Hess' Law and Enthalpy of Formation

Skills to Develop

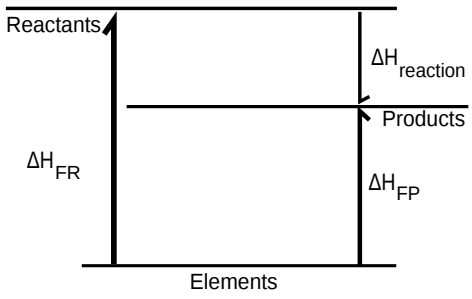
- Calculate enthalpies of reactions using Hess' Law
- Define "standard state"

What is Hess' Law?

Hess' Law is an early statement of the law of conservation of energy (1840). It says that the heat liberated by a process doesn't depend on how the process happens (only on the starting and ending states: in other words, it's a state function). Now we know we should really use enthalpy for this, not heat, because enthalpy is a state function, so this is true, while heat is a process. Hess' Law lets us break a reaction or process into a series of small, easily measured steps, and then we can add up the ΔH of the steps to find the change in enthalpy of the whole thing. Hess' law is a great way to think about chemical processes and make predictions. We'll see lots of applications of Hess' law, but right now let's start with finding reaction enthalpies using standard enthalpies of formation.

Standard Enthalpies of Formation

Standard enthalpies of formation help us predict reaction enthalpies for many reactions if the products and reactants are well-studied, even if the specific reaction is new. To do this, we imagine that we take the reactants and separate them into their pure elements in a **standard state**. The standard state is the element in its most stable form at room temperature and atmospheric pressure. Then we take the elements and recombine them to make the products. The reaction enthalpy is equal to the difference in the enthalpies of these processes.



A diagram showing the application of Hess' law to reaction enthalpy calculated from enthalpies of formation.

Example

Let's look at a specific example. Here are some enthalpies of formation (in kJ/mol of reaction):

$$C(s)+O_2(g)\rightarrow CO_2(g)\ \Delta H_f=-394\ kJ\$\$\$\$C(s)+2H_2(g)\rightarrow CH_4(g)\ \Delta H_f=-75\ kJ\$\$\$\$2H_2(g)+O_2(g)\rightarrow 2H_2O(l)\ \Delta H_f=-572\ kJ \quad (1)$$

Let's use these enthalpies of formation to calculate the enthalpy of combustion for 1 mol of methane. The reaction we want is

$$CH_4(g)+2O_2(g)\rightarrow CO_2(g)+2H_2O(g) \quad (2)$$

If we reverse a reaction, we change the sign on ΔH, and if we multiply the reaction by a constant coefficient, we multiply ΔH by the same coefficient. Let's combine the formation constant equations so they add up to the reaction we want:

CH_4 (g)	→	C(s) + $2H_2$(g)	ΔH_f = 75 kJ
C(s) + O_2(g)	→	CO_2(g)	ΔH_f = -394 kJ
$2H_2$(g) + O_2(g)	→	$2H_2O$(l)	ΔH_f = -572 kJ

That's almost right but we're missing the state of the water:

$$H_2O(l) \rightarrow H_2O(g)\ \Delta H = 44\ kJ \tag{3}$$

The full set is

$CH_4(g)$	→	$C(s) + 2H_2(g)$	ΔH_f = 75 kJ
$C(s) + O_2(g)$	→	$CO_2(g)$	ΔH_f = -394 kJ
$2H_2(g) + O_2(g)$	→	$2H_2O(l)$	ΔH_f = -572 kJ
$2H_2O(l)$	→	$2H_2O(g)$	ΔH_f = 88 kJ
$CH_4(g) + 2O_2(g)$	→	**$CO_2(g) + 2H_2O(g)$**	**ΔH = ?**

Note that everything but the desired reaction cancels. Now we just add up the enthalpies of each step, and we find that the enthalpy of combustion of 1 mole of methane is -803 kJ.

General Procedure for Hess' Law Calculations

Determine the equation for the desired process (the process for which you want to know the enthalpy change). Break it into steps for which you can look up the enthalpy changes. This probably means steps like formation from elements, and changes of state. (Later, we'll include other processes like ionization, etc) Arrange the steps so that everything cancels out leaving just the desired reaction. Make sure the coefficients on equations are correct (multiply the equation and ΔH by a constant if needed) and that all the components are in the correct state (like the example above, we had to convert from liquid water to gaseous water). Then just add it up!

Outside Links

- Khan Academy: Enthalpy of Formation (12 min)
- Khan Academy: Hess's Law and Reaction Enthalpy Change (16 min)
- Khan Academy: Hess's Law Example (12 min)
- CrashCourse Chemistry: Enthalpy (11 min)

Contributors and Attributions

- Emily V Eames (City College of San Francisco)