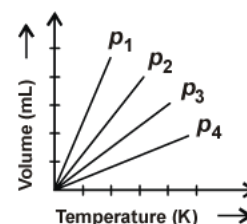


- Q1. A plot of volume ( $V$ ) versus temperature ( $T$ ) for a gas at constant pressure is a straight line passing through the origin. The plots at different values of pressure are shown in figure.

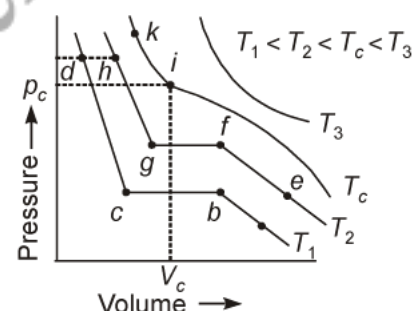


What is the correct order of pressure?

- Q2. Arrange the following species in increasing order of their van der Waals' radii:

Cl, H, O and N.

- Q3. What is the type of graph between  $\log(p)$  and  $\log(1/V)$  at constant temperature?
- Q4. The critical temperature ( $T_c$ ) and critical pressure ( $p_c$ ) of  $\text{CO}_2$  are  $30.98^\circ\text{C}$  and  $73\text{ atm}$ , respectively. Can  $\text{CO}_2(g)$  be liquefied at  $32^\circ\text{C}$  and  $80\text{ atm}$  pressure?
- Q5. What is the slope between  $pV$  and  $p$  at constant temperature?
- Q6. Critical temperature of  $\text{NH}_3$  and  $\text{SO}_2$  are  $405.0$  and  $430.3\text{ K}$ . Which one will have higher value of  $a$  and why?
- Q7. A  $2.0\text{ L}$  container at  $25^\circ\text{C}$  contains  $1.25\text{ mol}$  of oxygen and  $3.3\text{ mol}$  of carbon.
- What is the initial pressure in the flask?
  - If carbon and oxygen react to form  $\text{CO}$ , then what will be the final pressure in the container?
- Q8. Density of a gas is found to be  $10.92\text{ g dm}^{-3}$  at  $27^\circ\text{C}$  and  $4\text{ bar}$  pressure. What will be its density at STP?
- Q9. Isotherms of carbon dioxide at various temperatures are represented in the following figure.



Answer the following questions:

- In which state, will  $\text{CO}_2$  exist between the points  $a$  and  $b$  at temperature  $T_1$ ?
  - At what point, will  $\text{CO}_2$  start liquefying when temperature is  $T_1$ ?
  - At what point, will  $\text{CO}_2$  be completely liquefied when temperature is  $T_2$ ?
  - Will condensation take place when the temperature is  $T_3$ ?
- Q10. (a) A  $2.5\text{ L}$  flask contains  $0.25\text{ mol}$  each of nitrogen and sulphur dioxide gas at  $27^\circ\text{C}$ . Calculate the partial pressure exerted by each gas and also the total pressure.
- (b) Helium cannot be liquefied at room temperature, no matter what pressure we apply. Why?
- Q11. What will be the pressure exerted by a mixture of  $3.2\text{ g}$  of methane and  $4.4\text{ g}$  of carbon dioxide contained in a  $9\text{ dm}^3$  flask at  $27^\circ\text{C}$ ?

- Q12. At 25°C and 0.993 bar, 135 mL of a gas is collected over water, If the gas weighs 0.16 g and the aqueous tension at 25°C, is 0.0317. Calculate the molar mass of the gas.
- Q13. (a) What will be the effect of increasing the temperature of a liquid, on intermolecular forces operating between its particles? What will happen to the viscosity of a liquid if its temperature is increased?
- (b) What volume will be occupied by 9.8 g of CO<sub>2</sub> at 31.1°C and 1 bar pressure?  
[R = 0.083 bar LK<sup>-1</sup> mol<sup>-1</sup>]
- Q14. A flask of 4.0 dm<sup>3</sup> containing 4.0 dm<sup>3</sup> N<sub>2</sub> at 4.0 bar, was connected to a flask of 6.0 dm<sup>3</sup> containing 6.0 dm<sup>3</sup> helium at 6.0 bar and the gases were allowed to mix isothermally. What will be the total pressure of the resulting mixture?
- Q15. (a) The compressibility factor (compression factor) for one mole of van der Waals' at 0°C and 100 atm pressure is found to be 0.5. Assuming that the volume correction for the gaseous molecule is negligible, calculate the van der Waals' constant (a)
- (b) What will be the effect on the value of compressibility factor above Boyle's temperature for a real gas?
- Q16. (a) How does the surface tension of a liquid vary with increase in temperature?
- (b) Why the rain droplets are spherical in shape?
- (c) What will be the molar volume of nitrogen and argon at 273.15 K and 1 atm?
- Q17. Pay load is defined as the difference between the mass of displaced air and the mass of the balloon. Calculate the pay load when a balloon of radius 10 m, mass 100 kg is filled with helium at 1.66 bar 27°C. (Density of air = 1.2 kg m<sup>-3</sup> and R = 0.083 bar dm<sup>3</sup> K<sup>-1</sup> mol<sup>-1</sup>).
- Q18. (a) The balloon blown up at 5°C has volume of 480 mL. At this stage, the balloon is filled to 7.8 of its maximum stretching capacity.
- (i) What will be the volume of balloon at 30°C?
- (ii) Calculate the temperature at which the balloon will burst.
- (b) The sample of nitrogen occupies a volume of 320 cm<sup>3</sup> at STP. Calculate its volume at 546.3 K and 0.5 bar pressure.

**S1.** According to Boyle's law,  $V \propto \frac{1}{p}$  or  $pV = \text{constant}$

At a particular temperature,  $pV = \text{constant}$

Thus,  $p_1 V_1 = p_2 V_2 = p_3 V_3 = p_4 V_4$

As  $V_1 > V_2 > V_3 > V_4$

$\therefore p_1 < p_2 < p_3 < p_4$

**S2.** The van der Waals' radii increases as the number of energy shells increases and decreases as the nuclear charge increases. Since, H has only one energy shell and Cl has three, therefore the van der Waals' radius of H is smallest while that of Cl is largest. Further, both N and O have two energy shells but the nuclear charge on O (+ 8) is higher than that on N (+ 7), therefore the van der Waals' radius of N is greater than that of O. Hence, the overall increasing order is:

$$H < O < N < Cl.$$

**S3.** According to Boyle's law, at constant temperature.

$$p = \frac{k}{V}$$

On taking logarithm of both sides.

$$\log(p) = \log\left(\frac{1}{V}\right) + \log(k)$$

This equation can be compared to the equation of straight line,

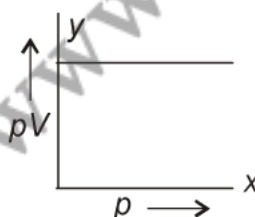
$$y = mx + c$$

Hence, a graph between  $\log(p)$  and  $\log\left(\frac{1}{V}\right)$  gives a straight line.



**S4.** The given  $T_c$  and  $P_c$  of  $\text{CO}_2$  gas are  $30.98^\circ\text{C}$  and  $73 \text{ atm}$ , respectively. This suggests that  $\text{CO}_2$  gas cannot be liquefied above  $30.98^\circ\text{C}$  and  $73 \text{ atm}$  pressure. However high temperature and pressure are being applied on  $\text{CO}_2$  gas. Hence,  $\text{CO}_2$  gas cannot be liquefied at  $32^\circ\text{C}$  and  $80 \text{ atm}$  pressure.

**S5.** Plot of  $pV$  against  $p$  is a straight line parallel to the x-axis, i.e., the slope will be zero.



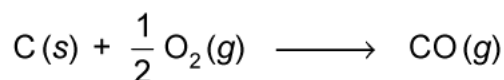
**S6.** Higher the critical temperature, more easily the gas is liquified, i.e., greater are the intermolecular forces and hence, greater is the value of  $a$ . Hence,  $\text{SO}_2$  has a greater value of  $a$ .

**S7.** (a) Given,  $V = 2.0 \text{ L}$ ,  $T = 25^\circ\text{C} = 25 + 273 = 298 \text{ K}$ ,  $N_{\text{O}_2} = 1.25 \text{ mol}$

As only gas exerts pressure, so pressure in the flask is entirely due to oxygen which can be calculated as:

$$pV = nRT$$
$$\Rightarrow p = \frac{nRT}{V} = \frac{1.25 \text{ mol} \times 0.082 \text{ L atm K}^{-1} \text{ mol}^{-1} \times 298 \text{ K}}{2.0 \text{ L}}$$
$$= 15.3 \text{ atm.}$$

(b) The reaction between C and O occurs for the formation of CO can be represented as:



As 1 mol of C reacts with  $\frac{1}{2}$  mole of  $\text{O}_2$ , therefore 3.3 mol of C will react with  $\frac{1}{2} \times 3.3 = 1.65 \text{ mol}$  of  $\text{O}_2$ .

But only 1.25 mol of  $\text{O}_2$  are available, therefore  $\text{O}_2$  will be the limiting reagent. Also,  $\frac{1}{2}$  mol of  $\text{O}_2$  produces 1 mol of CO.

$\therefore$  1.25 mol of  $\text{O}_2$  will produce  $= 1 \times 1.25 \times 2 = 2.50 \text{ mol}$

Now,

$$p = \frac{nRT}{V} = \frac{2.50 \text{ mol} \times 0.0821 \text{ L atm mol}^{-1} \text{ K}^{-1} \times 298 \text{ K}}{2 \text{ L}}$$
$$= 30.5 \text{ atm.}$$

**S8.** The density of a gas is given by,  $d = \frac{mp}{RT}$

For the same gas, at different temperatures and pressures

$$\frac{d_1}{d_2} = \frac{p_1}{T_1} \times \frac{T_2}{p_2} \quad \dots (i)$$

Here,  $d_1 = 10.92 \text{ g dm}^{-3}$ ,  $T_1 = 27^\circ\text{C}$  or  $300 \text{ K}$

$$p_1 = 4 \text{ bar}$$

At STP,  $d_2 = ?$   $T_2 = 0^\circ\text{C}$

or  $273 \text{ K}$ ,  $p_2 = 1 \text{ bar}$

$$\therefore \text{ From Eq. (i), we get } \frac{10.92}{d_2} = \frac{4 \text{ bar}}{300 \text{ K}} \times \frac{273 \text{ K}}{1 \text{ bar}}$$

$$\therefore d_2 = 3 \text{ g dm}^{-3}.$$

**S9.** (a)  $\text{CO}_2$  will exist in gaseous state between the points  $a$  and  $b$  at temperature  $T_1$ .

(b) At point  $b$ ,  $\text{CO}_2$  will start liquefying when temperature is  $T_1$ .

(c) At point  $g$ ,  $\text{CO}_2$  will be completely liquefied when temperature is  $T_2$ .

(d) No, condensation will not take place when the temperature is  $T_3$  because  $T_3 > T_c$ .

- S10.** (a) Given, number of moles,  $n = 0.25 \text{ mol}$ ,  
 Temperature,  $T = 27^\circ\text{C} = (273 + 27) \text{ K} = 300 \text{ K}$   
 and volume,  $V = 2.5 \text{ L} = 2.5 \times 10^{-3} \text{ m}^3$

Partial pressure of sulphur dioxide

$$p_{\text{SO}_2} = \frac{nRT}{V} = \frac{0.25 \text{ mol} \times 8.314 \text{ J mol}^{-1} \times 300 \text{ K}}{2.5 \times 10^{-3} \text{ m}^3}$$

$$= 249 \times 10^5 \text{ Nm}^{-2} = 2.49 \times 10^5 \text{ Pa}$$

Similarly,  $p_{\text{N}_2} = 2.49 \times 10^5 \text{ Pa}$

By using Dalton's law of partial pressures, we get

$$p_{\text{Total}} = p_{\text{N}_2} + p_{\text{SO}_2}$$

$$= 2.49 \times 10^5 \text{ Pa} + 2.49 \times 10^5 \text{ Pa} = 4.98 \times 10^5 \text{ Pa.}$$

- (b) Helium and few other gases (like hydrogen) has critical temperature lower than room temperature. So, they can be liquefied above room temperature, no matter how much pressure we apply.

**S11.** Moles of  $\text{CH}_4$ ,  $n_{\text{CH}_4} = \frac{\text{Mass of CH}_4}{\text{Molar mass of CH}_4}$  [ $\therefore$  Molar mass of  $\text{CH}_4 = 12 + 4 \times 1 = 16$ ]

$$= \frac{3.2}{16} = 0.2 \text{ mol}$$

Similarly, moles of  $\text{CO}_2$ ,  $n_{\text{CO}_2} = \frac{4.4}{44} = 0.1 \text{ mol}$  [ $\therefore$  Molar mass of  $\text{CH}_4 = 12 + 4 \times 1 = 16$ ]

$$\text{Total moles} = 0.2 + 0.1 = 0.3 \text{ mole}$$

On applying ideal equation,  $pV = nRT$

Pressure,  $p = \frac{nRT}{V} = \frac{0.3 \text{ mol} \times 0.0821 \text{ dm}^3 \text{ atm K}^{-1} \text{ mol}^{-1} \times 300 \text{ K}}{9 \text{ dm}^3}$

$$= 0.821 \text{ atm} = 8.314 \times 10^4 \text{ Pa.}$$

**S12.**  $p_{\text{Gas}} = p_{\text{Total}} - p_{\text{H}_2\text{O}}$   
 $p_{\text{Total}} = 0.993 \text{ bar}$ ,  $p_{\text{H}_2\text{O}} = 0.0317 \text{ bar}$   
 $\therefore p_{\text{Gas}} = 0.993 - 0.0317 = 0.9613 \text{ bar}$   
 Now,  $pV = nRT$

But,  $n = \frac{m}{M}$ ,  $pV = \frac{m}{M} RT$  or  $M = \frac{mRT}{pV}$

and  $m = 0.16 \text{ g}$

$$V = 135 \text{ mL} = \frac{135}{1000} \text{ dm}^3$$

$$T = 25 + 273 = 298 \text{ K}, \quad p = 0.9613 \text{ bar}$$

$$R = 0.083 \text{ bar dm}^3 \text{ mol}^{-1} \text{ K}^{-1}$$

Substituting the values,

$$M = \frac{(0.16 \text{ g}) \times (0.083 \text{ bar dm}^3 \text{ mol}^{-1} \text{ K}^{-1}) \times (298 \text{ K})}{0.9613 \text{ bar} \times 0.135 \text{ dm}^3}$$

$$= 30.49 \text{ g mol}^{-1}.$$

**S13.** (a) On increasing the temperature of a liquid, the kinetic energy of the molecules increases so that it can overcome the intermolecular forces of attraction and hence, the liquid can flow more easily, *i.e.*, viscosity of the liquid decreases.

(b) We know that,  $pV = nRT$

$$\Rightarrow V = \frac{nRT}{p} = \frac{W}{M} \cdot \frac{RT}{p}$$

$$\Rightarrow V = \frac{8.8 \text{ g}}{44 \text{ g mol}^{-1}} \times \frac{0.083 \text{ bar L K}^{-1} \text{ mol}^{-1} \times (273 + 31.1) \text{ K}}{1 \text{ bar}} = 5.05 \text{ L}$$

**S14.** Applying Boyle's law, (since, the temperature remains constant.)

For  $\text{N}_2$  gas,  $p'_{\text{N}_2} V' \text{ (after mixing)} = p_{\text{N}_2} V_{\text{N}_2} \text{ (before mixing)}$

$$p'_{\text{N}_2} = \frac{p_{\text{N}_2} V_{\text{N}_2}}{V'} = \frac{4 \times 4.0}{10} = \frac{16}{10} = 1.6 \text{ bar}$$

Similarly, for He gas,

$p'_{\text{He}} V' \text{ (after mixing)} = p_{\text{He}} V_{\text{He}} \text{ (before mixing)}$

$$p'_{\text{He}} = \frac{p_{\text{He}} V_{\text{He}}}{V'} = \frac{6 \times 6.0}{10} = \frac{36}{10} = 3.6 \text{ bar}$$

$\therefore$  Total pressure of mixture =  $1.6 + 3.6 = 5.2 \text{ bar}$ .

**S15.** (a)  $Z = \frac{pV}{nRT}$ , *i.e.*,  $0.5 = \frac{100 \times V}{1 \times 0.082 \times 273}$

$$V = 0.1119 \text{ L}$$

According to van der Waals' equation for 1 mole gas

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

Neglecting  $b$ ,  $\left( p + \frac{a}{V^2} \right) v = RT$

or  $pV + \frac{a}{V} = RT$

or  $\frac{pV}{RT} + \frac{a}{VRT} = 1$

or  $a = \left( 1 - \frac{pV}{RT} \right) VRT$

$$= (1 - 0.5) 0.1119 \times 0.082 \times 273$$

$$= 1.252 \text{ atm L}^2 \text{ mol}^{-2}.$$

(b) For a real gas, above Boyle's temperature, gas shows positive deviation and hence,  $Z > 1$ .

**S16.** (a) The surface tension of liquids generally decreases with increase in temperature and becomes zero at the critical temperature. It is obviously due to the fact that increase in temperature, the kinetic energy of the molecules increase and therefore, the intermolecular attraction decreases.

(b) Liquids have a property, called surface tension, due to which liquids tend to contract (to decrease) the surface area. For a given volume of a liquid, a sphere has the least surface area, hence the rain droplets tend to form spherical droplets.

(c) When temperature and pressure of a gas is 273.15 K (or 0°C) and 1 atm (or 1 bar or  $10^5$  Pa), such conditions are called standard temperature and pressure (STP) conditions.

Under these conditions, the volume occupied by 1 mol of each and every gas is 22.4 L.

Hence, the molar volume of  $N_2$  and Ar at 273.15 K and 1 atm is 22.4 L.

**S17.** Given, radius,

$$R_{\text{Balloon}} = 10$$

$$V_{\text{Balloon}} = \frac{4}{3} \pi r^3 = \frac{4}{3} \times \frac{22}{7} \times (10)^3 = 4190.476 \text{ m}^3$$

$$V_{\text{Balloon}} = V_{\text{Displaced air (by balloon)}}$$

$\therefore$  Mass of displaced air =  $V_{\text{Displaced air}} \times \text{Density of air}$

$$m_{\text{Displaced air}} = 4190.476 \text{ m}^3 \times 1.2 \text{ kg m}^{-3}$$

$$m = 5028.57 \text{ kg}$$

Mass of He filled in balloon,  $m_{\text{He}} = \frac{pVM}{RT}$   $\left( \because pV = \frac{mRT}{M} \right)$

$$m_{\text{He}} = \frac{1.66 \text{ bar} \times 4190.476 \times 10^3 \text{ dm}^3 \times 4 \times 10^{-3} \text{ kg mol}^{-1}}{0.083 \text{ bar dm}^3 \text{ K}^{-1} \text{ mol}^{-1} \times 300 \text{ K}}$$

$$m_{\text{He}} = 1117.46 \text{ kg}$$

Total mass of filled balloon,  $m_{\text{Balloon}} = \text{Mass of balloon} + \text{Mass of He filled in the balloon}$

$$m_{\text{Balloon}} = 100 \text{ kg} + 1117.46 \text{ kg} = 1217.46 \text{ kg}$$

Pay load = Mass of displaced air – Mass of balloon

$$= 5028.57 \text{ kg} - 1217.46 \text{ kg} = 3811.11 \text{ kg}.$$

**S18.** (a) 480 mL represents  $\frac{7}{8}$  of the balloon's maximum capacity.

Therefore, maximum volume of the balloon is  $480 \times \frac{8}{7} = 548.6 \text{ mL}$

Thus, if volume exceeds 548.6 mL, the balloon will burst.

(i) Calculation of the volume of balloon at 30°C.

By Charles' law, we get  $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

Here,  $V_1 = 480 \text{ mL}$ ,  $T_1 = 278 \text{ K}$ ,  $T_2 = 300 \text{ K}$ ,  $V_2 = ?$

Thus,  $V_2 = \frac{V_1 T_2}{T_1} = \frac{480 \times 300}{278} = 523.16 \text{ mL}$

(ii) Calculation of the temperature at which balloon will burst.

Here,  $V_1 = 480 \text{ mL}$ ,  $T_1 = 278 \text{ K}$ ,  $V_2 = 548.6 \text{ mL}$ ,  $T_2 = ?$

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

$$\begin{aligned} \therefore T_2 &= \frac{V_2 \times T_1}{V_1} = \frac{548.6 \times 278}{480} = 317.7 \text{ K} \\ &= 317.7 - 273 = 44.7^\circ\text{C} \end{aligned}$$

Thus, the balloon will burst as soon as temperature exceeds 44.7°C.

(b) Given,  $p_1 = 1.0 \text{ bar}$ ,  $p_2 = 0.5 \text{ bar}$ ,  $V_1 = 320 \text{ cm}^3$ ,  $T_1 = 273.15 \text{ K}$ ,  $T_2 = 546.3 \text{ K}$ ,  $V_2 = ?$

According to gas equation, we get

$$\frac{p_1 V_1}{T_1} = \frac{p_2 V_2}{T_2}$$

$$\therefore V_2 = \frac{p_1 V_1 \times T_2}{T_1 \times p_2} = \frac{1 \times 320 \times 546.3}{273.15 \times 0.5} = 1280 \text{ cm}^3.$$