This print-out should have 20 questions. Multiple-choice questions may continue on the next column or page – find all choices before answering.

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001 10.0 points

What is the pH of an aqueous solution that is 0.018 M $C_6H_5NH_2$ ($K_b = 4.3 \times 10^{-10}$) and 0.12 M $C_6H_5NH_3Cl$?

1. 4.63

2. 3.81 **correct**

3. 9.37

4. 5.46

5. 8.54

6. 10.19

7. 4.02

8. 2.87

Explanation:

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002 10.0 points

A buffer solution is made by dissolving 0.45 moles of a weak acid (HA) and 0.23 moles of KOH into 750 mL of solution. What is the pH of this buffer? $K_a = 2.6 \times 10^{-6}$ for HA.

Correct answer: 5.60433 pH.

Explanation:

$n_{HA} = 0.45$ mol $n_{KOH} = 0.23$ mol

$K_a = 2.6 \times 10^{-6}$ for HA

You must substract the 0.23 moles of KOH from the 0.45 moles of HA because the strong base will neutralize the weak acid. You therefore would *make* 0.23 moles of A^- and be *left* with 0.22 moles of HA. You can now plug this ratio into the equilibrium equation or in the Henderson-Hasselbalch equation to get pH.

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003 10.0 points

Which one of the following combinations is NOT a buffer solution?

1. HCN and NaCN

2. NH_3 and $(NH_4)_2SO_4$

3. CH_3COOH and $NaCH_3COO$

4. HBr and KBr **correct**

5. NH_3 and NH_4Br

Explanation:

A buffer must contain a weak acid/base conjugate pair. HBr/Br^- is a strong acid conjugate pair. CH_3COOH/CH_3COO^-, HCN/CN^- and NH_4^+/NH_3 are weak acid/base conjugate pairs.

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004 10.0 points

Which of the following mixtures will be a buffer when dissolved in a liter of water?

1. 0.1 mol $Ca(OH)_2$ and 0.3 mol HI

2. 0.3 mol NaCl and 0.3 mol HCl

3. 0.4 mol NH_3 and 0.4 mol HCl

4. 0.2 mol HBr and 0.1 mol NaOH

5. 0.2 mol HF and 0.1 mol NaOH **correct**

Explanation:

Eliminate answers that are obviously incorrect. The choice with "0.2 mol HBr" and "0.1 mol $Ca(OH)_2$" are strong acids and strong bases respectively; therefore, NOT buffers. The choice with "0.3 mol NaCl" is a combination of spectator ions and a strong acid; this does not form a buffer. Remaining for calculation are choices with "0.4 mol NH_3" and "0.2 mol HF". Now perform the neutralizaton calculations on the remaining possibilities:

Choice with 0.4 mol NH_3

	NH_3 +	H^+ ⇌	$NH_4{}^+$
Initial	0.4	0.4	0
Change	−0.4	−0.4	0.4
Final	0	0	0.4

Choice with 0.2 mol HF

$$HF + OH^- \rightleftharpoons F^- + H_2O$$

Initial	0.2	0.1	0	–
Change	−0.1	−0.1	0.1	–
Final	0.1	0	0.1	–

The choice with 0.2 mol HF has both weak acid and weak conjugate base left over, so it is the buffer solution.

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005 10.0 points

What is the equilibrium pH of a solution which is initially mixed at 0.200 M in formic acid and 0.00500 M in formate ion? $K_a = 1.8 \times 10^{-4}$ for formic acid.

1. 2.14

2. None of the other answers is correct

3. 11.86

4. 4.35

5. 2.40 **correct**

6. 5.34

Explanation:

$K_a = 1.8 \times 10^{-4}$ $[HA]_{ini} = 0.2$ M

$[A^-]_{ini} = 0.005$ M

You CANNOT use the assumption here! K_a is just a little too big and the 0.005 M concentration is too small. You must solve this fully using the quadratic equation.

	[HA]	$[H^+]$	$[A^-]$
Initial	0.2	0	0.005
Change	$-x$	$+x$	$+x$
Equil.	$0.2 - x$	x	$0.005 + x$

$$K_a = \frac{[H^+][A^-]}{[HA]}$$

$$0.00018 = \frac{(x)(0.005 + x)}{(0.2 - x)}$$

Using the quadratic equation $x = 0.00394514$, which is also the concentration of H^+:

$$pH = -\log(0.00394514) = 2.40394$$

Here's a good question for you. What is the pH of a plain 0.200 M formic acid solution? You should get 2.22, which is higher than the 2.14 choice given above (using the Henderson-Hassleback equation). How can the pH drop when you ADD conjugate BASE to an acid? It can't. Watch out for these borderline cases.

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006 10.0 points

What is the pH of a solution which is 0.600 M in dimethylamine ($(CH_3)_2NH$) and 0.400 M in dimethylamine hydrochloride $\left((CH_3)_2NH_2^+Cl^-\right)$? K_b for dimethylamine $= 7.4 \times 10^{-4}$.

1. 11.05 **correct**

2. 10.69

3. 10.78

4. 11.21

5. 2.95

6. 3.31

7. 10.87

Explanation:

$K_w = 1 \times 10^{-14}$ $K_a = 0.00074$

$[(CH_3)_2NH] = 0.6$ M $[(CH_3)_2NH_2^+] = 0.4$ M

$$K_{a,\ (CH_3)_2NH_2^+} = \frac{K_w}{K_{b,\ (CH_3)_2NH}}$$

Applying the Henderson-Hasselbalch equation,

$$\begin{aligned} pH &= pK_a + \log\left(\frac{[(CH_3)_2NH]}{[(CH_3)_2NH_2^+]}\right) \\ &= -\log\left(\frac{K_w}{K_a}\right) + \log\left(\frac{[(CH_3)_2NH]}{[(CH_3)_2NH_2^+]}\right) \\ &= -\log\left(\frac{1 \times 10^{-14}}{0.00074}\right) + \log\left(\frac{0.6}{0.4}\right) \\ &= 11.0453 \end{aligned}$$

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007 10.0 points

What would be the final pH if 0.0100 moles of solid NaOH were added to 100 mL of a buffer solution containing 0.600 molar formic acid (ionization constant $= 1.8 \times 10^{-4}$) and 0.300 molar sodium formate?

1. 3.44

2. 4.05

3. 3.84

4. None of these

5. 3.65 **correct**

Explanation:

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008 10.0 points

A buffer was prepared by mixing 0.200 mole of ammonia ($K_b = 1.8 \times 10^{-5}$) and 0.200 mole of ammonium chloride to form an aqueous solution with a total volume of 500 mL. To 250 mL of this solution was added 50.0 mL of 1.00 M HCl. What is the pH of this solution?

1. 9.35

2. 8.38

3. 7.87

4. 8.53

5. 8.18

6. 8.78 **correct**

7. 9.73

Explanation:

$[NH_3] = \dfrac{0.2 \text{ mol}}{500 \text{ mL}}$ $\qquad [HCl] = \dfrac{1.0 \text{ mol}}{1000 \text{ mL}}$

$[NH_4^+] = \dfrac{0.2 \text{ mol}}{500 \text{ mL}}$ $\qquad [Cl^-] = \dfrac{0.2 \text{ mol}}{500 \text{ mL}}$

$K_b = 1.8 \times 10^{-5}$

Initial condition (ini):

$$n_{NH_3} = 250 \text{ mL} \times \frac{0.2 \text{ mol}}{500 \text{ mL}} = 100 \text{ mmol}$$

$$n_{HCl} = 50 \text{ mL} \times \frac{1.0 \text{ mol}}{1000 \text{ mL}} = 50 \text{ mmol}$$

$$n_{NH_4^+} = 250 \text{ mL} \times \frac{0.2 \text{ mol}}{500 \text{ mL}} = 100 \text{ mmol}$$

$$n_{Cl^-} = 250 \text{ mL} \times \frac{0.2 \text{ mol}}{500 \text{ mL}} = 100 \text{ mmol}$$

	NH_3 +	HCl →	NH_4^+ +	Cl^-
ini, mmol	100	50	100	100
Δ, mmol	−50	−50	50	50
fin, mmol	50	0	150	150

Cl^- is a spectator ion. NH_4^+/NH_3 is a buffer system.

$$\begin{aligned} pH &= pK_a + \log\left(\frac{[NH_3]}{[NH_4^+]}\right) \\ &= -\log\left(\frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}}\right) + \log\left(\frac{50}{150}\right) \\ &= 8.77815 \end{aligned}$$

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009 10.0 points

A solution is 0.30 M in NH_3. What concentration of NH_4Cl would be required to achieve a buffer solution with a final pH of 9.0? $K_b = 1.8 \times 10^{-5}$ for NH_3.

1. 0.10 M

2. 0.30 M

3. 0.45 M

4. 0.20 M

5. 0.54 M **correct**

Explanation:

$[NH_3] = 0.30$ M $\qquad$ pH = 9

$K_b = 1.8 \times 10^{-5}$ for NH_3

$K_{a,\,NH_4^+} = \dfrac{K_w}{K_{b,\,NH_3}}$, so by the Henderson-Hasselbanch equation,

$$\begin{aligned} pH &= pK_a + \log\left(\frac{[NH_3]}{[NH_4^+]}\right) \\ &= pK_a + \log[NH_3] - \log[NH_4^+] \\ \log[NH_4^+] &= pK_a + \log[NH_3] - pH \end{aligned}$$

$$= -\log\left(\frac{K_\mathrm{w}}{K_\mathrm{a}}\right) + \log[\mathrm{NH_3}] - \mathrm{pH}$$
$$= -\log\left(\frac{1\times10^{-14}}{1.8\times10^{-5}}\right) + \log(0.3) - 9$$
$$= -0.267606$$

$$[\mathrm{NH_4^+}] = [\mathrm{NH_4Cl}] = 10^{-\mathrm{pH}} = 10^{-0.267606} = 0.54$$

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010 10.0 points

What is the pH at the half-stoichiometric point for the titration of 0.22 M HNO_2(aq) with 0.01 M KOH(aq)? For HNO_2, $K_\mathrm{a} = 4.3\times10^{-4}$.

1. 3.37 **correct**

2. 2.01

3. 7.00

4. 2.16

5. 2.31

Explanation:

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011 10.0 points

For the titration of 50.0 mL of 0.020 M aqueous salicylic acid with 0.020 M KOH(aq), calculate the pH after the addition of 55.0 mL of KOH(aq). For salycylic acid, $\mathrm{p}K_\mathrm{a} = 2.97$.

1. 10.98 **correct**

2. 11.26

3. 12.30

4. 7.00

5. 12.02

Explanation:

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012 10.0 points

Consider the titration of 50.0 mL of 0.0200 M HClO(aq) with 0.100 M NaOH(aq). What is the formula of the main species in the solution after the addition of 10.0 mL of base?

1. $HClO_2$

2. ClO_2

3. ClOH

4. NaOH

5. ClO^- **correct**

Explanation:

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013 10.0 points

50.0 mL of 0.0018 M aniline (a weak base) is titrated with 0.0048 M HNO_3. How many mL of the acid are required to reach the equivalence point?

1. 133 mL

2. 18.8 mL **correct**

3. Need to know the K_b of aniline.

4. Bad titration since HNO_3 is not a strong acid.

5. 4.21 mL

Explanation:

$V_\mathrm{aniline} = 50$ mL [Aniline] = 0.0018 M
$[\mathrm{HNO_3}] = 0.0048$ M

Aniline is a monobasic base (*i.e.*, it produces one OH^- in solution). Thus you can expect that aniline and HNO_3 will react in a one-to-one fashion.

With this ratio, we can determine how much HNO_3 will be required to react with all of the aniline.

First, convert 50.0 mL aniline into L of aniline:

$$50.0\text{ mL aniline}\left(\frac{1\text{ L}}{1000\text{ mL}}\right) = 0.0500\text{ L aniline}$$

Then use the ratio to determine the volume of HNO_3 needed:

$$(0.0500 \text{ L aniline})\left(\frac{0.0018 \text{ mol aniline}}{1 \text{ L aniline}}\right) \times \left(\frac{1 \text{ mol HNO}_3}{1 \text{ mol aniline}}\right)\left(\frac{1 \text{ L HNO}_3}{0.0048 \text{ mol HNO}_3}\right)$$
$$= 0.0188 \text{ L HNO}_3\left(\frac{1000 \text{ mL}}{1 \text{ L}}\right)$$
$$= 18.8 \text{ mL HNO}_3$$

014 10.0 points

When we titrate a weak base with a strong acid, the pH at the equivalence point will be

1. pH > 7.

2. pH < 7. **correct**

3. pH = 7.

Explanation:

At the equivalence point, all of the strong acid will be consumed, leaving only its very weak (to the point of being neutral) conjugate base behind. However, the weak base will be converted into its corresponding conjugate acid, which will *not* be neutral. The resulting solution will then be *acidic*, having a pH < 7.

015 10.0 points

What is the pH at the equivalence point in the titration of 10.0 mL of 0.31 M HZ with 0.200 M NaOH? $K_a = 3.4 \times 10^{-7}$ for HZ.

Correct answer: 9.777.

Explanation:

$V_{HZ} = 10$ mL [NaOH] = 0.2 M
[HZ] = 0.31 M $K_a = 3.4 \times 10^{-7}$

$$\text{HZ} + \text{NaOH} \rightarrow \text{Na}^+ + \text{Z}^- + \text{H}_2\text{O}$$

At the equivalence point all the HZ and NaOH have produced NA^+ (spectator ion) and Z^-.

Amount NaOH that reacts
= amount of Z^- produced
= amount of HZ that reacts
= (10 mL)(0.31 M) = 3.1 mmol

Volume of 0.2 M NaOH required

$$= \frac{3.1 \times 10^{-3} \text{ mol NaOH}}{\frac{0.2 \text{ mol NaOH}}{1 \text{ L NaOH(aq)}}} = 15.5 \text{ mL}$$

Total volume solution at endpoint = 25.5 mL

$$[\text{Z}^-] = \frac{3.1 \text{ mmol}}{25.5 \text{ mL}} = 0.121569 \text{ M}$$

Z^- is hydrolyzed:

	Z^-	$+ H_2O \rightleftharpoons$	HZ	+ OH^-
ini, M	0.121569		0	0
Δ, M	$-x$		x	x
eq, M	$0.121569 - x$		x	x

$$K_b = \frac{K_w}{K_{a\text{ HZ}}} = \frac{[\text{HZ}][\text{OH}^-]}{[\text{Z}^-]}$$

$$\frac{1 \times 10^{-14}}{3.4 \times 10^{-7}} = \frac{x^2}{0.121569 - x}$$

Assume $x \ll 0.121569$:

$$\frac{1 \times 10^{-14}}{3.4 \times 10^{-7}} = \frac{x^2}{0.121569}$$

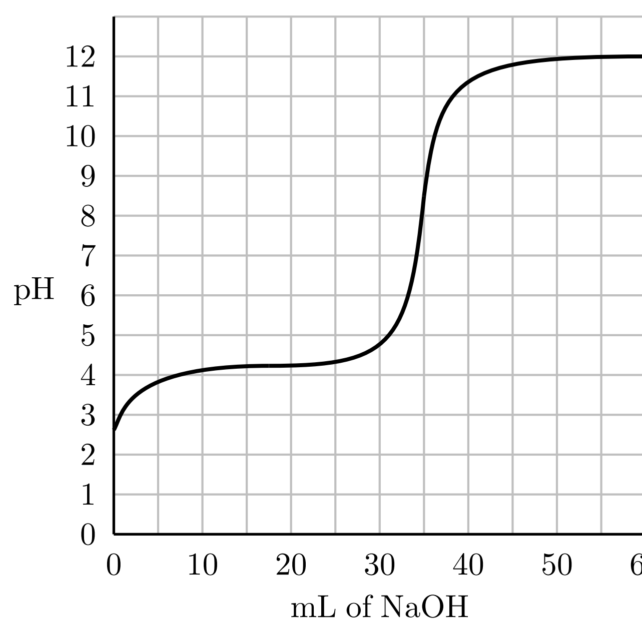
$$x = [\text{OH}^-] = 5.97959 \times 10^{-5} \text{ M}$$
$$(\ll 0.121569)$$

$$\text{pOH} = -\log(5.97959 \times 10^{-5}) = 4.22333$$

$$\text{pH} = 14 - 4.22333 = 9.777$$

016 (part 1 of 2) 10.0 points

Titration Curve



What is the pH at the equivalence point of this titration?

1. 4.23

2. 8.49 **correct**

3. 6.36

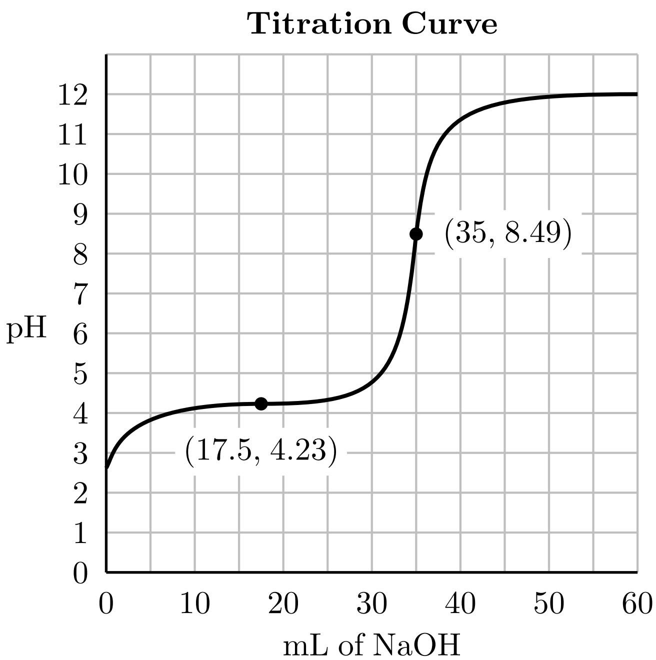
4. 10.25

5. 2.62

6. 5.08

7. 3.43

Explanation:
The inflection points are shown below.



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017 (part 2 of 2) 10.0 points

What is the pK_a of this acid?

1. 6.36

2. 3.43

3. 5.08

4. 10.25

5. 2.62

6. 8.49

7. 4.23 **correct**

Explanation:

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018 10.0 points

The acid form of an indicator is yellow and its anion is blue. The K_a of this indicator is 1×10^{-5}. What will be the approximate pH range over which this indicator changes color?

1. $9 < \text{pH} < 11$

2. $5 < \text{pH} < 7$

3. $4 < \text{pH} < 6$ **correct**

4. $8 < \text{pH} < 10$

5. $3 < \text{pH} < 5$

Explanation:
The pK_a of this indicator is 5, so the indicator will change colors around pH 5. Thus you would expect a color change between pH 4 and pH 6.

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019 10.0 points

The un-ionized form of an acid indicator is yellow and its anion is blue. The K_a of this indicator is 10^{-5}. What will be the color of the indicator in a solution of pH 3?

1. green

2. red

3. blue

4. yellow **correct**

5. orange

Explanation:
$K_a = 10^{-5}$ pH = 3

$$\underset{\text{yellow}}{\text{HIn}} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \underset{\text{blue}}{\text{In}^-}$$

$$K_a = 10^{-5}$$
$$\text{p}K_a = -\log\left(10^{-5}\right) = 5$$

The color change range is pH = pK_a ±1. At pH values above 6 the indicator will be ionized and at pH values below 4 the indicator will be un-ionized.

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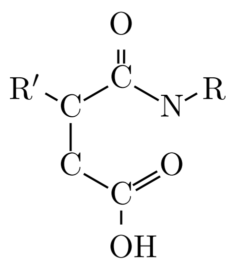
020 10.0 points



This is a structure of an aspartic acid sidechain on a polypeptide. The pK_a of aspartic acid is 3.86. If this polypeptide were in an aqueous solution with a pH of 7, the sidechain would have what charge?

1. negative **correct**

2. neutral

3. positive

4. no way to know

Explanation:

Since the pH is greater than the pKa, the acid group will be deprotonated leaving the sidechain with a charge of minus 1.