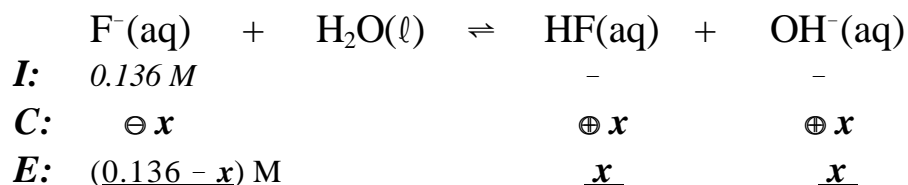


**PART A  
MC #10**

Given: 0.136 M NaF(aq), a *weak base* ( $\text{Na}^+$  is neutral and  $\text{F}^-$  is *basic*).  
 $K_a(\text{HF}) = 3.5 \times 10^{-4}$

To find: pH



$$K_c = \frac{[\text{HF}][\text{OH}^-]}{[\text{F}^-]} = K_b(\text{F}^-) = \frac{K_w}{K_a(\text{HF})}$$

$$\frac{(x)(x)}{0.136 - x} = \frac{1.0 \times 10^{-14}}{3.5 \times 10^{-4}} = 2.857 \times 10^{-11}$$

1<sup>st</sup> approximation:  $1.971 \times 10^{-6} \text{ M}$   
 2<sup>nd</sup> approximation:  **$1.971 \times 10^{-6} \text{ M}$**

$$x = \underline{1.971} \times 10^{-6} \text{ M} = [\text{OH}^-]$$

**pOH = -log [OH<sup>-</sup>]**

$$\text{pOH} = -\log(1.971 \times 10^{-6}) =$$

$$5.705$$

$$\therefore \text{pH} = 14.00 - \text{pOH} = 14.00 - 5.705 = 8.295 = \underline{\underline{8.30}}$$

**PART A  
MC #11**

Given: ① 1.0 M of HCOOH(aq) (a *weak acid*)  
 $\therefore$  **initial total** [HCOOH]  
 $=$  **dissociated** [H<sup>+</sup>] + **undissociated** [HCOOH] = 1.0 M

② pH = 1.88  
 $\therefore$  [H<sup>+</sup>] =  $10^{-\text{pH}} = 10^{-1.88} = \underline{1.318} \times 10^{-2} \text{ M}$

$$\% \text{ ionization of HCOOH} = \frac{\text{dissociated } [\text{H}^+]}{\text{initial total } [\text{HCOOH}]} \times 100\%$$

$$= \frac{1.318 \times 10^{-2} \text{ M}}{1.0 \text{ M}} \times 100\% = 1.318\% = 1.3\%$$

**PART A  
MC #12**

**strongest acid** (*i.e.* most  $\text{H}_3\text{O}^+$  produced in water)  
 → *highest*  $K_a$  (*i.e.* *largest* acid dissociation power)  
 → ∴ *smallest*  $\text{p}K_a$  ( $\text{p}K_a = -\log K_a$ )

- (a)  $\text{HClO}_2$  ( $\text{p}K_a = 2.00$ ) (∴ *smallest*  $\text{p}K_a$  → *largest*  $K_a$  → **strongest acid**)
- (b) HF ( $\text{p}K_a = 3.46$ )
- (c)  $\text{HNO}_2$  ( $\text{p}K_a = 3.34$ )
- (d) HCN ( $\text{p}K_a = 9.32$ )
- (e)  $\text{CH}_3\text{COOH}$  ( $\text{p}K_a = 4.75$ )

**Decreasing order of acidity:**

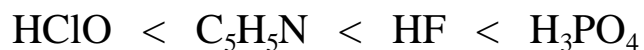


**PART A  
MC #13**

**highest pH** → *most basic*  
 → *smallest*  $[\text{H}_3\text{O}^+]$   
 → *lowest*  $K_a$   
 → ***largest*  $\text{p}K_a$**

- (a)  $\text{C}_5\text{H}_5\text{N}$  ( $\text{p}K_b = 8.74$ ) (∴  $\text{p}K_a(\text{C}_5\text{H}_5\text{NH}^+) = 5.26$ )
- (b) HF ( $\text{p}K_a = 3.46$ )
- (c)  $\text{H}_3\text{PO}_4$  ( $\text{p}K_a = 2.12$ )
- (d)  $\text{HClO}$  ( $\text{p}K_a = 7.52$ ) (∴ *largest*  $\text{p}K_a$ )
- (e) Since all are acids, the pH is the same for all solutions.

**Increasing order of acidity:**



**PART A**  
**MC #14**

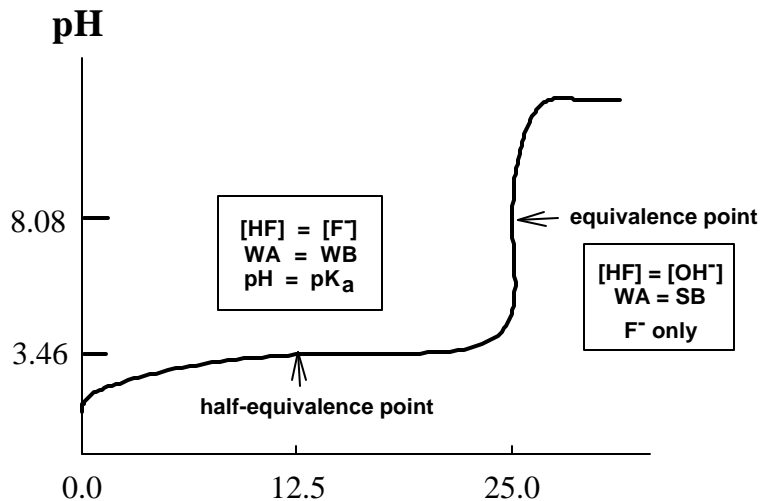
Given: 25.0 mL of 0.100 M HF(aq) (a *weak acid*)  
12.5 mL of 0.100 M NaOH(aq) (a *strong base*)  
 $\text{pK}_a(\text{HF}) = 3.46$

To find: major species present in solution

**WA + SB Titration**

<b>NIE:</b>	HF(aq)	+	OH <sup>-</sup> (aq)	→	F <sup>-</sup> (aq)	+	H <sub>2</sub> O(l)
<b>I:</b>	25.0 mL		12.5 mL		-		
	0.100 M		0.100 M		-		
	$2.50 \times 10^{-3} \text{ mol}$		$1.25 \times 10^{-3} \text{ mol}$		-		
<b>C:</b>	⊖ $1.25 \times 10^{-3} \text{ mol}$		⊖ $1.25 \times 10^{-3} \text{ mol}$		⊕ $1.25 \times 10^{-3} \text{ mol}$		
<b>E:</b>	<u><math>1.25 \times 10^{-3} \text{ mol}</math></u>		-			<u><math>1.25 \times 10^{-3} \text{ mol}</math></u>	

- ① **1 mole** HF(aq) reacts with **1 mole** NaOH(aq)  
∴ NaOH is the *limiting reactant*, and is being used up during the titration.  
50% of HF(aq) is converted into F<sup>-</sup>(aq) by the addition of OH<sup>-</sup>(aq).
- ② It reaches **half-neutralization** (or half-equivalence point).  
(i.e.  $n(\text{HF})$  *leftover* =  $n(\text{F}^-)$  *produced* =  $1.25 \times 10^{-3} \text{ mol}$  and  $\text{pH} = \text{pK}_a(\text{HF}) = 3.46$ )
- ③ It is a **1:1 buffer**. The major species present are:  
**HF(aq), F<sup>-</sup>(aq) and Na<sup>+</sup>(aq)**



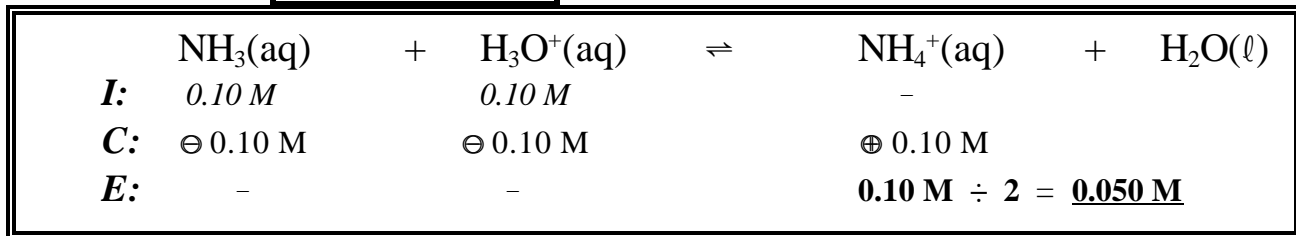
- ① **Before** equivalence point,  $\text{OH}^- (\text{aq})$  is the *limiting reactant*.
- ② **After** equivalence point,  $\text{OH}^- (\text{aq})$  is *in excess*, and  $\text{OH}^- (\text{aq})$  is one of the **major species**.
- ③  $\text{Na}^+ (\text{aq})$  is a spectator ion which is always present in the solution.

0.100 M NaOH (mL)

**PART A**  
**MC #15**

Given:  $[\text{NH}_3(\text{aq})] = 0.10 \text{ M}$  (a weak base)  
 $[\text{HCl}(\text{aq})] = 0.10 \text{ M}$  (a strong acid)

**WA + SB**  
**Titration**



0.10 M  $\text{NH}_3(\text{aq})$  reacts completely with 0.10 M  $\text{H}_3\text{O}^+(\text{aq})$  to produce 100%  $\text{NH}_4^+(\text{aq})$  with concentration of  $(0.10 \text{ M} \div 2) = 0.050 \text{ M}$ .

**At half-equivalence point** (i.e. half-neutralization),

- (1)  $n(\text{NH}_3) = n(\text{NH}_4^+)$  (i.e. WB = WA)
- (2)  $\text{pH} = \text{pK}_a = 9.26$
- (3) 1:1 buffer
- (4) major species:  $\text{NH}_3(\text{aq})$ ,  $\text{NH}_4^+(\text{aq})$ ,  $\text{Cl}^-(\text{aq})$

**At equivalence point** (i.e. complete neutralization),

- (1)  $n(\text{NH}_3) = n(\text{H}^+)$  (i.e. WB = SA)
- (2) all  $\text{NH}_3(\text{aq})$  is used up, only  $\text{NH}_4^+(\text{aq})$  produced *theoretically*
- (3)  $[\text{NH}_4^+] = 0.10 \text{ M} \div 2$  ( $\because$  volume double)
- (4) pH at equivalence point = 5.28
- (5) major species: 100%  $\text{NH}_4^+(\text{aq})$ ,  $\text{Cl}^-(\text{aq})$

(a) a solution of pH = 7.00

UNTRUE

(∴ Only a strong acid reacts with a strong base will give final pH of solution at 7.00.)

(b) a 0.10 M solution of  $\text{NH}_4\text{Cl}$  **UNTRUE**

(∴ **At equivalence point**,  $[\text{NH}_4\text{Cl}]$  is 0.050 M because the volume is doubled.

(c) **a 0.050 M solution of  $\text{NH}_4\text{Cl}$**  **TRUE**

(∴ The volume is doubled and the concentration is halved.)

(d) a 1:1 buffer solution **UNTRUE**

(∴  $[\text{NH}_3] \neq [\text{NH}_4^+]$  (i.e.  $\text{WB} \neq \text{WA}$ ), it is not at half-equivalence point.)

Since  $[\text{NH}_3] = [\text{H}^+]$  (i.e.  $\text{WB} = \text{SA}$ ), **all**  $\text{NH}_3(\text{aq})$  are *theoretically* converted into  $\text{NH}_4^+(\text{aq})$ .

(e) a solution containing equal concentrations of  $\text{NH}_3$  and  $\text{NH}_4\text{Cl}$  **UNTRUE**

(∴  $[\text{NH}_3] = [\text{NH}_4^+]$  *only* occurs at half-equivalence point.)