Worksheet – Hess' Law

Hess' Law

Hess' Law states that the heat evolved in a given process can be expressed as the sum of the heats of several processes that, when added, yield the process of interest. In other words, enthalpy is a state function. If the reactants and products are the same, it doesn't matter how the reaction is carried out.

- When you reverse a reaction, change the sign of ΔH.
- If you modify the coefficients of a reaction, multiply ΔH by the same factor.

1. Carbon tetrachloride, CCl_4, an organic solvent is prepared by the reaction of chlorine gas, Cl_2, with CS_2.
Using Hess' Law, determine the heat of reaction (ΔH) for the reaction:

$$\underline{CS_2\,(l)} + \underline{3\,Cl_2\,(g)} \rightarrow \underline{CCl_4\,(l)} + \underline{S_2Cl_2\,(l)}$$

given these data:

i. $CS_2\,(l) + 3\,O_2\,(g) \rightarrow CO_2\,(g) + 2\,SO_2\,(g)$ ΔH = -1077 kJ
ii. $S\,(s) + \frac{1}{2}\,Cl_2\,(g) \rightarrow \frac{1}{2}\,S_2Cl_2\,(l)$ ΔH = -30.1 kJ
iii. $C\,(gr) + 2\,Cl_2\,(g) \rightarrow CCl_4\,(l)$ ΔH = -135.4 kJ
iv. $S\,(s) + O_2\,(g) \rightarrow SO_2\,(g)$ ΔH = -296.9 kJ
v. $C\,(gr) + O_2\,(g) \rightarrow CO_2\,(g)$ ΔH = -393.5 kJ

$CS_2(l) + 3O_2(g) \rightarrow CO_2(g) + 2SO_2(g)$ −1077
$C + 2Cl_2(g) \rightarrow CCl_4(l)$ −135.4
$2(S(s) + \frac{1}{2}Cl_2(g) \rightarrow \frac{1}{2}S_2Cl_2(l))$ 2(−30.1)
$2SO_2 \rightarrow S(s) + O_2(g)$ 2(+296.9)
$CO_2 \rightarrow C + O_2$ + 393.5
−285.3 kJ

Heats of formation

Another way to solve for ΔH_{rxn} is to use the **standard enthalpies of formation**, ΔH°_f.
A **formation reaction** is one in which **1 mole of product** is formed from **elements** in their standard states.
The heat of formation of elements in their standard state is zero.

2. Which of previous (i – v) is/are formation reactions?

~~i~~ ~~ii~~ (iii) (iv) (v) elements → 1 mol product

compounds
reactants ½ mol
products = ½ ΔH°_f

The standard heats of formation can be used to determine the heat of reaction as follows:

$$\Delta H_{rxn} = \Sigma \Delta H^\circ_f \text{ products} - \Sigma \Delta H^\circ_f \text{ products}$$

3. Using the standard heats of formation, ΔH°_f, shown below, to calculate the ΔH of the following reaction:

$2\ F_2\ (g) + 2\ H_2O\ (l) \rightarrow 4\ HF\ (g) + O_2\ (g)$

	ΔH°_f (kJ)
H_2O	-286
HF	-267

$\Delta H = 4\Delta H_f^\circ HF - 2\ \Delta H^\circ f\ H_2O(l)$

$\Delta H = 4(-267) - 2(-286) = -1596\ kJ$

$\Delta H^\circ_f F_2 = \Delta H^\circ_f O_2 = 0$

4. Write the formation reaction for C_2H_6 (g)

$\Delta H = \Sigma \Delta H_f^\circ$ product $- \Delta H_f^\circ$ reactants

$2C(gr) + 3H_2(g) \rightarrow C_2H_6(g)$ 1 mol product

elements in standard state

5. Find the heat of formation of C_2H_6 (g) from the following data:

$2\ C_2H_6\ (g) + 7\ O_2\ (g) \rightarrow 4\ CO_2\ (g) + 6\ H_2O\ (l)$	$\Delta H = -3120$ kJ
$C\ (gr) + O_2\ (g) \rightarrow CO_2\ (g)$	$\Delta H = -394$ kJ
$H_2\ (g) + \frac{1}{2}\ O_2\ (g) \rightarrow H_2O\ (g)$	$\Delta H = -242$ kJ
$H_2O\ (l) \rightarrow H_2O\ (g)$	$\Delta H = 44$ kJ

$2C + 3H_2 \rightarrow C_2H_6$

$\frac{1}{2}(4CO_2 + 6H_2O \rightarrow 2C_2H_6 + 7O_2)$ $\frac{1}{2}(+3120)$

$2(C + O_2 \rightarrow CO_2)$ $2(-394)$

$3(H_2 + \frac{1}{2}O_2 \rightarrow H_2O(g))$ $3(-242)$

$3(H_2O(g) \rightarrow H_2O(l))$ $3(44)$

$-86\ kJ$

Bond energies

In many cases, the heats of formation and heats of reaction are not tabulated. This is especially true for newly synthesized compounds. Under these circumstances, **approximate** values for ΔH reaction can be calculated using **bond energies**.
Bond energies represent the energy required to break a given type of bond. Bond breakage is endothermic, so all bond energies are positive. Bond breakage is endothermic (+) and bond formation is exothermic (-).

These can be used as follows:

$$\Delta H_{rxn} = \Sigma \text{ bonds broken} - \Sigma \text{ bonds formed}$$

6. Use the reaction and ΔH from **question 3** to calculate the average bond energy (in kJ/mol) for an H-F bond, given the following bond energies:

$2F_2 + 2H_2O \rightarrow 4HF + O_2$

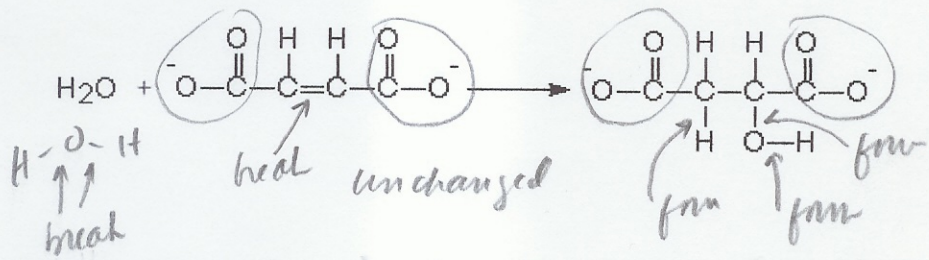
$\Delta H = -496 = 2(158) + 4(460) - 4(H\text{-}F) - 499$

$-4(H\text{-}F) = -2199$

$H\text{-}F = 537\ kJ$

bond	bond energy
H-H	436 kJ
O-H	460 kJ
F-F	156 kJ
O-O	142 kJ
O=O	499 kJ

7. In the Krebs cycle, fumaric acid is hydrated to malic acid. Calculate the heat of reaction in kJ/mol, using the approximate bond energies and reaction shown below.



bond	energy (kJ)
C-H	413
C=C	614
C-C	348
O-H	463
C=O	799
C-O	358
C≡C	839
H-H	436

$\Delta H = 2(O\text{-}H) + (C=C) - (C\text{-}H) - (C\text{-}O) - (O\text{-}H) + (C\text{-}C)$

$\Delta H = (463)+(614) - (413) + (358) - (348) = -42\ kJ$