

Name: _____

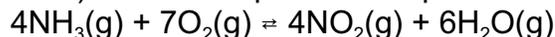
(4 points)

Chemistry 114 Fourth Hour Exam

Remember- Show all work for partial credit

1.

A. (3 points) Write an equilibrium expression for the reaction:



$$K = \frac{[\text{NO}_2]^4 [\text{H}_2\text{O}]^6}{[\text{NH}_3]^4 [\text{O}_2]^7}$$

B. (3 points) If the above reaction makes X moles of NO₂:How many moles of H₂O are formed?

$$X \text{ moles NO}_2 \times \frac{6 \text{ mole H}_2\text{O}}{4 \text{ moles NO}_2} = 1.5 X$$

How many moles of O₂ are consumed?

$$X \text{ moles NO}_2 \times \frac{7 \text{ moles O}_2}{4 \text{ moles NO}_2} = 1.75 X$$

How many mole of NH₃ are consumed?

$$X \text{ moles NO}_2 \times \frac{4 \text{ moles NH}_3}{4 \text{ moles NO}_2} = 1 X$$

C. (3 points) If you start with all reactants at 1M concentrations, and all products at 0M concentrations, fill in the following ICE table:

Initial 1 1 0 0 Change -X -1.75X X +1.5X Equilibrium 1-X 1-1.75X X 1.5X D. (3 points) Assume that K_{eq} for this reaction is 3.59x10⁻⁷. Write the equilibrium expression for the final system as the 1M initial concentrations of reactants come to equilibrium with products. (Do NOT try to solve!)

$$3.59 \times 10^{-7} = \frac{X^4 (1.5X)^6}{(1-X)^4 (1-1.75X)^7}$$

2. (12 points) I have a container in which the following chemical reaction has reached equilibrium: $2\text{HI}(\text{g}) \rightleftharpoons \text{H}_2(\text{g}) + \text{I}_2(\text{g})$

Will the equilibrium for this reaction shift to the right, left, or remain unchanged as :

HI is added? Right

H_2 is added? Left

I_2 is removed? Right

Ar(g) is added? No change

Volume of the container is doubled No change

The temperature is decreased (this reaction is exothermic) Right

3. (12 points)

A.) What is the pH and pOH of a solution that is $6.38 \times 10^{-3} \text{M}$ HCl.

$$\begin{aligned} \text{pH} &= -\log[\text{H}^+] \\ &= -\log(6.38 \times 10^{-3}) \\ &= -(-2.195) \\ &= 2.195 \end{aligned}$$

$$\begin{aligned} 14 &= \text{pH} + \text{pOH} \\ 14 &= 2.195 + X \\ 14 - 2.195 &= X \\ \text{pOH} &= X = 11.805 \end{aligned}$$

B.) If a solution has a pH of 8.46, what is the concentration of $\text{Ca}(\text{OH})_2$ in this solution?

$$\begin{aligned} 14 &= \text{pH} + \text{pOH} \\ 14 &= 8.46 + \text{pOH} \\ 14 - 8.46 &= \text{pOH} \\ \text{pOH} &= 5.54 \\ \text{pOH} &= 5.54 = -\log[\text{OH}^-] \\ -5.54 &= \log[\text{OH}^-] \\ 10^{-5.54} &= [\text{OH}^-] = 2.9 \times 10^{-6} \end{aligned}$$

But for every mole of $\text{Ca}(\text{OH})_2$ I get 2 moles of OH^- so:

$$\frac{2.9 \times 10^{-6} \text{ mole } \text{OH}^-}{\text{liter}} \times \frac{1 \text{ mole } \text{Ca}(\text{OH})_2}{2 \text{ moles } \text{OH}^-} = 1.45 \times 10^{-6} \text{ M } \text{Ca}(\text{OH})_2$$

C.) What is the pH of a $4.3 \times 10^{-8} \text{M}$ HClO_4 ?

This is a strong acid at a concentration less than 10^{-7} , so the $[\text{H}^+]$ due to water itself is significant in this problem. The quick and dirty (but not entirely correct) answer goes like this:

$$\begin{aligned} [\text{H}^+] \text{ due to water} &= 1 \times 10^{-7} \\ [\text{H}^+] \text{ due to strong acid} &= 4.38 \times 10^{-8} \\ \text{Total } [\text{H}^+] &= 1 \times 10^{-7} + 4.38 \times 10^{-8} = 1.438 \times 10^{-7} \\ \text{pH} &= -\log(1.438 \times 10^{-7}) \\ &= 6.84 \end{aligned}$$

4. (12 points) Are the following compounds going to acidic, basic, or neutral?

CH_3COOK Basic

NaCl Neutral

$\text{Sn}(\text{NO}_3)_4$ Acidic

PO_3 Acidic

5. (12 points) Define the following terms:

Arrhenius Acid

A compound that produces H^+ in an aqueous solution.

Brønsted-Lowry Base

A proton acceptor.

Lewis Acid

An electron acceptor.

Positional Entropy

Entropy calculated statistically from the number of states of a system.

Entropy

S, a measure of randomness.

Free Energy

Energy available to do work or $G = H + TS$.

6. (12 points) 50 mls of 0.1M Sodium Acetate is mixed with 75 mls of .3M Acetic Acid. What is the pH of this solution? (The pKa of acetic acid is 4.75)

First, find the concentrations of both species in the final mixed solution using $M_1V_1 = M_2V_2$:

$$[\text{Acetic Acid}] \quad 75(.3) = 125(X); X = 75(.3)/125 = .18\text{M}$$

$$[\text{Sodium Acetate}] \quad 50(.1) = 125(X); X = 50(.1)/125 = .04\text{M}$$

Recognizing that we have a mixture of an acid and its conjugate base we can use the Henderson Haselbach equation:

$$\text{pH} = \text{pKa} + \log \frac{[\text{A}^-]}{[\text{HA}]}$$

$$= 4.75 + \log (.04/.18)$$

$$= 4.10$$

7. (12 points) Given the values of ΔH and ΔS , which of the following changes will be spontaneous at constant T and P?

a. $\Delta H = + 25 \text{ kJ}$, $\Delta S = -100 \text{ J/K}$, $T=300 \text{ K}$

+ ΔH and - ΔS , never spontaneous at any temperature

b. $\Delta H = + 25 \text{ kJ}$, $\Delta S = +5.0 \text{ J/K}$, $T=300 \text{ K}$

$\Delta G = 25,000\text{J} - 300(5)\text{J} = 25000 - 1500 = +23,500$ Not spontaneous at this temp

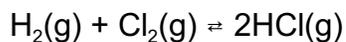
c. $\Delta H = -10 \text{ kJ}$, $\Delta S = +5.0 \text{ J/K}$, $T=298 \text{ K}$

- ΔH and + ΔS , always spontaneous at all temperatures

d. $\Delta H = -10 \text{ kJ}$, $\Delta S = -40 \text{ J/K}$, $T=200\text{K}$

$\Delta G = -10,000\text{J} - 200(-40)\text{J} = -10,000 + 8,000 = -2,000$ Spontaneous at this temp

8. (12 points) Consider the reaction



Calculate ΔH° , ΔS° , ΔG° , and K at 298K using the data below:

Substance/State	ΔH° (kJ/mol)	ΔS° (J/K·mol)	ΔG° (kJ/mol)
H ₂ (g)	0	131	0
H(g)	217	115	203
H ⁺ (aq)	0	0	0
Cl ₂ (g)	0	223	0
Cl ₂ (aq)	-23	121	7
Cl ⁻ (aq)	-167	57	-131
HCl(g)	-92	187	-95

$$\begin{aligned}\Delta H &= \text{Sum of Products} - \text{Sum of reactants} \\ &= 2(-92) - [1(0) + 1(0)] \\ &= -184 \text{ kJ/mol}\end{aligned}$$

$$\begin{aligned}\Delta S &= \text{Sum of Products} - \text{Sum of reactants} \\ &= 2(187) - [1(131) + 1(223)] \\ &= 20 \text{ J/mol}\end{aligned}$$

$$\begin{aligned}\Delta G &= \text{Sum of Products} - \text{Sum of reactants} \\ &= 2(-95) - [1(0) + 1(0)] \\ &= -190 \text{ kJ/mol}\end{aligned}$$

-or-

$$\begin{aligned}\Delta G &= \Delta H - T\Delta S \\ &= -184000 \text{ J/mol} - [298(20\text{J/mol})] \\ &= -184000 \text{ J/mol} - 5960\text{J/mol} \\ &= -189960\text{J/mol} = -190 \text{ kJ/mol}\end{aligned}$$

$$\begin{aligned}\Delta G &= -RT \ln(K) \\ -190000 \text{ J/mol} &= 8.314\text{J/K}\cdot\text{mol} \cdot 298\text{K} \ln(K) \\ -190000/[-8.314(298)] &= \ln K \\ +76.7 &= \ln K \\ K &= e^{+76.7} \\ K &= 2 \times 10^{33}\end{aligned}$$