

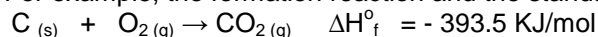
5.5: STANDARD ENTHALPIES OF FORMATION

$[\Delta H_f^\circ]$

A third method used to determine enthalpies of reaction is the **standard enthalpy of formation**.

Others are calorimetry and Hess's law.

For example, the formation reaction and the standard enthalpy of formation of carbon dioxide are:



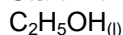
Steps for writing Formation Equations

1. Write one mole of product in the state that has been specified
2. Write the reactant elements in their standard states
3. Choose equation coefficients for the reactants to give a balanced equation yielding one mole of product.

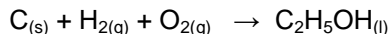
Example 1:

Write the equation for formation of liquid ethanol

1. Start with one mole of product



2. Write the reactant elements in their standard states



3. Balance the equation to yield one mole of product.

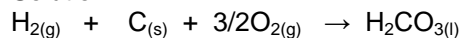


Example 2:

What is the formation equation for liquid carbonic acid?



Solution:



Example 3: -



ΔH_f° is always for 1 mole of the product being formed

In general $\Delta H_f^\circ = [\Delta H_f^\circ \text{ compounds formed}] - [\Delta H^\circ \text{ element}_1 + \Delta H^\circ \text{ element}_2 + \dots + \Delta H_n^\circ]$.

ENTHALPY OF REACTION ΔH_r° :-

Is equal to the difference between the sum of enthalpies of the formation of Products and that of the reactants.

$\Delta H_r^\circ = (\text{enthalpy of Product} - \text{enthalpy of Reactions})$

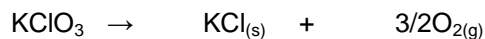
$$\Delta H_r^\circ = \sum \Delta H_f^\circ \text{ Products} - \sum \Delta H_f^\circ \text{ Reactants}$$

i.e. $aA + bB \rightarrow cC + dD$

$$\Delta H_r^\circ = [c(\Delta H_f^\circ C) + (d(\Delta H_f^\circ D))] - [a(\Delta H_f^\circ A) + (b(\Delta H_f^\circ B))]$$

Try this example: -

Calculate the enthalpy change, ΔH°_r , for the decomposition of potassium Chlorate.



Using Standard Enthalpies of formation ΔH°_f for elements

The standard enthalpy of formation of an element already in its standard state is zero.

Example, the standard enthalpies of formation of $\text{Fe}_{(s)}$, $\text{O}_{2(g)}$, and $\text{Br}_{2(l)}$ are all zero

However, standard molar enthalpies of formation help us to compare the stabilities of substances.e.g.,

$$\Delta H^\circ_{f(\text{graphite})} = 0 \text{ KJ/mol}$$

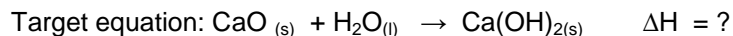
But diamond made from the same carbon is slightly less stable but with greater potential energy than graphite



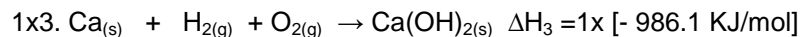
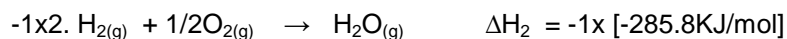
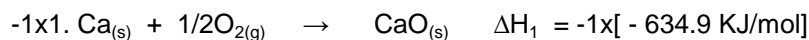
Using Enthalpies of formation to find ΔH

$$\Delta H = \sum n \Delta H^\circ_{(\text{products})} - \sum n \Delta H^\circ_{(\text{reactants})}$$

Example: What is the thermochemical equation for the reaction of lime (Calcium oxide) and water?



Use the tables of standard enthalpies in appendix C4 in the text and Hess's Law to find the enthalpy change for the target equation. (see p.333)



Multistep Energy Calculations Using Standard Enthalpies of Formation

This involves the use of many of the problem-solving skills that you have acquired

1. $\Delta H = q$
2. $\Delta H = n\Delta H_r$
3. $\Delta H = \sum n\Delta H^\circ_{(\text{products})} - \sum n\Delta H^\circ_{(\text{reactants})}$

Try sample problems on p. 336-338