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MWF 2pm - 3pm

Exam 3

Fall 2018

Nov 7, 2018

Wednesday 7:30 - 9:00 PM

BUR 106

Remember to refer to the Periodic Table handout that is separate from this exam copy.

NOTE: Please keep this exam copy intact (all pages still stapled - including this cover page). You must turn in **ALL** the materials that were distributed. This means that you turn in your exam copy (name and signature included), bubble sheet, periodic table handout, and all scratch paper. Please also have your UT ID card ready to show as well.

This print-out should have 25 questions. Multiple-choice questions may continue on the next column or page – find all choices before answering.

001 4.0 points

What is the molar solubility of CuBr in 0.5 M NaBr? The K_{sp} is 4.2×10^{-8} .

1. 4.20×10^{-7}
2. 4.20×10^{-8}
3. 8.40×10^{-8} **correct**
4. 3.48×10^{-3}
5. 2.05×10^{-4}

Explanation:

002 4.0 points

What is the molar solubility of $Zn(OH)_2$ in a solution buffered at pH of 8.0? ($Zn(OH)_2$, $K_{sp} = 3.0 \times 10^{-17}$)

1. 3.0×10^{-5} **correct**
2. 2.0×10^{-6}
3. 5.0×10^{-6}
4. 8.0×10^{-4}
5. 4.0×10^{-9}

Explanation:

pH=8 means that pOH=6 and $[OH^-]=10^{-6}$
 $K_{sp} = [Zn^{2+}][OH^-]^2$ (solve for Zn^{2+})
 $[Zn^{2+}] = 3.0 \times 10^{-17}/(10^{-6})^2$
 $[Zn^{2+}] = 3.0 \times 10^{-5}$ M

003 4.0 points

A metal hydroxide has the general formula, $M(OH)_3$, and a K_{sp} equal to 3.5×10^{-24} . Enough of the compound is added to water such that the solution is saturated. What is the pH of this solution?

1. 8.63

2. 9.14

3. 8.07

4. 7.00

5. 7.42

6. 7.80

7. 8.95

8. 6.22

9. 5.74

10. 8.26 **correct**

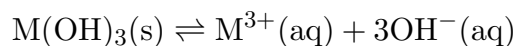
Explanation:

This is a 1:3 salt and therefore the molar solubility, $x = \sqrt[4]{\frac{K_{sp}}{27}}$.

$$x = \sqrt[4]{\frac{3.5 \times 10^{-24}}{27}}$$

$$x = 6.0 \times 10^{-7}$$

Remember that the dissociation is



and that the concentration of OH^- is $3x$ or 3 times the molar solubility. Therefore:

$$[OH^-] = 3(6.0 \times 10^{-7}) = 1.8 \times 10^{-6} \text{ M}$$

$$pOH = -\log(1.8 \times 10^{-6}) = 5.74$$

$$pH = 14 - 5.74 = 8.26$$

004 4.0 points

An indicator changes from red to blue going from its acidic (nonionized) form to its basic (ionized) form. The indicator has a K_a of 3.2×10^{-5} . What color would be visible in a solution with this indicator at pH 5.62?

1. blueish-purple

2. blue **correct**

3. purple

4. red

5. reddish-purple

Explanation:

The pK_a is 4.5 so that is where you would see purple (50/50 blend of red and blue). You would see all red for pH 3.5 and lower, and all blue for 5.5 and above. The 5.62 pH given is definitely beyond the transition range and you would see BLUE.

005 4.0 points

50.0 mL of an HCl solution with a pH of 3.5 neutralizes 200.0 mL of a Ca(OH)_2 solution. What is the molarity of the Ca(OH)_2 solution?

1. 3.95×10^{-5} M **correct**

2. 0.0342 M

3. 4.55×10^{-6} M

4. 1.69 M

5. 1.78×10^{-4} M

6. 7.91×10^{-5} M

7. 14.3 M

8. 2.11×10^{-5} M

9. 1.58×10^{-4} M

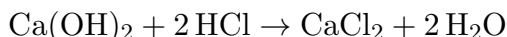
Explanation:

$V_{\text{HCl}} = 50 \text{ mL}$

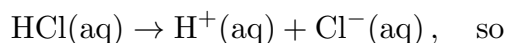
pH = 3.5

$V_{\text{Ca(OH)}_2} = 200 \text{ mL} = 0.2 \text{ L}$

The reaction is



Find the moles of HCl. This is a strong acid:



$$[\text{HCl}] = [\text{H}^+] = 10^{-\text{pH}}$$

$$= 10^{-3.5} = 0.000316228 \text{ M}$$

$$\begin{aligned} (50.0 \text{ mL}) \times \frac{1 \text{ L}}{1000 \text{ mL}} \\ \times \frac{0.000316228 \text{ mol HCl}}{1 \text{ L}} \\ = 1.58114 \times 10^{-5} \text{ mol HCl} \end{aligned}$$

Find the moles of Ca(OH)_2 :

$$\begin{aligned} (1.58114 \times 10^{-5} \text{ mol HCl}) \times \frac{1 \text{ mol Ca(OH)}_2}{2 \text{ mol HCl}} \\ = 7.90569 \times 10^{-6} \text{ mol Ca(OH)}_2 \end{aligned}$$

Thus the concentration is

$$\begin{aligned} [\text{Ca(OH)}_2] &= \frac{7.90569 \times 10^{-6} \text{ mol Ca(OH)}_2}{0.2 \text{ L}} \\ &= 3.95285 \times 10^{-5} \text{ M Ca(OH)}_2 \end{aligned}$$

006 4.0 points

What is the pH of a solution containing 100 mL of 0.5 M NH_3 and 300 mL of 0.1 M HCl? The pK_b of ammonia is 4.75.

1. 8.551

2. 7.000

3. 9.074 **correct**

4. 4.574

5. 9.949

6. 4.926

7. 9.426

Explanation:

Reaction is: $\text{NH}_3 + \text{H}^+ \rightarrow \text{NH}_4^+$

$\text{NH}_3 = 50 \text{ mmol}$

$\text{H}^+ = 30 \text{ mmol}$ (limiting reactant)

So after the reaction, the amounts of conjugates are:

$\text{NH}_3 = 20 \text{ mmol}$

$\text{NH}_4^+ = 30 \text{ mmol}$

Both are now in 400 mL of solution.

H-H Equation: $\text{pH} = \text{pKa} + \log(\text{base/acid})$

$$\text{pH} = 9.25 + \log(20/30) = 9.074$$

007 4.0 points

If a small amount of a strong base is added to a buffer made up of a weak acid HA and the sodium salt of its conjugate base NaA, the pH of the buffer solution does not change appreciably because

1. the strong base reacts with HA to give AOH and H^+ .
2. the strong base reacts with HA to give A^- , a weak base. **correct**
3. the strong base reacts with A^- to give HA, a weak acid.
4. the K_a of HA is changed.
5. no reaction occurs.

Explanation:

008 4.0 points

Consider the K_{sp} values of the following salts and indicate which of these is least soluble in water.

1. AB_4 $K_{sp} = 1 \times 10^{-35}$ **correct**
2. AB_3 $K_{sp} = 1 \times 10^{-24}$
3. AB $K_{sp} = 1 \times 10^{-8}$
4. AB_2 $K_{sp} = 1 \times 10^{-15}$

Explanation:

009 4.0 points

Which of the following shows the right equation for the ion product of lead(II) chlorate, $Pb(ClO_3)_2$?

1. $[Pb^{2+}][ClO_3^-]$
2. $[Pb^{2+}][ClO_3^-]^2$ **correct**
3. $[Pb^{2+}](2 * [ClO_3^-])^2$
4. $[Pb^{2+}]^2[ClO_3^-]$

5. $[Pb^{2+}]^2[ClO_3^-]^2$

Explanation:

$$K_{sp} = [Pb^{2+}][ClO_3^-]^2$$

010 4.0 points

Which of the following would be equal to K_{a3} for orthocarbonic acid, H_4CO_4 ?

1. $\frac{[H_2CO_4^{2-}]}{[H^+][HCO_4^{3-}]}$
2. $\frac{[H^+][HCO_4^{3-}]}{[H_2CO_4^{2-}]}$ **correct**
3. $\frac{[H_4CO_4]}{[H^+]^3[HCO_4^{3-}]}$
4. $\frac{[H^+]^3[HCO_4^{3-}]}{[H_2CO_4^{2-}]}$
5. $\frac{[H^+]^3[HCO_4^{3-}]}{[H_4CO_4]}$

Explanation:

K_{a3} for orthocarbonic acid, H_4CO_4 , describes the 3rd deprotonation event. $H_2CO_4^{2-} \leftrightarrow H^+ + HCO_4^{3-}$

011 4.0 points

For a triprotic acid, H_3A , with pK_a values of 2.5, 6.5 and 10.5, what species is present in highest concentration when 1M H_3A is buffered at pH 7?

1. H_2A^-
2. A^{3-}
3. HA^{2-} **correct**
4. H^+
5. H_3A

Explanation:

Since the pH is greater than pK_{a2} , but much closer to pK_{a2} than to pK_{a3} , the most abundant species in solution will be the second deprotonated species, HA^{2-} .

012 4.0 points

Which of the following indicators would be

most suitable for the titration of 0.10 M $(\text{CH}_3)_3\text{N}$ ($K_b = 6.5 \times 10^{-5}$) with 0.10 M HCl? Transition ranges in pH units are given in parentheses after the indicator name.

1. chlorophenol red (4.8 - 6.4) **correct**
2. bromocresol green (3.8 - 5.4)
3. thymol blue (1.2 - 2.8)
4. phenolphthalein (8.0 - 9.6)
5. alizarin yellow (10.1 - 12.0)
6. bromothymol blue (6.0 - 7.6)

Explanation:

At the equivalence point the solution will be 0.05 M $(\text{CH}_3)_3\text{NH}^+$ which has a $K_a = 1.54 \times 10^{-10}$. This translates to a equivalence point pH of 5.56. The chlorophenol red has 5.56 in the middle of its transition range.

013 4.0 points

What is the pH of a solution which is 0.600 M in dimethylamine $((\text{CH}_3)_2\text{NH})$ and 0.400 M in dimethylamine hydrochloride $((\text{CH}_3)_2\text{NH}_2^+\text{Cl}^-)$? K_b for dimethylamine = 7.4×10^{-4} .

1. 11.05 **correct**
2. 2.95
3. 10.69
4. 10.87
5. 10.78
6. 3.31
7. 11.21

Explanation:

$K_w = 1 \times 10^{-14}$ $K_a = 0.00074$
 $[(\text{CH}_3)_2\text{NH}] = 0.6 \text{ M}$ $[(\text{CH}_3)_2\text{NH}_2^+] = 0.4 \text{ M}$
 $K_a, (\text{CH}_3)_2\text{NH}_2^+ = \frac{K_w}{K_b, (\text{CH}_3)_2\text{NH}}$
 Applying the Henderson-Hasselbalch equa-

tion,

$$\begin{aligned} \text{pH} &= \text{p}K_a + \log \left(\frac{[(\text{CH}_3)_2\text{NH}]}{[(\text{CH}_3)_2\text{NH}_2^+]} \right) \\ &= -\log \left(\frac{K_w}{K_a} \right) + \log \left(\frac{[(\text{CH}_3)_2\text{NH}]}{[(\text{CH}_3)_2\text{NH}_2^+]} \right) \\ &= -\log \left(\frac{1 \times 10^{-14}}{0.00074} \right) + \log \left(\frac{0.6}{0.4} \right) \\ &= 11.0453 \end{aligned}$$

014 4.0 points

Consider 0.7 L of a buffer composed of 0.4 M HClO and 0.5 M NaClO? How many moles of strong base could this buffer withstand?

1. 0.28 **correct**
2. 0
3. 0.4
4. 0.35
5. 0.5

Explanation:

Any added strong base will react with the weak acid that is present (in this case HClO), of which there are 0.28 moles (0.7 L x 0.4 M = 0.28 moles).

015 4.0 points

What will be the pH at the equivalence point of a titration of 0.5 M acrylic acid with an equimolar solution of NaOH? Acrylic acid has a K_a of 5.6×10^{-5} .

1. 8.97
2. Not enough information is given.
3. 11.57
4. 8.82 **correct**
5. 7.00
6. 5.18

Explanation:

Equal volumes of the titrant and analyte will be used to reach the equivalence point. Regardless of the starting volume, at the equivalence point the volume will be double the starting value and all of the acrylic acid (HA) will have been converted to acrylate (A^-). The solution will be 0.25 M acrylate.

$$K_b = \frac{K_w}{K_a} = \frac{10^{-14}}{5.6 \times 10^{-5}} = 1.79 \times 10^{-10}$$

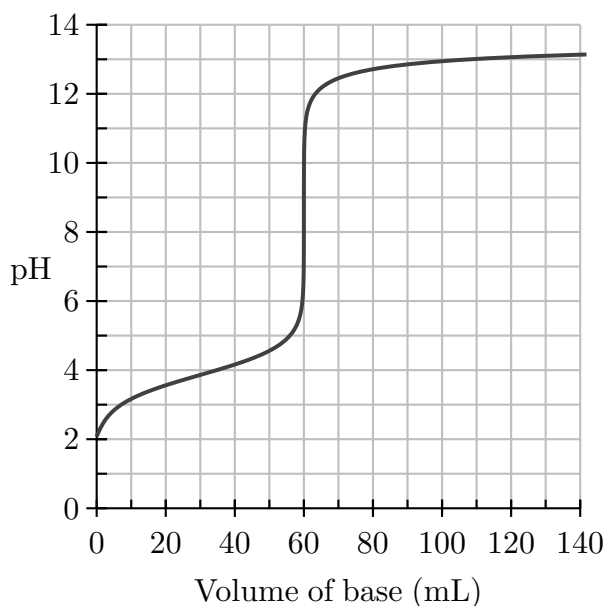
$$\begin{aligned} [OH^-] &= (K_b C_b)^{1/2} \\ &= (1.79 \times 10^{-10} \cdot 0.25)^{1/2} \\ &= 6.68 \times 10^{-6} \end{aligned}$$

$$\text{pOH} = 5.18$$

$$\text{pH} = 8.82$$

016 4.0 points

Consider the following titration curve and data table.



Acid	pKa
pyruvic	2.50
lactic	3.86
acetic	4.74
aconitic	1.84

What is the identity of the analyte?

1. aconitic

2. pyruvic

3. lactic **correct**

4. acetic

Explanation:**017 4.0 points**

The solubility product constant of barium sulfite ($BaSO_3$) is 8.0×10^{-7} . Will a precipitate of $BaSO_3$ form if 100 mL of a 4.0×10^{-4} molar solution of $BaCl_2$ is added to 100 mL of a 3.0×10^{-3} molar solution of Na_2SO_3 ?

1. yes, because the solubility product constant is exceeded

2. yes, because the solubility product constant is not exceeded

3. no, because the solubility product constant is not exceeded **correct**

4. no, because the solubility product constant is exceeded

Explanation:**018 4.0 points**

The K_{sp} of $Cd_3(PO_4)_2$ at 18 °C is 1.08×10^{-33} . What is its molar solubility at this temperature?

1. 3.3×10^{-17} M

2. 1.0×10^{-7} M **correct**

3. 6.5×10^{-11} M

4. 2.5×10^{-9} M

Explanation:

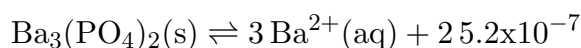
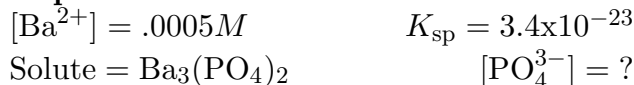
The equation for $K_{sp} = (2x)^2(3x)^3 = 108x^5$

$$x = \left(\frac{1.08 \times 10^{-33}}{108} \right)^{1/5} = 1 \times 10^{-7} \text{ M}$$

019 4.0 points

A solution in equilibrium with solid barium phosphate ($Ba_3(PO_4)_2$) is found to have a barium ion concentration of 0.0005 M and a K_{sp} of 3.4×10^{-23} . Calculate the concentration of phosphate ion.

1. 1.3×10^{-7}
2. 1.04×10^{-6}
3. 5.2×10^{-7} **correct**
4. 2.6×10^{-7}
5. 2.7×10^{-13}

Explanation:

$$K_{sp} = [Ba^{2+}]^3 [PO_4^{3-}]^2$$

$$[PO_4^{3-}]^2 = \frac{K_{sp}}{[Ba^{2+}]^3}$$

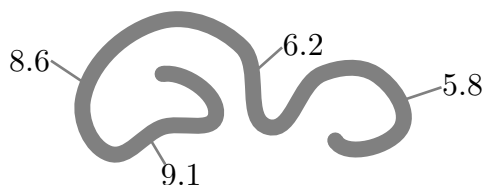
$$[PO_4^{3-}] = \sqrt{\frac{K_{sp}}{[Ba^{2+}]^3}}$$

$$= \sqrt{\frac{3.4 \times 10^{-23}}{(0.0005)^3}}$$

$$= 5.2 \times 10^{-7}$$

020 4.0 points

Consider the diagram below that represents a protein chain. The four positions shown with numbers are the pK_a 's of the carboxylic acid residues in the chain. If the protein is in a pH 7.4 buffer, what is the combined total charge on those four side chains?



1. **-2 correct**

2. +3
3. +2
4. -1
5. 0
6. -3
7. +1
8. -4
9. +4

Explanation:

The two residues with pK_a 's LESS than 7.4 will all be deprotonated and have a -1 charge each. The other two residue at will stay protonated and at zero. So the combined charge is -2 .

021 4.0 points

What will be the molar ratio of SALT : ACID (in nice WHOLE numbers) of sodium hypochlorite (SALT) and hypochlorous acid ($K_a = 3.5 \times 10^{-8}$) in a solution with a pH of 7.058?

1. 4 : 1
2. 7 : 2
3. 2 : 3
4. Correct ratio is not shown.
5. 1 : 3
6. 1 : 2
7. 9 : 2
8. 2 : 5 **correct**

Explanation:

$$K_a = 3.5 \times 10^{-8}$$

$$pH = 7.058$$

$$\text{pH} = \text{p}K_{\text{a}} + \log \frac{[\text{OCl}^-]}{[\text{HOCl}]}$$

$$\begin{aligned} \log \frac{[\text{OCl}^-]}{[\text{HOCl}]} &= \text{pH} - \text{p}K_{\text{a}} \\ &= 7.058 - [-\log(3.5 \times 10^{-8})] \\ &= -0.397932 \end{aligned}$$

$$\begin{aligned} \frac{\text{mol NaOCl}^-}{\text{mol HOCl}} &= \frac{[\text{OCl}^-]}{[\text{HOCl}]} = 10^{-0.397932} \\ &= 0.400007 \end{aligned}$$

022 4.0 points

What is K_{sp} for HgI_2 , if its molar solubility is $2.2 \times 10^{-10} \text{ mol/L}$?

1. 1.4×10^{-21}

2. 6.9×10^{-11}

3. 4.7×10^{-30}

4. 8.3×10^{-8}

5. 5.2×10^{-25}

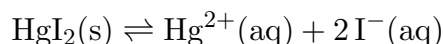
6. 1.7×10^{-14}

7. 4.3×10^{-29} **correct**

Explanation:

$$S = 2.2 \times 10^{-10} \text{ mol/L}$$

The solubility equilibrium is



$$[\text{Hg}^{2+}] = S = 2.2 \times 10^{-10} \text{ mol/L}$$

$$[\text{I}^{-}] = 2S = 4.4 \times 10^{-10} \text{ mol/L}$$

$$\begin{aligned} K_{\text{sp}} &= [\text{Hg}^{2+}][\text{I}^{-}]^2 \\ &= (2.2 \times 10^{-10})(4.4 \times 10^{-10})^2 \\ &= 4.2592 \times 10^{-29} \end{aligned}$$

023 4.0 points

A 100 mL buffer solution initially has 1.00 M HClO ($K_{\text{a}} = 3.0 \times 10^{-8}$) and 0.75 M NaClO . What is the new pH when this 100 mL of 0.10 M NaOH is added to the initial buffer solution?

1. 7.50 **correct**

2. 9.52

3. 11.00

4. 10.71

Explanation:

Consider the molarity (in mmol):

HClO	$+$	NaOH	\rightarrow	NaClO	$+$	H_2O
100		10		75		
-10		-10		10		
90		0		85		

$$\begin{aligned} \text{pH} &= \text{p}K_{\text{a}} + \log \left(\frac{85/200}{90/200} \right) \\ &= 7.52 + \log \left(\frac{85}{90} \right) \\ &= 7.49518 \end{aligned}$$

024 4.0 points

Which of the following mixtures would produce a buffer?

- I) 100 mL of 0.1 M NH_4Cl + 75 mL of 0.1 M NaOH
- II) 100 mL of 0.1 M NH_3 + 100 mL of 0.5 M HCl
- III) 100 mL of 0.1 M NaCl + 50 mL of 0.2 M HCl
- IV) 100 mL of 0.1 M HClO_2 + 100 mL of 0.1 M NaClO_2
- V) 100 mL of 0.1 M NH_3 + 75 mL of 0.1 M NaOH

1. I, IV, and V only

2. I, II, III, and IV only

3. I and IV only **correct**

4. II and III only

5. I only

Explanation:

A buffer contains a weak acid or weak base, plus the salt of that weak acid or base; or, a mixture which will have this composition after any acid-base reactions occur. You may have to calculate the number of moles of each species to determine the composition after any acid-base reaction.

025 4.0 points

You have been asked to create a buffer with a pH of 8.2. Which of these chemicals would you choose? You also have sodium hydroxide and hydrochloric acid to use as needed.

1. Formic acid $K_a = 1.77 \times 10^{-4}$

2. Propionic acid $K_a = 1.34 \times 10^{-5}$

3. Pyridine $K_b = 1.79 \times 10^{-9}$

4. Ammonia $K_b = 1.80 \times 10^{-5}$

5. Hypobromous acid $K_a = 2.00 \times 10^{-9}$ **correct**

Explanation:

The effective operating range of a buffer is +/- 1 pH unit either side of pKa. The best buffer has the closest value to the chosen pH. The pKa's are:

Hypobromous acid: 8.70

Propionic acid: 4.87

Pyridinium (use $K_a K_b = K_w$): 5.25

Formic acid: 3.75

Ammonium (use $K_a K_b = K_w$): 9.25

The closest one to 8.2 is hypobromous acid.