Individual Exam #2: Chemical and Physical Equilibrium

1) The Haber reaction ($N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$) is used commonly in industry, as ammonia is used in the production of fertilizer. Suppose we mixed 0.30 moles of N_2 and 0.25 moles of H_2 in a 1L container at 298 K. If the value of $K_p = 6.8\ x\ 10^{-5}$ at 298 K, how many moles of N_2, H_2, and NH_3 are present in the container at equilibrium? Round your answer to **three** decimal places.

Using Ideal Gas Law, $P_{N_2} = 7.34\ atm, P_{H_2} = 6.116\ atm, P_{NH_3} = 0\ atm$ initially.

$$N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$$

Reaction: $-x \quad -3x \quad +2x$

$$K_p = \frac{(2x)^2}{(7.34 - x)(6.116 - 3x)^3}$$

Using the given value of K_p, the technique of successive approximations, to iterate values of x to approach the real value of x:

X_0 = 0 atm

X_1 = 0.169 atm

X_2 = 0.1467 atm

X_3 = 0.1495 atm

X_4 = 0.1492 atm

X_5 = 0.1492 atm

Thus, x converges to 0.1492 atm. So final pressures are: $P_{N_2} = 7.1908\ atm, P_{H_2} = 5.6684\ atm, P_{NH_3} = 0.2984\ atm$

Using Ideal Gas Law again,

$$moles\ N_2 = 0.294$$

$$moles\ H_2 = 0.232$$

$$moles\ NH_3 = 0.00122$$

2) Silver chloride (AgCl) is an insoluble salt ($k_{sp} = 1.77x10^{-10}$) and is prepared by the reaction of sodium chloride (NaCl) and silver nitrate ($AgNO_3$) in solution. Once the silver chloride is prepared, ammonia can be added to the test tube, causing the dissolution of the precipitate and the formation of a dark blue complex known as diammine silver $[Ag(NH_3)_2]^+$. What is the solubility of silver chloride (in moles per liter) if we added $NH_{3(aq)}$ to the tube until it reached 1.50 M? The equilibrium constant of formation, K_f , for diamine silver is$1.6x10^7$. Give your answer to **three** decimal places.

$$Solvation\ of\ Silver\ chloride: AgCl_{(s)} \rightleftarrows Ag^+_{(aq)} + Cl^-((aq))$$

$$Formation\ of\ diammine\ silver: Ag^+_{(aq)} + 2NH_{3(aq)} \rightleftarrows [Ag(NH_3)_2]^+_{(aq)}$$

Combining these two reactions:

$$AgCl_{(s)} + 2NH_{3(aq)} \rightleftarrows [Ag(NH_3)_2]^+_{(aq)} + Cl^-_{(aq)}$$

Reaction: -x $\quad 1.5 - 2x \quad + x \quad + x$

$$K = K_{sp} * K_f = (1.77 * 10^{-10})(1.6 * 10^7) = 2.832 * 10^{-3}$$

$$K = \frac{x^2}{(1.5 - 2x)^2} = 2.832 * 10^{-3}$$

$$\frac{x}{1.5 - 2x} = 0.0532$$

$$x = 0.072\ M = solubility\ of\ AgCl$$

3) *Buffers (Adapted from Voet and Voet, Biochemistry, 5th ed.)*

Glycine is often used as a buffer in the purification of proteins. The neutral form of glycine can act as an acid, while the deprotonated form, glycine⁻, acts as the conjugate base. The pK_a of glycine is 9.60. Suppose you had to run a purification procedure using a glycine buffer, which at equilibrium is 0.1M in glycine and has a pH of 9.2. When you look for this in your storeroom, you find to your dismay that you only have large quantities of pH 9.0 and pH 10.0 solutions, both of which are also 0.1M in glycine at equilibrium. How much of the two should you mix to obtain 1.0L of your required solution?

Using the Henderson-Hasselbalch Equation ($pH = pK_a + \log\left(\frac{[glycine^-]}{[glycine]}\right)$, the ratio of glycine$^-$ to glycine can be found for any given pH.

At pH 9.2, $\frac{[glycine^-]}{[glycine]} = 0.398$
At pH 9.0, $\frac{[glycine^-]}{[glycine]} = 0.251$
At pH 10.0, $\frac{[glycine^-]}{[glycine]} = 2.51$
Since [glycine] = 0.1M for all the solutions, the [glycine$^-$] for each pH can be found:
At pH 9.0, [glycine$^-$] = 0.0251 M
At pH 10.0, [glycine$^-$] = 0.251 M

We want pH 9.2, where [glycine$^-$] = 0.0398 M. Using the two provided solutions, to reach a total volume of 1L,

$$0.0251x + 0.251(1 - x) = 0.0398$$
$$x = 0.935\ L$$

Thus, you should mix 935 mL of the pH 9 solution with 65 mL of the pH 10 solution.