

This print-out should have 49 questions. Multiple-choice questions may continue on the next column or page – find all choices before answering.

001 10.0 points

What is the pH at the point during a titration when 30 mL of 0.200 M HCl has been added to 31 mL of 0.0500 M NH₃ solution? HCl is strong acid, $K_b = 1.8 \times 10^{-5}$ for NH₃.

1. 1.25964
2. 1.10914
3. 0.979797
4. 0.938343
5. 1.07542
6. 1.20713
7. 1.24304
8. 1.00352
9. 1.13697
10. 1.33882

Correct answer: 1.13697 pH.

Explanation:

002 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. $3 < \text{pH} < 5$
2. $8 < \text{pH} < 10$
3. $4 < \text{pH} < 6$ **correct**
4. $9 < \text{pH} < 11$
5. $5 < \text{pH} < 7$

Explanation:

The $\text{p}K_a$ 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.

003 10.0 points

50.0 mL of 0.0018 M aniline (a weak base) is titrated with 0.0048 M HNO₃. How many mL

of the acid are required to reach the equivalence point?

1. Need to know the K_b of aniline.
2. 4.21 mL
3. 18.8 mL **correct**
4. Bad titration since HNO₃ is not a strong acid.
5. 133 mL

Explanation:

$$V_{\text{aniline}} = 50 \text{ mL} \quad [\text{Aniline}] = 0.0018 \text{ M}$$

$$[\text{HNO}_3] = 0.0048 \text{ M}$$

Aniline is a monobasic base (*i.e.*, it produces one OH⁻ in solution). Thus you can expect that aniline and HNO₃ will react in a one-to-one fashion.

With this ratio, we can determine how much HNO₃ 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 HNO₃ needed:

$$(0.0500 \text{ L aniline}) \left(\frac{0.0018 \text{ mol aniline}}{1 \text{ L aniline}} \right)$$

$$\times \left(\frac{1 \text{ mol HNO}_3}{1 \text{ mol aniline}} \right) \left(\frac{1 \text{ L HNO}_3}{0.0048 \text{ mol HNO}_3} \right)$$

$$= 0.0188 \text{ L HNO}_3 \left(\frac{1000 \text{ mL}}{1 \text{ L}} \right)$$

$$= 18.8 \text{ mL HNO}_3$$

004 10.0 points

A buffer was prepared by mixing 0.50 mole of HX acid and 0.50 mole of NaX to form an aqueous solution with a total volume of 1.00 liter. The pH of this buffer was 5.525. Then, to 400 mL of this buffer solution was added 25.0 mL of 2.0 M HCl. What is the pH of this new solution?

1. 4.00315
2. 4.70315
3. 5.30315
4. 5.20315
5. 3.80315

6. 5.00315
7. 4.90315
8. 4.80315
9. 4.40315
10. 4.20315

Correct answer: 5.30315.

Explanation:

$$\begin{array}{ll} n_{\text{HX}} = 0.50 \text{ mol} & n_{\text{NaX}} = 0.50 \text{ mol} \\ V_{\text{soln}} = 1.0 \text{ L} & pK_a = 5.525 \\ V_{\text{HCL}} = 25.0 \text{ mL} & [\text{HCl}] = 2.0 \text{ M} \end{array}$$

Notice that the pK_a of this acid (HX) must be 5.525 by definition. The HCl converts X^- into HX. 25 mL of 2.0 M is equal to 50 mmol of H^+ added. Take this amount AWAY from the X^- and add it TO the HX:

200 mmol initial of HX + 50 = 250 mmol HX

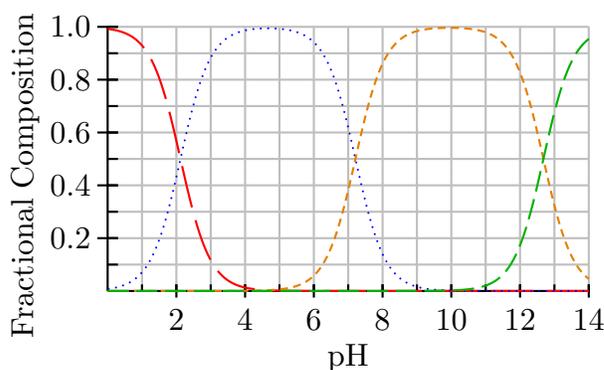
200 mmol initial of X^- (-50) = 150 mmol X^-

Now substitute into the Henderson-Hasselbalch equation:

$$\begin{aligned} \text{pH} &= 5.525 + \log\left(\frac{150}{250}\right) \\ &= 5.30315 \end{aligned}$$

005 10.0 points

Consider the fraction of species diagram for phosphoric acid.



What is the dominant species at $\text{pH} = 0.83$?

1. H_2PO_4^-
2. PO_4^{3-}
3. HPO_4^{2-}
4. None of these is correct

5. H_3PO_4 correct

Explanation:

The dominant species depends on the pH, where the highest pH is the most deprotonated species and the lowest pH is the most protonated species.

006 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. 4.35
2. None of the other answers is correct
3. 2.14
4. 5.34
5. 11.86
6. 2.40 correct

Explanation:

$$K_a = 1.8 \times 10^{-4} \quad [\text{HA}]_{\text{ini}} = 0.2 \text{ M} \\ [\text{A}^-]_{\text{ini}} = 0.005 \text{ M}$$

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	$0.005 + x$

$$K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]} \\ 0.00018 = \frac{(x)(0.005 + x)}{(0.2 - x)}$$

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

$$\text{pH} = -\log(0.00394514) = 2.40394$$

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-Hasselback equation). How can the pH drop when you ADD conjugate BASE to an acid? It can't. Watch out for these borderline cases.

007 10.0 points

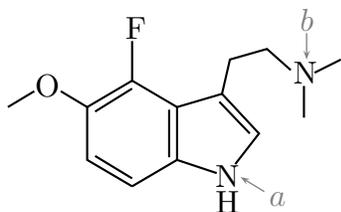
Calculate the ratio of the molarities of HPO_4^{2-} and H_2PO_4^- ions required to achieve buffering at $\text{pH} = 7.00$. For H_3PO_4 , $\text{p}K_{a1} = 2.12$, $\text{p}K_{a2} = 7.21$, and $\text{p}K_{a3} = 12.68$.

- 1.23
- 0.81
- 1.62
- 0.21
- 0.62 correct

Explanation:

008 10.0 points

Consider a large organic molecule seen on Exam 3 of CH301. This molecule has two functional groups that can participate in acid base chemistry. The amine group marked with arrow *a* has a $\text{p}K_a = 6.5$. The amine group marked with arrow *b* has a $\text{p}K_a = 10.6$.



As shown above, the molecular formula for this molecule is $\text{C}_{13}\text{H}_{17}\text{FN}_2\text{O}$. What is the molecular formula at $\text{pH} = 3.4$?

- $\text{C}_{13}\text{H}_{17}\text{FN}_2\text{O}$
- $[\text{C}_{13}\text{H}_{16}\text{FN}_2\text{O}]^-$
- $[\text{C}_{13}\text{H}_{15}\text{FN}_2\text{O}]^{2-}$



Explanation:

An organic functional group will be deprotonated when the $\text{pH} > \text{p}K_a$ ("more basic/proton-accepting environment") and protonated when the $\text{pH} < \text{p}K_a$ ("more acidic/proton-donating environment"). For an amine group, this means the functional group will be positively charged when the $\text{pH} < \text{p}K_a$. At $\text{pH} = 7.6$, amine group *a* will be neutral (deprotonated) and amine group *b* will be positive (protonated). This results in an overall formula of $[\text{C}_{13}\text{H}_{18}\text{FN}_2\text{O}]^+$. At $\text{pH} = 3.4$, both amine groups will be protonated, resulting in an overall formula of $[\text{C}_{13}\text{H}_{19}\text{FN}_2\text{O}]^{2+}$.

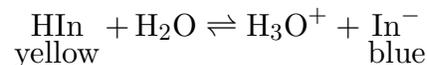
009 10.0 points

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 $\text{pH} 3$?

- green
- blue
- yellow correct
- red
- orange

Explanation:

$$K_a = 10^{-5} \qquad \text{pH} = 3$$



$$K_a = 10^{-5}$$

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

The color change range is $\text{pH} = \text{p}K_a \pm 1$. At pH values above 6 the indicator will be ionized and at pH values below 4 the indicator will be un-ionized.

010 10.0 points

Mixing the contents of which two solutions will result in an ideal buffer at pH = 4.63?

1. 50 mL 0.3 M HCl and 150 mL 0.1 M C₆H₅NH₂

2. 40 mL 0.2 M HCl and 160 mL 0.1 M C₆H₅NH₂ **correct**

3. 180 mL 0.3 M HCl and 60 mL 0.1 M C₅H₅N

4. 60 mL 0.3 M HCl and 180 mL 0.1 M C₅H₅N

5. 30 mL 0.3 M HCl and 180 mL 0.1 M C₅H₅N

Explanation:

Only two of these choices will actually give buffer solutions:

30 mL 0.3 M HCl and 180 mL 0.1 M C₅H₅N

40 mL 0.2 M HCl and 160 mL 0.1 M C₆H₅NH₂

While the buffer zone overlaps for these different solutions, the buffer capacity is higher at pH values closer to their pH. The C₅H₅N solution will have a better buffer in a pH range closer to its conjugate pK_a (5.26), while the C₆H₅NH₂ solution will be a better buffer closer to its conjugate pK_a (4.63).

011 10.0 points

A solution is 0.30 M in NH₃. What concentration of NH₄Cl would be required to achieve a buffer solution with a final pH of 9.0? K_b = 1.8 × 10⁻⁵ for NH₃.

1. 0.45 M

2. 0.54 M **correct**

3. 0.10 M

4. 0.30 M

5. 0.20 M

Explanation:

[NH₃] = 0.30 M

pH = 9

K_b = 1.8 × 10⁻⁵ for NH₃

K_{a, NH₄⁺} = $\frac{K_w}{K_{b, NH_3}}$, so by the Henderson-Hasselbanch equation,

$$\begin{aligned} \text{pH} &= \text{p}K_a + \log\left(\frac{[\text{NH}_3]}{[\text{NH}_4^+]}\right) \\ &= \text{p}K_a + \log[\text{NH}_3] - \log[\text{NH}_4^+] \\ \log[\text{NH}_4^+] &= \text{p}K_a + \log[\text{NH}_3] - \text{pH} \\ &= -\log\left(\frac{K_w}{K_a}\right) + \log[\text{NH}_3] - \text{pH} \\ &= -\log\left(\frac{1 \times 10^{-14}}{1.8 \times 10^{-5}}\right) \\ &\quad + \log(0.3) - 9 \\ &= -0.267606 \end{aligned}$$

$$\begin{aligned} [\text{NH}_4^+] &= [\text{NH}_4\text{Cl}] = 10^{-\text{pH}} \\ &= 10^{-0.267606} \\ &= 0.54 \end{aligned}$$

012 10.0 points

A 100 mL portion of 0.300 M acetic acid is being titrated with 0.200 M NaOH solution. What is the [H⁺] of the solution after 50.0 mL of the NaOH solution has been added? The ionization constant of acetic acid is 1.8 × 10⁻⁵.

1. 9.94 × 10⁻⁶

2. 8.95 × 10⁻⁶

3. 6.01 × 10⁻⁴

4. 3.63 × 10⁻⁵ **correct**

5. 1.21 × 10⁻⁵

Explanation:

V_{CH₃COOH} = 100 mL V_{NaOH} = 50 mL
 [CH₃COOH] = 0.300 M [NaOH] = 0.200 M
 K_a = 1.8 × 10⁻⁵

Initially,

n_{CH₃COOH} = (100 mL)(0.3 M) = 30 mmol

n_{NaOH} = (50 mL)(0.2 M) = 10 mmol

$$\text{CH}_3\text{COOH} + \text{NaOH} \rightarrow \text{CH}_3\text{COO}^- + \text{Na}^+ + \text{H}_2\text{O}$$

ini	30	10	0	0
Δ	-10	-10	10	10
fin	20	0	10	10

Na^+ is a spectator ion. CH_3COOH and CH_3COO^- form a buffer.

Total volume = 150 mL

$$\begin{aligned} \text{pH} &= \text{p}K_a + \log \left(\frac{[\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]} \right) \\ &= -\log(1.8 \times 10^{-5}) \\ &\quad + \log \left(\frac{10 \text{ mmol}/150 \text{ mL}}{20 \text{ mmol}/150 \text{ mL}} \right) \\ &= 4.4437 \\ [\text{H}_3\text{O}^+] &= 10^{-4.4437} = 3.6 \times 10^{-5} \text{ M} \end{aligned}$$

013 10.0 points

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

1. pH < 7. **correct**
2. pH = 7.
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.

014 10.0 points

Explain why the salt of a weak acid, as well as the acid itself, must be present to form a buffer solution.

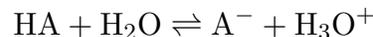
1. The anion from the salt is needed to partially neutralize added base.
2. The anion from the salt is needed to partially neutralize added acid. **correct**
3. The cation from the salt is needed to partially neutralize added acid.

4. Actually, a weak acid by itself is a buffer; no salt is needed.

5. The cation from the salt is needed to partially neutralize added base.

Explanation:

The salt of the acid provides the anion which is the conjugate base of the buffer system:



This anion A^- reacts with any added acid (H_3O^+) to prevent any appreciable change in pH.

015 10.0 points

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

1. NH_3 and $(\text{NH}_4)_2\text{SO}_4$
2. HBr and KBr **correct**
3. NH_3 and NH_4Br
4. CH_3COOH and NaCH_3COO
5. HCN and NaCN

Explanation:

A buffer must contain a weak acid/base conjugate pair. HBr/ Br^- is a strong acid conjugate pair. $\text{CH}_3\text{COOH}/\text{CH}_3\text{COO}^-$, HCN/ CN^- and $\text{NH}_4^+/\text{NH}_3$ are weak acid/base conjugate pairs.

016 10.0 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. 10.87
2. 10.78
3. 2.95

4. 11.05 correct

5. 10.69

6. 3.31

7. 11.21

Explanation:

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

$$[(\text{CH}_3)_2\text{NH}] = 0.6 \text{ M} \quad [(\text{CH}_3)_2\text{NH}_2^+] = 0.4 \text{ M}$$

$$K_{a, (\text{CH}_3)_2\text{NH}_2^+} = \frac{K_w}{K_{b, (\text{CH}_3)_2\text{NH}}}$$

Applying the Henderson-Hasselbalch equation,

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

017 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 840 mL of solution. What is the pH of this buffer? $K_a = 5 \times 10^{-6}$ for HA.

- 4.73956
- 4.94604
- 5.06473
- 4.96534
- 6.16058
- 5.46267
- 5.26343
- 5.32034
- 5.51505
- 5.69897

Correct answer: 5.32034 pH.

Explanation:

$$n_{\text{HA}} = 0.45 \text{ mol} \quad n_{\text{KOH}} = 0.23 \text{ mol}$$

$$K_a = 5 \times 10^{-6} \text{ for HA}$$

You must subtract 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.

018 10.0 points

You have a solution that is buffered at pH = 2.0 using H_3PO_4 and H_2PO_4^- ($\text{p}K_{a1} = 2.12$; $\text{p}K_{a2} = 7.21$; $\text{p}K_{a3} = 12.68$). You decide to titrate this buffer with a strong base. 15.0 mL are needed to reach the first equivalence point. What is the total volume of base that will have been added when the second equivalence point is reached?

1. > 30 mL correct

2. A second equivalence point in the titration will never be observed.

3. < 30 mL

4. 30 mL

Explanation:**019 10.0 points**

Which of the following indicators would be most suitable for the titration of 0.10 M lactic acid with 0.10 M KOH(aq)? For lactic acid, $\text{p}K_a = 3.08$.

1. phenol red ($\text{p}K_{\text{In}} = 7.9$) correct2. methyl orange ($\text{p}K_{\text{In}} = 3.4$)3. alizarin yellow ($\text{p}K_{\text{In}} = 11.2$)4. bromophenol blue ($\text{p}K_{\text{In}} = 3.9$)5. thymol blue ($\text{p}K_{\text{In}} = 1.7$)**Explanation:****020 10.0 points**

The equivalence point for the neutralization of 10^{-3} M NH_3 by HCl occurs at pH = 6.1. The equivalence point is accurately detected using the indicator DYNAMO RED when the

solution turns from clear to red. Which of the following statements about DYNAMO RED is FALSE?

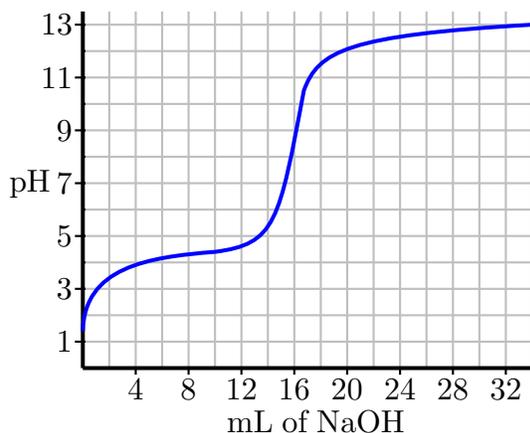
1. The concentration of of DYNAMO RED in the titrated solution must be much less than 10^{-3} M.
2. The acid form of DYNAMO RED is colored.
3. DYNAMO RED could be a diprotic acid.
4. The pK_a of DYNAMO RED must be much less than 4.0. **correct**
5. DYNAMO RED is an acid or a base in its nonionized state.

Explanation:

Acid-base indicators are weak acid-base systems and must be present in the solution in concentrations considerably less than the acid and base reactants. In this case, before HCl neutralizes NH_3 the indicator is in its colorless basic form. After the equivalence point the indicator is in its red acidic form. The color change occurs in the range $pH = pK_a \pm 1$; i.e., 5.1 to 7.1.

021 (part 1 of 3) 10.0 points

A sample of 100 mL of a weak acid (HA) solution was titrated with 0.040 M NaOH. The pH curve for this titration is shown.



What is the concentration of the original weak acid solution (the 100 mL) ?

1. 0.016 M
2. 0.0013 M
3. 0.0126 M
4. 0.0064 M **correct**
5. 0.0032 M

Explanation:

The equivalence point is 16 mL of NaOH
 $\times 0.040 \text{ M} = 0.00064 \text{ mol OH}^-$

mol acid = mol base

$$0.00064 \text{ mol} / 0.100 \text{ L} = 0.0064 \text{ M HA}$$

022 (part 2 of 3) 10.0 points

Which of these acids and bases could be the solution that was titrated?

1. Acrylic acid **correct**
2. Acetic acid
3. Hypoiodous acid
4. Nitric acid
5. Hydrofluoric acid
6. Trimethylamine

Explanation:

The half equivalence point is at 8 mL and the $pH = pK_a$ there which is about 4.25. The only acid that matches this value is acrylic acid.

023 (part 3 of 3) 10.0 points

Which of the following chemicals would be a suitable indicator for this titration experiment?

1. Methyl orange
2. Bromocresol green
3. Methyl red

4. Phenolphthalein **correct**

5. Bromothymol blue

Explanation:

Phenolphthalein is the best choice for an indicator which has a color change pH range that includes the end point pH around 8.5.

024 10.0 points

Lithium acetate is a common component of the buffer system used in gel electrophoresis of DNA and RNA. How many moles of lithium hydroxide do you need to add to a 650 mL 0.590 M CH_3COOH solution to make a buffer with a pH = 5.000?

- 0.180 mol
- 0.590 mol
- 0.384 mol
- 0.214 mol
- 0.151 mol
- 3.97 mol
- 0.247 mol **correct**
- 0.192 mol

Explanation:

You will be using the Henderson-Hasselbach equation to solve for the ratio of A^- : HA in moles.

$$\text{pH} = \text{p}K_a + \log\left(\frac{\text{mol}_{\text{A}^-}}{\text{mol}_{\text{HA}}}\right)$$

You are looking for this ratio when the pH = 5.00:

$$10^{5.00-4.74} = \frac{\text{mol}_{\text{A}^-}}{\text{mol}_{\text{HA}}}$$

$$1.82 = \frac{x}{\text{mol}_{\text{HA}} - x}$$

Solve for moles of HA to plug into this ratio:

$$0.384 \text{ mol HA} = 0.650 \text{ L} \times 0.590 \text{ M}$$

Then solve for x, which is equal to the moles

of LiOH:

$$0.2475 \text{ mol LiOH} = x$$

025 10.0 points

What molar ratio of sodium acetate to acetic acid (NaAc/HAc) should be used in preparing a buffer having a pH of 4.35? ($K_a = 1.8 \times 10^{-5}$ for acetic acid.)

- 0.40 : 1.0 **correct**
- 1.0 : 1.0
- 1.0 : 0.40
- 0.59 : 1.0
- 2.0 : 1.0

Explanation:

$$\text{pH} = 4.35 \qquad K_a = 1.8 \times 10^{-5}$$

$$\text{pH} = \text{p}K_a + \log\left(\frac{[\text{NaAc}]}{[\text{HAc}]}\right)$$

$$\begin{aligned} \log\left(\frac{[\text{NaAc}]}{[\text{HAc}]}\right) &= \text{pH} - \text{p}K_a \\ &= 4.35 - [-\log(1.8 \times 10^{-5})] \\ &= -0.394727 \end{aligned}$$

$$\frac{[\text{NaAc}]}{[\text{HAc}]} = 10^{-0.394727} = 0.40297$$

026 10.0 points

A buffer (pH 4.74) was prepared by mixing 1.00 mole of acetic acid and 1.00 mole of sodium acetate to form a 1.0 liter aqueous solution. To 100 mL of this solution, 10.0 mL of 2.00 molar NaOH was added. What is the new pH?

- 5.20
- 5.31
- 4.47
- 4.92 **correct**
- 4.72

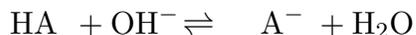
Explanation:

$$K_a = 1.8 \times 10^{-5} \quad \text{pH} = 4.74$$

Acetic acid is CH_3COOH and acetate is CH_3COO^- (Na^+ is a spectator so we just leave it out). OR for simplicity, we can just write acetic acid as HA and acetate as A^- .

100 mL of the buffer solution will contain 0.1 moles of HA and 0.1 moles of A^- .

10.0 mL of 2.0M NaOH contains 0.02 moles OH^- (again, Na^+ is just a spectator).



Initial	0.1	---	0.01	---
Change	0.1-0.02	-0.02	0.1+0.02	---
Final	0.08		0.12	

Use the Weak Acid Buffer Equation:

$$\begin{aligned} [\text{H}^+] &= K_a \left(\frac{[\text{HA}]}{[\text{A}^-]} \right) \\ &= 1.8 \times 10^{-5} \times \left(\frac{0.08}{0.12} \right) \\ &= 1.2 \times 10^{-5} \text{ M} \\ \text{pH} &= -\log(1.2 \times 10^{-5}) = 4.92 \end{aligned}$$

027 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. 8.38
2. 9.35
3. 7.87
4. 8.78 correct
5. 9.73
6. 8.53
7. 8.18

Explanation:

$$[\text{NH}_3] = \frac{0.2 \text{ mol}}{500 \text{ mL}} \quad [\text{HCl}] = \frac{1.0 \text{ mol}}{1000 \text{ mL}}$$

$$[\text{NH}_4^+] = \frac{0.2 \text{ mol}}{500 \text{ mL}} \quad [\text{Cl}^-] = \frac{0.2 \text{ mol}}{500 \text{ mL}}$$

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

Initial condition (ini):

$$n_{\text{NH}_3} = 250 \text{ mL} \times \frac{0.2 \text{ mol}}{500 \text{ mL}} = 100 \text{ mmol}$$

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

$$n_{\text{NH}_4^+} = 250 \text{ mL} \times \frac{0.2 \text{ mol}}{500 \text{ mL}} = 100 \text{ mmol}$$

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

	NH_3	$+$	HCl	\rightarrow	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. $\text{NH}_4^+/\text{NH}_3$ is a buffer system.

$$\begin{aligned} \text{pH} &= \text{p}K_a + \log \left(\frac{[\text{NH}_3]}{[\text{NH}_4^+]}\right) \\ &= -\log \left(\frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} \right) + \log \left(\frac{50}{150} \right) \\ &= 8.77815 \end{aligned}$$

028 10.0 points

The diprotic acid, H_2A has two $\text{p}K_a$ values equal to 3.90 and 7.10. You mix 0.3 moles NaHA into 334 mL pure water. What is the pH of this solution?

1. 7.10
2. 5.50 correct
3. 3.57
4. 4.85
5. 1.97
6. 3.90
7. 5.85

8. 4.05

Explanation:

Create a fraction of species diagram to confirm that the pH will be equal to the average of the two pK_a values.

$$\text{pH} = (3.90 + 7.10) / 2$$

$$\text{pH} = 5.50$$

029 10.0 points

A solution of an unknown weak acid (HA) has a ratio $\frac{[A^-]}{[HA]} = \frac{1}{2000}$ at a pH of 3. What would be this ratio at a pH of 5?

1. 2000

2. $\frac{10^{-3}}{2}$ 3. $\frac{1}{20}$ correct4. $\frac{20}{1}$ 5. $\frac{3}{5}$ **Explanation:**

$$\text{pH}_{\text{ini}} = 3.00 \qquad \text{pH}_{\text{fin}} = 5.00$$

$$\frac{[A^-]}{[HA]} = \frac{1}{2000}$$

$$\text{pH}_{\text{ini}} = \text{p}K_a + \log \left(\frac{[A^-]}{[HA]} \right)$$

$$\text{p}K_a = \text{pH}_{\text{ini}} - \log \left(\frac{[A^-]}{[HA]} \right)$$

$$= 3 - \log \left(\frac{1}{2000} \right)$$

$$\log \left(\frac{[A^-]}{[HA]} \right) = \text{pH}_{\text{fin}} - \text{p}K_a$$

$$= 5 - \left[3 - \log \left(\frac{1}{2000} \right) \right]$$

$$= -1.30103$$

$$\frac{[A^-]}{[HA]} = 10^{-1.30103} = 0.05 = \frac{1}{20}$$

030 10.0 points

To simulate blood conditions, a phosphate buffer system with a pH = 7.4 is desired. What mass of Na_2HPO_4 must be added to 1 L of 0.15 M $\text{NaH}_2\text{PO}_4(\text{aq})$ to prepare such a buffer?

1. 28.5829

2. 395.763

3. 122.027

4. 311.663

5. 62.6624

6. 151.709

7. 32.9802

8. 137.418

9. 270.438

10. 211.073

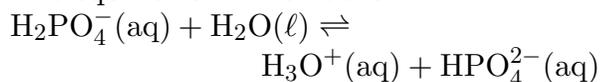
Correct answer: 32.9802 g.

Explanation:

$$[\text{NaH}_2\text{PO}_4] = 0.15 \text{ mol/L}$$

$$\text{MW}_{\text{Na}_2\text{HPO}_4} = 141.959 \text{ g/mol}$$

The equilibrium involved is



$$K_{a2} = \frac{[\text{H}_3\text{O}^+][\text{HPO}_4^{2-}]}{[\text{H}_2\text{PO}_4^-]}$$

Using Henderson-Hasselbalch equation,

$$\text{pH} = \text{p}K_a + \log \left(\frac{[\text{HPO}_4^{2-}]}{[\text{H}_2\text{PO}_4^-]} \right)$$

$$7.4 = 7.21 + \log \left(\frac{[\text{HPO}_4^{2-}]}{[\text{H}_2\text{PO}_4^-]} \right)$$

$$\log \left(\frac{[\text{HPO}_4^{2-}]}{[\text{H}_2\text{PO}_4^-]} \right) = 0.19$$

$$\frac{[\text{HPO}_4^{2-}]}{[\text{H}_2\text{PO}_4^-]} = 10^{0.19} \\ = 1.54882$$

$$[\text{Na}_2\text{HPO}_4] = \frac{[\text{HPO}_4^{2-}]}{[\text{H}_2\text{PO}_4^-]} [\text{NaH}_2\text{PO}_4] \\ = (1.54882)(0.15 \text{ mol/L}) \\ = 0.232322 \text{ mol/L}$$

$$\text{mol} = 0.232322 \text{ mol/L (1 L)} = 0.232322 \text{ mol}$$

$$\begin{aligned} m_{\text{Na}_2\text{HPO}_4} &= 0.232322 \text{ mol Na}_2\text{HPO}_4 \\ &\times \frac{141.959 \text{ Na}_2\text{HPO}_4}{1 \text{ mol Na}_2\text{HPO}_4} \\ &= 32.9802 \text{ g Na}_2\text{HPO}_4 \end{aligned}$$

031 10.0 points

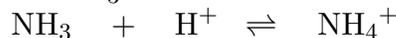
Which of the following mixtures will be a buffer when dissolved in a liter of water?

- 0.3 mol NaCl and 0.3 mol HCl
- 0.2 mol HBr and 0.1 mol NaOH
- 0.2 mol HF and 0.1 mol NaOH **correct**
- 0.4 mol NH₃ and 0.4 mol HCl
- 0.1 mol Ca(OH)₂ and 0.3 mol HI

Explanation:

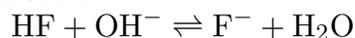
Eliminate answers that are obviously incorrect. The choice with “0.2 mol HBr” and “0.1 mol Ca(OH)₂” are strong acids and strong bases respectively; therefore, NOT buffers. The choice with “0.3 mol NaCl” is a combination of spectator ions and a strong acid; this does not form a buffer. Remaining for calculation are choices with “0.4 mol NH₃” and “0.2 mol HF”. Now perform the neutralization calculations on the remaining possibilities:

Choice with 0.4 mol NH₃



Initial	0.4	0.4	0
Change	-0.4	-0.4	0.4
Final	0	0	0.4

Choice with 0.2 mol HF



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.

032 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 salicylic acid, $pK_a = 2.97$.

- 11.26
- 10.98 **correct**
- 12.02
- 12.30
- 7.00

Explanation:**033 10.0 points**

Which solution is most basic?

- 0.10 M HClO and 0.10 M NaClO
- 0.10 M CH₃COOH and 0.20 M NaCH₃COO
- 0.10 M aq NH₃ and 0.10 M NH₄Cl **correct**
- 0.10 M CH₃COOH and 0.10 M NaCH₃COO
- 0.10 M aq NH₃ and 0.20 M NH₄Cl

Explanation:

All are buffers where

$$\text{pH} = \text{p}K_a + \log \left(\frac{[\text{base}]}{[\text{acid}]} \right)$$

034 10.0 points

A few drops of methyl red are added to an unknown solution. Which of the following represents the best conclusion that can be made if the unknown solution turns yellow?

- The solution has a pH greater than 6.0 **correct**

2. The solution has a pH less than 4.8
3. The solution has a pH value approximately between 4.8 and 6.0
4. The solution is extremely acidic
5. The solution is extremely basic

Explanation:

The solution will be red when the pH is less than 4.8. The solution will be orange between $\text{pH} = 4.8$ and $\text{pH} = 6.0$. The solution will be yellow when the pH is greater than 6.0.

035 10.0 points

What is the pH of an aqueous solution that is 0.018 M $\text{C}_6\text{H}_5\text{NH}_2$ ($K_b = 4.3 \times 10^{-10}$) and 0.12 M $\text{C}_6\text{H}_5\text{NH}_3\text{Cl}$?

1. 10.19
2. 9.37
3. 4.02
4. 4.63
5. 8.54
6. 2.87
7. 5.46
8. 3.81 correct

Explanation:

036 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 = 1.8×10^{-4}) and 0.300 molar sodium formate?

1. 3.65 correct
2. None of these

3. 4.05

4. 3.84

5. 3.44

Explanation:

037 10.0 points

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

1. ClO_2
2. ClOH
3. HClO_2
4. ClO^- correct
5. NaOH

Explanation:

038 (part 1 of 2) 10.0 points

For the next two questions, choose if the statement is true or false.

An acceptable weak acid buffer has a minimum 1:10 ratio of $\text{A}^-:\text{HA}$.

1. True correct
2. False

Explanation:

This statement is true. A buffer is acceptable between +/- 1 pH unit, corresponding to a 1:10 ratio minimum.

039 (part 2 of 2) 10.0 points

The end point of a titration curve has a similar pH value to the equivalence point, but not necessarily equal.

1. True correct
2. False

Explanation:

The end point is a qualitative assesment of the equivalence point being reached in a titration. Therefore, it does not have to have an identical pH.

040 10.0 points

How much NH_4Cl must be added to 2.0 liters of 0.200 M aqueous ammonia to give a solution with $\text{pH} = 8.2$? Assume no volume change due to the addition of NH_4Cl . K_b for NH_3 is 1.8×10^{-5} .

1. 146 g
2. 225 g
3. 243 g **correct**
4. 166 g
5. 123 g

Explanation:

$\text{pH} = 8.2$

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

Applying the Henderson-Hasselbalch equation,

$$K_a \text{NH}_4^+ = \frac{K_w}{K_b \text{NH}_3}$$

$$\begin{aligned} \text{pH} &= \text{p}K_a + \log\left(\frac{[\text{NH}_3]}{[\text{NH}_4^+]}\right) \\ &= \text{p}K_a + \log[\text{NH}_3] - \log[\text{NH}_4^+] \end{aligned}$$

$$\begin{aligned} \log[\text{NH}_4^+] &= \text{p}K_a + \log[\text{NH}_3] - \text{pH} \\ &= -\log\left(\frac{K_w}{K_b}\right) + \log[\text{NH}_3] - \text{pH} \\ &= -\log\left(\frac{1 \times 10^{-14}}{1.8 \times 10^{-5}}\right) \\ &\quad + \log(0.2) - 8.2 \\ &= 0.356303 \end{aligned}$$

$$[\text{NH}_4^+] = 10^{0.356303} = 2.27145 \text{ mol/L}$$

Add to 2 L soln 4.54 mol NH_4^+ ; i.e., 4.54 mol NH_4Cl (53.3 g/mol):

$$(2 \text{ L soln}) \cdot \frac{2.27145 \text{ mol NH}_4^+}{\text{L soln}} \cdot \frac{1 \text{ mol NH}_4\text{Cl}}{1 \text{ mol NH}_4^+} \cdot \frac{53.3 \text{ g NH}_4\text{Cl}}{1.0 \text{ mol NH}_4\text{Cl}} = 242.136 \text{ g NH}_4\text{Cl}$$

041 10.0 points

Calculate the ratio of $\frac{[\text{NH}_3]}{[\text{NH}_4^+]}$ that gives a solution of $\text{pH} = 10.60$.

1. $\frac{[\text{NH}_3]}{[\text{NH}_4^+]} = 22$ **correct**
2. $\frac{[\text{NH}_3]}{[\text{NH}_4^+]} = 0.045$
3. $\frac{[\text{NH}_3]}{[\text{NH}_4^+]} = 5.6 \times 10^{-6}$
4. $\frac{[\text{NH}_3]}{[\text{NH}_4^+]} = 5.6 \times 10^{-5}$
5. $\frac{[\text{NH}_3]}{[\text{NH}_4^+]} = 1.8 \times 10^{-5}$

Explanation:

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

$\text{pH} = 10.6$

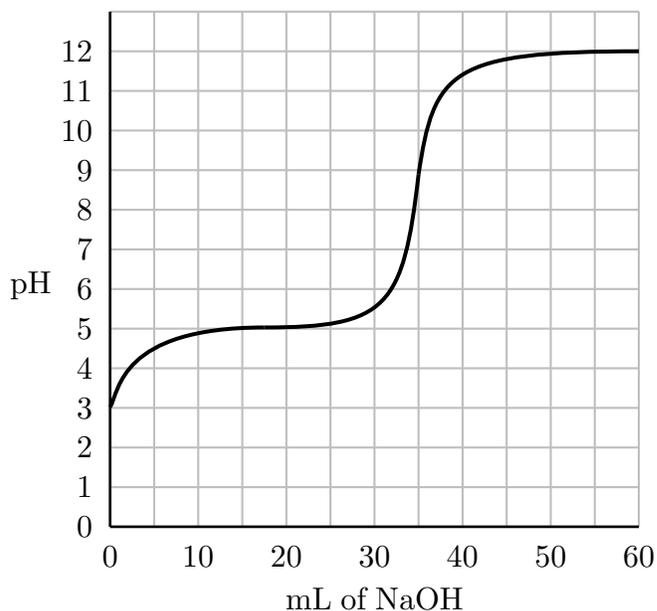
$$K_a[\text{NH}_4^+] = \frac{K_w}{K_b[\text{NH}_3]}, \text{ so}$$

$$\text{pH} = \text{p}K_a + \log\left(\frac{[\text{NH}_3]}{[\text{NH}_4^+]}\right)$$

$$\begin{aligned} \log\left(\frac{[\text{NH}_3]}{[\text{NH}_4^+]}\right) &= \text{pH} - \text{p}K_a \\ &= 10.6 - \left[-\log\left(\frac{1 \times 10^{-14}}{1.8 \times 10^{-5}}\right)\right] \\ &= 1.34473 \\ \frac{[\text{NH}_3]}{[\text{NH}_4^+]} &= 10^{1.34473} = 22.1171 \end{aligned}$$

042 (part 1 of 2) 10.0 points

Titration Curve



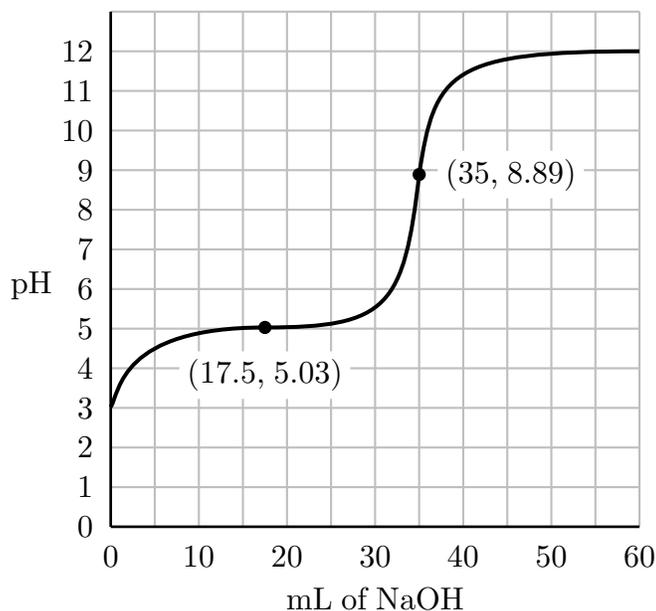
What is the pH at the equivalence point of this titration?

1. 6.96
2. 10.45
3. 8.89 correct
4. 3.02
5. 5.03
6. 4.03
7. 5.8

Explanation:

The inflection points are shown below.

Titration Curve



043 (part 2 of 2) 10.0 points
What is the pK_a of this acid?

1. 3.02
2. 5.8
3. 8.89
4. 5.03 correct
5. 10.45
6. 6.96
7. 4.03

Explanation:

044 10.0 points

A buffer is prepared by combining 300 mL 0.850 M butanoic acid (C_3H_7COOH) to 300 mL 0.640 M sodium hydroxide (NaOH). What is the pH of this solution?

1. 4.82
2. 5.22
3. 5.16

4. 4.34

5. 5.05

6. 4.70

7. 9.18

8. 5.30 correct

Explanation:

This is a simple Henderson-Hasselbach calculation.

$$\text{pH} = \text{p}K_a + \log \left(\frac{C_{A^-}}{C_{\text{HA}}} \right)$$

The $\text{p}K_a$ for butanoic acid is 4.82. The ratio of $C_{A^-}:C_{\text{HA}}$ must account for the dilution when making the solution.

$$\text{Moles of HA} = 0.850 \text{ M} \times 0.300 \text{ L} = 0.255$$

$$\text{Moles of NaOH} = 0.640 \times 0.300 \text{ L} = 0.192$$

In the partial neutralization reaction, NaOH is the limiting reagent. Solve for the moles of butanoic acid remaining:

$$\text{Moles HA} = 0.255 - 0.192 = 0.0630 \text{ mol}$$

The moles of A^- formed is equal to the amount of NaOH that reacted:

$$\text{Moles } A^- = 0.192 \text{ mol}$$

Because you are taking a ratio of molarity in the same volume, you can stick to moles for the Henderson-Hasselbach calculation:

$$\text{pH} = \text{p}K_a + \log \left(\frac{0.192}{0.0630} \right)$$

$$\text{pH} = 5.30$$

045 10.0 points

What is the pH at the half-stoichiometric point for the titration of 0.22 M $\text{HNO}_2(\text{aq})$ with 0.01 M $\text{KOH}(\text{aq})$? For HNO_2 , $K_a = 4.3 \times 10^{-4}$.

1. 7.00

2. 3.37 correct

3. 2.01

4. 2.31

5. 2.16

Explanation:**046 10.0 points**

Consider a buffer made of 0.10 M HF and 0.15 M KF with a volume of 450 mL. Which of the following statements is/are true?

- I) If 10 mL 0.1 M NaOH is added, the potassium from KF and hydroxide from NaOH will react to form potassium hydroxide
- II) If a small amount of NaOH is added, the solution will increase in pH a very small amount
- III) If HI is added, the fluoride anions will react with hydrogen ions
- IV) The $\text{p}K_a$ for HF represents the neutralization reaction that forms KF

1. III only

2. II and III correct

3. II only

4. I, II, and III

5. II, III, and IV

6. I, II, and IV

Explanation:

Statement one: neither K or Na will react in solution because they are spectator ions. You will not form a strong base (KOH) in this reaction. False.

Statement two: this is a buffer solution and this is exactly what will happen with a buffer. True.

Statement three: the neutralization reaction that occurs upon the addition of a strong acid will primarily involve the reaction between H^+ and F^- . True.

Statement four: the $\text{p}K_a$ for the HF reaction

represents the deprotonation (acid reaction) of the weak acid in water. False.

047 10.0 points

At the stoichiometric point in the titration of 0.130 M HCOOH(aq) with 0.130 M KOH(aq),

1. the pH is greater than 7. **correct**
2. the pH is 7.0.
3. the pH is less than 7.
4. [HCOOH] = 0.0650 M.
5. [HCO₂⁻] = 0.130 M.

Explanation:

048 10.0 points

Methyl orange is an indicator with a K_a of 3.4×10^{-4} . Its non-ionized form is red, while its ionized form is yellow. At a pH of exactly 4.00, this indicator's color would best be described as which of the following colors?

1. red-orange
2. orange
3. green
4. blue
5. yellow
6. cyan
7. red
8. purple
9. yellow-orange **correct**

Explanation:

The pK_a is 3.47 so that is where yellow and red mix 50/50 to give an orange color. You would see all red for any pH at or below 2.47, and see only yellow for any pH at or above 4.47. The pH 4.00 given is on the distinctly

basic side of the middle so it would have a yellow-orange color.

049 10.0 points

A solution has an initial concentration of 0.0100 M HClO ($K_a = 3.5 \times 10^{-8}$) and 0.0300 M NaClO. What is the pH after the addition of 0.0030 mol of solid NaOH to 1.00 L of this solution? Assume no volume change.

1. 8.02
2. 5.33
3. 8.13 **correct**
4. 9.06
5. 5.34

Explanation:

$$\begin{array}{l} [\text{HClO}] = 0.01 \text{ M} \\ K_a = 3.5 \times 10^{-8} \end{array} \qquad \begin{array}{l} [\text{ClO}^-] = 0.03 \text{ M} \\ \text{NaOH} = 0.0030 \text{ mol} \end{array}$$

	HClO	+ NaOH	→	Na ⁺	+ ClO ⁻	
					+ H ₂ O	
ini	0.01	0.003		0.03	0.03	
Δ	-0.003	-0.003		0.003	0.003	
fin	0.007	0		0.033	0.033	

Na⁺ is a spectator ion. HClO and OCl⁻ produce a buffer system.

$$\begin{aligned} \text{pH} &= \text{p}K_a + \log \left(\frac{[\text{ClO}^-]}{[\text{HClO}]} \right) \\ &= -\log (3.5 \times 10^{-8}) + \log \left(\frac{0.033}{0.007} \right) \\ &= 8.12935 \end{aligned}$$