In-Class Exercise for Chapter 19 – Solutions

1.) Estimate ΔH° for the following reaction given the data below:

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

Bond	Bond Energy (kJ/mol)	No. of bonds	No. of moles	Sign	ΔH (kJ/mol)
C-C	348	1	1	+	348
C-H	413	6	1	+	2478
O=O	497	1	7/2	+	1739.5
C=O	743	2	2	-	-2972
O-H	463	2	3	-	-2778
TOTAL					**-1180 kJ**

2.) Estimate ΔH° for the reaction given in problem 1 using the data below:

Compound	**ΔH (kJ/mol)**
$C_2H_{4(g)}$	52.3
$CO_{2(g)}$	-393.5
$H_2O_{(g)}$	-285.8

3(-285.8) + 2(-393.5) – 1(52.3) = **-1410 kJ**

3.) Determine ΔH° for the reaction using the data below:

$$ClO_{(g)} + O_{3(g)} \rightarrow Cl_{(g)} + 2O_{2(g)} \qquad \Delta H^o = -29.90\ kJ$$

$$2O_{3(g)} \rightarrow 3O_{2(g)} \qquad \Delta H^o = 24.18\ kJ$$

$$Cl_{(g)} + O_{3(g)} \rightarrow ClO_{(g)} + O_{2(g)}$$

$$Cl_{(g)} + 2O_{2(g)} \rightarrow ClO_{(g)} + O_{3(g)} \qquad \Delta H^o = 29.90\ kJ$$

$$2O_{3(g)} \rightarrow 3O_{2(g)} \qquad \Delta H^o = 24.18\ kJ$$

$$Cl_{(g)} + \cancel{2O_{2(g)}} + \cancel{2}O_{3(g)} \rightarrow ClO_{(g)} + \cancel{O_{3(g)}} + \cancel{3}O_{2(g)}$$

$$Cl_{(g)} + O_{3(g)} \rightarrow ClO_{(g)} + O_{2(g)} \qquad \Delta H^o = 29.90 + 24.18 = 54.08\ kJ$$

4.) Calculate the Gibb's free energy for the reaction using the information given below.

$$Fe_2O_{3(s)} + 3CO_{(g)} \rightarrow 2Fe_{(s)} + 3CO_{2(g)},\ \Delta H^o_{rxn} = -24.8\ kJ$$

Compound	**Entropy (J/K*mol)**
$Fe_2O_{3(s)}$	87.4
$CO_{(g)}$	197.6
$Fe_{(s)}$	27.3
$CO_{2(g)}$	213.6

Is the reaction endergonic or exogonic? Will it be spontaneous at high, low, or any temperature?

$$\Delta G^o_{rxn} = \Delta H^o_{rxn} - T\Delta S^o_{rxn}$$

$$\Delta S^o_{rxn} = \sum mS^o_{\text{products}} - \sum nS^o_{\text{reactants}}$$

$$\Delta S^o_{rxn} = \{2(27.3) + 3(213.6) - [87.4 + 3(197.6)]\} \, {}^{J}\!/_{K}$$

$$\Delta S^o_{rxn} = 15.2 \, {}^{J}\!/_{K}$$

$$\Delta G^o_{rxn} = -24.8kJ - 298K \times 15.2 \, {}^{J}\!/_{K} \times {}^{kJ}\!/_{1000J}$$

$$\Delta G^o_{rxn} = -29.3kJ \therefore exergonic$$

since $\Delta H^o_{rxn} < 0$ and $\Delta S^o_{rxn} > 0$ the reaction is spontaneous at all T

5.) Calculate the Gibb’s free energy for the reaction above given the following data:

Compound	**Free Energy (kJ/ mol)**
$Fe_2O_{3(s)}$	-742.2
$CO_{(g)}$	-137.2
$Fe_{(s)}$	0.0
$CO_{2(g)}$	-394.4

Is the reaction endergonic or exogonic? Will it be spontaneous at high, low, or any temperature?

$$\Delta G^o_{rxn} = \sum m\Delta G^o_{\text{f, products}} - \sum n\Delta G^o_{\text{f, reactants}}$$

$$\Delta G^o_{rxn} = \{2(0.0) + 3(-394.4) - [-747.2 + 3(-137.2)]\}kJ$$

$$\Delta G^o_{rxn} = -29.4kJ \therefore exergonic$$

without ΔH^o_{rxn} and ΔS^o_{rxn} we cannot determine the spontaneity dependence on T