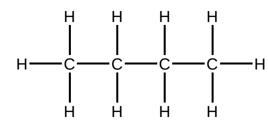
Chapter 3 - Homework

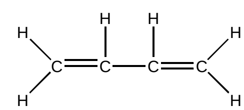
3.1 HydroFluorocarbons ($C_xH_yF_z$) are safer alternative refrigerants than are Chlorofluorocarbons ($C_xCl_yF_z$), which are a major source of ozone depletion in the upper atmosphere. A certain fluorocarbon liquid has an enthalpy of vaporization, $\Delta_{vap}H^o = 26.0$ kJ/mol. Calculate q, w, ΔH, and ΔU when 1.50 mol of the compound is vaporized at -23 ^oC and 750 torr.

3.2 Use Bond Enthalpies (below) to estimate the enthalpy change for the hydrogenation of 1,3-butadiene to form butane:

$C_4H_6(g) + 2\ H_2(g) \rightarrow C_4H_{10}(g)$

Bond	**BE**
H-H	436 kJ/mol
C-H	412
C-C	348
C=C	612

n-butane: 

1,3-butadiene: 

3.3 The enthalpy of formation of ethylbenzene [$C_6H_5(CH_2CH_3)$] is -12.6 kJ/mol.
Calculate the enthalpy of combustion of ethylbenzene

Note: $\Delta_fH^o(CO_2) = -393.5$ kJ/mol , $\Delta_fH^o(H_2O) = -285.8$ kJ/mol

3.4 Consider the formation of liquid methyl methanoate, $CH_3C(O)OCH_3(l)$, from the elements at 25 ^oC :

$3\ C(s) + 3\ H_2(g) + O_2(g) \rightarrow CH_3C(O)OCH_3(l)$.

For this reaction, $\Delta H^o = \Delta_fH^o(CH_3C(O)OCH_3) = -442$ kJ/mol

Calculate ΔU^o for this reaction at 25 ^oC.

3.5 The standard enthalpy of combustion of napthalene, $C_{10}H_8$, is -5157 kJ/mol. Calculate the enthalpy of formation of napthalene.

Note: $\Delta_fH^o(CO_2) = -393.5$ kJ/mol , $\Delta_fH^o(H_2O) = -285.8$ kJ/mol

3.6 The complete combustion of fumaric acid, HOOCCH ═ CHCOOH(s) [$C_4H_4O_4$], in a (constant volume) bomb calorimeter released 1333 kJ per mole of fumaric acid. The reaction is: HOOCCH=CHCOOH(s) + 3 O_2(g) $\rightarrow$ 4 CO_2(g) + 2 H_2O(l)

Calculate:

(a) The internal energy of combustion of fumaric acid at 25 ^{o}C.

(b) The enthalpy of combustion of fumaric acid at 25 ^{o}C.

(c) The enthalpy of formation of fumaric acid at 25 ^{o}C.

Note: $\Delta_f H^o(CO_2)$ = -393.5 kJ/mol , $\Delta_f H^o(H_2O)$ = -285.8 kJ/mol

3.7 Classify as endothermic or exothermic:
(a) a combustion reaction in which the enthalpy of combustion is -2020 kJ/mol
(b) a solution process for which Δ_{soln} = +4.0 kJ/mol
(c) vaporization
(d) fusion
(e) sublimation

3.8 Calculate the standard enthalpy of formation of N_2O_5 from the following data:
Hint: Start by writing the equation representing the formation of 1 mole of N_2O_5 from the elements in their standard states.

(1) 2 NO + O_2 $\rightarrow$ 2 NO_2 ΔH_1 = -114.1 kJ
(2) 4 NO_2 + O_2 $\rightarrow$ 2 N_2O_5 ΔH_2 = -110.2 kJ
(3) N_2 + O_2 $\rightarrow$ 2 NO ΔH_3 = +180.5 kJ

3.9 From the following enthalpies of reaction,

N_2 + O_2 $\rightarrow$ 2NO ΔH = +180.7 kJ

4N_2 + 2O_2 $\rightarrow$ 4N_2O ΔH = +326.2 kJ

Calculate ΔH for the reaction, 2N_2O + O_2 $\rightarrow$ 4NO.

3.10 Use the enthalpy of formation data below to calculate the Fuel Value $C_6H_6(l)$, in
kJ/g

Compound	$\Delta_f H^o$
CO_2	-393.5 kJ/mol
H_2O	-285.8
$C_6H_6(l)$	+49.0

3.11 The Dietary Value of glycine, $C_2H_5NO_2(s)$ [M=75 g/mol], is 13.0 kJ/g. Calculate the enthalpy of combustion of glycine, in kJ/mol.

3.12 When 210. grams of 1-propanol [$CH_3(OH)CH_2CH_3$, M = 60. g/mol] is combusted
$O_2(g)$ to form $CO_2(g)$ + $H_2O(l)$, the heat involved is q = ΔH = -7,070 kJ

Calculate the Fuel Value (FV) of 1-propanol, in kJ/g.