Gibbs Free Energy

1) Explain why reaction (I) has a negative ΔS° value whereas reaction (II) has a positive ΔS° value.

I) $2\ CH_3OH(g) + 3\ O_2(g) \rightarrow 2\ CO_2(g) + 4\ H_2O(l)$ $\quad \Delta S^\circ = -387$ J/molK
II) $2\ CH_3OH(g) + 3\ O_2(g) \rightarrow 2\ CO_2(g) + 4\ H_2O(g)$ $\quad \Delta S^\circ = 89$ J/molK

2) Which of the following reactions is thermodynamically favored at all temperatures? Justify your answer.

I) $2\ H_2O(l) \rightarrow 2\ H_2(g) + O_2(g)$ $\quad \Delta H^\circ = +572$ kJ
II) $8\ Fe(s) + 6\ O_2(g) \rightarrow 4\ Fe_2O_3(s)$ $\quad \Delta H^\circ = -3296.8$ kJ
II) $2\ CH_3OH(g) + 3\ O_2(g) \rightarrow 2\ CO_2(g) + 4\ H_2O(g)$ $\quad \Delta H^\circ = -1352$ kJ

3) Calculate ΔG° for the following reactions. Which of these processes are thermodynamically favored under standard conditions?
   a. $2\ H_2O(l) \rightarrow 2\ H_2(g) + O_2(g)$ $\quad \Delta H^\circ = +572$ kJ
   b. $8\ Fe(s) + 6\ O_2(g) \rightarrow 4\ Fe_2O_3(s)$ $\quad \Delta H^\circ = -3296.8$ kJ
   c. $2\ CH_3OH(g) + 3\ O_2(g) \rightarrow 2\ CO_2(g) + 4\ H_2O(l)$ $\quad \Delta H^\circ = -1538$ kJ

4) At what temperature does the reaction in question (3.a.) become thermodynamically favored? Assume that changes in temperature do not affect the ΔH and ΔS values.

5) How can the process in question (3.b.) be thermodynamically favored when the entropy of the system decreases so dramatically? Justify your answer.

6) The following questions pertain to the reaction below.

$$2\ H_2(g) + 2\ NO(g) \rightarrow N_2(g) + 2\ H_2O(g)$$

   a. Find the value of ΔG° for the reaction above using the ΔG_f° values in the appendix of your textbook.
   b. Is the reaction thermodynamically favorable under standard conditions? Justify your answer.
   c. Predict the sign associated with the ΔS value for the reaction above. Justify your answer.
   d. Predict the sign associated with the ΔH° value for the reaction above. Justify your answer.

7) The following questions pertain to the reaction below.

$$Al_2O_3(s) + 3\ H_2(g) \rightarrow 2\ Al(s) + 3\ H_2O(g)$$

a. Find the value of ΔG° for the reaction above using the ΔG_f° values in the appendix of your textbook.
b. Is the reaction thermodynamically favorable under standard conditions?
c. Predict the sign associated with the ΔS value for the reaction above. Justify your answer.

8) The following questions pertain to the reaction below.

$$2\ SO_2(g) + O_2(g) \rightarrow 2\ SO_3(g)$$

a. Predict the sign associated with the change in entropy, ΔS, for the reaction above. Justify your answer.
b. Does the change in entropy, ΔS, favour reactants or products? In other words, which condition would it prefer? Explain.
c. The ΔG value for this reaction is negative at low temperatures and positive at high temperatures. Predict the sign associated with the ΔH value for this reaction. Justify your answer.
d. Under what conditions does this process become thermodynamically favored? Justify your answer.

9) The following questions pertain to the reaction below.

$$Cu_2S(s) \rightarrow 2\ Cu(s) + S(s)$$

a. Predict the sign associated with the change in entropy, ΔS, for the reaction above. Justify your answer.
b. Does the change in entropy, ΔS, favour reactants or products? Explain.
c. The ΔG value for this reaction is positive at low temperatures and negative at high temperatures. Predict the sign associated with the ΔH value for this reaction. Justify your answer.
d. Under what conditions does this process become thermodynamically favored? Justify your answer.

10) For a reaction that occurred at 197.0°C, the enthalpy change, ΔH, was found to be +26.5 kJ/mol and the free energy change, ΔG, was found to be - 46 kJ/mol.
a. Find ΔS for this process as 197.0°C.
b. What is the principal force that is driving this reaction in the forward direction, ΔS or ΔH? Explain.
c. If the temperature of the system decreased dramatically, could this process become non-thermodynamically favored? Justify your answer.

11) For a certain process the entropy change, ΔS, was found to be -117.8 J/molK and the free energy change, ΔG, was found to be -27 kJ/mol at 320K.
a. Find ΔH for this process at 320 K.
b. What is the principal force that is driving this reaction in the forward direction, ΔS or ΔH? Explain.
c. If the temperature of the system increased dramatically, could this process become non-thermodynamically favored? Justify your answer.

12) For a certain reaction the enthalpy change, ΔH, was found to be - 423.5 kJ/mol and the free energy change, ΔG, was found to be - 349 kJ/mol, at 112°C. Find ΔS for the reaction at 112°C.

13) Predict the approximate magnitude of the ΔS value for the following reaction. Justify your prediction.

$$2\ CH_4(g) \rightarrow C_2H_6(g) + H_2(g)$$

14) At 375 K, the decomposition of copper (I) oxide is not a thermodynamically favored process. $2\ Cu_2O(s) \rightarrow 4\ Cu(s) + O_2(g)\ \ \Delta G_{375K} = 280.0$ kJ
If solid carbon is added to the system, the following reaction would occur.

$$2\ C(s) + O_2(g) \rightarrow 2\ CO(g)\ \ \Delta G_{375K} = -287.6 \text{ kJ}$$

Is the formation of pure copper from the decomposition of copper (I) oxide in the presence of solid carbon a thermodynamically favored process at 375 K? Justify your answer.

15) Due to the large positive value for ΔG°, zinc cannot be extracted from zinc sulfide through the following process. $ZnS(s) \rightarrow Zn(s) + S(s)\ \ \Delta G^\circ = 198.3$ kJ/mol
However, the combustion of sulfur is a thermodynamically favored process.

$$S(s) + O_2(g) \rightarrow SO_2(g)\ \ \Delta G^\circ = -300.1 \text{ kJ/mol}$$

Is the formation of pure zinc a thermodynamically favored process if the two reactions above are coupled? Justify your answer.