Lecture 16

Thermodynamics II

Calorimetry
Hess’s Law
Enthalpy or Formation
Four Methods for Finding $\Delta H$

1) Calculate it using average bond enthalpies
2) Measure it using a calorimeter.
3) Calculate it using Hess’s Law.
4) Calculate it using enthalpies of formation
Calorimetry

- Calorimetry
  - The measurement of heat transfer

- Calorimeter
  - A device used to determine the amount of heat transferred

Heat flow cannot be measured directly. We measure the temperature before and after a reaction, and use that data to calculate heat transfer.
Calorimetry

• Specific Heat Capacity \((c)\)
  
  • The amount of heat required to raise the temperature of 1 g of a substance by 1 K.
  
  • A large value for water (4.184 J/g\(\cdot\)K)
    – It takes a lot of heat to increase the temperature of water.
  
  • A small value for iron (0.45 J/g\(\cdot\)K)
    – It takes much less heat to raise the temperature of iron.

\[
q = m \cdot c \cdot \Delta T
\]

Heat lost or gained by a substance

Mass of that substance

Specific heat capacity of that substance

Temperature change of that substance

\(\Delta T = T_f - T_i\)
Coffee Cup Calorimeter

Thermometer

Stir Stick
Coffee Cup Calorimetry

When measuring the heat lost or gained in a chemical reaction that takes place in a coffee cup calorimeter, we use the solution as the basis for our calculations.

\[ q = m \cdot c \cdot \Delta T \]

- **Heat lost or gained by the solution in the reaction**
- **Mass of the solution**
- **Specific heat of the solution**: 4.184 J/g•K is often used if concentrations are low.
- **Temperature change of the solution**: \( \Delta T = T_f - T_i \)
Coffee Cup Calorimetry

• The reactants and products are the system.
• The water which they are dissolved in and calorimeter make up the surroundings.
• The heat lost or gained by the solution is equal to the heat lost or gained by the reaction.

E.g.) If you calculated $q_{\text{solution}}$ to be $+568$ J, the solution gained heat. This means the reaction was exothermic, as the heat gained by the solution is equal to the heat lost by the reaction ($q_{\text{rxn}} = -568$ J).
Ex1) A 97 g iron bar was heated to 143°C and placed in a coffee cup calorimeter containing 350.0 g of water at 25.00°C. The maximum temperature of the water was measured to be 28.40°C.

a) How much heat was gained by the water?
b) How much heat was lost by the iron bar?
c) What is the specific heat of iron?
Ex1) Calorimetry (cont.)

a) How much heat was gained by the water?

b) How much heat was lost by the iron bar?
Ex1) Calorimetry (cont.)

a) How much heat was gained by the water?

\[ q_w = mc\Delta T \]

b) How much heat was lost by the iron bar?
Ex1) Calorimetry (cont.)

a) How much heat was gained by the water?

\[ q_w = mc\Delta T \]

\[ q_w = (350.0 \text{ g})(4.184 \text{ J/g} \cdot \text{°C})(28.40 \text{ °C} - 25.00 \text{ °C}) \]

b) How much heat was lost by the iron bar?
Ex1) Calorimetry (cont.)

a) How much heat was gained by the water?

\[ q_w = mc\Delta T \]

\[ q_w = (350.0 \text{ g})(4.184 \text{ J/g} \cdot ^\circ \text{C})(28.40 ^\circ \text{C} - 25.00 ^\circ \text{C}) \]

\[ q_w = +4979 \text{ J} \]

Gaining heat is endothermic

b) How much heat was lost by the iron bar?
Ex1) Calorimetry (cont.)

a) How much heat was gained by the water?

\[ q_w = mc\Delta T \]

\[ q_w = (350.0 \text{ g})(4.184 \text{ J/g} \cdot ^\circ \text{C})(28.40 \ ^\circ \text{C} - 25.00 \ ^\circ \text{C}) \]

\[ q_w = +4979 \text{ J} \]

Gaining heat is endothermic

b) How much heat was lost by the iron bar?

\[ q_w = -q_{Fe} \]
Ex1) Calorimetry (cont.)

a) How much heat was gained by the water?

\[ q_w = mc\Delta T \]
\[ q_w = (350.0 \text{ g})(4.184 \text{ J/g} \cdot \text{°C})(28.40 \text{ °C} - 25.00 \text{ °C}) \]
\[ q_w = +4979 \text{ J} \]

Gaining heat is endothermic

b) How much heat was lost by the iron bar?

\[ q_w = -q_{Fe} \]
\[ q_{Fe} = -4979 \text{ J} \]

Losing heat is exothermic

As energy is conserved, the heat gained by the water is equal in magnitude to the heat lost by the iron bar.
c) What is the specific heat of iron?
c) What is the specific heat of iron?

\[ q_{Fe} = mc\Delta T \]
c) What is the specific heat of iron?

\[ q_{\text{Fe}} = mc\Delta T \]

\[ c = \frac{q_{\text{Fe}}}{m\Delta T} \]
c) What is the specific heat of iron?

\[ q_{Fe} = mc\Delta T \]

\[ c = \frac{q_{Fe}}{m\Delta T} \]

\[ c = \frac{-4979 \text{ J}}{(97 \text{ g})(28.40 \ ^\circ\text{C} - 143 \ ^\circ\text{C})} \]
c) What is the specific heat of iron?

\[ q_{Fe} = mc\Delta T \]

\[ c = \frac{q_{Fe}}{m\Delta T} \]

\[ c = \frac{-4979 \text{ J}}{(97 \text{ g})(28.40 ^\circ \text{C} - 143 ^\circ \text{C})} \]

\[ c = 0.45 \text{ J/g} \cdot ^\circ \text{C} \]
Ex2) When 20.0 mL of 1.00 \( M \) NaCl at 22.00\(^\circ\)C is mixed with 20.0 mL of 1.00 \( M \) AgNO\(_3\) at 22.00\(^\circ\)C in a calorimeter, a white precipitate forms and the temperature of the mixture reaches 29.60\(^\circ\)C. Assume that the specific heat capacity of the mixture is 4.184 J/g•K, its density is 1.00 g/mL, and the volumes are additive.

a) Write the net ionic equation for this reaction.

b) Calculate \( \Delta H \) for the reaction.
Ex2) When 20.0 mL of 1.00 M NaCl at 22.00°C is mixed with 20.0 mL of 1.00 M AgNO₃ at 22.00°C in a calorimeter, a white precipitate forms and the temperature of the mixture reaches 29.60°C. Assume that the specific heat capacity of the mixture is 4.184 J/g·K, its density is 1.00 g/mL, and the volumes are additive.

a) Write the net ionic equation for this reaction.

\[ \text{Ag}^+ (aq) + \text{Cl}^- (aq) \rightarrow \text{AgCl (s)} \]

b) Calculate $\Delta H$ for the reaction.
Ex2) Calorimetry (cont.)

Step 1) Calculate $q_{sol}$.

Step 2) Calculate $q_{rxn}$. 
Ex2) Calorimetry (cont.)

Step 1) Calculate $q_{sol}$.

$$q_{sol} = mc\Delta T$$

Step 2) Calculate $q_{rxn}$. 
Ex2) Calorimetry (cont.)

Step 1) Calculate $q_{sol}$.

$$ q_{sol} = mc\Delta T $$

$$ q_{sol} = (40.0 \text{ g})(4.184 \text{ J/g} \cdot ^\circ C)(29.60^\circ C - 22.00^\circ C) $$

Step 2) Calculate $q_{rxn}$. 
Step 1) Calculate $q_{sol}$.

$$q_{sol} = mc\Delta T$$

$$q_{sol} = (40.0\text{ g})(4.184\text{ J/g} \cdot ^\circ\text{C})(29.60^\circ\text{C} - 22.00^\circ\text{C})$$

$$q_{sol} = (40.0\text{ g})(4.184\text{ J/g} \cdot ^\circ\text{C})(7.60^\circ\text{C})$$

Step 2) Calculate $q_{rxn}$.
Ex2) Calorimetry (cont.)

Step 1) Calculate $q_{sol}$.

$$ q_{sol} = mc\Delta T $$

$$ q_{sol} = (40.0 \text{ g})(4.184 \text{ J/g} \cdot ^\circ \text{C})(29.60^\circ \text{C} - 22.00^\circ \text{C}) $$

$$ q_{sol} = (40.0 \text{ g})(4.184 \text{ J/g} \cdot ^\circ \text{C})(7.60^\circ \text{C}) $$

$$ q_{sol} = +1270 \text{ J} $$

Step 2) Calculate $q_{rxn}$.
Step 1) Calculate $q_{sol}$.

$$q_{sol} = mc\Delta T$$

$$q_{sol} = (40.0 \text{ g})(4.184 \text{ J/g·°C})(29.60°C - 22.00°C)$$

$$q_{sol} = (40.0 \text{ g})(4.184 \text{ J/g·°C})(7.60°C)$$

$$q_{sol} = +1270 \text{ J}$$

Step 2) Calculate $q_{rxn}$.

$$q_{rxn} = -q_{sol}$$

$$q_{rxn} = -1270 \text{ J}$$

The reaction is **exothermic** as heat is released to the water.
Step 3) Calculate Moles of Ag$^+$ or Cl$^-$. 

Step 4) Calculate $\Delta H_{rxn}$. 

Ex2) Calorimetry (cont.)
Ex2) Calorimetry (cont.)

Step 3) Calculate Moles of Ag\(^+\) or Cl\(^-\).

\[
0.0200 \text{ L} \times \frac{1.0 \text{ moles Ag}^+}{1 \text{ L}} = 0.0200 \text{ moles Ag}^+
\]

Step 4) Calculate \(\Delta H_{rxn}\).
Step 3) Calculate Moles of Ag\(^+\) or Cl\(^-\).

\[
0.0200 \text{ L} \times \frac{1.0 \text{ moles Ag}^+}{1 \text{ L}} = 0.0200 \text{ moles Ag}^+
\]

Step 4) Calculate \(\Delta H_{rxn}\).

\[
\Delta H_{rxn} = \frac{q_{rxn}}{n_{\text{AgCl}}}
\]

\[
\Delta H_{rxn} = \frac{-1270 \text{ J}}{0.0200 \text{ moles AgCl}}
\]
Ex2) Calorimetry (cont.)

Step 3) Calculate Moles of Ag⁺ or Cl⁻.

\[
0.0200 \text{ L} \times \frac{1.0 \text{ moles Ag}^+}{1 \text{ L}} = 0.0200 \text{ moles Ag}^+
\]

Step 4) Calculate \(\Delta H_{\text{rxn}}\).

\[
\Delta H_{\text{rxn}} = \frac{q_{\text{rxn}}}{n_{\text{AgCl}}}
\]

\[
\Delta H_{\text{rxn}} = \frac{-1270 \text{ J}}{0.0200 \text{ moles AgCl}}
\]

\[
\Delta H_{\text{rxn}} = -63500 \text{ J/mol} = -63.5 \text{ kJ/mol}
\]
\[ q_{\text{rxn}} \text{ vs. } \Delta H_{\text{rxn}} \]

\[ q_{\text{rxn}} = \text{The heat lost or gained in the experiment that took place in the calorimeter} \]

\[ \Delta H_{\text{rxn}} = \text{The heat lost or gained in the balanced chemical equation} \]
Four Methods for Finding $\Delta H$

1) Calculate it using average bond enthalpies.
2) Measure it using a calorimeter.
3) **Calculate it using Hess’s Law.**
4) Calculate it using enthalpies of formation.
Hess’s Law

If a reaction is carried out in a series of steps, the overall change in enthalpy will be equal to the sum of the enthalpy changes for the individual steps.

• The overall enthalpy change will be the same if a reaction is carried out in one step or in several steps.
**Ex1) Hess’s Law**

Ex1) Calculate \( \Delta H \) for the following reaction using the information in the table below.

\[
4 \text{NO}_2(g) + \text{O}_2(g) \rightarrow 2 \text{N}_2\text{O}_5(s)
\]

<table>
<thead>
<tr>
<th>Reaction</th>
<th>Chemical Equation</th>
<th>Enthalpy Change</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>( 2 \text{NO}_2(g) + \frac{3}{2}\text{O}_2(g) \rightarrow \text{N}_2\text{O}_5(s) )</td>
<td>( \Delta H = -223.7 \text{ kJ} )</td>
</tr>
<tr>
<td>2</td>
<td>( 2 \text{NO}_2(g) + \text{O}_2(g) \rightarrow 2 \text{NO}_2(g) )</td>
<td>( \Delta H = -114.2 \text{ kJ} )</td>
</tr>
</tbody>
</table>
Ex1) Hess’s Law (cont.)
Ex1) Hess’s Law (cont.)

| 2(rxn 1) | $4 \text{NO}_{(g)} + 3 \text{O}_2(g) \rightarrow 2 \text{N}_2\text{O}_5(s)$ | $\Delta H = 2(-223.7 \text{ kJ})$ |
Ex1) Hess’s Law (cont.)

| 2(rxn 1) | 4 NO\(_{(g)}\) + 3 O\(_{2(g)}\) \rightarrow 2 N\(_{2O5(s)}\) \quad \Delta H = 2(-223.7 \text{ kJ}) |
| -2(rxn 2) | 4 NO\(_{2(g)}\) \rightarrow 4 NO\(_{(g)}\) + 2 O\(_{2(g)}\) \quad \Delta H = -2(-114.2 \text{ kJ}) |
### Ex1) Hess’s Law (cont.)

<table>
<thead>
<tr>
<th>2(rxn 1)</th>
<th>$4 \text{NO}_{(g)} + 3 \text{O}_2(g) \rightarrow 2 \text{N}_2\text{O}_5(s)$</th>
<th>$\Delta H = 2(-223.7 \text{ kJ})$</th>
</tr>
</thead>
<tbody>
<tr>
<td>-2(rxn 2)</td>
<td>$4 \text{NO}<em>2(g) \rightarrow 4 \text{NO}</em>{(g)} + 2 \text{O}_2(g)$</td>
<td>$\Delta H = -2(-114.2 \text{ kJ})$</td>
</tr>
</tbody>
</table>
### Ex1) Hess’s Law (cont.)

<table>
<thead>
<tr>
<th>Reaction 1</th>
<th>Reaction 2</th>
<th>Combined ΔH</th>
</tr>
</thead>
<tbody>
<tr>
<td>2(rxn 1)</td>
<td>4 (\text{NO}_2(g)) (\rightarrow) 4 (\text{NO}(g)) + 2 (\text{O}_2(g))</td>
<td>(\Delta H = 2(-223.7 \text{ kJ}))</td>
</tr>
<tr>
<td>4 (\text{NO}(g)) + 3 (\text{O}_2(g)) (\rightarrow) 2 (\text{N}_2\text{O}_5(s))</td>
<td>(\Delta H = -2(-114.2 \text{ kJ}))</td>
<td></td>
</tr>
<tr>
<td>-2(rxn 2)</td>
<td>4 (\text{NO}_2(g)) (\rightarrow) 4 (\text{NO}(g)) + 2 (\text{O}_2(g))</td>
<td>(\Delta H = -219.0 \text{ kJ})</td>
</tr>
<tr>
<td>4 (\text{NO}_2(g)) + (\text{O}_2(g)) (\rightarrow) 2 (\text{N}_2\text{O}_5(s))</td>
<td>(\Delta H = -219.0 \text{ kJ})</td>
<td></td>
</tr>
</tbody>
</table>
Ex2) Hess’s Law

Ex2) Calculate $\Delta H$ for the following reaction using the information in the table below.

$$3 \text{ C}_\text{s}(s) + 4 \text{ H}_2\text{(g)} \rightarrow \text{C}_3\text{H}_8\text{(g)}$$

<table>
<thead>
<tr>
<th>Reaction</th>
<th>Chemical Equation</th>
<th>Enthalpy Change</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>$2 \text{ H}_2\text{(g)} + \text{ O}_2\text{(g)} \rightarrow 2 \text{ H}_2\text{O} \text{(l)}$</td>
<td>$\Delta H = -571.7 \text{ kJ}$</td>
</tr>
<tr>
<td>2</td>
<td>$\text{C}_3\text{H}_8\text{(g)} + 5\text{ O}_2\text{(g)} \rightarrow 3\text{CO}_2\text{(g)} + 4\text{H}_2\text{O} \text{(l)}$</td>
<td>$\Delta H = -2220.1 \text{ kJ}$</td>
</tr>
<tr>
<td>3</td>
<td>$\text{C} \text{(s)} + \text{ O}_2\text{(g)} \rightarrow \text{CO}_2\text{(g)}$</td>
<td>$\Delta H = -393.5 \text{ kJ}$</td>
</tr>
</tbody>
</table>
Ex2) Hess’s Law (cont.)

2(rxn 1) \[ 4 \text{H}_2(g) + 2 \text{O}_2(g) \rightarrow 4 \text{H}_2\text{O}(l) \]

\[ \Delta H = 2(-571.7 \text{ kJ}) \]
Ex2) Hess’s Law (cont.)

\[
\begin{array}{l}
\text{2(rxn 1)} & 4 \text{H}_2(\text{g}) + 2 \text{O}_2(\text{g}) \rightarrow 4 \text{H}_2\text{O}(\text{l}) & \Delta H = 2(-571.7 \text{ kJ}) \\
\text{-1(rxn 2)} & 3 \text{CO}_2(\text{g}) + 4 \text{H}_2\text{O}(\text{l}) \rightarrow \text{C}_3\text{H}_8(\text{g}) + 5 \text{O}_2(\text{g}) & \Delta H = -(-2220.1 \text{ kJ}) \\
\end{array}
\]
### Ex2) Hess’s Law (cont.)

<table>
<thead>
<tr>
<th>Reaction</th>
<th>Equation</th>
<th>ΔH</th>
</tr>
</thead>
<tbody>
<tr>
<td>2(rxn 1)</td>
<td>(4 \text{ H}_2(g) + 2 \text{ O}_2(g) \rightarrow 4 \text{ H}_2\text{O}(l))</td>
<td>(\Delta H = 2(-571.7 \text{ kJ}))</td>
</tr>
<tr>
<td>-1(rxn 2)</td>
<td>(3 \text{ CO}_2(g) + 4 \text{ H}_2\text{O}(l) \rightarrow \text{C}_3\text{H}_8(g) + 5 \text{ O}_2(g))</td>
<td>(\Delta H = -(2220.1 \text{ kJ}))</td>
</tr>
<tr>
<td>3(rxn 3)</td>
<td>(3 \text{ C}(s) + 3 \text{ O}_2(g) \rightarrow 3 \text{ CO}_2(g))</td>
<td>(\Delta H = 3(-393.5 \text{ kJ}))</td>
</tr>
</tbody>
</table>
### Ex2) Hess’s Law (cont.)

<table>
<thead>
<tr>
<th>Reaction</th>
<th>Balanced Equation</th>
<th>Change in Enthalpy ($\Delta H$)</th>
</tr>
</thead>
<tbody>
<tr>
<td>2(rxn 1)</td>
<td>$4 \text{H}_2(g) + 2 \text{O}_2(g) \rightarrow 4 \text{H}_2\text{O}(l)$</td>
<td>$\Delta H = 2(-571.7 \text{kJ})$</td>
</tr>
<tr>
<td>-1(rxn 2)</td>
<td>$3 \text{CO}_2(g) + 4 \text{H}_2\text{O}(l) \rightarrow \text{C}_3\text{H}_8(g) + 5 \text{O}_2(g)$</td>
<td>$\Delta H = -(-2220.1 \text{kJ})$</td>
</tr>
<tr>
<td>3(rxn 3)</td>
<td>$3 \text{C}(s) + 3 \text{O}_2(g) \rightarrow 3 \text{CO}_2(g)$</td>
<td>$\Delta H = 3(-393.5 \text{kJ})$</td>
</tr>
</tbody>
</table>

$3 \text{C}(s) + 4 \text{H}_2(g) \rightarrow \text{C}_3\text{H}_8(g)$

$\Delta H = -103.8 \text{kJ}$
Four Methods for Finding $\Delta H$

1) Calculate it using average bond enthalpies.
2) Measure it using a calorimeter.
3) Calculate it using Hess’s Law.
4) Calculate it using enthalpies of formation.
Standard Enthalpy of Formation

A hypothetical value that indicates how much heat would be lost or gained during the formation of one mole of a compound from the most common form of its elements in their standard states.

\[ \Delta H_f^0 \rightarrow \text{Formation} \]

standard state (25°C, 1 atm)
Heat of formation reactions are always written so that all reactants exist as they would under standard conditions, and there is one mole of product.
The $\Delta H_f^0$ value for the most stable form of any element in its standard state is zero.

<table>
<thead>
<tr>
<th>Element</th>
<th>$\Delta H_f^0$</th>
<th>Element</th>
<th>$\Delta H_f^0$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ca(s)</td>
<td>0</td>
<td>Cl$_2$(g)</td>
<td>0</td>
</tr>
<tr>
<td>Ag(s)</td>
<td>0</td>
<td>H$_2$(g)</td>
<td>0</td>
</tr>
<tr>
<td>Na(s)</td>
<td>0</td>
<td>Fe(s)</td>
<td>0</td>
</tr>
<tr>
<td>O$_2$(g)</td>
<td>0</td>
<td>N$_2$(g)</td>
<td>0</td>
</tr>
</tbody>
</table>
Formula for Calculating the Enthalpy of a Reaction under Standard Conditions

\[ \Delta H_{\text{rxn}}^0 = \sum n \Delta H_f^0 (\text{products}) - \sum n \Delta H_f^0 (\text{reactants}) \]

Sum of stoichiometric coefficients
Ex1) Enthalpy of Formation

Ex1) Find $\Delta H^\circ_{\text{rxn}}$ for the thermite reaction under standard conditions.

$$\text{Fe}_2\text{O}_3(s) + 2 \text{Al}(s) \rightarrow 2 \text{Fe}(s) + \text{Al}_2\text{O}_3(s)$$

$$\Delta H^\circ_{\text{rxn}} = \sum n \Delta H^\circ_f (\text{products}) - \sum n \Delta H^\circ_f (\text{reactants})$$

$$\Delta H^\circ_{\text{rxn}} = [2 \Delta H^\circ_f (\text{Fe}(s)) + 1 \Delta H^\circ_f (\text{Al}_2\text{O}_3(s))] - [2 \Delta H^\circ_f (\text{Al}(s)) + 1 \Delta H^\circ_f (\text{Fe}_2\text{O}_3(s))]$$

$$\Delta H^\circ_{\text{rxn}} = [2(0) + 1(-1675.7)] - [2(0) + 1(-821.4)]$$

$$\Delta H^\circ_{\text{rxn}} = -854.3 \text{ kJ/mol}$$
Ex2) Enthalpy of Formation

Ex2) Find $\Delta H_{\text{rxn}}^0$ for this reaction at standard conditions.

$$2 \text{H}_2\text{S}(g) + 3 \text{O}_2(g) \rightarrow 2 \text{SO}_2(g) + 2 \text{H}_2\text{O}(g)$$

$$\Delta H_{\text{rxn}}^0 = \sum n \Delta H_f^o \text{ (products)} - \sum n \Delta H_f^o \text{ (reactants)}$$
Ex2) Enthalpy of Formation

Ex2) Find $\Delta H^\circ_{\text{rxn}}$ for this reaction at standard conditions.

$$2 \text{H}_2\text{S}_{(g)} + 3 \text{O}_2_{(g)} \rightarrow 2 \text{SO}_2_{(g)} + 2 \text{H}_2\text{O}_{(g)}$$

$$\Delta H^\circ_{\text{rxn}} = \sum n \Delta H^\circ_f \text{ (products)} - \sum n \Delta H^\circ_f \text{ (reactants)}$$

$$\Delta H^\circ_{\text{rxn}} = [2 (\Delta H^\circ_f \text{(SO}_2_{(g)}) + 2 (\Delta H^\circ_f \text{(H}_2\text{O}_{(g)})]) - [2 (\Delta H^\circ_f \text{(H}_2\text{S}_{(g)}) + 3 (\Delta H^\circ_f \text{(O}_2_{(g)})])$$
Ex2) Enthalpy of Formation

Ex2) Find $\Delta H^o_{\text{rxn}}$ for this reaction at standard conditions.

$$2 \text{ H}_2\text{S}_{(g)} + 3 \text{ O}_2(g) \rightarrow 2 \text{ SO}_2(g) + 2 \text{ H}_2\text{O}_{(g)}$$

$$\Delta H^o_{\text{rxn}} = \sum n \Delta H_f^o \text{ (products)} - \sum n \Delta H_f^o \text{ (reactants)}$$

$$\Delta H^o_{\text{rxn}} = [2 (\Delta H_f^o (\text{SO}_2(g))) + 2 (\Delta H_f^o (\text{H}_2\text{O}_{(g)}))] - [2 (\Delta H_f^o (\text{H}_2\text{S}_{(g)}))) + 3 (\Delta H_f^o (\text{O}_2(g)))]$$

$$\Delta H^o_{\text{rxn}} = [2(-296.8) + 2(-241.8)] - [2(-20.2) + 3(0)]$$

$$\Delta H^o_{\text{rxn}} = -1036.8 \text{ kJ/mol}$$
## Enthalpy of Formation

<table>
<thead>
<tr>
<th>Reaction</th>
<th>$\Delta H_f^o$ kJ/mol</th>
<th>Elemental metal 1\textsuperscript{st} used by humans</th>
</tr>
</thead>
<tbody>
<tr>
<td>$2 \text{ Au}_\text{(s)} + \frac{3}{2} \text{ O}_2 \text{(g)} \rightarrow \text{ Au}_2\text{O}_3\text{(s)}$</td>
<td>$+131$</td>
<td>$\sim6000$ BCE</td>
</tr>
<tr>
<td>$2 \text{ Ag}_\text{(s)} + \frac{1}{2} \text{ O}_2 \text{(g)} \rightarrow \text{ Ag}_2\text{O}\text{(s)}$</td>
<td>$-31$</td>
<td>$\sim4000$ BCE</td>
</tr>
<tr>
<td>$\text{ Cu}_\text{(s)} + \frac{1}{2} \text{ O}_2 \text{(g)} \rightarrow \text{ CuO}\text{(s)}$</td>
<td>$-156$</td>
<td>$\sim4000$ BCE</td>
</tr>
<tr>
<td>$\text{ Pb}_\text{(s)} + \frac{1}{2} \text{ O}_2 \text{(g)} \rightarrow \text{ PbO}\text{(s)}$</td>
<td>$-217$</td>
<td>$\sim3500$ BCE</td>
</tr>
<tr>
<td>$\text{ Sn}_\text{(s)} + \text{ O}_2 \text{(g)} \rightarrow \text{ SnO}_2\text{(s)}$</td>
<td>$-581$</td>
<td>$\sim1800$ BCE</td>
</tr>
<tr>
<td>$2 \text{ Fe}_\text{(s)} + \frac{3}{2} \text{ O}_2 \text{(g)} \rightarrow \text{ Fe}_2\text{O}_3\text{(s)}$</td>
<td>$-822$</td>
<td>$\sim1200$ BCE</td>
</tr>
</tbody>
</table>
Ex3) Enthalpy of Formation

Ex3) Use the enthalpy of formation values below to determine $\Delta H_{\text{rxn}}$ for the reaction between $C_{(s)}$ and $Fe_2O_3_{(s)}$ to produce $Fe_{(s)}$.

$$2 \text{ Fe}_{(s)} + \frac{3}{2} \text{ O}_2_{(g)} \rightarrow \text{ Fe}_2\text{O}_3_{(s)}$$

$$\Delta H_f^o = -822 \text{ kJ/mol}$$

$$C_{(s)} + \text{ O}_2_{(g)} \rightarrow \text{ CO}_2_{(s)}$$

$$\Delta H_f^o = -394 \text{ kJ/mol}$$
Ex3) Enthalpy of Formation

Ex3) Use the enthalpy of formation values below to determine $\Delta H_{\text{rxn}}$ for the reaction between $C_{(s)}$ and $\text{Fe}_2\text{O}_3_{(s)}$ to produce $\text{Fe}_{(s)}$.

\[ \text{2 Fe}_{(s)} + \frac{3}{2} \text{O}_2\,(g) \rightarrow \text{Fe}_2\text{O}_3\,(s) \quad \Delta H_f^o = -822 \text{ kJ/mol} \]
\[ \text{C}_{(s)} + \text{O}_2\,(g) \rightarrow \text{CO}_2\,(s) \quad \Delta H_f^o = -394 \text{ kJ/mol} \]

\[
\begin{array}{c|c|c}
-2(\text{rxn 1}) & 2 \text{Fe}_2\text{O}_3\,(s) \rightarrow 4 \text{Fe}_{(s)} + 3 \text{O}_2\,(g) & \Delta H = -2(-822 \text{ kJ}) \\
\end{array}
\]
Ex3) Use the enthalpy of formation values below to determine $\Delta H_{\text{rxn}}$ for the reaction between $C_{(s)}$ and $\text{Fe}_2\text{O}_3_{(s)}$ to produce $\text{Fe}_{(s)}$.

\[
\begin{align*}
2 \text{Fe}_{(s)} + \frac{3}{2} \text{O}_2(g) & \rightarrow \text{Fe}_2\text{O}_3_{(s)} \quad \Delta H_f^o = -822 \text{ kJ/mol} \\
\text{C}_{(s)} + \text{O}_2(g) & \rightarrow \text{CO}_2(s) \quad \Delta H_f^o = -394 \text{ kJ/mol}
\end{align*}
\]

\[
\begin{array}{ccc}
-2 & (\text{rxn 1}) & 2 \text{Fe}_2\text{O}_3_{(s)} \rightarrow 4 \text{Fe}_{(s)} + 3 \text{O}_2(g) \quad \Delta H = -2(-822 \text{ kJ}) \\
3 & (\text{rxn 2}) & 3 \text{C}_{(s)} + 3 \text{O}_2(g) \rightarrow 3 \text{CO}_2(s) \quad \Delta H = 3(-394 \text{ kJ})
\end{array}
\]

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### Ex3) Enthalpy of Formation

Ex3) Use the enthalpy of formation values below to determine $\Delta H_{\text{rxn}}$ for the reaction between $C_{(s)}$ and $\text{Fe}_2\text{O}_3_{(s)}$ to produce $\text{Fe}_{(s)}$.

\[
\begin{align*}
2 \text{Fe}_{(s)} + \frac{3}{2} \text{O}_2_{(g)} & \rightarrow \text{Fe}_2\text{O}_3_{(s)} \quad \Delta H_f^o = -822 \text{ kJ/mol} \\
\text{C}_{(s)} + \text{O}_2_{(g)} & \rightarrow \text{CO}_2_{(s)} \quad \Delta H_f^o = -394 \text{ kJ/mol}
\end{align*}
\]

\[
\begin{array}{c|c|c}
\text{rxn 1} & \text{rxn 2} & \Delta H \\
\hline
-2 & 2 \text{Fe}_2\text{O}_3_{(s)} & 4 \text{Fe}_{(s)} + 3 \text{O}_2_{(g)} \quad \Delta H = -2(-822 \text{ kJ}) \\
3 & 3 \text{C}_{(s)} + 3 \text{O}_2_{(g)} & 3 \text{CO}_2_{(s)} \quad \Delta H = 3(-394 \text{ kJ}) \\
\end{array}
\]

\[
\begin{align*}
2 \text{Fe}_2\text{O}_3_{(s)} + 3 \text{C}_{(s)} & \rightarrow 4 \text{Fe}_{(s)} + 3 \text{CO}_2_{(s)} \quad \Delta H = +462 \text{ kJ}
\end{align*}
\]
Ex) Enthalpy and Stoichiometry

\[ 2 \text{H}_2\text{O}_2(aq) \rightarrow \text{O}_2(g) + 2 \text{H}_2\text{O}(l) \quad \Delta H = -186 \text{ kJ} \]

Ex) How much heat is released when 3.56 g \(\text{H}_2\text{O}_2(aq)\) decomposes into water and oxygen gas?

\[
3.56 \text{ g} \quad \text{H}_2\text{O}_2 \times \frac{1 \text{ mol} \quad \text{H}_2\text{O}_2}{34.02 \text{ g} \quad \text{H}_2\text{O}_2} \times \frac{-186 \text{ kJ}}{2 \text{ mol} \quad \text{H}_2\text{O}_2} = -9.73 \text{ kJ}
\]

There are 2 moles \(\text{H}_2\text{O}_2\) in the balanced equation.