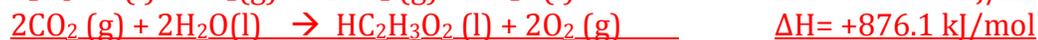
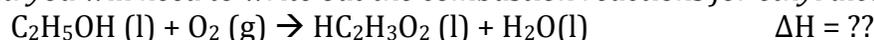


Reaction Energy Review Problems - **KEY**

1. The standard heat of combustion of liquid ethyl alcohol is -1366.8 kJ/mol and that of acetic acid is -876.1 kJ/mol. What is the heat of reaction for the oxidation of ethyl alcohol to acetic acid? (*Hint: you will need to write out the combustion reactions for ethyl alcohol and acetic acid*)



2. A calorimeter containing 1.00 liter of water at 23°C is warmed to 68°C when 5.00 grams of butter is burned. Calculate (a) the total number of joules absorbed by the water, (b) the number of Joules/g given off by the oxidation of butter fat, (c) the kJ/g.

(a) $q = mC_p\Delta T$ (1000 g) (4.184 J/g °C) (45°C) = **188280 J**

(b) $188280\text{J} \div 5 \text{ g} = \mathbf{37656 \text{ J/g}}$

(c) $37656 \text{ J/g} \div 1000 \text{ J/kJ} = \mathbf{37.656 \text{ kJ/g}}$

3. When a red-hot iron is dipped into a beaker of boiling water, 25 g of water is vaporized. How many kilojoules did the iron give up? (Heat of vaporization of water = 40.67 kJ/mol).

$$25 \text{ g} \times \frac{1 \text{ mole}}{18 \text{ g}} \times \frac{40.67 \text{ kJ}}{1 \text{ mole}} = \mathbf{56.486 \text{ kJ}}$$

4. How many joules are required to heat 150 grams of copper from 5°C to 100°C? The same quantity of heat is added to 150 grams of aluminum at 5°C. Which gets hotter, the copper or the aluminum? Specific heat of Al = 0.897 J/g °C; Specific heat of Cu = 0.385 J/g °C.

$$q = mC_p\Delta T$$

For copper: (150 g) (0.385 J/g K) (95) = **5486.25 J**

For aluminum $\Delta T = \frac{q}{mC_p} = \frac{5486.25 \text{ J}}{(150 \text{ g}) (0.897 \text{ J/g K})} = 40.77 \text{ }^\circ\text{C}$

So if both metals start out at 5 °C, **the copper gets hotter.**

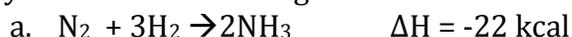
5. The heat of combustion of ethylene, C₂H₄, is -1411.2 kJ/mol. Determine the heat of formation of ethylene using any necessary data from your textbook.



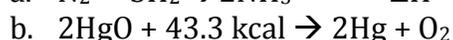
6. Calculate the ΔH_r for the reaction $2\text{NH}_3 (\text{g}) + \text{H}_2(\text{g}) + \text{Cl}_2 (\text{g}) \rightarrow 2\text{NH}_4\text{Cl} (\text{s})$

$$\Delta\text{H}_r = \Sigma\Delta\text{H}_{\text{products}} - \Sigma\Delta\text{H}_{\text{reactants}} = [2(-314.4 \text{ kJ/mol})] - [2(-45.9 \text{ kJ/mol}) + 0 + 0] = -628.8 + 91.8 = \mathbf{-537 \text{ kJ/mol}}$$

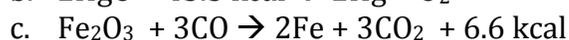
7. Identify each of the following reactions as endothermic or exothermic.



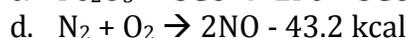
exothermic



endothermic



exothermic



neither - format incorrect

8. Given the equation: $\text{H}_2 (\text{g}) + \frac{1}{2} \text{O}_2 (\text{g}) \rightarrow \text{H}_2\text{O} (\text{g}) + 241.8 \text{ kJ}$. (a) How much heat is evolved in the formation of 2.5 moles of water? (b) How much heat is liberated when 1.75 moles of H_2 burn in excess oxygen? (c) Write a thermochemical equation corresponding to the reaction above for the combustion of 2 moles of H_2 . (d) How much heat is absorbed when 4 moles of water decompose in the reverse reaction? (e) How much heat is evolved for each gram of hydrogen burned?

$$(a) 2.5 \text{ moles} \times \frac{241.8 \text{ kJ}}{1 \text{ mole}} = \mathbf{604.5 \text{ kJ}}$$

$$(b) 1.75 \text{ moles} \times \frac{241.8 \text{ kJ}}{1 \text{ mole}} = \mathbf{423.15 \text{ kJ}}$$



$$(d) 4 \text{ moles} \times \frac{241.8 \text{ kJ}}{1 \text{ mole}} = \mathbf{967.2 \text{ kJ}}$$

$$(e) 1 \text{ g H}_2 \times \frac{1 \text{ mole}}{2 \text{ g}} \times \frac{241.8 \text{ kJ}}{1 \text{ mole}} = \mathbf{120.9 \text{ kJ}}$$

9. An acetylene torch delivers 250 kJ/min. If no loss to the surroundings, **how long will it take** this torch to vaporize 180 g of water at 25°C? The heat of vaporization for water is 40.67 kJ/mol.

$$q = mC_p\Delta T \quad (180 \text{ g}) (4.184 \text{ J/g } ^\circ\text{C}) (75 ^\circ\text{C}) = 56484 \text{ J}$$

$$180 \text{ g} \times \frac{1 \text{ mole}}{18 \text{ g}} \times \frac{40.67 \text{ kJ}}{1 \text{ mole}} = 406.7 \text{ kJ} = 406700 \text{ J}$$

$$\text{Total} = 463184 \text{ J} = 463.184 \text{ kJ} \times \frac{1 \text{ min}}{250 \text{ kJ}} = \mathbf{1.85 \text{ min}}$$

10. Determine the resulting temperature when 100 g of ice at 0°C is mixed with 300 g of water at 60°C. The heat of fusion of ice = 6.01 kJ/mol.

heat gained by ice = heat lost by water

$$(m)(\text{Heat of fusion}) + mC_p\Delta T = mC_p\Delta T$$

$$5.56 \text{ mole} \times 6.01 \text{ kJ/mol} + (100 \text{ g}) (2.06 \text{ J/g } ^\circ\text{C}) (T-0) = (300 \text{ g}) (4.184 \text{ J/g } ^\circ\text{C}) (60-T)$$

$$33389 \text{ J} + 206T = 75312 - 1255.2T$$

$$1461.2T = 41923$$

$$\mathbf{T = 28.7 ^\circ\text{C}}$$

11. Calculate the heat of decomposition of CaCO_3 into CaO and CO_2 . The heats of formation of CaCO_3 , CaO and CO_2 are respectively -1207.1 kJ/mol, -635.5 kJ/mol and -393.5 kJ/mol.

$$\Delta H_f = \sum \Delta H_{\text{products}} - \sum \Delta H_{\text{reactants}}$$

$$(-635.5 + -393.5) - (-1207.1) = \mathbf{+178.1 \text{ kJ/mol}}$$

12. When aluminum is oxidized to form 5.10 grams of aluminum oxide, 19.95 kilocalories of heat are liberated. (A) Calculate the heat of formation of aluminum oxide in kcal/mol. (B) Write the thermochemical equation for this process.

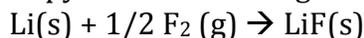
$$(a) 5.10 \text{ g Al}_2\text{O}_3 \times \frac{1 \text{ mole}}{102 \text{ g}} = 0.05 \text{ mole} \quad \frac{19.95 \text{ kcal}}{0.05 \text{ mole}} = \mathbf{399 \text{ kcal/mole}}$$



13. Examine the reactions listed below with their accompanying changes in enthalpy.

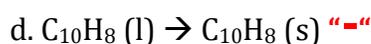
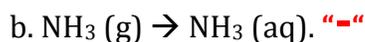
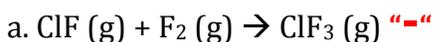
Sublimation:	$\text{Li(s)} \rightarrow \text{Li(g)}$	$\Delta H = 37.1 \text{ kcal}$
Dissociation:	$1/2 \text{ F}_2 \text{ (g)} \rightarrow \text{F (g)}$	$\Delta H = 18.3 \text{ kcal}$
Ionization:	$\text{Li(g)} \rightarrow \text{Li}^+(\text{g}) + \text{e}^-$	$\Delta H = 124 \text{ kcal}$
Electron affinity:	$\text{F(g)} + \text{e}^- \rightarrow \text{F}^-(\text{g})$	$\Delta H = -84 \text{ kcal}$
Crystal energy:	$\text{Li}^+(\text{g}) + \text{F}^-(\text{g}) \rightarrow \text{LiF(s)}$	$\Delta H = -246.6 \text{ kcal}$

What is the change in enthalpy for the following reaction:



Add all the individual ΔH : $37.1 + 18.3 + 124 + -84 + -246.6 = \underline{-151.2 \text{ kcal}}$

14. Predict the sign of ΔS system for each of the following changes. (i.e. is it "+" or "-")



15. Given ΔH_{system} , T , and ΔS_{system} , determine if each of the following processes or reactions is spontaneous or nonspontaneous. $\Delta G = \Delta H - T\Delta S$

a. $\Delta H_{\text{system}} = -75.9 \text{ kJ}$, $T = 273 \text{ K}$, $\Delta S_{\text{system}} = 138 \text{ J/K}$

$$-75900 \text{ J} - (273 \text{ K})(138 \text{ J/K}) =$$

$$-75900 \text{ J} - 37674 \text{ J} = \underline{-113574 \text{ J} (-113.574 \text{ kJ}) \text{ spontaneous}}$$

b. $\Delta H_{\text{system}} = -27.6 \text{ kJ}$, $T = 535 \text{ K}$, $\Delta S_{\text{system}} = -55.2 \text{ J/K}$

$$-27600 \text{ J} - (535 \text{ K})(-55.2 \text{ J/K}) =$$

$$-27600 \text{ J} + 29532 \text{ J} = \underline{1932 \text{ J} (1.932 \text{ kJ}) \text{ nonspontaneous}}$$

c. $\Delta H_{\text{system}} = 365 \text{ kJ}$, $T = 388 \text{ K}$, $\Delta S_{\text{system}} = -55.2 \text{ J/K}$

$$365000 \text{ J} - (388 \text{ K})(-55.2 \text{ J/K}) =$$

$$365000 \text{ J} + 21417.6 \text{ J} = \underline{343582.4 \text{ J} (343.5824 \text{ kJ}) \text{ nonspontaneous}}$$

16. Calculate the **temperature** at which $\Delta G_{\text{system}} = -34.7 \text{ kJ}$ if $\Delta H_{\text{system}} = -28.8 \text{ kJ}$ and $S_{\text{system}} = 22.2 \text{ J/K}$

$$\Delta G = \Delta H - T\Delta S$$

$$-T\Delta S = \Delta G - \Delta H$$

$$-T = \frac{\Delta G - \Delta H}{\Delta S}$$

$$\Delta S$$

$$\frac{-34.7 \text{ kJ} - (-28.8 \text{ kJ})}{0.0222 \text{ kJ/K}}$$

$$0.0222 \text{ kJ/K}$$

$$= \frac{-5.9 \text{ kJ}}{0.0222 \text{ kJ/K}} = -265.76 \text{ K} \rightarrow \underline{265.76 \text{ K}}$$

$$0.0222 \text{ kJ/K}$$