value of 'a', more easily the gas is liquefied.

57. The Vanderwaal's parameters for gases W, X, Y and Z are

Gas	A	B
W	4.0	0.027
Z	8.0	0.030

Y	6.0	0.032
Z	12.0	0.027

Which one of these gases has the highest critical temperature?

- (1) W                                      (2) X                                      (3) Y                                      (4) Z

$$T_c = \frac{8a}{27Rb} \text{ Thus } T_c \propto \frac{a}{b}$$

58. The correct order of normal boiling points of  $O_2, N_2, NH_3$  and  $CH_4$ , for whom the values of Vanderwaal's constant 'a' are 1.360, 1.390, 4.170 and  $2.253 \text{ L}^2 \cdot \text{atm} \cdot \text{mol}^{-2}$  respectively, is

- (1)  $O_2 < N_2 < NH_3 < CH_4$                                       (2)  $O_2 < N_2 < CH_4 < NH_3$   
 (3)  $NH_3 < CH_4 < N_2 < O_2$                                       (4)  $NH_3 < CH_4 < O_2 < N_2$

Boiling point  $\propto a$

59. If V is the volume of one molecules of a gas, then Vanderwaals constant 'b' is : ( $N_0$  = Avogadro's number)

- (1)  $\frac{4V}{N_0}$                                       (2) 4V                                      (3)  $\frac{N_0}{4V}$                                       (4)  $4VN_0$

We know  $6 = 4 \left( \frac{4}{3} \pi r^3 \right) N_0$

Now, if  $V = \frac{4}{3} \pi r^3$  then  $b = 4VN_0$

60. In Vanderwaal's equation of state for a non-ideal gas, the term that accounts for intermolecular forces of repulsion and volume of the molecules is

- (1) (V- b)                                      (2) RT                                      (3)  $\left( P + \frac{a}{V^2} \right)$                                       (4)  $(RT)^{-1}$

The volume correction factor takes into account the forces of repulsion and the volume of the molecules.

61. On heating vapours of  $S_8(g)$  they decomposes to  $S_2(g)$ . Due to this, the Vanderwaal's constant 'b' for the resulting gas.

- (1) increases                                      (2) decreases  
 (3) remains same                                      (4) changes unpredictably

As  $S_8$  decomposes to  $S_2$ , the size of the molecules decreases. Thus, 'b' decreases.

62. What is the compressibility factor (Z) for 0.02 mole of a Vanderwaals' gas at pressure of 0.1 atm. Assume the size of gas molecules is negligible.

Given  $RT = 20 \text{ L atm mol}^{-1}$  and  $a = 1000 \text{ atm L}^2 \text{ mol}^{-2}$

- (1) 2                                      (2) 1                                      (3) 0.02                                      (4) 0.5



If volume correction factor is ignored, then attraction dominates.

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

$$PV^2 + an^2 = nVRT$$

$$\text{Or } 0.1V^2 - 0.02V(20) + 1000(0.02)^2 = 0$$

$$\text{Or } 0.1V^2 - 0.4V + 0.4 = 0$$

$$\text{Or } V^2 - 4V + 4 = 0$$

$$\text{Or } (V - 2)^2 = 0 \Rightarrow V = 2$$

$$\therefore Z = \frac{PV}{nRT} = \frac{0.1 \times 2}{0.02 \times 20} = 0.5$$

63. One way of writing the equation of state for real gases is  $PV = RT \left[ 1 + \frac{B}{V} + \dots \right]$  where B is a constant. An approximate expression for B in terms of the van der Waals constant 'a' and 'b' is

(1)  $B = a - \frac{b}{RT}$       (2)  $B = b - \frac{a}{RT}$       (3)  $B = RT - \frac{a}{b}$       (4)  $B = \frac{b}{a}$

Vanderwaal's equation can be written as

$$Z = 1 + \left( b - \frac{a}{RT} \right) V_m$$

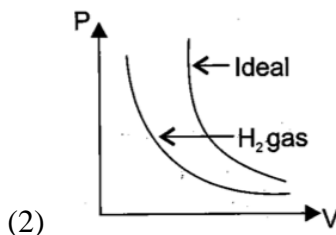
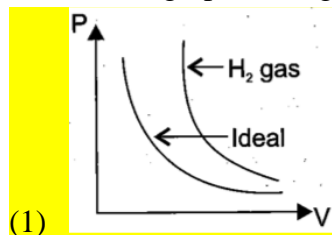
So by comparing coefficient of  $V_m = b - \frac{a}{RT}$

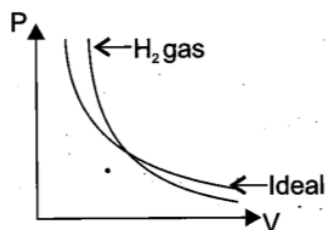
64. A real gas obeying Vanderwaal's equation will resemble ideal gas, if the:

(1) constants a and b are small      (2) a is large and b is small  
 (3) a is small and b is large      (4) constant a and b are large ]

If a and b are negligibly small then the Vanderwaal's equation of state transforms to ideal gas equation

65. The correct graph to H<sub>2</sub> gas is:





- (3) (4) All of these

For H<sub>2</sub> and He repulsive forces dominate even at low pressure. Thus, at high as well as low pressures  $V_{H_2} > V_{ideal\ gas}$

66. For the non-zero values of force of attraction between gas molecules, gas equation will be:

- (1)  $PV = nRT - \frac{n^2 a}{V}$  (2)  $PV = nRT + nbP$  (3)  $PV = nRT$  (4)  $P = \frac{nRT}{V - b}$

$$Z = \frac{PV}{nRT}; \quad V = \frac{0.9 \times 0.0821 \times 273}{9}$$

$$= 2.24 \text{ litre / mol}$$

∴ Volume of 1 mili – mole of gas = 2.24mL

67. Compressibility factor for H<sub>2</sub> behaving as real gas is:

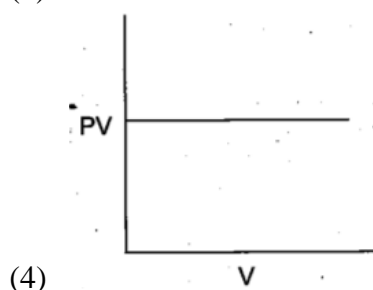
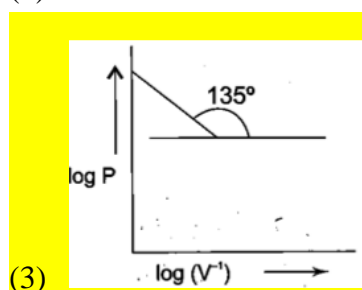
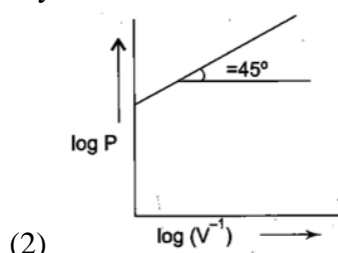
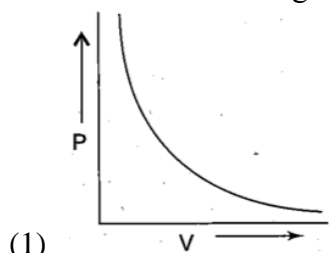
- (1) 1 (2)  $\left(1 - \frac{a}{RTV}\right)$  (3)  $\left(1 + \frac{Pb}{RT}\right)$  (4)  $\frac{RTV}{(1-a)}$

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

$$\therefore PV = nRT - \frac{an^2}{V}$$

### Gas Laws and ideal Gas Equation

68. Which of the following curve does not represent Boyle's law?



At constant temperature  $P \propto 1/V \Rightarrow PV = \text{constant}$

$$\therefore \log P + \log V = \log C.$$

$$\therefore \log P + \log V = \log C$$

$$\text{or, } \log P = \log \left( \frac{1}{V} \right) + \log C$$

69. A gas is heated from  $0^\circ\text{C}$  to  $100^\circ\text{C}$  at 1.0 atm pressure. If the initial volume of the gas is 10.0 L, its final volume would be :

- (1) 7.32 L                      (2) 10.00 L                      (3) 13.66 L                      (4) 20.00 L

$$\frac{10}{V_2} = \frac{273}{373} \Rightarrow V_2 = 13.66\text{L}$$

70. If the pressure of a gas contained in a closed vessel is increased by 0.4% when heated by  $1^\circ\text{C}$  its initial temperature must be:

- (1) 250 K                      (2)  $250^\circ\text{C}$                       (3)  $25^\circ\text{C}$                       (4) 25 K

$$\frac{100}{100.4} = \frac{T}{T+1} \Rightarrow T = \frac{100}{0.4} = 250\text{K}$$

71. An amount of 1.00 g of a gaseous compound of boron and hydrogen occupies 0.820 liter at 1.00 atm and at  $3^\circ\text{C}$ . The compound is ( $R = 0.0820$ ) liter atm mole<sup>-1</sup> K<sup>-1</sup>; at. wt: H = 1.0, B = 10.8).

- (1)  $\text{BH}_3$                       (2)  $\text{B}_4\text{H}_{10}$                       (3)  $\text{B}_2\text{H}_6$                       (4)  $\text{B}_3\text{H}_{12}$

$$PV = nRT$$

$$PV = \frac{1}{M}RT \Rightarrow M = \frac{0.082 \times 276}{0.82}$$

$$= 27.6 \Rightarrow \text{B}_2\text{H}_6$$

72. A  $0.5\text{ dm}^3$  flask contains gas A and  $1\text{ dm}^3$  flask contains gas B at the same temperature. If density of A =  $3\text{ g/dm}^3$  and that of B =  $1.5\text{ g/dm}^3$  and the molar mass of A =  $\frac{1}{2}$  of B, the ratio of pressure exerted by gases is:

- (1)  $\frac{P_A}{P_B} = 2$                       (2)  $\frac{P_A}{P_B} = 1$                       (3)  $\frac{P_A}{P_B} = 4$                       (4)  $\frac{P_A}{P_B} = 3$

$$P_A = \frac{3RT}{M_A}; \quad P_B = \frac{1.5RT}{M_B}$$

$$\frac{P_A}{P_B} = \frac{2M_B}{M_A} = \frac{2 \times 2M_A}{M_A} = 4$$

73. A and B are two identical vessels. A contains 15 g ethane at 1 atm and 298 K. the vessel B contains 75 g of a gas  $\text{X}_2$  at same temperature and pressure. The vapour density of  $\text{X}_2$  is:

- (1) 75                      (2) 150                      (3) 37.5                      (4) 45

Since, P, V, T are same therefore number of moles must also be the same

$$\frac{15}{30} = \frac{75}{M_B}$$

$$M_B = 150 \Rightarrow (V.D.)_B = \frac{150}{2} = 75$$

74. The density of neon will be highest at:

- (1) STP                      (2) 0°C, 2 atm                      (3) 273°C, 1 atm                      (4) 273°C, 2 atm

$$\rho \propto \frac{P}{T}$$

75. 3 moles of N<sub>2</sub> gas at 27°C and 3 atm pressure has same volume as 5 mole of H<sub>2</sub> gas at 127°C. Then the pressure of H<sub>2</sub> gas is

- (1)  $\frac{10}{3}$  atm                      (2)  $\frac{5}{3}$  atm                      (3) 20 atm                      (4)  $\frac{20}{3}$  atm

$$\frac{n_1 T_1}{P_1} = \frac{n_2 T_2}{P_2} \Rightarrow \frac{3 \times 300}{3} = \frac{5 \times 400}{P_2} \quad P_2 = \frac{20}{3} \text{ atm}$$

### ASSERTION & REASON

Each question has 5 choices (1), (2), (3), (4) and (E) out of which ONLY ONE is correct.

- (1) Statement-1 is true, statement-2 is true; statement-2 is a correct explanation for statement
- (2) Statement-1 is true, statement-2 is true; statement-2 is NOT a correct explanation for statement
- (3) Statement-1 is true, statement-2 is false
- (4) Statement-1 is false, statement-2 is true
- (5) Statement-1 and statement-2 both are false

1. Statement-1: Plot of P Vs. 1/V is a straight line for constant temperature and fixed amount of ideal gas.  
Statement-2: Pressure is directly proportional to volume.  
**(3)**
2. Statement-1: Absolute zero is a theoretically possible temperature at which the volume of the gas becomes zero  
Statement-2: The total kinetic energy of the molecules is zero at this temperature.  
**(2)**
3. Statement-1: In a container containing gas 'A' at temp 400 K, some more gas A at temp. 300 K is introduced the pressure of the system increases.  
Statement-2: increases in gaseous particle increases the number of collisions among the molecules.  
**(3)**
4. Statement-1: Gas with lower molar mass will effuse or diffuse faster.  
Statement-2: Total kinetic energy of any gas depends upon its molar mass.  
**(3)**

5. Statement-1: Pressure exerted by a mixture of gases is equal to the sum of their partial pressures.  
Statement-2: Reacting gases react to form a new gas having pressure equal to the sum of both.  
**(3)**
6. Statement-1: CH<sub>4</sub>, CO<sub>2</sub> has value of Z (compressibility factor) less than one, generally.  
Statement-2: Z < 1 is due to repulsive forces among the molecules.  
**(3)**
7. Statement-1: Critical temperature of the gas is the temperature at which it occupies 22.4 L of volume.  
Statement-2: Molar volume of every gas at NTP is 22.4 L.  
**(5)**
8. Statement-1: Excluded volume or co-volume equals to (v-nb) for n moles gas  
Statement-2: Co-volume depends on the effective size of gas molecules.  
**(4)**
9. Statement-1: Gases like N<sub>2</sub>, O<sub>2</sub> behave as ideal gases at high temperature and low pressure  
Statement-2: Molecular interaction diminished at high temperature and low pressure.  
**(1)**
10. Statement-1: Most probable velocity is like velocity possessed by maximum fraction of molecules at the same temperature.  
Statement-2: On collision, more and more molecules acquire higher speed at the same temperature.  
**(3)**
11. Statement-1: Noble gases can be liquefied.  
Statement-2: Attractive forces can exist between non-polar molecules.  
**(1)**
12. Statement-1: The diffusion rate of oxygen is smaller than that of nitrogen under same conditions of T and P.  
Statement-2: Molecular mass of nitrogen is smaller than that of oxygen.  
**(1)**

**PREVIOUS YEARS QUESTIONS**

1. If P.V.M.T and R are pressure, Volume, Molar mass, temperature and gas constant respectively, then for an ideal gas, the density is given by **[1989]**
- (1)  $\frac{RT}{PM}$                       (2)  $\frac{P}{RT}$                       (3)  $\frac{M}{V}$                       (4)  $\frac{PM}{RT}$
- $PV = nRT = \frac{m}{M} RT$
- Or  $PM = \frac{m}{V} RT = dRT \Rightarrow d = \frac{PM}{RT}$
2. Pressure remaining the same, the volume of a given mass of an ideal gas increases for every degree centigrade rise in temperature by definite fraction of its volume at **[1989]**
- (1) 0°C**                      (2) its critical temperature  
(3) absolute zero                      (4) its Boyle temperature

Charle's Law- The volume of the given mass of a gas increases or decreases by  $\frac{1}{273}$  of its volume at  $0^{\circ}\text{C}$  for each degree rise or fall of temperature at constant pressure.

$$V_t = V_0 \left( 1 + \frac{t}{273} \right) \text{ at constant Pressure}$$

3. Correct gas equation is:

(1)  $\frac{V_1 T_2}{P_1} = \frac{V_2 T_1}{P_2}$       (2)  $\frac{P_1 V_1}{P_2 V_2} = \frac{T_1}{T_2}$       (3)  $\frac{P_1 T_2}{V_1} = \frac{P_2 V_2}{T_2}$       (4)  $\frac{V_1 V_2}{T_1 T_2} = P_1 P_2$  [1989]

$$\frac{PV}{T} = \text{constant or } \frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

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

4. Root mean square velocity of a gas molecule is proportional to [1990]

(1)  $m^{1/2}$       (2)  $m^0$       (3)  $m^{-1/2}$       (4)  $m$

According to kinetic gas equation  $PV = \frac{1}{3} mNu^2$ ,  $u$  = root mean square velocity

$$\Rightarrow u^2 = \frac{3PV}{mN} \text{ or } u \propto \frac{1}{\sqrt{m}} \text{ i.e } u \propto m^{-\frac{1}{2}}$$

5. Absolute zero is defined as the temperature [1990]

- (1) At which all molecular motion ceases      (2) at which liquid helium boils  
 (3) at which ether boils      (4) all of the above

Absolute zero is the temperature at which kinetic energy of gas molecules becomes zero i.e. all molecular motion ceases.

6. In van der Waal's equation of state for a non-ideal gas, the term that accounts for intermolecular forces is [1990]

(1)  $(V - b)$       (2)  $(RT)^{-1}$       (3)  $\left( P + \frac{a}{V^2} \right)$       (4)  $RT$

Van der Waal's equation for 1 mole:

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

Here,  $\left( P + \frac{a}{V^2} \right)$  represent the intermolecular forces and  $(V - b)$  is the corrected volume

7. In a closed flask of 5 litres, 1.0 g of  $\text{H}_2$  is heated from 300 to 600K. Which statement is not correct? [1991]

- (1) Pressure of the gas increases      (2) The rate of collision increases

- (3) The number of moles of gas increases      (4) The energy of gaseous molecules increases  
The number of moles of gas do not changes.

8. At constant temperature, for a given mass of an ideal gas [1991]  
 (1) The ratio of pressure and volume always remains constant.  
 (2) Volume always remains constant.  
 (3) Pressure always remains constant  
 (4) The product of pressure and volume always remains constant.

According to Boyle's law at constant temperature  $P \propto \frac{1}{V}$  or  $PV = \text{constant}$

9. The root mean square speed at STP for the gases  $H_2, N_2, O_2$  and  $HBr$  are in the order: [1991]  
 (1)  $H_2 < N_2 < O_2 < HBr$       (2)  $HBr < O_2 < N_2 < H_2$   
 (3)  $H_2 < N_2 = O_2 < HBr$       (4)  $HBr < O_2 < H_2 < N_2$

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

$$\text{Or } u = \sqrt{3PV / M}$$

$$\text{At STP. } u \propto \sqrt{\frac{1}{M}}$$

i.e. higher will be the molar mass, lower will be the value of  $u_{rms}$ .

Molecular masses of  $H_2, N_2, O_2$  and  $HBr$  are 2, 28, 32 and 81. Hence the correct order of  $u_{rms}$  will be  $HBr < O_2 < N_2 < H_2$

10. A gas is said to behave like an ideal gas when the relation  $PV/T = \text{constant}$ . When do you expect a real gas to behave like an ideal gas? [1991]  
 (1) When the temperature low  
 (2) When both the temperature and pressure are low  
 (3) When both the temperature and pressure are high  
 (4) When the temperature is high and pressure is low

At low pressure and high temperature correction for 1 mole of a gas is negligible, i.e the effect of  $a/v^2$  and  $b$  becomes negligible. Thus, the gas equation becomes

$$PV = nRT$$

$$PV = RT \text{ or } \frac{PV}{RT} = 1 [\text{Ideal gas equation}]$$

$\therefore Z = 1$ , and gas shows ideal behaviour.

11. Which is not true in case of an ideal gas? [1992]  
 (1) It cannot be converted into a liquid  
 (2) There is no interaction between the molecules  
 (3) All molecules of the gas move with same speed

(4) At a given temperature, PV is proportional to the amount of the gas  
Molecules of an ideal gas move with different speeds.

12. The correct value of the gas constant 'R' is close to : [1992]

- (1) 0.082 litre-atmosphere K (2) 0.082 litre-atmosphere K<sup>-1</sup> mol<sup>-1</sup>  
 (3) 0.082 litre – atmosphere<sup>-1</sup> K mol<sup>-1</sup> (4) 0.082 litre<sup>-1</sup> atmosphere<sup>-1</sup> K mol  
 R = 0.082 litre atm K<sup>-1</sup>mol<sup>-1</sup>

13. A closed flask contains water in all its three states solid, liquid and vapour at 0<sup>o</sup>C. In this situation, the average kinetic energy of water molecules will be [1992]

- (1) the greatest in all the three states (2) the greatest in vapour state  
 (3) the greatest in the liquid state (4) the greatest in the solid state

In vapour state, the molecules are free to move with highest average kinetic energy.

14. An ideal gas can't be liquefied because [1992]

- (1) its critical temperature is always above 0<sup>o</sup>C  
 (2) its molecules are relatively smaller in size  
 (3) it solidifies before becoming a liquid  
 (4) force operated between its molecules are negligible

In an ideal gas, the intermolecular forces of attraction are negligible and hence, it cannot be liquefied.

15. Select one correct statement. In the gas equation PV = nRT [1992]

- (1) n is the number of molecules of a gas  
 (2) V denotes volume of one mole of the gas  
 (3) n moles of the gas have a volume V  
 (4) P is the pressure of the gas when only one mole of gas is present

In the equation PV = nRT, n moles of the gas have volume V.

16. Under what conditions will a pure sample of an ideal gas not only exhibit a pressure of 1 atm but also a concentration of 1 mole litre<sup>-1</sup>? (R = 0.082 litre atm mol<sup>-1</sup>deg<sup>-1</sup>) [1993]

- (1) At STP  
 (2) When V = 22.4 litres  
 (3) When T= 12K  
 (4) Impossible under any conditions

$$PV = nRT \text{ or } P = \frac{n}{V} RT = CRT$$

Hence 1 = 1 × 0.082 × T

$$\therefore T = \frac{1}{0.082} = 12K$$



17. When is deviation more in the behavior of a gas from the ideal gas equation  $PV = nRT$ ? [1993]  
 (1) At high temperature and low pressure (2) At low temperature and high pressure  
 (3) At high temperature and high pressure (4) At low temperature and low high pressure  
 At low temperature and high pressure.

18. At STP, 0.50 ml  $H_2$  gas and 10 mol He gas [1993]  
 (1) have equal average kinetic energies (2) have equal molecular speeds  
 (3) occupy equal volumes (4) have equal effusion rates

Average kinetic energy depends only on temperature  $\left( K.E. = \frac{3}{2} KT \right)$

19. Internal energy and pressure of a gas per unit volume are related as  
 (1)  $P = \frac{2}{3} E$  (2)  $P = \frac{3}{2} E$  (3)  $P = \frac{1}{2} E$  (4)  $P = 2E$

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

$$= \frac{2}{3} \cdot \frac{1}{2} M u^2 = \frac{2}{3} E \text{ or } P = \frac{2}{3} E \text{ per unit vol.}$$

20. The ratio among most probable velocity mean velocity and root mean square velocity is given by [1993]

- (1) 1 : 2 : 3 (2)  $1 : \sqrt{2} \cdot \sqrt{3}$  (3)  $\sqrt{2} : \sqrt{3} : \sqrt{8/\pi}$  (4)  $\sqrt{2} : \sqrt{8/\pi} : \sqrt{3}$

Most probable velocity ( $\alpha$ ) =  $\sqrt{\frac{2RT}{M}}$

Mean velocity ( $\bar{v}$ ) =  $\sqrt{\frac{8RT}{\pi M}}$

Root mean square velocity ( $u$ ) =  $\sqrt{\frac{3RT}{M}}$

$$\therefore \alpha : \bar{v} : u = \sqrt{\frac{2RT}{M}} : \sqrt{\frac{8RT}{\pi M}} : \sqrt{\frac{3RT}{M}}$$

$$= \sqrt{2} : \sqrt{\frac{8}{\pi}} : \sqrt{3}$$

21. The temperature of the gas is raised from  $27^{\circ}C$  to  $927^{\circ}C$ , the root mean square velocity is [1994]  
 (1)  $\sqrt{927/27}$  time the earlier value (2) same as before  
 (3) halved (4) doubled

$$u \propto \sqrt{T} \text{ or } u_1 / u_2 = \sqrt{T_1 / T_2}$$

$$= \sqrt{\frac{27+273}{927+273}} = \sqrt{\frac{300}{1200}} = \frac{1}{2}$$

$$\therefore u_2 = 2u_1$$

22. In a pair of immiscible liquids, a common solute dissolves in both and the equilibrium is reached. Then the concentration of the solute in upper layer is **[1994]**

- (1) In fixed ratio with that in the lower layer  
 (2) Same as the lower layer  
 (3) Lower than the lower layer  
 (4) Higher than the lower layer

By definition of Nernst distribution law. When a solute is shaken with two immiscible liquids, having solubility in both, the solute distributes itself between the two liquids in such a way that the ratio of its concentrations in two liquids is constant at a given temperature, provided the molecular state of the solute remains the same in both the liquids.

23. A liquid can exist only **[1994]**

- (1) between triple point and critical temperature  
 (2) at any temperature above the melting point  
 (3) between melting point and critical temperature  
 (4) between boiling and melting temperature

A substance exists as a liquid above its m. pt. and below its b. pt.

24. 500 mL of nitrogen at 27°C is cooled to -5°C at the same pressure. The new volume becomes **[1995]**

- (1) 326.32 mL      (2) 446.66 mL      (3) 546.32 mL      (4) 771.56 mL

Given initial volume ( $V_1$ ) = 500mL; Initial temperature ( $T_1$ ) = 27°C = 300K and final temperature ( $T_2$ ) = -5°C = 268K.

From Charle's law:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \text{ or } \frac{500}{300} = \frac{V_2}{268}$$

Where  $V_2$  = New volume of gas

$$V_2 = \frac{500}{300} \times 268 = 446.66ml$$

25. 600 c.c. of a gas at a pressure of 750 mm is compressed to 500 c.c. Taking the temperature to remain constant, the increase in pressure, is **[1995]**

- (1) 150 mm      (2) 250 mm      (3) 350 mm      (4) 450 mm

Given initial volume ( $V_1$ ) = 600c.c.; Initial pressure ( $P_1$ ) = 750 mm and final volume ( $V_2$ ) = 500c.c.

According to Boyle's law

$$P_1V_1 = P_2V_2$$

$$\Rightarrow 750 \times 600 = P_2 \times 500$$

$$\text{Or } P_2 = \frac{750 \times 600}{500} = 900 \text{ mm}$$

$$\begin{aligned} \text{Therefore increase in pressure} &= (900 - 750) \\ &= 150 \text{ mm} \end{aligned}$$

26. Cyclopropane and oxygen at partial pressure 170 torr and 570 torr respectively are mixed in a gas cylinder. What is the ratio of the number of moles of cyclopropane to the number of moles of oxygen ( $n_{C_3H_6} / n_{O_2}$ )? [1996]

$$(1) \frac{170 \times 42}{570 \times 32} = 0.39$$

$$(2) \frac{170}{42} \bigg/ \left( \frac{170}{42} + \frac{570}{32} \right) \approx 0.19$$

$$(3) \frac{170}{740} = 0.23$$

$$(4) \frac{170}{570} = 0.30$$

By Ideal gas equation

$$P_1V = n_1RT$$

$$n_1 \propto P_1 \text{ and } n_2 \propto P_2$$

$$\frac{n_1}{n_2} = \frac{P_1}{P_2} \Rightarrow \frac{n_1}{n_2} = \frac{170}{570} = 0.30$$

27. At which one of the following temperature-pressure conditions the deviation of a gas from ideal behavior is expected to be minimum? [1996]

- (1) 350K and 3 atm    (2) 552 K and 1 atm    (3) 250K and 4 atm    (4) 450K and 2 atm

At low pressure and high temperature, real gas nearly behave like ideal gas. Hence, deviation is minimum from ideal behaviour.

28. From a heated mixture of nitrogen, oxygen and carbon, two compounds (out of the many obtained) are isolated. The rates of diffusion of the two isolated compounds are almost identical. The two compounds are [1999]

- (1)  $N_2O$  and  $CO_2$     (2) CO and NO    (3)  $CO_2$  and  $NO_2$     (4)  $N_2O$  and CO

Rate of diffusion depend upon molecular weight

$$\frac{r_1}{r_2} \sqrt{\frac{M_2}{M_1}} \Rightarrow r_1 = r_2 \text{ if } M_1 = M_2$$

Hence compound are  $N_2O$  and  $CO_2$  as both have same molar mass i.e. 22.

29. If 500 mL of gas A at 400 torr and 666.6 mL of B at 600 torr are placed in a 3 litre flask, the pressure of the system will be [1999]

- (1) 200 torr    (2) 100 torr    (3) 550 torr    (4) 366 torr

Applying Boyle's law  $P_1V_1 = P_2V_2$  for both gases

$$\frac{500}{1000} \times 400 = P \times 3 \Rightarrow P = \frac{200}{3}$$

$$600 \times \frac{666.6}{1000} = P' \times 3 \Rightarrow P' = \frac{400}{3}$$

$$\Rightarrow P_T = P + P' = \frac{200}{3} + \frac{400}{3} = \frac{600}{3} = 200 \text{ torr}$$

30. A gaseous mixture contains  $H_2$  and  $O_2$  molar ratio 8 : 1. The ratio of  $H_2$  :  $O_2$  by weight in this mixture would be [1999]

- (1) 4 : 1                      (2) 1 : 8                      (3) 8 : 1                      (4) 1 : 2

$$\frac{\text{Moles of } H_2}{\text{Moles of } O_2} = \frac{8}{1} \text{ (Given)}$$

$$\left( \frac{M.W. \text{ of } O_2}{M.W. \text{ of } H_2} \right) \cdot \left( \frac{\text{weight of } H_2}{\text{weight of } O_2} \right) = \frac{8}{1}$$

$$\frac{\text{weight of } H_2}{\text{weight of } O_2} = \frac{8 \times 2}{32 \times 1} = \frac{1}{2}$$

31. Which of the following expression correctly represents the relationship between the average molar kinetic energy, KE, of CO and  $N_2$  molecules at the same temperature? [2000]

- (1)  $KE_{CO} < KE_{N_2}$   
 (2)  $KE_{CO} > KE_{N_2}$   
 (3)  $KE_{CO} = KE_{N_2}$   
 (4) Cannot be predicted unless volume of the gases given

$$\text{Average molar kinetic energy } \frac{3}{2} kT$$

As temperature is same hence average kinetic energy of CO and  $N_2$  will be same.

32. van der Waals real gas, act as an ideal gas, at which conditions? [2002]

- (1) High temperature, low pressure                      (2) Low temperature, high pressure  
 (3) High temperature, high pressure                      (4) Low temperature, low pressure

At higher temperature and low pressure real gas acts as an ideal gas.

33. The surface tension of which of the following liquid is maximum? [2005]

- (1)  $C_2H_5OH$                       (2)  $CH_3OH$                       (3)  $H_2O$                       (4)  $C_6H_6$

Due to intermolecular H-bonding the surface tension of  $H_2O$  is more than other given liquids. One  $H_2O$  molecule is joined with 4 another  $H_2O$  molecule through H-bond.

Hydrogen bonding is in order  $H_2O > C_2H_5OH > CH_3OH$ .

34. If a gas expands at constant temperature, it indicates that : [2008]

- (1) kinetic energy of molecules decreases
- (2) pressure of the gas increases
- (3) kinetic energy of molecules remains the same
- (4) number of the molecules of gas increases

At any constant temperature the K.E. of gaseous molecules remains same.

35. If a gas expands at constant temperature, it indicates that [CBSE AIPMT 2008]

- (1) kinetic energy of molecules decreases
- (2) pressure of the gas increases
- (3) kinetic energy of molecules remains the same
- (4) number of the molecules of gas increases

If a gas expands at a constant temperature, it indicates that the kinetic energy of molecules remains the same.

The average kinetic energy of a gas molecule is given by the expression:

$$\text{K.E} = \frac{3}{2}kT$$

Thus,  $\text{K.E} \propto T$

At constant temperature, kinetic energy is constant.

36. The density of gas is  $1.964 \text{ g dm}^{-3}$  at 273 K and 76 cm Hg. The gas is [AFMC 2008]

- (1)  $\text{CH}_4$
- (2)  $\text{C}_2\text{H}_6$
- (3)  $\text{CO}_2$
- (4) Xe

$$pV = nRT \text{ or } pV = \frac{w}{M}RT \text{ or } M = \frac{w}{V} \frac{RT}{p} \text{ or } M = d \frac{RT}{p}$$

$$\text{Here, } d = 1.964 \text{g/dm}^3 = 1.964 \times 10^{-3} \text{g/cc}$$

$$P = 76 \text{ cm} = 760 \text{ mm Hg} = 1 \text{ atm}$$

$$R = 0.0821 \text{ L atm K}^{-1} \text{ mol}^{-1}$$

$$= 82.1 \text{ cc atmK}^{-1}\text{mol}^{-1}$$

$$T = 273\text{K}$$

$$\therefore M = \frac{1.964 \times 10^{-3} \times 82.1 \times 273}{1} = 44$$

The molecular weight of  $\text{CO}_2$  is 44.

So, the gas is  $\text{CO}_2$ .

37. A cylinder of 5 L capacity, filled with air at NTP is connected with another evacuated cylinder of 30 L capacity. The resultant air pressure in both the cylinders will be [AMU 2008]

- (1) 10.8 cm of Hg
- (2) 14.9 cm of Hg
- (3) 21.8 cm of Hg
- (4) 38.8 cm of Hg

Boyle's law states that  $P_1V_1 = P_2V_2$

Here, at NTP, pressure = 76 cm of Hg =  $P_1$

$$V_1 = 5L$$

$$V_2 = 30L + 5L$$

Now

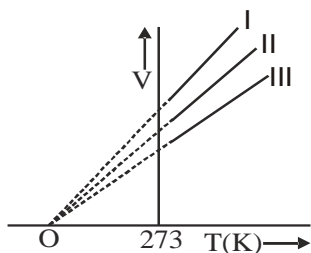
$$P_1 \times (5) = P_2 \times (35)$$

$$P_1 = 7P_2$$

$$P_2 = 76/7 = 10.8 \text{ cm of Hg}$$

So, pressure in both cylinder = 10.8cm

38. The volume temperature graphs of a given mass of an ideal gas at constant pressure are shown below. What is the correct order of pressure? **[DUMET 2008]**



- (1)  $p_3 > p_2 > p_1$       (2)  $p_1 > p_2 > p_3$       (3)  $p_2 > p_3 > p_1$       (4)  $p_2 > p_1 > p_3$

39. The average velocity (in cm/s) of hydrogen molecule at 27°C will be **[DUMET 2008]**

- (1)  $19.3 \times 10^4$       (2)  $17.8 \times 10^4$       (3)  $24.93 \times 10^9$       (4)  $17.8 \times 10^8$

Most probable velocity is  $V_{mps} = \sqrt{\frac{8RT}{\pi M}}$

$$R = 8.314$$

$$T = 27 + 273 = 300K$$

$M = 2g/mol$  molecular mass of hydrogen

$$\text{Most probable velocity will be } \sqrt{\frac{8 \times 8.314 \times 300}{3.14 \times 2 \times 10^{-3}}} = 1782.50m/s = 17.8 \times 10^4cm/s$$

40. The most probable velocity of a gas molecule at 298 K is 300 m/s. Its rms velocity, (in m/s) is  
**[Punjab PMET 2008]**  
 (1) 420 (2) 245 (3) 402 (4) 367

$$C_p = 0.8166 \times C$$

$$C = \frac{C_p}{0.8166} = \frac{300}{0.8166} = 367\text{K}$$

41. By the ideal gas law, the pressure of 0.60 mole  $\text{NH}_3$  gas in a 3.00 L vessel at  $25^\circ\text{C}$  is  
**[AFMC 2009]**  
 (1) 48.9 atm (2) 4.89 atm (3) 0.489 atm (4) 489 atm

$$\text{Pressure is given as } P = \frac{nRT}{V} = \frac{0.6 \times 0.082 \times (25 + 273)}{3} = 4.89\text{atm}$$

42. Critical temperature of  $\text{H}_2\text{O}$ ,  $\text{NH}_3$ ,  $\text{CO}_2$  and  $\text{O}_2$  are 647 K, 405.6 K, 304.10 K and 154.2K respectively. If the cooling starts from 500 K to their critical temperature, the gas that liquefies first is  
**[AMU 2009]**  
 (1)  $\text{H}_2\text{O}$  (2)  $\text{NH}_3$  (3)  $\text{CO}_2$  (4)  $\text{O}_2$

43. A gas can be liquified  
**[CPMT 2009]**  
 (1) at its critical temperature (2) above its critical temperature  
 (3) below its critical temperature (4) at  $0^\circ\text{C}$

A gas can be liquefied below its critical temperature.

For example, the critical temperature  $T_c$  for  $\text{CO}_2$ ,  $\text{O}_2$  and  $\text{H}_2$  are  $31.1^\circ\text{C}$ ,  $-118.8^\circ\text{C}$  and  $-240^\circ\text{C}$  respectively.

Thus,  $\text{CO}_2$  gas cannot be liquefied above  $31.1^\circ\text{C}$  however high the pressure may be applied.

44. The pressure exerted by 6.0 g of methane gas in a  $0.03 \text{ m}^3$  vessel at  $129^\circ\text{C}$  is (Atomic masses: C = 12.01, H = 1.01 and  $R = 8.314 \text{ JK}^{-1} \text{ mol}^{-1}$ )  
**[2010]**  
 (1) 31684Pa (2) 215216Pa (3) 13409Pa (4) 41648 Pa

$$P = \frac{nRT}{V} = \frac{w}{m} \frac{RT}{V}$$

$$= \frac{6}{16.05} \times \frac{8.314 \times 402}{0.03} \approx 41648\text{Pa}$$

45. What will be the partial pressure of He and  $\text{O}_2$  respectively, if 200 mL of He at 0.66 atm and 400 mL of  $\text{O}_2$  at 0.52 atm pressure are mixed in 400 mL vessel at  $20^\circ\text{C}$ ?  
**[AFMC 2010]**  
 (1) 0.33 and 0.56 (2) 0.33 and 0.52 (3) 0.38 and 0.52 (4) 0.25 and 0.45

Partial pressure is the pressure of individual gas when it is placed in a required vessel or in final vessel.

Given :- Initially volume of He is 200 ml at 0.66 atm after placing it in a vessel of volume 400 ml  
 $P_1V_1 = P_2V_2$

$$.66 \times 200 = P_2 \times 400$$

$P_2 = 0.33 \text{ atm}$  ----- (Partial pressure of He)

Initially  $O_2$  is present at  $0.25 \text{ atm}$  in  $400 \text{ ml}$  volume and then it is placed in  $400 \text{ ml}$  vessel hence the partial pressure of  $O_2$  is  $0.52 \text{ atm}$

Partial pressure of He =  $0.33 \text{ atm}$

Partial pressure of  $O_2 = 0.52 \text{ atm}$

46. Which of the following expression is true regarding gas laws? (w = weight; M = molecular mass)

[CPMT 2010]

$$(1) \frac{T_1}{T_2} = \frac{M_1 w_2}{M_2 w_1}$$

$$(2) \frac{T_1}{T_2} = \frac{M_2 w_1}{M_1 w_2}$$

$$(3) \frac{T_1}{T_2} = \frac{M_1 w_1}{M_2 w_2}$$

$$(4) \frac{T_2}{T_1} = \frac{M_1 w_1}{M_2 w_2}$$

According to ideal gas equation

$$PV = nRT$$

Where, P = Pressure

V = Volume

n = No. of mole

R = Gas constant

T = Temperature

$$n = \frac{w_1}{M_1}$$

Where  $w_1$  = weight of the gas

$m_1$  = Molar mass

Taking P, V and R constant for two gas 1 and 2

$$T_1 = \frac{P \times V \times M_1}{w_1 \times R}$$

$$T_2 = \frac{P \times V \times M_2}{w_2 \times R}$$

$$\frac{T_1}{T_2} = \frac{M_1 w_2}{M_2 w_1}$$

47. A mixture of gases having different molecular weights is separated by which method?

[CPMT 2010]

(1) **Atmolysis**

(2) Metathesis

(3) Ostwald and Walker method

(4) Reverse osmosis

A process for separating gases or vapors of different molecular weights by transmission through a porous substance is called atmolysis.



The exchange of bonds between two reacting chemical species is Metathesis method.

Ostwald and Walker's Dynamic Method (Gas Saturation Method): In this method, the relative lowering of vapor pressure can be calculated in an easy simple way Procedure.

A process by which a solvent passes through a porous membrane in the direction opposite to that for natural osmosis when subjected to a hydrostatic pressure greater than the osmotic pressure.

48. If  $p$ ,  $V$ ,  $M$ ,  $T$  and  $R$  are the symbols of pressure, volume, molecular weight, temperature and gas constant respectively, what is the equation of density of ideal gas? **[MP PMT 2010]**

- (1)  $\frac{RT}{pM}$       (2)  $\frac{p}{RT}$       (3)  $\frac{M}{V}$       (4)  $\frac{pM}{RT}$

49. The temperature at which 28 g of  $N_2$  will occupy a volume of 10.0 L at 2.46 atm is **[RPMT 2010]**

(1) 299.6 K      (2)  $0^\circ C$       (3) 273 K      (4)  $10^\circ C$

We know that,  $pV = nRT$ .

$$\text{Then } T = \frac{pV}{nR}$$

Given:-  $P = 2.46 \text{ atm}$ ,  $V = 10 \text{ L}$ ,  $n = 1 \text{ mol}$ .

$$R = 0.082 \text{ litres/atm/Kmol}$$

$$\therefore T = \frac{2.46 \text{ atm} \times 10 \text{ litre}}{1 \text{ mol} \times 0.082 \text{ litre/atm/Kmol}}$$

$$\text{So } T = 300 \text{ K}$$

Hence the temperature at which 28 g  $N_2$  occupies a volume of 10 litres at 2.46 atm is 300 K.

50. For one mole of an ideal gas, increasing the temperature from  $10^\circ C$  to  $20^\circ C$  **[KCET 2010]**

- (1) increases the average kinetic energy by two times  
 (2) increases the rms velocity of  $\sqrt{2}$  times  
 (3) increases the rms velocity by two times  
 (4) increases both the average kinetic energy and rms velocity, but not significantly

51. In which one of the following, does the given amount of chlorine exert the least pressure in a vessel of capacity  $1 \text{ dm}^3$  at  $273 \text{ K}$ ? **[KCET 2010]**

- (1) 0.0355 g      (2) 0.071 g  
 (3)  $6.023 \times 10^{21}$  molecules      (4) 0.02 mol

The ideal gas equation is:

$$PV = nRT$$

When V and T are the same,

$$P \propto n$$

Thus, when the number of moles, i.e., n is least, it will exert the least pressure.

$$(a) n = \frac{\text{wt}}{\text{mol.wt}} = \frac{0.0355}{35.5} = 1 \times 10^{-3} \text{ mol}$$

$$(b) n = \frac{0.071}{35.5} = 2 \times 10^{-3} \text{ mol}$$

$$(c) n = \frac{\text{number of molecules}}{N_A}$$

$$= \frac{6.023 \times 10^{21}}{6.023 \times 10^{23}} = 0.01 \text{ mol}$$

$$(d) n = 0.02 \text{ mol}$$

Thus, 0.0355 g chlorine will exert the least pressure.

52. Two gram of hydrogen diffuse from a container in 10 min. How many gram of oxygen would diffuse through the same container in the same time under similar conditions? **[Manipal 2010]**

- (1) 0.5 g                      (2) 4 g                      (3) 6 g                      **(4) 8 g**

$$\frac{r_{H_2}}{r_{O_2}} = \frac{V_{H_2}}{V_{O_2}}$$

$$\frac{n_{H_2}}{n_{O_2}} = \sqrt{\frac{M_{O_2}}{M_{H_2}}}$$

$$\Rightarrow \frac{\frac{2}{w}}{\frac{2}{2}} = \sqrt{\frac{32}{2}} = \sqrt{16} = 4$$

$$\Rightarrow \frac{32}{w} = 4$$

$$\Rightarrow w = 8 \text{ gms}$$

53. Vander Waals' equation of state is obeyed by real gases. For n moles of a real gas, the expression will be [Manipal 2010]

(1)  $\left(\frac{p}{n} + \frac{na}{V^2}\right)\left(\frac{V}{n-b}\right) = RT$

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

(3)  $\left(p + \frac{na}{V^2}\right)(nV - b) = nRT$

(4)  $\left(p + \frac{na}{V^2}\right)(nV - b) = nRT$

54. In Vander Waals' equation of state of the gas, the constant 'b' is a measure of [JCECE 2010]

- (1) intermolecular repulsions (2) intermolecular attraction  
 (3) volume occupied by the molecules (4) intermolecular collisions per unit volume

Vander waals equation of state is  $\left(P + \frac{an^2}{V^2}\right)(V - nb) = nRT$

Correct Answer: Option C

Explanation:

In the ideal gas equation, we neglect the intermolecular attractions and volume occupied by molecules. However, the real gas equation accounts for both of these. The parameter a is a measure of intermolecular attraction and the parameter b is a measure of the volume occupied by the molecules. In particular, if the volume of a single gas molecule is v, then  $b=4vN_A$ . Hence b is four times the molar volume of the gas.

55. The compressibility factor for an ideal gas is [OJEE 2010]

- (1) 1 (2) <1 (3) >1 (4) ∞

For an ideal gas,

$PV=nRT$

$Z=1$

The compressibility factor (Z), also known as the compression factor, is the ratio of the molar volume of a gas to the molar volume of an ideal gas at the same temperature and pressure for an ideal gas the compressibility factor is 1.

56. The density of gas A is twice that of B. Molecular mass of A is half of molecular mass of B. The ratio of partial pressure of A and B is [OJEE 2010]

- (1) 1/4 (2) 1/2 (3) 4/1 (4) 2/1

57. At what temperature, the rms velocity of gas shown at 50°C will be doubled? [OJEE 2010]

- (1) 626 K (2) 1019 K (3) 200°C (4) 1019°C

58. Given: rms velocity of hydrogen at 300K is  $1.9 \times 10^3$  m/s. The rms velocity of oxygen at 1200K will be [OJEE 2010]

- (1)  $7.6 \times 10^3$  m/s (2)  $3.8 \times 10^3$  m/s (3)  $0.95 \times 10^3$  m/s (4)  $0.475 \times 10^3$  m/s

59. The Vander Waals' constant 'a' for different gases have been given as

Gas	a(atm. L <sup>2</sup> . mol <sup>-2</sup> )
O <sub>2</sub>	1.36
N <sub>2</sub>	1.39
CH <sub>4</sub>	2.25
NH <sub>3</sub>	4.17

The gas that can be most easily liquefied is

[OJEE 2010]

- (1) O<sub>2</sub>                      (2) N<sub>2</sub>                      (3) CH<sub>4</sub>                      (4) NH<sub>3</sub>

The gas that can be most easily liquefied is ammonia NH<sub>3</sub> as it has highest value of the van der Waals' constant 'a'.

Note: The higher is the value of van der Waals' constant 'a', stronger are the intermolecular forces of attraction and greater is the ease of liquefaction

60. If temperature changes from 27°C to 127°C, the relative percentage change in rms velocity is

[OJEE 2010]

- (1) 1.56                      (2) 2.56                      (3) 15.6                      (4) 82.4

61. At identical temperature and pressure, the rate of diffusion of hydrogen gas is  $3\sqrt{3}$  times that of a hydrocarbon having molecular formula C<sub>n</sub>H<sub>2n-2</sub>. What is the value of n? [WB JEE 2010]

- (1) 1                      (2) 4                      (3) 3                      (4) 8

Given, rate of diffusion of hydrogen =  $r_{H_2}$

rate of diffusion of hydrocarbon =  $r_{HC}$

Also  $r_{H_2} = 3\sqrt{3}r_{HC} \rightarrow (1)$

Also P and T is constant.

We know by Graham's law of diffusion

$$r \propto \frac{1}{\sqrt{M}} \quad M \text{ is molar mass}$$

$$\frac{r_{H_2}}{r_{HC}} = \sqrt{\frac{M_{HC}}{M_{H_2}}}$$

$$\frac{r_{H_2}}{r_{HC}} = \sqrt{\frac{M_{HC}}{2}}$$

From 1st equation,

$$3\sqrt{3} = \sqrt{\frac{M_{HC}}{2}} \text{ [squaring on both sides]}$$

$$27 = M_{HC}/2$$

$$M_{HC} = 54$$

Molar mass of hydrocarbon  $C_nH_{2n-2} = 54\text{g/mole}$

We know

$$1C = 12\text{g/mole, } 1H = 1\text{g/mole}$$

$$12 \times n + 1 \times (2n - 2) = 54$$

$$n = 4$$

62.

63. The real gas most closely approaches the behaviour of an ideal gas at [BVP 2010]

- (1) 15 atm and 200K (2) 1 atm and 273 K (3) 0.5 atm and 500K (4) 15 atm and 500K

An ideal gas is a hypothetical gas whose pressure, volume and temperature behaviour is completely described by the ideal gas equation. Actually no gas is ideal or perfect in nature. All gases are real gases. Real gases do not obey the ideal gas laws exactly under all conditions of temperature and pressure.

Real gases deviate from ideal behaviour because of mainly two assumptions of "Kinetic theory of gases".

(i)The volume of a gas particle is negligible compared to the volume of the container (while the real gas particle has some significant volume).

(ii)There is no interaction between gaseous particles (while attraction forces exist between real gas particles).

So at lowest pressure and highest temperature, a real gas most closely approaches the behaviour of an ideal gas.

64. Average KE of  $CO_2$  at  $27^\circ C$  is  $E$ . The average kinetic energy of  $N_2$  at the same temperature will be [BVP 2010]

- (1)  $E$  (2)  $22E$  (3)  $E/22$  (4)  $E/\sqrt{2}$

Kinetic energy is the energy possessed by an object by virtue of its motion.

Average KE of  $\text{CO}_2$  at  $27^\circ\text{C}$  is  $E$ . The average kinetic energy of  $\text{N}_2$  at the same temperature will be  $E$

Kinetic energy of 1 mole of gas is  $\frac{3}{2}RT$ .

Thus, the average kinetic energy of 1 mole of any gas is directly proportional to its absolute temperature and independent of its molar mass.

65. Equation of state is **[Kerala CEE 2011]**

- (1)  $M \times V = w p R T$     (2)  $\frac{M T}{R} = \frac{w V}{p}$     (3)  $\frac{w R}{p} = \frac{V M}{T}$     (4) None of these

66. A gaseous mixture was prepared by taking equal mole of  $\text{CO}$  and  $\text{N}_2$ . If the total pressure of the mixture was found 1 atmosphere, the partial pressure of the nitrogen ( $\text{N}_2$ ) in the mixture is

**[2011]**

- (1) 0.5 atm    (2) 0.8 atm    (3) 0.9 atm    (4) 1 atm

Given  $n_{\text{CO}} = n_{\text{N}_2}$

$$P_{\text{CO}} + P_{\text{N}_2} = 1 \text{ atm}$$

Partial pressure of a gas = mole fraction of gas  $\times$  total pressure

$$\therefore P_{\text{N}_2} = \frac{n_{\text{N}_2}}{n_{\text{CO}} + n_{\text{N}_2}} \times 1 = \frac{n_{\text{N}_2}}{2n_{\text{N}_2}} \times 1$$

$$= \frac{1}{2} = 0.5 \text{ atm}$$

67. Two gases  $A$  and  $B$  having the same volume diffuse through a porous partition in 20 and 10 seconds respectively. The molecular mass of  $A$  is 49 u. Molecular mass of  $B$  will be : **[2011]**

- (1) 50.00 u    (2) 12.25 u    (3) 6.50 u    (4) 25.00 u

$$\frac{r_A}{r_B} = \sqrt{\frac{M_B}{M_A}}$$

$$\frac{\frac{V}{20}}{\frac{V}{10}} \sqrt{\frac{M_B}{49}} \Rightarrow \frac{1}{2} = \sqrt{\frac{M_B}{49}}$$

$$M_B = \frac{1}{4} \times 49 = 12.25 \text{ u}$$

68. A bubble of air is underwater at temperature  $15^\circ\text{C}$  and the pressure 1.5 bar. If the bubble rises to the surface where the temperature is  $25^\circ\text{C}$  and the pressure is 1.0 bar, what will happen to the volume of the bubble? **[2011M]**

- (1) Volume will become greater by a factor of 1.6.  
 (2) Volume will become greater by a factor of 1.1.

(3) Volume will become smaller by a factor of 0.70.

(4) Volume will become greater by a factor of 2.5.

Given

$$P_1 = 1.5 \text{ bar}, T_1 = 273 + 15 = 288 \text{ K}, V_1 = V$$

$$P_2 = 1.0 \text{ bar}, T_2 = 273 + 25 = 298 \text{ K}, V_2 = ?$$

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

$$\frac{1.5 \times V}{288} = \frac{1 \times V_2}{298}$$

$V_2 = 1.55V$  i.e. volume of bubble will be almost 1.6 time to initial volume of bubble.

69. A gaseous mixture was prepared by taking equal moles of CO and N<sub>2</sub>. If the total pressure of the mixture was found 1 atm, the partial pressure of the nitrogen (N<sub>2</sub>) in the mixture is

[CBSE AIPMT 2011]

- (1) 1 atm                      (2) 0.5 atm                      (3) 0.8 atm                      (4) 0.9 atm

According to Dalton's Law of partial pressure.

The partial pressure of gas  $\propto$  Mole fraction of gas (X gas)

$$X_{\text{CO}} = \frac{n_{\text{CO}}}{n_{\text{CO}} + n_{\text{N}_2}} = \frac{n_{\text{CO}}}{2n_{\text{CO}}} \quad [ \because n_{\text{CO}} = n_{\text{N}_2} ]$$

$$\Rightarrow X_{\text{CO}} = \frac{1}{2} = 0.5$$

$$\text{Partial pressure of N}_2 = X_{\text{N}_2} \cdot P_T = [1 - X_{\text{CO}}] \cdot P = [1 - 0.5] \times [1] = 0.5 \text{ atm}$$

70. Equal weights of CH<sub>4</sub> and H<sub>2</sub> are mixed in a container at 25°C. Fraction of total pressure exerted by methane is

[DUMET 2011]

- (1)  $\frac{1}{2}$                       (2)  $\frac{1}{3}$                       (3)  $\frac{1}{9}$                       (4)  $\frac{8}{9}$

71. A bubble of gas released at the bottom of a lake increases to eight times its original volume when it reaches the surface. Assuming that atmospheric pressure is equivalent to the pressure exerted by a column of water, 10 m height, the depth of the lake is

[Kerala CEE 2011]

- (1) 80 m                      (2) 90 m                      (3) 40 m                      (4) 70 m

Let us assume that the bubble is  $h$  meters below the surface of the water, and  $P_0$  is the atmospheric pressure.

Pressure only due to water at a depth of  $h$  meters would be equal to  $\rho gh$ . Also, atmospheric pressure would be acting on the bubble. So, the total pressure on the gas bubble at that depth  $= P_0 + \rho gh = \rho g \times 10 + \rho gh = \rho g(h + 10)$ .

At the surface of water, only atmospheric pressure would be acting on the bubble i.e.  $P = P_0 = \rho g \times 10$

As  $PV = \text{constant}$ ,  $\rho g(h + 10) \times V = \rho g10 \times 8V \rightarrow h + 10 = 80 \rightarrow h = 70$  meters.

72. Three different gases X, Y and Z of molecular masses 2, 16 and 64 were enclosed in a vessel at constant temperature till equilibrium is reached. Which of the following statement is correct? [Kerala CEE 2011]
- (1) Gas Z will be at the top of the vessel
  - (2) Gas Y will be at the top of the vessel
  - (3) Gas X will be at the bottom and Z will be at the top
  - (4) Gases will form homogeneous mixture

73. By what factor does the average velocity of a gaseous molecule increase when the temperature (in Kelvin) is doubled? [2011]
- (1) 2.0
  - (2) 2.8
  - (3) 4.0
  - (4) 1.4

$$\text{Average velocity} = \sqrt{\frac{8RT}{\pi M}}$$

$$\text{i.e., } v \propto \sqrt{T}$$

$$\therefore \frac{V_2}{V_1} = \sqrt{\frac{2T}{T}} = 1.41$$

74. 50 mL of each gas A and of gas B takes 150 and 200 seconds respectively for effusing through a pin hole under the similar condition. If molecular mass of gas B is 36, the molecular mass of gas A will be : [2012]
- (1) 96
  - (2) 128
  - (3) 32
  - (4) 64

$$\frac{V_A}{t_A} / \frac{V_B}{t_B} \sqrt{\frac{M_B}{M_A}}$$

$$\frac{t_B}{t_A} \Rightarrow \frac{200}{150} = \sqrt{\frac{36}{M_A}} \Rightarrow \frac{4}{3} = \sqrt{\frac{36}{M_A}}$$

$$\Rightarrow \frac{16}{9} = \frac{36}{M_A} \Rightarrow M_A = \frac{81}{4} = 20.25$$

75. For real gases van der Waals' equation is written as [CBSE Screening 2012]



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

where 'a' and 'b' are vander Waals' constants.

The sets of gases are:

(I) O<sub>2</sub>, CO<sub>2</sub>, H<sub>2</sub> and He (II) CH<sub>4</sub>, O<sub>2</sub> and H<sub>2</sub>

The gases given in set - I in increasing order of 'b' and gases given in set - II in decreasing order of 'a' are arranged below. Select the correct order from the following:

- (1) (I) He < H<sub>2</sub> < CO<sub>2</sub> < O<sub>2</sub> (II) CH<sub>4</sub> > H<sub>2</sub> > O<sub>2</sub> (2) (I) O<sub>2</sub> < He < H<sub>2</sub> < CO<sub>2</sub> (II) H<sub>2</sub> > O<sub>2</sub> > CH<sub>4</sub>  
 (3) (I) H<sub>2</sub> < He < O<sub>2</sub> < CO<sub>2</sub> (II) CH<sub>4</sub> > O<sub>2</sub> > H<sub>2</sub> (4) (I) He < H<sub>2</sub> < O<sub>2</sub> < CO<sub>2</sub> (II) CH<sub>4</sub> > O<sub>2</sub> > H<sub>2</sub>

76. A certain gas takes three times as long to effuse out as helium. Its molecular mass will be: [2012 M]

- (1) 27 u (2) 36 u (3) 64 u (4) 9 u

According to Graham's law of diffusion

$$r \propto \frac{1}{d} \sqrt{\frac{1}{M}} \Rightarrow \frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}}$$

$$\frac{3r_1}{r_1} = \sqrt{\frac{M_2}{4}} \Rightarrow 9 = \frac{M_2}{4}$$

$$M_2 = 36u$$

77. Dipole-induced dipole interactions are present in which of the following pairs : [NEET 2013]

- (1) Cl<sub>2</sub> and CCl<sub>4</sub> (2) HCl and He atoms (3) SiF<sub>4</sub> and He atoms (4) H<sub>2</sub>O and alcohol  
 HCl is polar ( $\mu \neq 0$ ) and He is non-polar ( $\mu = 0$ ) thus gives dipole-induced dipole interaction.

78. Maximum deviation from ideal gas is expected from : [NEET 2013]

- (1) N<sub>2</sub>(g) (2) CH<sub>4</sub>(g) (3) NH<sub>3</sub>(g) (4) H<sub>2</sub>(g)

Higher the critical temperature more easily will be the gas liquify. Now since most easily liquefiable gas show larger deviation, NH<sub>3</sub> will show maximum deviation from ideal behaviour.

79. What is the density of N<sub>2</sub> gas at 227<sup>o</sup>C and 5.00 atm pressure? ( $R = 0.0821 \text{ atm K}^{-1} \text{ mol}^{-1}$ )

- (1) 0.29 g/ml (2) 1.40 g/ml (3) 2.81 g/ml (4) 3.41 g/ml [NEET Kar. 2013]

$$\text{Density} = \frac{PM}{RT} = \frac{5 \times 28}{0.0821 \times 500} = 3.41 \text{ g / mL}$$

80. Equal masses of H<sub>2</sub>, O<sub>2</sub> and methane have been taken in a container of volume V at temperature 27<sup>o</sup>C in identical conditions. The ratio of the volumes of gases H<sub>2</sub> : O<sub>2</sub> : methane would be : [2014]

- (1) 8 : 16 : 1 (2) 16 : 8 : 1 (3) 16 : 1 : 2 (4) 8 : 1 : 2

According to Avogadro's law "At same temperature and pressure. Volume  $\propto$  no. of moles"

$$n_{H_2} = \frac{w}{2}; n_{O_2} = \frac{w}{32}, n_{CH_4} = \frac{w}{16}$$

$$\therefore V_{H_2} : V_{O_2} : V_{CH_4} = n_{H_2} : n_{O_2} : n_{CH_4}$$

$$= \frac{w}{2} : \frac{w}{32} : \frac{w}{16} = 16 : 1 : 2$$

81. A gas such as carbon monoxide would be most likely to obey the ideal gas law at [2015 RS]

- (1) high temperatures and low pressures      (2) low temperature and high pressure  
 (3) high temperature and low pressure      (4) low temperatures and low pressure

At high temperature and low pressure

82. A mixture of gases contains H<sub>2</sub> and O<sub>2</sub> gases in the ratio of 1: 4 (w/w). What is the molar ratio of the two gases in the mixture? [AIPMT - 2015]

- (1) 4:1      (2) 16:1      (3) 2:1      (4) 1:4

83. Equal moles of hydrogen and oxygen gases are placed in a container with a pin-hole through which both can escape. What fraction of the oxygen escapes in the time required for one-half of the hydrogen to escape? [2016]

- (1) 1/8      (2) 1/4      (3) 3/8      (4) 1/2

Given  $n_{H_2} = n_{O_2}$  and  $t_{H_2} = t_{O_2}$

According to Graham's law diffusion for two different gases.

$$\frac{r_{H_2}}{r_{O_2}} = \frac{v_1 / t_1}{v_2 / t_2} \Rightarrow \sqrt{\frac{M_{O_2}}{M_{H_2}}} = \sqrt{\frac{32}{2}}$$

$$\frac{1/2}{1/x} = \sqrt{16} = 4$$

$$\frac{x}{2} = 4$$

$$\therefore x = 8$$

$\therefore$  Fraction of O<sub>2</sub> = 1/8

84. The correct factor 'a' to the ideal gas equation corresponds to [2018]

- (1) Density of the gas molecules  
 (2) Volume of the gas molecules  
 (3) Forces of attraction between the gas molecules  
 (4) Electric field present between the gas molecules

In real gas equation, van der Wall constant.

85. Given van der Waals constants for NH<sub>3</sub>, H<sub>2</sub>, O<sub>2</sub> and CO<sub>2</sub> are respectively 4.17, 0.244, 1.36 and 3.59, which one of the following gases is most easily liquefied? [2018]

- (1) NH<sub>3</sub>      (2) H<sub>2</sub>      (3) CO<sub>2</sub>      (4) O<sub>2</sub>

Van der Waal constant 'a', signifies intermolecular force of attraction.

Higher is the value of 'a' easier will be the liquefaction of gas

86. The volume occupied by 1.8 g of water vapour at 374°C and 1 bar pressure will be [Use  $R = 0.083$  bar LK<sup>-1</sup> mol<sup>-1</sup>] [NEET Odisha 2019]

(1) 5.37 L                      (2) 96.66 L                      (3) 55.87 L                      (4) 3.10 L

According to ideal gas equation

$$PV = nRT$$

$$V = \frac{W}{M} \left( \frac{RT}{P} \right)$$

$$V = \frac{1.8}{18} \times \frac{0.083 \times 647}{1} = 5.37L$$

87. A gas at 350 K and 15 bar has molar volume 20 percent smaller than that for an ideal gas under the same conditions. The correct option about the gas and its compressibility factor (Z) is [2019]

(1)  $Z > 1$  and attractive force are dominant                      (2)  $Z > 1$  and repulsive forces are dominant

(3)  $Z < 1$  and attractive force are dominant                      (4)  $Z < 1$  and repulsive forces are dominant

Compressibility factor,  $Z = \frac{PV}{nRT}$

Given: At 350K and 15 bar

Molar volume < volume of ideal gas

∴  $Z < 1$

Therefore, attractive forces are dominant and the gas can be compressed easily.

88. A mixture of N<sub>2</sub> and Ar gases in a cylinder contains 7 g of N<sub>2</sub> and 8 g of Ar. If the total pressure of the mixture of the gases in the cylinder is 27 bar, the partial pressure of N<sub>2</sub> is: [2020]

[Use atomic masses (in g mol<sup>-1</sup>): N = 14, Ar = 40]

(1) 12 bar                      (2) 15 bar                      (3) 18 bar                      (4) 9 bar

$$n_{N_2} = \frac{7}{28} = \frac{1}{4} = 0.25$$

$$n_{Ar} = \frac{8}{40} = \frac{1}{5} = 0.20$$

Now, applying Dalton's law of partial pressure,

$$p_{N_2} = \text{mole fraction of } N_2 \cdot P_{Total}$$

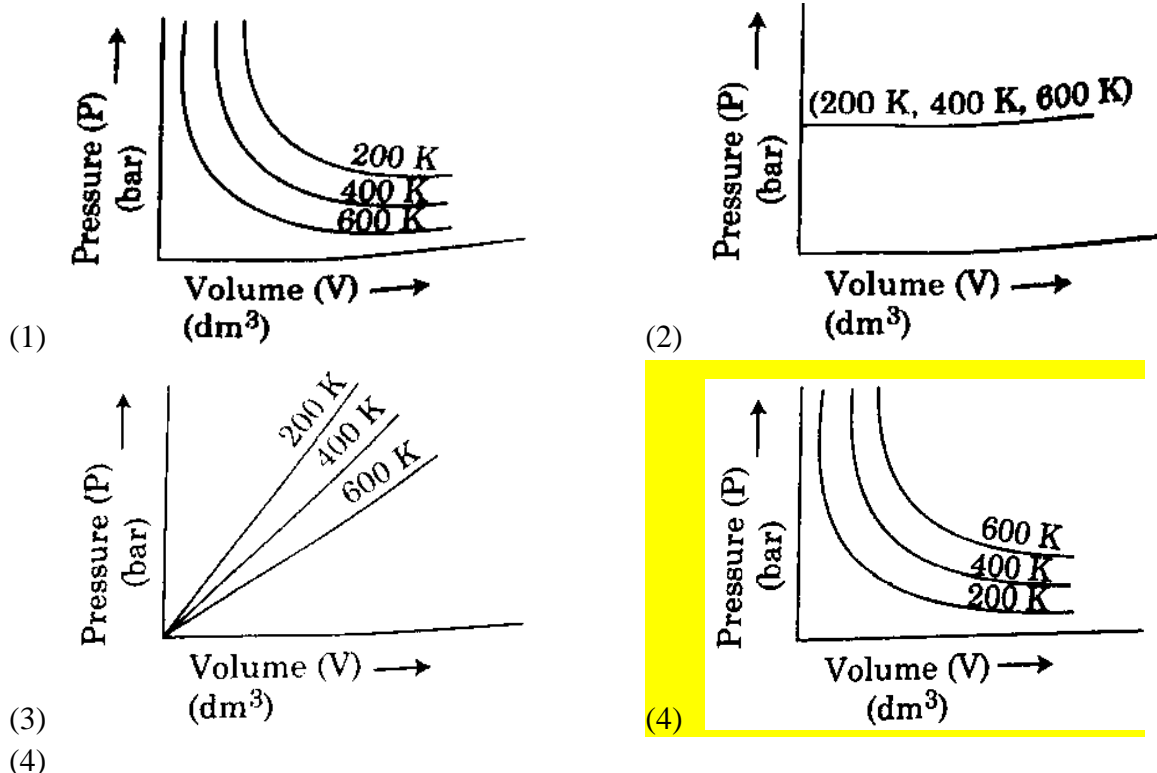
$$\frac{0.25}{0.45} \times 27 = \frac{5}{9} \times 27 = 15bar$$

89. The minimum pressure required to compress 600 dm<sup>3</sup> of a gas at 1 bar to 150 dm<sup>3</sup> at 40°C is [NEET (Oct.)2020]

(1) 4.0 bar                      (2) 0.2 bar                      (3) 1.0 bar                      (4) 2.5 bar

90. Choose the correct option for the total pressure (in atm) in a mixture of 4gO<sub>2</sub> and 2gH<sub>2</sub> confined in a total volume of one litre at 0°C is [Given, R = 0.082 L atm mol<sup>-1</sup>K<sup>-1</sup>, T=273K] [NEET 2021]  
 (1) 2.518 (2) 2.602 (3) 25.18 (4) 26.02

91. Choose the correct option for graphical representation of Boyle's law, which shows a graph of pressure vs volume of a gas at different temperature [NEET 2021]



At a given temperature, P vs V will be hyperbola and as the temperature increases (PV) will increase.

92. For real gases, van der Waals equation is written as  $\left(p + \frac{an^2}{V^2}\right)(v - nb) = nRT$  where 'a' and 'b' are van der Waals constants. Two sets of gases are  
 (I) O<sub>2</sub>, CO<sub>2</sub>, H<sub>2</sub> and He  
 (II) CH<sub>4</sub>, O<sub>2</sub> and H<sub>2</sub>

The gases given in set – I in increasing order of 'b' and gases given in set – II in decreasing order of 'a', are arranged below. Select the correct order from the following [NEET 2021]

- (1) (I) HE < H<sub>2</sub> < CO<sub>2</sub> < O<sub>2</sub> (II) CH<sub>4</sub> > H<sub>2</sub> > O<sub>2</sub>  
 (2) (I) O<sub>2</sub> < He < H<sub>2</sub> < CO<sub>2</sub> (II) H<sub>2</sub> > O<sub>2</sub> > CH<sub>4</sub>  
 (3) (I) H<sub>2</sub> < He < O<sub>2</sub> < CO<sub>2</sub> (II) CH<sub>4</sub> > O<sub>2</sub> > H<sub>2</sub>  
 (4) (I) H<sub>2</sub> < O<sub>2</sub> < H<sub>2</sub> < CO<sub>2</sub> (II) O<sub>2</sub> > CH<sub>4</sub> > H<sub>2</sub>

Set – I: O<sub>2</sub>, CO<sub>2</sub>, H<sub>2</sub>, He for 'b' values

Set – II:  $CH_4, O_2, H_2$  for 'a' values

Value of van der Waals constant 'b' increases with increase in molecular volume. Clearly, the increasing order of molecular volume (size of molecule) is:

$He < H_2 < O_2 < CO_2$  Value of 'a' increases with increase in intermolecular attraction.  $CH_4$  is a polar molecule, thus, it will possess highest value of 'a'.  $O_2$  and  $H_2$ , both are non-polar molecules. The van der Waals force  $\propto$  molecular mass. Hence, the correct order of value of 'b' is :  $CH_4 > O_2 > H_2$ .

From the given options, (3) is most suitable.