2. Answer the following questions related to hydrocarbons.

(a) Determine the empirical formula of a hydrocarbon that contains 85.7 percent carbon by mass.

\[
\begin{align*}
n_C &= \frac{85.7 \text{ g C}}{12.01 \text{ g C}} \times \frac{1 \text{ mol C}}{7.14 \text{ mol C}} = 7.14 \text{ mol C} \\
n_H &= \frac{14.3 \text{ g H}}{1.008 \text{ g H}} \times \frac{1 \text{ mol H}}{14.2 \text{ mol H}} = 14.2 \text{ mol H} \\
\frac{7.14 \text{ mol C}}{7.14} : \frac{14.2 \text{ mol H}}{7.14} \\
1 \text{ mol C} : 1.99 \text{ mol H}
\end{align*}
\]

The empirical formula is \( \text{CH}_2 \)

1 point for moles of \( C \) and moles of \( H \)

1 point for ratio of moles of \( C \) to moles of \( H \)

1 point for correct formula

(b) The density of the hydrocarbon in part (a) is 2.0 g L\(^{-1}\) at 50\(^\circ\)C and 0.948 atm.

(i) Calculate the molar mass of the hydrocarbon.

\[
\begin{align*}
PV &= nRT \\
\text{molar mass} &= \frac{\text{mass}}{\text{molar mass}} \times \frac{RT}{P} = \text{density} \times \frac{RT}{P} \\
\text{molar mass} &= 2.0 \text{ g L}^{-1} \times \frac{0.0821 \text{ L atm mol K}}{0.948 \text{ atm}} \\
\text{molar mass} &= 56 \text{ g mol}^{-1}
\end{align*}
\]

1 point for correct substitution

1 point for the answer

(ii) Determine the molecular formula of the hydrocarbon.

\[
\begin{align*}
\text{empirical mass} \times n &= \text{molar mass} \\
\text{empirical mass for CH}_2 &= 14 \text{ g mol}^{-1} \\
14 \text{ g mol}^{-1} \times n &= 56 \text{ g mol}^{-1} \\
n &= 4 \\
The \text{ molecular formula is C}_4\text{H}_8.
\end{align*}
\]

1 point for correct formula
Question 2 (cont’d.)

(c) Two flasks are connected by a stopcock as shown below. The 5.0 L flask contains CH₄ at a pressure of 3.0 atm, and the 1.0 L flask contains C₂H₆ at a pressure of 0.55 atm. Calculate the total pressure of the system after the stopcock is opened. Assume that the temperature remains constant.

\[
\begin{align*}
\text{CH}_4 & \quad \text{C}_2\text{H}_6 \\
5.0 \text{ L} & \quad 1.0 \text{ L}
\end{align*}
\]

\[
P_f \text{ of CH}_4 = \frac{P \cdot V}{V_f} = \frac{(3.0 \text{ atm})(5.0 \text{ L})}{6.0 \text{ L}} = 2.5 \text{ atm CH}_4
\]

1 point for final pressure of CH₄ or C₂H₆

\[
P_f \text{ of C}_2\text{H}_6 = \frac{P \cdot V}{V_f} = \frac{(0.55 \text{ atm})(1.0 \text{ L})}{6.0 \text{ L}} = 0.092 \text{ atm C}_2\text{H}_6
\]

1 point for the total pressure

\[
P_T = P_f \text{ CH}_4 + P_f \text{ C}_2\text{H}_6 = 2.5 \text{ atm} + 0.092 \text{ atm} = 2.6 \text{ atm}
\]

(d) Octane, C₈H₁₈(l), has a density of 0.703 g mL⁻¹ at 20°C. A 255 mL sample of C₈H₁₈(l) measured at 20°C reacts completely with excess oxygen as represented by the equation below.

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

Calculate the total number of moles of gaseous products formed.

\[
n_{\text{products}} = 255 \text{ mL C}_8\text{H}_{18} \times \frac{0.703 \text{ g C}_8\text{H}_{18}}{1 \text{ mL C}_8\text{H}_{18}} \times \frac{1 \text{ mol C}_8\text{H}_{18}}{114 \text{ g C}_8\text{H}_{18}} = 34 \text{ mol products}
\]

\[
\frac{34 \text{ mol products}}{2 \text{ mol C}_8\text{H}_{18}} = 17 \text{ mol products}
\]

1 point for substitution of any of these conversion factors

1 point for the correct answer