Reference Energy State

Reference energy state - elements are defined as the reference point at which the potential energy is shown to be zero.
Therefore: E_p of $H_{2(g)} = 0$ kJ

Ex. $H_{2(g)} + 1/2O_{2(g)} \longrightarrow H_2O_{(g)}$ $\Delta H_f = -285.8$ kJ

*allows us to describe the enthalpy change for a formation reaction from zero to a final value

Thermal Stability

Thermal Stability - the tendency of a compound to resist decomposition when heated.

- the more endothermic the simple decomposition (sd), the more stable the compound.

Ex. $H^{o}_{(sd)} = + 280.7$ kJ/mol　　　$H^{o}_{(sd)} = + 577.6$ kJ/mol
SnO　　　SnO_2

Therefore is more stable.

*Normally not given the H_{sd}, but given the H_f

Which is more stable, ammonia or butane?

NH_3　　　C_4H_{10}

$H_f = -45.9$ kJ/mol　　　$H_f = -125.6$ kJ/mol

$H_{sd} = 45.9$ kJ/mol　　　$H_{sd} = 125.6$ kJ/mol

$CO_2 \longrightarrow H_f = -393.5$ kJ/mol

$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$ $\Delta H_f = -393.5$ kJ

$CO_{2(g)} \rightarrow C_{(s)} + O_{2(g)}$ $\Delta H_{SD} = 393.5$ kJ

Predicting ΔH_r Using Formation Reactions

The Standard Enthalpy Change (ΔH°_r) for a reaction can be found by writing the formation equation and corresponding standard enthalpy change for each compound in the given equation and then applying Hess' Law.

Ex. $CaO_{(s)} + H_2O_{(l)} \longrightarrow Ca(OH)_{2(s)}$ $\Delta H_r = ?$

Step 1: Write formation equations (with standard enthalpy change) each compound in the given equation.

① $Ca_{(s)} + \frac{1}{2}O_{2(g)} \longrightarrow CaO_{(s)}$ $\Delta H_r = -634.9\ kJ$

② $H_{2(g)} + \frac{1}{2}O_{2(g)} \longrightarrow H_2O_{(l)}$ $\Delta H_r = -285.8\ kJ$

③ $Ca_{(s)} + O_{2(g)} + H_{2(g)} \longrightarrow Ca(OH)_{2(s)}$ $\Delta H_r = -986.1\ kJ$

Step 2: Apply Hess' Law

Rev ①

④ $CaO_{(s)} \longrightarrow Ca_{(s)} + \frac{1}{2}O_{2(g)}$ $\Delta H_r = 634.9\ kJ$

Rev ②

⑤ $H_2O_{(l)} \longrightarrow H_{2(g)} + \frac{1}{2}O_{2(g)}$ $\Delta H_r = 285.8\ kJ$

③ + ④ + ⑤

$CaO_{(s)} + H_2O_{(l)} \longrightarrow Ca(OH)_{2(s)}$ $\boxed{\Delta H_r = -65.4\ kJ}$

Enthalpies of Formation to Predict ΔH_r

$$\Delta H_r = \Delta H_f + (-\Delta H_f) + (-\Delta H_f)$$

$Ca(OH)_2$ CaO H_2O

$$\Delta H_r = \Delta H_f - (\Delta H_f + \Delta H_f)$$

$Ca(OH)_2$ CaO H_2O

$$\Delta H_r = \Delta H_{fp} - \Delta H_{fr}$$

products reactants

$$\Delta H_r = \Sigma n H_{fp} - \Sigma n H_{fr}$$

knowing that $\Delta H = nH$

Ex. What is the standard molar enthalpy of combustion of methane fuel?

$$CH_{4(g)} + 2O_{2(g)} \longrightarrow CO_{2(g)} + 2H_2O_{(g)}$$

$$\Delta H_r = \sum nH_{fp} - \sum nH_{fr}$$

CO_2

$$\Delta H_r = \left[(1\,mol)\left(-393.5\,\frac{kJ}{mol}\right) + (2\,mol)\left(-241.8\,\frac{kJ}{mol}\right)\right] - \left[(1\,mol)\left(-74.4\,\frac{kJ}{mol}\right) + (2\,mol)\left(0\,\frac{kJ}{mol}\right)\right]$$