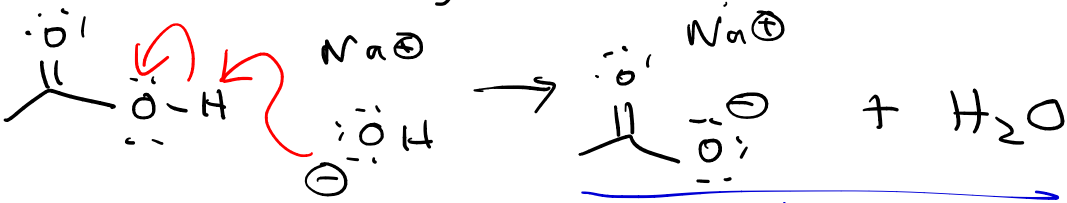
6/3/20

Problem 4: pH/pOH of a sol'n of a conjugate of a weak acid

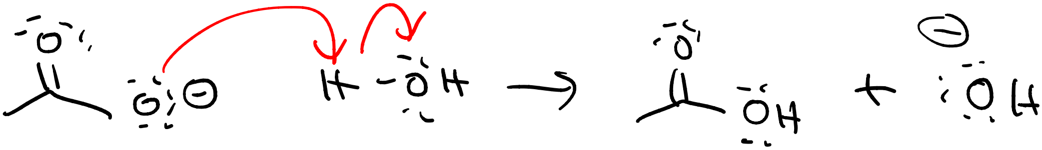
Suppose a 0.75 M aqueous sol'n of acetic acid is exactly neutralized by NaOH without changing the volume of the solution. Given the pK_a of acetic acid is 4.76, calculate the pH of the resulting solution.



This is a neutralization equation, which is not represented by either K_a or K_b

This is the same as sodium acetate dissolved in water.

It is assumed in this problem that acetic acid is completely neutralized; then, the resulting salt is allowed to reach equilibrium.



$$B^- + H_2O \rightarrow HB + OH^-$$

$$K_B = \frac{[HB][OH^-]}{[B^-]} = \frac{[CH_3CH_2OH]\,[OH^-]}{[CH_3CH_2O^-]}$$

$K_a = K_w / K_B$

This statement is only true of conjugate acid/base pairs.

$$K_B = K_w / K_a$$

$$pK_B = pK_w - pK_a$$

assume 25°C

$$pK_B = 14 - 4.76 = 9.24$$

$$K_B = 10^{-pK_B} = 10^{-9.24} = 5.75 \times 10^{-10}$$

equal to concentration of acid since fully neutralized

	$CH_3CH_2O^-$	CH_3CH_2OH	OH^-
I	0.75	0	0
C	$-x$	$+x$	$+x$
E	$0.75 \cancel{-x}$	x	x

assume x is small

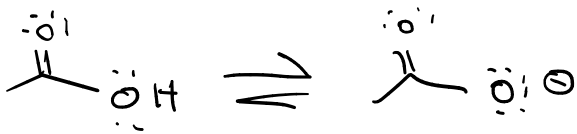
$$K_B = \frac{[HB][OH^-]}{[B^-]} = \frac{(x)(x)}{.75} = 5.75 \times 10^{-10}$$

$$x^2 = (.75)(9.75 \times 10^{-10})$$
$$= 4.32 \times 10^{-10}$$
$$x = 2.08 \times 10^{-5} = [OH^-]$$

$$[H^+] = K_w/[OH^-] = \frac{1.0 \times 10^{-14}}{2.08 \times 10^{-5}}$$
$$= 4.81 \times 10^{-10}$$
$$pH = -\log_{10}(4.81 \times 10^{-10}) = \underline{9.32}$$

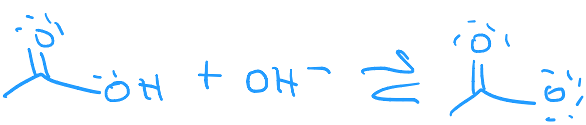
Since this is a neutralization of a weak acid with a strong base, it makes sense the neutralized sol'n should be basic.

(neutral ≠ neutralized)

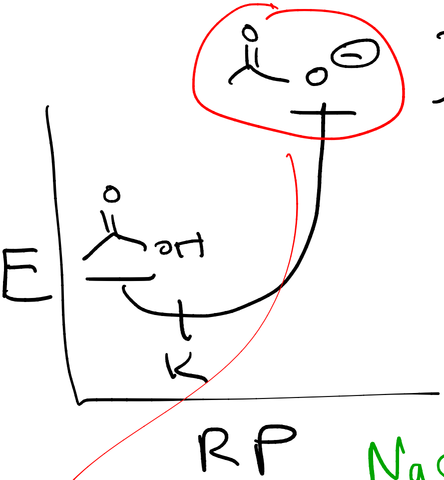


$+ \cancel{H^+}$ K_A

$$+ \quad \cancel{H^+}(aq) + OH^-(aq) \rightleftharpoons H_2O(l) \quad 1/K_w$$



$+ H_2O$ $K = K_a/K_w$



$K_a = 1.76 \times 10^{-5}$ → weak acid

$K_a < 1$ → ∴ $\Delta E > 0$

therefore

$NaOCH_2CH_3$ "=" $CH_3CH_2OH + NaOH$

The substance sodium acetate in solution would be the same result as forcing acetic acid to dissociate using NaOH, → Neutralized

this means that, in terms of the dissociation of acetic acid, the solution is not at all at equilibrium, since there's no reactants so no reverse rxn, and the system is not at its lowest point in energy.

Therefore, a solution of acetate will react with water to produce hydroxide in order to be able to establish equilibrium, → Not neutral

neutral: not pH 7 (only true at one temperature)

$[H^+] = [OH^-]$ (just like pure water)

neutralized: moles acid = mole base*

* adjusted for stoichiometry

A neutralized solution is only neutral if the acid and base inolved are equal in strength.

If a weak acid is neutralized by a strong base, there will effectively be a low $[H^+]$ due to weak dissociation of the acid and a high $[OH^-]$ due to it being a strong base. Since there is more OH^- than H^+, the solution is basic.