Name ______________________
Period ___ Date ___/___/___

6 • Energy and Chemical Reactions

HESS'S LAW CALCULATIONS

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.

Heats of Formation: Write the formation equations for the following. [See Table 6.2 on page 270 of text.]

Compound	Formation Equation	ΔH_f ($kJ \cdot mol^{-1}$)
$CH_4(g)$	$C(s) + 2H_2(g) \rightarrow CH_4(g)$	-74.8
$H_2O(l)$	$H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l)$	-285.8
$H_2O(g)$	$H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(g)$	-241.8
$CO_2(g)$	$C(s) + O_2(g) \rightarrow CO_2(g)$	-393.5
$C_2H_6(g)$	$2C(s) + 3H_2(g) \rightarrow C_2H_6(g)$	-84.7
$C_3H_8(g)$	$3C(s) + 4H_2(g) \rightarrow C_3H_8(g)$	???
$C_4H_{10}(g)$	$4C(s) + 5H_2(g) \rightarrow C_4H_{10}(g)$	-125.6

Example in class:
Calculate the $\Delta H_{combustion}$ for CH_4:

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

$CH_4 \rightarrow C + 2H_2$ $-(-74.8)$ 74.8

$C + O_2 \rightarrow CO_2$ -393.5 -393.5

$2H_2 + O_2 \rightarrow 2H_2O(g)$ $2(-241.8)$ -483.6

-802.3 kJ/mol

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

$$C_2H_6 + \tfrac{7}{2}O_2 \rightarrow 2CO_2 + 3H_2O(g)$$

$C_2H_6 \rightarrow 2C + 3H_2$ $-(-84.7)$ 84.7

$2C + 2O_2 \rightarrow 2CO_2$ $2(-393.5)$ -787

$3H_2 + \frac{3}{2}O_2 \rightarrow 3H_2O$ $3(-241.8)$ -725.4

-1427.7 kJ/mol

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

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

Use these two reactions:

$N_2(g) + O_2(g) \rightarrow 2NO(g)$ .... ΔH = ~~180 kJ~~

$N_2(g) + 2O_2(g) \rightarrow 2NO_2(g)$ .... ΔH = 68 kJ

$2NO \rightarrow N_2 + O_2$ -180 kJ

$2NO + O_2 \rightarrow 2NO_2$ -112 kJ

c) 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? change sign

If a compound has a **coefficient of 3**…
what do you do to the equation? mult by 3 What do you do to the ΔH_f? x 3

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

$$\Delta H_{rxn} = \Sigma \Delta H_f \text{ PRODUCTS} - \Sigma \Delta H_f \text{ REACTANTS}$$

"Σ" = sum

Compound	ΔH_f ($kJ \cdot mol^{-1}$)
$H_2O(g)$	-241.8
$CO_2(g)$	-393.5
$C_4H_{10}(g)$	-125.6

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

$$C_4H_{10} + \frac{13}{2}O_2 \rightarrow 4CO_2 + 5H_2O$$

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

$$= 4(-393.5) + 5(-241.8) - (-125.6)$$

$$= -2657.4 \text{ kJ/mol}$$

d) The heat of combustion of propane, C_3H_8, is -2220 $kJ \cdot mol^{-1}$.
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.

$$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$$

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

$$-2220 \frac{kJ}{mol} = 3(-393.5) + 4(-241.8) - x$$

$$x = 3(-393.5) + 4(-241.8) + 2220$$

$$= -1180.5 - 967.2 + 2220 = \boxed{72.3 \text{ kJ/mol}}$$