SUMMARY #1 TITRATION CALCULATIONS SUMMARY

Any time you add a strong acid or strong base, set up a MOLE TABLE. The results of the mole table should tell you exactly what step to do next. There are three possible scenarios:

<u>Scenario A</u>: Before you've reached the equivalence point.

mole table:	HA	+	NaOH	------>	H_2O	+	A^-
final results:	x		0		-		y

** *Since you have a concentration of both HA and A^- left at the end of the ice table, you have a common ion or buffer situation. In a titration, you would be in the midst of the buffer region. So, to see how the system will reestablish equilibrium after the addition of a strong acid or strong base, do an ICE table or H.H. Eq.*

	HA	H^+	+	A^-
I	x			y
C				
E				

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

<u>Scenario B</u>: At the equivalence point.

mole table:	HA	+	NaOH	------>	H_2O	+	A^-
final results:	0		0		-		x

** *Since you have no original HA and no extra NaOH, the only species present is the conjugate base. The conjugate base will HYDROLYSIS. For weak acid or weak base titrations, at the equivalence point, the conjugate partner will always hydrolyze. For a weak acid, the pH at the equivalence point will always be greater than 7 and for a weak base, the pH at the equivalence point will be below 7.*

	A^-	+	H_2O	HA	+	OH^-
I	y		-	0		0
C						
E						

** *Don't forget to use the appropriate K value. If you make OH^-, use K_b. If you make H_3O^+, use K_a!!*

<u>Scenario C</u>: Beyond the equivalence point

mole table:	HA	+	NaOH	------>	H_2O	+	A^-
final results:	0		x		-		y

** *At this point, you've neutralized the acid and now have EXTRA strong base. At this point, you also have some A^- left too which will hydrolyze. But, since their is OH^- present as well, the A^- will hydrolyze to a very small extent and will not significantly affect the pH. It is the left over strong base that will primarily determine the pH. Take the - log of it's concentration and get the pH.*

$pOH = -\log [OH^-] = -\log [x]$ and $pH = 14 - pOH$