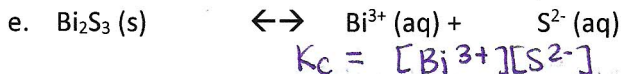
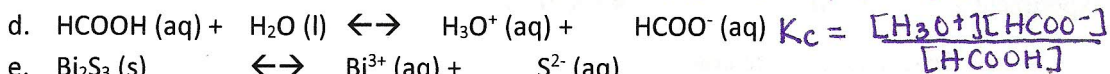
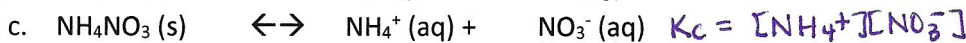
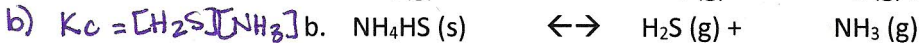
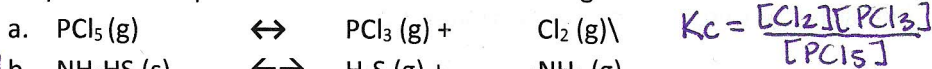
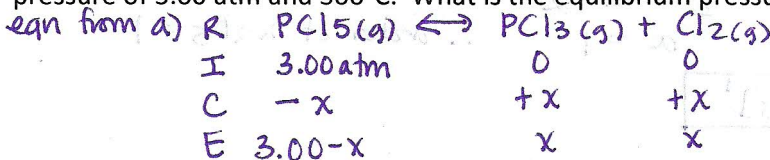


### More Equilibrium Practice

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



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  $\text{PCl}_5$  with a pressure of 3.00 atm and 300°C. What is the equilibrium pressure of chlorine gas.  $(\text{Cl}_2)_{eq} = ?$



$$11.5 = \frac{[\text{PCl}_3][\text{Cl}_2]}{[\text{PCl}_5]} = \frac{x^2}{3.00-x}$$

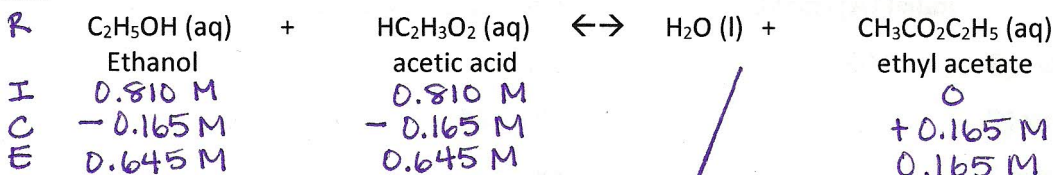
$$11.5(3.00-x) = x^2$$

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

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

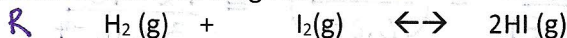
$x = -13.9696$  or  $x = 2.4696$ , so:  
 $(\text{PCl}_5)_{eq} = 3.00 - 2.4696 = .53 \text{ atm}$   
 $(\text{PCl}_3)_{eq} = (\text{Cl}_2)_{eq} = \boxed{2.47 \text{ atm}}$

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.

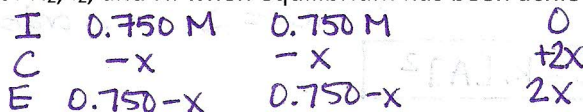


$$K_c = \frac{[\text{CH}_3\text{CO}_2\text{C}_2\text{H}_5]}{[\text{C}_2\text{H}_5\text{OH}][\text{HC}_2\text{H}_3\text{O}_2]} = \frac{.165}{0.645^2} = \boxed{0.397}$$

4. The equilibrium constant for the following reaction



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  $\text{H}_2$ ,  $\text{I}_2$ , and  $\text{HI}$  when equilibrium has been achieved?



$[\text{HI}]_{eq} = 2(.725) = 1.45 \text{ M}$   
 $[\text{H}_2]_{eq} = [\text{I}_2]_{eq} = .750 - .725 = .025 \text{ M}$

$$K_c = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]} = \frac{(2x)^2}{(.750-x)^2} = 57.85$$

Take  $\sqrt{\quad}$  of both sides

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

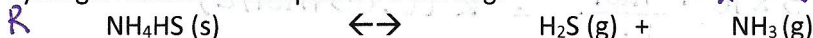
$$2x = 57.85(0.750-x)$$

$$2x = 43.388 - 57.85x$$

$$59.85x = 43.388$$

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

5. Ammonium hydrogen sulfide decomposes on heating.



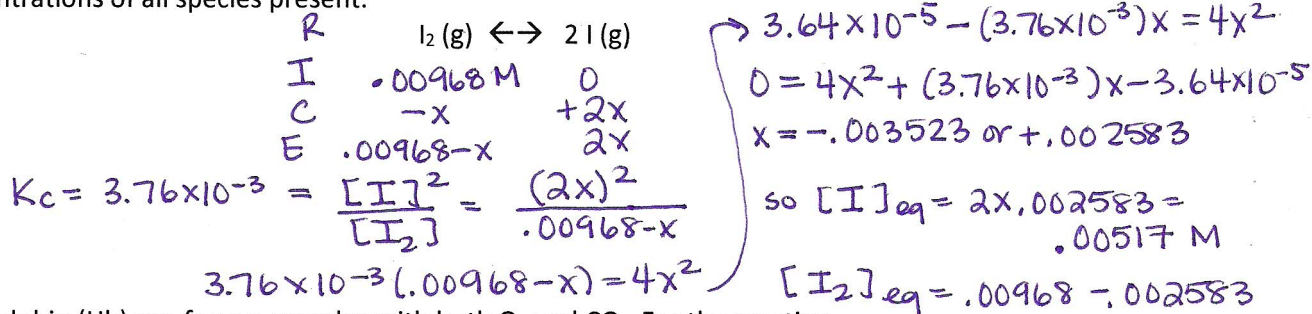
If  $K_{eq}$  is 0.11 at 25°C when the partial pressures are expressed in atmospheres what is the total pressure in the flask at equilibrium?



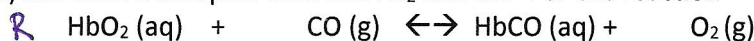
$$K_{eq} = 0.11 = (\text{H}_2\text{S})(\text{NH}_3) = x^2$$

$$.33 \text{ atm} = x, \text{ so total } P = .33 + .33 = .66$$

6. The equilibrium constant for the dissociation of iodine molecules to iodine atoms is  $3.76 \times 10^{-3}$  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.



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



at body temperature, K is about  $2.0 \times 10^2$ . If the ratio  $\frac{[HbCO]}{[HbO_2]}$  comes close to 1.0, death is

probable. What partial pressure of CO in the air is likely to be fatal? Assume the partial pressure of  $O_2$  is 0.20 atm.

$$K = 2.0 \times 10^2 = \frac{[HbCO] \cdot O_2}{[HbO_2] \cdot CO}$$

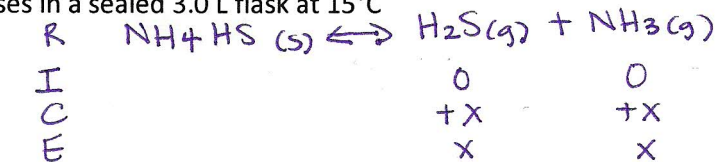
$$2.0 \times 10^2 = \frac{(O_2)}{(CO)} = \frac{.20}{(CO)}$$

$$2.0 \times 10^2 (CO) = .20$$

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

8.  $K_{eq}$  for the decomposition of ammonium hydrogen sulfide <sup>see 1b) for reaction</sup> is  $1.8 \times 10^{-4}$  at  $15^\circ C$ .

- 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  $15^\circ C$



$$1.8 \times 10^{-4} = [H_2S][NH_3] = x^2$$

$$\sqrt{1.8 \times 10^{-4}} = x = .013 \text{ M} = [NH_3]_{eq} = [H_2S]_{eq}$$

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

same concentrations!  $.013 \text{ M} = [H_2S] = [NH_3]$