Name: ___________________________

Chem 101A
Lab Quiz 4

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1. For the reaction $2\ NH_3(g) \rightleftharpoons N_2(g) + 3\ H_2(g)$, when the equilibrium mixture is heated, the amount of NH_3 present decreases. State whether this reaction is exothermic or endothermic and explain your answer. (3 pts.)

Heating adds heat. This is shifting the reaction toward the right by using up NH_3. Therefore, heat is behaving like a reactant. This corresponds to an endothermic reaction.

2. For a 0.400 M solution of an unknown acid, the pH was found to be 2.35. Determine K_a for the acid. (3 pts.)

2 sig figs

$$[H^+] = 10^{-2.35} = 0.004467\ M = x$$

$$K_a = \frac{[H^+][A^-]}{[HA]_0 - x} = \frac{x^2}{0.400 - x}$$

$$= \frac{(0.004467)^2}{0.3955}$$

$$= 5.0 \times 10^{-5}$$

Given the following data:

Solution studied: 5.00 mL 6.0×10^{-3} M $Fe(NO_3)_3$
2.00 mL 6.0×10^{-3} M KSCN
3.00 mL H_2O

$$[Fe^{3+}]_0 = \frac{5.00\text{ mL}(6.0\times10^{-3}\text{ M})}{10.00\text{ mL}} = 3.0\times10^{-3}\text{ M}$$

$$[SCN^-]_0 = \frac{2.00\text{ mL}(6.0\times10^{-3}\text{ M})}{10.00\text{ mL}} = 1.2\times10^{-3}\text{ M}$$

Absorbance of solution at 450 nm: 0.870
In the equation $A = \varepsilon\ell c$, $\varepsilon\ell = 4.2 \times 10^3\ M^{-1}$

a. For the equilibrium that produces $FeSCN^{2+}$ from Fe^{3+} and SCN^-, calculate the equilibrium constant, K. (3 pts.)

$$K = \frac{[FeSCN^{2+}]}{[Fe^{3+}][SCN^-]} = \frac{x}{(3.0\times10^{-3}-x)(1.2\times10^{-3}-x)}$$

$$[FeSCN^{2+}] = \frac{A}{\varepsilon\ell} = \frac{0.870}{4.2\times10^3\ M^{-1}} = 2.071\times10^{-4}\text{ M} = x$$

$$K = \frac{2.071\times10^{-4}}{(2.793\times10^{-3})(9.93\times10^{-4})} = 7.47\times10^{1} \approx 7\times10^{1}$$

b. Determine the percentage of Fe^{3+} that reacted in this solution. (3 pts.)

$$\%\ Fe^{3+}\text{ reacted} = \frac{x}{[Fe^{3+}]_0}\times100\%$$

$$= \frac{2.071\times10^{-4}\text{ M}}{3.0\times10^{-3}\text{ M}}\times100\%$$

$$= 6.9\%$$

4. From the following data for KCl:

$\Delta H_{CLE} = -715$ kJ/mol $\Delta H_{\text{hydration energy}} = -684$ kJ/mol $\Delta H_f = -435.5$ kJ/mol

a. ~~Draw a partial~~ Born-Haber energy level diagram showing the energies of [K(s) + 1/2 Cl_2(g)], [K^+(g) + Cl^-(g)], [KCl(s)], and [K^+(aq) + Cl^-(aq)]. (3)

zero by def. → [K(s) + 1/2 Cl_2(g)]

E (kJ)

279.5 —— $K^+_{(g)} + Cl^-_{(g)}$

0 —— $K_{(s)} + \frac{1}{2} Cl_{2(g)}$

-404.5 —— $K^+_{(aq)} + Cl^-_{(aq)}$

-435.5 —— $KCl_{(s)}$

b. Determine ΔH_{soln} for KCl(s). (2)

ΔH_{SOLN} is for the process $KCl_{(s)} \rightarrow K^+_{(aq)} + Cl^-_{(aq)}$

$\Delta H_{SOLN} = (-404.5 - -435.5)$ kJ for each mole of KCl

$\Delta H_{SOLN} = 31$ kJ/mol

$$\ln P = -\frac{\Delta H_{vap}}{R}\left(\frac{1}{T}\right) + C$$

For an unknown liquid, a graph of lnP *vs*. 1/T gives a straight line with a slope of -3.67×10^3 K. Determine the value of ΔH_{vap} for this liquid. ($R = 8.3145\ J{\cdot}mol^{-1}{\cdot}K^{-1}$). (3)

$$-3.67 \times 10^3\ K = \frac{-\Delta H_{vap}}{R}$$

$$\Delta H_{vap} = (3.67 \times 10^3\ K)\ R$$

$$= (3.67 \times 10^3\ K)\left(8.3145 \times 10^{-3}\ \frac{kJ}{mol \cdot K}\right)$$

$$\Delta H_{vap} = 30.5\ \frac{kJ}{mol}$$