Name: KEY General Chemistry

Practice Test

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[redacted] 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. $Ca(Cl)_2$
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 (buffer)
b. stong 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? ($Ka = 4.9 \times 10^{-10}$)

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 Kb?

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

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

$$\frac{[B]}{[A]} = 10^{pH - pka} = 10^{5.00 - (-\log(7.2\times10^{-4}))} = 72$$

15. What will the pOH of a titration of a weak base and strong acid (titrant) be just after the equivalence point has been reach 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), though 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? ($K_a = 3.27 \times 10^{-4}$)

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

$$K_a = 3.27\times10^{-4} = \frac{[H^+]^2}{[C_{18}H_{21}NO_3]_i - \cancel{[H^+]}} \text{ ign.}$$

$$\therefore [H^+] = \sqrt{(3.27\times10^{-4})(0.0010570016)}$$

$$[H^+] = 5.879111525\times10^{-4}\ M$$

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

$$= -\log(5.879111525\times10^{-4}\ M)$$

$$\boxed{pH = 3.231}$$

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? ($Ka = 4.3 \times 10^{-7}$)

$$\left.\begin{array}{l} mol\ H_2CO_{3\,i} = (0.15\,M)(0.150\,L) = 0.0225\,mol \\ mol\ HCO^-_{3\,i} = (0.30\,M)(0.150\,L) = 0.045\,mol \end{array}\right\}$$ moles present in buffer at equilibrium before we add KOH.

$mol\ OH^-_i = (1.0\,M)(0.025\,L) = 0.025\,mol$

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

mol	H_2CO_3	$+\ OH^- \longrightarrow$	$HCO_3^- + H_2O$
i	0.0225	0.025	0.045
c	−0.0225	−0.0225	+0.0225
f	∅	0.0025	0.0675

The strong base present will dominate and dictate the pH.

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

$\boxed{pH = 12.15 \text{ buffer died}}$

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? ($Ka = 1.8 \times 10^{-5}$)

* Weak Acid & OH⁻ are 1:1 ratio.

@ equivalence point

$mol\ Acid_i = mol\ Base_{added}$

$(0.12\,M)(0.045\,L) = (1.0\,M)(V_{added})$

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

$\therefore$ $\boxed{\text{5.4 mL of LiOH needed to reach eq. pt.}}$

mol	$C_2H_4O_2$	$+\ OH^- \longrightarrow$	$C_2H_3O_2^- + H_2O$
i	0.0054	0.0054	∅
c	−0.0054	−0.0054	+0.0054
f	∅	∅	0.0054

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

$$[OH^-] = \sqrt{\left(\frac{0.0054\,mol}{0.0504\,L}\right)\left(\frac{10^{-14}}{1.8\times10^{-5}}\right)}$$

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

$pH = -pOH + 14$
$= \log(7.715167498 \times 10^{-6}\,M) + 14$

$\boxed{pH = 8.89}$

nobody cares

19. ~~Hydromorphone (Dilaudid) is a class II analgesic narcotic that is eight times stronger than morphine and belongs to the opioid family~~. If the pH of a 0.0107M solution if Dilaudid is 11.07, determine the pK_a of the solution.

pH > 7 ∴ we are dealing with a species that acts as a base around H_2O.

$$B + H_2O \overset{K_b}{\rightleftharpoons} BH^+ + OH^-$$

	B + H_2O	BH^+	OH^-
[I]	0.0107	∅	∅
[C]	$-x$	$+x$	$+x$
[F]	$0.0107-x$	x	x

note $x = [OH^-]_{eq} = 10^{-pOH} = 10^{-(14-11.07)}$

∴ $x = 0.0011748976$ M

$$K_b = \frac{x^2}{(0.0107-x)} = \frac{K_w}{K_a}$$

$$\therefore K_a = \frac{K_w(0.0107-x)}{x^2}$$

$$pK_a = -\log\left[\frac{K_w(0.0107-x)}{x^2}\right]$$

$$pK_a = -\log\left[\frac{10^{-14}(0.0107-0.0011748976)}{(0.0011748976)^2}\right]$$

$$\boxed{pK_a = 10.16}$$

20. A 0.10M solution of HCl is added from a buret to a 35-mL of 0.10M NH_3. What is the pH of the solution after 32-mL of HCl is added to the flask?

First find equivalence pt!

mol $Base_i$ = Mol $Acid_{added}$

$(0.10M)(0.035L) = (0.10M)(V_{Acid\,added})$

$V_{Acid\,added} = 0.035\ L = 35\ mL$

∴ 32 mL is before eq. pt.

mol	NH_3	+ H^+	→ NH_4^+
i	0.0035	0.0032	∅
c	−0.0032	−0.0032	+0.0032
f	0.0003	∅	0.0032

Buffer!

$$pH = pKa + \log\left(\frac{Base}{Acid}\right)$$

$$= -\log\left(\frac{10^{-14}}{1.8\times10^{-5}}\right) + \log\left(\frac{0.0003\ mol}{0.0032\ mol}\right)$$

$$\therefore \boxed{pH = 8.23}$$

Additional Test Questions:

1. Which of the follow are true regarding a buffer solution?
    a. A buffer is not affected by dilution (it's pH)
    b. All buffers have a common ion effect
    c. Buffers have the highest capacity when their pH is equal to the pKa of the acid in solution
    d. All of the above

2. A 0.20M of an electrolyte has a pH of 4.00, the electrolyte is:
    a. A strong acid
    b. A strong base
    c. A weak acid
    d. A weak base

3. Which of the following salts would have a pH less than 7.00?
    a. NH_4Cl
    b. NaOH
    c. HCl
    d. KF

4. 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

    a. 1.4×10^{-5}

    b. 8.6×10^{-10}

    c. 5.2×10^{-8}

    d. 3.5×10^{-6}

    e. 4.6×10^{-9}

5. 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 $K_{sp} = 1.2 \times 10^{-11}$.

    a. 11

    b. 10

    c. 9

    d. 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 $Q_{sp} > K_{sp}$

$$Al(OH)_3\,(s) \underset{}{\overset{K_{sp}}{\rightleftharpoons}} Al^{3+}\,(aq) + 3OH^-\,(aq)$$

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

← ignore solids!

$$[Al^{3+}]_i = [Al(NO_3)_3]_i = 0.0250\ M$$

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

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

$$[OH^-]_{diluted} = \frac{(0.000200\,M)(25\,mL)}{(125+25\ mL)} = 3.\bar{3} \times 10^{-5}\ M$$

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

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

$$\boxed{\begin{array}{l} Q_{sp} < K_{sp} \\ \therefore \text{No precipitate} \end{array}}$$