# 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*)

   $C_2H_5OH$ (l) + $O_2$ (g) → $HC_2H_3O_2$ (l) + $H_2O$(l) $\Delta H$ = ??

   $C_2H_5OH$ (l) + $3O_2$(g) → $2CO_2$ (g) + $3H_2O$(l) $\Delta H$= -1366.8 kJ/mol
   $2CO_2$ (g) + $2H_2O$(l) → $HC_2H_3O_2$ (l) + $2O_2$ (g) $\Delta H$= +876.1 kJ/mol
   $C_2H_5OH$ (l) + $O_2$ (g) → $HC_2H_3O_2$ (l) + $H_2O$(l) **$\Delta H$= -490.7 kJ/mol**

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) 188280J ÷5 g= **37656 J/g**
   (c) 37656 J/g ÷1000 J/kJ = **37.656 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).

   $25\text{ g} \times \frac{1\text{ mole}}{18\text{ g}} \times \frac{40.67\text{ kJ}}{1\text{ mole}} =$ **56.486 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 ºC; Specific heat of Cu = 0.385 J/g ºC.

   $q = mC_p\Delta T$
   For copper: (150 g) (0.385 J/g K) (95) = **5486.25 J**

   For aluminum $\Delta T = \frac{q}{mC_p} = \frac{5486.25\text{ J}}{(150\text{ g})(0.897\text{ J/g K})} = 40.77$ °C
   So if both metals start out at 5 °C, **the copper gets hotter**.

5. The heat of combustion of ethylene, $C_2H_4$, is –1411.2 kJ/mol. Determine the heat of formation of ethylene using any necessary data from your textbook.

   2(C (s) + $O_2$ (g) → $CO_2$ (g) $\Delta H^\circ_f$ = -393.5 kJ/mol) -787
   2($H_2$ (g)+ ½ $O_2$ (g) → $H_2O$ (l) $\Delta H^\circ_f$ = -285.8 kJ/mol) -571.6
   $2CO_2$ (g) + $2H_2O$ (l)→ $C_2H_4$ (l) + $3O_2$ (g) $\Delta H_c$ = +1411.2 kJ/mol
   2C (s) + $2H_2$ (g) → $C_2H_4$ (l) **$\Delta H_f$ = +52.6 kJ/mol**

6. Calculate the $\Delta H_r$ for the reaction $2NH_3$ (g) + $H_2$(g) + $Cl_2$ (g) → $2NH_4Cl$ (s)

   $\Delta H_r = \Sigma\Delta H_{products} - \Sigma\Delta H_{reactants}$ = [2(-314.4 kJ/mol)]- [2(-45.9 kJ/mol) + 0 + 0] =
   -628.8 + 91.8 = **-537 kJ/mol**

7. Identify each of the following reactions as endothermic or exothermic.
   a. $N_2$ + $3H_2$ →$2NH_3$ $\Delta H$ = -22 kcal **exothermic**
   b. 2HgO + 43.3 kcal → 2Hg + $O_2$ **endothermic**
   c. $Fe_2O_3$ + 3CO → 2Fe + $3CO_2$ + 6.6 kcal **exothermic**
   d. $N_2$ + $O_2$ → 2NO - 43.2 kcal **neither – format incorrect**

8. Given the equation: $H_2$ (g) + ½ $O_2$ (g) → $H_2O$ (g) + 241.8 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 a absorbed when 4 moles of water decompose in the reverse reaction? (e) How much heat is evolved for each gram of hydrogen burned?

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

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

**(c) $2H_2$ (g) + $O_2$ (g) → $2H_2O$ (l) + 483.65 kJ**

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

(e) $1 \text{ g } H_2 \times \frac{1 \text{ mole}}{2 \text{ g}} \times \frac{241.8 \text{ kJ}}{1 \text{ mole}} =$ **120.9 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.

$q = mC_p\Delta T$ (180 g) (4.184 J/g °C) (75 °C) = 56484 J

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

Total = 463184 J = 463.184 kJ $\times \frac{1 \text{ min}}{250 \text{ kJ}} =$ **1.85 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)(Heat of fusion) + $mC_p\Delta T$ = $mC_p\Delta T$

5.56 mole x 6.01 kJ/mol + (100 g) (2.06 J/g °C) (T-0) = (300 g) (4.184 J/g °C) (60-T)

33389 J + 206T = 75312 – 1255.2T

1461.2T = 41923 **T= 28.7 °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.

$\Delta H_r = \Sigma\Delta H_{products} - \Sigma\Delta H_{reactants}$

(-635.5 + -393.5) – (-1207.1) = **+178.1 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) $5.10 \text{ g } Al_2O_3 \times \frac{1\text{mole}}{102 \text{ g}} = 0.05 \text{ mole}$ $\frac{19.95 \text{ kcal}}{0.05 \text{ mole}} =$ **399 kcal/mole**

(b) **$2Al$ (s) + $3/2O_2$ (g) → $Al_2O_3$ (s) + 399 kcal**

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

| | | |
|---|---|---|
| Sublimation: | $Li(s) \rightarrow Li(g)$ | $\Delta H$= 37.1 kcal |
| Dissociation: | $1/2\ F_2\ (g) \rightarrow F\ (g)$ | $\Delta H$= 18.3 kcal |
| Ionization: | $Li(g) \rightarrow Li^+(g) + e^-$ | $\Delta H$= 124 kcal |
| Electron affinity: | $F(g) + e^- \rightarrow F^-(g)$ | $\Delta H$= -84 kcal |
| Crystal energy: | $Li^+(g) + F^-(g) \rightarrow LiF(s)$ | $\Delta H$= -246.6 kcal |

What is the change in enthalpy for the following reaction:

$Li(s) + 1/2\ F_2\ (g) \rightarrow LiF(s)$

Add all the individual $\Delta H$: 37.1 + 18.3 + 124 + -84 + -246.6 = **-151.2 kcal**

14. Predict the sign of $\Delta S$ system for each of the following changes. (i.e. is it "+" or "-"?)

a. $ClF\ (g) + F_2\ (g) \rightarrow ClF_3\ (g)$ **"-"**

b. $NH_3\ (g) \rightarrow NH_3\ (aq)$. **"-"**

c. $CH_3OH\ (l) \rightarrow CH_3OH\ (aq)$ **"+"**

d. $C_{10}H_8\ (l) \rightarrow C_{10}H_8\ (s)$ **"-"**

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

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

-75900 J – (273 K) (138 J/K) =

-75900 J – 37674 J = **-113574 J (-113.574 kJ) spontaneous**

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

-27600 J – (535 K) (-55.2 J/K) =

-27600 J + 29532 J = **1932 J (1.932 kJ) nonspontaneous**

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

365000 J – (388 K) (-55.2 J/K) =

365000 J + 21417.6 J = **343582.4 J (343.5824 kJ) nonspontaneous**

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

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

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

$-T = \dfrac{\Delta G - \Delta H}{\Delta S} \qquad \dfrac{-34.7\ kJ - (-28.8\ kJ)}{0.0222\ kJ/K} = \dfrac{-5.9\ kJ}{0.0222\ kJ/K} = -265.76\ K \rightarrow \underline{\mathbf{265.76\ K}}$