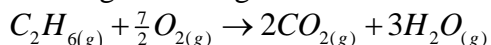


In-Class Exercise for Chapter 19 – Solutions

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



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

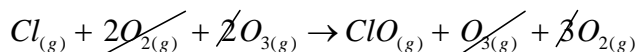
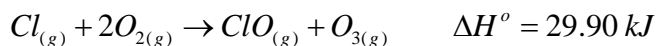
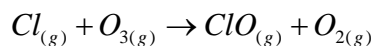
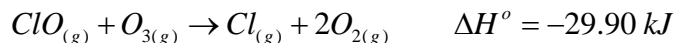
2.) Estimate ΔH° for the reaction given in problem 1 using the data below:

Compound ΔH (kJ/mol)

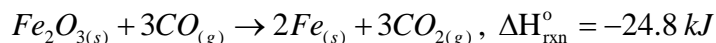


$$3(-285.8) + 2(-393.5) - 1(52.3) = \mathbf{-1410 \text{ kJ}}$$

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



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



Compound Entropy (J/Kmol*)**



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

$$\Delta G_{rxn}^{\circ} = \Delta H_{rxn}^{\circ} - T\Delta S_{rxn}^{\circ}$$

$$\Delta S_{rxn}^{\circ} = \sum mS_{products}^{\circ} - \sum nS_{reactants}^{\circ}$$

$$\Delta S_{rxn}^{\circ} = \{2(27.3) + 3(213.6) - [87.4 + 3(197.6)]\} J/K$$

$$\Delta S_{rxn}^{\circ} = 15.2 J/K$$

$$\Delta G_{rxn}^{\circ} = -24.8 kJ - 298 K \times 15.2 J/K \times kJ/1000 J$$

$$\Delta G_{rxn}^{\circ} = -29.3 kJ \therefore \text{exergonic}$$

since $\Delta H_{rxn}^{\circ} < 0$ and $\Delta S_{rxn}^{\circ} > 0$ 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?

$$\Delta G_{rxn}^{\circ} = \sum m\Delta G_{f, products}^{\circ} - \sum n\Delta G_{f, reactants}^{\circ}$$

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

$$\Delta G_{rxn}^{\circ} = -29.4 kJ \therefore \text{exergonic}$$

without ΔH_{rxn}° and ΔS_{rxn}° we cannot determine the spontaneity dependence on T