

Practice MCQS: Acid Base Equilibria

- 1) When 200. mL of 2.0M NaOH(aq) is added to 500. mL of 1.0M HCl (aq), the pH of the resulting mixture is closest to

(A) 1.0 (B) 3.0 (C) 7.0 (D) 13.0

0.4 mol NaOH

0.5 mol HCl

$$\frac{0.1 \text{ mol H}^+}{0.7 \text{ L}} = 0.14 \text{ M} \approx 1.0 \text{ pH}$$

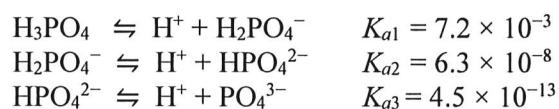
- 2) An acetate buffer solution is prepared by combining 50. mL of 0.20 M acetic acid, HC₂H₃O₂(aq), and 50. mL of 0.20 M sodium acetate, NaC₂H₃O₂(aq). A 5.0 mL sample of 0.10 M NaOH(aq) is added to the buffer solution. Which of the following is a correct pairing of the acetate species present in greater concentration and of the pH of the solution after the NaOH(aq) is added? (The pK_a of acetic acid is 4.7.)

Base will reduce [HC₂H₃O₂] and slightly raise pH

| Acetate Species | pH |
|--------------------------------------------------------------|-------|
| A) HC ₂ H ₃ O ₂ | < 4.7 |
| B) HC ₂ H ₃ O ₂ | > 4.7 |
| C) C ₂ H ₃ O ₂ ⁻ | < 4.7 |
| D) C ₂ H ₃ O ₂ ⁻ | > 4.7 |

- 3) A solution is prepared by mixing 50 mL of 1M NaH₂PO₄ with 50 mL of 1M Na₂HPO₄. On the basis of the information which of the following species is present in the solution at the lowest concentration?

A) Na⁺ B) HPO₄²⁻ C) H₂PO₄⁻ (D) PO₄³⁻



very small

- 4) Acid-dissociation constants of two acids are listed in the table. A 20. mL sample of a 0.10 M solution of each acid is titrated to the equivalence point with 20. mL of 0.10 M NaOH. Which of the following is a true statement about the pH of the solutions at the equivalence point?

| Solution | Acid | K _a |
|----------|-----------------------------------|-------------------------|
| 1 | CH ₃ CO ₂ H | 1.75 × 10 ⁻⁵ |
| 2 | CF ₃ CO ₂ H | 1.0 × 10 ⁰ |

- A) Solution 1 has a higher pH at the equivalence point because CH₃CO₂H is the stronger acid. *is not*
 B) Solution 1 has a higher pH at the equivalence point because CH₃CO₂H has the stronger conjugate base.
 C) Solution 1 has a lower pH at the equivalence point because CH₃CO₂H is the stronger acid.
 D) Solution 1 has a lower pH at the equivalence point because CH₃CO₂H has the stronger conjugate base.

- 5) Which of the following accounts for the observation that the pH of pure water at 37°C is 6.8 ?

- A) At 37°C water is naturally acidic. *like magic?*
 B) At 37°C the autoionization constant for water, K_w, is larger than it is at 25°C.
 C) At 37°C water has a lower density than it does at 25°C; therefore, [H⁺] is greater. *what? no*
 D) At 37°C water ionizes to a lesser extent than it does at 25°C.

- 6) The dissociation of the weak acid HF in water is represented by the equation: HF(aq) + H₂O(l) ⇌ H₃O⁺(aq) + F⁻(aq)
 Adding a 1.0 mL sample of which of the following would increase the percent ionization of HF(aq) in 10 mL of a solution of 1.0 M HF ?

- A) 1.0 M KF B) 1.0 M H₂SO₄ C) 10.0 M HF (D) Distilled water

A B Shift left Shift left nothing Shift + right

- 7) HX(aq) + Y⁻(aq) ⇌ HY(aq) + X⁻(aq) *K_{eq} > 1* A solution of a salt of a weak acid HY is added to a solution of another weak acid HX. Based on the information given above, which of the following species is the strongest base?

- A) HX(aq) (B) Y⁻(aq) C) HY(aq) D) X⁻(aq) *right side favored*

- 8) H₂C₂O₄ + 2 H₂O ⇌ 2 H₃O⁺ + C₂O₄²⁻ Oxalic acid, H₂C₂O₄, is a diprotic acid with K₁ = 5 × 10⁻² and K₂ = 5 × 10⁻⁵. Which of the following is equal to the equilibrium constant for the reaction represented above?

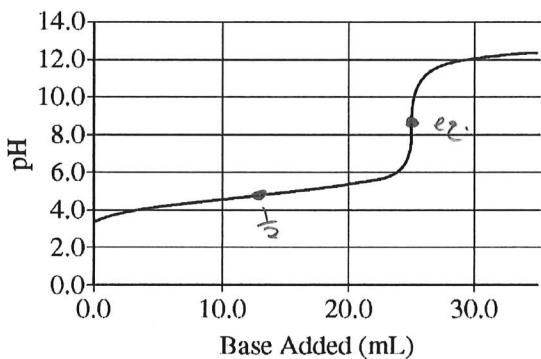
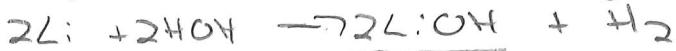
- A) 1000 B) 1 × 10⁻³ (C) 2.5 × 10⁻⁶ D) 2.5 × 10⁻⁷

$$\begin{aligned} K_1 \cdot K_2 &= K_3 \\ (5 \times 10^{-2}) \cdot (5 \times 10^{-5}) &= 2.5 \times 10^{-6} \end{aligned}$$

Questions 9-12 refer to the following information.

A 0.35 g sample of Li(s) is placed in an Erlenmeyer flask containing 100 mL of water at 25°C. A balloon is placed over the mouth of the flask to collect the hydrogen gas that is generated. After all of the Li(s) has reacted with H₂O(l), the solution in the flask is added to a clean, dry buret and used to titrate an aqueous solution of a monoprotic acid.

The pH curve for this titration is shown in the diagram.



- 9) Which of the following changes will most likely increase the rate of reaction between Li(s) and water?

- A) Using 125 mL of water instead of 100 mL
- B) Using a 0.25 g sample of Li(s) instead of a 0.35 g sample
- C) Using a 0.35 g sample of Li(s) cut into small pieces *more surface area*
- D) Decreasing the water temperature before adding the Li(s)

- 10) What will be the effect on the amount of gas produced if the experiment is repeated using 0.35 g of K(s) instead of 0.35 g of Li(s) ?

- A) No gas will be produced when K(s) is used.
- B) Some gas will be produced but less than the amount of gas produced with Li(s).
- C) Equal quantities of gas will be produced with the two metals.
- D) More gas will be produced with K(s) than with Li(s).

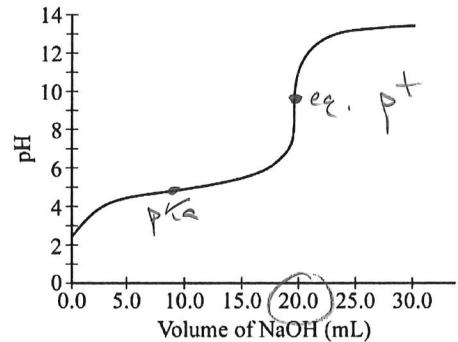
- 11) On the basis of the pH curve, the pK_a value of the acid is closest to A) 4 B) 5 C) 8 D) 12

- 12) Which of the following is the balanced net-ionic equation for the reaction between Li(s) and water?

- A) 2 Li(s) + 2 H⁺(aq) + 2 OH⁻(aq) → 2 Li⁺(aq) + 2 OH⁻(aq) + H₂(g)
- B) 2 Li(s) + 2 H₂O(l) → 2 Li⁺(aq) + 2 OH⁻(aq) + H₂(g)
- C) 2 Li(s) + 2 H₂O(l) → 2 LiOH(s) + H₂(g)
- D) 2 Li(s) + 2 H₂O(l) → 2 LiH(s) + H₂(g)

- 13) Data collected during the titration of a 20.0 mL sample of a 0.10 M solution of a monoprotic acid with a solution of NaOH of unknown concentration are plotted in the graph. Based on the data, which of the following are the approximate pK_a of the acid and the molar concentration of the NaOH?

| | pK _a | [NaOH] |
|------------------------------------|-----------------|---------|
| A | 4.7 | 0.050 M |
| <input checked="" type="radio"/> B | 4.7 | 0.10 M |
| <input checked="" type="radio"/> C | 9.3 | 0.050 M |
| <input checked="" type="radio"/> D | 9.3 | 0.10 M |



- 14) In a saturated solution of Zn(OH)₂ at 25°C , the value of [OH⁻] is 2.0×10^{-6} M.

What is the value of the solubility-product constant, K_{sp}, for Zn(OH)₂ at 25°C ?

- A) 4.0×10^{-18}
- B) 8.0×10^{-18}
- C) 1.6×10^{-17}
- D) 4.0×10^{-12}

$$K_{sp} = [\text{Zn}^{2+}][\text{OH}^-]^2 = [1.0 \times 10^{-6}]^2 [2.0 \times 10^{-6}]^2 = 4.0 \times 10^{-18}$$

- 15) Based on the K_{sp} values in the table, a saturated solution of which of the following compounds has the highest [Cl⁻] ?

- A) PbCl₂
- B) CuCl
- C) AgCl
- D) Hg₂Cl₂

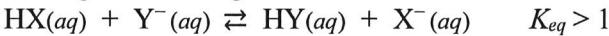
largest K_{sp}

| Compound | K _{sp} |
|---------------------------------|-----------------------|
| PbCl ₂ | 1.2×10^{-5} |
| CuCl | 1.6×10^{-7} |
| AgCl | 1.8×10^{-10} |
| Hg ₂ Cl ₂ | 1.4×10^{-18} |

most soluble

- 16) Based on the information given below, which of the following is the strongest acid?

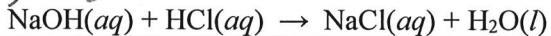
- A) HX(aq)
- B) Y⁻(aq)
- C) HY(aq)
- D) X⁻(aq)



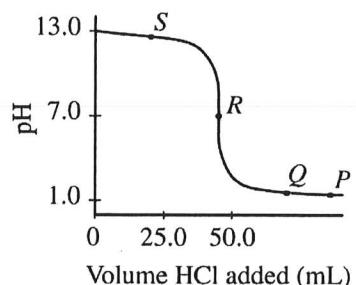
right side favored

Questions 17-18 refer to the following information.

strong strong



To determine the concentration of a $\text{NaOH}(aq)$ solution, a student titrated a 50. mL sample with 0.10 M $\text{HCl}(aq)$. The reaction is represented by the equation above. The titration is monitored using a pH meter, and the experimental results are plotted in the graph below.



- 17) At the point labeled *equivalence P* on the pH curve, which of the following ions are present in the reaction mixture at a concentration greater than 0.01 M?

A) Na^+ and Cl^- only B) Na^+ , Cl^- , and H^+ only C) Na^+ , Cl^- , and OH^- only D) Na^+ , Cl^- , H^+ , and OH^-

- H⁺ and OH⁻ have been neutralized*
- 18) One student titrated the $\text{NaOH}(aq)$ with 1.0 M $\text{HCl}(aq)$ instead of 0.10 M $\text{HCl}(aq)$. How would the student's titration curve differ from the original curve?

A) The initial pH would be 11 instead of 13. *Base didn't change*

B) The pH at the equivalence point would be 5 instead of 7. *no*

C) The pH far beyond the equivalence point would be higher than in the original curve.

D) The pH far beyond the equivalence point would be lower than in the original curve. *1 M HCl = 0 pH*

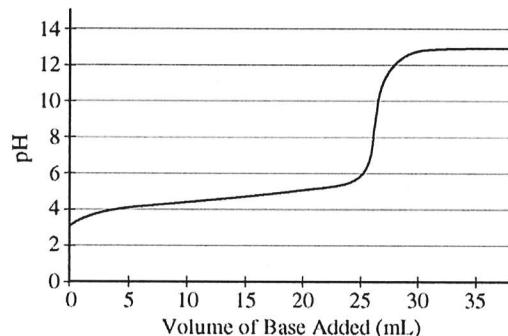
- 19) A student performs an acid-base titration and plots the experimental results in the graph. Which of the following statements best explains the experimental findings?

A) A strong acid was titrated with a strong base, as evidenced by the equivalence point at pH = 7. *not*

B) A strong acid was titrated with a strong base, as evidenced by the equivalence point at pH > 7. *would be 7*

C) A weak acid was titrated with a strong base, as evidenced by the equivalence point at pH > 7.

D) A weak acid was titrated with a weak base, as evidenced by the equivalence point at pH approximately 7. *not*

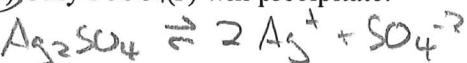


- 20) A 1.0 L solution of $\text{AgNO}_3(aq)$ and $\text{Pb}(\text{NO}_3)_2(aq)$ has a Ag^+ concentration of 0.020 M and a Pb^{2+} concentration of 0.0010 M. A 0.0010 mol sample of $\text{K}_2\text{SO}_4(s)$ is added to the solution. Based on the information in the table above, which of the following will occur? (Assume that the volume change of the solution is negligible.)

A) No precipitate will form. B) Only $\text{Ag}_2\text{SO}_4(s)$ will precipitate.

C) Only $\text{PbSO}_4(s)$ will precipitate. D) Both $\text{Ag}_2\text{SO}_4(s)$ and $\text{PbSO}_4(s)$ will precipitate.

| Compound | K_{sp} at 298 K |
|--------------------------|--------------------|
| Ag_2SO_4 | 1×10^{-5} |
| PbSO_4 | 1×10^{-8} |



$$Q = [\text{Ag}^+]^2 [\text{SO}_4^{2-}]$$

$$= [0.020]^2 [0.010]$$

$$= 4 \times 10^{-7} < 1 \times 10^{-5} \text{ no precip } \begin{matrix} (\text{shift left}) \\ (\text{right}) \end{matrix}$$



$$Q = [\text{Pb}^{2+}] [\text{SO}_4^{2-}]$$

$$= [0.0010]^2$$

$$= 1 \times 10^{-6} > 1 \times 10^{-8}$$

will precip
(shift left left)

- 21) The pH of a 0.01 M $\text{HNO}_2(aq)$ solution is in which of the following ranges? (For $\text{HNO}_2(aq)$, $K_a = 4.0 \times 10^{-4}$)

A) Between 1 and 2

B) Between 2 and 3

C) Between 4 and 5

D) Between 6 and 7



$$K_a = \frac{[\text{H}^+][\text{NO}_2^-]}{[\text{HNO}_2]}$$

$$4.0 \times 10^{-4} = \frac{x^2}{0.01}$$

$$x = 0.002 = [\text{H}^+]$$

$$\text{pH} = -\log(0.002) = 2.7$$

- 22) Which of the following correctly ranks the three monoprotic acids listed in the table from the weakest to the strongest?

- A) X < Y < Z
B) X < Z < Y
C) Y < Z < X
D) Z < Y < X

$$\begin{aligned} [\text{H}^+] &= 10^{-2.3} \\ &= 0.005 \\ &\sim 100\% \text{ ionized} \end{aligned}$$

| Acid | Concentration | pH |
|------|---------------|-----|
| X | 0.005 M | 2.3 |
| Y | 2.0 M | 2.8 |
| Z | 3.0 M | 2.8 |

$$[\text{H}^+] \rightarrow 0.005$$

Y has a higher % ionization

- 23) Three saturated solutions (X, Y, and Z) are prepared at 25°C. Based on the information in the table above, which of the following lists the solutions in order of increasing $[\text{Ag}^+]$?

- A) X < Z < Y
B) Y < X < Z
C) Z < Y < X
D) Z < X < Y

| Solution | Solute | K_{sp} at 25°C |
|----------|--------|-----------------------|
| X | AgBr | 5.0×10^{-13} |
| Y | AgCl | 1.8×10^{-10} |
| Z | AgI | 8.3×10^{-17} |

lowest to highest K_{sp}

- 24) What are the relative strengths of the acids and bases in the reaction represented by the equation below?

Acid Strength

- A) $\text{HClO}_2 < \text{HCOOH}$
B) $\text{HClO}_2 < \text{HCOOH}$
C) $\text{HClO}_2 > \text{HCOOH}$
D) $\text{HClO}_2 > \text{HCOOH}$

Base Strength

- $\text{ClO}_2^- < \text{HCOO}^-$
 $\text{ClO}_2^- > \text{HCOO}^-$
 $\text{ClO}_2^- > \text{HCOO}^-$
 $\text{ClO}_2^- < \text{HCOO}^-$



stronger acid $K_{eq} < 1$

left favored

stronger base

- 25) A student prepares a solution by combining 100 mL of 0.30 M $\text{HNO}_2(aq)$ and 100 mL of 0.30 M $\text{KNO}_2(aq)$. Which of the following equations represents the reaction that best helps to explain why adding a few drops of 1.0 M $\text{HCl}(aq)$ does not significantly change the pH of the solution?

- A) $\text{K}^+(aq) + \text{Cl}^-(aq) \rightarrow \text{KCl}(s)$
B) $\text{HNO}_2(aq) \rightarrow \text{H}^+(aq) + \text{NO}_2^-(aq)$
C) $\text{H}^+(aq) + \text{OH}^-(aq) \rightarrow \text{H}_2\text{O}(l)$
D) $\text{H}^+(aq) + \text{NO}_2^-(aq) \rightarrow \text{HNO}_2(aq)$

Base portion of buffer reacts with H^+

Practice FRQs Acid/Base Equil.

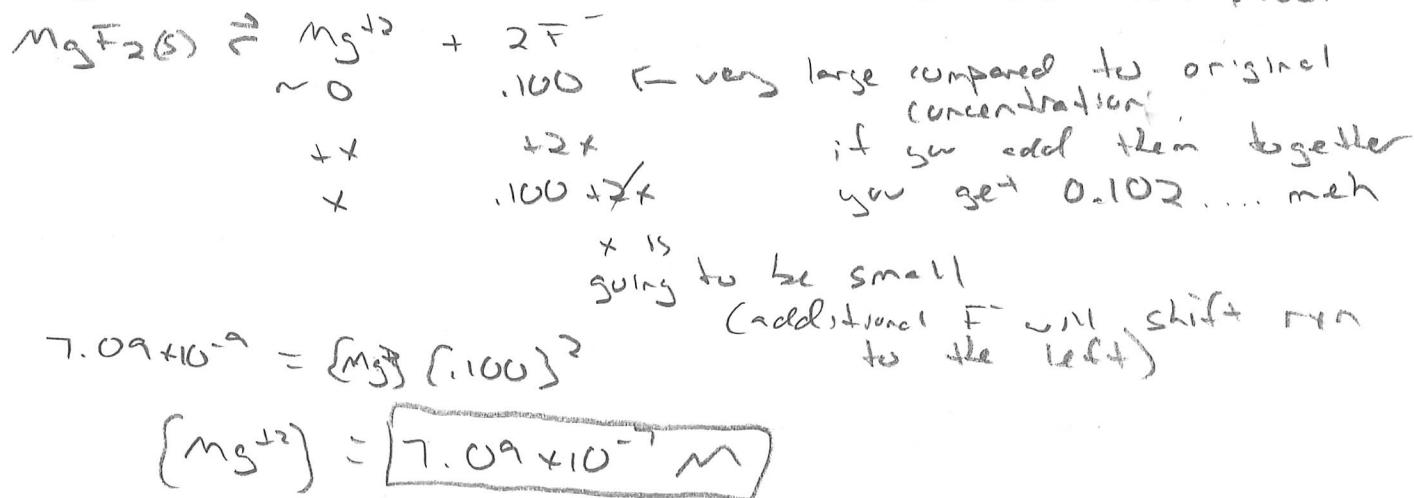
- 1) In a saturated solution of MgF_2 at $18^\circ C$, the concentration of Mg^{2+} is $1.21 \times 10^{-3} M$. The equilibrium is represented by the equation: $MgF_2(s) \rightleftharpoons Mg^{2+}(aq) + 2 F^-(aq)$

1994 #1
ratio

- a) Write the expression for the solubility-product constant, K_{sp} , and calculate its value at $18^\circ C$. (2pts)

$$K_{sp} = [Mg^{2+}][F^-]^2 = [1.21 \times 10^{-3}][2.42 \times 10^{-3}]^2 = 7.09 \times 10^{-9}$$

- b) Calculate the equilibrium concentration of Mg^{2+} in 1.000 L of saturated MgF_2 solution at $18^\circ C$ to which 0.100 mole of solid KF has been added. KF dissolves completely. Assume the volume change is negligible. (2pts)



- c) Predict whether a precipitate of MgF_2 will form when 100.0 mL of a $3.00 \times 10^{-3} M$ $Mg(NO_3)_2$ solution is mixed with 200.0 mL of $2.00 \times 10^{-3} M$ NaF solution at $18^\circ C$. Calculations to support your prediction must be shown.

$$0.1000 \text{ L } Mg(NO_3)_2 \left| \frac{3.00 \times 10^{-3} \text{ mol}}{1 \text{ L}} \right. = 3.00 \times 10^{-4} \text{ mol} / .300 \text{ L} \quad (3 \text{ pts})$$

$$= 1.00 \times 10^{-3} M$$

$$0.2000 \text{ L } NaF \left| \frac{2.00 \times 10^{-3} \text{ mol}}{1 \text{ L}} \right. = 4.00 \times 10^{-4} \text{ mol} / .300 \text{ L} \quad = 1.33 \times 10^{-3} M$$

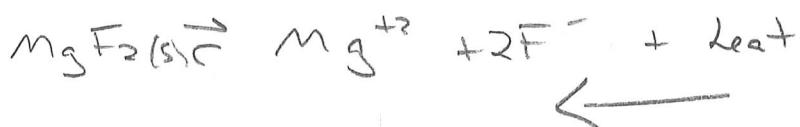
$$Q = [1.00 \times 10^{-3}][1.33 \times 10^{-3}]^2 = 1.77 \times 10^{-9} < K_{sp}$$

reaction will shift right (soluble side)
no precipitate

- d) At $27^\circ C$ the concentration of Mg^{2+} in a saturated solution of MgF_2 is $1.17 \times 10^{-3} M$. Is the dissolving of MgF_2 in water endothermic or exothermic? Give an explanation to support your conclusion. (2pts)

$$K_{sp} = [1.17 \times 10^{-3}][2.34 \times 10^{-3}]^2 = 6.41 \times 10^{-9}$$

- $K_{sp} @ 27^\circ C$ is less than $K_{sp} @ 18^\circ C$
- More soluble @ $18^\circ C$
- Reaction is exothermic because adding heat causes a shift to the left (solid side)

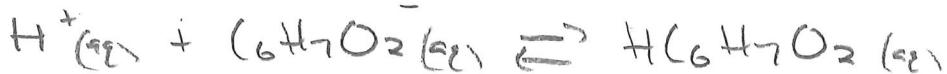




2) Potassium sorbate, $KC_6H_7O_2$ (molar mass 150. g/mol) is commonly added to diet soft drinks as a preservative. A stock solution of $KC_6H_7O_2(aq)$ of known concentration must be prepared. A student titrates 45.00 mL of the stock solution with 1.25 M $HCl(aq)$ using both an indicator and a pH meter. The value of K_a for sorbic acid, $HC_6H_7O_2$, is 1.7×10^{-5} .

a) Write the net-ionic equation for the reaction between $KC_6H_7O_2(aq)$ and $HCl(aq)$. (1pt)

2015 #3



b) A total of 29.95 mL of 1.25 M $HCl(aq)$ is required to reach the equivalence point. (2pts)

Calculate $[KC_6H_7O_2]$ in the stock solution.

$$\frac{0.02995 \text{ L } HCl}{1 \text{ L } HCl} \left| \begin{array}{c} 1.25 \text{ mol } HCl \\ \hline 1 \text{ mol } HCl \end{array} \right| \left| \begin{array}{c} 1 \text{ mol } KC_6H_7O_2 \\ \hline 1 \text{ mol } HCl \end{array} \right| = 0.832 \text{ M } KC_6H_7O_2$$

$\cancel{\Delta} \text{ or } M_a V_a = M_b V_b \text{ (1:1 ratio)}$

c) The pH at the equivalence point of the titration is measured to be 2.54. Which of the following indicators would be the best choice for determining the end point of the titration? Justify your answer. (2pts)

Thymol Blue - pK_a closest to equivalence point

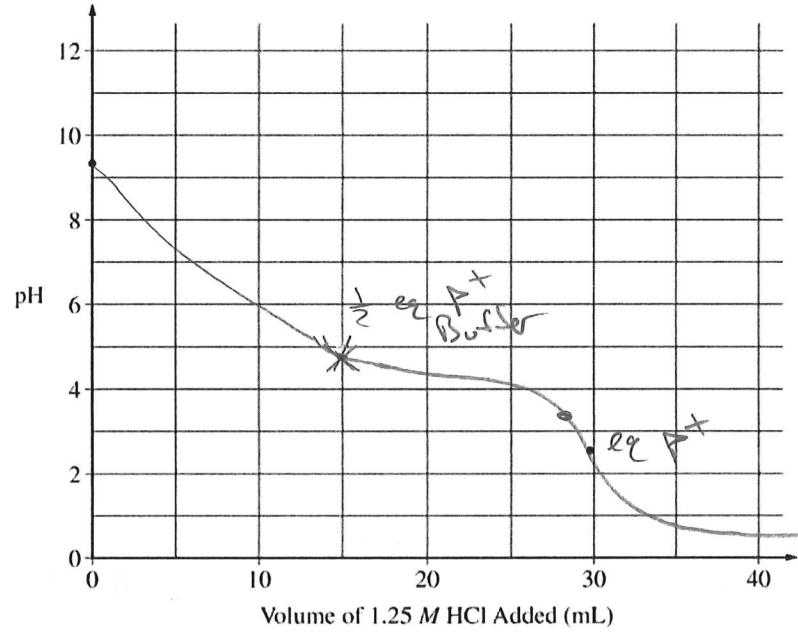
d) Calculate the pH at the half-equivalence point. (1pt)

pK_a

$$pH = pK_a = -\log(1.7 \times 10^{-5}) = 4.77$$

| Indicator | pK_a |
|------------------|--------|
| Phenolphthalein | 9.3 |
| Bromothymol blue | 7.0 |
| Methyl red | 5.0 |
| Thymol blue | 2.0 |
| Methyl violet | 0.80 |

e) The initial pH and the equivalence point are plotted on the graph. Accurately sketch the titration curve on the graph below. Mark the position of the half-equivalence point on the curve with an X. (3pts)



f) The pH of the soft drink is 3.37 after the addition of the $KC_6H_7O_2(aq)$. Which species, $HC_6H_7O_2$ or $C_6H_7O_2^-$, has a higher concentration in the soft drink? Justify your answer. (1pt)

3.37 is lower than ideal buffer range ($pH = 4.77$)

so $HC_6H_7O_2$ will have a higher concentration.

OR

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

$$3.37 = 4.77 + \log \frac{[A^-]}{[HA]}$$

$$\log \frac{[A^-]}{[HA]} = -1.40 \quad \frac{[A^-]}{[HA]} = 0.398 \quad \text{less greater}$$

A student is instructed to prepare 100.0 mL of 1.250 M NaOH from a stock solution of 5.000 M NaOH. 2011B 5
The student follows the proper safety guidelines.

- a) Calculate the volume of 5.000 M NaOH needed to accurately prepare 100.0 mL of 1.250 M NaOH solution.

$$M_1 V_1 = M_2 V_2$$

$$(5.000)(V_1) = (1.250)(100.0)$$

$$V_1 = 25.00 \text{ mL}$$

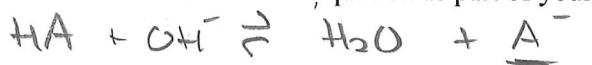
- b) Describe the steps in a procedure to prepare 100.0 mL of 1.250 M NaOH solution using 5.000 M NaOH and equipment selected from the list below.

| | | | |
|----------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------|------------------------|-------------------------------------------|-------------------------|
| Balance | 25 mL Erlenmeyer flask | 100 mL graduated cylinder | 100 mL volumetric flask |
| 50 mL buret | 100 mL Florence flask | 25 mL pipet | 100 mL beaker |
| Eyedropper | Drying oven | Wash bottle of distilled H ₂ O | Crucible |
| <ul style="list-style-type: none"> Add 25 mL of 5.000 M NaOH to a 100 mL vol flask using the 50 mL buret. (Or measure 25 mL in the grad cylinder & rinse w/ water) Use distilled water to fill the flask to the 100 mL line. | | | |

- c) The student is given 50.0 mL of a 1.00 M solution of a weak, monoprotic acid, HA. The solution is titrated with the 1.250 M NaOH to the endpoint. (Assume that the endpoint is at the equivalence point.)

- i) Explain why the solution is basic at the equivalence point of the titration.

Include a chemical equation as part of your explanation.



The A⁻ formed during the titration will react w/
H₂O to produce more OH⁻ ions.
This will make the solution basic

- ii) Identify the indicator in the table below that would be best for the titration. Justify your choice.

Phenolphthalein - basic pH @ equivalence pt.

| Indicator | pK _a |
|------------------|-----------------|
| Methyl red | 5 |
| Bromothymol blue | 7 |
| Phenolphthalein | 9 |

- d) The student is given another 50.0 mL sample of 1.00 M HA, which the student adds to the solution that had been titrated to the endpoint in part (c). The result is a solution with a pH of 5.0.

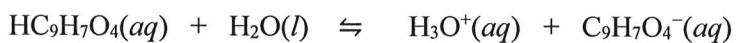
- i) What is the value of the acid-dissociation constant, K_a, for the weak acid? Explain your reasoning.

50 mL more acid will push the reaction to the half equivalence point. So [HA] = [A⁻]

$$\text{pH} = \text{pK}_a \rightarrow \text{K}_a = 10^{-5.0} = 1 \times 10^{-5}$$

- ii) Explain why the addition of a few drops of 1.250 M NaOH to the resulting solution does not appreciably change its pH.

The solution is a buffer. It will resist changing its pH as the HA reacts with the OH⁻.

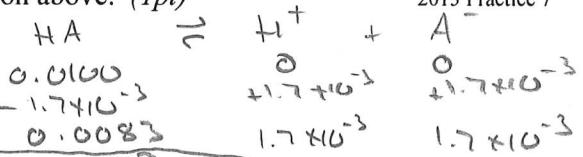


The molecular formula of acetylsalicylic acid, also known as aspirin, is $\text{HC}_9\text{H}_7\text{O}_4$. The dissociation of $\text{HC}_9\text{H}_7\text{O}_4(aq)$ is represented by the equation above. The pH of $0.0100 M \text{HC}_9\text{H}_7\text{O}_4(aq)$ is measured to be 2.78.

- a) Write the expression for the equilibrium constant, K_a , for the reaction above. (1pt)

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{C}_9\text{H}_7\text{O}_4^-]}{[\text{HC}_9\text{H}_7\text{O}_4]}$$

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- b) Calculate the value of K_a for acetylsalicylic acid. (2pts)

$$[\text{H}_3\text{O}^+] = 10^{-2.78} = 1.7 \times 10^{-3} M$$

~~not small enough to ignore~~

$$K_a = \frac{(1.7 \times 10^{-3})^2}{[0.0100 - 1.7 \times 10^{-3}]} = 3.5 \times 10^{-4}$$

- c) An aqueous solution of aspirin is buffered to have equal concentrations of $\text{HC}_9\text{H}_7\text{O}_4(aq)$ and $\text{C}_9\text{H}_7\text{O}_4^-(aq)$.

Calculate the pH of the solution. (1pt)

$$\text{pH} = \text{p}K_a = -\log(3.5 \times 10^{-4}) = 3.46$$

A volume of 30.0 mL of $0.10 M \text{NH}_3(aq)$ is titrated with $0.20 M \text{HCl}(aq)$. The value of the base-dissociation constant, K_b , for NH_3 in water is 1.8×10^{-5} at 25°C.

- a) Write the net-ionic equation for the reaction of $\text{NH}_3(aq)$ with $\text{HCl}(aq)$. (1pt)



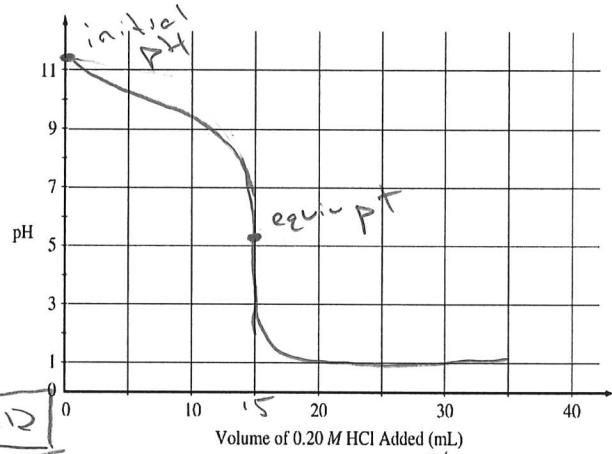
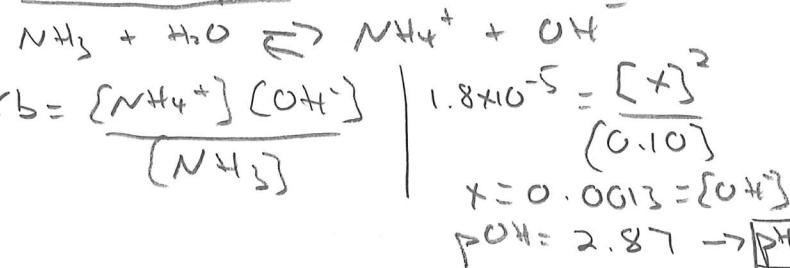
~~strong~~

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- b) Using the axes provided, sketch the titration curve

that results when a total of 40.0 mL of
 $0.20 M \text{HCl}(aq)$ is added dropwise to the
30.0 mL volume of $0.10 M \text{NH}_3(aq)$. (3pts)

initial pH:

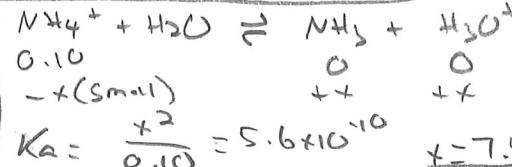


equivalence point:

$$M_a V_a = M_b V_b$$

$$(0.20)(V_a) = (0.10)(30.0)$$

$$V_a = 15 \text{ mL}$$



$$K_a = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.6 \times 10^{-10}$$

- c) From the table, select the most appropriate indicator for the titration. Justify your choice. (2pts)

Methyl Red - $\text{p}K_a$ is closest to equivalence pH.

| Indicator | $\text{p}K_a$ |
|------------------|---------------|
| Methyl Red | 5.5 |
| Bromothymol Blue | 7.1 |
| Phenolphthalein | 8.7 |

- d) If equal volumes of $0.10 M \text{NH}_3(aq)$ and $0.10 M \text{NH}_4\text{Cl}(aq)$ are mixed, is the resulting solution acidic, neutral, or basic? Explain. (2pts) Buffer

$$K_b \text{ of } \text{NH}_3 = 1.8 \times 10^{-5} \quad K_a \text{ of } \text{NH}_4^+ = 5.6 \times 10^{-10}$$

Higher K_b means the solution will be basic
(NH_3 is a stronger base than NH_4^+ is an acid)

