a. $\frac{[NH_4^+][OH^-]}{[NH_4OH]} = K_{eq}$

b. $K_{eq} = \frac{[H^+][C_2H_3O_2^-]}{[HC_2H_3O_2]}$

c. $K_{eq} = \frac{[H_2O][NH_4^+][C_2H_3O_2^-]}{[NH_4OH][HC_2H_3O_2]}$

d. $K_{eq} = \frac{[Na^+][Cl^-][HCN]}{[Na^+][CN^-][H^+][Cl^-]}$

e. $\frac{[Al^{3+}][Cl^-]}{[AlCl_3]} = K_{eq}$

f. $\frac{[H^+][HS^-]}{[H_2S]} = K_{eq}$

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g. $K_{eq} = \frac{[FeS][H^+]^2}{[Fe^{2+}][H_2S]}$

## CHEMICAL EQUILIBRIUM

1. Write the expression for the equilibrium constant for each of the following equilibria:

a) $NH_4OH \leftrightarrow NH_4^+ + OH^-$ Keq =

b) $HC_2H_3O_2 \leftrightarrow H^+ + C_2H_3O_2^-$ Keq =

c) $NH_4OH + HC_2H_3O_2 \leftrightarrow H_2O + NH_4^+ + C_2H_3O_2^-$ Keq =

d) $Na^+ + CN^- + H^+ + Cl^- \leftrightarrow Na^+ + Cl^- + HCN$ Keq =

e) $AlCl_3 \leftrightarrow Al^{+++} + 3Cl^-$ Keq =

f) $H_2S \leftrightarrow H^+ + HS^-$ Keq =

g) $Fe^{++} + H_2S \leftrightarrow FeS + 2H^+$ Keq =

h) $4NH_3 + 5O_2 \leftrightarrow 4NO + 6H_2O$ Keq =

h. $K_{eq} = \frac{[NO]^4[H_2O]^6}{[NH_3]^4[O_2]^5}$

2. If an equilibrium reaction releases energy in the form of heat in the forward reaction, what is the effect of increasing the temperature? If the equilibrium system releases heat in the reverse reaction, what is the effect of an increase in temperature?

Forward rxn increasing temp- reaction shifts left reverse rxn increase temp rxn shifts right

3. What effect would an increase in pressure have on the following? ...a decrease in pressure?

| | $N_2$ | + $3H_2$ | $\leftrightarrow$ $2NH_3$ | |
|---|---|---|---|---|
| increase pressure | ↓ | ↓ | ↑ | shift right |
| decrease pressure | ↑ | ↑ | ↓ | left |

4. Explain the effect of each of the following factors on the equilibrium:

$N_2 + 3H_2 \leftrightarrow 2NH_3 + \text{heat}$

a. increase temperature shift left
b. add more $H_2$ shift right
c. add more $NH_3$ shift left
d. decrease pressure shift left
e. placing all reactants into 1/2 the volume shift right

5. a. At a given temperature, the $K_{eq}$ for the gas phase reaction: $H_2 + I_2 \leftrightarrow 2HI$ is 4 X 10 -6. If the concentrations of $H_2$ and $I_2$ at equilibrium are 2 X 10 -4, find the concentration of HI.

$K_{eq} = \frac{[HI]^2}{[H_2][I_2]}$

$4 \times 10^{-7} = [HI]$ ~~4x10^-4 Keq~~

b. At a given temperature, the reaction $CO(g) + H_2O(g) \leftrightarrow H_2(g) + CO_2(g)$, all gases, produced the following concentrations: $H_2O = 0.5$ M ; $CO = 0.2$ M ; $H_2 = 3.2$ M; $CO_2 = 4.2$ M. Find the value of $K_{eq}$ at that temperature.

$K_{eq} = \frac{[CO_2][H_2]}{[CO][H_2O]} = \frac{4.2M(3.2M)}{.2M(.5M)} = [4 \times 10^6 M]$ $[2 \times 10^{-4} M]$

$134 = K_{eq}$

c. If the temperature of the reactants in part b is changed, the $K_{eq}$ becomes 2.4 X 10 -3. If all of the concentrations except CO are then adjusted to the values given in b, what is the new concentration of CO?

$2.4 \times 10^{-3} = \frac{(4.2)^2 3.2}{(.5)}$ ← should X be on the bottom? oops!

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$X = [CO] = 7.5 \times 10^{-5} M$

6. Given: $2NH_3 + H_2SO_4 \leftrightarrow (NH_4)_2SO_4$
   a. What is the equilibrium constant expression? $K_{eq} = \frac{[(NH_4)_2SO_4]}{[NH_3]^2[H_2SO_4]}$
   b. If the concentration (in moles per liter) after the reaction of $NH_3$, $H_2SO_4$, and $(NH_4)_2SO_4$ are 2,3,4, respectively, what is $K_{eq}$? $K_{eq} = \frac{4M}{(2M)^2(3)} = .33$
   c. If the concentration of $NH_3$ is increased, what happens to the equilibrium? shift right
   d. If pressure is added, what happens to the equilibrium? shift right
   e. If heat is added, what happens to $K_{eq}$? decreases assuming reaction is endothermic

7. The following reaction goes to completion because of the formation of un-ionized water:
   $2NO + H_2 \longrightarrow N_2O + H_2O + 87{,}030$ cal
   What effect would an increase in temperature have on its rate?
   the reaction would shift left

8. The equilibrium constant for the reaction $N_2 + O_2 \leftrightarrow 2NO$ is $2.6 \times 10^{-4}$ at 4000°C. What would be the concentration in moles per liter of nitrogen and oxygen at equilibrium if 0.25 mole of NO were placed in a 1 liter container and then heated to 4000°C? $2.6\times10^{-4} = \frac{[NO]^2}{[N_2][O_2]}$

9. One liter of HI was heated to 500°C under constant pressure until equilibrium was established according to the equation $2HI \leftrightarrow H_2 + I_2$. Analysis showed the following concentrations: $[H_2] = 0.42$ M, $[I_2] = = 0.42$ M, $[HI] = 3.52$ M Calculate the value of $K_{eq}$. $K_{eq} = \frac{[H_2][I_2]}{[HI]^2} = \frac{.42(.42)}{3.52^2} = .014$

$2.6\times10^{-4}\frac{[.25]^2}{x^2}$

$x^2 = 240$

$[N_2] = [O_2] = 16M$

10. Three reactions exist in equilibrium:

| | Keq |
|---|---|
| $N_2 + 3H_2 \leftrightarrow 2NH_3$ | $1.7 \times 10^{-6}$ |
| $Cl_2 + H_2 \leftrightarrow 2HCl$ | $3.4 \times 10^{-2}$ |
| $F_2 + H_2 \leftrightarrow 2HF$ | $5.1 \times 10^{-4}$ |

   a. Which reaction has more products present at equilibrium? HCl
   b. Which reaction has more reactants present at equilibrium? $NH_3$

11. The following is the potential energy diagram for the reversible reaction
   $N_2 + O_2 \leftrightarrow 2NO$
   a. Is the forward or reverse reaction favored by an increase in temperature? Explain. reverse
   b. Explain which reaction is favored because of energy of activation. forward
   c. Explain the effect of a catalyst on this curve. lower activation energy curve would lower