This print-out should have 8 questions. Multiple-choice questions may continue on the next column or page – find all choices before answering.

ChemPrin3e T11 28
001 10.0 points

Which of the following mixtures gives a buffer with a pH less than 7.0? For acetic acid, $K_a = 1.8 \times 10^{-5}$ and for NH_3, $K_b = 1.8 \times 10^{-5}$.

1. 10 mL of 0.1 M aqueous acetic acid + 10 mL of 0.1 M NaOH(aq)

2. 10 mL of 0.1 M NH_3(aq) + 10 mL of 0.1 M HCl(aq)

3. 10 mL of 0.1 M aqueous acetic acid + 5.0 mL of 0.1 M NaOH(aq) **correct**

4. 10 mL of 0.1 M NH_3(aq) + 5.0 mL of 0.1 M HCl(aq)

5. 10 mL of 0.1 M aqueous acetic acid + 10 mL of 0.1 M NH_3(aq)

Explanation:

Sparks buffer 003
002 10.0 points

What is the pH of a solution made to be 0.5 M in HCN and 0.3 M in NaCN? K_a for HCN is 4.0×10^{-10}.

1. 9.17 **correct**

2. 6.67

3. 6.67×10^{-10}

4. 9.6

5. 4.83

6. 7.33

Explanation:

Msci 19 0730
003 10.0 points

A 100 mL portion of 0.3 M acetic acid is being titrated with 0.2 M NaOH solution. What is the pH of the solution after 100 mL of the NaOH solution has been added? The ionization constant of acetic acid is 1.8×10^{-5}.

1. pH = 4

2. pH = 4.32

3. pH = 5.27

4. pH = 5.05 **correct**

5. pH = 4.71

Explanation:

$V_{CH_3COOH} = 100$ mL $[CH_3COOH] = 0.3$ M
$V_{NaOH} = 100$ mL $[NaOH] = 0.2$ M
$K_a = 1.8 \times 10^{-5}$

For CH_3COOH, (0.3 M)(0.1 L) = 0.03 mol
For NaOH, (0.2 M)(0.1 L) = 0.02 mol

CH_3COOH +	NaOH	$\rightarrow NaCH_3COO + H_2O$
0.03 mol	0.02 mol	0 mol
−0.02 mol	−0.02 mol	+0.02 mol
0.01 mol	0 mol	0.02 mol

$CH_3COOH \rightarrow$	CH_3COO^- +	H^+
0.01 mol	0.02 mol	
0.2 L	0.2 L	
0.05 M	0.1 M	x

Thus

$$K_a = \frac{[CH_3COO^-][H^+]}{[CH_3COOH]}$$

$$1.8 \times 10^{-5} = \frac{0.1\,x}{0.05}$$

$$x = [H^+] = \frac{K_a\,[CH_3COOH]}{[CH_3COO^-]} = \frac{(1.8 \times 10^{-5})(0.05)}{0.1} = 9 \times 10^{-6}$$

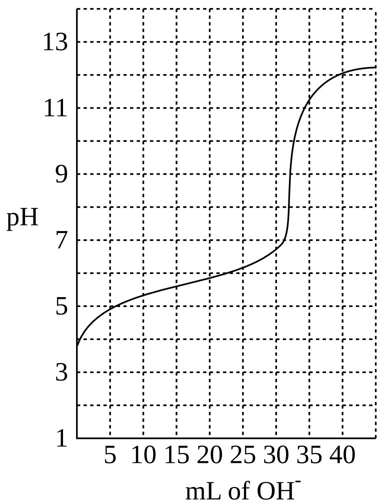
Thus

$$\text{pH} = -\log[H^+] = 5.04576$$

Msci 19 0751

004 10.0 points

What is the pK_a of the acid titrated in this pH curve?



1. 6.8

2. 9.0

3. 5.6 **correct**

4. 5.9

5. 4.7

Explanation:

Msci 19 0402

005 10.0 points

The ionization constants for an imaginary weak acid H_3A are $K_1 = 1.0 \times 10^{-3}$, $K_2 = 1.0 \times 10^{-7}$, and $K_3 = 1.0 \times 10^{-11}$. Would a water solution of NaH_2A be acidic, basic, or neutral? *Hint:* Write the chemical equation and the equilibrium constant for the two possible reactions.

1. neutral

2. acidic because of further ionization of H_2A^- **correct**

3. acidic because of the hydrolysis of H_2A^-

4. basic because of the further ionization of H_2A^-

5. basic because of the hydrolysis of H_2A^-

Explanation:

$$H_2A^- \rightleftharpoons H^+ + HA^{2-}$$

$$K_2 = 10^{-7}$$

$$H_2A^- + H_2O \rightleftharpoons H_3A + OH^-$$

$$K_{b3} = \frac{K_w}{K_{a1}} = \frac{10^{-14}}{10^{-2}} = 10^{-11}$$

PH 10 77 78

006 10.0 points

What is the pH of 0.15 M $NaHSO_3$(aq) if $K_{a1} = 0.015$, $K_{a2} = 1.2 \times 10^{-7}$, p$K_{a1} = 1.81$ and p$K_{a2} = 6.91$?

1. 3.02

2. 7.82

3. 4.36 **correct**

4. None of these

5. 6.92

6. @@@

Explanation:

p$K_{a1} = 1.81$ p$K_{a2} = 6.91$

$M = 0.15$ M

This is a salt of a polyprotic acid. The salt will dissociate into solution. The cation is an extremely weak acid and does not affect the equilibrium. The anion can then either protonate or deprotonate; the extent to which these processes occur is determined by the relative values of pK_{a1} and pK_{a2}. The pH is

$$\begin{aligned} \text{pH} &= \frac{1}{2}(\text{p}K_{a1} + \text{p}K_{a2}) \\ &= \frac{1}{2}(1.81 + 6.91) \\ &= 4.36\,. \end{aligned}$$

Note the pH of a salt solution of a polyprotic acid is independent of the concentration of the salt as long as it is not extremely dilute.

Msci 46 0014

007 10.0 points

What is the molar solubility of CaF_2? ($K_{sp} = 3.9 \times 10^{-11}$.)

1. 3.9×10^{-11}

2. 3.4×10^{-4}

3. 6.2×10^{-6}

4. 4.4×10^{-6}

5. 2.1×10^{-4} **correct**

Explanation:

$$CaF_2 \rightleftharpoons Ca^{2+} + 2\,F^-$$

$$\begin{aligned} K_{sp} &= [Ca^{2+}]\,[F^-]^2 \\ 3.9 \times 10^{-11} &= (x)\,(2x)^2 \\ &= 4\,x^3 \\ x &= 2.1 \times 10^{-4} \end{aligned}$$

Msci 18 0890

008 10.0 points

A buffer (pH 9.25) was prepared by mixing 1.00 mole of ammonia and 1.00 mole of ammonium chloride to form an aqueous solution with a total volume of 1.00 L. To 500 mL of this solution was added 50.0 mL of 1.00 M HCl. What is the pH of this solution?

1. 9.17 **correct**

2. 9.49

3. 9.83

4. 8.97

5. 8.71

Explanation:

$[NH_3] = 1$ M $\qquad [NH_4^+] = 1$ M

$[HCl] = 1$ M $\qquad pH_{ini} = 9.25$

Initial condition (ini):

$n_{NH_3} = 500 \times 1.0 = 500$ mmol

$n_{NH_4^+} = 500 \times 1.0 = 500$ mmol

$n_{HCl} = 50 \times 1.0 = 50$ mmol

$n_{Cl^-} = 500 \times 1.0 = 500$ mmol

	NH_3 +	HCl →	NH_4^+ +	Cl^-
ini	500	50	500	500
Δ	−50	−50	50	50
fin	450	0	550	550

Cl^- are spectator ion. NH_4^+/NH_3 is a buffer system. Since $[NH_3] = [NH_4^+]$ in the original buffer $pK_a = pH_{ini} = 9.25$, and

$$\begin{aligned} pH_{fin} &= pK_a + \log\left(\frac{[NH_3]}{[NH_4^+]}\right) \\ &= -\log\left(\frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}}\right) + \log\left(\frac{450}{550}\right) \\ &= 9.16812 \end{aligned}$$