Decompose 1.1 g of H$_2$O$_2$ in a sealed flask with a volume of 2.50 L. What is the final pressure of O$_2$ and H$_2$O at 25 °C?

Solution

Strategy:

• Calculate the # of moles of H$_2$O$_2$ consumed and then # of moles of O$_2$ and H$_2$O produced.
• Then calculate P from n, R, T, and V.

P of O$_2$ = \( \frac{nRT}{V} \)

= \( \frac{(0.016 \text{ mol})(0.0821 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(298 \text{ K})}{2.50 \text{ L}} \)

P of O$_2$ = 0.16 atm

P of H$_2$O = 0.32 atm
Dalton’s Law of Partial Pressures

\[ 2 \text{H}_2\text{O}_2(\text{liq}) \rightarrow 2 \text{H}_2\text{O}(g) + \text{O}_2(g) \]
\[ \begin{align*}
0.32 \text{ atm} & \quad 0.16 \text{ atm}
\end{align*} \]

What is the total pressure in the flask?

\[ P_{\text{total}} \text{ in gas mixture} = P_A + P_B + \ldots \]
where there are separate components A, B, ...

Therefore, \[ P_{\text{total}} = P(\text{H}_2\text{O}) + P(\text{O}_2) = 0.48 \text{ atm} \]

Dalton’s Law:
Total P is the sum of PARTIAL pressures.

Dalton’s Law of Partial Pressures

Problem 12.34: Octane burns to give \( \text{CO}_2(g) \) and \( \text{H}_2\text{O}(g) \). A 0.095 g sample of octane burns in just enough \( \text{O}_2 \) for complete reaction. What is the pressure of \( \text{O}_2 \) in a 4.75 L flask at 22 °C flask before reaction?

\[ 2 \text{C}_8\text{H}_{18}(g) + 25 \text{O}_2(g) \rightarrow 18 \text{H}_2\text{O}(g) + 16 \text{CO}_2(g) \]

Solution:
1) compute the number of moles of octane
2) compute the number of moles of oxygen
3) use \( P = \frac{nRT}{V} \)

Dalton’s Law of Partial Pressures

Problem 12.34: Octane burns to give \( \text{CO}_2(g) \) and \( \text{H}_2\text{O}(g) \). A 0.095 g sample of octane burns in just enough \( \text{O}_2 \) for complete reaction. What is the total pressure in the 4.75 L flask after reaction?

\[ 2 \text{C}_8\text{H}_{18}(g) + 25 \text{O}_2(g) \rightarrow 18 \text{H}_2\text{O}(g) + 16 \text{CO}_2(g) \]

Solution:
1) compute the number of moles of octane
2) compute the number of moles of water
3) use \( P = \frac{nRT}{V} \)

Dalton’s Law of Partial Pressures

Problem 12.34: Octane burns to give \( \text{CO}_2(g) \) and \( \text{H}_2\text{O}(g) \). A 0.095 g sample of octane burns in just enough \( \text{O}_2 \) for complete reaction. What is the total pressure in the 4.75 L flask after reaction?

\[ 2 \text{C}_8\text{H}_{18}(g) + 25 \text{O}_2(g) \rightarrow 18 \text{H}_2\text{O}(g) + 16 \text{CO}_2(g) \]

Solution:
1) compute the number of moles of octane
2) compute the number of moles of oxygen
3) use \( P = \frac{nRT}{V} \) for both
4) add the two partial pressures
Dalton’s Law of Partial Pressures

Problem 12.34: Octane burns to give CO\textsubscript{2}(g) and H\textsubscript{2}O(g). A 0.095 g sample of octane burns in just enough O\textsubscript{2} for complete reaction. What is the total pressure in the 4.75 L flask after reaction?

\[
2 \text{C}_8\text{H}_{18}(g) + 25 \text{O}_2(g) \rightarrow 18 \text{H}_2\text{O}(g) + 16 \text{CO}_2(g)
\]

Solution:
1) compute the number of moles of water
2) compute the number of moles of CO\textsubscript{2}
3) use \(P = nRT/V\) for both
4) add the two partial pressures

GAS DENSITY

\[
PV = nRT
\]

mass & m = n\cdot M \text{ where } M \text{ is molar mass}

so \(PV = (m/M)RT\)

and thus density is

\[
d = \frac{m}{V} = \frac{P \cdot M}{R \cdot T}
\]

40 miles above the earth’s surface \(T = 250 \text{ K}\) and \(P = 0.2 \text{ mm Hg}\). What is the density of the air there? (Assume M ~ 29 g/mol)

Solution:
1) convert P to units of atm
2) plug into the formula

The density of air at 15 °C and 1.00 atm is 1.23 g/L. What is the average molar mass of the gases that make up air?

1. Calculate the average # of moles in 1 L of air:
   \[
   V = 1.00 \text{ L}
   \]
   \[
   P = 1.00 \text{ atm}
   \]
   \[
   T = 288 \text{ K}
   \]
   \[
   n = \frac{PV}{RT} = 0.0423 \text{ mol}
   \]

2. Calculate the average molar mass:
   \[
   \text{mass/mol} = \frac{1.23 \text{ g}}{0.0423 \text{ mol}} = 29.1 \text{ g/mol}
   \]