NCERT Exercise

Question 1:
What will be the minimum pressure required to compress 500 dm^3 of air at 1 bar to 200 dm^3 at 30°C?

Solution 1:
Given,
Initial pressure, p_1 = 1 bar
Initial volume, V_1= 500 dm^3
Final volume, V_2 = 200 dm^3
Since the temperature remains constant, the final pressure (p_2) can be calculated using Boyle's law.
According to Boyle's law,

$$p_1V_1 = p_2V_2$$

$$\Rightarrow p_2 = \frac{p_1V_1}{V_2}$$

$$= \frac{1\times500}{200}\text{bar}$$

$$= 2.5 \text{ bar}$$

Therefore, the minimum pressure required is 2.5 bar.

Question 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 answer vessel of volume 180 mL at 35 °C. What would be its pressure?

Solution 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,

$$p_1V_1 = p_2V_2$$

$$p_2 = \frac{p_1V_1}{V_2}$$

$$= \frac{1.2\times120}{180}\text{bar}$$

$$= 0.8 \text{ bar}$$

Study Materials

[NCERT Solutions for Class 6 to 12 (Math & Science)]
[Revision Notes for Class 6 to 12 (Math & Science)]
[RD Sharma Solutions for Class 6 to 12 Mathematics]
[RS Aggarwal Solutions for Class 6, 7 & 10 Mathematics]
[Important Questions for Class 6 to 12 (Math & Science)]
[CBSE Sample Papers for Class 9, 10 & 12 (Math & Science)]
[Important Formula for Class 6 to 12 Math]
[CBSE Syllabus for Class 6 to 12]
[Lakhmir Singh Solutions for Class 9 & 10]
[Previous Year Question Paper]
[CBSE Class 12 Previous Year Question Paper]
[CBSE Class 10 Previous Year Question Paper]
[JEE Main & Advanced Question Paper]
[NEET Previous Year Question Paper]

Therefore, the pressure would be 0.8 bar.

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

Solution 3:
The equation of state is given by,
$pV = nRT$ ………… (i)
Where,
$p \to$ Pressure of gas
$V \to$ Volume of gas
$n \to$ Number of moles of gas
$R \to$ Gas constant
$T \to$ Temperature of gas
From equation (i) we have,
$$\frac{n}{V} = \frac{p}{RT}$$
Replacing n with $\frac{m}{M}$, we have
$$\frac{m}{MV} = \frac{p}{RT} \text{ …….. (ii)}$$
Where,
$m \to$ Mass of gas
$M \to$ Molar mass of gas
But, $\frac{m}{V} = d$ (d = density of gas)
Thus, from equation (ii), we have
$$\frac{d}{M} = \frac{p}{RT}$$
$$\Rightarrow d = \left(\frac{M}{RT}\right)p$$
Molar mass (M) of gas is always constant and therefore, at constant temperature $(T), \frac{M}{RT}$ = constant,
d = (constant)p
$\Rightarrow d \propto p$
Hence, at a given temperature, the density (d) of gas is proportional to its pressure (p)

Question 4:
At 0ºC, the density of 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?

Solution 4:
Density (d) of 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,
$$d_1 = d_2$$
$$\therefore M_1 p_1 = M_2 p_2$$
Given,
p_1 = 2 bar
p_2 = 5 bar
Molecular mass of nitrogen, M_2 = 28 g/mol
Now, $M_1 = \dfrac{M_2 p_2}{p_1}$
$$= \frac{28 \times 5}{2}$$
$$= 70\,\text{g/mol}$$
Hence, the molecular mass of the oxide is 70 g/mol.

Question 5:
Pressure of 1 g of an ideal gas A 27 ºC 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.

Solution 5:
For ideal gas A, the ideal gas equation is given by,
$$p_B V = n_B RT \;\;\ldots\ldots\ldots \text{(ii)}$$
Where, p_B and n_B represent the pressure and number of moles of gas B.
[V and T are constants for gases A and B]

From equation (i), we have

$$p_A V = \frac{m_A}{M_A} RT \Rightarrow \frac{p_A M_A}{m_A} = \frac{RT}{V} \quad \ldots\ldots \text{(iii)}$$

From equation (ii), we have

$$p_B V = \frac{m_B}{M_B} RT \Rightarrow \frac{p_B M_B}{m_B} = \frac{RT}{V} \quad \ldots\ldots\text{(iv)}$$

Where, M_A and M_B are the molecular masses of gases A and B respectively.
Now, from equations (iii) and (iv), we have

$$\frac{p_A M_B}{m_B} = \frac{p_B M_B}{m_B} \quad \ldots\ldots \text{(v)}$$

Given,

$m_A = 1g$

$p_A = 2\,bar$

$m_B = 2g$

$p_B = (3\text{-}2) = 1\,bar$

(Since total pressure is 3 bar)

Substituting these values in equation (v), we have

$$\frac{2 \times M_A}{1} = \frac{1 \times M_B}{1}$$

$$\Rightarrow 4M_A = M_B$$

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

$$4M_A = M_B$$

Question 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?

Solution 6:

The reaction of aluminum with caustic soda can be represented as:

$$2Al + 2NaOH + 2H_2 0 \rightarrow 2NaAlO_2 + 3H_2$$

$2 \times 27 g \qquad\qquad\qquad\qquad 3 \times 22400\, mL$

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

$\therefore$ 0.15 g 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 be V_2 at $p_2 = 0.987$ atm (since 1 bar = 0.987 atm) and $T_2 = 20^oC$ = (273.15 + 20) K = 293.15 K.

$$\frac{p_1V_1}{T_1} = \frac{p_2V_2}{T_2}$$

$$\Rightarrow V_2 = \frac{p_1V_1T_2}{p_2T_1}$$

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

$$= 202.98 \text{ mL}$$

$$= 203 \text{ mL}$$

Therefore, 203 mL of dihydrogen will be released.

Question 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 d^3 flask at 27 ^oC?

Solution 7:

It is known that,

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

For methane (CH_4),

$$p_{CH_4} = \frac{3.2}{16} \times \frac{8.314 \times 300}{9 \times 10^{-3}} \begin{bmatrix} \text{Since } 9\, dm^3 = 9 \times 10^{-3} m^3 \\ 27^oC = 300K \end{bmatrix}$$

$$= 5.543 \times 10^4 \text{Pa}$$

For carbon dioxide (CO_2),

$$p_{CO_4} = \frac{4.4}{44} \times \frac{8.314 \times 300}{9 \times 10^{-3}}$$

$$= 2.771 \times 10^4 \text{Pa}$$

The pressure exerted by the mixture can be obtained as:

$$p = p_{CH_4} + p_{CO_2}$$

$$= \left(5.543 \times 10^4 + 2.771 \times 10^4\right) Pa$$

$$= 8.314 \times 10^4 \, Pa$$

Hence, the total pressure exerted by the mixture is $= 8.314 \times 10^4 \, Pa$.

Question 8:
What will be the pressure of he gaseous mixture when 0.5 L of H_2 at 0.8 bar and 2.0 L of dioxygen at 0.7 bar are introduced in a 1L vessel at 27°C?

Solution 8:
Let the partial pressure of H_2 in the vessel be p_{H_2}.
Now,
$p_1 = 0.8\,\text{bar}$ $\quad$ $p_2 = p_{H_2} = ?$
$V_1 = 0.5\text{L}$ $\quad$ $V_2 = 1\text{L}$
It is known that,

$$p_1V_1 = p_2V_2$$

$$\Rightarrow p_2 = \frac{p_1V_1}{V_2}$$

$$\Rightarrow p_{H_2} = \frac{0.8 \times 0.5}{1}$$

$$= 0.4\,bar$$

Now, let the partial pressure of O_2 in the vessel be p_{O_2}.
$p_1 = 0.7\,\text{bar}$ $\quad$ $p_2 = p_{O_2} = ?$
$V_1 = 2.0\,\text{L}$ $\quad$ $V_2 = 1\text{L}$

$$p_1V_1 = p_2V_2$$

$$\Rightarrow p_2 = \frac{p_1V_1}{V_2}$$

$$\Rightarrow p_{O_2} = \frac{0.7 \times 20}{1}$$

$$= 0.4\,bar$$

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

$$p_{total} = p_{H_2} + p_{O_2}$$

$$= 0.4 + 1.4$$

$$= 1.8\ bar$$

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

Question 9:
Density of a gas is found to be 5.46 g/dm^3 at 27 °C at 2 bar pressure. What will be its density at STP?

Solution9:
Given,
$d_1 = 5.46\, g / dm^3$
$p_1 = 2\, bar$
$T_1 = 27^o C = (27 + 273) K = 300\, K$
$p_2 = 1 bar$
$T_2 = 273\, K$
$d_2 = ?$
The density (d_2) of the gas at STP can be calculated using the equation,

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

$$\therefore \frac{d_1}{d_2} = \frac{\frac{Mp_1}{RT_1}}{\frac{Mp_2}{RT_2}}$$

$$\Rightarrow \frac{d_1}{d_2} = \frac{p_1 T_2}{p_2 T_1}$$

$$\Rightarrow d_2 = \frac{p_2 T_1 d_1}{p_1 T_2}$$

$$= \frac{1 \times 300 \times 5.46}{2 \times 273}$$

$$= 3 g dm^{-3}$$

Hence, the density of the gas at STP will be 3 g dm^{-3}.

Question 10:
34.05 mL of phosphorus vapour weighs 0.0625 g at 546 ^oC and 0.1 bar pressure. What is the molar mass of phosphorus?

Solution10:
Given,
p = 0.1 bar
V = 34.05 mL = 34.05 × 10^{-3} L = 34.05 × 10^{-3} dm^{-3}
R =0.083 bar dm^3 K^{-1} mol^{-1}
T = 546^{o}C = (546 + 273) K = 819 K
The number of mass (n) can be calculated using the ideal gas equation as:

$$pV = nRT$$
$$\Rightarrow n = \frac{pV}{RT}$$
$$= \frac{0.1 \times 34.05 \times 10^{-3}}{0.083 \times 819}$$
$$= 5.01 \times 10^{-5}\ mol$$

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

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

Question 11:
A student forgot to add the reaction mixture to the round bottomed flask at 27 ^{o}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 ^{o}C. What fraction of air would have been expelled out?

Solution11:
Let the volume of the round bottomed flask be V.
Then, the volume of air inside the flask at 27 ^{o}C is V.
Now,
V_1 = V
T_1 = 27 ^{o}C = 300 K
V_2 = ?
T_2 = 477 ^{o}C = 750 K
According to Charles's law,

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$
$$\Rightarrow V_2 = \frac{V_1 T_2}{T_1}$$
$$= \frac{750V}{300}$$
$$= 2.5V$$

Therefore, volume of air expelled out = 2.5 V – V = 1.5 V

Hence, fraction of air expelled out $= \frac{1.5V}{2.5V} = \frac{3}{5}$

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

Solution12:
Given,
n = 4.0 mol
V = 5 dm^3
p = 3.32 bar
R = 0.083 bar dm^3 K^{-1} mol^{-1}
The temperature (T) can be calculated using the ideal gas equation as:

$$pV = nRT$$

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

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

$$= 50\,K$$

Hence, the required temperature is 50 K.

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

Solution13:
Molar mass of dinitrogen (N_2) = 28 g mol^{-1}
Thus, 1.4 g of $N_2 = \frac{1.4}{28} = 0.05\,mol$
= 0.05 × 6.02 $\times 10^{23}$ number of molecules
= 3.01 $\times 10^{23}$ number of molecules
Now, 1 molecule of N_2 contains 14 electrons.
Therefore, 3.01 × 10^{23} molecules of N_2 contains = 14 × 3.01 × 1023
= 4.214 $\times 10^{23}$ electrons

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

Solution14:
Avogadro number = 6.02 × 10^{23}
Thus, time required

$= \frac{6.02\times10^{23}}{10^{10}}s$

$=6.02\times10^{23}s$

$= \frac{6.02\times10^{23}}{60\times60\times24\times365}$ years

$=1.909\times10^{6}$ years

Hence, the time taken would be $=1.909\times10^{6}$ years .

Question 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 bar dm^3 K^{-1} mol^{-1}).

Solution 15:

Given,

Mass of dioxygen (O_2) = 8 g

Thus, number of moles of $O_2 = \frac{8}{32} = 0.25$ mole

Mass of dihydrogen (H_2) = 4 g

$H_2 = \frac{4}{2} = 2$ mole

Therefore, total number of moles in the mixture = 0.25 + 2 2.25 mole

Given,

V = 1 dm3

n = 2.25 mol

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

T = 27°C = 300 K

Total pressure (p) can be calculated as:

$pV = nRT$

$\Rightarrow p = \frac{nRT}{V}$

$= \frac{225\times0.083\times300}{1}$

$= 56.025\, bar$

Hence, the total pressure of the mixture is 56.025 bar.

Question 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 27°C. (Density of air = 1.2 kg m^{-3}. And R = 0.083 bar dm^{-3} K^{-1} mol^{-1}).

Solution16:
Given,
Radius of the balloon, r = 10 m

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

$= \frac{4}{3} \times \frac{22}{7} \times 10^{23}$

$= 4190.5\, m^3$ (approx)

Thus, the volume of the displaced air is 4190.5 m^3.
Given,
Density of air = 1.2 kg m^{-3}
Then, mass of displaced air = 4190.5 $\times$ 1.2 kg
= 5028.6 kg
Now, mass of helium (m) inside the balloon is given by,

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

Here,
M = 4 $\times$ 10^{-3} kg mol^{-1}
p = 1.66 bar
V = Volume of the balloon
  = 4190.5 m^3
R = 0.083 bar dm^3 K^{-1} mol^{-1}
T = 27°C = 300 K

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

$= 1117.5\, kg$ (approx)

Now, total mass of the balloon filled with helium = (100 + 1117.5) kg
= 1217.5 kg
Hence, pay load = (5028.6 – 1217.5) kg
=3811.1 kg
Hence, the pay load of the balloon is 3811.1 kg.

Question 17:
Calculate the volume occupied by 8.8 g of CO_2 at 31.1°C and 1 bar pressure.
R = 0.083 bar L K^{-1} mol^{-1}.

Solution 17:
It is known that,

$pV = \frac{m}{N}RT$

$\Rightarrow V = \frac{mRT}{Mp}$

Here,
m = 8.8 g
R = 0.083 bar LK^{-1} mol^{-1}
T = 31.1°C = 304.1 K
M = 44 g
p = 1 bar

Thus, Volume (V) $= \frac{8.8 \times 0.083 \times 304.1}{44 \times 1}$
= 5.04806 L
= 5.05 L

Hence, the volume occupied is 5.05 L.

Question 18:

2.9 g of 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?

Solution 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 gas can be calculated as:

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

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

According to the equation,

$$\frac{0.184}{2}\times\frac{R\times 290}{p}=\frac{2.9}{M}\times\frac{R\times 368}{p}$$
$$\Rightarrow \frac{0.184\times 290}{2}=\frac{2.9\times 368}{M}$$
$$\Rightarrow M=\frac{2.9\times 368\times 2}{0.184\times 290}$$
$$=40\,g\,mol^{-1}$$
Hence, the molar mass of the gas is 40 g mol^{-1}.

Question 19:
A mixture of dihydrogen and dioxygen atone bar pressure contains 20% by weight of dihydrogen. Calculate the partial pressure of dihydrogen.

Solution 19:
Let the weight of dihydrogen be 20 g and the weight of dioxygen be 80 g.
Then, the number of moles of dihydrogen, $n_{H_2}=\frac{20}{2}=10\,moles$ and the number of moles of dioxygen, $n_{O_2}=\frac{80}{32}=2.5\,moles$.
Given,
Total pressure of the mixture, $P_{total}=1$ bar
Then, partial pressure of dihydrogen,
$$p_{H_2}=\frac{n_{H_2}}{n_{H_2}+n_{O_2}}\times P_{total}$$
$$=\frac{10}{10+2.5}\times 1$$
$$=0.8\,bar$$
Hence, the partial pressure of dihydrogen is 0.8 bar.

Question 20:
What would be the SI units for the quantity pV^2T^2/n?

Solution 20:
The SI units 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.

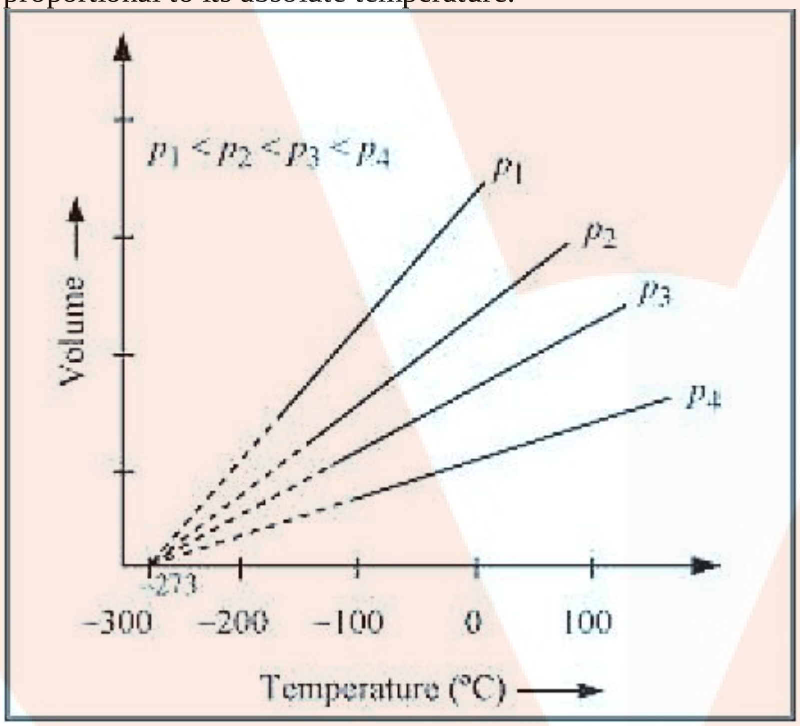
Therefore, the SI unit for quantity $\frac{pV^2T^2}{n}$ is given by,

$$= \frac{(\text{Nm}^{-2})(\text{m}^3)^2(\text{K})^2}{\text{mol}}$$

$$= \text{Nm}^4\text{K}^2\,\text{mol}^{-1}$$

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

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



It was found that for all gases (at any given pressure), the plots of volume vs. temperature (in°C) is a straight line. If this line is extended to zero volume, then it intersects the temperature-axis at –273°C. In other words, the volume of any gas at 273°C is zero. This is because all gases get liquefied before reaching a temperature of 273°C. Hence, it can be concluded that – 273°C is the lowest possible temperature.

Question 22:
Critical temperature for carbon dioxide and methane are 31.1 °C and – 81.9 °C respectively.

Which of these has stronger intermolecular forces and why?

Solution 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.

Question 23:
Explain the physical significance of Van der Waals parameters.

Solution23:
The vander waals equation is an equation of state for a fluid composed of particles that have a non-zero volume and a pair wise attractive inter-particle force(Vander waals force) The equation is

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

Physical significance of ‘a’:
‘a’ is a measure of the magnitude of intermolecular attractive forces within a gas.
Physical significance of ‘b’:
‘b’ is a measure of the volume of a gas molecule.
V is the total volume of the container containing the fluid.