

CHE 1302

Basic Principles of Modern Chemistry II

Week 10

Hello and welcome to the weekly resources for Chemistry 1302! This resource covers topics typically taught by professors during the 10th week of classes.

On our website, <https://baylor.edu/tutoring>, you'll find the following links:

“Online Study Guide Resources” – If you don't see the topics you're learning right now, click here to find the weekly resources for the rest of the semester!

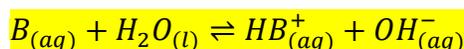
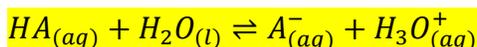
“How to Participate in Group Tutoring” - See if there is a Chemistry 1302 group tutoring session being hosted this semester – these are weekly question/answer sessions taught by our master tutors!

You can also view tutoring times for your course or schedule a private 30-minute appointment! Check out the website to learn more. You can also give us a call at (254)710-4135, or drop in. Our hours are Monday-Thursday 9 am – 8 pm on class days.

KEY WORDS: Acid/Base Equilibria, Buffered Solutions, Titrations

TOPIC OF THE WEEK: Setting up an I/C/E table for Acids/Bases

The first thing you need to know for an I/C/E table is the **reaction equation**. Remember the definitions of acids and bases and use them to write these:



Note: These charges aren't the same for all acids/bases! An acid can start out positive and become neutral when it loses a proton, or a base can start out negative and become neutral when it gains a proton.

The way the I/C/E table gets filled out depends on the strength of the acid/base. If dealing with a **strong acid/base**, it will dissociate completely. This means that the **concentration of H₃O⁺/OH⁻ at equilibrium will be the same as the concentration of the acid/base at the beginning!** A strong acid is so good at donating its proton that none of it will be left at the end of the equilibrium. Since you know the ending concentration of H₃O⁺, you can calculate the change in concentration of H₃O⁺ and fill out the rest of the I/C/E table! But in most cases, what

you'll need to find is the pH, which is equal to $-\log[H_3O^+]$, using the concentration at equilibrium.

For weak acids/bases, not all of the acid will lose its proton. See Highlight 1 for more details on setting up these problems.

Highlight 1: Buffer Solutions

If the **acid/base is weak**, fill out the **I/C/E table** as follows:

I:

If given **initial concentrations**, put them here. You can assume that hydronium and hydroxide have initial concentrations of zero (unless the solution is very dilute, less than 10^{-6} M!)

If given **percent dissociation**, $\% = \frac{[A]_f}{[HA]_i} * 100$. The percent dissociation is how much acid has lost a proton, as compared to the original amount of acid. Solve for $[HA]_i$ so that you can use it in this row: $[HA]_i = \frac{[A]_f}{\%} * 100$

E:

If given **pH or pOH**, use them to calculate $[H_3O^+]$ and $[OH^-]$. Remember, p means "negative log."

A special case of equilibrium calculations: **buffer calculations**

The definition of a buffer is a solution of weak acid (or base) and its conjugate base (or acid). The reason that it can be useful is that it **only slightly changes pH** on addition of a small amount of strong acid or base.

Some **examples**: NH_3/NH_4^+ , CH_3COOH/CH_3COO^- , $H_2PO_4^-/HPO_4^{2-}$, etc.

If given a buffer and asked to calculate its pH, you will need to use an **I/C/E table** as shown above. Alternatively, you can use the **Henderson-Hasselbach equation**, which is derived from I/C/E tables as follows:

1. Write the equation for K_a : $K_a = \frac{[H_3O^+][A^-]}{[HA]}$
2. Solve for $[H_3O^+]$, which you will use to calculate pH: $[H_3O^+] = \frac{K_a[HA]}{[A^-]}$
3. Take the negative log of both sides. Use this log property: $\log(ab) = \log(a) + \log(b)$
 - a. $-\log[H_3O^+] = -\log(K_a) - \log\left(\frac{[HA]}{[A^-]}\right)$
4. Write in terms of pH and pKa. The sign can be switched, as long as what's inside the parenthesis is inverted, as shown below. Use this log property: $a\log(b) = \log(b^a)$
 - a. $pH = pK_a + \log\left(\frac{[A^-]}{[HA]}\right)$

Remember: only use this equation if the solution is a combination of a weak acid/weak base! If strong acid/base, solve as shown in Topic of the Week.

Why would a combination of weak acid/base resist pH change? Look back at the dissociation equation for an acid. Hydronium is a stronger acid than HA, so the equilibrium favors the left side of the equation. In other words, **hydronium is better at donating its proton than HA is at donating its proton**. So, when the $HA + H_2O \rightleftharpoons A^- + H_3O^+$ equilibrium is present, **adding hydronium doesn't matter as much** as it would otherwise, because a lot of it just gets turned into HA. Don't worry about these details too much, but do know the **definition of a buffer!**

Highlight 2: Special Cases: Acid/Base Equilibria

You may come across these special situations that change different pieces of the method of setting up an equilibrium problem!

Very dilute solution: If the initial concentration of the solution is less than about 10^{-6} M, you can't ignore the initial concentrations of hydronium and hydroxide that naturally occur in water.

Polyprotic acid: If the acid is polyprotic (has multiple protons to donate), you have to consider each dissociation separately. If it is a strong acid, use the number of protons to calculate the pH! For example, H_2SO_4 has 2 protons that can be donated. So, if a solution is 0.1 M in H_2SO_4 , it will be 0.2 M in H_3O^+ . If it is a weak acid, I/C/E tables get more complicated: [CHEMISTRY 201: Calculating pH of a weak polyprotic acid - YouTube](#)

Highlight 3: Titrations

Titrations are a way that **chemists can figure out the concentration of a solution experimentally**. Say that we have an acidic solution, but we're not sure what concentration it is. In a titration, we would add base until all of the acid has reacted—and then we'd know, based on the amount of base (haha!), how much acid we started out with.

But how to tell that all of the acid has reacted? This is where **acid-base indicators** come in. An indicator is a weak acid/base solution (remember, weak acids/bases have both acid and base at equilibrium). Once you've added enough base, the base that you add will start reacting with the indicator. The special thing about indicators is that, when the base begins reacting with them, they'll change color. So, your **equivalence point** (amount of base you had to add to react with all the acid) is just before your **endpoint** (when the solution changes color because the indicator is reacting).

To calculate the **molarity of the unknown** solution:

1. Use the **volume of acid/base added**, and the **molarity of its solution**, to calculate the **number of moles of acid/base added**.
2. This is equal to the number of **moles of acid/base initially present in the solution**. Use this, and the **volume of the unknown solution**, to calculate **its initial molarity**.

To predict an **equivalence point** of a solution with given molarity:

1. Use given molarity, and initial volume, to find number of moles of acid/base initially present.
2. This is the number of moles of base/acid that will need to be added. Use this, and the molarity of the solution you're adding, to calculate the volume that will need to be added.

To calculate the **pH**:

1. Write out the reaction between the base/acid present initially and the acid/base being added:
 - a. *initially present + added* \rightleftharpoons H_2O + *conjugate of initially present*
 - b. Ex. $CH_3COOH + OH^- \rightleftharpoons H_2O + CH_3COO^-$
 - c. Figure out how many moles of acid/base were initially present, using given molarity/volume
 - d. Figure out how many moles of acid/base were added, using given molarity/volume
 - e. Solve the limiting reagent problem to figure out how much of each component is left
2. If there is leftover initially present base/acid, you are before equivalence point. Set up an I/C/E table with this base/acid and its conjugate, and fill in initial concentrations calculated in Step 1.

Here is an excellent resource that goes over strong and weak titrations—before, after, and at equivalence point: [17.4: Titrations and pH Curves - Chemistry LibreTexts](#)

Check Your Understanding

1. What is the pH of a solution of 0.1 M HCl in water?
 2. A 1.00 L buffer solution contains 0.100 mol CH_3COOH and 0.100 mol CH_3COONa . Use the Henderson-Hasselbach equation to determine the pH of the buffer solution. Calculate the new pH after adding 0.010 mol of solid NaOH to the buffer. (K_a for CH_3COOH is 1.8×10^{-5} ; ignore any slight changes in volume that might occur upon addition of the base.)
 3. What volume of 0.1 M NaOH would be required to titrate 35.0 mL of 0.231 M $H_2C_2O_4$?
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Things You May Struggle With

1. Whenever using I/C/E tables, make sure you're using molarities (mol/L), as opposed to moles.
2. Don't use the "x is small" rule if you know x, or if you're given pH or pOH.
3. If the acid is polyprotic (has more than one proton available for donation), make sure you consider all protons when calculating pH in dissociation!
4. Calculating pH during a titration can be some of the trickiest questions in this course. Try using the resource linked above to fill in a table like this one:

CALCULATION INSTRUCTIONS	WEAK ACID/BASE	STRONG ACID/BASE
BEFORE EQUIVALENCE POINT		
AT EQUIVALENCE POINT		
AFTER EQUIVALENCE POINT		

That's all this week! Please reach out if you have any questions and don't forget to visit the Tutoring Center website for further information at www.baylor.edu/tutoring. Answers to Check Your Learning are below.

1. 1
2. 4.79
3. 0.162 L (problem taken from <https://www.blogger.com/blogin.g?blogspotURL=https://lineyounnoullits.blogspot.com/2021/08/acid-base-titration-worksheet-answers.html&type=blog>)