CH302
Unit 6 Activity: Acids, Bases, and Salts

Part I
You have had a lot of practice using ionization constants and calculating the concentration of species in aqueous solutions. Recently we have added practice in calculating pHs of solutions. Here is a review question:

1. The pH of a 0.2 M aqueous solution of crotonic acid is 2.69. What is the K_a of crotonic acid?
   pH=2.69; $[H^+]= 10^{-pH} = 2.042 \times 10^{-3}$

   Assume 1 L

Reaction	HA (aq) →	H^+(aq) +	A-(aq)
Initial	0.2 mol	0	0
Change	-2.402×10^{-3} mol	$+2.402 \times 10^{-3}$ mol	$+2.402 \times 10^{-3}$ mol
Equilibrium	$0.2 - 2.402 \times 10^{-3}$ ~ 0.2 mol 0.2 mol/1L = 0.2 M	2.402×10^{-3} mol 2.402×10^{-3} mol/1L = 2.402×10^{-3} M	2.402×10^{-3} mol 2.402×10^{-3} mol/1L = 2.402×10^{-3} M

$$\boxed{k_a = \frac{[H^+][A^-]}{[HA]} = \frac{[2.042 \times 10^{-3}][2.042 \times 10^{-3}]}{[0.2]} = 2.08 \times 10^{-5}}$$

2. What is the percent ionization of the acid in this solution?

$$\boxed{\%_{ionization} = \frac{[A^-]}{C_{HA}} * 100\% = \frac{[2.042 \times 10^{-3}]}{0.2} * 100\% = 0.12 * 100\% = 1.2\%}$$

Part II: Acid-Base Reactions: Strong Acid + Strong Base

Up to this point we have focused on how an acid will behave in water or how a base will behave in water. Now we want to look at reactions involving both acids and bases together in a solution. Consider this reaction:

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$$

1. If equal parts of equal concentration acid and base are mixed, is the resulting solution acidic, basic, or neutral?
   Neutral, the conjugates of strong acids and bases do not form weak acid or base
2. Now let's think about a reaction with specific amounts and concentrations: You start with 200 mL of a 0.1 M HCl solution and add it to 500 mL of a 0.1 M NaOH solution.

   a. What is the pH of the original HCl solution?

   $$\boxed{pH = -\log[0.1] = 1}$$

   b. What is the pH of the original NaOH solution?

   $$\boxed{\begin{aligned} &pOH = -\log[0.1] = 1 \\ &pOH + pH = 14 \\ &pH = 14 - pOH = 14 - 1 = 13 \end{aligned}}$$

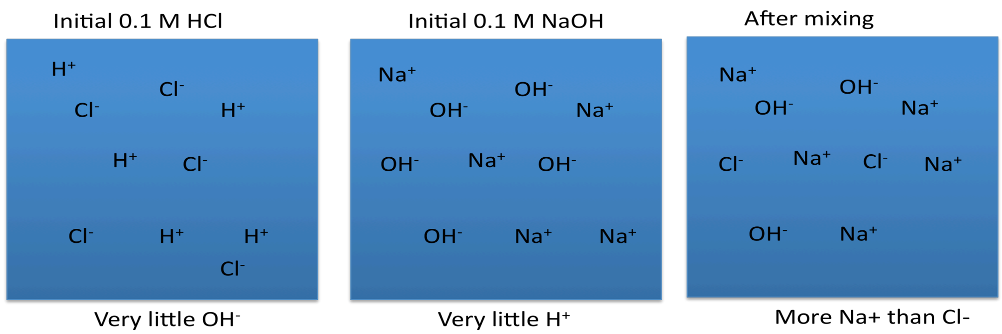
   c. What is the pH of the new solution after these are mixed?

   $$\boxed{0.200\ L * 0.1 \frac{mol}{L} = 0.02\ mol\ HCl; \quad 0.500\ L * 0.1 \frac{mol}{L} = 0.05\ mol\ NaOH}$$

Reaction	HCl (aq) +	NaOH (aq) →	NaCl(aq) +	H_2O (l)
Initial	0.02	0.05	0	N/A
Change	-0.02	-0.02	+0.02	N/A
Equilibrium	0	0.03	0.2	N/A

$$\boxed{\begin{aligned} &[NaOH] = \frac{0.03\ mol\ NaOH}{0.7\ L} = 0.043\ M; \\ &pOH = -\log(0.043) = 1.37 \\ &pH = 14 - 1.37 = 12.63 \end{aligned}}$$

d. Draw pictures showing a molecular view of each of these solutions – the original HCl solution, original NaOH solution, and the solution after the reaction has occurred.

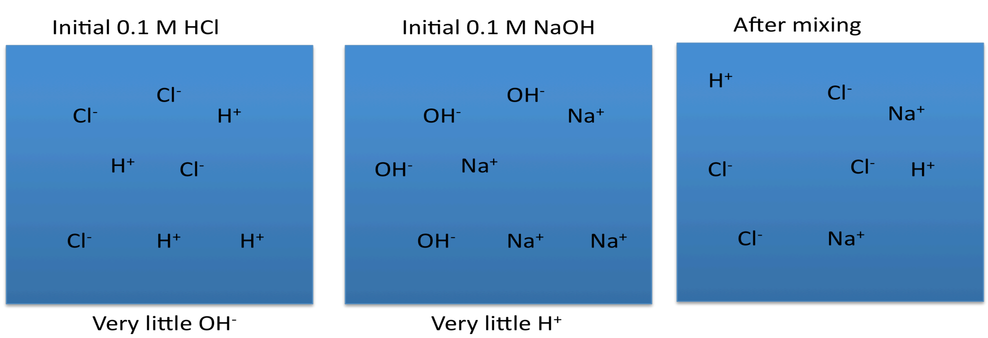


3. Consider the same reaction with these starting amounts: 2 parts 0.1 M HCl mixed with 1 part 0.1 M NaOH.

   a. Will this solution be acidic, basic, or neutral?

      Acidic

   b. Draw a picture showing a molecular view of the resulting solution.



   c. What is the pH of the resulting solution?

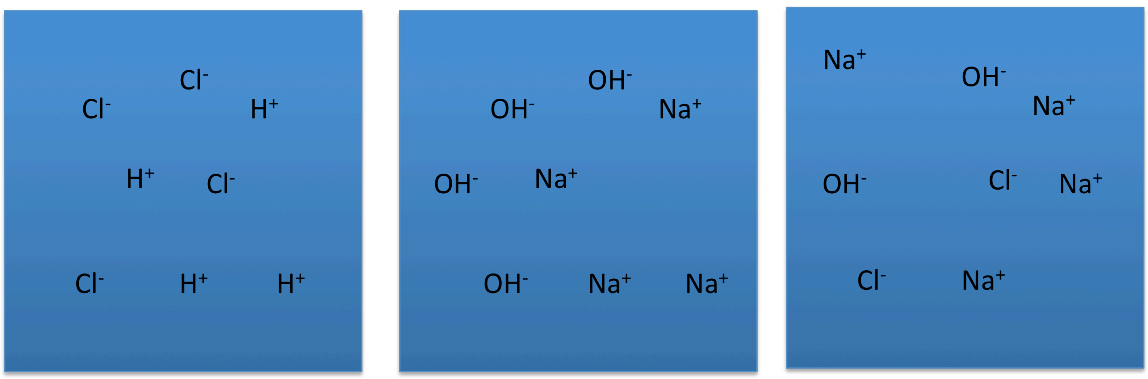
      Assume 2 L HCl and 1 L NaOH.

Reaction	HCl (aq) +	NaOH (aq) →	NaCl(aq) +	H_20 (l)
Initial	0.2	0.1	0	N/A
Change	-0.1	-0.1	+0.1	N/A
Equilibrium	0.1	0	0.1	N/A

$$\boxed{[HCl] = \frac{0.1\ mol\ NaOH}{3\ L} = 0.033\ M; \quad pH = -\log(0.033) = 1.48}$$

4. Consider the same reaction with these starting amounts: 1 part 0.1 M HCl mixed with 2 parts 0.1 M NaOH.
   a. Will this solution be acidic, basic, or neutral?
      Basic
   b. Draw a picture showing a molecular view of the resulting solution.

Initial 0.1 M HCl | Initial 0.1 M NaOH | After mixing



Very little OH^- | Very little H^+

   c. What is the pH of the resulting solution?

      Assume 1 L HCl and 2 L NaOH.

Reaction	HCl (aq) +	NaOH (aq) →	NaCl(aq) +	H_20 (l)
Initial	0.1	0.2	0	N/A
Change	-0.1	-0.1	+0.1	N/A
Equilibrium	0.0	0.1	0.1	N/A

$$\boxed{\begin{aligned} &[NaOH] = \frac{0.1\ mol\ NaOH}{3\ L} = 0.033\ M; \\ &pOH = -\log(0.033) = 1.48 \\ &pH = 14 - 1.48 = 12.52 \end{aligned}}$$

Part III: Weak Acid-Strong Base Reactions

The reactions in part II make use of our basic stoichiometry. As we move on to consider reactions involving weak acids and bases, we will still need that stoichiometry, but now we also must remember that not only can the original reactant acids and bases change the pH of the solution, but the resulting salt may contain ions that can act as an acid or base. This was not an issue in the strong acid-strong base situation, because their conjugate acids/bases (the ions in the salt) would be exceedingly weak and have essentially no effect on the solution pH.

For part III, consider this reaction:

$$CH_3COOH\ (aq) + NaOH(aq) \rightarrow NaCH_3COO\ (aq) + H_2O\ (l)$$

$$K_a \text{ for } CH_3COOH = 1.8 \times 10^{-5}$$

1. Is a 0.1 M CH_3COOH solution acidic, basic, or neutral? Acidic
2. If 1 part 0.1 M CH_3COOH is mixed with 1 part 0.1 M NaOH:
   a. Is the resulting solution acidic, basic, or neutral? Basic
   b. Calculate the pH of the resulting solution.

Assume 1 L CH_3COOH and 1 L NaOH

Reaction	$CH_3COOH(aq)$ +	NaOH (aq) →	$NaCH_3COO(aq)$ +	H_20 (l)
Initial	0.1 mol	0.1 mol	0 mol	N/A
Change	-0.1 mol	-0.1 mol	+0.1 mol	N/A
Equilibrium	0.0 mol	0.0 mol	0.1 mol	N/A

Now, we work a weak base problem with $NaCH_3COO$

Reaction	$NaCH_3COO$ (aq) +	H_20 (l) →	$CH_3COOH(aq)$ +	OH- (aq)
Initial	0.1 mol	N/A	0	0
Change	-x	N/A	+x	+x
Equilibrium	0.1-x ~ 0.1 mol = 0.1 mol/2L = 0.05 M	N/A	x	x

$$k_a \times k_b = k_w; \quad k_b = \frac{k_w}{k_a} = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.56 \times 10^{-10}$$

$$k_b = \frac{[CH_3COOH][OH]}{[NaCH_3COO]} = \frac{[x][x]}{[0.05]} = \frac{x^2}{0.05}$$

$$5.56 \times 10^{-10} = \frac{x^2}{0.05}; \quad x^2 = 2.78 \times 10^{-11}$$

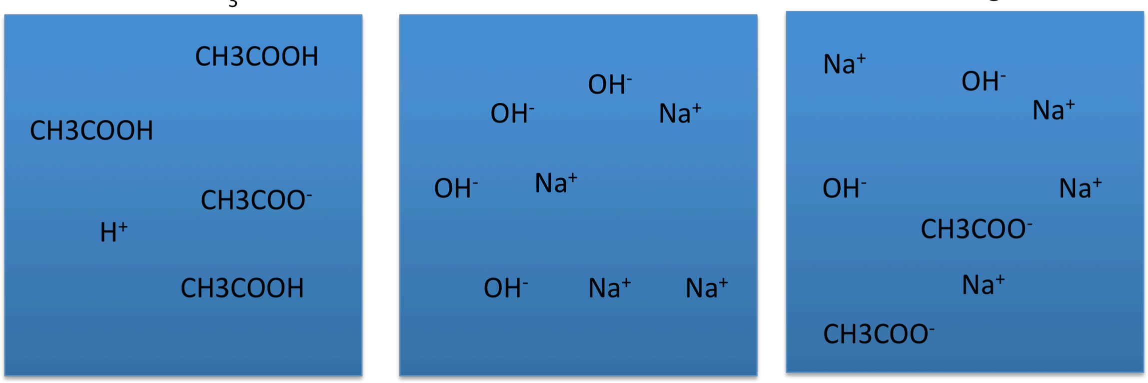
$$x = \sqrt{2.78 \times 10^{-11}} = 5.27 \times 10^{-6} = [OH]$$

$$pOH = -\log(5.27 \times 10^{-6}) = 5.3$$

$$pH = 14 - 5.3 = 8.7$$

3. If 1 part 0.1 M CH_3COOH is mixed with 2 parts 0.1 M NaOH, is the resulting solution acidic, basic, or neutral? Basic
   a. Compare the pH to the pH in the previous question – Predict: will it be higher, lower, or the same? Higher than 2b
   b. Draw a picture showing a molecular view of the resulting solution.

Initial 0.1 M CH_3COOH | Initial 0.1 M NaOH | After mixing



Small amount of dissociation

   c. Calculate the pH of this solution.

Assume 1 L CH_3COOH and 2 L NaOH

Reaction	$CH_3COOH(aq)$ +	NaOH (aq) →	$NaCH_3COO(aq)$ +	H_20 (l)
Initial	0.1 mol	0.2 mol	0 mol	N/A
Change	-0.1 mol	-0.1 mol	+0.1 mol	N/A
Equilibrium	0.0 mol	0.1 mol	0.1 mol	N/A

Now, we work a strong base problem with NaOH

$$\boxed{\begin{array}{l} [NaOH] = \dfrac{0.1\ mol\ NaOH}{3\ L} = 0.033\ M; \\ pOH = -\log(0.033) = 1.47 \\ pH = 14 - 1.5 = 12.5 \end{array}}$$

Part IV: pH of salt solutions

1. The pH of a solution of a soluble salt will be:
   A) Neutral
   B) Basic
   C) Acidic
   D) Any of the above, depends on the salt

2. The pH of a 0.1 M aqueous solution of $NaCH_3COO$ will be:
   A) Neutral
   B) Basic
   C) Acidic

3. Can you calculate the pH of a 0.1 M aqueous solution of $NaCH_3COO$?

Assume 1 L

Reaction	$NaCH_3COO$ (aq) +	H_20 (l) →	CH_3COOH(aq) +	OH- (aq)
Initial	0.1 mol	N/A	0	0
Change	-x	N/A	+x	+x
Equilibrium	0.1-x ~ 0.1 mol = 0.1 mol/1L = 0.1 M	N/A	x	x

$$\boxed{\begin{aligned} & k_a \times k_b = k_w; \quad k_b = \frac{k_w}{k_a} = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.56 \times 10^{-10} \\ & k_b = \frac{[CH_3COOH][OH]}{[NaCH_3COO]} = \frac{[x][x]}{[0.1]} = \frac{x^2}{0.1} \\ & 5.56 \times 10^{-10} = \frac{x^2}{0.1} \\ & x^2 = 5.56 \times 10^{-11} \\ & x = \sqrt{5.56 \times 10^{-11}} = 7.46 \times 10^{-6} = [OH] \\ & pOH = -\log(7.46 \times 10^{-6}) = 5.1 \\ & pH = 14 - 5.1 = 8.9 \end{aligned}}$$

4. For each salt, indicate whether a solution of this salt would be acidic, basic, or neutral.
   A. NaCl - Neutral
   B. NaF - Basic
   C. NH_4Cl - Acidic

5. Can you generalize what you did above, indicating acidic, basic, or neutral in each blank?
   A salt of a weak acid and strong base will make a __basic__ solution.
   A salt of a strong acid and weak base will make a __acidic__ solution.
   A salt of a strong acid and strong base will make a _neutral_ solution.

Part V: Practice what you have learned! Try this weak base + strong acid problem. You should be able to do this problem using the concepts you have learned. K_b for NH_3 = 1.8 x 10^{-5}

1. Write the chemical reaction and calculate the pH when a 0.1 M solution of ammonia is mixed with a 0.1 M solution of hydrochloric acid.
   $NH_3(aq) + HCl(aq) \rightarrow NH_4^+(aq) + Cl^-(aq)$
2. Before doing any calculations, predict whether the resulting solution would be acidic, basic, or neutral.
   Acidic
3. Calculate the pH.
   Assume 1 L NH_3 and 1 L HCl

Reaction	$HCl(aq)$ +	$NH_3(aq)$ →	$NH_4^+(aq)$ +	Cl^- (l)
Initial	0.1 mol	0.1 mol	0 mol	0 mol
Change	-0.1 mol	-0.1 mol	+0.1 mol	+0.1 mol
Equilibrium	0.0 mol	0.0 mol	0.1 mol	0.1 mol

Now, we work a weak acid problem with $NaCH_3COO$

Reaction	NH_4^+ (aq) +	H_20 (l) →	$NH_3(aq)$ +	H^+ (aq)
Initial	0.1 mol	N/A	0	0
Change	-x	N/A	+x	+x
Equilibrium	0.1-x ~ 0.1 mol = 0.1 mol/2L = 0.05 M	N/A	x	x

$$k_a \times k_b = k_w; \quad k_a = \frac{k_w}{k_b} = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.56 \times 10^{-10}$$

$$k_b = \frac{[NH_3][H^+]}{[NH_4^+]} = \frac{[x][x]}{[0.0.5]} = \frac{x^2}{0.0.5}$$

$$5.56 \times 10^{-10} = \frac{x^2}{0.05}$$

$$x^2 = 2.78 \times 10^{-11}$$

$$x = \sqrt{2.78 \times 10^{-11}} = 5.27 \times 10^{-6} = [H^+]$$

$$pH = -\log(5.27 \times 10^{-6}) = 5.3$$