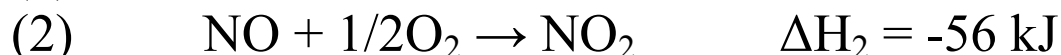
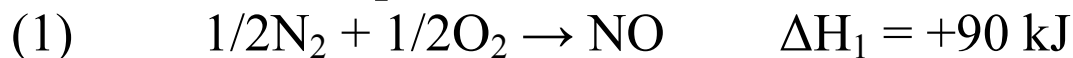


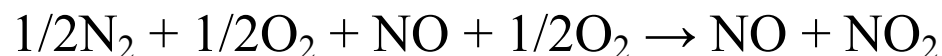
Hess's Law of Additivity of Reactions

A scientist named Hess discovered that chemical reactions that, when added, equal another chemical equation (called the 'target' equation), their enthalpies (ΔH) sum to that of the 'target' equation. Here is an example to illustrate:

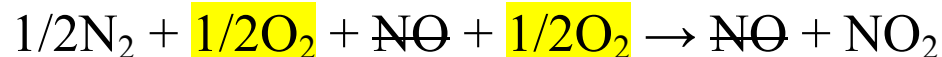
These three reactions have been carefully studied in the lab. Their enthalpies have been determined:



If you look closely, you can see that equations (1) + (2) equal equation (3). When you add equations, simply write all reactants and products together in one equation, then simplify. In this case, here is (1) + (2)



You can cancel the NO (it's on both sides), and combine the O₂'s on the left:



This gives the simplified equation:



This is the same as equation (3).

Since (1) + (2) = (3), then $\Delta H_1 + \Delta H_2$ should equal ΔH_3 :
 $+90 \text{ kJ} + -56 \text{ kJ} = +34 \text{ kJ}$. $+34 \text{ kJ}$ is ΔH_3 .

A practical application of this concept involves calculating enthalpies of reactions that are difficult to study in the lab (some reasons are: chem's are expensive; reaction is dangerous or unpredictable).

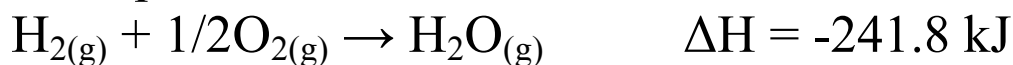
Since literally thousands of reactions have been studied, with their enthalpies available in reference sources, it is possible to find 2 or more of these reactions that 'sum up' to some difficult reaction. The difficult reaction's enthalpy can then be calculated indirectly, without ever actually performing the reaction itself.

There is one small catch! Reactions that are used from reference sources must often be 'modified' in order to use them to solve a particular problem.

In order to modify these equations to obtain enthalpies of reactions, two basic laws about enthalpies must often be used. They deal with reversing a reaction and multiplying a reaction by a coefficient. The laws are as follows:

Law 1: When a chemical reaction is reversed, the sign of its ΔH changes.

Example: This reaction is well known:



Therefore, we also know the reverse reaction:

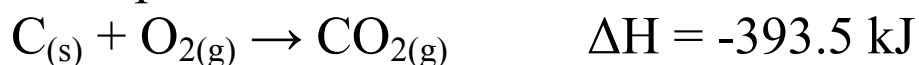


(This makes perfect sense. If you reverse a reaction, the

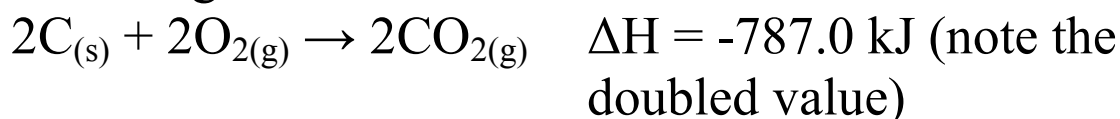
energy flow will also be reversed.)

Law 2: When a chemical reaction is multiplied by a coefficient, the same is true of its enthalpy.

Example: This reaction is well known:



When multiplied by a coefficient, (in this case 2), the following is also true:



(This makes perfect sense. If you double the amount of chemical, you double the energy change.)

When you are asked to determine the enthalpy of some reaction, you won't be expected to go to references sources and find equations (from several 100 possibilities) that sum up to your 'target' equation. You will be given a list of equations to choose from, called 'known' equations. You will have to determine which ones to use and how they must be modified (reversed and/or multiplied by coefficients) in order to sum up to your target equation.

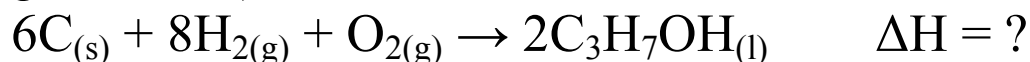
Following a set of systematic steps can help you get to the answer as quickly as possible.

Instead of looking at the 'whole' target equation and trying to solve it in one step, concentrate on one chemical at a time.

- 1) Choose the first chemical of your target equation.
- 2) Find a 'known' equation that has this chemical in it (be aware of states, they will be important in some questions).
[If the chemical appears in more than one 'known' equation, skip this chemical, it will be accounted for in the end.]
- 3) Modify the 'known' equation so the chemical is on the correct side AND with the correct coefficient. Make the same change to its ΔH (change sign and/or multiply).
- 4) Repeat steps 2 and 3 for all remaining chemicals.
- 5) Add up all of your modified 'known' equations, verifying they do in fact add up to your target equation. Add up all ΔH 's.
- 6) The sum of the ΔH 's will be the ΔH of your target equation.

Here is an example:

Target equation (the equation for which you are asked to get the ΔH):



Calculate its ΔH using the following known equations:



Our target equation has 4 chemicals in it. Each chemical will be obtained one at a time. The end result should be the target equation.

First, get the ' $6\text{C}_{(s)}$ ' reactant. C exists in known equation (2):

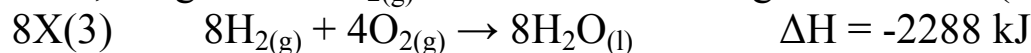


'C' is on the correct side (reactant), but it needs a coefficient of '6', therefore we will multiply known equation (2) by 6 [noted as 6X(2)]:



(notice how ΔH was also multiplied by 6).

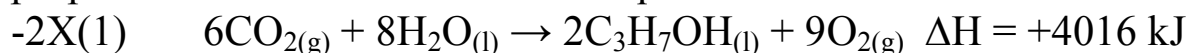
Next, we get the ' $8\text{H}_{2(g)}$ ' reactant. We will get it from 8X(3):



Next, we need ' $\text{O}_{2(g)}$ ' reactant. Oxygen exists in all three known equations. It will be skipped, as it should be properly accounted for in the end.

Lastly, we need ' $2\text{C}_3\text{H}_7\text{OH}_{(l)}$ ' product. We will get it from -2X(1)

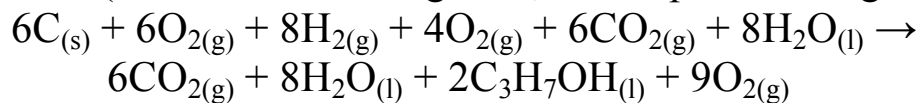
[known equation (1) needs to be flipped AND multiplied by 2 to get the propanol on the correct side of the equation with the correct coefficient]



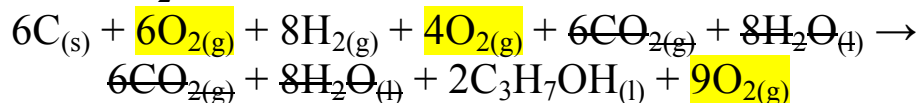
Next, the sum of the three modified known equations will be checked to make sure that they add up to the target equation. They are rewritten here to be more clear:



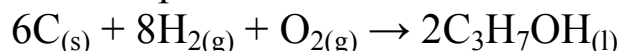
Sum (all the reactants together, all the products together):



The $\text{CO}_{2(\text{g})}$'s and $\text{H}_2\text{O}_{(\text{l})}$'s 'cancel' themselves out, and the O_2 's simplify to one O_2 on the reactant side:



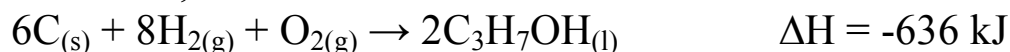
This simplifies to:



This is exactly our target equation! That means that the sum of the ΔH 's of the three modified equations must be the ΔH of the target equation:

$$\Delta\text{H} = (-2364 - 2288 + 4016) \text{ kJ} = -636 \text{ kJ}$$

Therefore, the final answer looks like this:



There are 2 more examples on pages 324-325 you can read if you want to see more.

HW: P326 #1-3; P329 #4-5; P330 #1-4