

Acid Base Equilibrium Review

Proof of true understanding of acid base equilibrium culminates in the ability to find pH of any solution or combination of solutions. The ability to determine pH of a multitude of solutions encompasses many of the basic knowledge pieces that make up the solution equilibrium component of Big Idea 6 in the AP Chemistry course description. Also keep in mind that titration curves, which graph the relationship between pH and concentrations, are created by series of pH calculations. Titrations are a very important, yet often rushed topic in AP Chemistry, so let's make sure you see the big picture.

You should be able to find the pH of:

1. A strong acid or base
2. A weak acid or base
3. A salt solution
4. A buffer solution
5. A mixture of a strong acid or base with any of the others (aka Invader Problem)

Let's start by making sure you can identify which type of solution is present before we review the math. For each of these solutions, identify which of the solution types (#1-5 above) are present.

1. NH_3
2. HClO_4
3. $\text{HC}_2\text{H}_3\text{O}_2$
4. NH_4Cl
5. $\text{HNO}_3 + \text{NH}_3$
6. $\text{HNO}_2 + \text{NaNO}_2$
7. NaCHO_2
8. $\text{C}_2\text{H}_5\text{NH}_2 + \text{C}_2\text{H}_5\text{NH}_3\text{Cl}$
9. HF
10. $\text{KOH} + \text{HC}_2\text{H}_3\text{O}_2$

Now, let's review the math to calculate the pH for each type of solution.

STRONG ACID OR BASE

To find the pH of a strong acid or base, take the $-\log$ of the acid or base concentration. The $-\log$ of a strong acid concentration will give the pH. The $-\log$ of a strong base concentration will give the pOH. The only reason we can use the concentration of the molarity directly is because these acids and bases are strong. Strong acids and bases dissociate 100% so for every one mole of acid or base there is one mole of hydronium ions (H^+) and hydroxide ions (OH^-) respectively.

Ex 1) Calculate the pH of 0.450M HCl

Ex 2) Calculate the pH of 0.710M KOH

WEAK ACID OR BASE

To find the pH of a weak acid or base, write a Lowry Brønsted reaction by adding the weak acid or weak base to water. Put that reaction in a RICE table. The given concentration is the “initial” concentration of the acid or base reactant. Solve for the RICE table “x” by writing the equilibrium expression, plugging in the values from the equilibrium line of the RICE table, and setting it equal to the given K_a or K_b . To find the pH/pOH, take the $-\log[x]$ since “x” represents the hydronium ion concentration in acids and the hydroxide ion concentration in bases.

Ex 3) Calculate the pH of a 0.150M solution of ammonia. The K_b of ammonia is 1.80×10^{-5} .

Ex 4) Calculate the pH of a 0.227M solution of hydrocyanic acid (HCN). The $K_a = 6.2 \times 10^{-10}$.

SALT SOLUTION (HYDROLYSIS)

To find the pH of a salt solution, first determine which ion within the salt will undergo hydrolysis. In other words, which ion from the salt will react with water to form a compound that won't significantly dissociate? A quick way to figure that out look at each ion in the salt and ask yourself, “if this is paired with an H^+ or OH^- will that substance be ‘strong’ or ‘weak’?” We're looking for the one that makes a *weak* product because that product will not re-ionize and will thus change the H^+/OH^- concentration in the solution. Once you've identified the ‘weak’ ion, write an equation for that ion splitting a water molecule and use that as the Reaction line in a RICE table. Proceed with solving for pH as you do with any other weak acid or base problem.

One other important component of hydrolysis problems is the conversion of K_a to K_b or vice versa. The question prompts given on the AP exam will provide the K_a or K_b of the *conjugate* acid or base, not the K of the ion you are using in the hydrolysis equation. The K_a or K_b of the hydrolyzing ion needs to be determined using $K_w = K_a \times K_b$. *This is a necessary step for every hydrolysis problem!*

Ex 5) Calculate the pH of a 0.95M $KC_3H_5O_3$ solution (potassium lactate).
The K_a , for lactic acid, $HC_3H_5O_3$ is 7.1×10^{-12}

BUFFER SOLUTION

Method 1: If a solution contains a weak acid and its conjugate base, or a weak base and its conjugate acid, then a buffer is present. Using the Henderson-Hasselbach (HH) equation is one way to determine pH of a buffer solution. There is a pH version and a pOH version of the equation. The acid version of this equation is provided to students on the AP Equation insert of the AP Chemistry exam. You are allowed access to that formula sheet for the duration of the test. You can also program this equation into your calculator!

$$pH = pK_a + \log \frac{[A^-]}{[HA]}$$

$$pOH = pK_b + \log \frac{[HB^+]}{[B]}$$

Using two versions of the HH equation prevents you from getting the concentrations of the acids/bases and their conjugates confused when plugging in to the log ratio. Always put the ION concentration in the numerator and the acid/base concentration in the denominator. Say "ION ON TOP!" over and over to help you remember.

*I will
Always
use 2nd
method*

$$pH = pK_a + \log \frac{[ion]}{[acid]}$$

$$pOH = pK_b + \log \frac{[ion]}{[base]}$$

Ex 6a) Calculate the pH of a solution which is 0.53M in $HC_6H_4NO_2$ and 0.50M $NaC_6H_4NO_2$. The K_a for nicotinic acid is 1.7×10^{-5} .

Ex 7a) Calculate the pH of a solution which is 0.245M in NH_3 and 0.245M in NH_4Cl . The K_b for ammonia is 1.8×10^{-5} .

Method 2: Another method for solving buffer problems using an equation similar to HH but without the logarithms. (Actually, when deriving the HH equation, this equation occurs just one step prior to taking the logarithms that generated the HH equation). In this approach we use a modified K_a expression that isolates the $[H^+]$, rather than the K_a :

$$[H^+] = K_a \frac{[acid]}{[base]}$$

So, in your buffer system, start by identifying the component that is easily recognizable as either a weak acid or base. That concentration is the one that will go in the [acid] or [base] brackets. Then, the conjugate of that weak acid/base is what will go in the remaining bracket. Always plug in the K_a and you will always be solving for the $[H^+]$. If they give you a K_b , just convert it to a K_a with $K_w = K_a \times K_b$ and then plug in the K_a . There is no need to consider a "basic" version of this equation because we are almost always interested in the $[H^+]$ even if it's a basic solution. Once you have the $[H^+]$, convert it to pH with the standard $-\log[H^+]$ and you're done.

Ex 6b) Calculate the pH of a solution which is 0.53M in $HC_6H_4NO_2$ and 0.50M $NaC_6H_4NO_2$. The K_a for nicotinic acid is 1.7×10^{-5} .

Ex 7b) Calculate the pH of a solution which is 0.245M in NH_3 and 0.245M in NH_4Cl . The K_b for ammonia is 1.8×10^{-5} .

MIXTURE OF STRONG ACID/BASE WITH ANYTHING ELSE

Treat the addition of a strong acid or base to a system already at equilibrium as an “invader.” This invasion creates a war. The war is represented by the one-way (\rightarrow) stoichiometric reaction between the invader and the system at equilibrium. Once the war is over, the stoichiometric survivors remaining will then determine the pH of the solution. This is a multistep process that will be shown in the three examples provided for you below, with each example having a different outcome.

Ex 8) 125mL of a 0.525M HCl is mixed with 125mL of a buffer system containing 1M acetic acid and 1M sodium acetate. Find the pH of the solution. The K_a of acetic acid is 1.8×10^{-5} .

Before you start, since this is a buffer, make sure you can identify the components:

The acid is _____ and the base is _____

Step 1: Write the invasion equation as the strong acid enters the buffer system. (Think: which component of the buffer will neutralize this invader? And note, this is NOT an equilibrium reaction \rightarrow reaction only)

X I used Ding Acid Invader Base all to mmol \Rightarrow then used $[H^+] = K_a \frac{\text{Acid} \pm \text{Invader}}{\text{Base} \pm \text{Invader}}$

Step 2: Figure out the stoichiometric survivors using a quick mole stoichiometry table such as:

Initial moles			
Change in moles			
Final moles			

Step 3: Based on the survivors, make a plan. Are there both [acid] and [base] components of a buffer? Or, is there just one survivor? If it's still a buffer, plug the final moles of each piece into either HH or buffer equation and solve for pH. If there is only one survivor left, find its pH appropriately as an acid, base or salt.

Ex 9) 125mL of a 0.500M NaOH is mixed with 125mL of 0.500M acetic acid. Find the pH of the solution.
The K_a of acetic acid is 1.8×10^{-5} .



Ex 10) 200mL of a 0.500M NaOH is mixed with 125mL of 0.500M acetic acid. Find the pH of the solution.
The K_a of acetic acid is 1.8×10^{-5} .