CHEM1109 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\(_2\)O(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’:

   \[ \text{O}_2(g) \rightarrow \text{O}_2(g) \quad \text{and} \quad \text{H}_2(g) \rightarrow \text{H}_2(g) \]

   Neither process does anything so the enthalpy change must be zero.

Model 2: Calculating the Enthalpy of Reaction using \(\Delta_f H\)
1. \(\Delta_{\text{rxn}}H^\circ = -882 \text{ kJ mol}^{-1}\)
2. \(\Delta_{\text{rxn}}H^\circ = -100 \text{ kJ mol}^{-1}\)
3. \(\Delta_{\text{rxn}}H^\circ = -57 \text{ kJ mol}^{-1}\)
4. The reaction involves making a N-N bond, with no bonds being broken. It is exothermic.
5. \(\Delta_{\text{rxn}}H^\circ = -28.5 \text{ kJ mol}^{-1}\). This value is exactly half that for reaction A.
6. \(\Delta_{\text{rxn}}H^\circ = +57 \text{ kJ mol}^{-1}\). This value is equal to -1 times the value for reaction A and -\(\frac{1}{2}\) times the value for reaction B.
7. Reaction C involves breaking a N-N bond, with no bonds being made. It is endothermic.
8. \(\Delta_{\text{rxn}}H^\circ = +28.5 \text{ kJ mol}^{-1}\). 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
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, \(\Delta H^\circ\), for the decomposition of 1.00 mol of this compound under standard conditions.

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

\[
(3 \times 12.01 \text{ (C)} + 5 \times 1.008 \text{ (H)} + 3 \times 14.01 \text{ (N)} + 9 \times 16.00 \text{ (O)}) \text{ g mol}^{-1} = 227.1 \text{ g mol}^{-1}
\]

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:

\[
\Delta H^\circ = 15.6 \text{ kJ} / 0.0110 \text{ mol} = -1420 \text{ kJ mol}^{-1}
\]

Answer: \(-1420 \text{ kJ mol}^{-1}\)

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

**Data:**

<table>
<thead>
<tr>
<th></th>
<th>(\Delta H^\circ) (kJ mol(^{-1}))</th>
</tr>
</thead>
<tbody>
<tr>
<td>(H_2O(g))</td>
<td>-242</td>
</tr>
<tr>
<td>(CO_2(g))</td>
<td>-394</td>
</tr>
</tbody>
</table>

The balanced reaction above is for the decomposition of 4 mol of nitroglycerine. Hence, \(\Delta_{rxn}H^\circ = 4 \times -1420 \text{ kJ mol}^{-1} = -5670 \text{ kJ mol}^{-1}\).

Using \(\Delta_{rxn}H^\circ = \Sigma m\Delta_fH^\circ(\text{products}) - \Sigma n\Delta_fH^\circ(\text{reactants})\), the enthalpy change for the above reaction is:

\[
\Delta_{rxn}H^\circ = [12\Delta_fH^\circ(\text{CO}_2(g)) + 10\Delta_fH^\circ(\text{H}_2\text{O}(g))] - [4\Delta_fH^\circ(\text{C}_3\text{H}_5(\text{NO}_3)_3(l))]
\]

Hence:

\[-5670 \text{ kJ mol}^{-1} = [(12 \times -394 + 10 \times -242) \text{ kJ mol}^{-1}] - [4\Delta_fH^\circ(\text{C}_3\text{H}_5(\text{NO}_3)_3(l))]\]

\[
\Delta_fH^\circ(\text{C}_3\text{H}_5(\text{NO}_3)_3(l)) = -370. \text{ kJ mol}^{-1}
\]

Answer: \(-370. \text{ kJ mol}^{-1}\)
Ammonium perchlorate mixed with powdered aluminium powers the space shuttle booster rockets:

$$2\text{NH}_4\text{ClO}_4(\text{s}) + 2\text{Al}(\text{s}) \rightarrow \text{Al}_2\text{O}_3(\text{s}) + 2\text{HCl}(\text{g}) + 2\text{NO}(\text{g}) + 3\text{H}_2\text{O}(\text{g})$$

Given the following thermochemical data, how much heat would be released per gram of Al(s)?

- $\Delta H^\circ_{f} (\text{H}_2\text{O}(\text{l})) = -285.1 \text{ kJ mol}^{-1}$
- $\Delta H^\circ_{f} (\text{Al}_2\text{O}_3(\text{s})) = -1669.8 \text{ kJ mol}^{-1}$
- $\Delta H^\circ_{f} (\text{NO}(\text{g})) = 90.4 \text{ kJ mol}^{-1}$
- $\Delta H^\circ_{f} (\text{NH}_4\text{ClO}_4(\text{s})) = -290.6 \text{ kJ mol}^{-1}$
- $\Delta H^\circ_{f} (\text{HCl}(\text{g})) = -92.3 \text{ kJ mol}^{-1}$
- $\Delta H^\circ_{\text{vap}} (\text{H}_2\text{O}) = 44.1 \text{ kJ mol}^{-1}$

Using $\Delta_{\text{rxn}} H^\circ = \sum m\Delta_f H^\circ (\text{products}) - \sum n\Delta_f H^\circ (\text{reactants})$:

\[
\Delta_{\text{rxn}} H^\circ = \sum \Delta_f H^\circ (\text{Al}_2\text{O}_3(\text{s})) + 2\Delta_f H^\circ (\text{HCl}(\text{g})) + 2\Delta_f H^\circ (\text{NO}(\text{g})) + 3\Delta_f H^\circ (\text{H}_2\text{O}(\text{g})) \\
- \sum 2\Delta_f H^\circ (\text{NH}_4\text{ClO}_4(\text{s})) + 2\Delta_f H^\circ (\text{Al}(\text{s})) \\
= \{-1669.8 + 2(-92.3) + 2(90.4) + 3(-285.1 + 44.1)\} - \{2(-290.6) + 2(0)\} \\
= -1815.4 \text{ kJ mol}^{-1}
\]

In this calculation, $\Delta_f H^\circ (\text{Al}(\text{s})) = 0$ for an element in its standard state and $\Delta_f H^\circ (\text{H}_2\text{O}(\text{g})) = \Delta_f H^\circ (\text{H}_2\text{O}(\text{l})) + \Delta_{\text{vap}} H^\circ (\text{H}_2\text{O})$ have been used.

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

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$^{-1}$
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.

\[ 2\text{Al(s)} + \text{Fe}_2\text{O}_3(s) \rightarrow \text{Al}_2\text{O}_3(s) + 2\text{Fe(l)} \]

Assume that the values in the table are independent of temperature.

<table>
<thead>
<tr>
<th>Substance</th>
<th>Enthalpy of formation, (\Delta fH^o) kJ mol(^{-1})</th>
<th>Molar heat capacity, (C_p) J K(^{-1}) mol(^{-1})</th>
<th>Melting point °C</th>
<th>Enthalpy of fusion kJ mol(^{-1})</th>
</tr>
</thead>
<tbody>
<tr>
<td>Al</td>
<td>0</td>
<td>24</td>
<td>660</td>
<td>11</td>
</tr>
<tr>
<td>Al(_2)O(_3)</td>
<td>–1676</td>
<td>79</td>
<td>2054</td>
<td>109</td>
</tr>
<tr>
<td>Fe</td>
<td>0</td>
<td>25</td>
<td>1535</td>
<td>14</td>
</tr>
<tr>
<td>Fe(_2)O(_3)</td>
<td>–824</td>
<td>104</td>
<td>1565</td>
<td>138</td>
</tr>
</tbody>
</table>

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.

\[ \Delta H = \sum \Delta fH(\text{products}) - \sum \Delta fH(\text{reactants}) \]
\[ = \Delta fH(\text{Al}_2\text{O}_3(s)) + 2\Delta fH(\text{Fe(s)}) - (2\Delta fH(\text{Al(s)}) + \Delta fH(\text{Fe}_2\text{O}_3(s))) \]
\[ = [(-1676 + 2 \times 0) - (-824 + 2 \times 0)] \text{kJ mol}^{-1} \]
\[ = -852 \text{ kJ mol}^{-1} \]

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

\[ \Delta H = n_{\text{Fe(s)}} \times C_p(\text{Fe(s)}) \times \Delta T \]
\[ = (2 \text{ mol}) \times (25 \text{ J K}^{-1} \text{ mol}^{-1}) \times (1535–25) \text{ K} = 75.5 \text{ kJ} \]

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

\[ \Delta H = n_{\text{Al}_2\text{O}_3(s)} \times C_p(\text{Al}_2\text{O}_3(s)) \times \Delta T \]
\[ = (1 \text{ mol}) \times (79 \text{ J K}^{-1} \text{ mol}^{-1}) \times (1535–25) \text{ K} = 119 \text{ kJ} \]

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

\[ \Delta H = 2 \times \Delta_{\text{fus}} H^o = (2 \text{ mol}) \times (14 \text{ kJ mol}^{-1}) = 28 \text{ kJ} \]

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.