

Name: KEY
Practice Test

General Chemistry

1. The conjugate acid of $C_4H_5O_6^-$ is:

- a. $C_4H_4O_6^{2-}$
- b. $C_4H_6O_6$
- c. $C_4H_7O_6^+$
- d. none of the above

2. Water has amphiprotic properties and acts like a base when reacted with:

- a. NO_3^-
- b. NH_3
- c. NH_4^+
- d. SO_4^{2-}

3. Which of the following salts produces an acidic solution:

- a. NaCN
- b. NH_4Cl
- c. NaF
- d. $Ca(NO_3)_2$

4. According to the Bronsted-Lowry theory an acid is a(n) _____ while according to the Lewis definition a base is a(n) _____.

- a. proton donor, electron-pair acceptor
- b. electron-pair donor, proton donor
- c. proton acceptor, electron-pair donor
- d. proton donor, electron-pair donor

5. Based on your knowledge of trends in acid strength, which is the correct order of strongest to weakest bases. (weakest to strongest acid)

- a. $ClO^- > BrO^- > IO^-$
- b. $IO^- > BrO^- > ClO^-$
- c. $HOCl > HOBr > HOI$
- d. $BrO^- > IO^- > ClO^-$

6. Suppose you wanted to make a buffer solution that would decrease the change in pH with the addition of a strong base. What would the best option for this proposed buffer.

- a. HF and NaF (HF is stronger than NH_4^+)
- b. NH_3 and NH_4Cl
- c. HCl and NaCl
- d. NaOH and NaBr → not buffers.

7. Which of the following salts would be added to a solution to reduce the hydronium ion concentration.

- a. $\text{Ca}(\text{Cl})_2$
- b. NH_4Br
- c. HNO_3
- d. NaCN

8. A strong base is added to a solution containing an aqueous weak acid. Before the titration has reached the equivalence point and after some base has been added, what is in the beaker.

- a. weak acid and its conjugate salt ~~strong acid~~ (buffer)
- b. strong base and its conjugate salt
- d. just a salt
- e. strong acid from the acid base reaction

9. Which if the following species can act as a Lewis Acid:

- a. NH_3
- b. F^-
- c. BF_3
- d. H_2O

10. A buffer's capacity will increase if you _____.

- a. add a strong acid or base to the buffer
- b. add water to the buffer
- c. increase the buffer's molarity
- d. you cannot increase the buffer's capacity.

11. What is the pH of a solution containing .15 M HCN and .15 M NaCN?

($K_a = 4.9 \times 10^{-10}$)

- a. 9.31
- b. 4.54
- c. 10.31
- d. 4.34

12. The percent dissociation of a .10M weak base is 4.5%, what is the K_b ?

- a. 1.025×10^{-5}
- b. 2.025×10^{-4}
- c. 2.025×10^{-5}
- d. 1.025×10^{-4}

13. What mass of KOH is necessary to prepare 800.0 mL of a solution having a pH of 11.56?

- a. 0.16g
- b. 0.23g
- c. 1.02g
- d. 0.98g

14. In lab you are asked to prepare a buffer solution with a pH of 5.00. In the hood you are given sodium fluoride and hydrofluoric acid. What ratio of base to acid is required to obtain the desired pH?

- a. 34:1
- b. 72:1**
- c. 54:1
- d. 43:1

$$pH = pK_a + \log\left(\frac{[B]}{[A]}\right)$$

$$\frac{[B]}{[A]} = 10^{pH - pK_a} = 10^{5.00 - (-\log(7.2 \times 10^{-4}))}$$

$$= 72$$

15. What will the pOH of a titration of a weak base and strong acid (titrant) be just after the equivalence point has been reached?

- a. below 7
- b. above 7**
- c. 7 exactly
- d. not enough information given

Only conjugate acid will be present! $\therefore pH < 7$
 $pOH > 7$

16. Aspirin (acetylsalicylic acid $C_{18}H_{21}NO_3$) is a weak acid that moves from the stomach (pH=2), through the intestinal mucosa into the blood stream (pH=7.4), and finally to the blood-brain barrier to treat headaches, minor pains, or prevent blood clots. Much of the effectiveness of aspirin is due its acid-base properties and its ability to dissociate and adjust its equilibrium in different pH environments. If someone takes 75mg of acetylsalicylic acid with one cup (237mL) of water, what is the pH of the solution taken? ($K_a = 3.27 \times 10^{-4}$)

$$[C_{18}H_{21}NO_3]_i = \frac{[75 \text{ mg} \cdot \frac{1 \text{ g}}{1000 \text{ mg}} \cdot \frac{1 \text{ mol}}{299.39 \text{ g}}]}{0.237 \text{ L}} = 0.0010570016 \text{ M}$$

$$K_a = 3.27 \times 10^{-4} = \frac{[H^+]^2}{[C_{18}H_{21}NO_3]_i - [H^+]} \rightarrow \text{ign.}$$

$$\therefore [H^+] = \sqrt{(3.27 \times 10^{-4})(0.0010570016)}$$

$$[H^+] = 5.879111525 \times 10^{-4} \text{ M}$$

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

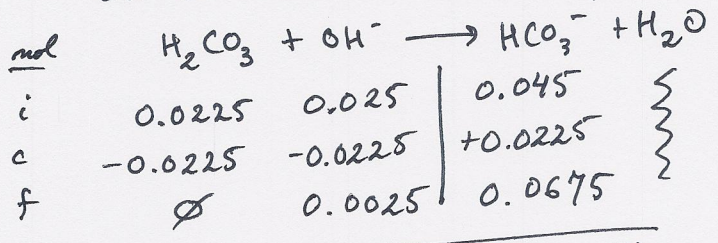
$$= -\log(5.879111525 \times 10^{-4} \text{ M})$$

$$pH = 3.231$$

17. A 150-mL solution containing 0.15M carbonic acid and 0.30M sodium bicarbonate forms a buffer in solution. 25-mL of 1.0M potassium hydroxide is slowly added to the solution. What is the new pH of the solution? Did the buffer survive? ($K_a = 4.3 \times 10^{-7}$)

$$\begin{aligned} \text{mol } H_2CO_3 &= (0.15M)(0.150L) = 0.0225 \text{ mol} \\ \text{mol } HCO_3^- &= (0.30M)(0.150L) = 0.045 \text{ mol} \\ \text{mol } OH^- &= (1.0M)(0.025L) = 0.025 \text{ mol} \end{aligned} \left. \begin{array}{l} \text{moles present} \\ \text{in buffer at equilibrium} \\ \text{before we add KOH.} \end{array} \right\}$$

OH^- will react with the acid-part of the buffer:



The strong base present will dominate and dictate the pH.

$$\begin{aligned} \therefore pH &= 14 - pOH \\ &= 14 + \log [OH^-] \\ &= 14 + \log \left(\frac{0.0025 \text{ mol}}{0.175 L} \right) \end{aligned}$$

pH = 12.15
buffer died

18. A weak acid-strong base titration is done with 45-mL solution of .12M $C_2H_4O_2$ and 1.0M LiOH. How many mL of LiOH are needed to reach the equivalence point? What is the pH of the solution at the equivalence point? ($K_a = 1.8 \times 10^{-5}$)

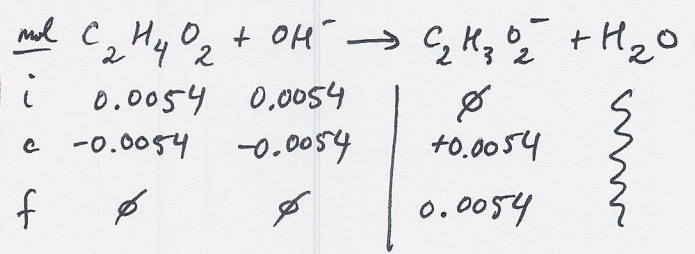
* Weak Acid & OH^- are 1:1 ratio.

@ equivalence point

$$\begin{aligned} \text{mol Acid}_i &= \text{mol Base added} \\ (0.12M)(0.045L) &= (1.0M)(V_{\text{added}}) \end{aligned}$$

$$\therefore V_{\text{added}} = 0.0054 L$$

\therefore 5.4 mL of LiOH needed to reach eq. pt.



$$K_b = \frac{K_w}{K_a} = \frac{[OH^-]^2}{[C_2H_3O_2^-]_i - [OH^-]_{\text{ign.}}}$$

$$[OH^-] = \sqrt{\left(\frac{0.0054 \text{ mol}}{0.0504 L} \right) \left(\frac{10^{-14}}{1.8 \times 10^{-5}} \right)}$$

$$[OH^-] = 7.715167498 \times 10^{-6} M$$

$$\begin{aligned} pH &= -pOH + 14 \\ &= \log(7.715167498 \times 10^{-6} M) + 14 \end{aligned}$$

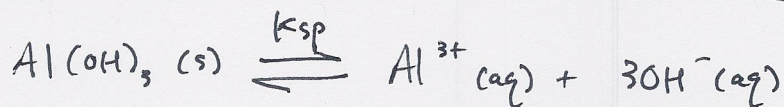
pH = 8.89

Additional Test Questions:

- Which of the follow are true regarding a buffer solution?
 - A buffer is not affected by dilution (it's pH)
 - All buffers have a common ion effect
 - Buffers have the highest capacity when their pH is equal to the pKa of the acid in solution
 - All of the above
- A 0.20M of an electrolyte has a pH of 4.00, the electrolyte is:
 - A strong acid
 - A strong base
 - A weak acid
 - A weak base
- Which of the following salts would have a pH less than 7.00?
 - NH_4Cl
 - NaOH
 - HCl
 - KF
- Calculate the solubility product constant for lead(II) iodide if 0.0024 mole of I ion is present in 2.0 L of a saturated lead(II) iodide solution
 - 1.4×10^{-5}
 - 8.6×10^{-10}
 - 5.2×10^{-8}
 - 3.5×10^{-6}
 - 4.6×10^{-9}
- Calculate the pH of a solution necessary to just begin the precipitation of magnesium hydroxide when the concentration of magnesium ion = 0.001 M. For magnesium hydroxide $K_{sp} = 1.2 \times 10^{-11}$.
 - 11
 - 10
 - 9
 - 8

6. Will a precipitate form when 125 ml of 0.0250 M aluminum nitrate and 25.0 ml of 0.000100 M calcium hydroxide are mixed together? Why? K_{sp} of aluminum hydroxide = 3.7×10^{-15}

Precipitate will form iff $Q_{sp} > K_{sp}$



$$Q_{sp} = \frac{[\text{Al}^{3+}][\text{OH}^-]^3}{1}$$

← ignore solids!

$$[\text{Al}^{3+}]_i = [\text{Al}(\text{NO}_3)_3]_i = 0.0250 \text{ M}$$

$$[\text{Al}^{3+}]_{\text{diluted}} = \frac{(0.0250 \text{ M})(125 \text{ mL})}{(125 + 25 \text{ mL})} = 0.0208\bar{3} \text{ M}$$

$$[\text{OH}^-]_i = 2 \cdot [\text{Ca}(\text{OH})_2] = 0.000200 \text{ M}$$

$$[\text{OH}^-]_{\text{diluted}} = \frac{(0.000200 \text{ M})(25 \text{ mL})}{(125 + 25 \text{ mL})} = 3.3 \times 10^{-5} \text{ M}$$

$$Q_{sp} = (0.0208\bar{3})(3.3 \times 10^{-5})^3$$

$$Q_{sp} = 7.716 \times 10^{-16}$$

$$Q_{sp} < K_{sp}$$

∴ No precipitate