Worksheet 10 – 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: Endothermic and Exothermic Processes
1. (a) Endothermic  (b) Endothermic  (c) Exothermic

Model 2: Enthalpy of Atomization ($\Delta_{\text{atom}}H$) and Enthalpy of Atom Combination ($\Delta_{\text{ac}}H$)
1. $\Delta_{\text{atom}}H = -\Delta_{\text{ac}}H$
2. $\Delta H = 0 \text{ kJ}$

Model 3: Enthalpy of Reaction using $\Delta_{\text{atom}}H$ and $\Delta_{\text{ac}}H$
1. It involves bond breaking only, which requires energy.
2. It involves bond making only, which releases energy.
3. $\Delta H = (239 + 435) \text{ kJ mol}^{-1} + (2 \times -431) \text{ kJ mol}^{-1} = -188 \text{ kJ mol}^{-1}$
4. $\Delta_{\text{rxn}}H = \Delta_{\text{ac}}H \text{ (products)} - \Delta_{\text{ac}}H \text{ (reactants)}$
5. $\Delta_{\text{rxn}}H = \Delta_{\text{atom}}H \text{ (reactants)} - \Delta_{\text{atom}}H \text{ (products)}$
6. $\Delta_{\text{rxn}}H$ will be positive. If the bonds that need to be broken in the reactants are stronger than those formed in the products, the reaction is endothermic.

Model 4: Enthalpy of Reaction using $\Delta_iH$
1. $\Delta_iH^\circ (\text{O}_2(\text{g}))$ refers to the formation of $\text{O}_2(\text{g})$ from the element in its standard state (i.e. $\text{O}_2(\text{g})$):
   $$\text{O}_2(\text{g}) \rightarrow \text{O}_2(\text{g})$$
   $\Delta_iH^\circ (\text{H}_2(\text{g}))$ refers to the formation of $\text{H}_2(\text{g})$ from the element in its standard state (i.e. $\text{H}_2(\text{g})$):
   $$\text{H}_2(\text{g}) \rightarrow \text{H}_2(\text{g})$$
   Both processes involve no change and so $\Delta H^\circ = 0$.
2. $\Delta_{\text{rxn}}H^\circ = -892 \text{ kJ mol}^{-1}$
3. $\Delta_{\text{rxn}}H^\circ = ((-1096) - (-602 + -394)) \text{ kJ mol}^{-1} = -100 \text{ kJ mol}^{-1}$
Diborane (B\(_2\)H\(_6\)) is a highly reactive compound, which was once considered as a possible rocket fuel for the US space program. Calculate the heat of formation of diborane at 298 K from the following reactions.

<table>
<thead>
<tr>
<th>Reaction</th>
<th>(\Delta H_f) (kJ mol(^{-1}))</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>2B(s) + (\frac{3}{2})O(_2)(g) (\rightarrow) B(_2)O(_3)(s)</td>
</tr>
<tr>
<td>2</td>
<td>B(_2)H(_6)(g) + 3O(_2)(g) (\rightarrow) B(_2)O(_3)(s) + 3H(_2)O(g)</td>
</tr>
<tr>
<td>3</td>
<td>H(_2)(g) + (\frac{1}{2})O(_2)(g) (\rightarrow) H(_2)O(l)</td>
</tr>
<tr>
<td>4</td>
<td>H(_2)O(l) (\rightarrow) H(_2)O(g)</td>
</tr>
</tbody>
</table>

The heat of formation of a compound refers to its formation from the elements in their standard states.

In the reactions above, (1) and (3) correspond to the formation of B\(_2\)O\(_3\)(s) and H\(_2\)O(l) from their elements so:

\[
\Delta_f H^0 (B_2O_3(s)) = -1273 \text{ kJ mol}^{-1} \quad \text{and} \quad \Delta_f H^0 (H_2O(l)) = -286 \text{ kJ mol}^{-1}
\]

Formation of H\(_2\)O(g) from its elements corresponds to formation of H\(_2\)O(l) [\(-286 \text{ kJ mol}^{-1}\)] followed by vaporization of H\(_2\)O(l) [\(+44 \text{ kJ mol}^{-1}\) (reaction (4))]. The heat of formation of H\(_2\)O(g) is therefore:

\[
\Delta_f H^0 (H_2O (g)) = -286 + 44 = -242 \text{ kJ mol}^{-1}
\]

Reaction (2) corresponds to the combustion of B\(_2\)H\(_6\)(s). The heat of combustion is given by:

\[
\Delta_{rxn} H^0 = \sum m\Delta_f H^0 (\text{products}) - \sum n\Delta_f H^0 (\text{reactants})
\]

\[
= [\Delta_f H^0 (B_2O_3(s)) + 3\Delta_f H^0 (H_2O(g))] - [\Delta_f H^0 (B_2H_6(s)) + 3\Delta_f H^0 (O_2(g))]\text{All of}
\]

\[
= -2035 \text{ kJ mol}^{-1}
\]

The \(\Delta_f H^0\) values except that for B\(_2\)H\(_6\)(s) are known from above – the \(\Delta_f H^0 (O_2(g))\) is zero because it is an element in its standard state. Substituting these values in gives:

\[
[-1273 + 3 \times (-242)] - [\Delta_f H^0 (B_2H_6(s)) + 0] = -2035 \text{ kJ mol}^{-1}
\]

or

\[
\Delta_f H^0 (B_2H_6(s)) = [-1273 + 3 \times (-242)] + 2035 = +36 \text{ kJ mol}^{-1}
\]

Answer: \(+36 \text{ kJ mol}^{-1}\)
The current “petrochemical economy” is based on the combustion of fossil fuels, of which octane is a typical example.

\[
2\text{C}_8\text{H}_{18}(l) + 25\text{O}_2(g) \rightarrow 16\text{CO}_2(g) + 18\text{H}_2\text{O}(l)
\]

Calculate the heat of combustion of octane using the supplied heat of formation data.
Data: \(\text{C}_8\text{H}_{18}(l): -249.9 \text{ kJ mol}^{-1}\); \(\text{CO}_2(g): -393.5 \text{ kJ mol}^{-1}\); \(\text{H}_2\text{O}(l): -285.8 \text{ kJ mol}^{-1}\)

Using \(\Delta_{\text{rxn}}H^\circ = \Sigma m\Delta_f H^\circ(\text{products}) - \Sigma n\Delta_f H^\circ(\text{reactants})\), the heat of the reaction as written is:

\[
\Delta_{\text{rxn}}H^\circ = [16\Delta_f H^\circ(\text{CO}_2(g)) + 18\Delta_f H^\circ(\text{H}_2\text{O}(l))] - [2\Delta_f H^\circ(\text{C}_8\text{H}_{18}(l))]
\]

\[
= [(16 \times -393.5) + (18 \times -285.8)] - [(2 \times -249.9)]
\]

\[
= -10940 \text{ kJ mol}^{-1}
\]

where \(\Delta_f H^\circ(\text{O}_2(g)) = 0\) has been used for the formation of an element in its standard state.

The enthalpy of combustion is defined per mole of fuel. The above reaction is for the combustion of two moles of \(\text{C}_8\text{H}_{18}\).

Hence, \(\Delta_{\text{comb}}H^\circ = \frac{1}{2} \times -10940 = -5470 \text{ kJ mol}^{-1}\)

Answer: \(-5470 \text{ kJ mol}^{-1}\)