Answer:

1) A 68.0 mL volume of 0.25 M HBr is titrated with 0.50 M KOH. Calculate the pH after addition of 34.0 mL of KOH at 25 ∘C.

In neutralization reactions, a base and an acid react to form an ionic salt and water. There is a rule that when a strong base and a strong acid react, the pH of their salt is always neutral which is at pH 7. However, this is only true if equal amounts of acid and base are consumed and that there is no excess. Otherwise, the excess acidity or basicity will adjust the total pH.

Strong acids are the following: HCl, HBr, HI, HClO4, HClO3, HNO3 and H2SO4. Strong bases are KOH, LiOH, NaOH, Ca(OH)2, Sr(OH)2 and Ba(OH)2. Therefore, we can already establish that both of the reactants are strong. The complete reaction is

$HBr + KOH \Rightarrow KBr + H_2O$

So, 1 mole of HBr would require 1 mol of KOH, and vice versa. Let'scompute the amount of the initial reactants:

mol HBr: (0.25 mol/L)*(0.68 L) = 0.17 mol HBr

mol KOH: (0.5 mol/L)*(0.34 L) = 0.17 mol KOH

There are equal amounts of acid and base. Thus, pH of the KBr solution is neutral at pH 7.

2) Consider the titration of 50.0 mL of 0.20 M NH_3 (Kb=1.8×10−5) with 0.20 M HNO_3. Calculate the pH after addition of 50.0 mL of the titrant at 25 ∘C.

The reaction between NH_3 and HNO_3 is:
$NH_3(aq) + HNO_3(aq) \dashrightarrow NH^{4+}(aq) + NO^{3-}(aq)$
In the titration, you have added equal moles of HNO_3 as you had NH_3 initially. So, at this point in the titration, you now have a solution of NH^{4+}.

Now, NH4+ is a weak acid, and you have the equilibrium:
$NH^{4+}(aq) \dashrightarrow NH_3(aq) + H^+(aq)$
$Ka = [NH_3][H^+]/[NH^{4+}]$
Ka for this can be calculated from Kb for NH_3 since Ka × Kb = 1.0×10^{-14}. So,
$Ka = 1.0\times10^{-14} / 1.8\times10^{-5} = 5.6\times10^{-10}$

Now, in that solution, let $[NH_3] = [H^+] = x$, and $[NH^{4+}] = 0.10$ M. Then,
$5.6\times10^{-10} = x^2 / 0.1$
$x = [H+] = 7.4\times10^{-6}$
pH = 5.13

3) A 30.0-mL volume of 0.50 M CH3COOH (Ka=1.8×10−5) was titrated with 0.50 M NaOH. Calculate the pH after addition of 30.0 mL of NaOH at 25 ∘C.

Answer provided by Socratic.org

The only species you'll have in solution after you do the titration is acetate ion (and Na+ ions, but they don't mess with the pH). The dissociation was:

$$CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$$

So what you have at the end is:

$$CH3COO^-(aq) + H_2O(l) \rightleftharpoons CH_3COOH(aq) + OH^-(aq)$$

You have a weak base in water. So you need the K_b. At $25\,^\circ C$:

$$K_w = K_a \cdot K_b = 1 \times 10^{-14}$$

$$\Rightarrow K_b = 5.56 \times 10^{-10}$$

ICE Table

$$CH_3COO^-(aq) + H_2O(l) \rightleftharpoons CH_3COOH(aq) + OH^-(aq)$$

I	0.25 M	$-$	0 M	0 M
C	$-x$	$-$	$+x$	$+x$
E	$0.25 - x$	$-$	x	x

Assume x is much less than 0.25 because $K_b << 10^{-5}$.

$$K_b = \frac{x^2}{0.25 - x} \approx \frac{x^2}{0.25\text{ M}},$$

where $x = [OH^-]$

$$\rightarrow [OH^-] = 1.179 \times 10^{-5}\text{M}$$

$$\rightarrow [H^+] = \frac{K_w}{[OH^-]} = 8.48 \times 10^{-10}\text{M}$$

$$\rightarrow pH = 9.07$$