Gases and Stoichiometry

\[ 2 \text{H}_2\text{O}_2(\text{liq}) \rightarrow 2 \text{H}_2\text{O}(g) + \text{O}_2(g) \]

Decompose 1.1 g of \( \text{H}_2\text{O}_2 \) in a sealed flask with a volume of 2.50 L. What is the final pressure of \( \text{O}_2 \) and \( \text{H}_2\text{O} \) at 25 °C?

Solution

Strategy:

• Calculate the \# of moles of \( \text{H}_2\text{O}_2 \) consumed and then \# of moles of \( \text{O}_2 \) and \( \text{H}_2\text{O} \) produced.

• Then calculate \( P \) from \( n, R, T, \) and \( V \).

# mol \( \text{H}_2\text{O}_2 \) = \( \frac{1.1 \text{ g}}{34.0 \text{ g/mol}} \) = 0.032 mol

# mol of \( \text{O}_2 \) = \( \frac{\# \text{ mol} \text{H}_2\text{O}_2}{2} \) = 0.016 mol

\[ P = \frac{nRT}{V} \, \text{so} \]
Gases and Stoichiometry

2 H₂O₂(liq) → 2 H₂O(g) + O₂(g)

P of O₂ = nRT/V
= (0.016 mol)(0.0821 L·atm/K·mol)(298 K)
  2.50 L

P of O₂ = 0.16 atm

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Gases and Stoichiometry

2 H₂O₂(liq) → 2 H₂O(g) + O₂(g)

What is the P of H₂O? We could calculate as above. But consider the ideal gas law,
PV = nRT.
P = kn at a fixed T and V {from P = (RT/V)n}.

There are 2 times as many moles of H₂O as moles of O₂. P is proportional to n. Therefore, P of H₂O is twice that of O₂.

P of O₂ = 0.16 atm so P of H₂O = 0.32 atm

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Dalton’s Law of Partial Pressures

$$2 \text{H}_2\text{O}_2(\text{liq}) \rightarrow 2 \text{H}_2\text{O}(g) + \text{O}_2(g)$$

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<td>0.32 atm</td>
<td>0.16 atm</td>
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What is the total pressure in the flask?

$$P_{\text{total}} \text{ in gas mixture} = P_A + P_B + ...$$

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.

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Dalton’s Law of Partial Pressures

Problem 12.34: Octane burns to give CO$_2$(g) and H$_2$O(g). A 0.095 g sample of octane burns in just enough O$_2$ for complete reaction. What is the pressure of 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 $$P = \frac{nRT}{V}$$

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Dalton’s Law of Partial Pressures

Problem 12.34: Octane burns to give CO$_2$(g) and H$_2$O(g). A 0.095 g sample of octane burns in just enough O$_2$ for complete reaction. What is the pressure of H$_2$O(g) in a 4.75 L flask at 30 °C 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 = nRT/V

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Dalton’s Law of Partial Pressures

Problem 12.34: Octane burns to give CO$_2$(g) and H$_2$O(g). A 0.095 g sample of octane burns in just enough O$_2$ for complete reaction. What is the total pressure in the 4.75 L 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 = nRT/V for both
4) add the two partial pressures

Apr. 26, 2006
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 = \(m = n \cdot M\) where \(M\) is molar mass

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

and thus density is

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

Knowing density, P & T gives molar mass

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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 \( M \sim 29 \text{ g/mol} \))

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

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**GAS DENSITY**

The density of air at 15 \( ^\circ \text{C} \) 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 number 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:
   \[ \text{mass/mol} = \frac{1.23 \text{ g}}{0.0423 \text{ mol}} = 29.1 \text{ g/mol} \]