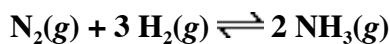


Gibbs Free Energy Problems
From the web and edited by S.L. Shultz

1. Calculate ΔH° and ΔS° for the following reaction and decide in which direction each of these factors will drive the reaction.

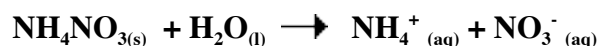
(HINT: Use Hess's Law: ΔH°_f , or S°_f , or $G^\circ_f = \sum H, \text{ or } S, \text{ or } G (\text{products}) - \sum H, S, G (\text{reactants})$)



Use these following formation values at standard temperatures:

Compound	ΔH°_f (kJ/mol)	S° (J/mol-K)
$\text{N}_2(\text{g})$	0	191.61
$\text{H}_2(\text{g})$	0	130.68
$\text{NH}_3(\text{g})$	-46.11	192.45

2a. Calculate ΔH° and ΔS° for the following reaction:



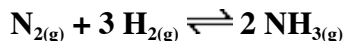
Compound	ΔH°_f (kJ/mol)	S° (J/mol-K)
$\text{NH}_4\text{NO}_3(\text{s})$	-365.56	151.08
$\text{NH}_4^+(\text{aq})$	-132.51	113.4
$\text{NO}_3^-(\text{aq})$	-205.0	146.4

2b. Use the results of this calculation to determine the value of ΔG° for this reaction at 25°C, and explain why NH_4NO_3 spontaneously dissolves in water at room temperature.

3. Use the following values of ΔH° and ΔS to predict whether the following reaction is spontaneous at 25°C:

$$\Delta H^\circ = -92.22 \text{ kJ} \quad (\text{favorable})$$

$$\Delta S^\circ = -198.75 \text{ J/K} \quad (\text{unfavorable})$$

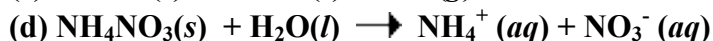
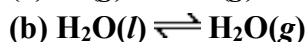
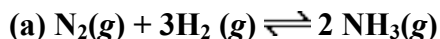


4. Predict whether the following reaction is still spontaneous at 500°C:



Assume that the values of ΔH° and ΔS° used in Problem 3 are still valid at this temperature.

5. Which of the following processes will lead to an increase in the entropy of the system?



6. Use $\Delta G^\circ = 32.96 \text{ kJ}$ to calculate the equilibrium constant for the following at 25°C, where R is the universal Gas Law Constant which is $R = 8.314 \text{ J/K}\cdot\text{mol}$. Make sure you convert to J or kJ depending on your preference in the units in the problem.

