

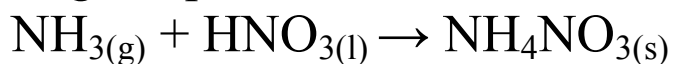
## SCH 4U

### The use of Formation Equations with Hess' Law

Using formation equations in the context of Hess' Law to calculate the enthalpy of a target reaction leads to many simplifications. The technique, along with an explanation of the simplifications, follows.

Main Idea: A formation equation can be used as a 'known' equation when calculating the enthalpy change of some target equation.

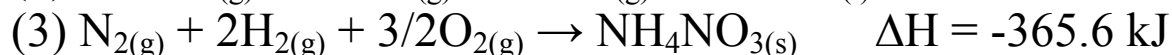
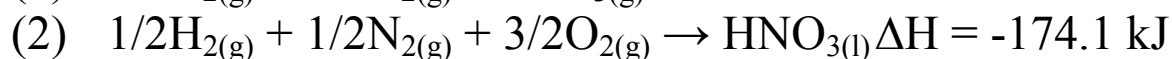
Consider this target equation:



Normally, you would be given a list of known equations that contain the three chemicals that are present in the target equation above. Your job would be to manipulate those known equations so that their sum would give you the target equation.

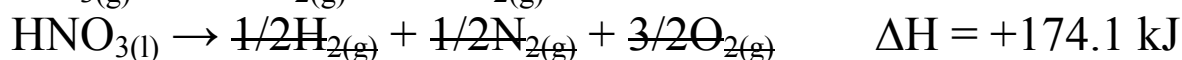
Now, your known equations will be the formation equations of all of the chemicals in your target equation. Their enthalpies will be the enthalpies of formation ( $\Delta H_f$ ) from appendix C6 (p799) of your textbook.

In the above target equation, there are three chemicals ( $\text{NH}_{3(g)}$ ,  $\text{HNO}_{3(l)}$ ,  $\text{NH}_4\text{NO}_{3(s)}$ ); therefore we need the following three known equations, which are the formation equations of these three chemicals

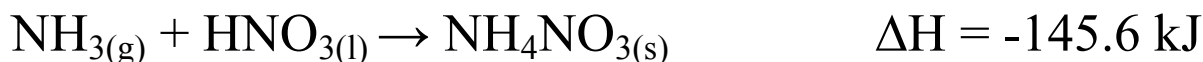


Now, we manipulate these three known equations so that their sum equals the target equation. In this case:

-1X(1); -1X(2); 1X(3); which is:



SUM:



It is important to take note of the following, because these things always happen when you use formation equations as known equations:

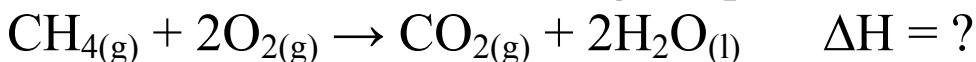
- All of the elements cancel each other out
- You always ‘flip’ a formation equation for a reactant (because the substance starts as a product in the original formation equation)
- You never ‘flip’ a formation equation for a product
- You always multiply a formation equation by the coefficient in front of the compound in the target (this was not apparent in the above example as all of the coefficients were 1)
- You can ignore any elements in their standard state in

a target equation since its  $\Delta H_f = 0$  kJ (this was not apparent in the above example)

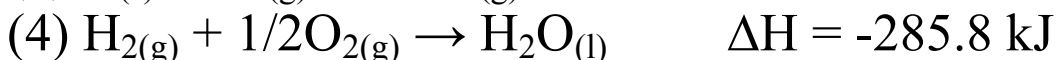
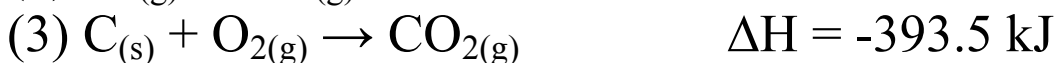
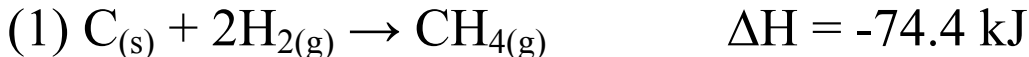
Since you always do the same thing with formation equations (as explained above), it is not necessary to actually write down the formation equations, you need only 'manipulate' the  $\Delta H_f$ 's.

Here is a concise example:

Calculate the  $\Delta H$  for this target equation:



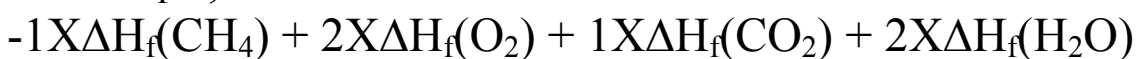
In its full detail, you would write the following four formation equations:



(NOTE: always get the correct  $\Delta H$  for water based on its state [see the two entries in the table on page 800 for water]).

You would then do -1X(1); 2X(2); 1X(3); 2X(4)

You don't have to do all of this with the formation equations above, you need only do the same work with the  $\Delta H_f$ 's, or:



or

$$-(-74.4) + 2(0) + -393.5 + 2(-285.8) = -890.7 \text{ kJ}$$

You will ALWAYS add the  $\Delta H_f$  of the products, multiplied by their coefficients and subtract the  $\Delta H_f$  of the reactants, multiplied by their coefficients. You will ALWAYS ignore the  $\Delta H_f$  of elements in their standard state, because they are always 0 kJ. Or, in summary:

$$\Delta H = \sum n\Delta H_f \text{ products} - \sum n\Delta H_f \text{ reactants}$$

Work on the following homework:

P335 #2-5; P338 #7-9; P339 #1-4