Q1. What will be the minimum pressure required to compress 500 dm³ of air at 1 bar to 200 dm³ at 30°C?

Answer:

Initial pressure, P_1 = 1 bar

Initial volume, V_1 = 500

dm^3

Final volume, V_2 = 200
dm^3
As the temperature remains the same, the final pressure (P_2) can be calculated with the help of Boyle's law.

According to Boyle's law,

$P_1V_1 = P_2V_2$

P_2 =

$\frac{P_1V_1}{V_2}$

=

$\frac{1 \times 500}{200}$
= 2.5 bar

∴ the *minimum pressure required to compress is* 2.5 bar.

Q2. 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?

Answer:

Initial pressure, P_1 = 1.2 bar

Initial volume, V_1 = 120 mL

Final volume, V_2 = 180 mL

As the temperature remains the same, the final pressure (P_2) can be calculated with the help of Boyle's law.

According to Boyle's law,

$P_1V_1 = P_2V_2$

$P_2 =$

$\frac{P_1V_1}{V_2}$

=

$\frac{1.2 \times 120}{180}$

= 0.8 bar

Therefore, the min pressure required is 0.8 bar.

Q3. Using the equation of state pV=nRT, show that at a given temperature density of a gas is proportional to gas pressure p.

Answer:

The equation of state is given by

pV = nRT ……..(1)

Where, p = Pressure

V = Volume

N = Number of moles

R = Gas constant

T = Temperature

$\frac{n}{V}$

=

$\frac{p}{RT}$

Replace n with

$\frac{m}{M}$

, therefore,

$\frac{m}{MV}$
=
$\frac{p}{RT}$
........(2)

Where, m = Mass

M = Molar mass

But,

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

Where, d = Density

Therefore, from equation (2), we get

$\frac{d}{M}$
=
$\frac{p}{RT}$

d = (
$\frac{M}{RT}$
) p

d
$\propto$
P

Therefore, at a given temperature, the density of the gas (d) is proportional to its pressure (p).

Q4. At 0°C, the density of a certain oxide of a gas at 2 bar is the same as that of dinitrogen at 5 bar. What is the molecular mass of the oxide?

Answer:

The density (d) of the substance at temperature (T) can be given by

d =

$\frac{Mp}{RT}$

Now, the density of oxide (d_1) is given as

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

Where M_1 = Mass of the oxide

p_1 = Pressure of the oxide

The density of dinitrogen gas (d_2) is given as,

d_2 = $\frac{M_1 p_2}{RT}$

Where M_2 = Mass of the oxide

p_2 = Pressure of the oxide

According to the 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 = $\frac{M_2 p_2}{p_1}$ = $\frac{28 \times 5}{2}$

= 70 g/mol

Therefore, the molecular mass of the oxide is 70 g/mol.

Q5. The 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 the same temperature, the pressure becomes 3 bar. Find a relationship between their molecular masses.

Answer:

For ideal gas A, the ideal gas equation is given by,

$$p_X V = n_X RT$$

......(1)

Where

p_X

and

n_X

epresents the pressure and number of moles of gas X.
For ideal gas Y, the ideal gas equation is given by

$$p_Y V = n_Y RT$$

......(2)

Where

p_Y

and

n_Y

represents the pressure and number of moles of gas Y.
[V and T are constants for gases X and Y]
From equation (1),

$$p_X V = \frac{m_X}{M_X}$$

RT

$\frac{p_X M_X}{m_X}$

=

$\frac{RT}{V}$

......(3)

From equation (2),

$p_Y V = \frac{m_Y}{M_Y}$
RT
$\frac{p_Y M_Y}{m_Y}$
=
$\frac{RT}{V}$
...... (4)

Where

M_X
and
M_Y

are the molecular masses of gases X and Y, respectively.
Now, from equations (3) and (4),

$\frac{p_X M_X}{m_X}$
=
$\frac{p_Y M_Y}{m_Y}$
..... (5)

Given,

m_X
= 1 g
p_X
= 2 bar
m_Y
= 2 g
p_Y

= (3 – 2) = 1 bar (Since total pressure is 3 bar.)
Substituting these values in equation (5),

$\frac{2 \times M_X}{1}$
=
$\frac{1 \times M_Y}{2}$
4

M_X
=
M_Y

Therefore, the relationship between the molecular masses of X and Y is

4

M_X
=
M_Y

Q6. The drain cleaner, Drainex, contains small bits of aluminium 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 aluminium reacts?

Answer:

The reaction of aluminium with caustic soda is given below.

$2Al + 2NaOH + 2H_2O$

$\rightarrow$
$2NaAlO_2 + 3H_2$
At Standard Temperature Pressure (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

= 186.67 mL of H_2

At Standard Temperature Pressure,

p_1
= 1 atm
V_1
= 186.67 mL
T_1
= 273.15 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°

C = (273.15 + 20) K = 293.15 K.
Now,

$\frac{p_1V_1}{T_1}$
=
$\frac{p_2V_2}{T_2}$
$V_2 = \frac{p_1V_1T_2}{p_2T_1}$

=

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

= 202.98 mL

= 203 mL

Hence, 203 mL of dihydrogen will be released.

Q7. 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 dm^3 flask at 27 °C?

Answer:

It is known that,

p =

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

For methane (CH_4),

p_{CH_4}
=

$\frac{3.2}{16}$
×
$\frac{8.314 \times 300}{9 \times 10^{-3}}$

[Since 9 dm³ =

9×10^{-3}
m^3
27^{o}

27∘
C = 300 K]
= 5.543 ×

10^4

Pa
For carbon dioxide (CO_2),

p_{CO_2}
=

$\frac{4.4}{44}$
×
$\frac{8.314 \times 300}{9 \times 10^{-3}}$

= 2.771 ×

10^4
Pa

The total pressure exerted by the mixture can be calculated as

p =

p_{CH_4}
+
p_{CO_2}

= (5.543 ×

10^4
+ 2.771 ×
10^4
) Pa

= 8.314 ×

10^4
Pa

Q8. What will be the pressure of the 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?

Answer:

Let the partial pressure of

H_2
in the container be
p_{H_2}

Now,

p_1
= 0.8 bar
p_2
=
p_{H_2}
V_1
= 0.5 L
V_2
= 1 L
It is known that,

p_1
V_1
=
p_2
V_2
p_2
=
$\frac{p_1 \times V_1}{V_2}$
p_{H_2}
=
$\frac{0.8 \times 0.5}{1}$
= 0.4 bar

Now, let the partial pressure of O_2 in the container be

p_{O_2}
.

Now,

p_1
= 0.7 bar
p_2
=
p_{O_2}
V_1
= 2.0 L
V_2
= 1 L
p_1
V_1
=
p_2
V_2
p_2
=
$\frac{p_1 \times V_1}{V_2}$
p_{O_2}
=
$\frac{0.7 \times 20}{1}$
= 1.4 bar

The total pressure of the gas mixture in the container can be obtained as

p_{total}
=
p_{H_2}
+
p_{O_2}
= 0.4 + 1.4

= 1.8 bar

Q9. The 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?

Answer:

Given,

d_1 = 5.46 g/dm^3

p_1 = 2 bar

T_1 =

$$27^{\circ}$$

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 =

$$\frac{Mp}{RT}$$

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

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

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}

Q10. 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?

Answer:

Given,

p = 0.1 bar

V = 34.05 mL = 34.05 × 10^{-3}

dm^3

R = 0.083 bar

dm^3
at K^{-1} mol^{-1}

T =

546^oC

= (546 + 273) K = 819 K
The no. of moles (n) can be calculated using the ideal gas equation as
pV = nRT

n =

$\frac{pV}{RT}$

=

$\frac{0.1\times34.05\times10^{-3}}{0.083\times819}$

Therefore, molar mass of phosphorus =

$\frac{0.0625}{5.01\times10^{-5}}$

= 125 g mol^{-1}

Q11. A student forgot to add the reaction mixture to the container at

27^o

C, but instead, he placed the container on the flame. After a lapse of time, he came to know about his mistake and using a pyrometer, he found the temp of the container

477°

C. What fraction of air would have been expelled out?

Answer:

Let the volume of the container be V.

The volume of the air inside the container at

27°

C is V.

Now,

$V_1 = V$

$T_1 =$

27°

C = 300 K $V_2 = ?$

$T_2 =$

477°

C = 750 K

According to Charles's law,

$\frac{V_1}{T_1}$

=

$\frac{V_2}{T_2}$

V_1

=

$\frac{V_1 T_2}{T_1}$

=

$\frac{750V}{300}$

= 2.5 V

Therefore, the volume of air expelled out

= 2.5 V – V = 1.5 V

Hence, the fraction of air expelled out

=

$\frac{1.5V}{2.5V}$

=

$\frac{3}{5}$

Q12. Calculate the temperature of 4.0 mol of a gas occupying 5 dm^3 at 3.32 bar. (R = 0.083 bar dm^3 K^{-1} mol^{-1}).

Given,

N= 4.0 mol

V = 5

dm^3

p = 3.32 bar

R = 0.083 bar

dm^3

at K^{-1} mol^{-1}

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

pV = nRT

T =

$\frac{pV}{nR}$

=

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

= 50 K

Therefore, the required temp is 50 K.

Q13. Calculate the total number of electrons present in 1.4 g of dinitrogen gas.

Answer:

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 × 10^{23} no. of molecules

= 3.01 × 10^{23} no. of molecules

Now, 1 molecule of N_2 has 14 electrons.

Therefore, 3.01 × 10^{23} molecules of N_2 contains

= 14 × 3.01 × 1023

= 4.214 × 10^{23} electrons

Q14. How much time would it take to distribute one Avogadro number of wheat grains, if 10^{10} grains are distributed each second?

Answer:

Avogadro no. = 6.02 × 10^{23}

Therefore, the time taken

=

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

= 6.02 × 10^{13} s

=

$$\frac{6.02 \times 10^{23}}{60 \times 60 \times 24 \times 365} years$$

= 1.909×10^6 years

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

Q15. Calculate the total pressure in a mixture of 8 g of dioxygen and 4 g of dihydrogen confined in a vessel of 1 dm³ at 27°C. R = 0.083 bar dm³ K⁻¹ mol⁻¹

Answer:

Given:

Mass of O_2 = 8 g

No. of moles

=

$$\frac{8}{32}$$

= 0.25 mole

Mass of H_2 = 4 g

No. of moles

=

$$\frac{4}{2}$$

= 2 mole

Hence, the total no. of moles in the mixture

= 0.25 + 2

= 2.25 mole

Given:

V = 1

dm^3

n = 2.25 mol

R = 0.083 bar

dm^3

at K^{-1} mol^{-1}
T =

27°

C = 300 K

Total pressure :
pV = nRT

p =

$\frac{nRT}{V}$

=

$\frac{225\ \times\ 0.083\ \times\ 300}{1}$

= 56.025 bar

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

Q16. Payload is defined as the difference between the mass of displaced air and the mass of the balloon. Calculate the payload 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 dm^3 K^{-1} mol^{-1})

Answer:

Given:

r = 10 m

Therefore, the volume of the balloon

=

$\frac{4}{3}$
πr^3

=

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

= 4190.5 m^3 (approx.)

Therefore, the volume of the displaced air

= 4190.5 × 1.2 kg

= 5028.6 kg

Mass of helium,

=

$\frac{MpV}{RT}$

Where, M = 4 × 10^{-3} kg mol^{-1}

p = 1.66 bar

V = Volume of the balloon

= 4190.5 m^3

R = 0.083 0.083 bar

dm^3

at 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 with helium

= (100 + 1117.5) kg

= 1217.5 kg

Therefore, payload

= (5028.6 – 1217.5)

= 3811.1 kg

Therefore, the payload of the balloon is 3811.1 kg.

Q17. Calculate the volume occupied by 8.8 g of CO_2 at 31.1°C and 1 bar pressure. R = 0.083 bar dm^3 K^{-1} mol^{-1}.

Answer:

pVM = mRT

V =

$$\frac{mRT}{Mp}$$

Given:

m = 8.8 g

R = 0.083 bar

$$dm^3$$

at K^{-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

Therefore, the volume occupied is 5.05 L.

Q18. 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?

Answer:

Volume,

V =

$$\frac{mRT}{Mp}$$

=

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

Let M be the molar mass of the unknown gas.

The volume occupied by the unknown gas is

=

$$\frac{mRT}{Mp}$$

=

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

According to the question,

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

=

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

$$\frac{0.184 \times 290}{2}$$

=

$$\frac{2.9 \times 368}{M}$$

M =

$$\frac{2.9 \times 368 \times 2}{0.184 \times 290}$$

= 40 g mol^{-1}

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

Q19. A mixture of dihydrogen and dioxygen at one bar pressure contains 20% by weight of dihydrogen. Calculate the partial pressure of dihydrogen.

Answer:

Let the weight of dihydrogen be 20 g.

Let the weight of dioxygen be 80 g.

No. of moles of dihydrogen (nH_2)

=

$$\frac{20}{2}$$

= 10 moles

No. of moles of dioxygen (nO_2)

=

$$\frac{80}{32}$$

= 2.5 moles

Given:

p_{total} = 1 bar

Therefore, partial pressure of dihydrogen (pH_2)

=

$$\frac{n_{H_2}}{n_{H_2} + n_{O_2}}$$

× p_{total}

=

$$\frac{10}{10 + 2.5} \times 1$$

= 0.8 bar

Therefore, the partial pressure of dihydrogen is 0.8 bar.

Q20. What will be the SI unit for the quantity

$\frac{pV^2T^2}{n}$

?

Answer:

SI unit of pressure, p =

Nm^{-2}

SI unit of volume, V =

m^3

SI unit of temp, T = K

SI unit of the number of moles, n = mol

Hence, SI unit of

$\frac{pV^2T^2}{n}$

is

=

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

=

$Nm^4K^2mol^{-1}$

Q21. In terms of Charles' law explain why

-273°

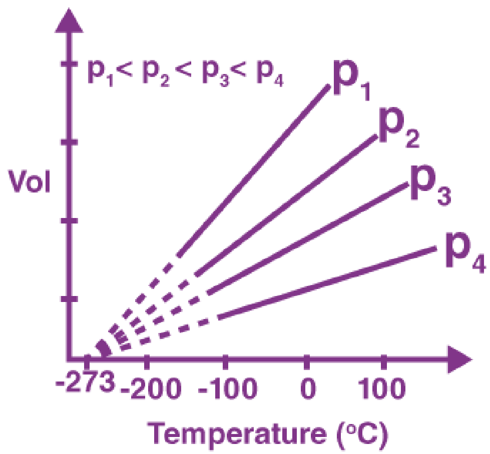
C is the lowest possible temperature.

Answer:

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

Charles found that for all gases, at any given pressure, the graph of volume vs temperature (in Celsius) is a straight line, and on extending to zero volume, each line intercepts the temperature axis at – 273.15°C.

We can see that the volume of the gas at – 273.15°C will be zero. This means that gas will not exist. In fact, all the gases get liquified before this temperature is reached.



Q22. The 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?

Answer:

If the critical temperature of a gas is higher, then it is easier to liquefy. That is, the intermolecular forces of attraction among the molecules of gas are directly proportional to its critical temperature.

Therefore, in CO_2 intermolecular forces of attraction are stronger.

Q23. Explain the physical significance of Van der Waals parameters.

Answer:

After accounting for pressure and volume corrections, the van der Waals equation is

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

The van der Waals constants or parameters are a and b.

The relevance of a and b is crucial here.

The magnitude of intermolecular attractive forces within the gas is measured by the value of ‘a,’ which is independent of temperature and pressure.

The volume occupied by the molecule is represented by ‘b,’ while the total volume occupied by the molecules is represented by ‘nb.’