THERMOCHEMISTRY

Hess's Law

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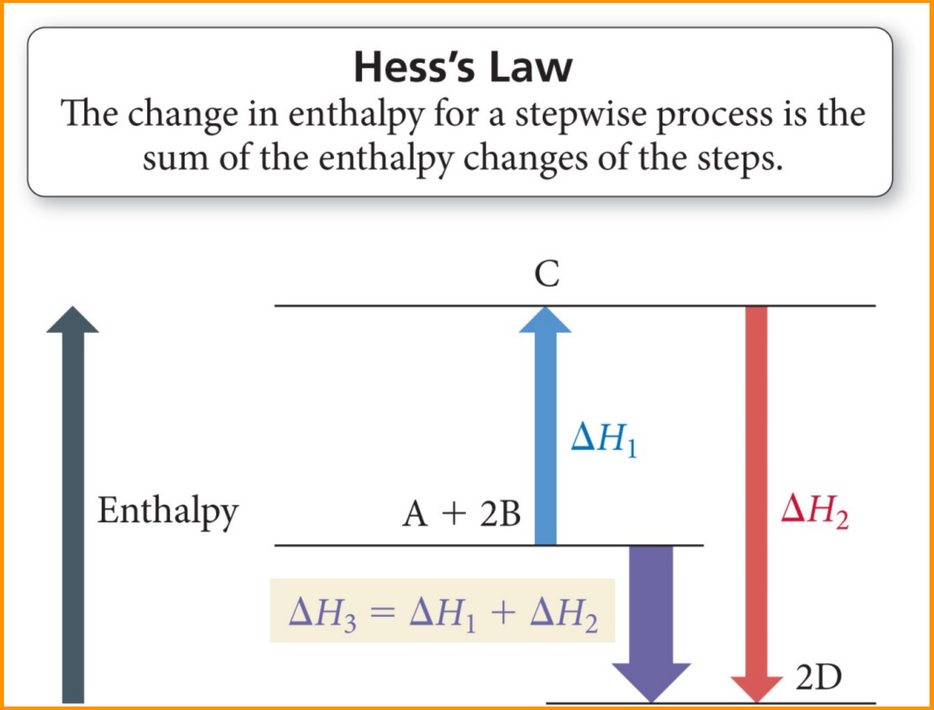
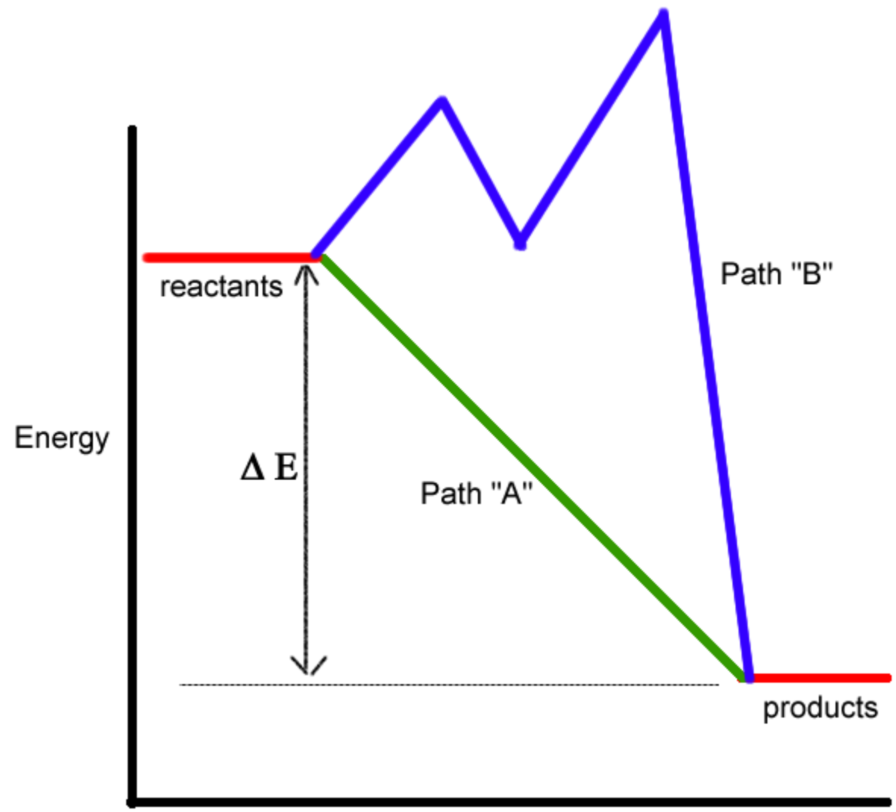
"In going from a particular set of reactants to a particular set of products, the change in enthalpy is the same whether the reaction takes place in one step or a series of steps."



Hess's Law



Relationships Involving ΔH_{rxn}

- **When reaction is multiplied by a factor, ΔH_{rxn} is multiplied by that factor.**
  - Because ΔH_{rxn} is extensive,
  - $C(s) + O_2(g) \rightarrow CO_2(g)$
    - **ΔH = −393.5 kJ**
  - $2\ C(s) + 2\ O_2(g) \rightarrow 2\ CO_2(g)$
    - **ΔH = 2(−393.5 kJ) = −787.0 kJ.**

- **If a reaction is reversed, then the sign of ΔH is changed.**

$$CO_2(g) \rightarrow C(s) + O_2(g) \qquad \Delta H = +393.5\ \text{kJ}$$

Standard Conditions

- The standard state is the state of a material at a defined set of conditions.
  - Pure gas at exactly 1 atm pressure
  - Pure solid or liquid in its most stable form at exactly 1 atm pressure and temperature of interest
    - Usually 25 °C
  - Substance in a solution with concentration 1 M

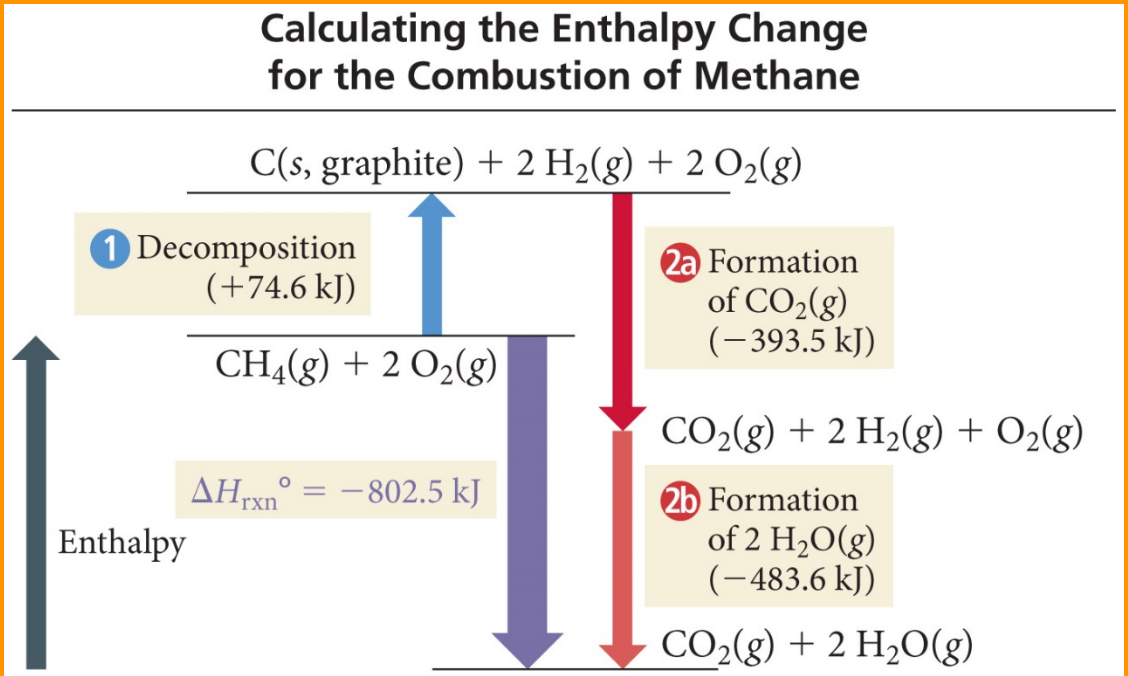
Standard Conditions

- The **standard enthalpy change, $\Delta H°$,** is the enthalpy change when all reactants and products are in their standard states.

Standard Conditions

- **The standard enthalpy of formation, ΔH_f°, is the enthalpy change for the reaction forming 1 mole of a pure compound from its constituent elements.**
  - **The elements must be in their standard states.**
  - **The ΔH_f° for a <u>pure element in its standard state</u> = 0 kJ/mol.**

$CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$

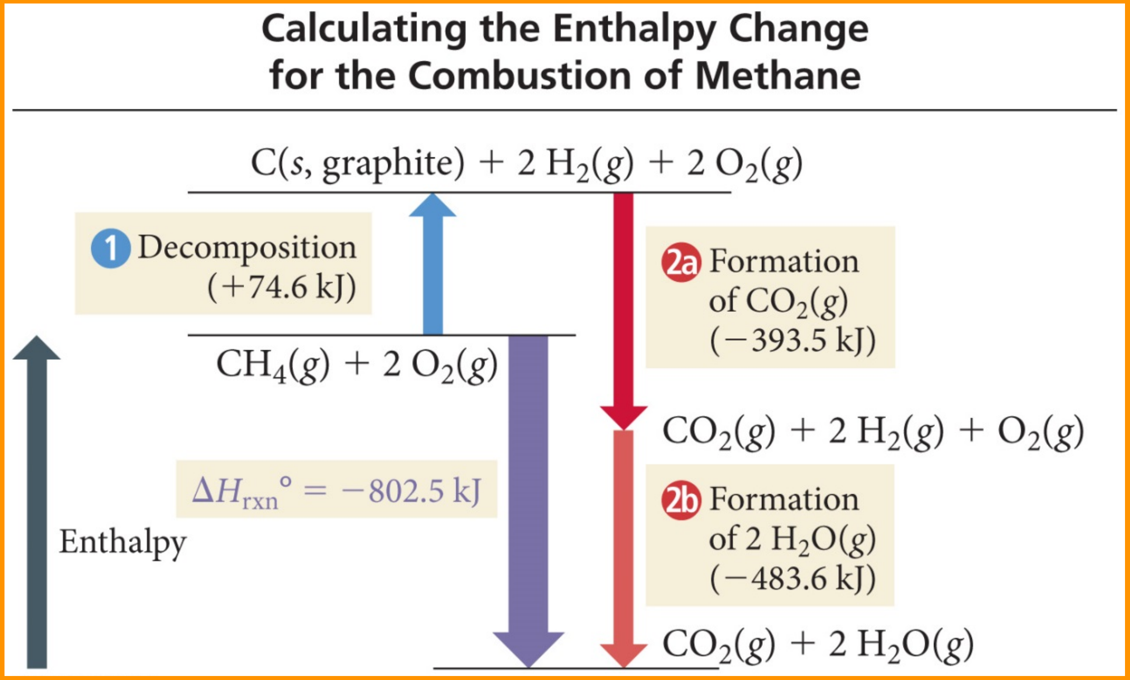


$C(s,\ \text{graphite}) + 2\ H_2(g) \rightarrow CH_4(g)$ $\quad \Delta H_f^\circ = -74.6$ kJ/mol CH_4

$C(s,\ \text{graphite}) + O_2(g) \rightarrow CO_2(g)$ $\quad \Delta H_f^\circ = -393.5$ kJ/mol CO_2

$H_2(g) + \frac{1}{2}\ O_2(g) \rightarrow H_2O(g)$ $\quad \Delta H_f^\circ = -241.8$ kJ/mol H_2O

$CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$



$CH_4(g) \rightarrow$ ~~C(s, graphite)~~ + ~~$2\ H_2(g)$~~ $\Delta H° = +74.6$ kJ

~~C(s, graphite)~~ $+ O_2(g) \rightarrow CO_2(g)$ $\Delta H_f° = -393.5$ kJ/mol CO

~~$2\ H_2(g)$~~ $+ O_2(g) \rightarrow 2\ H_2O(g)$ $\Delta H° = -483.6$ kJ

$CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g)$ $\Delta H° = -802.5$ kJ

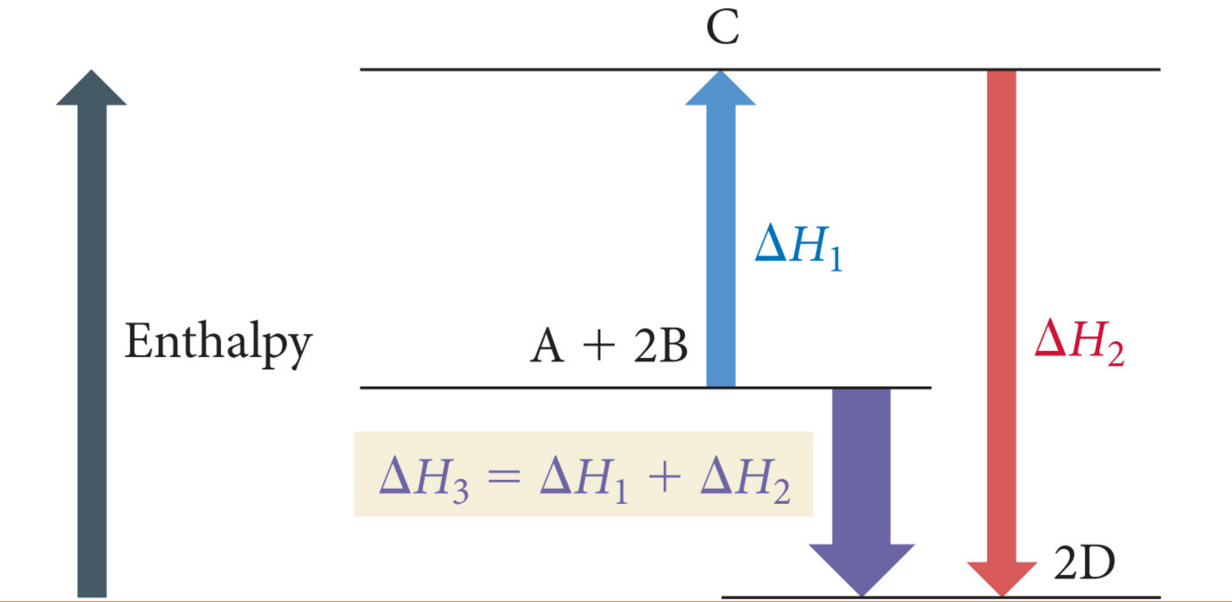
Use Hess's Law to determine Δ*H* for the following target reaction.

$\frac{1}{2}\ N_2\ (g) + \frac{1}{2}\ O_2\ (g) \rightarrow NO\ (g)$ $\Delta H = 90.3$ kJ

$NO\ (g) + \frac{1}{2}\ Cl_2\ (g) \rightarrow NOCl\ (g)$ $\Delta H = -38.6$ kJ

$2\ NOCl\ (g) \rightarrow N_2\ (g) + O_2\ (g) + Cl_2\ (g)$ $\Delta H = ?$

a) –51.7 kJ
b) 51.7 kJ
c) –103.4 kJ
d) 103.4 kJ
e) 142.0 kJ



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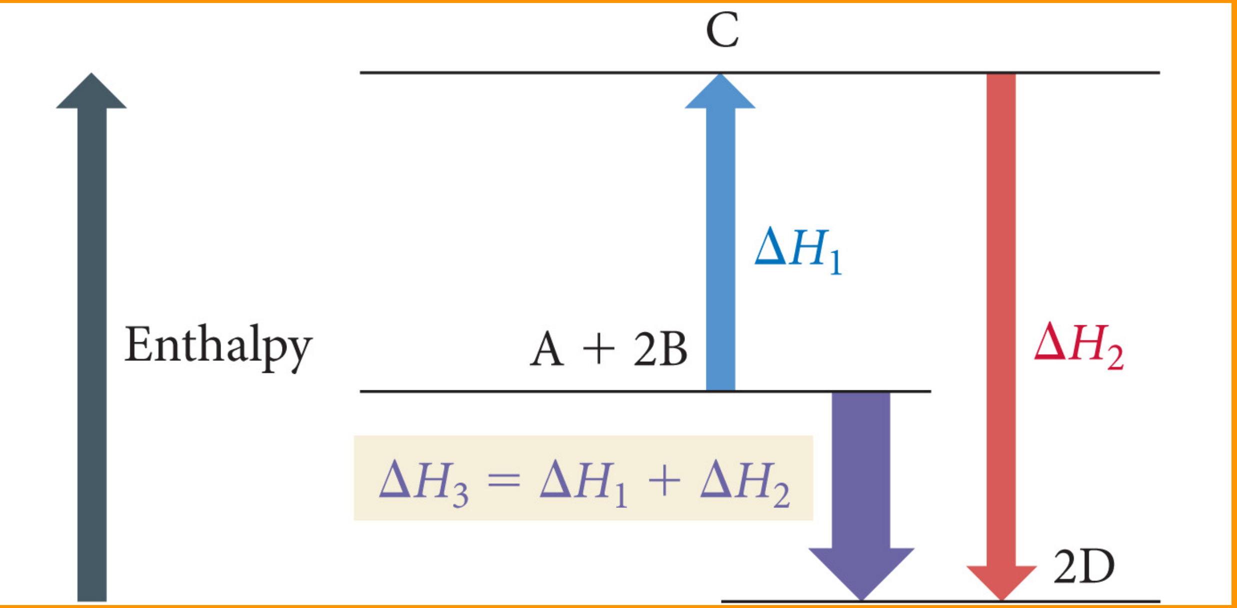

$FeO (s) + CO (g) \rightarrow Fe (s) + CO_2 (g)$

Calculate the standard enthalpy change for this Rx from these Rx's of Iron Oxides w/ CO:

[1] $3Fe_2O_3 + CO (g) \rightarrow 2Fe_3O_4 + CO_2 (g)$ $\Delta H° = -47$ kJ

[2] $Fe_2O_3 + 3CO (g) \rightarrow 2Fe (s) + 3CO_2 (g)$ $\Delta H° = -25$ kJ

[3] $Fe_3O_4 + CO (g) \rightarrow 3FeO (s) + CO_2 (g)$ $\Delta H° = 19$ kJ

A. -53 kJ

B. -3 kJ

C. -41 kJ

D. 22 kJ

E. -11 kJ

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