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Problem 25.25 (RHK)

In an experiment, 1.35 mol *of oxygen* $\left(\mathrm{O_2}\right)$ *are heated at constant pressure starting at* $11.0^{0}\mathrm{C}$. *We have to calculate the amount of heat that must be added to the gas to double its volume.*

Solution:

Let the initial volume of the gas be $V\ \mathrm{m}^3$ and its pressure be p Pa. The initial temperature of the gas is

$T_i = \left(273.16 + 11.0\right)\ \mathrm{K} = 284.16\ \mathrm{K}.$

As the amount of oxygen gas is 1.35 mol, using the ideal gas equation of state we find

$$pV = 1.35 \times R \times T_i = 1.35 \times 8.315 \times 284.16\ \mathrm{J}$$
$$= 3{,}189.7\ \mathrm{J}.$$

As the volume of the gas is doubled at constant pressure, the final temperature of the gas will be

$T_f = 2 \times T_i = 568.3\ \mathrm{K}.$

For a diatomic gas the internal energy of n mol at temperature T is

$$E = \frac{5}{2} nRT,$$

therefore, the change in the internal energy of the gas on heating will be

$$\Delta E = \frac{5}{2} \times 1.35 \times 8.315 \times (568.32 - 284.16) \text{ J}$$
$$= 7974 \text{ J}.$$

We next calculate the work done on the gas during its expansion of its volume from V to $2V$. It is given by

$$W = -\int_{V}^{2V} pdV = -pV = -3189.7 \text{ J}.$$

The fist law of thermodynamics states that

$$\Delta E = Q + W,$$

where Q is the heat energy supplied to the system during its change of state.

Therefore,

$$Q = \Delta E - W = (7914 + 3189.7) \text{ J=11,163 J}.$$