Enthalpy \[ H = E + PV \]

= heat released/absorbed at constant pressure

Thermochemical Eq =

chemical Eq + \( \Delta H \)

\[ 2H_2(g) + O_2(g) \rightarrow 2H_2O(l) \quad \Delta H_{298}^\circ = -571.7 \text{ KJ} \]

(exothermic)

\(-571.7 \text{ KJ/mol } H_2\)

\(-571.7 \text{ KJ/mol } O_2\)

\(-571.7 \text{ KJ/mol } H_2O\)
Physical changes also:

\[ H_2O (s) \rightarrow H_2O (l) \quad \Delta \text{H}_{273} = +6.01 \text{kJ} \text{ endothermic} \]

Stoichiometric calculations include enthalpy:

\[ C \text{ (graphite)} + O_2(g) \rightarrow CO_2 (g) \]

\[ \Delta H = -393.5 \text{ KJ} \]

How much heat is released upon combustion of 1 Kg of graphite?

\[ 1000 \text{ g} \times \frac{\text{mol}}{12.01 \text{ g}} \times \frac{393.5 \text{ KJ}}{\text{mol c}} \]

\[ = -3.28 \times 10^4 \text{ KJ} \]
\[ \Delta H^\circ : \text{standard enthalpy change} \]
\[ P = 1 \text{ atm} \]
\( (\text{if } T \text{ not specified, we assume } 25^\circ \text{C}) \)

Standard enthalpy of combustion:

- Enthalpy of burning 1 mol of a substance in \( O_2 \) at 1 atm

\[
\text{CH}_4 (g) + 2O_2 (g) \rightarrow CO_2 (g) + 2H_2O (l) \]
\[ \Delta H^\circ = -890.4 \text{ KJ} \]

St. enthalpy of combustion: \( \nabla \)
\[ 2 \text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{H}_2\text{O}(\ell) \]

\[ \Delta H^\circ = -571.6 \text{ KJ} \]

St. enthalpy of combustion of H\(_2\): \[ -571.6 \text{ KJ} = \frac{\ldots \text{ KJ}}{2 \text{ mol H}_2} \]

Example

\[ \Delta H^\circ \text{comb (methane)} = -890.4 \text{ KJ} \]

\[ \text{(propane)} = -2219.9 \text{ KJ} \]

Compare the heats given off by burning 1 g of methane or propane.
\[
\text{CH}_4: \frac{1.00 \text{g}}{16.04 \text{g/mol}} \times 890.4 \frac{\text{kJ}}{\text{mol}} = 55.5 \text{kJ}
\]

\[
\text{C}_3\text{H}_8: \frac{1.00 \text{g}}{44.10 \text{g/mol}} \times 2219.9 \frac{\text{kJ}}{\text{mol}} = 50.3 \text{kJ}
\]

---

Standard Enthalpy of formation

\( = H \) of formation of 1 mol of the substance from the most stable form of its elements

e.g. C: most stable: graphite

\( \text{C (graphite)} \rightarrow \text{C (diamond)} \)

\( \Delta H^\circ = +1.90 \text{kJ} \)
\[ \Delta H_f^0 \]

\[ 2C_{(\text{graphite})} + O_2(g) \rightarrow 2CO(g) \]

\[ \Delta H^0 = -221.0 \text{ KJ} \]

\[ \Delta H_f^0 (CO) = -\frac{221.0 \text{ KJ}}{2 \text{ mol}} \]

\[ = -110.5 \text{ KJ/mol} \]

These data are tabulated.

Example:

\[ \Delta H_f^0 (NH_3(g)) = -46.1 \text{ KJ/mol} \]

How many KJ are given off when 10.0 g \( N_2 \) react with excess \( N_2(g) + 3H_2(g) \rightarrow 2NH_3(g) \)

\[ \Delta H^0 = -92.2 \text{ KJ} \]
\[ 10.0 \text{g N}_2 \frac{\text{mol}}{28.01 \text{g mol N}_2} \times 92.2 \text{KJ mol}^{-1} \]

\[ = 32.9 \text{KJ} \]

\[ E: \text{internal energy} \]

\[ H = E + PV \]

"State functions"

(completely determined by \( T, P \) of the system)

(they do not depend on the history of the system)

\( q, w \): not state functions
Hess's law

The $\Delta H$ for a reaction is the same whether the reaction takes place in one or several steps.

e.g. $C \rightarrow CO_2$

one step:

$C + O_2 \rightarrow CO_2 \quad \Delta H^0 = -393.5 \text{ KJ}$

two steps

$C + 1/2 O_2 \rightarrow CO \quad \Delta H^0 = -110.5$

$CO + 1/2 O_2 \rightarrow CO_2 \quad \Delta H^0 = -283.0 \text{ KJ}$

sum $-393.5$

Total enthalpy change is the same.
So: we can add up the thermochemical equations

\[
\begin{align*}
C + \frac{1}{2}O_2 & \rightarrow CO \quad \Delta H^\circ = -110.5 \\
CO + \frac{1}{2}O_2 & \rightarrow CO_2 \quad \Delta H^\circ = -283.0
\end{align*}
\]

\[
C + CO + O_2 \rightarrow CO + CO_2 \quad \Delta H^\circ = -393.5
\]

Example: Calculate \( \Delta H \) for combustion of 1 mol C\((g)\) given that:

\[
C(\text{graph}) \rightarrow C(\text{g}) \quad \Delta H^\circ = 716.7 \text{ KJ}
\]

(sublimation of C)

\[
C(\text{graph}) + O_2 \rightarrow CO_2(\text{g}) \quad \Delta H^\circ = -393.5 \text{ KJ}
\]
When you flip a reaction, change the sign of ΔH.
When you multiply a reaction by a number, multiply also ΔH by same number.

We want ΔH for this:

\[ \text{C}(g) + \text{O}_2 \rightarrow \text{CO}_2(g) \]

Flip the 1st rxn and add to the 2nd:

\[ \text{C}(g) \rightarrow \text{C(graphite)} \quad \Delta H^\circ = -716.7 \text{ kJ} \]
\[ \text{C(graphite)} + \text{O}_2 \rightarrow \text{CO}_2 \quad \Delta H^\circ = -393.5 \text{ kJ} \]
\[ \text{C}(g) + \text{O}_2 \rightarrow \text{CO}_2 \quad \Delta H^\circ = -1110.2 \text{ kJ} \]
We can get the $\Delta H^\circ$ if we know all the $\Delta H_f^\circ$ of the reactants & products

Simple formula:

$\Delta H^\circ = \text{Sum of } \Delta H_f^\circ \text{ of products} - \text{Sum of } \Delta H_f^\circ \text{ of reactants}$

[this can be proven using Hess’s law: write down the reaction of formation of each component, then combine them]
example: \[ 2 \text{H}_2\text{S}(g) + \text{SO}_2(g) \rightarrow 3 \text{S}(s) + 2\text{H}_2\text{O}(l) \]

Appendix D: \( \Delta H^\circ_f \)

- \( \Delta H^\circ_f (\text{H}_2\text{O}(l)) = -285.8 \text{ kJ} \)
- \( \Delta H^\circ_f (\text{H}_2\text{S}(g)) = -20.15 \text{ kJ} \)
- \( \Delta H^\circ_f (\text{SO}_2(g)) = -296.4 \text{ kJ} \)
- \( \Delta H^\circ_f (\text{S}(s)) = 0 \)

\[ \Delta H^\circ = 2 \times \Delta H^\circ_f (\text{H}_2\text{O}) - 2 \times \Delta H^\circ_f (\text{H}_2\text{S}) - \Delta H^\circ_f (\text{SO}_2) = -234.9 \text{ kJ} \]

For elements in their most stable form: \( \Delta H^\circ_f = 0 \)
How we measure $\Delta H$ experimentally.

"Calorimetry"

Constant Volume

$(q = \Delta E)$

Constant Pressure

$(q = \Delta H)$

more common

reaction
water

Temperature change

From this we extract $q$
Constant P calorimeter:

\[ q \text{ measurement} \]

\[ \Delta T \]

Relate \( q \) to \( \Delta T \):

Heat capacity of an object \((C)\) : amount of heat required to raise the \( T \) by 1 K (=1°C)

\[ q = C \cdot \Delta T \]
$G'$ depends on the size/mass of the object
To eliminate the size dependence
$c$ (molar heat capacity): for $1 \text{ mol}$
$s$ (specific heat): for $1 \text{ g}$

\[
G' = N \cdot c = m \cdot s
\]

$J/K, \quad J/\text{mol} \cdot K, \quad J/\text{mol} \cdot g$