A Guide to Titrations and Finding the pH

Strong Acid/Strong Base:

Here, there's nothing special happening – we either have excess acid or base, unless we're at the equivalence point where the pH=7

- <u>0mL:</u> Here, you just have a solution of a strong acid with a known concentration.
  - $[HA]=[H^+]$ because it's a strong acid
  - $-\log[H^+]=pH$
- <u>Prior to Equivalence point:</u> Each mole of base neutralizes a mole of acid, but there is still an excess of acid.
  - Subtract moles of base added from the initial moles of acid and find new concentration of acid
  - Still, $[HA]=[H^+]$ because it's a strong acid
  - $-\log[H^+]=pH$
- <u>At Equivalence Point:</u> We've added sufficient base to exactly neutralize the acid, leaving just the salt present.
  - pH=7 because neither conjugate is able to interact with water to act as an acid or base
- <u>Past Equivalence Point:</u> Since all acid was neutralized at the equivalence point, any base added past that point is just making a solution of the strong base.
  - Subtract initial moles of acid from moles of base added to determine excess moles of base. Find concentration of base which is $[OH^-]$. $pOH=-\log[OH^-]$, $pH=14-pOH$

Weak Acid/Strong Base:

(Just invert pH and pOH completely throughout if you have a weak base/strong acid titration).

- <u>0mL:</u> Here, you just have a solution of weak acid with a known concentration.
  - $HA \leftrightarrow H^+ + A^-$, so $K_a = \frac{[H^+][A^-]}{[HA]}$.
  - Since all H^+ and A^- come from the dissociation of HA, $[H^+]=[A^-]$
  - If we set both as x, $K_a = \frac{x^2}{[HA]-x} \approx \frac{x^2}{[HA]}$ and pH=-log(x)
- <u>Prior to equivalence point:</u> Each mole of base neutralizes a mole of acid – creating a mole of base – but there is still an excess of acid.
  - mol A⁻= (mol OH⁻)ₐ_dded
  - mol HA = (mol HA)ᵢₙᵢₜᵢₐₗ – (mol OH⁻)ₐ_dded
  - Use Henderson-Hasselback here $pH = pK_a + \log\left(\frac{mol\ A^-}{mol\ HA}\right)$
    - HH on the equation sheet says concentration, but since they're always in the same solution, they have the same volume and that part cancels out!
- <u>Half-equivalence point:</u> When the volume of base added is exactly half of what is needed to neutralize the acid initially present. This means that half of the acid has been turned to salt, yielding an equal number of moles of acid and salt ([A⁻]=[HA] and mol A⁻=mol HA).
  - $pH = pK_a + \log\left(\frac{mol\ A^-}{mol\ HA}\right) = pK_a + \log(1) = pK_a$
  - At this point, pH ALWAYS equals pK_a. This is important enough to remember because it comes up all the time.
- <u>Equivalence Point:</u> The equivalence point is the conjugate base's one moment of glory. Here, we have a solution of just the conjugate base. Find the concentration of it by dividing the moles of salt (equal to the initial moles of acid) and dividing by the new volume. Then you just have a solution of a weak base with a known concentration.
  - Find K_b of the conjugate using $K_b = \frac{10^{-14}}{K_a}$
  - Then find new concentration of that salt by $[A^-] = \frac{mol\ HA_{initial}}{V_{new}}$
  - $A^- + H_2O \leftrightarrow HA + OH^-$, so $K_b = \frac{[HA][OH^-]}{[A^-]}$.
  - Since all HA and OH⁻ come from the reaction of A⁻ with water, [HA]=[OH⁻]
  - If we set both as x, $K_b = \frac{x^2}{[A^-]}$ and pOH=-log(x)
  - Subtract 14-pOH=pH
- <u>Past Equivalence Point:</u> Since all acid was neutralized at the equivalence point, any base added past that point is just making a solution of the strong base.
  - Subtract initial moles of acid from moles of base added to determine excess moles of base. Find concentration of base. pOH=-log[OH⁻], pH=14-pOH