More Equilibrium Practice

1. Write an equilibrium expression for each of the following unbalanced reactions:

   a. PCl_5 (g) ↔ PCl_3 (g) + Cl_2 (g)\ $K_c = \frac{[Cl_2][PCl_3]}{[PCl_5]}$

   b) $K_c = [H_2S][NH_3]$ b. NH_4HS (s) ↔ H_2S (g) + NH_3 (g)

   c. NH_4NO_3 (s) ↔ NH_4^+ (aq) + NO_3^- (aq) $K_c = [NH_4^+][NO_3^-]$

   d. HCOOH (aq) + H_2O (l) ↔ H_3O^+ (aq) + $HCOO^-$ (aq) $K_c = \frac{[H_3O^+][HCOO^-]}{[HCOOH]}$

   e. Bi_2S_3 (s) ↔ Bi^{3+} (aq) + S^{2-} (aq)

   $K_c = [Bi^{3+}][S^{2-}]$

2. The equilibrium constant K_{eq} for the following reaction is 11.5 at 300°C when the amounts of reactant and products are given in atmospheres. Suppose a tank initially contains PCl_5 with a pressure of 3.00 atm and 300°C. What is the equilibrium pressure of chlorine gas. $(Cl_2)_{eq} = ?$

eqn from a)

	$PCl_5(g)$ ↔	$PCl_3(g)$ +	$Cl_2(g)$
R			
I	3.00 atm	0	0
C	−x	+x	+x
E	3.00−x	x	x

$$11.5 = \frac{[PCl_3][Cl_2]}{[PCl_5]} = \frac{x^2}{3.00-x}$$

$$11.5(3.00-x) = x^2$$

$$34.5 - 11.5x = x^2$$

$$0 = x^2 + 11.5x - 34.5$$

→ $x = -13.9696$ or $x = 2.4696$, so:

$(PCl_5)_{eq} = 3.00 - 2.4696 = .53$ atm

$(PCl_3)_{eq} = (Cl_2)_{eq} = \boxed{2.47 \text{ atm}}$

3. An aqueous solution of ethanol and acetic acid, each with a concentration 0.810 M, is heated to 125°C. At equilibrium, the acetic acid concentration is 0.645 M. Calculate K_{eq} at 125°C for the reaction.

R	C_2H_5OH (aq) +	$HC_2H_3O_2$ (aq) ↔	H_2O (l) +	$CH_3CO_2C_2H_5$ (aq)
	Ethanol	acetic acid		ethyl acetate
I	0.810 M	0.810 M	/	0
C	−0.165 M	−0.165 M		+0.165 M
E	0.645 M	0.645 M		0.165 M

$$K_c = \frac{[CH_3CO_2C_2H_5]}{[C_2H_5OH][HC_2H_3O_2]} = \frac{.165}{0.645^2} = \boxed{0.397}$$

4. The equilibrium constant for the following reaction

R H_2 (g) + I_2(g) ↔ 2HI (g)

is determined to be 57.85 at 450°C. If 1.50 mol of each reactant is placed in a 2.00 L flask at 450°C what are the concentrations of H_2, I_2, and HI when equilibrium has been achieved?

	H_2	I_2	HI
I	0.750 M	0.750 M	0
C	−x	−x	+2x
E	0.750−x	0.750−x	2x

$$K_c = \frac{[HI]^2}{[H_2][I_2]} = \frac{(2x)^2}{(.750-x)^2} = 57.85$$

Take √ of both sides →

$$\frac{2x}{0.750-x} = \sqrt{57.85}$$

$$2x = \sqrt{57.85}(.750-x)$$

$$2x = 43.388 - 57.85x$$

$$59.85x = 43.388$$

$$x = .725 \text{ M}$$

→ $[HI]_{eq} = 2(.725) = 1.45$ M

$[H_2]_{eq} = [I_2]_{eq} = .750 - .725 = .25$ M

5. Ammonium hydrogen sulfide decomposes on heating.

R NH_4HS (s) ↔ H_2S (g) + NH_3 (g)

If K_{eq} is 0.11 at 25°C when the partial pressures are expressed in atmospheres what is the total pressure in the flask at equilibrium?

	NH_4HS	H_2S	NH_3
I	/	0	0
C		+x	+x
E		x	x

$K_{eq} = 0.11 = (H_2S)(NH_3) = x^2$

$.33$ atm $= x$, so total P = .33 + .33 = .66

6. The equilibrium constant for the dissociation of iodine molecules to iodine atoms is 3.76×10^{-3} at 1000K. Suppose 0.150 mol of I_2 is placed in a 15.5 L flask at 1000K. What are equilibrium concentrations of all species present.

	$I_2(g)$ $\leftrightarrow$	$2I(g)$
R		
I	.00968 M	0
C	−x	+2x
E	.00968−x	2x

$K_c = 3.76\times10^{-3} = \frac{[I]^2}{[I_2]} = \frac{(2x)^2}{.00968-x}$

$3.76\times10^{-3}(.00968-x) = 4x^2$

$\rightarrow 3.64\times10^{-5} - (3.76\times10^{-3})x = 4x^2$

$0 = 4x^2 + (3.76\times10^{-3})x - 3.64\times10^{-5}$

$x = -.003523$ or $+.002583$

so $[I]_{eq} = 2\times .002583 = .00517$ M

$[I_2]_{eq} = .00968 - .002583 = .00451$ M

7. Hemoglobin (Hb) can form a complex with both O_2 and CO. For the reaction

R $HbO_2(aq)$ + $CO(g)$ $\leftrightarrow$ $HbCO(aq)$ + $O_2(g)$

at body temperature, K is about 2.0×10^2. If the ratio $\frac{[HbCO]}{[HbO_2]}$ comes close to 1.0, death is probable. What partial pressure of CO in the air is likelv to be fatal? Assume the partial pressure of O_2 is 0.20 atm.

$K = 2.0\times10^2 = \frac{\overset{1}{\cancel{[HbCO]}}\ O_2}{\cancel{[HbO_2]}\ CO}$

$2.0\times10^2 = \frac{(O_2)}{(CO)} = \frac{.20}{(CO)}$

$\rightarrow 2.0\times10^2(CO) = .20$

$\boxed{(CO) = .0010 \text{ atm}}$

8. K_{eq} for the decomposition of ammonium hydrogen sulfide (NH4HS) is 1.8×10^{-4} at 15°C. see 1b) for reaction

   a. What are the equilibrium concentrations of NH_3 and H_2S when 5.00 grams of pure salt decomposes in a sealed 3.0 L flask at 15°C

	$NH_4HS(s) \leftrightarrow$	$H_2S(g)$ +	$NH_3(g)$
R			
I		0	0
C		+x	+x
E		x	x

$1.8\times10^{-4} = [H_2S][NH_3] = x^2$

$\sqrt{1.8\times10^{-4}} = x = .013$ M $= [NH_3]_{eq} = [H_2S]_{eq}$

   b. What are the equilibrium concentrations when 10.0 g of the pure salt decomposes in the sealed flask?

Same concentrations! .013 M = $[H_2S] = [NH_3]$