HClO   $K_a = 2.9 \times 10^{-8}$

Given 0.50M HClO and 0.50M NaClO, what would be the pH of the soln if 10.0 mL of 8.0M HCl were added to 1L of the buffer

Since two concentrations are being used to calculate a ratio, moles can be used to calculate the ratio instead since both components are in the same solution (same volume),

mole HClO = (1L)(0.50M) + (0.0100L)(8.0M)
mole NaClO = (1L)(0.50M) - (0.0100L)(8.0M)

mole HClO = 0.50 + 0.080 = 0.58
mol NaClO = 0.50 + 0.080 = 0.42
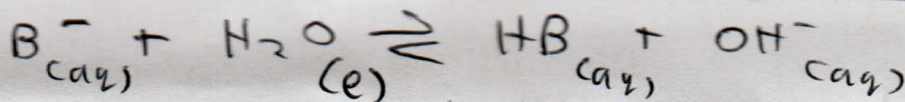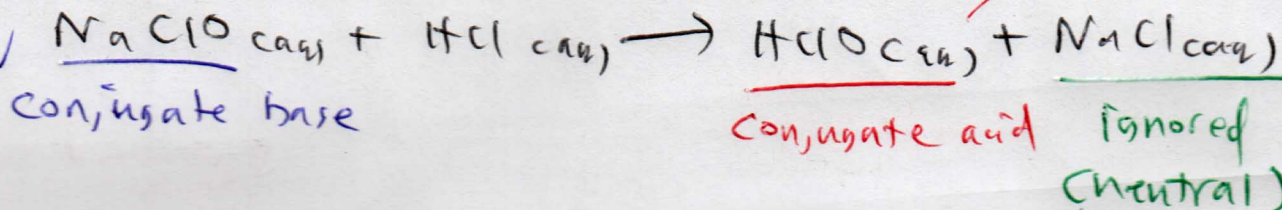
added because it formed

$pK_a = -\log_{10}(2.9 \times 10^{-8}) = 7.54$

$pH = pK_a + \log_{10}\frac{[A^-]}{[HA]} = 7.54 + \log\frac{0.42}{.58}$

$= 7.54 + (-0.14)$

$= 7.40$

subtracted because it reacted

$NaClO_{(aq)} + HCl_{(aq)} \rightarrow HClO_{(aq)} + NaCl_{(aq)}$

conjugate base            conjugate acid   ignored (neutral)

$B^-_{(aq)} + H_2O_{(e)} \rightleftharpoons HB_{(aq)} + OH^-_{(aq)}$

base dissociation constant    $K_B = \frac{[HB][OH^-]}{[B^-]}$

l2

$$K_a \cdot K_B = \frac{[H^+][A^-]}{[HA]}, \frac{[HA][OH^-]}{[B^-]}$$

If HA + B$^-$ are conjugates, the HA = HB, A$^-$ = B$^-$

$$= \frac{[H^+][A^-]}{[HA]}, \frac{[HA][OH^-]}{[A^-]} = [H^+][OH^-] = K_w$$

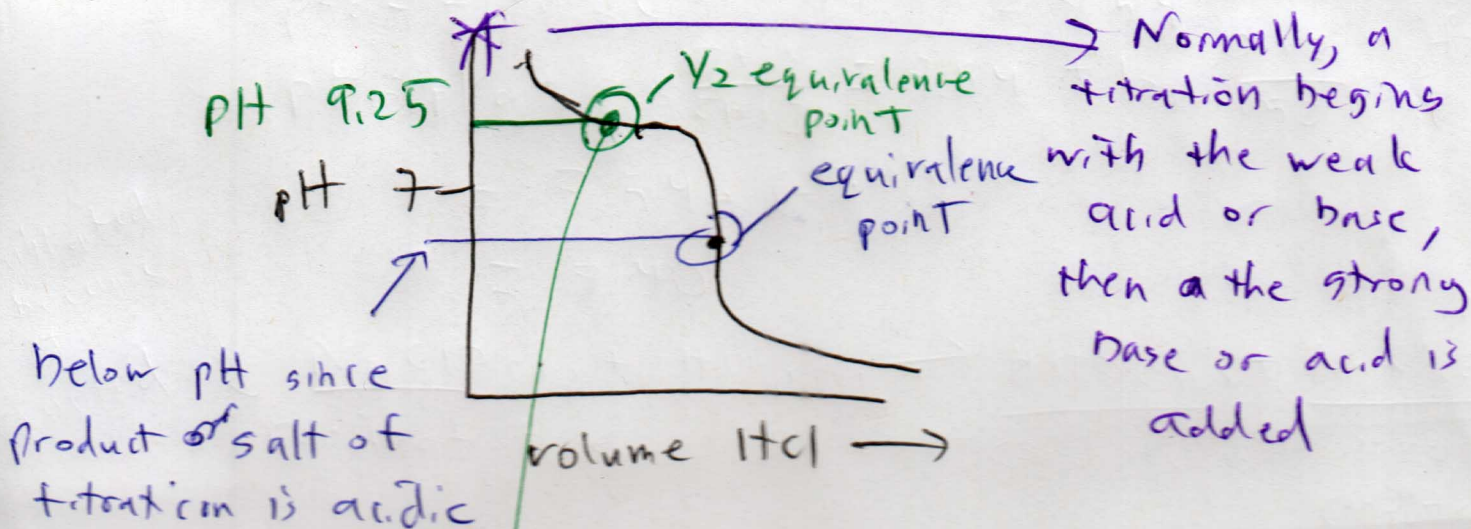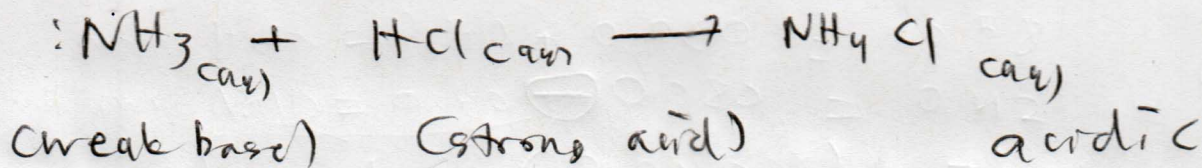$$\boxed{K_A \cdot K_B = K_w} \quad \text{(in water)}$$

when $K_w = 1.0 \times 10^{-14}$

$$-\log_{10}(K_A \cdot K_B = K_w)$$

$$-\log_{10}(K_A \cdot K_B) = -\log_{10} K_w$$

$$\cancel{\text{por}} -\log_{10} K_A + (-\log_{10} K_B) = -\log_{10} K_w$$

$$pK_a + pK_B = pK_w = 14$$

$$:NH_{3\,(aq)} + HCl_{(aq)} \longrightarrow NH_4Cl_{(aq)}$$

(weak base)     (strong acid)          acidic

pH 9.25

pH 7

below pH since
product of salt of
titration is acidic



½ equivalence
point

equivalence
point

volume HCl →

Normally, a
titration begins
with the weak
acid or base,
then a the strong
base or acid is
added

$$[:NH_3] = [NH_4^+]$$

$$K_B = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

$$K_B = [OH^-] \frac{[NH_3]}{[NH_3]}$$

$$pK_B = pOH$$

Can be eliminated
it concentrations
do not change much
when equilibrium is
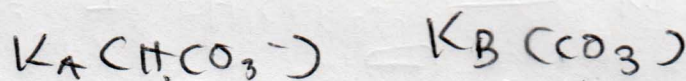established,

$$pOH + pH = pKw = 14$$

$$\boxed{pK_B = pKw - pH}$$

$$pK_B = 14 - 9.25 = 4.75$$

$$pK_B = \frac{[HB][OH]}{[A]}$$

$$pK_B = 1 - \log\left(\frac{[HB][OH]}{[A]}\right)$$

$$= pOH - \log\left(\frac{[HB]}{[B^-]}\right)$$

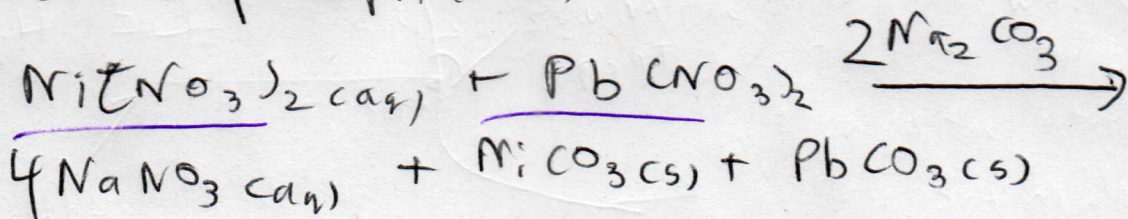$$pOH = pK_B + \log\frac{[HB]}{[B^-]}$$

$$HCO_3^- + CO_3^{-2}$$

$$K_A(H_2CO_3^-) \quad K_B(CO_3)$$

$$pK_a(H_2CO_3^-) = 14 - pK_B(CO_3)$$

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Selective precipitation

assume equal concentration

$$Ni(NO_3)_2\,(aq) + Pb(NO_3)_2 \xrightarrow{2Na_2CO_3}$$

$$4NaNO_3\,(aq) + NiCO_3\,(s) + PbCO_3\,(s)$$

$$Ksp(NiCO_3) = 1.3 \times 10^{-7}$$

$$Ksp(PbCO_3) = 7.4 \times 10^{-14}$$

$$Ksp(NiCO_3) >> Ksp(PbCO_3)$$

If $CO_3^{-2}$ is carefully added to a mixture of Ni(II) nitrate and Pb(II) nitrate, the $Pb^{+2}$ will be precipitated out of sol'n first since the Ksp value for lead is so much smaller than for nickel. Since the Ksp for nickel won't yet be reached, the $Pb^{+2}$ can be isolated from the $Ni^{+2}$,