Q1 M1-142-13

The lowest pressure attainable in the laboratory is $5.0 \times 10^{-18}$ Pa at 20 °C. How many gas molecules are there per $m^3$ at this pressure?

A) $1.2\times10^3$
B) $2.3\times10^3$
C) $4.4\times10^6$
D) $3.1\times10^5$
E) $5.6\times10^{-3}$

$$pV = NkT$$

$$\frac{N}{V} = \frac{P}{kT} = \frac{5\times10^{-18}}{1.38\times10^{-23}\times(20+273)} = 1.2\times10^{3}\ \frac{\text{molecule}}{m^3}$$

Q2 M1-132-15

Two moles of a monatomic ideal gas with an RMS speed of 254 m/s are contained in a tank that has a volume of 0.150 $m^3$. If the molar mass of the gas is 0.390 kg/mole, what is the pressure of the gas?

A) $1.12 \times10^5$ Pa
B) $7.17 \times10^5$ Pa
C) $2.22 \times10^4$ Pa.
D) $3.25 \times10^6$ Pa.
E) $6.87 \times10^4$ Pa.

$$V_{rms} = \sqrt{\frac{3RT}{M}} \Rightarrow T = \frac{M V_{rms}^2}{3R}$$

$$pV = nRT \Rightarrow p = \frac{nRT}{V} = \frac{n\cancel{R}}{V}\frac{M V_{rms}^2}{3\cancel{R}}$$

$$p = \frac{nMV_{rms}^2}{3V} = \frac{2(0.390)(254)^2}{3(0.150)} = 1.12\times10^{5}\ \text{Pa}$$

Q3 M1-122-13

An ideal gas initially at a pressure of 1.2 atm and temperature 74 °C undergoes an isothermal expansion to twice its original volume. During the expansion, the gas absorbs 20 kJ of heat. Find the number of moles for this gas?

A) 10
B) 12
C) 15
D) 18
E) 20

Isothermal $\Rightarrow \Delta E_{int} = 0 = Q - W$

$$\Rightarrow Q = W = nRT \ln\frac{V_f}{V_i}$$

$$n = \frac{Q}{RT\ln\frac{V_f}{V_i}} = \frac{20\times10^{3}}{(8.31)(74+273)\ln 2} = 10 \text{ moles}$$

Q4 M1-132-13

When an amount of heat of 35.1 J was added to a particular ideal gas, the volume of the gas changed from 50.0 $cm^3$ to 100 $cm^3$ while the pressure remained at 1.00 atm. If the quantity of gas present was 2.00 $\times 10^{-3}$ mol, find the value of specific heats $C_V$ and $C_p$(in J/mol.K), respectively.

A) 49.5 and 57.8
B) 57.8 and 49.5
C) 26.1 and 34.4
D) 51.1 and 61.5
E) 29.5 and 37.8

$Q = nC_p \Delta T$ constant

$pV = nRT \Rightarrow p\Delta V = nR\Delta T$

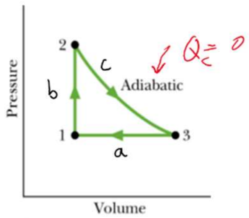
$C_p = \frac{Q}{n\Delta T} = \frac{Q}{\frac{p\Delta V}{R}} = \frac{RQ}{p\Delta V} = \frac{(8.31)(35.1)}{(1.01\times10^5)(100-50)\times10^{-6}}$

$C_p = 57.8 \frac{J}{mol\cdot K}$ $C_p = C_V + R \Rightarrow C_V = 49.4 \frac{J}{mol\cdot K}$

Q5 M1-122-09

The figure shows a cycle undergone by 1.0 mole of an ideal diatomic gas. The temperatures are $T_1$ = 400 K, $T_2$ = 700 K, and $T_3$ = 555 K. Calculate the net work done in one cycle.

A) 1.7 kJ by the gas
B) 1.7 kJ on the gas
C) 3.8 kJ on the gas
D) 3.8 kJ by the gas
E) 0.52 kJ by the gas




$\Delta E_{int} = Q - W$

$0 = Q - W$ (cycle)

$W = Q = Q_a + Q_b + Q_c$

$W = nC_p\Delta T_{31} + nC_V\Delta T_{12} + 0$

$W = \frac{7}{2}R(T_1 - T_3) + \frac{5}{2}R(T_2 - T_1)$

$W = 8.31\left[\frac{7}{2}(400-555) + \frac{5}{2}(700-400)\right] = +1.7$ kJ

by

Q6 M1-122-14

An ideal gas with a volume $V_0$ and a pressure $P_0$ undergoes a free expansion to volume $V_1$ and pressure $P_1$ where $V_1 = 32V_0$. The gas is then compressed adiabatically to the original volume $V_0$ and pressure $4P_0$. The ratio of specific heats, $\gamma$, of the ideal gas is:

A) 7/5
B) 2/5
C) 3/5
D) 1/5
E) 9/5

free expansion $\Delta E_{int} = 0 \Rightarrow T = \text{constant}$

$pV = nRT = \text{constant}$

$p_1V_1 = P_0V_0 \Rightarrow p_1 = \frac{P_0V_0}{V_1} = \frac{P_0V_0}{32V_0} = \frac{P_0}{32}$

Adiabatic expansion

$pV^\gamma = \text{constant}$

$P_1V_1^\gamma = P_2V_2^\gamma$ $P_2 = 4P_0$ $V_2 = V_0$

$\frac{P_0}{32}(32V_0)^\gamma = 4P_0V_0^\gamma$

$\frac{32^\gamma}{32} = 4 \Rightarrow 32^\gamma = 128 \Rightarrow \gamma \ln 32 = \ln 128$

$\gamma = \frac{\ln 128}{\ln 32} = 1.4 = \frac{7}{5}$