#420265

Topic: Gas laws

Molar volume is the volume occupied by 1 mol of any (ideal) gas at standard temperature and pressure (STP : 1 atmospheric pressure, 0^oC). Show that it is 22.4 litres.

Solution

The ideal gas equation relating pressure (P), volume (V), and absolute temperature (T) is given as:

PV = nRT

Where,

R is the universal gas constant $= 8.314\, J\, mol^{-1}K^{-1}$

n = Number of moles = 1

T = Standard temperature = 273 K

P = Standard pressure $=$ 1 atm $= 1.013 \times 10^5 Nm^{-2}$

$\therefore V = nRT/P$

$= 1 \times 8.314 \times 273/(1.013 \times 10^5)$

$= 0.0224 m^3$

$= 22.4$ litres

Hence, the molar volume of a gas at STP is 22.4 litres.

#420282

Topic: Gas laws

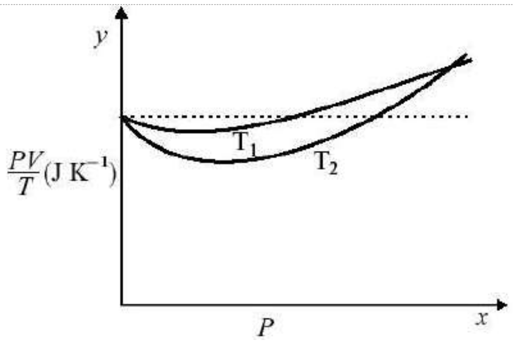

Figure 13.8 shows plot of PV/T versus P for $1.00 \times 10^{-3} kg$ of oxygen gas at two different temperatures.

(a) What does the dotted plot signify?

(b) Which is true : $T_1 > T_2\ or\ T_1 < T_2$?

(c) What is the value of PV / T where the curves meet on the y-axis?

(d) If we obtained similar plots for $1.00 \times 10^{-3} kg$ of hydrogen, would we get the same value of PV/T at the point where the curves meet on the y-axis? If not, what mass of hydrogen yields the same value of PV/T (for low pressure high temperature region of the plot) ? (Molecular mass of $H_2 = 2.02\, u$, of $O_2 = 32.0\, u$, $R = 8.31\, J\, mol^{-1} K^{-1}$.)

Solution

(a) The dotted plot in the graph signifies the ideal behaviour of the gas, i.e., the ratio PV/T is equal.

μR (μ is the number of moles and R is the universal gas constant) is a constant quantity. It is not dependent on the pressure of the gas.

(b) The dotted plot in the given graph represents an ideal gas. The curve of the gas at temperature T_1 is closer to the dotted plot than the curve of the gas at temperature T_2. A real gas approaches the behaviour of an ideal gas when its temperature increases. Therefore, $T_1 > T_2$ is true for the given plot.

(c) The value of the ratio PV/T, where the two curves meet, is μR. This is because the ideal gas equation is given as:

$PV = \mu RT$

$PV/T = \mu R$

Where,

P is the pressure

T is the temperature

V is the volume

μ is the number of moles

R is the universal constant

Molecular mass of oxygen $= 32.0g$

Mass of oxygen $= 1 \times 10^{-3} kg = 1g$

$R = 8.314\, J\, mol^{-1} K^{-1}$

$\therefore PV/T = (1/32) \times 8.314$

$= 0.26\, JK^{-1}$

Therefore, the value of the ratio PV/T, where the curves meet on the y-axis, is $0.26\, JK^{-1}$.

(d) If we obtain similar plots for $1.00 \times 10^{-3} kg$kg of hydrogen, then we will not get the same value of PV/T at the point where the curves meet the y-axis. This is because the molecular mass of hydrogen (2.02 u) is different from that of oxygen (32.0 u).

We have:

$PV/T = 0.26 JK^{-1}$

$R = 8.314\, J\, mol^{-1} K^{-1}$

Molecular mass (M) of H_2 = 2.02 u

$PV/T = \mu R$at constant temperature

Where, $\mu = mM$

$m =$ Mass of H_2

$\therefore m = (PV/T) \times (M/R)$

$= 0.26 \times 2.02/8.31$

$= 6.3 \times 10^{-2} g = 6.3 \times 10^{-5} kg$

Hence, $6.3 \times 10^{-5} kg$of H_2 will yield the same value of PV/T.

#420290

Topic: Gas laws

An oxygen cylinder of volume 30 litres has an initial gauge pressure of 15 atm and a temperature of 27^oC. After some oxygen is withdrawn from the cylinder the gauge pressure drops to 11 atm and its temperature drops to 17^oC. Estimate the mass of oxygen taken out of the cylinder.

($R = 8.31\, J\, mol^{-1}\, K^{-1}$, molecular mass of $O_2 = 32\, u$).

Solution

Volume of oxygen, $V_1 =$ 30 litres $= 30 \times 10^{-3} m^3$

Gauge pressure, $P_1 =$ 15 atm $= 15 \times 1.013 \times 10^5 Pa$

Temperature, $T_1 = 27^o C = 300K$

Universal gas constant, $R = 8.314\, J\, mol^{-1} K^{-1}$

Let the initial number of moles of oxygen gas in the cylinder be n_1 .

The gas equation is given as:

$P_1 V_1 = n_1 R T_1$

$\therefore n_1 = P_1 V_1 / R T_1$

$= (15.195 \times 10^5 \times 30 \times 10^{-3})/(8.314 \times 300) = 18.276$

But $n_1 = m_1/M$

Where,

$m_1 =$ Initial mass of oxygen

M = Molecular mass of oxygen = 32 g

$\therefore m_1 = N_1 M = 18.276 \times 32 = 584.84g$

After some oxygen is withdrawn from the cylinder, the pressure and temperature reduces.

Volume, $V_2 =$ 30 litres $= 30 \times 10^{-3} m^3$

Gauge pressure, $P_2 =$ 11 atm $= 11 \times 1.013 \times 10^5 Pa$

Temperature, $T_2 = 17^o C = 290K$

Let n_2 be the number of moles of oxygen left in the cylinder.

The gas equation is given as:

$P_2 V_2 = n_2 R T_2$

$\therefore n_2 = P_2 V_2 / R T_2$

$= (11.143 \times 10^5 \times 30 \times 10^{-30})/(8.314 \times 290) = 13.86$

But $n_2 = m_2/M$

Where,

m_2 is the mass of oxygen remaining in the cylinder

$\therefore m_2 = n_2 \times M = 13.86 \times 32 = 453.1g$

The mass of oxygen taken out of the cylinder is given by the relation:

Initial mass of oxygen in the cylinder – Final mass of oxygen in the cylinder

$= m_1 - m_2$

$= 584.84g – 453.1g$

$= 131.74g$

$= 0.131kg$

Therefore, 0.131 kg of oxygen is taken out of the cylinder.

#420297
Topic: Gas laws

An air bubble of volume 1.0 cm^3 rises from the bottom of a lake 40 m deep at a temperature of 12 $^\circ C$. To what volume does it grow when it reaches the surface which is at a temperature of 35$^\circ C$?

Solution

Volume of the air bubble, $V_1 = 1.0cm^3 = 1.0 \times 10^{-6}m^3$

Bubble rises to height, $d = 40m$

Temperature at a depth of 40 m, $T_1 = 12^oC = 285K$

Temperature at the surface of the lake, $T_2 = 35^oC = 308K$

The pressure on the surface of the lake: $P_2 = 1atm = 1 \times 1.103 \times 10^5 Pa$

The pressure at the depth of 40 m: $P_1 = 1\,atm + d\rho g$

Where,

ρ is the density of water $= 10^3 kg/m^3$

g is the acceleration due to gravity $= 9.8m/s^2$

$\therefore P_1 = 1.103 \times 10^5 + 40 \times 10^3 \times 9.8 = 493300Pa$

We have $\dfrac{P_1V_1}{T_1} = \dfrac{P_2V_2}{T_2}$

Where, V_2 is the volume of the air bubble when it reaches the surface.

$V_2 = \dfrac{P_1V_1T_2}{T_1P_2}$

$= \dfrac{493300 \times 1 \times 10^{-6} \times 308}{285 \times 1.013 \times 10^5}$

$= 5.263 \times 10^{-6}m^3$ or $5.263\,cm^3$

Therefore, when the air bubble reaches the surface, its volume becomes $5.263\,cm^3$.

#420307

Topic: Gas laws

Estimate the total number of air molecules (inclusive of oxygen, nitrogen, water vapour and other constituents) in a room of capacity 25.0 m^3 at a temperatured of 27 $^\circ C$ and 1 atm pressure.

Solution

Volume of the room, $V = 25.0m^3$

Temperature of the room, $T = 27^0C = 300K$

Pressure in the room, $P = 1\,atm = 1 \times 1.013 \times 10^5 Pa$

The ideal gas equation relating pressure (P), Volume (V), and absolute temperature (T) can be written as:

$PV = k_BNT$

Where,

k_B is Boltzmann constant $= 1.38 \times 10^{-23} m^2 kg\, s^{-2} K^{-1}$.

N is the number of air molecules in the room.

$\therefore N = \dfrac{PV}{k_BT}$

$= \dfrac{1.013 \times 10^5 \times 25}{1.38 \times 10^{-23} \times 300}$

$= 6.11 \times 10^{26}$ molecules

Therefore, the total number of air molecules in the given room is 6.11×10^{26}.

#420318

Topic: Gas laws

Estimate the average thermal energy of a helium atom at

(i) room temperature (27 $^\circ C$),

(ii) the temperature on the surface of the Sun (6000 K),

(iii) the temperature of 10 million kelvin (the typical core temperature in the case of a star).

Solution

(i) At room temperature, $T = 27^oC = 300K$

Average thermal energy $= (3/2)kT$

Where, k is Boltzmann constant $= 1.38 \times 10^{-23} m^2 kg\, s^{-2} K^{-1}$

$\therefore (3/2)kT = (3/2) \times 1.38 \times 10^{-38} \times 300$

$= 6.21 \times 10^{-21} J$

Hence, the average thermal energy of a helium atom at room temperature 27^oC is $6.21 \times 10^{-21} J$.

(ii) On the surface of the sun, $T = 6000K$

Average thermal energy $= (3/2)kT$

$= (3/2) \times 1.38 \times 10^{-38} \times 6000$

$= 1.241 \times 10^{-19} J$

Hence, the average thermal energy of a helium atom on the surface of the sun is $1.241 \times 10^{-19} J$.

(iii) At temperature, $T = 10^7 K$

Average thermal energy $= (3/2)kT$

$= (3/2) \times 1.38 \times 10^{-23} \times 10^7$

$= 2.07 \times 10^{-16} J$

Hence, the average thermal energy of a helium atom at the core of a star is $2.07 \times 10^{-16} J$.

#422939

Topic: Gas laws

What will be the minimum pressure required to compress 500 dm^3 of air at 1 bar to 200 dm^3 at 30^0 C?

Solution

The minimum pressure required to compress air is $P' = \dfrac{PV}{V'} = \dfrac{1 \times 500}{200} = 2.5\ bar$

#422942

Topic: Gas laws

A vessel of 120 mL capacity contain amount of gas at 35 $^\circ C$ and 1.2 bar pressure The gas is transferred to another vessel of volume 180 mL at 35 $^\circ C$. What would be its pressure?

Solution

The temperature and the amount of gas is constant. Hence, pressure is calculated from Boyle's law.

$P' = \dfrac{PV}{V'} = \dfrac{1.2 \times 120}{180} = 0.8 bar$

#422945

Topic: Gas laws

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

Solution

$PV = nRT = \dfrac{w}{M} RT$

Hence, $PM = \dfrac{w}{V} RT = \rho RT$

$\rho \propto P$

Here, $\rho = \dfrac{w}{V}$

$n = \dfrac{w}{M}$

ρ is the density, P is the pressure, V is the volume, n is the number of moles, R is the ideal gas constant, T is the temperature and M is the molar mass.

#422948

Topic: Gas laws

At $0^{\circ}C$ 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

The expression for density is $\rho = \dfrac{MP}{RT}$.

When the temperature is constant, $M' = \dfrac{MP}{P'}$.

Here, M and P are the molar mass and pressures of nitrogen and M' and P' are the molar mass and pressures of oxide.

Substitute values in the above equation.

$M' = \dfrac{28 \times 5}{2} = 70\ u$

#422951

Topic: Gas laws

Pressure of 1 g of an ideal gas A at 27 $^{\circ}C$ 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 ba

Find a relationship between their molecular masses?

Solution

Let M and M' be the molar masses of ideal gases A and B respectively.

The number of moles of gas A and B are $\dfrac{1}{M}$ and $\dfrac{2}{M'}$ respectively.

Let P and P' be the pressures of gases A and B respectively.

$P = 2\ bar$

$P + P' = 3\ bar$

$P' = 3 - 2 = 1\ bar$

The ideal gas equations for two gases A and B are,

$PV = nRT$......(i)

$P'V = n'RT$......(ii)

Divide equation (i) by equation (ii).

$\dfrac{P}{P'} = \dfrac{n}{n'} = \dfrac{1 \times M'}{2 \times M} = \dfrac{M'}{2M}$

$\dfrac{M'}{M} = 2\dfrac{P}{P'} = \dfrac{2 \times 2}{1} = 4$

$M' = 4M$

#422954

Topic: Gas laws

The drain cleaner, Drainex contains small bits of aluminum which react with caustic soda to produce dihydrogen. What volume of dihydrogen at $20^{\circ}C$ and one bar will be realeased when 0.15g of aluminum reacts?

Solution

The balanced chemical equation is $2Al + 2NaOH + 2H_2O \rightarrow 2NaAlO_2 + 3H_2$

54 g (2 mol) of Al will give 3 moles (or 67.2 L) of hydrogen.

0.15 g of Al will give $\dfrac{67.2}{54} \times 0.15 = 0.189L$of hydrogen.

The volume of hydrogen is at 1 bar and 273 K. The volume at 1 bar and 20^0C will be, $V' = \dfrac{PVT'}{P'T} = \dfrac{1 \times 0.189 \times 293}{1 \times 273} = 0.2028L = 202.8\ ml$

#422960

Topic: Gas laws

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 $^{\circ}C$?

Solution

The number of moles of methane is $\dfrac{3.2}{16} = 0.2$.

The pressure exerted by methane is $P = \dfrac{nRT}{V} = \dfrac{0.2 \times 8.314 \times 300}{9 \times 10^{-3}} = 5.54 \times 10^4 Pa$

The number of moles of CO_2 is $\dfrac{4.4}{44} = 0.1$.

The pressure exerted by CO_2 is $P' = \dfrac{n'RT}{V} = \dfrac{0.1 \times 8.314 \times 300}{9 \times 10^{-3}} = 2.77 \times 10^4 Pa$

The total pressure is $P + P' = 5.54 \times 10^4 Pa + 2.77 \times 10^4 Pa = 8.31 \times 10^4 Pa$

#422965

Topic: Gas laws

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 $^\circ C$?

Solution

(i) The partial pressure of hydrogen gas

$$P' = \frac{PV}{V'} = \frac{0.8 \times 0.5}{1} = 0.40 \text{bar}$$

(ii) The partial pressure of oxygen gas

$$P'' = \frac{PV}{V'} = \frac{0.7 \times 2.0}{1} = 1.40 \text{bar}$$

(iii) Total pressure of gaseous mixture is $P + P'' = 0.40 + 1.40 = 1.80$bar

#422968

Topic: Gas laws

Density of a gas is found to be 5.46 dm^3 at 27 $^\circ C$ at 2 bar pressure. What will be its density at STP?

Solution

The expression for density is $d = \dfrac{PM}{RT}$.

For a given gas, R and M are constant.

$$d \propto P/T$$

$$d' = \frac{dP'T}{PT'} = \frac{5.46 \times 1 \times 300}{2 \times 273} = 3.0\ g\ dm^{-3}$$

#422972

Topic: Gas laws

34.05 mL of phosphorus vapour weighs 0.0625 g at 546 $^\circ C$ and 0.1 bar pressure. What is the molar mass of phosphorus?

Solution

The molar mass of phosphorus, $M = \dfrac{wRT}{PV} = \dfrac{0.0625 \times 0.08314 \times 819}{1 \times 34.05 \times 10^{-3}} = 124.98$

#422975

Topic: Gas laws

A student forgot to add the reaction mixture to the round bottomed flask at 27 $^\circ C$ 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 $477^\circ C$. What fraction of air would have been expelled out?

Solution

The initial temperature $T = 27^oC = 27 + 273K = 300K$

The final temperature $T' = 477^oC = 477 + 273K = 750K$

The inital volume is V L and the final volume is V' L.

$$\frac{V}{V'} = \frac{T}{T'}$$
$$\frac{V}{V'} = \frac{300}{750}$$
$$\frac{V}{V'} = \frac{2}{5'}$$

The fraction of air expelled is $1 - \dfrac{V}{V'} = 1 - \dfrac{2}{5} = \dfrac{3}{5}$

#422977

Topic: Gas laws

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})

Solution

From ideal gas equation,

$$T = \frac{PV}{nR} = \frac{3.32 \times 5}{4.0 \times 0.083} = 50\ K$$

Thus, the temperature of the gas is 50 K.

#422985

Topic: Gas laws

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 $^\circ C$. (R = 0.083 bar dm^3 K^{-1} mol^{-1})

Solution

The partial pressure of oxygen is $P = \dfrac{nRT}{V} = \dfrac{8 \times 0.083 \times 300}{32 \times 1} = 0.225\ bar.$

The partial pressure of hydrogen gas is $P' = \dfrac{n'RT}{V'} = \dfrac{2 \times 0.083 \times 300}{1} = 49.8\ bar.$

Total pressure of the mixture$= 6.225 + 49.8 = 56.025\ bar.$

#422987

Topic: Gas laws

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 of mass 100 kg is fille with helium at 1.66 bar at 27 $^\circ C$. (Density of air = 1.2 kg m^3 and R = 0.083 bar dm 3 K^{-1} mol^{-1})

Solution

The volume of the balloon is $V = \dfrac{4}{3}\pi r^3$.

The radius of balloon is 10 m.

Hence, the volume of the balloon is $V = \dfrac{4}{3} \times 3.1416 \times (10)^3 = 4186.7 m^3$

The mass of displaced air is obtained from the product of volume and density. It is $4186.7 \times 1.2 = 5024.04 kg$.

The number of moles of gas present are $n = \dfrac{PV}{RT} = \dfrac{1.666 \times 4186.7 \times 10^3}{0.083 \times 300} = 279.11 \times 10^3.$

Note: Here, the unit of volume is changed from m^3 to dm^3.

$1m^3 = 1000dm^3$.

Mass of helium present is obtained by multiplying the number of moles with molar mass. It is $279.11 \times 10^3 \times 4 = 1116.44 \times 10^3 g = 1116.4$kg.

The mass of filled balloon is the sum of the mass of the empty ballon and the mass of He. It is $100 + 1116.4 = 1216.4$kg.

Pay load $=$ mass of displaced air $-$ mass of balloon $= 5024.04 - 1216.44 = 3807.6$kg

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#422992

Topic: Gas laws

Calculate the volume occupied by 8.8 g of CO$_2$ at 31.1$^\circ C$ and 1 bar pressure. (R = 0.083 bar L K^{-1} mol^{-1})

Solution

The number of moles is $n = \dfrac{8.8}{44} = 0.2$

The volume occupied is $V = \dfrac{nRT}{P}$.

$V = \dfrac{0.2 \times 0.083 \times 304.1}{1} = 5.05L$

#422994

Topic: Gas laws

2.9 g of gas at $95^\circ C$ occupied the same volume as 0.184 g of dihydrogen at $17^\circ C$ at the same pressure. What is the molar mass of the gas?

Solution

Let M be the molar mass of the gas.

The number of moles of the gas, $n = \dfrac{2.9}{M}$.

The volume of the gas is $V = \dfrac{nRT}{P}$.

$V = \dfrac{2.9R \times 368}{MP}$

The number of moles of dihydrogen are $n' = \dfrac{0.184}{2} = 0.092$

The volume of dihydrogen gas is $= \dfrac{0.092 \times R \times 290}{P}$.

At same pressure P, the volume of gas is equal to volume of dihydrogen.

$0.092R \times 290 = \dfrac{2.9R \times 368}{M}$

$M = \dfrac{2.9 \times 368}{0.092 \times 290} = 40$

Thus, the molar mass of the gas is $40g/mol$.

#422996

Topic: Gas laws

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

Solution

In 100 g of mixture, the mass of hydrogen and oxygen will be 20 and 80 g respectively.

The number of moles of hydrogen are $n = \dfrac{20}{2} = 10$

The number of moles of oxygen are $n' = \dfrac{80}{32} = 2.5$

The mole fraction of hydrogen is $\dfrac{10}{10 + 2.5} = 0.8$

Partial pressure of hydrogen is $0.8 \times P = 0.8 \times 1 = 0.8\ bar$

#423000

Topic: Gas laws

In terms of Charles' law, explain why - 273 $^\circ C$ is the lowest possible temperature.

Solution

At $0\ K$ or -273^oC, the volume of the gas becomes equal to zero . The gas ceases to exists. Hence, -273^oC is the lowest possible temperature.

#423001

Topic: Liquefaction of gases and critical temperature

Critical temperature for carbon dioxide and methane are $31.1^\circ C$ and $-81.9^\circ C$ respectively. Which of these has stronger intermolecular forces and why?

Solution

Higher is the critical temperature, more easily the gas can be liquefied and greater are the intermolecular forces of attraction. Hence, CO_2 has stronger intermolecular forces than H_2O.

#423002

Topic: Behaviour of real gases - Deviations from ideal behaviour

Explain the physical significance of Van der Waals parameters.

Solution

The van der Waals constant 'a' represents the magnitude of intermolecular forces of attraction and the Van der Waals constant 'b' represents the effective size of the molecules