

## 6 • Thermochemistry

### 6.3, 6.4 NOTES

The enthalpy of the reactants,  $H_{\text{reactants}}$  and the enthalpy of the products,  $H_{\text{products}}$  depend on the bonding of the reactants and products... nothing else. So, the  $\Delta H_{\text{reaction}}$  only depends on the initial and final state of the reaction, not how you got from one state to another state. It is called a "state function".

Practically speaking, if we can find several equations that "add up" to the equation we want, the  $\Delta H_{\text{reactions}}$  will add up to the overall  $\Delta H$ . This is called Hess's Law.

**Standard Enthalpy of Formation ( $\Delta H^\circ_f$ ):** Change in enthalpy that accompanies the formation of 1 mole of a compound from its elements with all substances in their standard states.

Conventional definitions for Standard States:

For a compound:

- The standard state of a gaseous substance is a pressure of exactly 1 atmosphere.
- For a pure substance in a condensed state ( liquid or solid ), the standard state is the pure liquid or solid.
- For a substance present in solution, the standard state is a concentration of exactly 1 M.

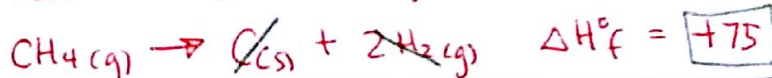
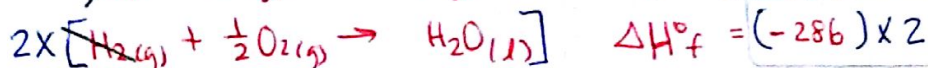
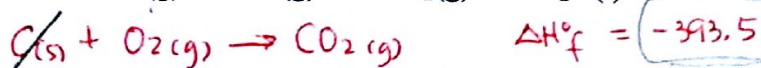
For an element:

- The standard state of an element is the form in which the element exists under conditions of 1 atm and 25°C.

[These constants can be found in Appendix 4 on pages A19-A22 of Zumdahl.]

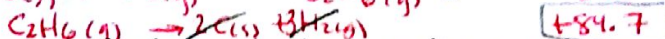
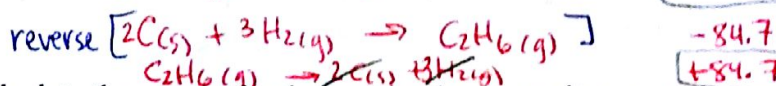
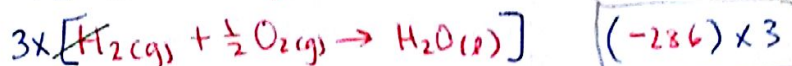
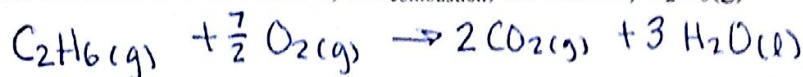
| Compound                           | Formation Equation  | $\Delta H^\circ_f$ (kJ/mol) |
|------------------------------------|---|-----------------------------|
| CH <sub>4</sub> (g)                | C(s) + 2H <sub>2</sub> (g) → CH <sub>4</sub> (g)                            | -75                         |
| H <sub>2</sub> O(l)                | H <sub>2</sub> (g) + $\frac{1}{2}$ O <sub>2</sub> (g) → H <sub>2</sub> O(l) | -286                        |
| H <sub>2</sub> O(g)                | H <sub>2</sub> (g) + $\frac{1}{2}$ O <sub>2</sub> (g) → H <sub>2</sub> O(g) | -242                        |
| CO <sub>2</sub> (g)                | C(s) + O <sub>2</sub> (g) → CO <sub>2</sub> (g)                             | -393.5                      |
| C <sub>2</sub> H <sub>6</sub> (g)  | 2C(s) + 3H <sub>2</sub> (g) → C <sub>2</sub> H <sub>6</sub> (g)             | -84.7                       |
| C <sub>3</sub> H <sub>8</sub> (g)  | 3C(s) + 4H <sub>2</sub> (g) → C <sub>3</sub> H <sub>8</sub> (g)             | -104                        |
| C <sub>4</sub> H <sub>10</sub> (g) | 4C(s) + 5H <sub>2</sub> (g) → C <sub>4</sub> H <sub>10</sub> (g)            | -125.6                      |

a) Calculate the  $\Delta H_{\text{combustion}}$  for CH<sub>4</sub>:



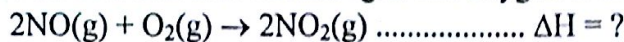
$$\Delta H = -890.5 \frac{\text{kJ}}{\text{mol}}$$

b) Calculate the heat of combustion,  $\Delta H_{\text{combustion}}$ , for ethane,  $\text{C}_2\text{H}_6(\text{g})$

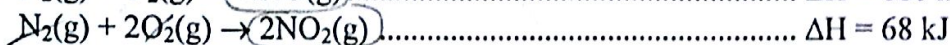


$$\Delta H = -1560.3 \frac{\text{kJ}}{\text{mol}}$$

c) Calculate the energy for the reaction between nitrogen and oxygen to form nitrogen dioxide:



Use these two reactions:



$$\Delta H = -112 \frac{\text{kJ}}{\text{mol}}$$

d) Notice that we do the same thing each time.

If a compound is a **reactant**...

what do you do to the equation? REVERSE IT What do you do to the  $\Delta H_f$ ?  $\times (-1)$

If a compound has a **coefficient of 3**...

what do you do to the equation? TRIPLE IT What do you do to the  $\Delta H_f$ ?  $\times 3$

Write the "shortcut version" of Hess's Law (when using  $H_f^\circ$ ):

$$\Delta H_{\text{rxn}} = \sum n_p \Delta H_f^\circ (\text{products}) - \sum n_r \Delta H_f^\circ (\text{reactants})$$

e) Use this shortcut to calculate the  $\Delta H_{\text{combustion}}$  of butane,  $\text{C}_4\text{H}_{10}(\text{g})$ .



$$\Delta H_{\text{combustion}} = 4 (\Delta H_f \text{CO}_2) + 5 (\Delta H_f \text{H}_2\text{O}) - (\Delta H_f \text{C}_4\text{H}_{10} + \frac{13}{2} (\Delta H_f \text{O}_2))$$

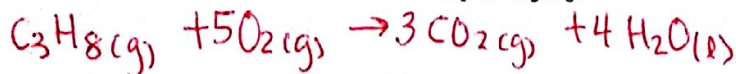
$$= 4(-393.5) + 5(-286) - [(-125.6) + \frac{13}{2}(0)]$$

$$= -2878.4 \frac{\text{kJ}}{\text{mol}}$$

f) The heat of combustion of propane,  $\text{C}_3\text{H}_8$ , is  $-2220.0 \text{ kJ/mol}$ .

Set up the shortcut for the calculation of the  $\Delta H_{\text{combustion}}$  of propane.

Use this information to calculate the  $\Delta H_f$  of  $\text{C}_3\text{H}_8$ .



$$\Delta H_{\text{combustion}} = 3 (\Delta H_f \text{CO}_2) + 4 (\Delta H_f \text{H}_2\text{O}) - \Delta H_f \text{C}_3\text{H}_8$$

$$-2220.0 = 3(-393.5) + 4(-286) - X$$

$$X = -104.5 \frac{\text{kJ}}{\text{mol}}$$