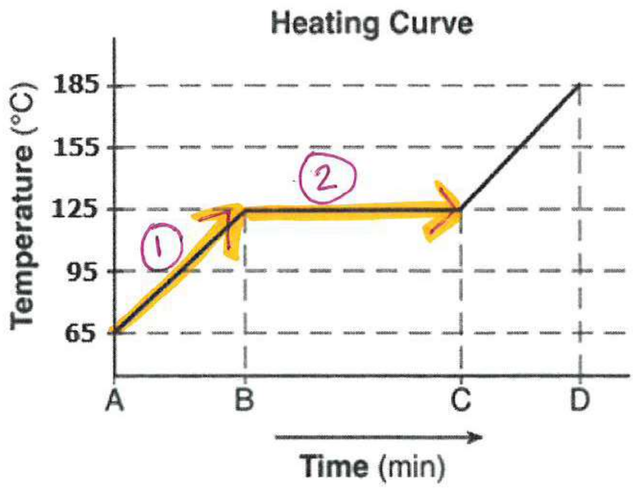
Unit 4: AP Quiz Free Response Practice (10 points)

1. Octane, C_8H_{18}, is a hydrocarbon and an alkane that is a common component in gasoline. A 1.00 mole sample of octane is a liquid at 65°C. The sample is heated uniformly to 185°C. The heating curve for the sample at standard pressure is shown below.



a. Determine the boiling point of octane at standard pressure. (1 point)

b. Calculate the amount of heat needed to completely vaporize 1.00 mol of the sample of octane originally at 65°C. The molar heat capacity of the substance in the liquid phase is 255 $\boxed{J/(mol\ °C)}$, and the heat of vaporization of the substance is 41.0 kJ/mol. (2 points)

$\hookrightarrow nC\Delta T$

A second 1.00 mol sample of liquid octane is combusted with excess oxygen gas in a bomb calorimeter.

c. Write a balanced equation for the complete combustion of octane liquid ~~gas~~, which yields $CO_2(g)$ and $H_2O(l)$. (1 point)

d. Using the heat of formation data in the table below, calculate the ΔH° for the reaction in part (c). (2 points)

Substance	$\Delta H_f°$ (kJ/mol)
$C_8H_{18}(l)$	-250.0
$CO_2(g)$	-393.5
$H_2O(l)$	-285.3

e. Is the amount of heat required to completely vaporize 1.00 mol of liquid octane originally at 65°C greater than, less than, or equal to the amount of heat released in the combustion of 1.00 mole of liquid octane? Justify your answer. (2 points)

f. All the heat evolved in the combustion of 1.00 mole of liquid octane is transferred to a sample of liquid water. The temperature of the water increases by 72°C. What was the mass of the sample of water in kilograms? (2 points)

(a.) 125°C (1pt)

(b.) 1st, heat $C_8H_{18}(l)$: 65°C → 125°C ⇒ $q = nC\Delta T$

$q = (1.00\ mol)(255\ \frac{J}{mol°C})(125°C - 65°C) = 15{,}000\ J = 15\ kJ$ (1pt)

2nd, vaporize C_8H_{18} $(l \to g)$ ⇒ $q = n\Delta H_{vap}$

$q = (1.00\ mol)(41.0\ kJ/mol) = 41.0\ kJ$ (1pt)

$\Rightarrow q_{total} = 15\ kJ + 41.0\ kJ = \boxed{56\ kJ}$

(c) $2\,C_8H_{18(l)} + 25\,O_{2(g)} \rightarrow 16\,CO_{2(g)} + 18\,H_2O_{(l)}$ (1 pt)

(d) $\Delta H^\circ_{comb} = \Sigma \Delta H_f^\circ(pr) - \Sigma \Delta H_f^\circ(re)$

$= [16 \cdot CO_2 + 18 \cdot H_2O] - [2 \cdot C_8H_{18} + 25 \cdot O_2]$

$= [16(-393.5) + 18(-285.3)] - [2(-250.0) + 25 \cdot \emptyset]$ ← 1 pt

$= -11{,}431.4 + 500.0$

$= \boxed{-10{,}931.4 \text{ kJ/mol}}$ ← 1 pt

(e) $1.00 \text{ mol } C_8H_{18} \times \dfrac{1 \text{ mol}_{rxn}}{2 \text{ mol } C_8H_{18}} \times \dfrac{-10{,}931.4 \text{ kJ}}{1 \text{ mol}_{rxn}} = \underbrace{-5{,}470 \text{ kJ}}_{\text{heat released in combustion}}$] 1 pt

$\overbrace{\underbrace{56 \text{ kJ (part b)}}_{\text{heat needed to vaporize 1.00 mol}} < 5{,}470 \text{ kJ}}^{1\text{ pt}}$

@ 65°C is definitely less than energy released in combustion!

(f) $q_{r_{comb}} = -q_{r_{H_2O}} = 5{,}470 \text{ kJ} = m(\text{in kg!}) \times 4.18 \text{ J/g°C} \times 72°C$

$\Rightarrow m = \dfrac{(5{,}470 \text{ kJ})}{(4.18 \text{ J/g°C})(72°C)} = \boxed{18 \text{ kg}}$