Tips on solving acid/base equilibrium problems:

The most important concept for this portion of CH302 is being able to identify what species exist in a solution, and what is left in solution if a neutralization reaction occurs (via addition of acid or base):
You'll work the following types of problems:
strong acid
strong base
weak acid
weak base
buffer
amphiprotic (amphoteric)

Try not to let the different notation (or lack thereof) in the textbook discourage you from using it as a resource. We use the notation (HA, A-, B, BH+) because it helps make things simpler for most students, but you don't have to strictly abide by it to correctly identify species as weak acids, weak bases, strong acids, etc. What you do need to do, though, is to come up w a system that works for you.

Generally speaking, acids produce H+ ions in water, and bases produce OH- ions when put in water.

STRONG ACIDS AND STRONG BASES:
You'll need to memorize these. There are 7 strong acids, and 8 strong bases. They ionize completely in water (K is infinity), and these types of problems are the easiest to solve, since [H+] = Ca and [OH-] = Cb, where C is concentration.

WEAK ACIDS:
Noted HA, BH+
Almost all acids, weak and strong, are given by "such and such acid". Examples include acetic acid, hydrochloric acid, citric acid. You'll need to know which are strong (Ka = infinity), and which are weak (all acids that are not strong are weak). Another indication of whether an acid is weak or strong is whether or not a Ka is given for the acid -- if no Ka is given for HNO3, it is because HNO3 (nitric acid) is a strong acid.
Monoprotic weak acids are generally referred to as HA. Diprotic acids are H2A (fully protonated form), triprotic acids H3A (again fully protonated form), and so on.....
The acids that are most difficult to identify are the BH+ forms of acids-- this is because they are not always given by the name " such and such acid." However, they are positively charged, so that is a helpful way to identify

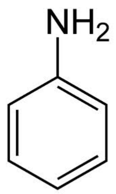
such species.

To solve a weak acid problem, we typically use the following:
[H+] = squareroot(CaKa) - this is an approximation though, and cannot be used if Ka is too large or if the solution is too dilute.

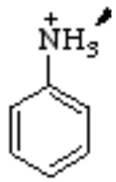
WEAK BASES:
Noted by A- and B
Identifying weak bases can also be tricky, but here are some helpful hints. The bases labeled as A-, the conjugate base of an acid HA, are anions (have negative charges), and often end in "-ate" or even "ite". Often, they are added to solution as salts that ionize when in water, such as sodium nitrite. Don't be fooled, or thrown off, by the sodium, it is just a spectator ion that is not participating in the equilibrium.
NaNO2 → Na+ + NO2-
The nitrite ion (NO2-) is free in solution, and acts as a weak base, even though the problem might not explicitly state nitrite ion.
For example, I'll list some weak acids, and their conjugate bases (HA, A-):
acetic acid, acetate ion
nitrous acid, nitrite ion
benzoic acid, benzoate ion
In a word problem, these species may be present and given as neutral salts that are strong electrolytes. If a problem asks you to find the pH of a 0.1 M solution of lithium acetate, you need to identify lithium acetate as a strong electrolyte that will ionize in water to form lithium ions and acetate ions. From there, recognize the "–ate" ending of acetate and solve a weak base problem.

Bases that are denoted by B can be recognized because they are neutral (have no charge), and are not called "acids" (weak acids HA, H2A, etc are neutral as well, but are typically called "such and such acid"). Often, they are ammonia (NH3, a weak base), or ammonia-like (replacing one hydrogen in ammonia with a carbon-containing group).

As an example, let's look at aniline (a weak base) which is ammonia where one hydrogen has been replaced with a phenyl ring. Its conjugate acid, anilinium, has one more proton than the base form.



Aniline, **B** (also known as phenylamine) – a weak base



Anilinium (also known as phenylammonium ion) **BH+** - a weak acid

What is the acid form of ammonia (NH_3)?

To solve a weak base problem, we typically use the following:
[OH-] = squareroot(CbKb) - this is an approximation though, and cannot be used if Kb is too large or if the solution is too dilute.

A note for the weak acid/weak base problem: If you are given a Kb in the problem, it doesn't mean you actually have a weak base in solution. You might have a weak acid.
For example, if you were told you had 0.5 M ammonium chloride (NH_4^+) in solution, and were given the Kb for ammonium and asked to solve for pH, you shouldn't do a Kb problem, because you have the acid form (BH+) of ammonia in solution, which is a weak acid. You need to do a Ka problem, which is fine because KaKb = Kw, so finding Ka for ammonium ion is trivial.

BUFFERS:
Buffers exist when a solution contains a weak conjugate acid base pair. This means that a weak acid, HA, and it's conjugate base, A-, are both present in the solution (the same is true for a weak base B and it's conjugate acid BH+). Without both species present, the solution will not be a buffer (without the presence of the weak acid/conjugate weak base equilibrium, the solution will not be buffered against pH changes upon addition of a strong acid or base, hence the term buffer).
For buffer systems, [H+] = Ka (Ca/Cb) and [OH-]= Kb (Cb/Ca)

If strong acid or base is added to a buffer system, a reaction will occur (a strong acid will react with the weak base, and a strong base will react with the weak acid, and stoichiometry will apply). After an addition of a strong acid or base to a buffer system some possibilities will exist for the solution:

1. both weak acid and it's weak conjugate base remain (still a buffer)
2. strong acid/base is left over (all of the weak base/acid is reacted away, do a strong acid or strong base problem)
3. Only weak acid or weak base is left
(do a weak acid or weak base problem, where you need to be concerned with additive volumes)

AMPHIPROTIC (AMPHOTERIC):
These species can act as both an acid and a base when placed in water; you will only work these types of problems when **only** the amphoteric species is in solution.

Consider the diprotic acid, H2A. It will yield only one amphoteric species, HA-.
If you determined only HA- is in the solution, [H+] = (Ka1Ka2)^0.5 = sqroot(Ka1Ka2).
However, if you have both H2A and HA- in solution, you do a buffer problem, where Ka1 is the required equilibrium constant.
If you have HA- and A2- in solution, you do a buffer problem where Ka2 is the appropriate equilibrium constant.