1. Decide whether each of the following processes are endothermic or exothermic:

a. condensing steam into water  
   exothermic
b. burning a candle  
   exothermic
c. melting ice cream  
   endothermic
d. cooling hot coffee  
   exothermic
e. formation of snow flakes  
   exothermic
f. heating iron to form iron (II) oxide (from Reactions lab)  
   endothermic

2. a. How many kJ are represented by $3.44 \times 10^4$ cal of heat?
   
   $3.44 \times 10^4 \text{ cal} \times \frac{4.184 \text{ J}}{1 \text{ cal}} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 144 \text{ kJ}$

b. If a reaction releases 70.8 kJ, how many nutritional calories does it generate?

   $70.8 \text{ kJ} \times \frac{1 \text{ kcal}}{4.184 \text{ kJ}} = 16.9 \text{ kcal or nutritional cal}$

   since $4.184 \text{ J} = 1 \text{ cal}$, $4.184 \text{ kJ} = 1 \text{ kcal}$

   $1 \text{ kcal} = 1 \text{ nutritional cal}$

c. How can you determine the amount of heat exchanged in a reaction?

   You can do the reaction in a container immersed in a known amount of water. If you measure the change in the temperature of the water you can calculate the energy released or absorbed since it takes 1 calorie (or 4.184 J) to raise one gram of water by 1 degree centigrade (or 1 cal absorbed for each 1 degree drop per gram).

3. Use the following equation to answer the questions that follow it:

   $2 \text{ H}_2\text{O (l)} \rightarrow 2 \text{ H}_2 (g) + \text{ O}_2 (g) \quad \Delta H = +571.6 \text{ kJ}$

a. Is this process exothermic or endothermic and why?

   Endothermic since $\Delta H$ is positive which means products contain more energy than the reactants.

b. How many kJ are transferred when 25.0 g of water are decomposed?

   $25.0 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.0 \text{ g H}_2\text{O}} \times \frac{571.6 \text{ kJ}}{2 \text{ mol H}_2\text{O}} = 397 \text{ kJ}$

c. How many g of hydrogen are produced when 775 J of energy are used?

   $775 \text{ J} \times \frac{1 \text{ kJ}}{1000 \text{ J}} \times \frac{2 \text{ mol H}_2}{571.6 \text{ kJ}} \times \frac{2.00 \text{ g H}_2}{1 \text{ mol H}_2} = 5.4 \times 10^{-3} \text{ g H}_2$
d. How many mol of water are decomposed if 450 kJ are used?

\[
450 \text{ kJ} \times \frac{2 \text{ mol H}_2\text{O}}{571.6 \text{ kJ}} = 1.57 \text{ mol H}_2\text{O}
\]

4. Use the following equation to answer the questions that follow it:

\[
\text{CH}_4 (g) + 2 \text{ O}_2 (g) \rightarrow \text{CO}_2 (g) + 2 \text{ H}_2\text{O(l)} \quad \Delta H = -890.4 \text{ kJ}
\]

methane

a. Is this process exothermic or endothermic and why?

Exothermic because \(\Delta H = -\) so the products have less energy than the reactants and energy was released.

b. How many moles of methane are required to transfer \(4.66 \times 10^3 \text{ kJ}\)?

\[
4.66 \times 10^3 \text{ kJ} \times \frac{1 \text{ mol CH}_4}{890.4 \text{ kJ}} = 5.23 \text{ mol CH}_4
\]

c. If you start with 10.0 g of methane and 20 g of oxygen gas, how much energy will be transferred?

\[
\begin{align*}
10.0 \text{ g CH}_4 \times \frac{1 \text{ mol CH}_4}{16.0 \text{ g CH}_4} \times \frac{890.4 \text{ kJ}}{1 \text{ mol CH}_4} &= 557 \text{ kJ} \\
20.0 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{32.0 \text{ g O}_2} \times \frac{890.4 \text{ kJ}}{2 \text{ mol O}_2} &= 2.8 \times 10^2 \text{ kJ}
\end{align*}
\]

The answer is \(2.8 \times 10^2 \text{ kJ}\) because this is a limiting reactant situation and you select the least amount of energy released.
5. Use the graph below to address the questions that follow it:

![Reaction Progress Graph](image)

a. Determine the activation energy for this reaction.
   \[ 580 - 335 \text{ kcal} = 245 \text{ kcal} \]

b. Is the reaction endothermic or exothermic? Explain how you know.
   Exothermic because the final products have less energy than the reactants so energy must have been released.

c. Determine the amount of energy absorbed or released in the process.
   \[ 335 - 280 \text{ kcal} = 55 \text{ kcal} \]

d. On the axes above, sketch the graph of the reaction if a catalyst is added.
   See graph above