Standard Enthalpies of Formation

$\triangle H_f$

* The standard enthalpy of formation is the quantity of energy associated with the formation of one mole of a substance from its elements in their standard states (at SATP)

Symbol: $\triangle H_f$

Unit: kJ/mol

Calculating ΔH_f

* According to Hess's Law, the enthalpies of known equations may be used to calculate the enthalpy of an unknown reaction.
  * $H_2O_{(g)} + C_{(s)} \longrightarrow H_{2(g)} + CO_{(g)} \quad \Delta H = ?$
* Use the formation equations for each of the products and reactants to create the target equation.

Example

* $H_2O_{(g)} + C_{(s)} \longrightarrow H_{2(g)} + CO_{(g)}$ $\Delta H = ?$
  * Consider:
  * $H_{2(g)} + 1/2\ O_{2(g)} \longrightarrow H_2O_{(g)}$ $\Delta H_f(H_2O) = -241.8kJ/mol$
  * $C_{(s)} + 1/2\ O_{2(g)} \longrightarrow CO_{(g)}$ $\Delta H_f(CO) = -110.5kJ/mol$

Equation 1

$\boxed{H_2O_{(g)}} + C_{(s)} \rightarrow H_{2(g)} + CO_{(g)}$

$H_{2(g)} + 1/2\ O_{2(g)} \rightarrow \boxed{H_2O_{(g)}}$ $\Delta H_f(H_2O) = -241.8 kJ/mol$

* H_2O on wrong side, reverse and change sign on $\triangle$ H

$H_2O_{(g)} \rightarrow H_{2(g)} + 1/2\ O_{2(g)}$ $\Delta H_f(H_2O) = -1(-241.8) kJ/mol$

Equation 1

$\boxed{H_2O_{(g)}} + C_{(s)} \rightarrow H_{2(g)} + CO_{(g)}$

$H_{2(g)} + 1/2\ O_{2(g)} \rightarrow \boxed{H_2O_{(g)}}$ $\Delta H_f(H_2O) = -241.8 kJ/mol$

* H_2O on wrong side, reverse and change sign on △ H

$H_2O_{(g)} \rightarrow H_{2(g)} + 1/2\ O_{2(g)}$ $\Delta H_f(H_2O) = -1(-241.8) kJ/mol$

Equation 2

$H_2O_{(g)} + C_{(s)} \rightarrow H_{2(g)} + CO_{(g)}$

$C_{(s)} + 1/2\ O_{2(g)} \rightarrow CO_{(g)}$ $\Delta H_f(CO) = -110.5 kJ/mol$

* Coefficients and sides match, △ H stays the same

$C_{(s)} + 1/2\ O_{2(g)} \rightarrow CO_{(g)}$ $\Delta H_f(CO) = -110.5 kJ/mol$

Equation 1

$\boxed{H_2O_{(g)}} + C_{(s)} \rightarrow H_{2(g)} + CO_{(g)}$

$H_{2(g)} + 1/2\ O_{2(g)} \rightarrow \boxed{H_2O_{(g)}}$ $\Delta H_f(H_2O) = -241.8 kJ/mol$

* H_2O on wrong side, reverse and change sign on ΔH

$H_2O_{(g)} \rightarrow H_{2(g)} + 1/2\ O_{2(g)}$ $\Delta H_f(H_2O) = -1(-241.8) kJ/mol$

Equation 2

$H_2O_{(g)} + C_{(s)} \rightarrow H_{2(g)} + CO_{(g)}$

$C_{(s)} + 1/2\ O_{2(g)} \rightarrow CO_{(g)}$ $\Delta H_f(CO) = -110.5 kJ/mol$

* Coefficients and sides match, ΔH stays the same

$C_{(s)} + 1/2\ O_{2(g)} \rightarrow CO_{(g)}$ $\Delta H_f(CO) = -110.5 kJ/mol$

Combine Equations

$H_2O_{(g)} \rightarrow H_{2(g)} + \cancel{1/2\ O_{2(g)}}$ $\Delta H_f(H_2O) = -1(-241.8) kJ/mol$

$C_{(s)} + \cancel{1/2\ O_{2(g)}} \rightarrow CO_{(g)}$ $\Delta H_f(CO) = -110.5 kJ/mol$

$H_2O_{(g)} + C_{(s)} \rightarrow H_{2(g)} + CO_{(g)}$ $\Delta H_f = 131.3 kJ$

Using Hess's Law

* When all enthalpies of formation are known, the enthalpy of a reaction is equal to the sum of the enthalpies of formation of the products minus the sum of the enthalpies of formation of the reactants.

$$\Delta H_{rxn} = \Sigma\, n\, \Delta H_{f\ (products)} - \Sigma\, n\, \Delta H_{f\ (reactants)}$$

n = the amount in moles of each product or reactant

Example

* Calculate the heat of combustion of methanol.

$CH_3OH_{(l)} + 3/2\ O_{2\,(g)} \rightarrow CO_{2\,(g)} + 2\ H_2O_{(g)}$ $\Delta H^\circ comb = ?$

$\triangle H_{CH_3OH} = -239.9$ kJ

$\triangle H_{O_2} = 0$ kJ

$\triangle H_{CO_2} = -393.5$ kJ

$\triangle H_{H_2O} = -241.8$ kJ

All atoms in their standard states have $H_f = 0$

Solution

* $\Delta H_{rxn} = \Sigma\, n\, \Delta H_{f\,(products)} - \Sigma\, n\, \Delta H_{f\,(reactants)}$

$CH_3OH_{(l)} + 3/2\ O_{2\,(g)} \longrightarrow CO_{2\,(g)} + 2\ H_2O_{(g)}$ $\Delta H^{o}comb = ?$

ΔH_{rxn} = [(1)(CO_2) + (2)(H_2O)] - [(1)(CH_3OH) + (3/2)(O_2)]
ΔH_{rxn} = [(1)(-393.5) + (2)(-241.8)] - [(1)(-239.2) + (3/2)(0)]
ΔH_{rxn} = [-877.1] - [-239.2]
ΔH_{rxn} = -637.9 kJ

$\triangle H_{CH3OH}$ = -239.9 kJ
$\triangle H_{O2}$ = 0 kJ
$\triangle H_{CO2}$ = -393.5 kJ
$\triangle H_{H2O}$ = -241.8 kJ

The $\triangle H_{comb}$ is -637.9 kJ.

Homework

* Pg 323 #51-56
* Pg 324 # 1,4,6,11