Culver City H.S. • AP Chemistry

Name Fontijn
Period ___ Date ___/___/___

6 • Thermochemistry

6.3,6.4 NOTES

The enthalpy of the reactants, $H_{reactants}$ and the enthalpy of the products, $H_{products}$ depend on the bonding of the reactants and products… nothing else. So, the $\Delta H_{reaction}$ 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 $\Delta H_{reactions}$ will add up to the overall ΔH. This is called Hess's Law.

Standard Enthalpy of Formation ($\Delta H°_f$): 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	$\Delta H°_f$ (kJ/mol)
$CH_4(g)$	$C(s) + 2H_2(g) \rightarrow CH_4(g)$	-75
$H_2O(l)$	$H_{2(g)} + \frac{1}{2} O_{2(g)} \rightarrow H_2O_{(l)}$	-286
$H_2O(g)$	$H_{2(g)} + \frac{1}{2} O_{2(g)} \rightarrow H_2O_{(g)}$	-242
$CO_2(g)$	$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$	-393.5
$C_2H_6(g)$	$2C_{(s)} + 3H_{2(g)} \rightarrow C_2H_{6(g)}$	-84.7
$C_3H_8(g)$	$3C_{(s)} + 4H_{2(g)} \rightarrow C_3H_{8(g)}$	-104
$C_4H_{10}(g)$	$4C_{(s)} + 5H_{2(g)} \rightarrow C_4H_{10(g)}$	-125.6

a) Calculate the $\Delta H_{combustion}$ for CH_4:

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

$$\cancel{C}_{(s)} + O_{2(g)} \rightarrow CO_{2(g)} \qquad \Delta H°_f = \boxed{-393.5}$$

$$2\times[\cancel{H_2}_{(g)} + \tfrac{1}{2}O_{2(g)} \rightarrow H_2O_{(l)}] \qquad \Delta H°_f = \boxed{(-286)\times 2}$$

reverse $[C_{(s)} + 2H_{2(g)} \rightarrow CH_{4(g)}] \qquad \Delta H°_f = -75$

$$CH_{4(g)} \rightarrow \cancel{C}_{(s)} + \cancel{2H_2}_{(g)} \qquad \Delta H°_f = \boxed{+75}$$

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

b) Calculate the heat of combustion, $\Delta H_{combustion}$, for ethane, $C_2H_6(g)$

$C_2H_{6(g)} + \frac{7}{2} O_{2(g)} \rightarrow 2 CO_{2(g)} + 3 H_2O_{(l)}$

$2\times[\cancel{C}_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}]$ $\boxed{(-393.5)\times 2}$

$3\times[\cancel{H}_{2(g)} + \frac{1}{2} O_{2(g)} \rightarrow H_2O_{(l)}]$ $\boxed{(-286)\times 3}$

reverse $[2C_{(s)} + 3H_{2(g)} \rightarrow C_2H_{6(g)}]$ -84.7

$C_2H_{6(g)} \rightarrow \cancel{2C_{(s)}} + \cancel{3H_{2(g)}}$ $\boxed{+84.7}$

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

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

$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$ .... $\Delta H = ?$

Use these two reactions:

reverse $N_2(g) + O_2(g) \rightarrow 2NO(g)$.. $\Delta H \overset{-}{=} 180$ kJ

$\cancel{N_2}(g) + 2\cancel{O_2}(g) \rightarrow 2NO_2(g)$.. $\Delta H = 68$ kJ

$2NO_{(g)} \rightarrow \cancel{N_{2(g)}} + \cancel{O_{2(g)}}$ $\Delta H = -180$ kJ

$\boxed{\Delta H = -112\ kJ}/mol$

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 ΔH_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 ΔH_f? $\times 3$

Write the "shortcut version" of Hess's Law (when using H°_f):

$$\Delta H_{rxn} = \Sigma\ n_p \Delta H^\circ_f\ (products) - \Sigma n_r \Delta H^\circ_f\ (reactants)$$

e) Use this shortcut to calculate the $\Delta H_{combustion}$ of butane, $C_4H_{10}(g)$.

$C_4H_{10(g)} + \frac{13}{2} O_{2(g)} \rightarrow 4 CO_{2(g)} + 5 H_2O_{(l)}$

$\Delta H_{combustion} = 4(\Delta H_f\ CO_2) + 5(\Delta H_f\ H_2O) - (\Delta H_f\ C_4H_{10} + \frac{13}{2}(\Delta H_f\ O_2))$

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

$= \boxed{-2878.4\ kJ/mol}$

f) The heat of combustion of propane, C_3H_8, is –2220.0 kJ/mol.
Set up the shortcut for the calculation of the $\Delta H_{combustion}$ of propane.
Use this information to calculate the ΔH_f of C_3H_8.

$C_3H_{8(g)} + 5O_{2(g)} \rightarrow 3 CO_{2(g)} + 4 H_2O_{(l)}$

$\Delta H_{combustion} = 3(\Delta H_f\ CO_2) + 4(\Delta H_f\ H_2O) - \Delta H_f\ C_3H_8$

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

$x = \boxed{-104.5\ kJ/mol}$