# CHEM1612 Worksheet 2 – Answers to Critical Thinking Questions

The worksheets are available in the tutorials and form an integral part of the learning outcomes and experience for this unit.

## Model 1: Forming a Compound from its Elements

1.   (a)   $C(s) + 2H_2(g) \rightarrow CH_4(g)$         (b)   $H_2(g) + \frac{1}{2} O_2(g) \rightarrow H_2O(l)$

2.   The enthalpy change when an element in its standard state is formed from the element in its standard state is zero. For example, for $O_2(g)$ and $H_2(g)$, the enthalpy of formation corresponds to the 'reactions':

$O_2(g) \rightarrow O_2(g)$         and         $H_2(g) \rightarrow H_2(g)$
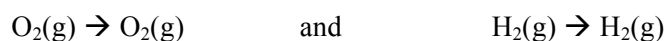
Neither process does *anything* so the enthalpy change must be zero.

## Model 2: Calculating the Enthalpy of Reaction using $\Delta_f H$

1.   $\Delta_{rxn}H° = -892$ kJ mol$^{-1}$

2.   $\Delta_{rxn}H° = -100$ kJ mol$^{-1}$

3.   $\Delta_{rxn}H° = -57$ kJ mol$^{-1}$

4.   The reaction involves *making* a N-N bond, with no bonds being broken. It is exothermic.

5.   $\Delta_{rxn}H° = -28.5$ kJ mol$^{-1}$. This value is exactly half that for reaction A.

6.   $\Delta_{rxn}H° = +57$ kJ mol$^{-1}$. This value is equal to -1 times the value for reaction A and -2 times the value for reaction B.

7.   Reaction C involves *breaking* a N-N bond, with no bonds being made. It is endothermic.

8.   $\Delta_{rxn}H° = +28.5$ kJ mol$^{-1}$. It is $\frac{1}{2}$ the value for reaction C since reaction D is $\frac{1}{2}$ of reaction C.

**Key to success: practice further by completing this week's tutorial homework**

**Key to even greater success: practice even further by completing this week's suggested exam questions**

**Marks
4**

• Nitroglycerine, $C_3H_5(NO_3)_3$, decomposes to form $N_2$, $O_2$, $CO_2$ and $H_2O$ according to the following equation.

$$4C_3H_5(NO_3)_3(l) \rightarrow 6N_2(g) + O_2(g) + 12CO_2(g) + 10H_2O(g)$$
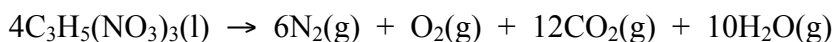
If 15.6 kJ of energy is evolved by the decomposition of 2.50 g of nitroglycerine at 1 atm and 25 °C, calculate the enthalpy change, $\Delta H°$, for the decomposition of 1.00 mol of this compound under standard conditions.

**The molar mass of $C_3H_5(NO_3)_3$ is:**

**$(3 \times 12.01$ (C) $+ 5 \times 1.008$ (H) $+ 3 \times 14.01$ (N) $+ 9 \times 16.00$ (O)) g mol$^{-1}$**

**$= 227.1$ g mol$^{-1}$**

**2.50 g therefore corresponds to:**

**$$\text{number of moles} = \frac{\text{mass}}{\text{molar mass}} = \frac{2.50 \text{ g}}{227.1 \text{ g mol}^{-1}} = 0.0110 \text{ mol}$$**

**As this amount leads to 15.6 kJ being evolved, the enthalpy change for the decomposition of 1.00 mol is:**

**$\Delta H° = 15.6$ kJ / $0.0110$ mol $= -1420$ kJ mol$^{-1}$**

Answer: **–1420 kJ mol$^{-1}$**

Hence calculate the enthalpy of formation of nitroglycerine under standard conditions.

Data:

| | $\Delta_f H°$ (kJ mol$^{-1}$) |
|---|---|
| $H_2O(g)$ | –242 |
| $CO_2(g)$ | –394 |

**The balanced reaction above is for the decomposition of 4 mol of nitroglycerine. Hence, $\Delta_{rxn}H° = 4 \times$ -1420 kJ mol$^{-1}$ = -5680 kJ mol$^{-1}$.**

**Using $\Delta_{rxn}H° = \Sigma m\Delta_f H°$(products) - $\Sigma n\Delta_f H°$(reactants), the enthalpy change for the above reaction is:**

**$\Delta_{rxn}H° = [12\Delta_f H°(CO_2(g)) + 10\Delta_f H°(H_2O(g))]$ - $[4\Delta_f H°(C_3H_5(NO_3)_3(l))]$**

**Hence:**

**-5680 kJ mol$^{-1}$ = $[(12 \times$ -394 $+ 10 \times$ -242) kJ mol$^{-1}]$ - $[4\Delta_f H°(C_3H_5(NO_3)_3(l))]$**

**$\Delta_f H°(C_3H_5(NO_3)_3(l)) = -367$ kJ mol$^{-1}$**

Answer: **–367 kJ mol$^{-1}$**

**Marks**
**3**

• Ammonium perchlorate mixed with powdered aluminium powers the space shuttle booster rockets:

$$2NH_4ClO_4(s) + 2Al(s) \rightarrow Al_2O_3(s) + 2HCl(g) + 2NO(g) + 3H_2O(g)$$
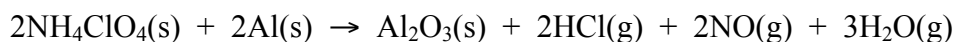
Given the following thermochemical data, how much heat would be released per gram of Al(s)?

$\Delta H_f^\circ$ (H$_2$O(l)) = −285.1 kJ mol$^{-1}$       $\Delta H_f^\circ$ (Al$_2$O$_3$(s)) = −1669.8 kJ mol$^{-1}$

$\Delta H_f^\circ$ (NO(g)) = 90.4 kJ mol$^{-1}$         $\Delta H_f^\circ$ (NH$_4$ClO$_4$(s)) = −290.6 kJ mol$^{-1}$

$\Delta H_f^\circ$ (HCl(g)) = −92.3 kJ mol$^{-1}$       $\Delta H_{vap}^\circ$ (H$_2$O) = 44.1 kJ mol$^{-1}$

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**Using $\Delta_{rxn}H^0 = \sum m\Delta_f H^0(\text{products}) - \sum n\Delta_f H^0(\text{reac tan ts})$:**

$$\Delta_{rxn}H^0 = \sum \Delta_f H°(Al_2O_3(s)) + 2\Delta_f H°(HCl(g)) + 2\Delta_f H°NO(g) + 3\Delta_f H°(H_2O(g))$$

$$- \sum 2\Delta_f H°(NH_4ClO_4(s)) + 2\Delta_f H°(Al(s))$$

$$= [(-1669.8) + 2(-92.3) + 2(90.4) + 3(-285.1 + 44.1)] - [2(-290.6) + 2(0)]$$

$$= -1815.4 \text{ kJ mol}^{-1}$$

**In this calculation, $\Delta_f H°(Al(s))= 0$ for an element in its standard state and $\Delta_f H°(H_2O(g)) = \Delta_f H°(H_2O(l)) + \Delta_{vap}H°(H_2O)$ have been used.**

**As written, this enthalpy change is for the reaction of two moles of Al(s). Therefore, per mole of Al(s), $\Delta H° = \frac{1}{2} \times$ -1815.4 kJ mol$^{-1}$ =907.7 kJ mol$^{-1}$.**

**As the atomic mass of aluminium is 26.98 g mol$^{-1}$, the heat released per gram of Al is:**

$$q = \frac{-907.7 \text{ kJ mol}^{-1}}{26.98 \text{ g mol}^{-1}} = 33.64 \text{ kJ g}^{-1}$$

Answer: **33.64 kJ g$^{-1}$**

**Marks
6**

- The thermite reaction is written below. Show that the heat released in this reaction is sufficient for the iron to be produced as molten metal.

$$2Al(s) + Fe_2O_3(s) \rightarrow Al_2O_3(s) + 2Fe(l)$$
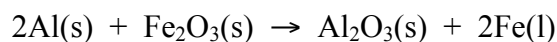
Assume that the values in the table are independent of temperature.

| Substance | Enthalpy of formation, $\Delta_f H^\circ$ kJ mol$^{-1}$ | Molar heat capacity, $C_p$ J K$^{-1}$ mol$^{-1}$ | Melting point $^\circ$C | Enthalpy of fusion kJ mol$^{-1}$ |
|---|---|---|---|---|
| Al | 0 | 24 | 660 | 11 |
| Al$_2$O$_3$ | –1676 | 79 | 2054 | 109 |
| Fe | 0 | 25 | 1535 | 14 |
| Fe$_2$O$_3$ | –824 | 104 | 1565 | 138 |

**Assume 1 mol of reactants at initial temperature of 25 °C. Need to show that $\Delta H$ for the reaction is *greater* than the amount of energy required to melt 2 mol of Fe(s) and heat all the products (2 mol of Fe(s) + 1 mol of Al$_2$O$_3$(s)) to the melting point of Fe.**

$\Delta H = \sum\Delta_f H(\textbf{products}) - \sum\Delta_f H(\textbf{reactants})$

$= \Delta_f H(\textbf{Al}_2\textbf{O}_3\textbf{(s)}) + 2\Delta_f H(\textbf{Fe(s)}) - (2\Delta_f H(\textbf{Al(s)}) + \Delta_f H(\textbf{Fe}_2\textbf{O}_3\textbf{(s)}))$

$= [(-1676 + 2 \times 0) - (-824 + 2 \times 0)] \textbf{ kJ mol}^{-1}$

$= -852 \textbf{ kJ mol}^{-1}$

**$\Delta H$ to heat 2 mol of Fe(s) to its melting point**

$\Delta H = n_{\textbf{Fe(s)}} \times C_p(\textbf{Fe(s)}) \times \Delta T$

$= (2 \textbf{ mol}) \times (25 \textbf{ J K}^{-1} \textbf{ mol}^{-1}) \times (1535-25) \textbf{ K} = 75.5 \textbf{ kJ}$

**$\Delta H$ to heat 1 mol of Al$_2$O$_3$(s) to melting point of Fe(s)**

$\Delta H = n_{\textbf{Al}_2\textbf{O}_3\textbf{(s)}} \times C_p(\textbf{Al}_2\textbf{O}_3\textbf{(s)}) \times \Delta T$

$= (1 \textbf{ mol}) \times (79 \textbf{ J K}^{-1} \textbf{ mol}^{-1}) \times (1535-25) \textbf{ K} = 119 \textbf{ kJ}$

**$\Delta H$ to melt 2 mol of Fe(s)**

$\Delta H = 2 \times \Delta_{\textbf{fus}}H^\circ = (2 \textbf{ mol}) \times (14 \textbf{ kJ mol}^{-1}) = 28 \textbf{ kJ}$

**Total energy required to melt the iron = (75.5 + 119 + 28) kJ = +222.5 kJ.**

**The energy generated by the reaction is more than enough to melt the iron.**