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:

\[
C_6H_4(OH)_2(aq) \rightarrow C_6H_4O_2(aq) + H_2(g) \quad \Delta H = +177.4 \text{ kJ}
\]

\[
H_2O_2(aq) \rightarrow H_2(g) O_2(g) \quad \Delta H = +191.2 \text{ kJ (reversing = sign change)}
\]

\[
2H_2(g) + O_2(g) \rightarrow 2H_2O(g) \quad \Delta H = -483.6 \text{ kJ (don’t forget the 2!)}
\]

\[
2H_2O(g) \rightarrow 2H_2O(l) \quad \Delta H = -87.6 \text{ kJ (don’t forget the 2!)}
\]

\[
C_6H_4(OH)_2(aq) + H_2O_2(aq) \rightarrow C_6H_4O_2(aq) + 2H_2O(g) \quad \Delta H = -202.6 \text{ kJ}
\]

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.

\[
\Delta H^\circ_f(CH_4) = -74.87 \text{ kJ/mole}; \quad \Delta H^\circ_f(C_2H_6) = -84.667; \quad \Delta H^\circ_f(C_3H_8) = -103.85 \text{ kJ}
\]

\[
\Delta H^\circ_f(CO_2) = -393.51 \text{ kJ/mole}; \quad \Delta H^\circ_f(H_2O(l)) = -285.840 \text{ kJ/mole}; \quad [\Delta H^\circ_f(H_2O(g)) = -241.826 \text{ kJ/mole}.] \quad \text{(Below, all bracketed values are for use of water as a vapor=gas.)}
\]

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

\[
CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O \quad \Delta H = -890.32 \text{ kJ} \quad [\Delta H = -802.28 \text{ kJ}]
\]

\[
C_2H_6 + \frac{7}{2} O_2 \rightarrow 2CO_2 + 3H_2O \quad \Delta H = -1559.873 \text{ kJ} \quad [\Delta H = -1427.81 \text{ kJ}]
\]

\[
C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O \quad \Delta H = -2220.04 \text{ kJ} \quad [\Delta H = -2043.95 \text{ kJ}]
\]

From the ideal gas law, given our data, we have:

\[
\frac{PV}{RT} = \frac{(744 \text{ mmHg})(215L)}{(0.08206 \frac{L \cdot \text{ atm}}{K \cdot \text{ mol}})(297.65K)} = 8.617 \text{ moles.}
\]

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 CH₄; 0.879 moles C₂H₆; and 0.672 moles C₃H₈, 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, plus some other thermochemical data. Let’s start with the balanced equation:
\[ 2\text{H}_2\text{S} + 3\text{O}_2 \rightarrow 2\text{H}_2\text{O}(l) + 2\text{SO}_2 \text{ (all other species gases)} \quad \Delta H^\circ = -1125 \text{ kJ} \]

a) We are asked how much heat would be evolved in producing one liter of SO\(_2\) at 1.00 atm and 25°C (= 298.15 K). From PV = nRT, we can get: moles of SO\(_2\) = n = \[ \frac{(1\text{ atm})(100\text{ L})}{(0.08206\text{ L} \cdot \text{atm} / \text{K} \cdot \text{mol})(298.15\text{ K})} = 4.09 \times 10^{-2} \text{ mole SO}_2 \] Our thermochemical equation tells us that we obtain 1125 kJ for every 2 moles of SO\(_2\), or 562.5 kJ/mol SO\(_2\), so for 0.0409 moles, we would get: \((0.0409 \text{ moles})(562.5 \text{ kJ/mol}) = 23 \text{ kJ}\).

b) We have a molar heat capacity of 30.2 J/(mol°C). We also have a \(\Delta T = 475 \degree C\) and the same number of moles as before \((4.09 \times 10^{-2})\). To perform the heating we are asked to do requires \(q = \text{(molar heat capacity)}(\Delta T)(n) = (30.2 \text{ J/(mol°C)})(475 \degree C)(0.0409 \text{ moles}) = 586.7 \text{ J, which represents } \frac{586.7\text{ J}}{23000\text{ J}}(100) = 2.55\%\).

4) [6.120] We have, as given:

\[ 4 \text{KO}_2(s) + 2 \text{H}_2\text{O}(l) \rightarrow 4 \text{KOH}(s) + 3 \text{O}_2(g) \]

\[ M_m: \quad 71.097 \quad 18.02 \quad 56.105 \quad 31.9998 \]

and are told to assume that this can be modeled by:

\[ \text{C}_6\text{H}_{12}\text{O}_6 + 6 \text{O}_2 \rightarrow 6 \text{CO}_2 + 6 \text{H}_2\text{O}(l) \quad \Delta H = -2803.04 \text{ kJ/mol Glucose} \]

\[ \Delta H^\circ_f: \quad -1273 \quad 0.0 \quad 6(-393.5) \quad 6(-285.840) \]

We need 150 kcal of energy, which equals 627.6 kJ of energy so:

? moles Glucose = 627.6 kJ \left[(1 \text{ mole glucose})/2803.04 \text{ kJ}\right] = 0.224 \text{ moles glucose} which requires \(6(0.224 \text{ moles}) = 1.344 \text{ moles oxygen. For every 3 moles of oxygen produced, 4 moles of KO}_2 \text{ are consumed, so we need (4/3)(1.344 moles) = 1.792 moles KO}_2 \text{ or 127.4 g.}\]