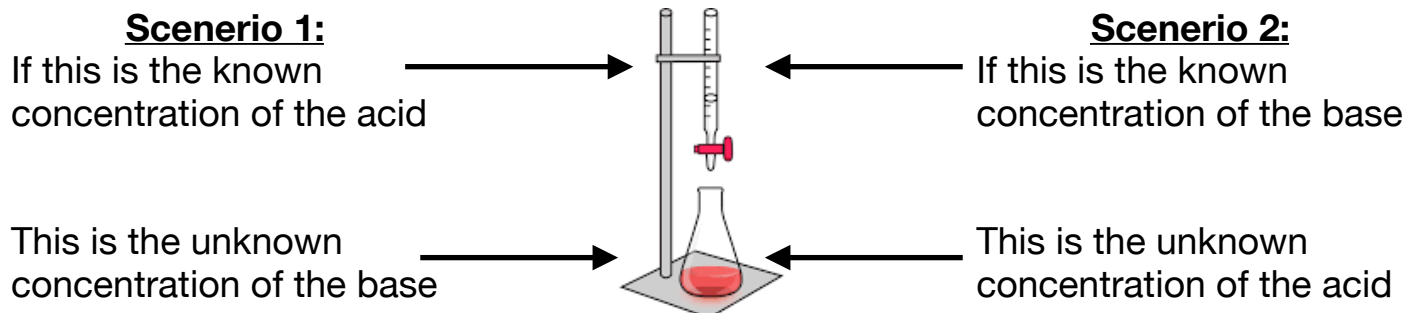
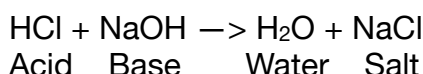


Titration:

- Titrations are used to determine the concentration of an unknown acid or base.
- When we do a titration the unknown concentration goes in a flask called the Analyte and the known concentration goes in a burette called a Titrant.



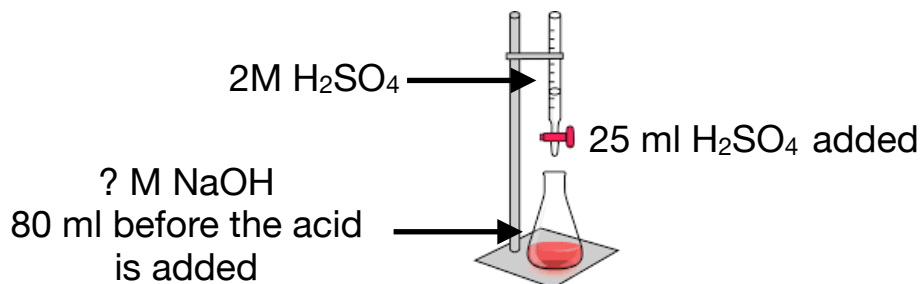
- To determine the concentration of the unknown we will keep track of how much of the acid or base we are adding from the burette into the flask until the moles of acid and moles of base are equal to each other (This is only for strong acids mixed with strong bases - we will talk about weak acids and bases later).
- We will know when the moles of the acid and base are equal to each other because a color change will happen that will be from a few drops of an indicator we will add to the flask. When the moles of acid are equal to the moles of base it is called being at the **equivalence point**.
- Anytime an acid and base are mixed together they will make a **salt** and **water**.



Example-1:

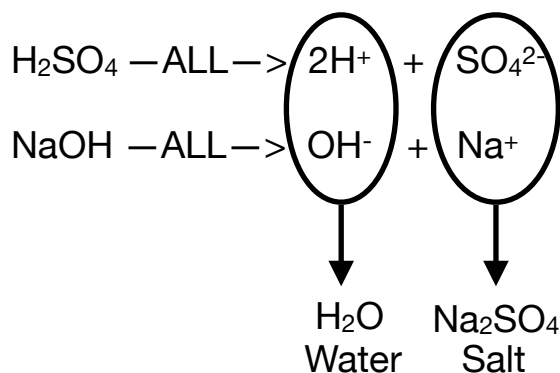
When 80 ml of NaOH is titrated with 2M, 25 ml of a H₂SO₄ solution is required to reach the equivalence point. Find the concentration of NaOH.

— Lets make a picture to see what is going on here —



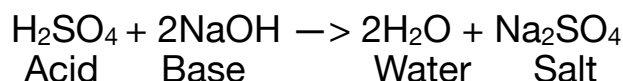
— You can think of what is going on in terms of the chemical reaction that is taking place in one of three ways. —

1st way (Arrhenius Acid/Base):



BEWARE: Remember to cross charges to get the chemical formula for Na_2SO_4 .

3rd way (A Single Reaction):



— Let's use the 3rd way to do this question. At the equivalence point the moles of acid are equal to the moles of base. Therefore we will use the moles of the acid to convert to moles of base using the balanced chemical reaction. —

$$(2 \text{ moles/l}) (0.025 \text{ l}) = 0.05 \text{ moles H}_2\text{SO}_4$$

$$(0.05 \text{ moles H}_2\text{SO}_4) (2 \text{ moles NaOH} / 1 \text{ mole H}_2\text{SO}_4) = 0.1 \text{ mole NaOH}$$

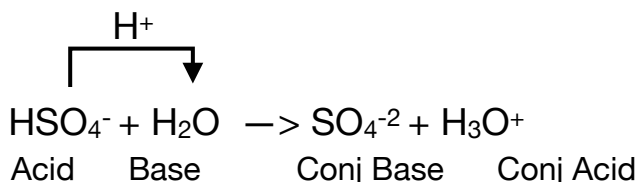
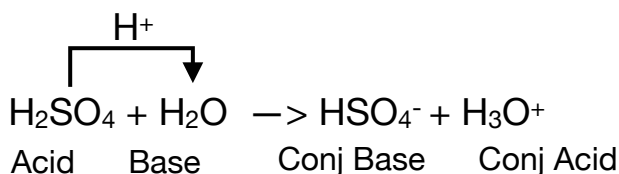
— Now that we know we had 0.1 mole of NaOH in the flask (before the acid was added) we divide by the volume of NaOH we had in the flask (in liters), which was 80 ml. —

$$0.1 / .08 = 1.25 \text{ M}$$

BEWARE: This is the concentration of NaOH before we added the acid. Once the acid is added it will decrease the concentration of NaOH until it is all gone and neutralized by the acid resulting in a pH of 7.

— Sometimes H_2SO_4 will only donate one Hydrogen instead of two. Whether H_2SO_4 will donate two hydrogens or one depends on the pH of the solution due to comparing the K_w of water and the second K_a of the acid (each H in H_2SO_4 has it's own K_a value).—

2nd way (conjugated Acid/Base):



BEWARE: HSO_4^- is a conj base, but it's K_a value is larger than water, which causes it to be an acid in water. Most of the time polyprotic acids only donate 1 Hydrogen not 2. Bases with 2 OH will almost always donate 2 OH like Ca_2OH .

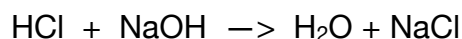
Example-2

When 15 ml of a HCl solution is titrated with 3M, 20 ml of the NaOH solution is required to reach the equivalence point. Find the concentration of HCl.

— First find the number of moles of NaOH it took to neutralize all of the acid —

$$(3)(0.02) = 0.06 \text{ moles of NaOH}$$

— Now write out the reaction —



$$(0.06 \text{ moles NaOH}) (1 \text{ mol HCl} / 1 \text{ mol NaOH}) = 0.06 \text{ moles of HCl}$$

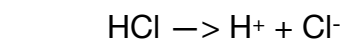
— Now that we know we had 0.06 mole of HCl in the flask (before the base was added) we divide by the volume of NaOH we had in the flask (in liters), which was 15 ml. —

$$0.06 \text{ moles HCl} / 0.015 \text{ l} = 4 \text{ M}$$

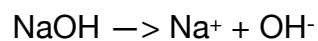
Example-3

If there is 50 ml of a 0.1 M HCl in a flask determine the pH of the **solution** when 0 ml, 400 ml, 500 ml, and 800 ml of 0.02 M NaOH is added.

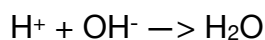
— This question ask for the pH at various points during a titration, so let's separate each part of the question into an A, B, C, and D. We will also use the following set of reaction for each part in the question —



Use this to find out how much H^+ is made



Use this to find out how out how much OH^- is made



Use this to find out how much water can be made by neutralization

— The equations above come from the *Example-1*. Although it is possible to use the 3rd way to answer this question, using the 1st way is a little easier to work with. —

— In this question the hydrogen concentration will decrease (pH increase) for two reasons. The first reason is because when the two solutions are mixed together the amount of water goes up, which is part of the solvent, that causes the concentration to go down. The second reason is because when moles of OH^- are added to moles of H^+ , some of the moles of H^+ are removed through a neutralization reaction.

PART A

— When 0 moles of a bases are added to the flask the concentration of the acid is 0.1 M, so to find the pH use $\text{pH} = -\log [\text{H}]$ —

$$\text{pH} = -\log (0.1) \quad \text{pH} = 1$$

PART B

— Once the base is added to the acid it will start to neutralize the acid. To find the pH we need to know how much acid will be left over after the base that was added. This involves finding limiting and excess reactant, but given the equation $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$ everything is a 1:1 ratio so we can just subtract the hydrogen moles by the hydroxide moles to find the limiting reactant. —

— First find how many moles of H we have from HCl —

$$(0.1 \text{ moles HCl} / 1) (0.05 \text{ L}) = 0.005 \text{ moles HCl}$$

$$(0.005 \text{ moles HCl}) (1 \text{ mol H} / 1 \text{ mol HCl}) = 0.005 \text{ moles H}$$

— Now find how many moles of OH we have from NaOH —

$$(0.01 \text{ moles NaOH} / 1) (0.4 \text{ L}) = 0.004 \text{ moles NaOH}$$

$$(0.004 \text{ moles NaOH}) (1 \text{ mol OH} / 1 \text{ mol NaOH}) = 0.004 \text{ moles OH}$$

— Now subtract the moles of H by the moles of OH to see how many moles of H will be left over —

$$0.005 \text{ moles H} - 0.004 \text{ moles OH} = 0.001 \text{ moles H}$$

— Now divide by the total liters of solution and find the pH using $\text{pH} = -\log [\text{H}]$ —

$$0.05 \text{ L from HCl} + 0.4 \text{ L from NaOH} = 0.45 \text{ L of solution}$$

$$0.001 / 0.45 = 0.0022$$

$$\text{pH} = -\log [0.0022]$$

$$\text{pH} = 2.63$$

PART C

— Just like in part B we will find the moles of H or OH left over —

$$(0.1 \text{ moles HCl} / 1) (0.05 \text{ L}) = 0.005 \text{ moles HCl}$$

$$(0.005 \text{ moles HCl}) (1 \text{ mol H} / 1 \text{ mol HCl}) = 0.005 \text{ moles H}$$

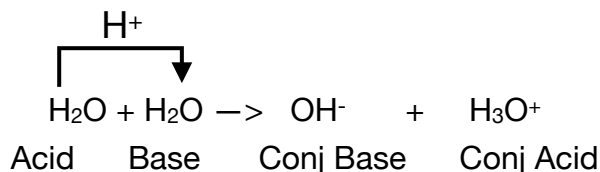
$$(0.01 \text{ moles NaOH} / 1) (0.5 \text{ L}) = 0.005 \text{ moles NaOH}$$

$$(0.004 \text{ moles NaOH}) (1 \text{ mol OH} / 1 \text{ mol NaOH}) = 0.005 \text{ moles OH}$$

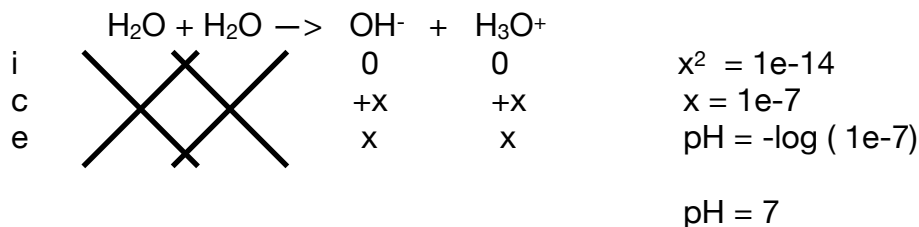
$$0.005 \text{ moles H} - 0.005 \text{ moles OH} = 0 \text{ moles H}$$

— This means that there was not a limiting reactant and both reactant are using up equally to make H_2O . —

Water can react with itself as an acid and base to form a conj acid/base in the following reaction.



In part C of this question there is nothing left but water and the K_w for water is $1\text{e-}14$, so do an ICE table to find the concentration of H^+ (H^+ is the same as H_3O^+).



BEWARE: The K_w for water is $1\text{e-}14$ at 25°C . If the temperature changes the K will change and the pH will change.

PART D

$$(0.1 \text{ moles HCl} / 1) (0.05 \text{ L}) = 0.005 \text{ moles HCl}$$

$$(0.005 \text{ moles HCl}) (1 \text{ mol H} / 1 \text{ mol HCl}) = 0.005 \text{ moles H}$$

$$(0.01 \text{ moles NaOH} / 1) (0.8 \text{ L}) = 0.008 \text{ moles NaOH}$$

$$(0.004 \text{ moles NaOH}) (1 \text{ mol OH} / 1 \text{ mol NaOH}) = 0.008 \text{ moles OH}$$

$$0.008 \text{ moles OH} - 0.005 \text{ moles H} = 0.003 \text{ moles OH}$$

—Remember the total liters of the solution is now 850 ml or 0.85 l —

$$0.003 \text{ moles OH} / 0.85 \text{ l} = 0.0035 \text{ OH}$$

—Because there is an excess amount of OH and not H we can find the pOH concentration. —

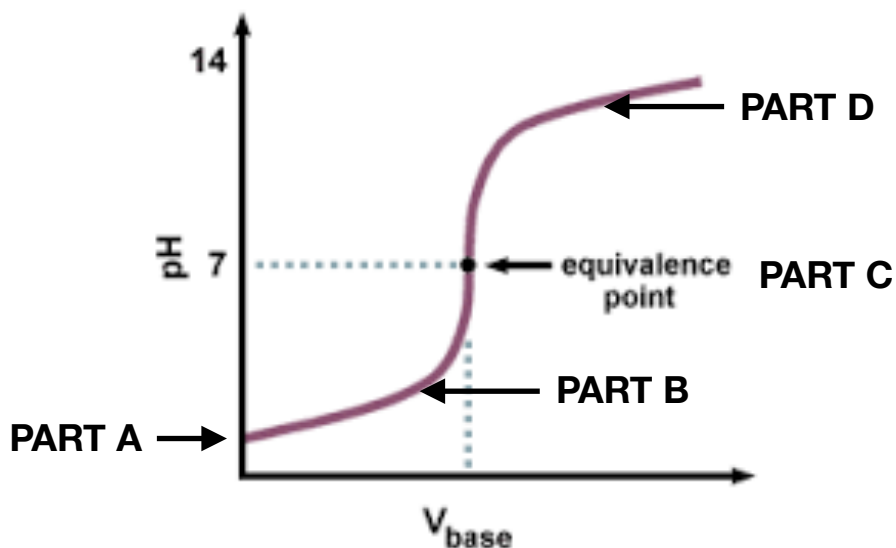
$$\text{pOH} = -\log (0.0035)$$

$$\text{pOH} = 2.45$$

$$\text{pH} + \text{pOH} = 14$$

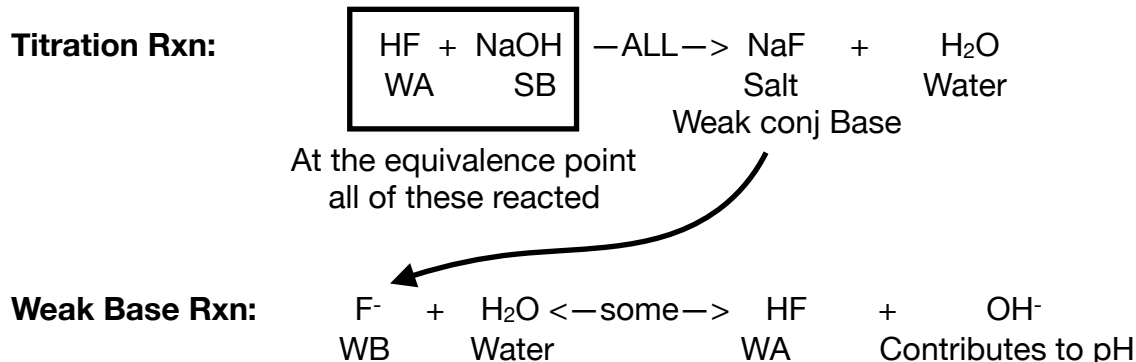
$$\text{pH} = 11.56$$

We can make a graph of *Example-3* for how the concentration of acid changes as more NaOH was added to it. If we make a graph it will look something like this:



- From looking at the graph you can see how the pH increases as more base is added to the flask. When the pH is below 7 we have excess amount of hydrogen and then the pH is above 7 we have excess amount of hydroxide (OH).
- When doing this in a lab a color change will happen when the solution goes from being acidic (below 7) to being basic (above 7). Notice that the curve shoots up very quick as it moves from an acid to a base and it is not linear.

- When a weak acid and a strong base are mixed together the equivalence point does not occur at a pH of 7. This is because a weak acid will end up producing a weak base and that newly formed weak base will contribute to the pH.
- Consider the following weak acid strong base titration:



- To figure how much OH is made from the F weak base we need to set up an ICE table.
- Weak bases titrated with strong acid will have a pH that is lower than 7 at the equivalence point. Weak acids titrated with a strong base will have a pH that is greater than 7 at the equivalence point.

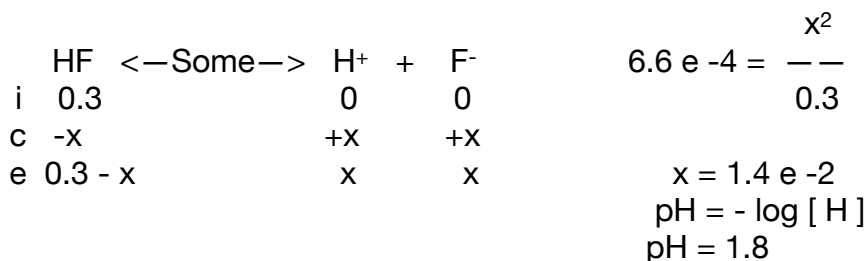
Example-4

25 ml of a 0.3 M solution of HF is titrated with 0.3 M solution of NaOH. Find the pH after 0 ml, 10 ml, 12.5 ml, 25 ml, and 26 ml of NaOH has been added to HF.
 $K_a = 6.6 \times 10^{-4}$

— This question involves mixing a certain amount of base into a weak acid. So let's separate this question into 5 parts. Part A, B, C, D, and E. —

PART A

— For the first part of this question There is only HF and no NaOH. To find the pH set up an ICE table for the weak acid. —



PART B

— If we have two solution being mixed together, then the concentrations of the acid and base will change. This is similar to a dilution problem. So first we will find the new concentration of both HF and NaOH. —

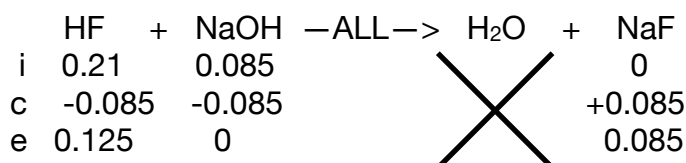
$$(0.3 \text{ moles} / 1 \text{ l}) (0.025 \text{ l}) = 0.0075 \text{ moles HF}$$

$$(0.3 \text{ moles} / 1 \text{ l}) (0.01 \text{ l}) = 0.003 \text{ moles NaOH}$$

$$(0.0075 \text{ moles HF}) / (0.035 \text{ l}) = 0.21 \text{ M HF}$$

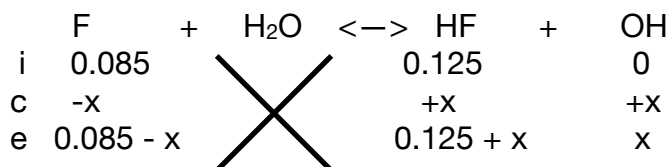
$$(0.003 \text{ moles NaOH}) / (0.035 \text{ l}) = 0.085 \text{ M NaOH}$$

— Now we will set up an acid/base reaction of NaOH and HF. —



— The way you know the change is equal to 0.085 is because this reaction goes to completion, which means either the HF or NaOH will run out and because NaOH is the smaller concentration it will run out first.

Now we need to set up a second reaction that involves HF and NaF to see how those will contribute to the pH of the solution. —



— NaF is a basic salt and Na is a spectator ion. You do not have to write spectator ions in the reaction if you do not want to. Notice that the problem gives us the K_a for HF and we have a reaction with F. Therefore we need to convert from a K_a to a K_b by dividing 1e-14 by 6.6 e -4. —

$$1.51 \text{ e}^{-11} = \frac{(0.125 + x)(x)}{(0.085 - x)}$$

To make the math work out easier you can get rid of the +x and -x because the K is very small.

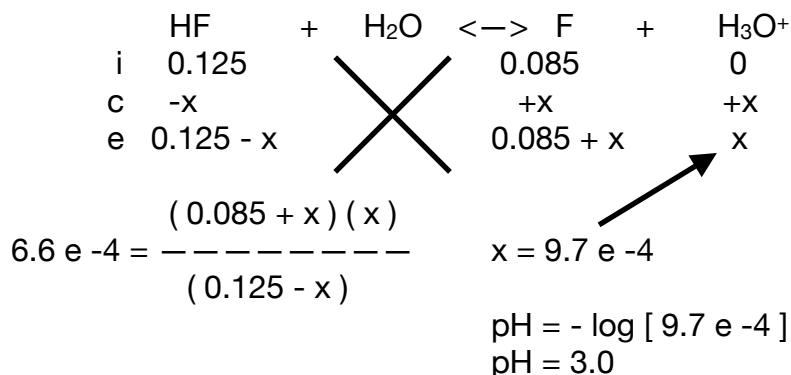
$$x = 1.03 \text{ e}^{-11}$$

$$\text{pOH} = -\log [1.03 \text{ e}^{-11}]$$

$$\text{pOH} = 10.98$$

$$\text{pH} = 3.0$$

— You can also do this question without having to convert from a K_a to a K_b . From the first ICE table in part B we have NaF and HF. So if we write HF as a reactant we can use K_a instead of K_b . You will see later that sometimes we have no choice but to use a K_b , but for part B we can use either. —



PART C

—For part C we need to find the new concentration all over again just like part B. —

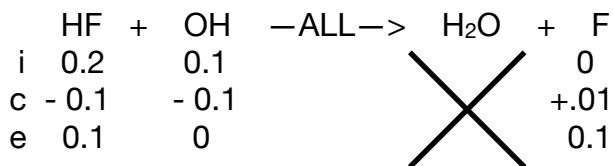
$$(0.3 \text{ moles} / 1 \text{ l}) (0.025 \text{ l}) = 0.0075 \text{ moles HF}$$

$$(0.3 \text{ moles} / 1 \text{ l}) (0.0125 \text{ l}) = 0.00375 \text{ moles NaOH}$$

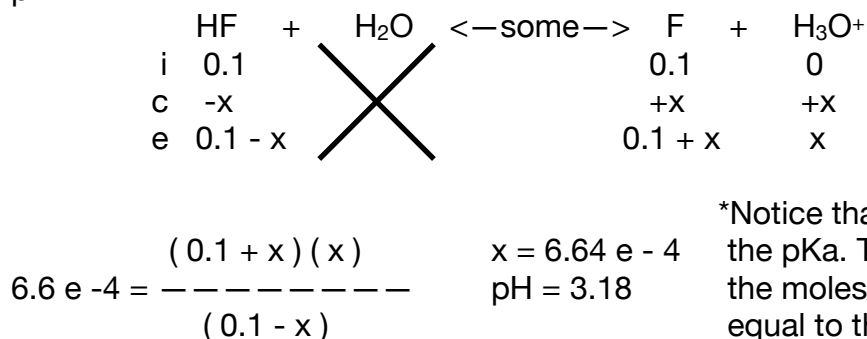
$$(0.0075 \text{ moles HF}) / (0.0375 \text{ l}) = 0.2 \text{ M HF}$$

$$(0.00375 \text{ moles NaOH}) / (0.0375 \text{ l}) = 0.1 \text{ M NaOH}$$

—Now we will set up an ICE table. —



—Now that all of the NaOH has been neutralized by HF, set up a reaction where HF can act as an Arrhenius acid to produce a conjugate base that will contribute to the pH.—



Notice that the pH is equal to the pKa. This always occurs when the moles of the weak acid are equal to the moles of conj base.

PART D

— Just like the parts B and C find the concentration of HF and NaOH. —

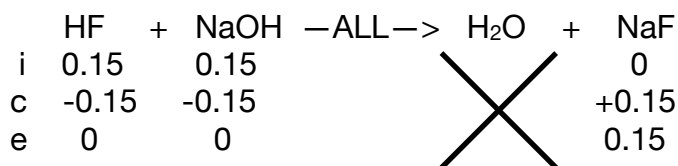
$$(0.3 \text{ moles} / 1 \text{ l}) (0.05 \text{ l}) = 0.0075 \text{ moles HF}$$

$$(0.3 \text{ moles} / 1 \text{ l}) (0.05 \text{ l}) = 0.0075 \text{ moles NaOH}$$

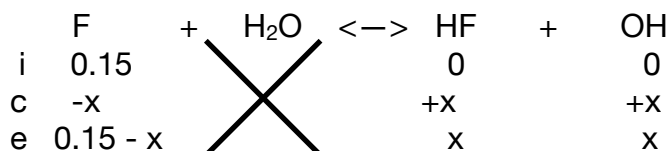
$$(0.0075 \text{ moles HF}) / (0.05 \text{ l}) = 0.15 \text{ M HF}$$

$$(0.0075 \text{ moles NaOH}) / (0.05 \text{ l}) = 0.15 \text{ M NaOH}$$

— Notice that the moles of HF and NaOH are equal. This is at the equivalence point for this titration. Just like before set up an ICE table. —



— The only chemical that now contribute to the pH is the basic salt NaF, so write a chemical reaction for NaF (or just F).



— Notice that we can not use a K_a here because we ran out of HF when all of it reacted with NaOH, so this time you have to convert to a K_b . —

$$1.51 \text{ e } -11 = \frac{x^2}{(0.15 - x)}$$

$$x = 1.5 \text{ e } -6$$

$$\text{pOH} = -\log [1.5 \text{ e } -6]$$

$$\text{pOH} = 5.82$$

$$\text{pH} = 8.17$$

— Notice that at the equivalence the solution is basic. —

PART E

—Just like parts B, C, and D find the number of moles of HF and NaOH.—

$$(0.3 \text{ moles} / 1 \text{ l}) (0.025 \text{ l}) = 0.0075 \text{ moles HF}$$

$$(0.3 \text{ moles} / 1 \text{ l}) (0.026 \text{ l}) = 0.0078 \text{ moles NaOH}$$

$$(0.00785 \text{ moles HF}) / (0.051 \text{ l}) = 0.147 \text{ M HF}$$

$$(0.0078 \text{ moles NaOH}) / (0.051 \text{ l}) = 0.152 \text{ M NaOH}$$

—Now make an ICE table for HF and NaOH.—

	HF	+	NaOH	—ALL—>	H ₂ O	+	NaF
i	0.147		0.152				0
c	-0.147		-0.147				+0.147
e	0		0.005				0.147

—This time all of the weak acid was neutralized by NaOH and we have excess NaOH, but we still have a basic salt, NaF. So set up another ICE table that includes NaF (or F) and NaOH. Because this ICE table has F as a reactant we need to use Kb instead of Ka.—

	F	+	H ₂ O	<—>	HF	+	OH
i	0.147				0		0.005
c	-x				+x		+x
e	0.147 - x				x		0.005 + x

$$1.51 \text{ e}^{-11} = \frac{(0.005 + x)x}{(0.147 - x)}$$

$$x = 4.4 \text{ e}^{-10}$$

0.005 + a small is a small number is still 0.005

$$\begin{aligned} \text{OH} &= [0.005 + 4.4 \text{ e}^{-10}] \\ \text{pOH} &= -\log [0.005] \\ \text{pOH} &= 2.3 \\ \text{pH} &= 11.7 \end{aligned}$$

- Now that we worked out all of the parts of *Example-4* let's see what this would look like if we graphed it.

