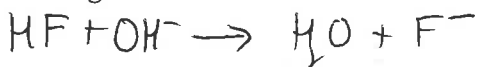


Case #2

Weak acid + Strong base

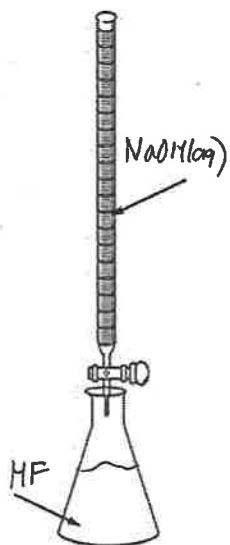
A 37.50 mL sample of 0.200 M HF  
( $K_a = 6.9 \times 10^{-4}$ ) is titrated with 0.125 M NaOH.

(a) Write the net ionic reaction that takes place during the titration.



(b) Calculate the volume of NaOH needed to completely neutralize the HF.

same as #1: 60.0 mL  
or 0.060 L

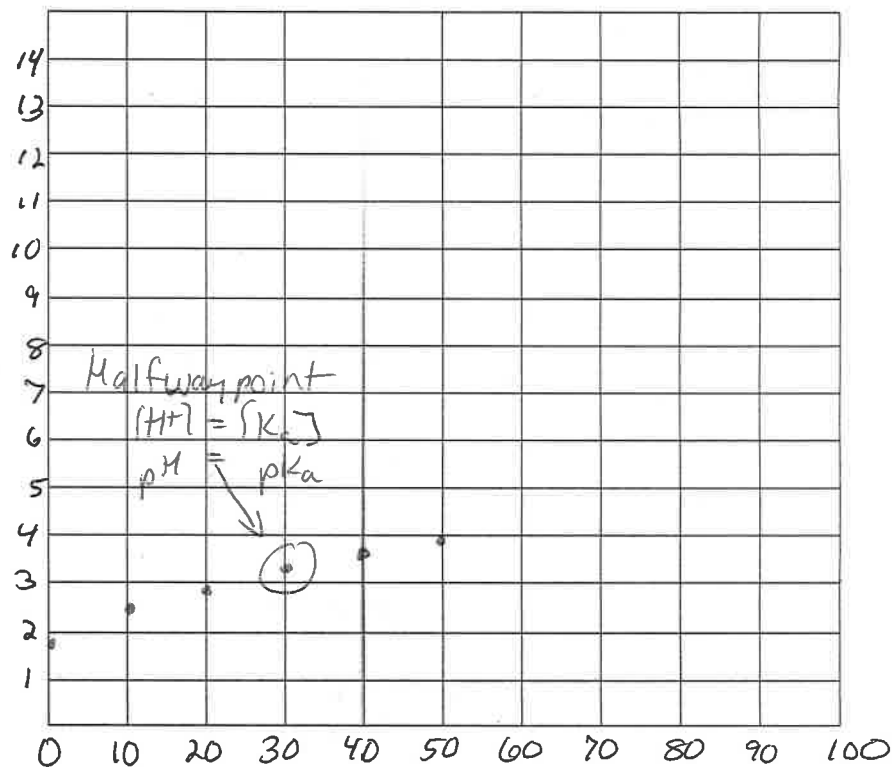


Use the following chart to track the pH of the HCl solution as the titration proceeds. Plot the pH on the graph provided.

Volume NaOH added	Moles OH-added	Moles HF remaining	Moles F <sup>-</sup> formed (from reaction)	[H <sup>+</sup> ]	pH
0	0	0.0075	0	$0.0117 \text{ M}$	1.93
10	0.00125	0.00625	0.00125	0.00345	2.46
20	0.0025	0.005	0.0025	0.00138	2.86
30	0.00375	0.00375	0.00375	$[\text{H}^+] = K_a!$	3.16
40	0.005	0.0025	0.005	$3.45 \times 10^{-4}$	3.46
50	0.00625	0.00125	0.00625	$1.38 \times 10^{-4}$	3.86
* 60	0.0075	0	0.0075		
70 excess OH <sup>-</sup>	0.0125				
75 excess	<del>0.0125</del>				
80 excess	0.025				
85					
90	0.0375				
100					

\*\* At the endpoint, solution is now a weak base solution.

pH vs. volume NaOH added:



$$K_a = \frac{[\text{H}^+][\text{F}^-]}{[\text{HF}]}$$

$$[\text{H}^+] = K_a \cdot \frac{\text{moles HF}}{\text{moles F}^-}$$

↑  
During titration

Here you can use moles instead of molarity

\* At the beginning, we  $K_a$  & behavior equation to find pH

$$6.9 \times 10^{-4} = \frac{x^2}{0.200 - x}$$

↑  
~0.200

$x = 0.0117 \text{ M}$