Dougherty Valley HS Chemistry
Equilibrium – Mixed Practice and Ksp

Worksheet #7

Name: **Period:** **Seat#:**

Directions: Complete the following chart by choosing from the following options:
Equilibrium Shift: *left, right, no change* Δ [ ] / Temp: *increase, decrease, no change* Δ Keq: *no, yes*

$A_{(g)} + B_{(aq)} \leftrightarrow C_{(s)}$ **$\Delta H_{rxn} = 453$ kJ**

Stressor	Eq. Shift	Δ [A]	Δ [B]	Δ [C]	Δ Temp	Δ Keq
1) The pressure of A in the reaction chamber is increased						
2) The temperature of the reaction is increased by 20°C						
3) A catalyst is added to the system						
4) As the reaction progresses, more of compound B is steadily added to the reaction chamber						
5) An inhibitor is added to the reaction chamber						
6) Argon gas is added to the reaction chamber, doubling the pressure						

7) Why are solids and liquids not included in Keq expressions?

8) Write the equilibrium constant expression for K_{eq}
$4H_3O^+_{(aq)} + 2Cl^-_{(aq)} + MnO_{2(s)} \leftrightarrow Mn^{2+}_{(aq)} + 6H_2O_{(l)} + Cl_{2\ g)}$

9) Suppose that for the unbalanced reaction: ____$N_{2\ (g)}$ + ____$Cl_{2\ (g)}$ ↔ ____$NCl_{3\ (g)}$ it is determined that, at a particular temperature, the equilibrium concentrations are $[N_2]$ = 0.000104 M, $[Cl_2]$ = 0.000201 M, and $[NCl_3]$ = 0.141 M. Calculate that value of K_{eq} for the reaction at this temperature.

10) Gaseous phosphorus pentachloride decomposes to chlorine gas and gaseous phosphorus trichloride. In a certain experiment, at a temperature where K_{eq} = 8.96E-2, the equilibrium concentrations of PCl_5 and Cl_2 were found to be 5.67E-3M and 0.233M, respectively. Calculate the concentration of PCl_3 present at equilibrium.

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11) $C_2H_4\,_{(g)} + H_2\,_{(g)} \leftrightarrow C_2H_6\,_{(g)}$ $K_c = 0.99$ The initial concentrations of $C_2H_4\,_{(g)}$, is 0.335M and that of hydrogen is 0.526M, and there is no C_2H_6 at the start. Which direction will the reaction shift to reach equilibrium (show the calculation – Hint: K vs Q)? What is the concentration for each substance at equilibrium? *[C_2H_4] = 0.236 M, [H_2] = 0.427 M, [C_2H_6] = 0.0995 M*

12) $2NO\,_{(g)} + 2H_2\,_{(g)} \leftrightarrow N_2\,_{(g)} + 2H_2O\,_{(g)}$ Determine the value of the equilibrium constant, K_c, for the reaction. Initially a mixture of 0.100 M NO, 0.050 M H_2, 0.100 M H_2O was allowed to reach equilibrium (initially there was no N_2). At equilibrium the concentration of NO was found to be 0.062 M *6.5E2*

13) $N_2O_4\,_{(g)} \leftrightarrow 2NO_2\,_{(g)}$ A reaction flask is charged with 3.00 atm of dinitrogen tetraoxide gas and 2.00 atm of nitrogen dioxide gas. At 25°C, the gases are allowed to reach equilibrium. The pressure of the nitrogen dioxide was found to have decreased by 0.952 atm. Estimate the value of K_p for this system. *3.16*

14) $HSO_4^-{}_{(aq)} + H_2O_{(l)} \leftrightarrow H_3O^+{}_{(aq)} + SO_4^{2-}{}_{(aq)}$ K = 1.2×10^{-5}

The initial concentrations are $[HSO_4^-] = 0.50M$, $[H_3O^+] = 0.020M$, $[SO_4^{2-}] = 0M$.

a. Which way would the reaction shift to reach equilibrium? (show the calculation – Hint: K vs Q)

b. What are the equilibrium concentrations of the products and the reactants?

15) For the unbalanced reaction: ____$N_{2\,(g)}$ + ____ $H_{2\,(g)} \leftrightarrow$ ____$NH_{3\,(g)}$ show your calculations for K_{eq} given the following equilibrium concentrations during each of the three experiments below:

Experiment #1	$[N_2]$ = 0.921 M	$[H_2]$ = 0.763 M	$[NH_3]$ = 0.157 M
Experiment #2	$[N_2]$ = 0.399 M	$[H_2]$ = 1.197 M	$[NH_3]$ =0.203 M
Experiment #3	$[N_2]$ = 2.59 M	$[H_2]$ = 2.77 M	$[NH_3]$ =1.82 M

What does the value of K_{eq} tell you about the relative concentrations of the reactants versus the products at equilibrium? Explain why. (Hint: Large K_{eq} vs Small K_{eq})

Ksp – Solubility Product Constant

The solubility product constant, Ksp, is the equilibrium constant for a solid substance dissolving in an aqueous solution. It represents the level at which a solute dissolves in solution. The more soluble a substance is, the higher the Ksp value it has.

In order to determine the Ksp of a substance you need to write the dissociation reaction first. You write the Ksp equation the same way as a normal Keq - products over reactants, and solids do not factor into the equations.

Do not forget to balance your dissociation reactions! The coefficients are the exponents in your Ksp equations so the balancing matters!

Examples:

$Sn(OH)_2$ (s) ↔ Sn^{2+} (aq) + $2OH^-$ (aq) $K_{sp} = [Sn^{2+}][OH^-]^2$

Ag_2CrO_4 (s) ↔ $2Ag^+$ (aq) + CrO_4^{2-} (aq) $K_{sp} = [Ag^+]^2 [CrO_4^{2-}]$

$Fe(OH)_3$ (s) ↔ Fe^{3+} (aq) + $3OH^-$ (aq) $K_{sp} = [Fe^{3+}] [OH^-]^3$

Balanced Dissociation Equation	Ksp Expression
16) Lead (II) Chloride (s)	
17) Silver dichromate (s)	
18) Strontium phosphate (s)	
19) $Cu_3(PO_4)_2$ (s)	
20) CuSCN (s)	
21) CdS (s)	

22) Explain the differences between equilibrium constant (Keq) and solubility product constant (Ksp)

23) If iron (III) phosphate (s) has a Ksp value of 1.3E-22, do you expect it to be considered "soluble" or "insoluble" on an old fashioned solubility chart like we used first semester (R-25)? Justify your answer.