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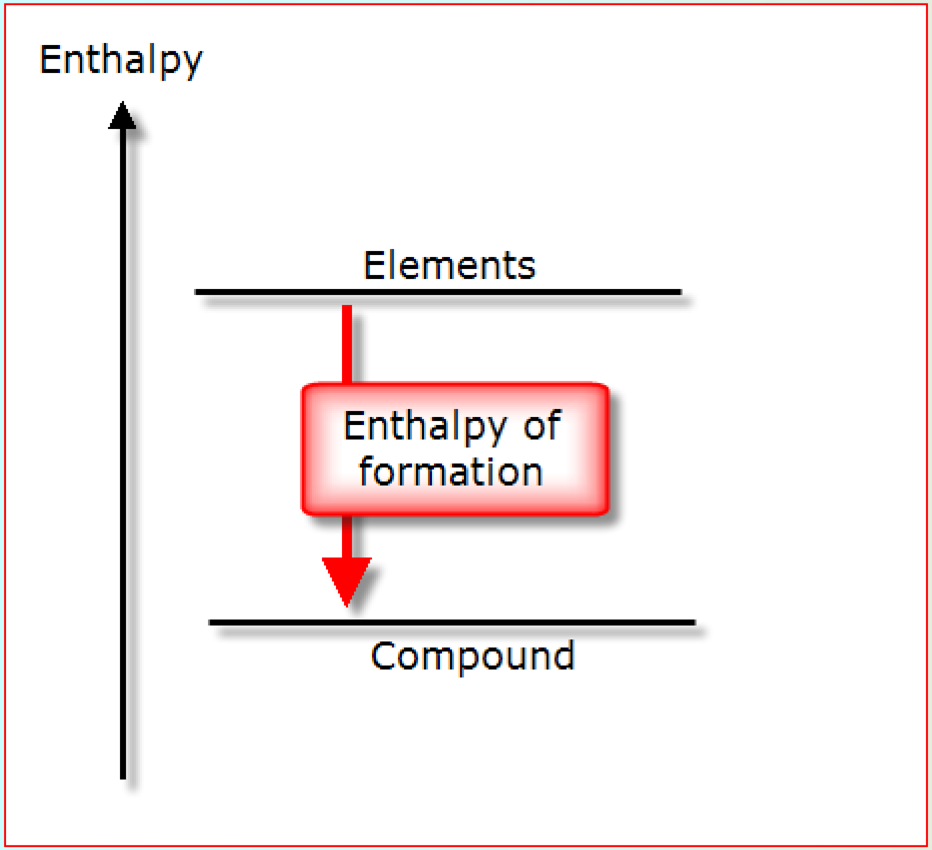
# Enthapies of Formation




**In this picture, we represent an exothermic reaction: the enthalpy of formation is negative.**

## Introduction

We can calculate the enthalpy changes for many reactions from tabulated ΔH values.

A particularly important process used for tabulating the enthalpies is the formation of a compound from its constituent elements.

The enthalpy changed associated with this process is called the **enthalpy formation** and is labelled

$$\Delta H_f$$

where the subscript "f" indicates that the substance has been formed from its elements.

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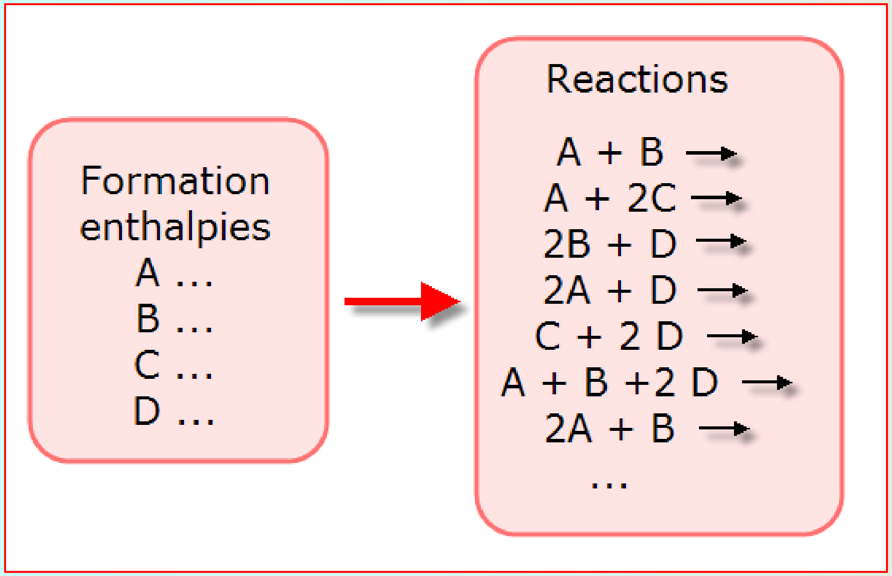


## Introduction

The idea behind the enthalpies of formation can be seen as a solution to the search for the minimum set of tabulated enthalpies of reaction to know the enthalpies of all possible reactions:

•QUESTION: Which is the least set of values needed to know the enthalpies of all possible reactions?

•ANSWER: One value per compound

•SOLUTION: The formation enthalpies of all compounds. With this, we can calculate the enthalpy of any possible combination or reaction.

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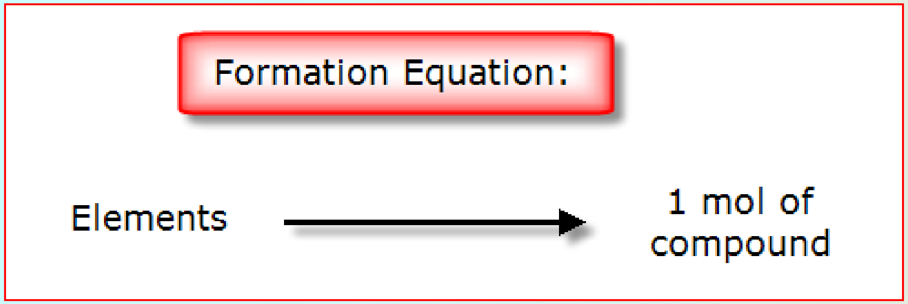


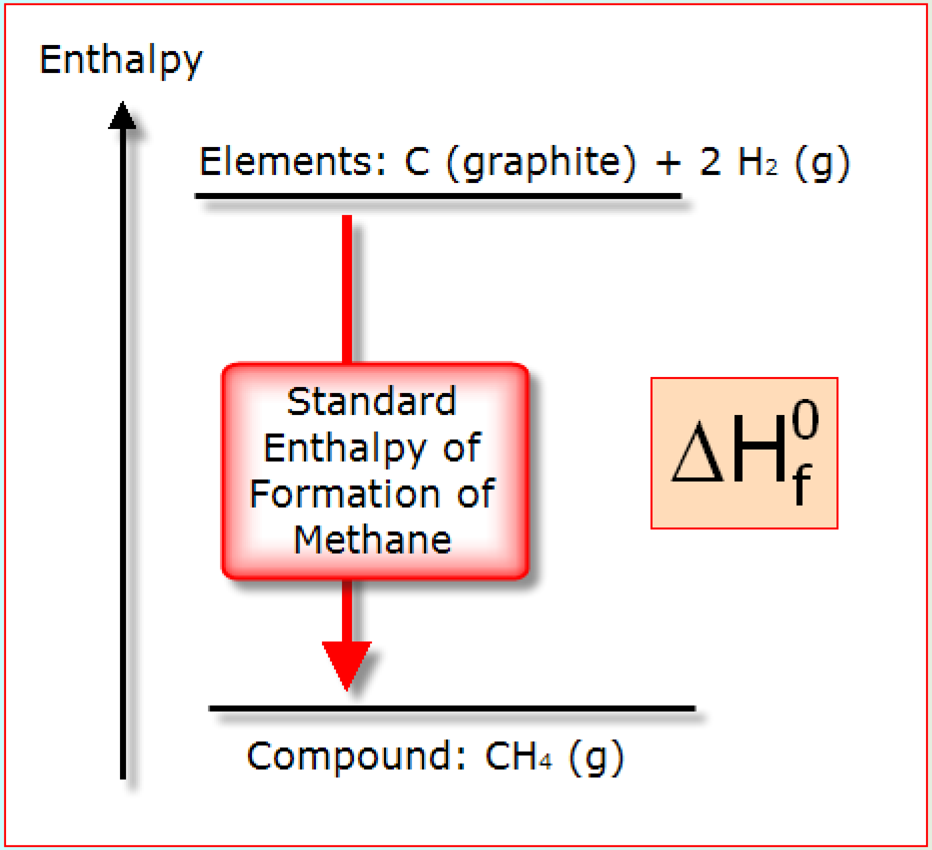


## Standard Enthalpy of Formation

The standard enthalpy of formation of a compound

$\Delta H_f^0$

is the change in enthalpy for the reaction that forms 1 mol of the compound from its elements, with all substances in their standard states.

A right superscript zero indicates the standard states

## Standard State

The magnitude of the enthalpy change depends on the conditions of temperature, pressure, and state (gas, liquid, solid, crystalline form).

It is convenient to define a set of conditions, called a **standard state**, at which the enthalpies are measured and tabulated.

The standard state of a substance is its pure form at atmospheric pressure (1 atm) and a temperature of 25 ºC (298 K)

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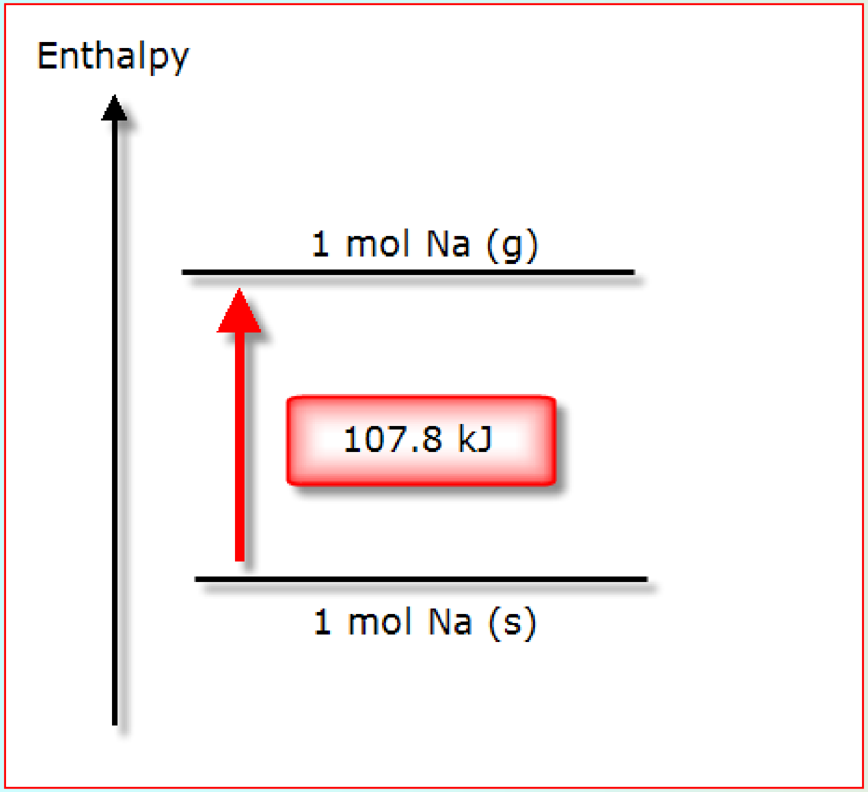


**Standard Enthalpies of Formation of Elements**

An element in its standard state is assigned a standard enthalpy change of formation

$$\Delta H_f^0 \text{ (element, standard state)} = 0$$

of zero.

For example, note that in the case of sodium

$$\Delta H_f^0 \,[Na(s)] = 0$$

but

$$\Delta H_f^0 \,[Na(g)] = 107.8 \text{ kJ/mol}$$

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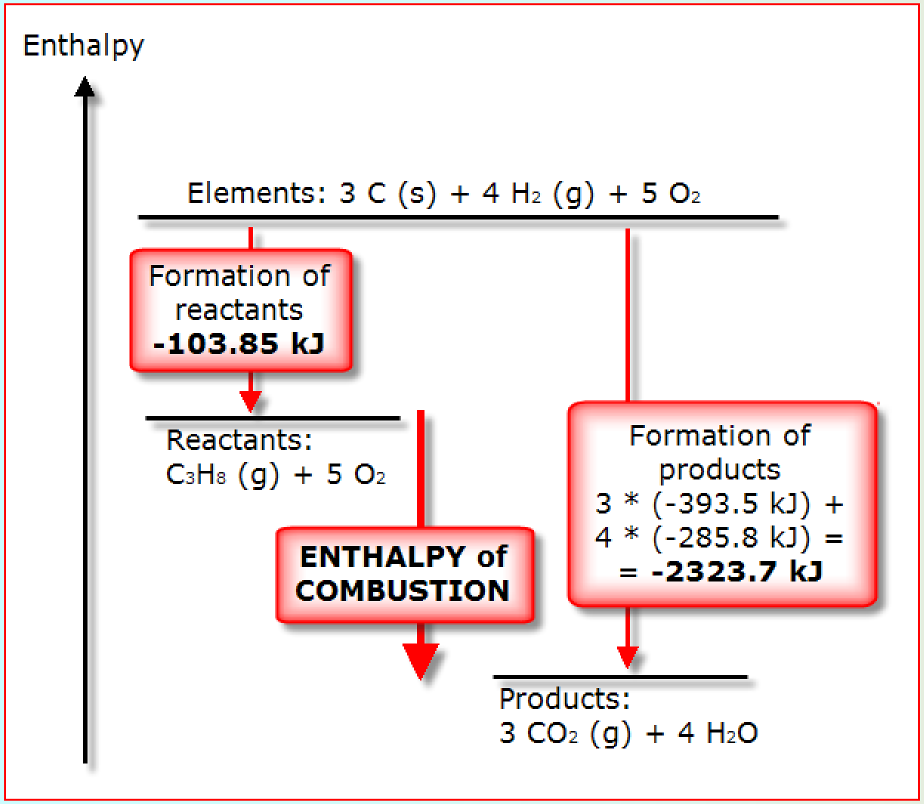


## Using Enthalpies of Formation to Calculate Enthalpies of Reaction

The enthapies of formation are usefeul because we can use them to calculate the standard enthapy change for any reaction.

Let's consider the combustion of propane:

$$C_3H_8\ (g) + 5\ O_2\ (g) \rightarrow 3\ CO_2\ (g) + 4\ H_2O\ (l)$$

If we know the standard formation enthalpies for the compounds involved in this reaction, we can calculate the enthalpy of the combustion reaction.

Standard formation enthalpies:

$$\Delta H_f^0\ [C_3H_8\ (g)] = -103.85\ kJ/mol$$

$$\Delta H_f^0\ [CO_2\ (g)] = -393.5\ kJ/mol$$

$$\Delta H_f^0\ [H_2O\ (l)] = -285.8\ kJ/mol$$

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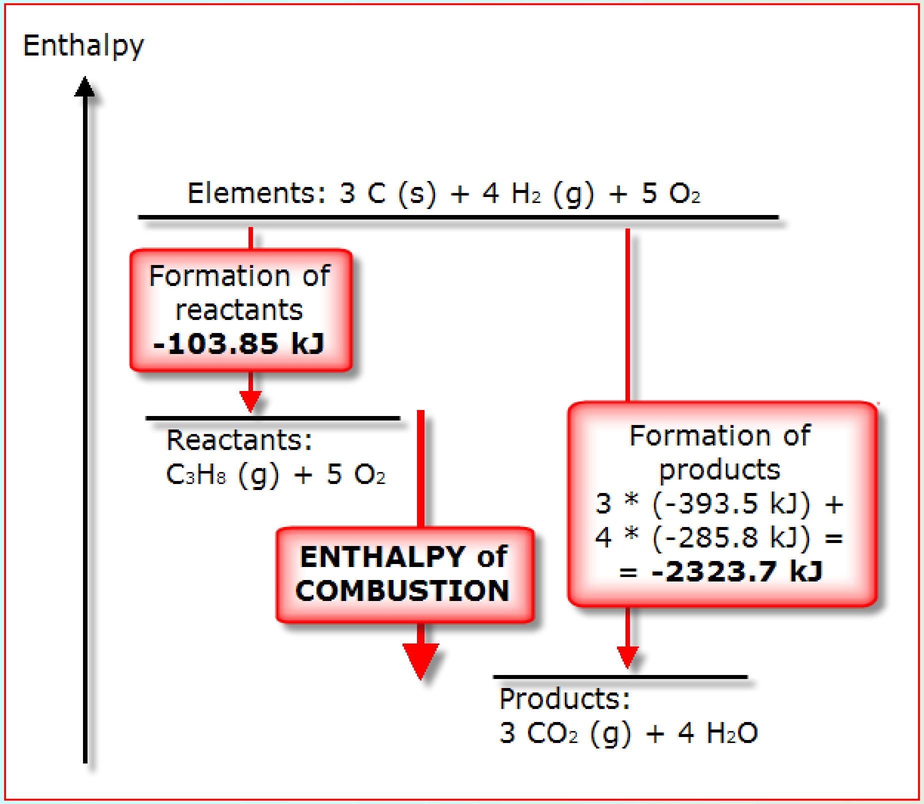
# Enthapies of Formation




$C_3H_8\ (g) + 5\ O_2\ (g) \rightarrow 3\ CO_2\ (g) + 4\ H_2O\ (l)$

## Using Enthalpies of Formation to Calculate Enthalpies of Reaction

Let's write the equations of formation in order to apply Hess's Law by combining them in order to get the enthalpy of combustion:

$3\ C\ (s) + 4\ H_2\ (g) \rightarrow C_3H_8\ (g) \qquad \Delta H_1 = -103.85\ kJ$

$C\ (s) + O_2\ (g) \rightarrow CO_2\ (g) \qquad \Delta H_2 = -393.5\ kJ$

$H_2\ (g) + 0.5\ O_2\ (g) \rightarrow H_2O\ (l) \qquad \Delta H_3 = -285.8\ kJ$

----------

$C_3H_8\ (g) + 5\ O_2(g) \rightarrow 3\ CO_2\ (g) + 4\ H_2O\ (l)$

We see that the formation equation of the **reactant needs to be reversed**:

$3\ C\ (s) + 4\ H_2\ (g) \rightarrow C_3H_8\ (g) \qquad \Delta H_1 = -103.85\ kJ$

$C\ (s) + O_2\ (g) \rightarrow CO_2\ (g) \qquad \Delta H_2 = -393.5\ kJ$

$H_2\ (g) + 0.5\ O_2\ (g) \rightarrow H_2O\ (l) \qquad \Delta H_3 = -285.8\ kJ$

reversed

----------

$C_3H_8\ (g) + 5\ O_2(g) \rightarrow 3\ CO_2\ (g) + 4\ H_2O\ (l)$

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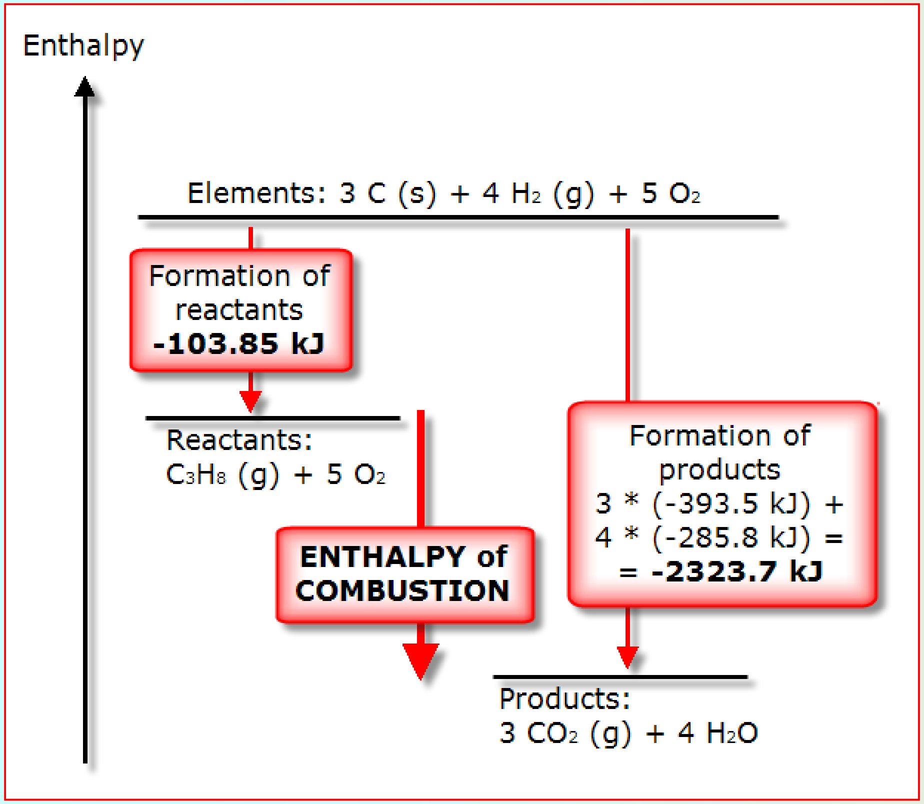


$$C_3H_8\ (g) + 5\ O_2\ (g) \rightarrow 3\ CO_2\ (g) + 4\ H_2O\ (l)$$

## Using Enthalpies of Formation to Calculate Enthalpies of Reaction

The formation equations of the products need to be multiplied by the number of moles of each product:

$C_3H_8\ (g) \rightarrow 3\ C\ (s) + 4\ H_2\ (g) \qquad \Delta H_1 = +103.85\ kJ$

$C\ (s) + O_2\ (g) \rightarrow CO_2\ (g) \qquad \Delta H_2 = -393.5\ kJ$

$H_2\ (g) + 0.5\ O_2\ (g) \rightarrow H_2O\ (l) \qquad \Delta H_3 = -285.8\ kJ$

* 3    * 4

----------

$C_3H_8\ (g) + 5\ O_2(g) \rightarrow 3\ CO_2\ (g) + 4\ H_2O\ (l)$

By summation we get the enthalpy of combustion of propane:

$C_3H_8\ (g) \rightarrow \cancel{3\ C\ (s)} + \cancel{4\ H_2\ (g)} \qquad \Delta H_1 = +103.85\ kJ$

$3 * [\cancel{C\ (s)} + O_2\ (g) \rightarrow CO_2\ (g) \qquad \Delta H_2 = -393.5\ kJ]$

$4 * [\cancel{H_2\ (g)} + 0.5\ O_2\ (g) \rightarrow H_2O\ (l) \qquad \Delta H_3 = -285.8\ kJ]$

----------

$C_3H_8\ (g) + 5\ O_2(g) \rightarrow 3\ CO_2\ (g) + 4\ H_2O\ (l) \quad \Delta H_{comb} = -2220\ kJ$

# Enthapies of Formation

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**Using Enthalpies of Formation to Calculate Enthalpies of Reaction**

We can break down any reaction into formation reactions.

This way, we obtain the following general expression:

$$\Delta H^0_{reaction} = \Sigma\ n * \Delta H^o_f\ (products) - \Sigma\ m * \Delta H^o_f\ (reactants)$$