In-class problem linked to lecture pages 20-29:

One mole of an ideal gas is contained in a volume of $0.01\ m^3$ and held at a constant temperature of 27°C. The ideal gas law is $pV=\mu RT$, where μ is given in moles and $R = 8.31\ J/(mole\text{-}K)$.

(a) Locate this state on a graph of pressure versus volume.

(b) How much heat must be transferred to the gas if it is to expand to twice its initial volume while its temperature is kept constant?

Physics 301

In-class problem linked to lecture pages 20-29:

(a) $PV = \mu RT$

$\mu = 1$ mole

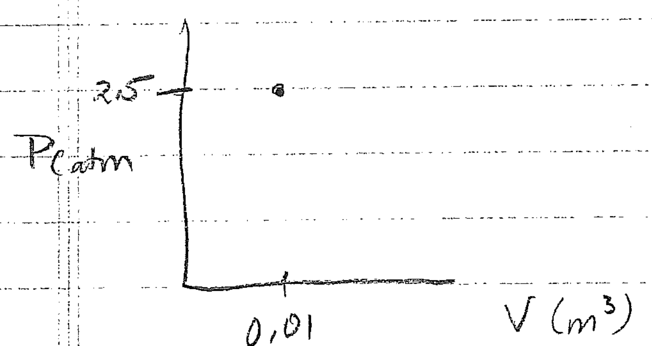
R = universal gas constant $= 8.31 \dfrac{J}{g\text{-mole-}^\circ K}$

$T = 27^\circ C = 300.15 K$

$V = 10^{-2}\ m^3$

So $P = \dfrac{(1\ \text{mole})\left(8.31 \dfrac{J}{\text{mole}\ K}\right)(300.15\ K)}{10^{-2}\ m^3} = ?$

$= 2.5 \times 10^5 \dfrac{J\text{-}m^3}{g} = 2.5 \times 10^5 \dfrac{N}{m^2} = 2.5$ atm



(b) $\Delta E = \Delta Q - \Delta W + \mu \Delta N$

$\Delta N = 0$ (no particle transfer)

$\Delta E = 0$ (no change in internal energy, because T=const.

So $\Delta Q = \Delta W$

$\Delta W = P \Delta V$

$dW = P dV = \mu \frac{RT \, dV}{V}$

$$\Delta W = \mu RT \int_{V_1}^{V_2} \frac{dV}{V} = \mu RT \ln\left(\frac{V_2}{V_1}\right) = \mu RT \ln 2$$

$$= 1 (8.31)(300.15) \ln 2 = 1.73 \times 10^3 \text{ J}$$

Thus 1.73×10^3 J of heat must be added to the system