A nitrogen gas from a 24.0 L container with the pressure of 2atm and an oxygen gas from a 12.0 L with a pressure of 2atm and 273K were mixed together in a 10.0 L container. what are the partial pressure exerted by each gas in the mixture and what is the total pressure? use ideal gas formula and mole fraction.

Solution:

for N_2: $P_{N_2} V_{N_2} = n_{N_2} RT \Rightarrow$

$$n = \frac{2\,atm \times 24.0\,L}{0{,}082057\left(\frac{l \cdot atm}{moles \cdot K}\right) \times 273K} = 2.143 \text{ moles};$$

for O_2: $$n = \frac{2\,atm \times 12.0\,L}{0{,}082057\left(\frac{l \cdot atm}{moles \cdot K}\right) \cdot 273K} = 1{,}071 \text{ moles}.$$

A mixture of ideal gases is formed after mixing of N_2 and O_2:

the partial pressure:

$$pi(N_2) = \frac{2{,}143\,moles \cdot 0{,}082057\,(l \cdot atm / moles \cdot K) \times 273K}{10\,L} = 4{,}801\,atm$$

$$pi(O_2) = \frac{1{,}071\,moles \cdot 0{,}082057\,(l \cdot atm / moles \cdot K) \times 273K}{10.0\,L} = 2.399\,atm$$

Due to Dalton's Law:

$P_{total} = pi(N_2) + pi(O_2) = 7{,}2\,atm$

Answer: $pi(N_2) = 4.801\,atm$

$pi(O_2) = 2.399\,atm$

$P_{total} = 7{,}200\,atm$