**Name** _____ **AP Chemistry**

**HW 1: Due 3/11/15. Complete all free response and multiple choice questions. One question will be graded. Show all work. Box and clearly label all final free response answers.**

| Substance | Standard Entropy Joule °C$^{-1}$ mole$^{-1}$ |
|---|---|
| $N_2(g)$ | 192 |
| $H_2(g)$ | 130 |
| $NH_3(g)$ | 192 |

Ammonia can be produced by the following reaction: $N_2(g) + 3\ H_2(g) \Longleftrightarrow 2\ NH_3(g)$

The Gibbs free energy of formation $\Delta G_f°$ of $NH_3(g)$ is -16.5 kilojoules per mole.

(a) Calculate the value for $\Delta H°$ for the reaction above at 298 K.

(b) Can the yield of ammonia be increased by raising the temperature? Explain.

(c) What is the equilibrium constant for the reaction above at 298 K?

(d) If 235 milliliters of $H_2$ gas measured at 25°C and 570 millimeters Hg were completely converted to ammonia and the ammonia were dissolved in sufficient water to make 0.500 liter of solution, what would be the molarity of the resulting solution?

a. $\Delta S$: $N_2 + 3H_2 \rightarrow 2NH_3$

$192 + 3(130) \rightarrow 2(192)$

$582 \rightarrow 384$

$\Delta S = 384 - 582$

$\Delta S = -198\ J/K$

$\Delta G = \Delta H - T\Delta S$

$-33 = \Delta H - [(298)(-0.198)]$

$\boxed{\Delta H = -92.0\ KJ}$

b. No, the reaction is exothermic so an increase in temperature would decrease the amount of product. The reaction would shift its equillibrium position to the left.

c. $\Delta G = -RT\ln K$

$-33000 = (-8.314)(298)\ln K$

$13.3 = \ln K$

$\boxed{K = 6.08 \times 10^5}$

d. $n = \frac{PV}{RT}$ $570 \div 760 = 0.75$ atm

$n = \frac{(0.75)(0.235)}{(0.08206)(298)} = 0.0072$ moles $H_2$

$0.0072\ mol\ H_2 \left| \frac{2\ mol\ NH_3}{3\ mol\ H_2} \right. = 0.0048\ mol\ NH_3$ $M = \frac{0.0048}{0.50} =$

$\boxed{M = 0.0096M}$

| Compound | $\Delta H°_f$ (kilocalories/ mole) | S° (calories/mole K) |
|---|---|---|
| $H_2O$ (l) | -68.3 | 16.7 |
| $CO_2$ (g) | -94.1 | 51.1 |
| $O_2$ (g) | 0.0 | 49.0 |
| $C_3H_8$ (g) | ? | 64.5 |

When 1.000 gram of propane gas, $C_3H_8$, is burned at 25°C and 1.00 atmosphere, $H_2O$(l) and $CO_2$(g) are formed with the evolution of 50.33 kilojoules.

(a) Write a balanced equation for the combustion reaction.

(b) Calculate the molar enthalpy of combustion, $\Delta H°_{comb}$, of propane.

(c) Calculate the standard molar enthalpy of formation, $\Delta H°_f$, of propane gas.

(d) Calculate the entropy change, $\Delta S°_{comb}$, for the combustion reaction and account of the sign of $\Delta S°_{comb}$.

a. $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$

b. $\frac{1.000\,g}{50.33\,kJ} = \frac{44.11}{x}$ $\boxed{x = -2220.\,kJ}$

$-2220.\,kJ \div 4.184 = -530.6$ Kcal

c. $\Delta H = 3CO_2 + 4H_2O \rightarrow C_3H_8 + 5O_2$

$-530.6 = 3(-94.1) + 4(-68.3) \rightarrow C_3H_8 + 5O_2$

$-530.6 = -555.5 - x$

$\boxed{x = -24.9 \text{ Kcal/mol}}$

d. $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$

$64.5 + 5(49.0) \rightarrow 3(51.1) + 4(16.7)$

$309.5 \rightarrow 220.1$

$\Delta S = 220.1 - 309.5$

$\boxed{\Delta S = -89.4 \text{ cal/K}}$

The reason for the decrease in entropy is 6 moles of gas reactant goes to 3 moles of ~~pr~~ gas product. The positional probability decreases.