

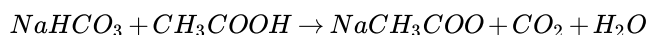
15.E: Acid-Base Equilibria (Exercises)

These are homework exercises to accompany the Textmap created for "Principles of Modern Chemistry" by Oxtoby et al. Complementary General Chemistry question banks can be found for other Textmaps and can be accessed [here](#).

Q3

Vinegar contains acetic acid CH_3COOH . What species serves as a base when vinegar is mixed with baking soda, sodium bicarbonate, during the preparation of bread?

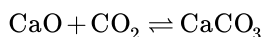
Solution



NaHCO_3 serves as the base.

Q7

- Thinking of acid-base reaction in terms of oxide donors and oxide acceptors, is the base oxide donors or oxide acceptors?
- Identify the acid and base in the reaction:



Solution

- The base is the oxide donor. We can distinctly see this in the autoionization of H_2O , where OH^- counts as the basic part
- CaO is the base, and CO_2 is the acid.

Q5

Baking soda known as NaHCO_3 is formed by adding water and carbon dioxide to sodium carbonate.

- Write a balanced equation for this chemical reaction
- Is this a Brønsted-Lowry acid-base reaction? What is a Brønsted-Lowry acid? What is a Brønsted-Lowry base?

Solution

- $\text{Na}_2\text{CO}_3 + \text{H}_2\text{O} + \text{CO}_2 \rightarrow 2\text{NaHCO}_3$
- This is not a Brønsted-Lowry acid-base reaction because it does not involve a transfer of an H^+ ion. A Brønsted-Lowry acid is a proton donor, and a Brønsted-Lowry base is a proton acceptor.

Q9

Identify each of the following oxides as an acid or base anhydride:

- CaO
- P_2O_5

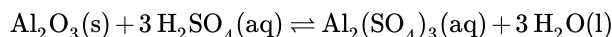
Solution

- CaO is the base anhydride of calcium hydroxide $\text{Ca}(\text{OH})_2$.
- P_2O_5 is the acid anhydride of phosphoric acid H_3PO_4 .

Q11

$\text{Al}(\text{III})$ oxide is amphoteric. What is the balanced chemical equation of $\text{Al}(\text{III})$ oxide react with aqueous H_2SO_4 ? What is the balanced equation of it reacts with KOH ?

Solution



Q13

The $[\text{H}_3\text{O}^+]$ concentration in a glass of orange juice is 3.96×10^{-5} M. What is the juice's pH?

Solution

$$\text{pH} = -\log[\text{H}_3\text{O}^+]$$

Since the concentration of hydronium ions is given, the pH calculation is as follows:

$$\text{pH} = -\log[3.96 \times 10^{-5}] = 4.4023$$

pH is a measure of hydrogen ions in a solution. This concentration defines the acidity or alkalinity of a solution.

Q17

The pK_w of an unknown salty water at 25°C is 13.665. This differs from the usual K_w of 14.00 at this temperature because dissolved salts make this unknown salty water a non-ideal solution. If the pH in the salty water is 7.8, what are the concentrations of H_3O^+ and OH^- in the salty water at 25°C ?

Solution

$\text{pH} = -\log[\text{H}_3\text{O}^+]$, hence,

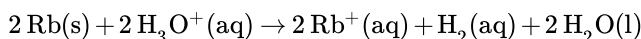
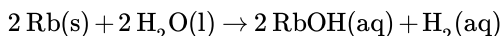
$$[\text{H}_3\text{O}^+] = 10^{-7.8} = 1.5849 \times 10^{-8} \text{ M}$$

Since $\text{pK}_w = [\text{OH}^-][\text{H}_3\text{O}^+]$,

$$[\text{OH}^-] = \frac{10^{-13.665}}{1.5849 \times 10^{-8} \text{ M}} = 1.3645 \times 10^{-6} \text{ M}$$

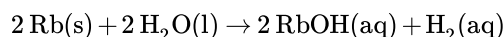
Q19

When rubidium (Rb) solid is added to water, there is an instantaneous and vigorous reaction (i.e., an explosion) as this video demonstrates. Based on this information, which of these two equations is a more accurate representation the reaction?



Solution

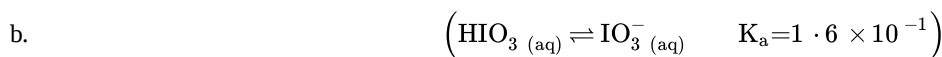
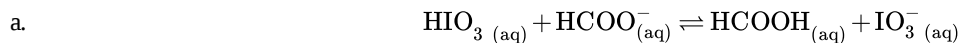
The equation:



Is a better representation of Rb being placed into water. It represents the direct interaction between the alkali metal and water. We know from the problem that the reaction is fast, vigorous, and forceful. The second equation represents a reaction in equilibrium, one that we would expect to be slow as there are only $1 \times 10^{-7} \text{ mol/L}$ of H_3O^+ in a 1L of water. Even if the reaction were vigorous at the low concentration of hydronium ions, there would not be enough of them to keep up with the speed of the reaction (they would become a limiting reactant), hence slowing or stopping the reaction from proceeding.

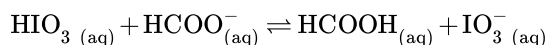
Q23

In the following chemical equation determine which species is the strongest acid and which is the strongest base, using the Brønsted–Lowry definition. At equilibrium, is there a greater concentration of reactants or products present?

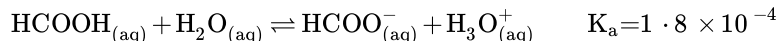
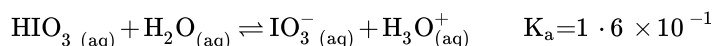


Solution

A Brønsted–Lowry acid is the species that donates protons in a solution. When comparing two different weak acids in solution, like:



We can compare their abilities to donate protons to see which one is the stronger of the two weak acids.



Seeing that $K_a = 1 \cdot 6 \times 10^{-1} > K_a = 1 \cdot 8 \times 10^{-4}$, we know HIO_3 is the stronger acid.

A Brønsted–Lowry base is the species that accepts protons. So in this case, we must examine which of the two weak acids has a stronger conjugate base, which means we must find the K_b for the reactions of the conjugate bases.

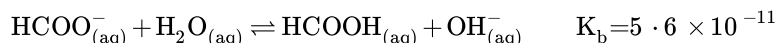
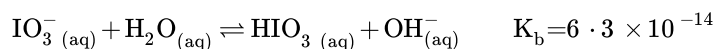
We know that:

$$K_w = K_a \times K_b$$

and

$$K_w = 1.0 \times 10^{-7}$$

So we can find K_b by dividing K_w by K_a which ultimately gives us:



Seeing that $K_b = 5 \cdot 6 \times 10^{-11} > K_b = 6 \cdot 3 \times 10^{-14}$, we know HCOO^- is the stronger base.

To determine whether there is a greater concentration of reactants or products present, the K value for the overall reaction must be determined. The overall reaction is the product of the first given reaction and the reverse of the second given reaction. Dividing the first value for K_a by the second gives

$$K = 888$$

$$K > 1$$

which indicates that at equilibrium, there is a greater concentration of products than reactants.

Q27

Acetic acid gives vinegar a sour taste and strong aroma. Its K_a value is 1.75×10^{-5} . What is the pH of the solution if 0.59 grams of acetic acid is dissolved in 40 mL of water?

Solution

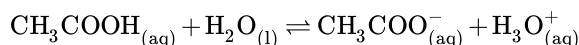
First, convert grams of acetic acid to moles.

$$(0.59 \text{ g } \text{CH}_3\text{COOH}) \left(\frac{1 \text{ mol}}{60.05 \text{ g } \text{CH}_3\text{COOH}} \right) = 0.0098 \text{ mol}$$

Then, find the molarity of acetic acid by dividing the number of moles of acetic acid by the number of liters of water.

$$\frac{0.0098 \text{ mol}}{0.04 \text{ L water}} = 0.246 \text{ M}$$

Using the molarity and K_a , construct and solve an ICE table to find out how much the acetic acid dissociates.



	CH_3COOH	H_2O	CH_3COO^-	H_3O^+

I	0.246	---	0	0
C	-x	---	+x	+x
E	0.246-x	---	x	x

The acid dissociation constant works in the below equation:

$$K_a = \frac{[H_3O^+][CH_3COO^-]}{[CH_3COOH]}$$

Plug in the final concentration values from the ICE table and solve for x.

$$1.75 \times 10^{-5} = \frac{x^2}{0.246 - x}$$

$$x = 0.0021$$

Use the calculated value of x to calculate the concentration of H_3O^+ , which in this case is equal to x. Then, plug this concentration into the equation for pH.

$$pH = -\log[0.0021]$$

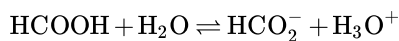
$$pH = 2.6848$$

Q29

- A student prepares a solution of 0.60M of formic acid carefully in a water bath that remains constant at 25°C, determine the pH of the solution. (K_a of formic acid: 1.8×10^{-4})
- How many grams of trichloroacetic acid should be dissolved per liter of deionized water so that the solution of trichloroacetic acid would have the same pH as that of the formic acid solution in a)? (K_a of trichloroacetic acid is 2.2×10^{-1})

Solution

Construct an ICE table based on the equation:



	$HCOOH$	H_2O	HCO_2^-	H_3O^+
I	0.6	-	0	0
C	-x	-	+x	+x
E	0.6-x	-	x	x

K_a can then be equated to an algebraic expression:

$$K_a = \frac{x^2}{0.6 - x}$$

$$1.8 \times 10^{-4} = \frac{x^2}{0.6 - x}$$

$$x = 0.0103026$$

or

$$x = -0.0104$$

x must be positive. Thus, $x = 0.0103026$.

$$[\text{H}_3\text{O}^+]$$

$$\text{pH} = -\log(0.0103026) = 1.987$$

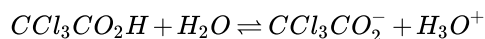
Because we know the pH we want to attain (1.987), we have start by first finding $[\text{H}_3\text{O}^+]$:

$$-\log[\text{H}_3\text{O}^+] = 1.987$$

$$[\text{H}_3\text{O}^+] = 10^{-1.987}$$

$$[\text{H}_3\text{O}^+] = 0.0103026$$

An ICE table can also be constructed for the reaction:



	$\text{CCl}_3\text{CO}_2\text{H}$	H_2O	$\text{CCl}_3\text{CO}_2^-$	H_3O^+
I	x	-	0	0
C	-0.010303	-	+0.010303	+0.010303
E	x-0.010303	-	0.010303	0.010303

$$K_a = \frac{(0.010303)^2}{(x - 0.010303)}$$

$$2.2 \times 10^{-1} = \frac{(0.010303)^2}{(x - 0.010303)}$$

$$x = 0.0107855\text{M}$$

To calculate the mass of trichloroacetic acid, we can calculate the molarity of trichloroacetic acid by its molar mass.

$$\text{Mass of trichloroacetic acid} = 0.0107855 \text{ M} \times \left(35.5 \frac{\text{g}}{\text{mol}} \times 3 + 12 \frac{\text{g}}{\text{mol}} \times 2 + 16 \frac{\text{g}}{\text{mol}} \times 2 + 1 \frac{\text{g}}{\text{mol}} \right) = 1.76 \text{ g}$$

Q41

Rank each of the 0.2 M solution below in an order of increasing pH: NH_4I , KF , HCl , KCl , KOH .

Solution



Q43

0.040 mol of Diethylamine ($\text{C}_4\text{H}_{11}\text{N}$, $\text{pK}_b = 11.09$) is titrated with 0.015 mol of HCl in a 1.00L wash bottle, calculate the pH value of the solution.

Solution

Because equivalence point is not reached yet, we can employ [Henderson-Hasselbalch equation](#).

$$pOH \approx pK_b + \log \frac{[BH^+]}{[B]}$$

$$pOH \approx (11.09) + \log \frac{[C_4H_{12}N^+]}{[C_4H_{11}N]}$$

$$pOH \approx (11.09) + \log \frac{0.015}{0.040 - 15}$$

$$pOH \approx 10.87$$

At room temperature $K_w = 14$ and

$$\begin{aligned} pH &= pK_w - pOH \\ &= 14.00 - 10.87 \\ &= 3.13 \end{aligned}$$

Q45

Prepare a Hypochlorous acid/Hypochlorite buffer at pH 7.

- Suppose you only have 0.5 mol of lithium hypochlorite, but an infinite supply of hypochlorous acid ($pK_a = 7.53$) and water. How many moles of hypochlorous acid should you use, assuming you use all 0.5 mol of lithium hypochlorite, and dilute the buffer to 100 mL?
- Suppose Professor Güntherfærd's arm was doused in about 0.12 moles total HCl . Will this buffer be enough to bring her arm up to a pH over 6.5?

Solution

Using the [Henderson–Hasselbalch approximation](#):

$$pH \approx pK_a + \log \frac{[ClO^-]_o}{[HClO]_o}$$

It's easy to re-arrange the equation to solve for $[HClO]_o$. Multiplying by 100 mL yields the moles of acid added.

$$pH \approx pK_a + \log \frac{[ClO^-]_o}{[HClO]_o}$$

$$7 \approx pK_a + \log \frac{[ClO^-]_o}{[HClO]_o}$$

$$7 \approx 7.53 + \log \frac{5}{[HClO]_o}$$

$$[HClO]_o \approx 16.94$$

$$\text{Moles } HClO \approx 1.694$$

Q49

A student is given 500 mL of a 0.500 M acetic acid solution and wants to create a pH 5.0 buffer. How many mL of 1 M NaOH must be added to the original solution? Acetic acid has a pK_a of 4.756.

Solution

We first use the [Henderson-Hasselbach approximation](#) to determine the required ratio of base to acid in the solution:

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

$$5.00 = 4.756 + \log\left(\frac{[base]}{[acid]}\right)$$

$$0.244 = \log\left(\frac{[base]}{[acid]}\right)$$

$$1.75 = \frac{[base]}{[acid]}$$

Next, we will determine the molar amount of base required to get a pH of 5.00. In order to simplify things we will first solve for the molar quantities as of each using the simple ratio above in relation to 1.75:

$$\begin{aligned}\text{Acetic Acid in 500 mL solution} &= (0.5 \text{ M}) \times (0.5 \text{ L}) \\ &= 0.25 \text{ mol}\end{aligned}$$

$$\text{NaOH in 500 mL solution} = (0.25 \text{ mol}) \times (1.75) + 0.25 = 0.6875 + \text{mol}$$

Now we'll determine the volume of 1M NaOH needed to raise the pH:

$$\begin{aligned}\left(\frac{1 \text{ mol}}{1 \text{ L}}\right) &= \left(\frac{0.6875 \text{ mol}}{x}\right) \\ s &= 0.6875 \text{ L NaOH}\end{aligned}$$

Now we'll check to make sure everything is right:

$$\text{Acetic Acid} = \left(\frac{0.25}{0.9375}\right) = 0.2667 \text{ M}$$

$$\text{NaOH} = \left(\frac{0.6875 - 0.25}{0.9375}\right) = 0.4667 \text{ M}$$

$$5.00 = 4.756 + \log\left(\frac{[base]}{[acid]}\right)$$

$$pH = 4.756 + \log\left(\frac{[0.4667]}{[0.2667]}\right)$$

$$pH = 4.756 + 0.243 = 4.999$$

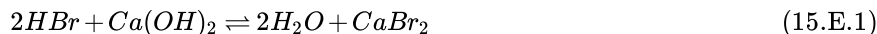
We get a pH of 4.999 which is about 5.00 and isn't exactly 5.00 due to rounding early in the problem, so the problem was done correctly.

Q51

0.15M of HBr is added into 50mL of 0.1M Ca(OH)_2 for the titration.

- What is the pH of the solution before HBr is added?
- What is the pH of the solution at the point when it needs 1 mL of HBr to neutralize the solution?
- What is the pH of the solution when it is titrated 1 mL past neutralization?

Solution



a) Ca(OH)_2 is strong base. They dissociate completely.

$$[\text{OH}^-] = 2[\text{Ca(OH)}_2] = 0.2\text{M}$$

$$\begin{aligned}\text{pH} &= \text{pK}_w - \text{pOH} \\ \text{pH} &= 14 + \log(0.2 \text{ M}) \\ \text{pH} &= 13.3\end{aligned}$$

b) Because $\text{Ca}(\text{OH})_2$ and HBr are both strong base and strong acid, at equilibrium, the pH is 7.00.

$$\text{mole}_{\text{Ca}(\text{OH})_2} = (\text{Volume})(\text{Molarity}) \quad (15.E.2)$$

$$\text{mole}_{\text{Ca}(\text{OH})_2} = (0.05\text{L})(0.1\text{M}) \quad (15.E.3)$$

$$\text{mole}_{\text{Ca}(\text{OH})_2} = 0.005 \text{ mol} \quad (15.E.4)$$

@ equilibrium;

$$\text{mole}_{\text{HBr}} = 2 \text{ mole}_{\text{Ca}(\text{OH})_2} \quad (15.E.5)$$

$$\text{mole}_{\text{HBr}} = 2(0.005 \text{ mol}) = 0.01 \text{ mol} \quad (15.E.6)$$

$$\text{Volume}_{\text{HBr}} = \frac{\text{mole}_{\text{HBr}}}{\text{molarity}_{\text{HBr}}} = \frac{0.01 \text{ mol}}{0.15\text{M}} = 0.0667 \text{ L} = 66.7 \text{ mL} \quad (15.E.7)$$

$$\text{Total Volume at equilibrium} = 66.7\text{mL} + 50\text{mL} = 116.7\text{mL} \quad (15.E.8)$$

$$\text{Total Volume 1mL short of equilibrium} = 116.7\text{mL} - 1\text{mL} = 115.7\text{mL} \quad (15.E.9)$$

$$\text{Volume}_{\text{HBr}} \text{ 1mL short of equilibrium} = 66.7\text{mL} - 1\text{mL} = 65.7 \text{ mL} \quad (15.E.10)$$

$$\text{mols of OH}^- \text{ not neutralized by HBr} = 0.01 \text{ mol} - 0.0657 \text{ L} \times (0.15 \text{ M}) = 1.45 \times 10^{-4} \text{ mol} \quad (15.E.11)$$

$$\text{pH} = 14 + \log\left(\frac{1.45 \times 10^{-4} \text{ mol}}{0.1157 \text{ L}}\right) = 11.1 \quad (15.E.12)$$

c)

At 1mL after equilibrium, $\text{Ca}(\text{OH})_2$ has been neutralized by HBr . Only HBr exists in the solution.

$$\text{mole}_{\text{HBr}} = (0.001\text{L})(0.15\text{M}) = 1.5 \times 10^{-4} \text{ mol} \quad (15.E.13)$$

$$\text{Volume} = \text{Total Volume} + 1\text{mL} = 117.7\text{mL} = 0.1177 \text{ L} \quad (15.E.14)$$

$$\text{M}_{\text{HBr}} \text{ in the solution} = \frac{1.5 \times 10^{-4} \text{ mol}}{0.1177 \text{ L}} = 1.274 \times 10^{-3} \text{ M} \quad (15.E.15)$$

$$\text{M}_{\text{HBr}} \text{ in the solution} = \text{M}_{\text{H}^+} \text{ in the solution} \quad (15.E.16)$$

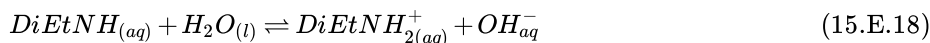
$$\text{pH} = -\log[\text{H}^+] = -\log(1.274 \times 10^{-3} \text{ M}) = 2.89 \quad (15.E.17)$$

Q55

K_b at 25°C for Diethylamine $((\text{C}_2\text{H}_5)_2\text{NH})$ is 1.3×10^{-3} . Consider the titration of 50.00 mL of a 0.1000 M solution of Diethylamine with 0.100 M HCl added with the following volumes: 0, 10.00, 50.00 mL. Calculate pH for each solutions. At an unknown volume beyond 50.00 mL, the pH is 3.90. Find the corresponding amount volume of HCl needed to obtain that pH.

Solution

When HCl volume = 0.00 mL.



$$\frac{[\text{DiEtNH}_2^+][\text{OH}^-]}{[\text{DiEtNH}]} = 1.3 \times 10^{-3} \quad (15.E.19)$$

$$[\text{OH}^-] = [\text{DiEtNH}_2^+] = y$$

$$[\text{DiEtNH}] = 0.1000 - y$$

$$\frac{y^2}{0.1000 - y} = 1.3 \times 10^{-3} \quad (15.E.20)$$

$$y = 0.01077\text{M} = [\text{OH}^-]$$

$$\text{pOH} = 1.97$$

$$\text{pH} = 12.03$$

When HCl volume = 10.00 mL.

$$[\text{DiEtNH}_2^+] = \frac{(0.1000\text{M})(0.01\text{L})}{(0.050 + 0.010)\text{L}} = 0.0167\text{ M} \quad (15.E.21)$$

$$[\text{DiEtNH}] = \frac{(0.1000\text{M})(0.050\text{L}) - (0.1000\text{M})(0.01\text{L})}{(0.050 + 0.010)\text{L}} = 0.0667\text{ M} \quad (15.E.22)$$

Plug it back to Henerson Hasselbalch equation

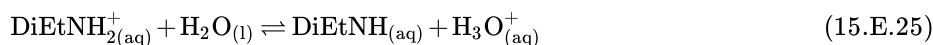
$$\text{pOH} = \text{pK}_b + \log\left(\frac{[\text{BH}^+]}{[\text{B}]}\right) \quad (15.E.23)$$

$$\text{pOH} = -\log(1.3 \times 10^{-3}) + \log\left(\frac{[0.0167\text{ M}]}{[0.0667\text{ M}]}\right)$$

$$\text{pH} = 14.00 - \text{pOH} = 14.00 - 2.28 = 11.72 \quad (15.E.24)$$

When HCl volume = 50 mL.

The titration is at the equivalence point. At equivalence, the reaction consists of 100 mL of 0.050 mol DiEtNH_2^+



$$K_a = \frac{1.00 \times 10^{-14}}{1.3 \times 10^{-3}} = 7.69 \times 10^{-12}$$

$$K_a = 7.69 \times 10^{-12} = \frac{x^2}{(0.0500 - x)}; x = [\text{H}_3\text{O}^+] \quad (15.E.26)$$

$$x = 1.96 \times 10^{-6}; \text{pH} = 5.71$$

Beyond the equivalence point

Beyond the equivalence point, solution behaves like HCl .

Given $\text{pH} = 3.90$.

$$10^{-3.90} = \frac{(z - 0.050\text{ L})}{0.100\text{ L} + z} \times 0.1000\text{ M} \quad (15.E.27)$$

$$z = 0.0502\text{ L}$$

Q57

Sodium Bicarbonate (NaHCO_3) is a very weak base when dissolved in water. Some amount of sodium bicarbonate is dissolved in 125 mL of a 0.25 M solution of HNO_3 . The 168 mL of 0.15 M NaOH was used to titrate the solution. How many grams of sodium bicarbonate were added?

Solution

We are titrating an acid with two bases so solve for the amount of acid the NaOH neutralizes and the remaining moles of acid will be the number of moles of sodium bicarbonate.

$$\text{Moles HNO}_3 = \frac{0.25\text{ moles}}{1\text{ L}} \times 0.125\text{ L} = 0.03125\text{ moles}$$

$$\text{Moles NaOH} = \frac{0.15 \text{ moles}}{1 \text{ L}} \times 0.168 \text{ L} = 0.0252 \text{ moles}$$

$$0.03125 - 0.0252 = 0.00605 \text{ moles sodium bicarbonate}$$

Now just multiply by sodium bicarbonate's molar mass ($84.007 \frac{\text{g}}{\text{mol}}$) to find the mass of sodium bicarbonate added

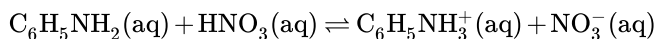
$$0.00605 \times 84.007 = 0.51\text{g}$$

Q59

What is the molarity of a HNO_3 water solution if it requires 31.80 mL of such solution to titrate 0.0662 g Aniline in 100 mL aqueous solution to equivalence point? What will the pH value be at the equivalence point if $K_b(\text{Aniline}) = 3.8 \times 10^{-10}$?

Solution

Aniline and HNO_3 react with a one-to-one stoichiometry



Therefore

$$M(\text{HNO}_3) = \frac{0.0662 \text{ g}}{93.13 \text{ g/mol}} \times \frac{1}{0.03180 \text{ L}} = 0.0224 \text{ M}$$

Suppose at the equivalence point all Aniline is converted to its conjugate acid, then its concentration equals

$$[\text{C}_6\text{H}_5\text{NH}_3^+] = \frac{0.0662 \text{ g}}{(93.13 \text{ g/mol}) 0.1318 \text{ L}} = 0.00539 \text{ M}$$

Also as some Aniline's conjugate acid reacts with water,

$$K_a = \frac{K_w}{K_b} = \frac{1.0 \times 10^{-14}}{3.8 \times 10^{-10}} = 2.63 \times 10^{-5} = \frac{[\text{H}_3\text{O}^+][\text{C}_6\text{H}_5\text{NH}_2]}{[\text{C}_6\text{H}_5\text{NH}_3^+]} = \frac{x^2}{0.00539 - x}$$

Therefore,

$$x = [\text{H}_3\text{O}^+] = 3.636 \times 10^{-4} \text{ M}$$

so $\text{pH} = 3.44$.

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