Chemical Potential energy - Enthalpy

ΔH change in enthalpy

Choice of zero potential energy is important.

- every element in its natural state at standard conditions (25°C and one atmosphere of pressure) is given a zero potential energy
- therefore there are over 100 zero points in thermodynamics.

In order to determine the potential energy of any compound, simply compair its energy to that of its constituent elements.

eg. $C(s) + O_2(g) \longrightarrow CO_2(g)$

$\Delta H = -393.5 kJ$

potential energy $\quad 0 + 0 \longrightarrow -393.5 \text{ kJ}$

↖assigned zero point ↑ $\quad$ ↑comes from measurments

this is an example of a <u>formation</u> reaction (formation reaction is when compounds are formed directly from its constituent elements.)

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eg Acetone  $\Delta H^\circ_{(CH_3)_2CO}$

↖heat of formation

enthalpy potential energy $\quad$ formula subscript present

This means two things

→ using element zero points, this value <u>is</u> the potential energy of acetone

→ $$\underbrace{3C(s) + 3H_2(g) + \tfrac{1}{2}O_2(g)}_{\text{elements only!}} \longrightarrow \underbrace{1(CH_3)_2CO(l)}_{\text{one mole of product}} \quad \underbrace{\Delta H = -248.1kJ}_{\text{heat of reaction}}$$

↓ change in enthalpy

← equivalent to the potential energy of acetone because the elements state at zero

eg2 $\Delta H^\circ_{AgCl(s)} = -127.0\,kJ$

$$\therefore \quad Ag(s) + \tfrac{1}{2}Cl_2(g) \longrightarrow AgCl(s) \quad \Delta H = -127.0kJ$$

Heat Summation Rule - Hess' law abbreviated

For any reaction

$$\Delta H = \underbrace{\left[\underset{\text{"the sum of"}}{\Sigma}\ \underset{\text{potential energy}}{\Delta H^\circ}\text{products}\right]}_{\text{final}} - \underbrace{\left[\Sigma\, \Delta H^\circ \text{reactants}\right]}_{\text{initial}}$$

eg. the combustion of octane

$$C_8H_{18}(l) + \tfrac{25}{2}O_2(g) \longrightarrow 8CO_2(g) + 9H_2O(l)$$

$$\Delta H = ? \quad \left[8\Delta H^\circ_{CO_2(g)} + 9\Delta H^\circ_{H_2O(l)}\right] - \left[\Delta H^\circ_{C_8H_8(l)} + \tfrac{25}{2}\Delta H^\circ_{O_2(g)}\right]$$

$$\Delta H = \left[8(-393.5kJ) + 9(-285.8kJ)\right] - \left[(-250.1kJ) + \tfrac{25}{2}(0)\right]$$

$$\Delta H = -5470.1kJ$$

eg2. alternate use of the heat summation rule.

The heat of combustion for butone is -2877.4 kJ
given: $\Delta H^\circ_{CO_2(g)} = -393.5\text{kJ}$ and $\Delta H^\circ_{H_2O(l)}$ -285.8kJ
find the heat of formation for butane

$$C_4H_{10}\,(l) + \tfrac{13}{2}O_2\,(g) \longrightarrow 4\,CO_2\,(g) + 5H_2O(l)$$

$\Delta H^\circ = -2877.4\,\text{kJ}$ (change in enthalpy

$$\Delta H = \left[4\Delta H^\circ_{CO_2(g)} + 5\Delta H^\circ_{H_2O(l)}\right] - \left[\Delta H^\circ_{C_4H_{10}(l)} + \tfrac{13}{2}\Delta H^\circ_{O_2(g)}\right]$$

$$-2877.4\text{kJ} = \left[4(-393.5\text{kJ}) + 5(-285.8\text{kJ})\right] - \left[\Delta H^\circ_{C_4H_{10}(l)} + \tfrac{13}{2}(0)\right]$$

$$\Delta H^\circ_{C_4H_{10}(l)} = -125.6\,\text{kJ}$$