64 g of each O_2 and CH_4 are mixed in a container of volume 41 dm^3 at 25 °C.

If total pressure of gaseous mixture is 2721 torr, find partial pressure of O_2 and CH_4.

Solution: We will assume that O_2 and CH_4 behave as ideal gases. Then, according to the Dalton's law, partial pressure of gaseous components in the ideal gas mixture can be calculated as:

$p_i = P_{\Sigma} \cdot x_i$, where the p_i, x_i – partial pressure and mole fraction of individual gas component in a gas mixture, respectively; P_{Σ} – total pressure of the gas mixture;

$x_i = \frac{n_i}{\sum n_i} = \frac{m_i / M_i}{\sum \frac{m_i}{M_i}}$; where n_i is the substance amount of individual gas component, moles;

m_i, M_i are the mass and molar weight of individual gas component, grams and g/mole, respectively;

M_{CH4} = 12 + 4 = 16 g/mole; M_{O2} = 16·2 = 32 g/mole;

Then, $x_{O_2} = \frac{64/32}{\frac{64}{32} + \frac{64}{16}} = 0.333$; $x_{CH_4} = \frac{64/16}{\frac{64}{32} + \frac{64}{16}} = 0.667$;

$p_{O_2} = 2721 \cdot 0.333 = 906$ torr; $p_{CH_4} = 2721 \cdot 0.667 = 1815$ torr;

Answer: Partial pressures of O_2 and CH_4 are equal to 906 torr and 1815 torr, respectively.