

NCERT Solutions for Class-XI Chemistry

Chapter-5 NCERT Chemistry Class 11

1. What will be the minimum pressure required to compress 500 dm³ of air at 1 bar to 200 dm³ at 30°C?

1. Given that $P_1 = 1$ bar.

$$V_1 = 500 \text{ dm}^3$$

$$V_2 = 200 \text{ dm}^3$$

As temperature is constant, the final pressure can be calculated by Boyle's law,

$$P_1 V_1 = P_2 V_2$$

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

$$\Rightarrow P_2 = \frac{1 \text{ bar} \times 500 \text{ dm}^3}{200 \text{ dm}^3}$$

$$\Rightarrow P_2 = 2.5 \text{ bar}$$

Therefore, minimum pressure required is 2.5 bar.

2. A vessel of 120 mL capacity contains a certain amount of gas at 35 °C and 1.2 bar pressure. The gas is transferred to another vessel of volume 180 mL at 35 °C. What would be its pressure?

2. Given,

Initial pressure, $p_1 = 1.2$ bar

Initial volume, $V_1 = 120$ mL

Final volume, $V_2 = 180$ mL

Since the temperature remains constant, the final pressure (p_2) can be calculated using Boyle's law.

According to Boyle's law,

$$P_1 V_1 = p_2 V_2$$

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

$$= \frac{1.2 \times 120}{180} \text{ bar}$$

Therefore, the pressure would be 0.8 bar.

3. Using the equation of state $pV = nRT$; show that at a given temperature density of a gas is proportional to gas pressure.

3. The equation of state is given by,

$$pV = nRT \dots\dots\dots (i)$$

Where,

p- Pressure of gas

V- Volume of gas

n- Number of moles of gas

R- Gas constant

T -Temperature of gas

From equation (i) we have,

$$\frac{n}{V} = \frac{P}{RT}$$

Replacing $n = m/M$ we have

$$\Rightarrow \frac{m}{MV} = \frac{P}{RT}$$

because the number of moles is given as: $n = \frac{\text{given mass}}{\text{molar}}$

Where,

M - Mass of gas

M - Molar mass of gas

$$\text{But } \frac{m}{V} = d$$

(where, d = density of gas)

Thus, from equation (ii), we have

$$\Rightarrow \frac{d}{M} = \frac{P}{RT}$$

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

Molar mass (M) of a gas is always constant and therefore, at constant temperature T, Hence, at a given temperature, the density (d) of gas is proportional to its pressure (p).

4. At 0°C, the density of a certain oxide of a gas at 2 bar is same as that of dinitrogen at 5 bar. What is the molecular mass of the oxide?
4. Density (d) of the substance at temperature (T) can be given by the expression,

$$d = \frac{Mp}{RT}$$

Now, density of oxide (d_1) is given by,

$$d_1 = \frac{M_1 p_1}{RT}$$

Where, M_1 and p_1 are the mass and pressure of the oxide respectively.

Density of dinitrogen gas (d_2) is given by,

$$d_2 = \frac{M_2 p_2}{RT}$$

Where, M_2 and p_2 are the mass and pressure of the oxide respectively.

According to the given question,

5. Pressure of 1 g of an ideal gas A at 27 °C is found to be 2 bar. When 2 g of another ideal gas B is introduced in the same flask at same temperature the pressure becomes 3 bar. Find a relationship between their molecular masses.
5. For ideal gas A, the ideal gas equation is given by,
 $\Rightarrow P_a V = n_a RT$
Where, P_a and n_a represent the pressure and number of moles of gas A.
For ideal gas B, the ideal gas equation is given by,
 $\Rightarrow P_b V = n_b RT$

Where, P_a and n_b represent the pressure and number of moles of gas B.
[V and T are constants for gases A and B]

Therefore we have,

$$\Rightarrow \frac{P_a M_a}{m_a} = \frac{P_b M_b}{m_b}$$

Where, M_a and M_b are the molecular masses of gases A and B respectively.

Now, we have

Substituting the values

$$m_a = 1 \text{ g}$$

$$P_a = 2 \text{ bar}$$

$$m_b = 2 \text{ g}$$

$$P_b = 1 \text{ bar.}$$

Given,

(Since total pressure is 3 bar)

Substituting these values in equation we have

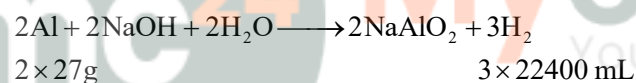
$$\Rightarrow \frac{2 \times M_a}{1} = \frac{1 \times M_b}{2}$$

Thus, a relationship between the molecular masses of A and B is given by

$$\Rightarrow 4M_a = M_b$$

6. The drain cleaner, Drainex contains small bits of aluminum which react with caustic soda to produce dihydrogen. What volume of dihydrogen at 20 °C and one bar will be released when 0.15g of aluminum reacts?

6. The reaction of aluminium with caustic soda can be represented as:



At STP (273.15 K and 1 atm), 54 g (2×27 g) of Al gives 3×22400 mL of H_2 ..

\therefore 0.15g Al gives $\frac{3 \times 22400 \times 0.15}{54}$ mL of H_2 i.e., 186.67 mL of H_2 .

At STP,

$$P_1 = 1 \text{ atm}$$

$$V_1 = 186.67 \text{ mL}$$

$$T_1 = 273.15 \text{ K}$$

Let the volume of dihydrogen V_2 be at $p_2 = 0.987$ atm (since 1 bar = 0.987 atm) and $T_2 = 20^\circ\text{C} = (273.15 + 20) \text{ K} = 293.15 \text{ K}$..

Now,

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

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

$$= \frac{1 \times 186.67 \times 293.15}{0.987 \times 273.15}$$

$$= 202.98 \text{ mL}$$

$$= 203 \text{ mL}$$

Therefore, 203 mL of dihydrogen will be released.

7. What will be the pressure exerted by a mixture of 3.2 g of methane and 4.4 g of carbon dioxide contained in a 9 dm³ flask at 27 °C?

7. It is known that,

$$\Rightarrow P = \frac{mRT}{MV}$$

For methane (CH₄),

$$\Rightarrow P = \frac{3.2\text{g} \times 8.314 \frac{\text{J}}{\text{mol} \cdot \text{K}} \times 300\text{K}}{16 \frac{\text{g}}{\text{mol}} \times 9 \text{ dm}^3 \times 0.01}$$

$$\Rightarrow P = 55430 \text{ Pa}$$

For carbon dioxide

$$\Rightarrow P = \frac{4.4\text{g} \times 8.314 \frac{\text{J}}{\text{mol} \cdot \text{K}} \times 300\text{K}}{44 \frac{\text{g}}{\text{mol}} \times 9 \text{ dm}^3 \times 0.01}$$

$$\Rightarrow P = 27710 \text{ Pa}$$

Total pressure exerted by the mixture can be obtained as addition of both partial pressures.

Hence, the total pressure exerted by the mixture is 8.314×10^4 Pa.

8. What will be the pressure of the gaseous mixture when 0.5 L of H₂ at 0.8 bar and 2.0 L of dioxygen at 0.7 bar are introduced in a 1L vessel at 27°C?

8. Let the partial pressure of H₂ in the vessel be P_{H₂}.

Now,

$$P_1 = 0.8 \text{ bar} \quad p_2 = P_{\text{H}_2}$$

$$V_1 = 0.5 \text{ L} \quad V_2 = 1 \text{ L} = ?$$

It is known that,

$$p_1 V_1 = p_2 V_2$$

$$\Rightarrow p_2 = \frac{p_1 V_1}{V_2}$$

$$\Rightarrow P_{\text{H}_2} = \frac{0.8 \times 0.5}{1}$$

$$= 0.4 \text{ bar}$$

Now, let the partial pressure of O₂ in the vessel be P_{O₂}.

Now,

$$p_1 = 0.7 \text{ bar} \quad p_2 = P_{\text{O}_2} = ?$$

$$V_1 = 2.0 \text{ L} \quad V_2 = 1 \text{ L}$$

$$p_1 V_1 = p_2 V_2$$

$$\Rightarrow p_2 = \frac{p_1 V_1}{V_2}$$

$$\Rightarrow P_{\text{O}_2} = \frac{0.7 \times 2.0}{1}$$

$$= 0.4 \text{ bar}$$

Total pressure of the gas mixture in the vessel can be obtained as:

$$P_{\text{total}} = p_{\text{H}_2} + p_{\text{O}_2}$$

$$= 0.4 + 1.4$$

$$= 1.8 \text{ bar}$$

Hence, the total pressure of the gaseous mixture in the vessel is 1.8 bar.

9. Density of a gas is found to be 5.46 g/dm³ at 27 °C at 2 bar pressure. What will be its density at STP?

9. We know that the density expression is

$$d = Mp/RT$$

where,

d – density of the gas

p-Pressure of gas

V- Volume of gas

n- Number of moles of gas

R- Gas constant

T -Temperature of gas

For a single component,

$$\Rightarrow \frac{d_1}{d_2} = \frac{P_1 T_2}{P_2 T_1}$$

$$\Rightarrow d_2 = \frac{1 \text{ bar} \times 300 \text{ K} \times 5.46 \text{ g / dm}^3}{2 \text{ bar} \times 273 \text{ K}}$$

$$d_2 = 3 \text{ g/dm}^3$$

10. 34.05 mL of phosphorus vapour weighs 0.0625 g at 546 °C and 0.1 bar pressure. What is the molar mass of phosphorus?

10. Given, p =

$$0.1 \text{ bar}$$

$$V =$$

$$= 34.05$$

$$\text{mL} =$$

$$34.05 \times$$

$$10^{-3} \text{ L} =$$

$$34.05 \times$$

$$10^{-3} \text{ dm}^3$$

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

$$T = 546^\circ\text{C} = (546 + 273) \text{ K} = 819 \text{ K}$$

The number of moles (n) can be calculated using the ideal gas equation as:

$$pV = nRT$$

$$\Rightarrow n = \frac{pV}{RT}$$

$$= \frac{0.1 \times 34.05 \times 10^{-3}}{0.083 \times 819}$$

$$= 5.01 \times 10^{-5} \text{ mol}$$

Therefore, molar mass of phosphorus = $\frac{0.0625}{5.01 \times 10^{-5}} = 1247.5 \text{ g mol}^{-1}$

Hence, the molar mass of phosphorus is $1247.5 \text{ g mol}^{-1}$.

11. A student forgot to add the reaction mixture to the round bottomed flask at 27°C but instead he/she placed the flask on the flame. After a lapse of time, he realized his mistake, and using a pyrometer he found the temperature of the flask was 477°C . What fraction of air would have been expelled out?

11. Given:

Initial temperature, $T_1 = 27^\circ\text{C} = (27 + 273) \text{ K} = 300\text{K}$

Final temperature, $T_2 = 477^\circ\text{C} (477 + 273) \text{ K} = 750\text{K}$

Let the volume of the round bottomed flask be V .

Then, the volume of air inside the flask at 300K is, $V_1 = V$.

Let the volume of air inside the flask at 750K is $= V_2$

According to Charles's law we know that: -

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

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

$$= \frac{750V}{300}$$

$$(\because V_1 = V)$$

$$= 2.5V$$

\therefore Volume of air expelled out $= 2.5V - V = 1.5V$

Hence, the fraction of the air expelled out is given by:

$$\frac{1.5V}{2.5V} = \frac{3.5}{5}$$

\therefore The fraction of the air expelled out is : $3/5$.

12. Calculate the temperature of 4.0 mol of a gas occupying 5 dm^3 at 3.32 bar . ($R = 0.083 \text{ bar dm}^3 \text{ K}^{-1} \text{ mol}^{-1}$).

12. Given, $n =$

4.0 mol $V =$

5 dm^3 $p =$

3.32 bar

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

The temperature (T) can be calculated using the ideal gas equation as:

$$pV = nRT$$

$$\Rightarrow T = \frac{pV}{nR}$$

$$= \frac{3.32 \times 5}{4 \times 0.083}$$

$$= 50\text{K}$$

Hence, the required temperature is 50 K .

13. Calculate the total number of electrons present in 1.4 g of dinitrogen gas.

13. Given:

Mass of dinitrogen = 1.4 g

We know that molar mass of dinitrogen = 28 g mol^{-1}

\therefore The number of moles of dinitrogen in 1.4g = $1.4/28 = 0.05 \text{ mol}$

Now, the number of molecules in 1.4g of dinitrogen:

$$= 0.05 \times 6.022 \times 10^{23}$$

$$= 3.01 \times 10^{23}$$

We know that 1 molecule of dinitrogen contains 14 electrons.

$\therefore 3.01 \times 10^{23}$ molecules of dinitrogen contains:

$$= 14 \times 3.01 \times 10^{23}$$

$$= 4.214 \times 10^{23} \text{ electrons.}$$

Hence, the total number of electrons present in 1.4 g of dinitrogen gas is 4.214×10^{23} .

14. How much time would it take to distribute one Avogadro number of wheat grains, if 10^{10} grains are distributed each second?

14. Avogadro number = 6.02×10^{23}

Thus, time required

$$= \frac{6.02 \times 10^{23}}{10^{10}} \text{ s}$$

$$= 6.02 \times 10^{13} \text{ s}$$

$$= \frac{6.02 \times 10^{13}}{60 \times 60 \times 24 \times 365} \text{ years}$$

$$= 1.909 \times 10^6 \text{ years}$$

Hence, the time taken would be 1.909×10^6 years.

15. Calculate the total pressure in a mixture of 8 g of dioxygen and 4 g of dihydrogen confined in a vessel of 1 dm^3 at 27°C . $R = 0.083 \text{ bar dm}^3 \text{ K}^{-1} \text{ mol}^{-1}$.

15. Given:

Mass of dioxygen = 8 g

Mass of dihydrogen = 4 g

Now, molar mass of dioxygen = 32g

Thus, number of moles of dioxygen = $8/32 = 0.25$ mole

Also, molar mass of dihydrogen = 2g

Thus, number of moles of dihydrogen = $4/2 = 2$ mole

Therefore, total number of moles in the mixture:

$$= 0.25 + 2 = 2.25 \text{ mole}$$

Given: -

$V = \text{volume} = 1 \text{ dm}^3$,

$n = \text{number of moles} = 2.25 \text{ mol}$

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

$T = \text{temperature} = 27^\circ\text{C} = 300 \text{ K}$

Let the total pressure be P .

From ideal gas equation we know that –

$$PV = nRT$$

$$\begin{aligned} \Rightarrow P &= \frac{nRT}{V} \\ &= \frac{2.25 \times 0.083 \times 300}{1} \\ &= 56.025 \text{ bar.} \end{aligned}$$

Hence, total pressure of the mixture is 56.025 bar.

- 16.** 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 at 27°C. (Density of air = 1.2 kg m⁻³ and R = 0.083 bar dm³ K⁻¹ mol⁻¹).

- 16.** Given,

Radius of the balloon, r = 10 m

$$\therefore \text{Volume of the balloon} = \frac{4}{3}\pi r^3$$

$$= \frac{4}{3} \times \frac{22}{7} \times 10^3$$

$$= 4190.5 \text{ m}^3 \text{ (approx)}$$

Thus, the volume of the displaced air is 4190.5 m³.

Given,

Density of air = 1.2 kg m⁻³

Then, mass of displaced air = 4190.5 × 1.2 kg

$$= 5028.6 \text{ kg}$$

Now, mass of helium (m) inside the balloon is given by,

$$m = \frac{MpV}{RT}$$

$$M = 4 \times 10^{-3} \text{ kg mol}^{-1}$$

$$p = 1.66 \text{ bar}$$

V = Volume of the balloon

$$= 4190.5 \text{ m}^3$$

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

$$T = 27^\circ\text{C} = 300\text{K}$$

$$\text{Then, } m = \frac{4 \times 10^{-3} \times 1.66 \times 4190.5 \times 10^3}{0.083 \times 300}$$

$$= 1117.5 \text{ kg (approx)}$$

Now, total mass of the balloon filled with helium = (100 + 1117.5) kg

$$= 1217.5 \text{ kg}$$

Hence, pay load = (5028.6 – 1217.5) kg

$$= 3811.1 \text{ kg}$$

Hence, the pay load of the balloon is 3811.1 kg.

- 17.** Calculate the volume occupied by 8.8 g of CO₂ at 31.1°C and 1 bar pressure.

$$R = 0.083 \text{ bar L K}^{-1} \text{ mol}^{-1}.$$

- 17.** From ideal gas equation we know that: -

$$PV = nRT$$

$$\Rightarrow PV = \frac{m}{M} \times RT$$

Where, P = pressure = 1 bar

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

T = temperature = $31.1 + 273 = 304.1 \text{ K}$

m = mass of $\text{CO}_2 = 8.8 \text{ g}$

M = molar mass $\text{CO}_2 = 44 \text{ g}$

Let the volume be V

$$\therefore V = \frac{mRT}{PM}$$

$$= \frac{8.8 \times 0.083 \times 304.1}{1 \times 44}$$

$$= 5.05 \text{ L}$$

\therefore The volume occupied is 5.05 L.

18. 2.9 g of a gas at 95°C occupied the same volume as 0.184 g of dihydrogen at 17°C , at the same pressure. What is the molar mass of the gas?

18. Volume (V) occupied by dihydrogen is given by,

$$V = \frac{m}{M} \frac{RT}{p}$$

$$= \frac{0.184}{2} \times \frac{R \times 290}{p}$$

Let M be the molar mass of the unknown gas. Volume (V) occupied by the unknown can be calculated as:

$$V = \frac{m}{M} \frac{RT}{p}$$

$$= \frac{2.9}{M} \times \frac{R \times 368}{p}$$

According to the question,

$$\frac{0.184}{2} \times \frac{R \times 290}{p} = \frac{2.9}{M} \times \frac{R \times 368}{p}$$

$$\Rightarrow \frac{0.184 \times 290}{2} = \frac{2.9 \times 368}{M}$$

$$\Rightarrow M = \frac{2.9 \times 368 \times 2}{0.184 \times 290}$$

$$= 40 \text{ g mol}^{-1}$$

Hence, the molar mass of the gas is 40 g mol^{-1} .

19. A mixture of dihydrogen and dioxygen at one bar pressure contains 20% by weight of dihydrogen. Calculate the partial pressure of dihydrogen.

19. Let the weight of dihydrogen be 20 g and the weight of dioxygen be 80 g.

Then, the number of moles of dihydrogen:

$$n(\text{H}_2) = 20/2 = 10 \text{ moles}$$

The number of moles of dioxygen:

$$n(\text{O}_2) = 80/32 = 2.5 \text{ moles}$$

Given that,

Total pressure of the mixture, $p(\text{total}) = 1 \text{ bar}$

Then, partial pressure of dihydrogen:

$$\begin{aligned} P(\text{H}_2) &= \frac{n(\text{H}_2)}{n(\text{H}_2) + n(\text{O}_2)} p(\text{total}) \\ &= \frac{10}{10 + 2.5} \times 1 \\ &= 0.8 \text{ bar} \end{aligned}$$

Hence, the partial pressure of dihydrogen is 0.8 bar.

20. What would be the SI unit for the quantity pV^2T^2/n ?

20. The SI unit for pressure, p is Nm^{-2} .

The SI unit for volume, V is m^3 .

The SI unit for temperature, T is K .

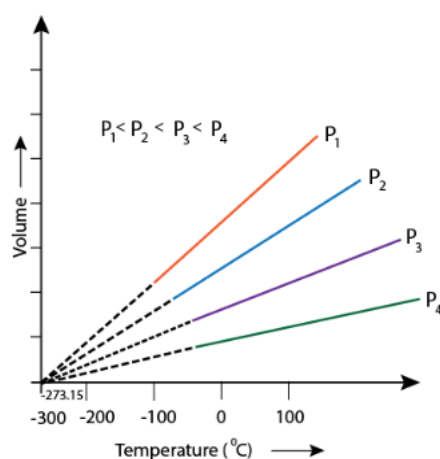
The SI unit for the number of moles, n is mol .

Therefore, the SI unit for quantity $\frac{pV^2T^2}{n}$ is given by,

$$\begin{aligned} &\frac{(\text{Nm}^{-2})(\text{m}^3)^2 (\text{K})^2}{\text{mol}} \\ &= \text{Nm}^4\text{K}^2 \text{mol}^{-1} \end{aligned}$$

21. In terms of Charles' law explain why -273°C is the lowest possible temperature.

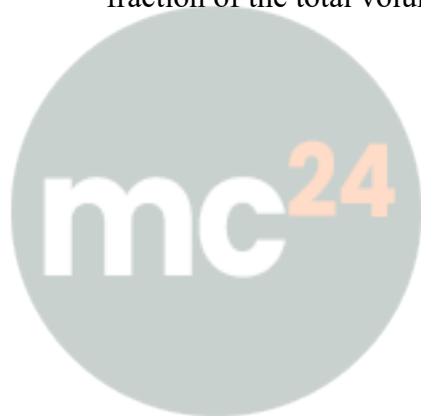
21. Charles' law states that at constant pressure, the volume of a fixed mass of gas is directly proportional to its absolute temperature.



Volume Vs Temperature graph ($^\circ\text{C}$)

From the above graph between V and T we can find that at constant pressure if temperature is lowered, there is also a decrease in the volume at constant pressure. The volume should therefore become zero at -273°C . Any further lowering is impossible as it will imply negative volume. Hence, -273°C is the lowest possible temperature.

22. Critical temperature for carbon dioxide and methane are $31.1\text{ }^{\circ}\text{C}$ and $-81.9\text{ }^{\circ}\text{C}$ respectively. Which of these has stronger intermolecular forces and why?
22. Higher is the critical temperature of a gas, easier is its liquefaction. This means that the intermolecular forces of attraction between the molecules of a gas are directly proportional to its critical temperature. Hence, intermolecular forces of attraction are stronger in the case of CO_2 .
23. Explain the physical significance of Van der Waals parameters.
23. Physical significance of 'a':
'a' is a measure of the magnitude of intermolecular attractive forces within a gas. This implies that if a molecule has a large value of a then the intermolecular forces of attraction between its molecules will be large and hence it could be easily liquefied.
- Physical significance of 'b':
'b' is a measure of the volume of a gas molecule. It gives us the volume occupied by all the gaseous molecules contained in a given vessel which was earlier neglected according to kinetic theory and this correction term becomes more significant at high pressure as the molecules are themselves incompressible and occupy an appreciable fraction of the total volume.



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