1) Well, we are given a fairly simple reaction for which we need to get the heat. Using the data we have in the problem, we can go this way:

\[ \text{C}_6\text{H}_4(\text{OH})_2(\text{aq}) \rightarrow \text{C}_6\text{H}_4\text{O}_2(\text{aq}) + \text{H}_2(\text{g}) \quad \Delta H = +177.4 \text{ kJ} \]
\[ \text{H}_2\text{O}_2(\text{aq}) \rightarrow \text{H}_2(\text{g}) + \text{O}_2(\text{g}) \quad \Delta H = +191.2 \text{ kJ} \quad \text{(reversing = sign change)} \]
\[ 2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{g}) \quad \Delta H = -483.6 \text{ kJ} \quad \text{(don’t forget the 2!)} \]
\[ 2\text{H}_2\text{O}(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{l}) \quad \Delta H = -87.6 \text{ kJ} \quad \text{(don’t forget the 2!)} \]

\[ \text{C}_6\text{H}_4(\text{OH})_2(\text{aq}) + \text{H}_2\text{O}_2(\text{aq}) \rightarrow \text{C}_6\text{H}_4\text{O}_2(\text{aq}) + 2\text{H}_2\text{O}(\text{g}) \quad \Delta H = -202.6 \text{ kJ} \]

2) We need to collect some data, and then get the individual combustion heats. The ideal gas law will give us our moles, and we can go from there.

\[ \Delta H^\circ_f(\text{CH}_4) = -74.87 \text{ kJ/mole}; \quad \Delta H^\circ_f(\text{C}_2\text{H}_6) = -84.667; \quad \Delta H^\circ_f(\text{C}_3\text{H}_8) = -103.85 \text{ kJ} \]
\[ \Delta H^\circ_f(\text{CO}_2) = -393.51 \text{ kJ/mole}; \quad \Delta H^\circ_f(\text{H}_2\text{O}(\text{l})) = -285.840 \text{ kJ/mole}; \quad \Delta H^\circ_f(\text{H}_2\text{O}(\text{g})) = -241.826 \text{ kJ/mole.} \] (Below, all bracketed values are for use of water as a vapor=gas.)

The individual combustion equations, with heats, are: (all components except water are gases):

\[ \text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O} \quad \Delta H = -890.32 \text{ kJ} \quad [\Delta H = -802.28 \text{ kJ}] \]
\[ \text{C}_2\text{H}_6 + \frac{7}{2}\text{O}_2 \rightarrow 2\text{CO}_2 + 3\text{H}_2\text{O} \quad \Delta H = -1559.873 \text{ kJ} \quad [\Delta H = -1427.81 \text{ kJ}] \]
\[ \text{C}_3\text{H}_8 + 5\text{O}_2 \rightarrow 3\text{CO}_2 + 4\text{H}_2\text{O} \quad \Delta H = -2220.04 \text{ kJ} \quad [\Delta H = -2043.95 \text{ kJ}] \]

From the ideal gas law, given our data, we have: \[ n = \frac{PV}{RT} = \]

\[ \frac{\left(\frac{744 \text{ mmHg}}{760 \text{ mmHg/atm}}\right)\left(215 \text{L}\right)}{\left(0.08206 \text{L \cdot atm / K \cdot mol}\right)\left(297.65 \text{K}\right)} = 8.617 \text{ moles.} \]

Given the total number of moles, and the mole fractions of each of our components, we can start to get our heat value. We have 7.07 moles \text{CH}_4; \ 0.879 \text{ moles C}_2\text{H}_6; \text{ and 0.672 moles C}_3\text{H}_8, \text{ so our } \Delta H_{\text{reaction}} = (7.07 \text{ mol})(-890.32 \text{ kJ/mol}) + (0.879 \text{ mole})(-1559.873 \text{ kJ/mol}) + (0.672 \text{ mole})(-2220.04 \text{ kJ/mol}) = -9158 \text{ kJ} \quad [\Delta H_{\text{reaction}} = -8301 \text{ kJ}] \]

3) We are given the balanced equation, and asked to get \(\Delta H^\circ\). Let’s do it.
\[
2\text{K} + 2\text{H}_2\text{O}(l) \rightarrow 2\text{K}^+ + 2\text{OH}^- + \text{H}_2
\]
\[
\begin{array}{c|c|c|c|c|c}
\text{moles} & -285.84 & -252.4 & -230.0 & 0.0 & \text{(kJ/mol)} \\
\end{array}
\]
\[\Delta H^\circ = 2(-252.4 + -230.0) - 2(-285.8) = -393.2 \text{ kJ}\] but this is for 2 moles K, so on a kJ/mol potassium basis, we are looking at \(-196.6 \text{ kJ}\).

Since we only have 5.00 g of K, we only have \[\frac{5g}{39.0983 \text{ g/mol}} = 0.128 \text{ moles},\] which leads to -25.16 kJ of heat being available. All this heat is being used to heat the 1.00 kg of water from 24.0°C to some higher final temperature. We can use \[q = \text{C} \times \text{m} \times \Delta T\] to get there.

25160 J = (4.184 J/g°C)(1000 g)(\Delta T), which leads to \[\Delta T = 6.01 °\text{C},\] and a final temperature of 24.0 + 6.01 = 30.01 °C.

4) We have, as given:
\[4 \text{ KO}_2(s) + 2 \text{ H}_2\text{O}(l) \rightarrow 4 \text{ KOH}(s) + 3 \text{ O}_2(g)\]
\[\begin{array}{c|c|c|c|c}
\text{M}_\text{m} & 71.097 & 18.02 & 56.105 & 31.9998 \\
\end{array}\]
and are told to assume that this can be modeled by:
\[\text{C}_6\text{H}_12\text{O}_6 + 6 \text{ O}_2 \rightarrow 6 \text{ CO}_2 + 6 \text{ H}_2\text{O}(l)\]
\[\Delta H = -2803.04 \text{ kJ/mol Glucose}\]
\[\Delta H^\circ_f: -1273 \quad 0.0 \quad 6(-393.5) \quad 6(-285.840)\]
We need 150 kcal of energy, which equals 627.6 kJ of energy so:
\[? \text{ moles Glucose} = \frac{627.6 \text{ kJ}}{(1 \text{ mole glucose)/2803.04 kJ})} = 0.224 \text{ moles glucose which requires } 6(0.224 \text{ moles}) = 1.344 \text{ moles oxygen. For every } 3 \text{ moles of oxygen produced, } 4 \text{ moles of KO}_2 \text{ are consumed, so we need } (4/3)(1.344 \text{ moles}) = 1.792 \text{ moles KO}_2 \text{ or } 127.4 \text{ g.}\]

5) Given that a negative \(\Delta H\) means an exothermic reaction (heat given off), it seems to me that this also means that heat \textbf{is being produced}. Therefore, if heat is being produced, it is, in many ways, just like a \textbf{product} of the reaction. In fact, next semester when we deal with equilibrium, we will treat \[A + B \rightarrow C + D \quad \Delta H < 0\] as if we could write the thermochemical equation in this manner:
\[A + B \rightarrow C + D + \Delta H.\]
(By the way, for an endothermic reaction, we could write:
\[A + B + \Delta H \rightarrow C + D\]
making heat a reactant.)