Name ___________________________
PHYS 0111 (Gen Chem 2)

Test 4
Spring 2005

A few equations:

$K = \dfrac{[\text{products}]}{[\text{reactants}]}$

$Q = \dfrac{[\text{products}]_o}{[\text{reactants}]_o}$

$pH = -\log[H_3O^+]$

$pOH = -\log[OH^-]$

$pK_w = -\log(K_w)$

$pK_a = -\log(K_a)$

$K_w = [H_3O^+][OH^-]$

$pK_w = pH + pOH$

$pH = pK_a + \log[A^-]/[HA]$

A few constants:

$K_w = 10^{-14}$

$pK_w = 14$

K_a values for a few acids

Acid	K_a	pK_a
HSO_4^-	1.2×10^{-2}	1.92
$HClO_2$	1.2×10^{-2}	1.92
H_3PO_4	7.5×10^{-3}	2.12
$CClH_2CO_2H$	1.35×10^{-3}	2.780
HF	7.2×10^{-4}	3.14
HNO_2	4.0×10^{-4}	3.40
CH_3CO_2H	1.8×10^{-5}	4.74
$[Al(H_2O)_6]^{3+}$	1.4×10^{-5}	4.85
$H_2PO_4^-$	6.2×10^{-8}	7.21
HOCl	3.5×10^{-8}	7.46
HCN	6.2×10^{-10}	9.21
NH_4^+	5.6×10^{-10}	9.25
HPO_4^{2-}	4.8×10^{-13}	12.32

1. _______

2. _______

3. _______

4. _______

5. _______

6. _______

7. _______

8. _______

1. (10 pts.) Determine the solubility of $BaCO_3$. For $BaCO_3$, $K_{sp} = 5.1 \times 10^{-9}$.

	$BaCO_3$ (s)	$\rightleftharpoons$	Ba^{2+}(aq)	+	CO_3^{2-}(aq)
init	solid		0		0
change	solid		+ x		+ x
equilib	solid		x		x

at equilibrium (a saturated solution) $K_{sp} = [Ba^{2+}][CO_3^{2-}]$

$5.1 \times 10^{-9} = x \cdot x$

$x = (5.1 \times 10^{-9})^{1/2}$

$x = 7.14 \times 10^{-5}$

Solubility of $BaCO_3$ (of the concentration of a saturated solution) is 7.1×10^{-5} M

2. a. (6 pts.) Using the information provided below, determine the K for the following reaction.

$$Cu(OH)_2(aq) + 2\,H_3O^+(aq) \rightleftharpoons Cu^{2+}(aq) + 4\,H_2O(l)$$

$2\,H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$ $\qquad K_w = 10^{-14}$

$Cu(OH)_2(s) \rightleftharpoons Cu^{2+}(aq) + 2\,OH^-(aq)$ $\qquad K_{sp} = 2.2 \times 10^{-22}$

$$\begin{array}{ll} & Cu(OH)_2(s) \rightleftharpoons Cu^{2+}(aq) + 2\,OH^-(aq) \\ +\,2 & \left[\, H_3O^+(aq) + OH^-(aq) \rightleftharpoons 2\,H_2O(l) \,\right] \\ \hline & Cu(OH)_2(aq) + 2\,H_3O^+(aq) \rightleftharpoons Cu^{2+}(aq) + 4\,H_2O(l) \end{array}$$

$K_{sp} = 2.2 \times 10^{-22}$

1/K_w because the rxn was turned around

$K = \dfrac{1}{(10^{-14})^2} = 10^{28}$

squared because the reaction was multiplied by two

$K = (2.2 \times 10^{-22})(10^{28})$

$K = 2.2 \times 10^{6}$

when rxns are added K's are multiplied

b. (6 pts.) Is $Cu(OH)_2$ considered soluble in water?

No, the K_{sp} is to small.

c. (4 pts.) Will $Cu(OH)_2$ dissolve in aqueous nitric acid? Explain.

Yes, since nitric acid is a strong acid, it will produce H^+ (H_3O^+ ions) that will react with the $Cu(OH)_2$, and the Cu^{2+} ions will go into solution as their OH^- counterions react with the protons H^+ (H_3O^+ ions).

3. a. (10 pts.) Determine the concentration of OH^- required to precipitate $Mg(OH)_2$ from a 0.105 M $Mg(NO_3)_2$ solution. For $Mg(OH)_2$ $K_{sp} = 8.9 \times 10^{-2}$. (This is the wrong K_{sp}. It should be 8.9×10^{-12}.)

With wrong K_{sp}

For a precipitate to form, Q ≥ K. So, when

$[Mg^{2+}][OH^-]^2 \geq 8.9 \times 10^{-2}$ a precipitate will form
$(0.105\ M)[OH^-]^2 \geq 8.9 \times 10^{-2}$
$[OH^-]^2 \geq (8.9 \times 10^{-2})/(0.105)$
$[OH^-] \geq [8.9 \times 10^{-2}/0.105]^{1/2}$
$[OH^-] \geq 0.92066$

When $[OH^-] \geq 0.92$ M $Mg(OH)_2$ will precipitate

With correct K_{sp}

For a precipitate to form, Q ≥ K. So, when

$[Mg^{2+}][OH^-]^2 \geq 8.9 \times 10^{-12}$ a precipitate will form
$(0.105\ M)[OH^-]^2 \geq 8.9 \times 10^{-12}$
$[OH^-]^2 \geq (8.9 \times 10^{-12})/(0.105)$
$[OH^-] \geq [8.9 \times 10^{-12}/0.105]^{1/2}$
$[OH^-] \geq 9.2066 \times 10^{-6}$

When $[OH^-] \geq 9.2 \times 10^{-6}$ M $Mg(OH)_2$ will precipitate

4. A solution was made by combining 20.0 mL of a 0.0340 M KOH solution with 100.0 mL of a solution that has an initial NH_4Cl concentration of 0.100 M and an initial NH_3 concentration of 0.110 M.

(10 pts.) Determine the pH of the resulting solution, and make certain to write any balanced chemical equations that are needed to determine the pH.

Strong base (SB) mixed with a weak acid (WA). SB reacts with WA in a reaction that essentially goes to completion.

	NH_4Cl (aq)	+	KOH (aq)	$\longrightarrow$	H_2O (l)	+	NH_3 (aq)	+ KCl (aq)
or	NH_4^+ (aq)	+	OH^- (aq)	$\longrightarrow$	H_2O (l)	+	NH_3 (aq)	
start	0.1 L x 0.1 M		0.02 L x 0.034 M				0.1 L x 0.11 M	
	0.01 mol		0.00068 mol				0.011 mol	
react	– 0.00068		– 0.00068				+ 0.00068	
end	0.00932 mol		0 mol				0.01168 mol	

After the neutralization reaction is complete, there is a weak acid and its conjugate base present in solution.

$\boxed{K_a \text{ version}}$	$NH_4^+(aq)$ +	$H_2O(l)$ $\rightleftharpoons$	$NH_3(aq)$ +	$H_3O^+(aq)$
init	$\frac{0.00932}{0.12\ L}$		$\frac{0.01168}{0.12\ L}$	0
change	– x		+ x	+ x
equilib	0.07767 – x		0.09733 + x	x

$$5.6 \times 10^{-10} = \frac{x\ (0.09733 + x)}{0.07767 - x}$$

small x approx

$5.6 \times 10^{-10} = x\ (0.09733)/0.07767$

$x = 4.4685 \times 10^{-10}$ $pH = -\log(4.4685 \times 10^{-10}) = 9.35$

buffer version

$pH = pK_a + \log [A-]/[HA]$

$pH = -\log(5.6 \times 10^{-10}) + \log(0.01168/0.00932)$

$pH = 9.35$

5. (10 pts.) A solution was prepared by dissolving 0.10 mol of HCl and 0.10 mol of NaCl in 250 mL of water. Is this solution a buffer? Explain.

No, this solution is not a buffer. This solution only contains a strong acid and a non-basic chloride ion. So, there is no weak base present to consume acid when acid is added, and there is no undissociated acid present to replenish protons that are consumed when base is added.

6. (10 pts.) Suggest two acid-base conjugate pairs that could be used to make a buffer that has a pH of approximately 7.3. Refer to the table on the cover page for a list of acids and K_a (pK_a) values.

pick two pairs that have a pK_a close to the desired pH.

an $H_2PO_4^-/HPO_4^{2-}$ buffer and an $HOCl/OCl^-$ buffer

7. The pH of a buffer (solution 1) that is 0.40 M in H_3PO_4 and 0.40 M in $H_2PO_4^-$ is the same as the pH of a buffer (solution 2) that is 0.10 M in H_3PO_4 and 0.10 M in $H_2PO_4^-$.

a. (6 pts.) Which of these solutions has a higher capacity for absorbing protons? Explain.

The 0.40 M $H_3PO_4/H_2PO_4^-$ buffer. Solution 1 has more base present than Solution 2. In other words, the Solution 2 will run out of base before Solution 1 does.

b. When 0.0010 mol of OH^- is added to 100 mL of each of the solutions described above, the pH of the solutions will change.

i. (4 pts.) Will the pH of the solutions decrease or increase? Explain.

The pH of the solution will increase slightly. When OH^- is added, H^+ will be consumed. Some, but not all of the H^+ will be replaced by the acid H_3PO_4. As $[H^+]$ decreases, pH increases. Another way of thinking of it is as follows: as $[H^+]$ decreases, the solution becomes more basic. The pH of a basic solution is higher than the pH of an acidic solution, so as a solution becomes more basic, its pH increases. Yet another way of thinking of it... a base, OH^-, is being added to the solution. Bases increase the pH of a solution.

ii. (4 pts.) The pH of which solution will change more?

Because solution two is more dilute, it has a lower capacity for absorbing acid or base. Thus, the pH of solution two would change more.

8. (2 pts each) At the end of a titration, the following chemicals remained in solution. Will the solutions be acidic, basic, or neutral?

a. $KClO_3$ **neutral** (K^+ no A/B prop. ClO_3^- is the conjugate base of a strong acid. Therefore, it is not a base.)

b. CH_3CO_2Na **basic** (Na^+ no A/B prop. $CH_3CO_2^-$ is the conjugate base of a weak acid. Therefore, it is a weak base.)

c. NH_4NO_3 **acidic** (NH_4^+ is a weak acid. NO_3^- is the conjugate base of a strong acid. Therefore, it is not a base.)

d. Na_2SO_4 **basic** (Na^+ no A/B prop. SO_4^{2-} is the conjugate base of the weak acid HSO_4^-. Therefore, SO_4^{2-} is a weak base.) Yes, I know that this one is very hard, but it is only worth 2 pts.