AP Chemistry Problem Set Chapter 15 **Name__________________**

30 points – 5 points for completion, 3 random essay problems will be graded, each worth 5 points. Each multiple choice must be answered (1 point each). Staple this sheet to the front of your essay responses.

Multiple Choice. Please indicate your multiple choice answers below.

1. ________ 2. ________ 3. ________ 4. ________ 5. ________

6. ________ 7. ________ 8. ________ 9. ________ 10. ________

Use these answers for questions 1-3. (1989)

Ionization Constants	
CH_3COOH	1.8×10^{-5}
NH_3	1.8×10^{-5}
H_2CO_3	$K_1 = 4 \times 10^{-7}$
H_2CO_3	$K_2 = 4 \times 10^{-11}$

a. a solution with a pH less than 7 that is not a buffer solution
b. a buffer solution with a pH between 4 and 7
c. a buffer solution with a pH between 7 and 10
d. a solution with a pH greater than 7 that is not a buffer solution
e. a solution with a pH of 7

1. A solution prepared to be initially 1 M in NaCl and 1 M in HCl.
2. A solution prepared to be initially 1 M in Na_2CO_3 and 1 M in CH_3COONa
3. A solution prepared to be initially 0.5 M in CH_3COOH and 1 M in CH_3COONa

Question 4-7 refer to aqueous solutions containing 1:1 mole ratios of the following pairs of substances. Assume all concentrations are 1 M. (1999)

a. NH_3 and NH_4Cl
b. H_3PO_4 and NaH_2PO_4
c. HCl and NaCl
d. NaOH and NH_3
e. NH_3 and $HC_2H_3O_2$ (acetic acid)

4. The solution with the lowest pH
5. The most nearly neutral solution
6. A buffer at a pH > 8
7. A buffer at a pH < 6

Acid	Acid Dissociation Constant, K_a
H_3PO_4	7×10^{-3}
$H_2PO_4^-$	8×10^{-8}
HPO_4^{2-}	5×10^{-13}

8. On the basis of the information above, a buffer with a pH = 9 can best be made by using: (1984)

a. pure NaH_2PO_4
b. $H_3PO_4 + H_2PO_4^-$
c. $H_2PO_4^- + PO_4^{3-}$
d. $H_2PO_4^- + HPO_4^{2-}$
e. $HPO_4^{2-} + PO_4^{3-}$

9. In the titration of a weak acid of unknown concentration with a standard solution of a strong base, a pH meter was used to follow the progress of the titration. Which of the following is true for this experiment? (1989)

a. The pH is 7 at the equivalence point.
b. The pH at the equivalence point depends on the indicator used.
c. The graph of pH versus volume of base added rises gradually at first and then much more rapidly.
d. The graph of pH versus volume of base added shows no sharp rise.
e. The $[H^+]$ at the equivalence point equals the ionization constant of the acid.

10. When phenolphthalein is used as the indicator in a titration of an HCl solution with a solution of NaOH, the indicator undergoes a color change from clear to red at the end point of the titration. This color change occurs abruptly because: (1989)

a. phenolphthalein is a very strong acid that is capable of rapid dissociation
b. the solution being titrated undergoes a large pH change near the end point of the titration
c. phenolphthalein undergoes an irreversible reaction in basic solution
d. OH^- acts as a catalyst for the decomposition of phenolphthalein
e. phenolphthalein is involved in the rate-determining step of the reaction between H_3O^+ and OH^-

Essays

1. Titrations

Consider the titration of 100.0 mL of 0.100 M H_2NNH_2 ($K_b = 3.0 \times 10^{-6}$) by 0.200 M HNO_3. Calculate the pH of the resulting solution after the following volumes of HNO_3 have been added.

a. 0.0 mL
b. 25.0 mL
c. 40.0 mL
d. 50.0 mL
e. 100.0 mL

2. Solubility Product

a. A saturated solution of silver arsenate (Ag_3AsO_4) contains 8.5×10^{-6} g Ag_3AsO_4 per mL. Calculate the K_{sp} of silver arsenate. Assume there are no other reactions but the K_{sp} reaction.

b. The solubility of silver chromate is (Ag_2CrO_4) in water is 2.7×10^{-3} g/100. mL. Calculate the K_{sp} of silver chromate. Assume there are no other reactions but the K_{sp} reaction.

c. The solubility product value for lead(II) iodide is 1.4×10^{-8} at 25°C. Calculate the solubility of PbI_2.

3. Buffers

Calculate the pH change when 1.0 mL of 1.0 M HCl is added to 0.100 L of a solution of:

a. 0.10 M acetic acid and 0.10 M sodium acetate
b. 0.010 M acetic acid and 0.010 M sodium acetate
c. 0.0010 M acetic acid and 0.0010 M sodium acetate

A buffer consists of 0.20 M propanoic acid ($Ka = 1.4 \times 10^{-5}$) and 0.30 M sodium propanoate.

d. Calculate the pH of this buffer.
e. Calculate the pH after the addition of 3.0 mL of the 1.0 M HCl to 0.010 L of the buffer.

4. Solubility Equilibria

a. A solution is prepared by mixing 75.0 mL of 0.020 M $BaCl_2$ and 125 mL of 0.040 M Na_2SO_4. Calculate the concentrations of Ba^{2+} and SO_4^{2-} at equilibrium ($K_{sp} = 1.5 \times 10^{-9}$).

b. A solution consists of 2.5×10^{-4} M Na_3PO_4. What is the minimum concentration of $AgNO_3$ that would cause precipitation of solid Ag_3PO_4 ($K_{sp} = 1.8 \times 10^{-18}$)?

5.

$$NH_3(aq) + H_2O(l) \Longleftrightarrow NH_4^+(aq) + OH^-(aq)$$

In aqueous solution, ammonia reacts as represented above. In 0.0180 *M* $NH_3(aq)$ at 25°C, the hydroxide ion concentration, $[OH^-]$, is 5.60×10^{-4} *M*. In answering the following, assume that temperature is constant at 25°C and that volumes are additive.

a. Write the equilibrium-constant expression for the reaction represented above.
b. Determine the pH of 0.0180 *M* $NH_3(aq)$.
c. Determine the value of the base ionization constant, K_b, for $NH_3(aq)$.
d. Determine the percent ionization of NH_3 in 0.0180 *M* $NH_3(aq)$.
e. In an experiment, a 20.0 mL sample of 0.0180 *M* $NH_3(aq)$ was placed in a flask and titrated to the equivalence point and beyond using 0.0120 *M* $HCl(aq)$.
    i. Determine the volume of 0.0120 *M* $HCl(aq)$ that was added to reach the equivalence point.
    ii. Determine the pH of the solution in the flask after a total of 15.0 mL of 0.0120 *M* $HCl(aq)$ was added.
    iii. Determine the pH of the solution in the flask after a total of 40.0 mL of 0.0120 *M* $HCl(aq)$ was added.

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Multiple Choice. Please indicate your multiple choice answers below.

1. 40% A	2. 26% D	3. 36% B	4. 52% C	5. 41% E
6. 33% A	7. 35% B	8. 32% D	9. 50% C	10. 66% B

Use these answers for questions 1-3.

Ionization Constants	
CH_3COOH	1.8×10^{-5}
NH_3	1.8×10^{-5}
H_2CO_3	$K_1 = 4 \times 10^{-7}$
H_2CO_3	$K_2 = 4 \times 10^{-11}$

a. a solution with a pH less than 7 that is not a buffer solution
b. a buffer solution with a pH between 4 and 7
c. a buffer solution with a pH between 7 and 10
d. a solution with a pH greater than 7 that is not a buffer solution
e. a solution with a pH of 7

1. A solution prepared to be initially 1 M in NaCl and 1 M in HCl. **1989 40% A**
2. A solution prepared to be initially 1 M in Na_2CO_3 and 1 M in CH_3COONa **1989 26% D**
3. A solution prepared to be initially 0.5 M in CH_3COOH and 1 M in CH_3COONa **1989 36% B**

Question 4-7 refer to aqueous solutions containing 1:1 mole ratios of the following pairs of substances. Assume all concentrations are 1 M.

a. NH_3 and NH_4Cl
b. H_3PO_4 and NaH_2PO_4
c. HCl and NaCl
d. NaOH and NH_3
e. NH_3 and $HC_2H_3O_2$ (acetic acid)

4. The solution with the lowest pH **1999 52% C**
5. The most nearly neutral solution **1999 41% E**
6. A buffer at a pH > 8 **1999 33% A**
7. A buffer at a pH < 6 **1999 35% B**

Acid	Acid Dissociation Constant, K_a
H_3PO_4	7×10^{-3}
$H_2PO_4^-$	8×10^{-8}
HPO_4^{2-}	5×10^{-13}

8. On the basis of the information above, a buffer with a pH = 9 can best be made by using **1984 32% D**

a. pure NaH_2PO_4
b. $H_3PO_4 + H_2PO_4^-$
c. $H_2PO_4^- + PO_4^{3-}$
d. $H_2PO_4^- + HPO_4^{2-}$
e. $HPO_4^{2-} + PO_4^{3-}$

9. In the titration of a weak acid of unknown concentration with a standard solution of a strong base, a pH meter was used to follow the progress of the titration. Which of the following is true for this experiment?

a. The pH is 7 at the equivalence point.
b. The pH at the equivalence point depends on the indicator used.
c. The graph of pH versus volume of base added rises gradually at first and then much more rapidly. 1989 50%
d. The graph of pH versus volume of base added shows no sharp rise.
e. The $[H^+]$ at the equivalence point equals the ionization constant of the acid.

10. When phenolphthalein is used as the indicator in a titration of an HCl solution with a solution of NaOH, the indicator undergoes a color change from clear to red at the end point of the titration. This color change occurs abruptly because:

a. phenolphthalein is a very strong acid that is capable of rapid dissociation
b. the solution being titrated undergoes a large pH change near the end point of the titration 1989 66%
c. phenolphthalein undergoes an irreversible reaction in basic solution
d. OH^- acts as a catalyst for the decomposition of phenolphthalein
e. phenolphthalein is involved in the rate-determining step of the reaction between H_3O^+ and OH^-

Essays

1. Titrations

Consider the titration of 100.0 mL of 0.100 M H_2NNH_2 ($K_b = 3.0 \times 10^{-6}$) by 0.200 M HNO_3. Calculate the pH of the resulting solution after the following volumes of HNO_3 have been added.

a. 0.0 mL **pH = 10.74**
b. 25.0 mL **pH = 8.5**
c. 40.0 mL **pH = 7.9**
d. 50.0 mL **pH = 4.8**
e. 100.0 mL **pH = 1.3**

2. Solubility Product

a. A saturated solution of silver arsenate (Ag_3AsO_4) contains 8.5×10^{-6} g Ag_3AsO_4 per mL. Calculate the K_{sp} of silver arsenate. Assume there are no other reactions but the K_{sp} reaction. **$K_{sp} = 3.09 \times 10^{-18}$**

b. The solubility of silver chromate is (Ag_2CrO_4) in water is 2.7×10^{-3} g/100. mL. Calculate the K_{sp} of silver chromate. Assume there are no other reactions but the K_{sp} reaction. **$K_{sp} = 2.16 \times 10^{-12}$**

c. The solubility product value for lead(II) iodide is 1.4×10^{-8} at 25°C. Calculate the solubility of PbI_2. **Solubility = 1.52×10^{-3}**

3. Buffers

Calculate the pH change when 1.0 mL of 1.0 M HCl is added to 0.100 L of a solution of:

a. 0.10 M acetic acid and 0.10 M sodium acetate **ΔpH = -0.087**
b. 0.010 M acetic acid and 0.010 M sodium acetate **ΔpH = -1.52**
c. 0.0010 M acetic acid and 0.0010 M sodium acetate **ΔpH = -2.69**

A buffer consists of 0.20 M propanoic acid ($K_a = 1.4 \times 10^{-5}$) and 0.30 M sodium propanoate.

d. Calculate the pH of this buffer. **pH = 5.03**
e. Calculate the pH after the addition of 3.0 mL of the 1.0 M HCl to 0.010 L of the buffer. **pH = 2.63**

4. Solubility Equilibria

a. A solution is prepared by mixing 75.0 mL of 0.020 M $BaCl_2$ and 125 mL of 0.040 M Na_2SO_4. Calculate the concentrations of Ba^{2+} and SO_4^{2-} at equilibrium ($K_{sp} = 1.5 \times 10^{-9}$).
$[Ba^{2+}] = 8.6 \times 10^{-8}$ $[SO_4^{2-}] = 1.75 \times 10^{-2}$

b. A solution consists of 2.5×10^{-4} M Na_3PO_4. What is the minimum concentration of $AgNO_3$ that would cause precipitation of solid Ag_3PO_4 ($K_{sp} = 1.8 \times 10^{-18}$)? **$[AgNO_3] = 1.93 \times 10^{-5}$ M**

5. AP Question - 1999

$$NH_3(aq) + H_2O(l) \Longleftrightarrow NH_4^+(aq) + OH^-(aq)$$

In aqueous solution, ammonia reacts as represented above. In 0.0180 *M* $NH_3(aq)$ at 25°C, the hydroxide ion concentration, $[OH^-]$, is 5.60×10^{-4} *M*. In answering the following, assume that temperature is constant at 25°C and that volumes are additive.

f. Write the equilibrium-constant expression for the reaction represented above. $\boldsymbol{K = \dfrac{[NH_4^+][OH^-]}{[NH_3]}}$

g. Determine the pH of 0.0180 *M* $NH_3(aq)$. **pH = 10.7**

h. Determine the value of the base ionization constant, K_b, for $NH_3(aq)$. **$K_b = 1.74 \times 10^{-5}$**

i. Determine the percent ionization of NH_3 in 0.0180 *M* $NH_3(aq)$. **3.11%**

j. In an experiment, a 20.0 mL sample of 0.0180 *M* $NH_3(aq)$ was placed in a flask and titrated to the equivalence point and beyond using 0.0120 *M* HCl(*aq*).

   i. Determine the volume of 0.0120 *M* HCl(*aq*) that was added to reach the equivalence point. **30.0 mL**

   ii. Determine the pH of the solution in the flask after a total of 15.0 mL of 0.0120 *M* HCl(*aq*) was added. **pH = 9.3**

   iii. Determine the pH of the solution in the flask after a total of 40.0 mL of 0.0120 *M* HCl(*aq*) was added. **pH = 2.7**