Lecture I

Solutions and Dilutions

PROBLEM1: We have a stock solution of HCl with a concentration of 1M. We want to prepare 100mL of HCl 0.1M. Which is the volume of the stock solution needed? How many mL of water should we add to obtain the final volume?

Solution:

C_1=1M

C_2=0.1M

V_1=x

V_2=100mL

$$C_1V_1 = C_2V_2$$

$$V_1 = \frac{0.1M \cdot 0.1L}{1M} = 0.01L = 10mL$$

It means that we take 10mL of the initial stock (1M) HCl and we will add 90mL of water in order to have the final volume of 100mL.

PROBLEM 2: How much of a 0.4 mg/mL stock solution of insulin solution is required to provide a patient with 10 micrograms (μg) of insulin?

Solution:

C_1 (stock solution) = 0.4 mg/mL=400 μg/mL

C_2V_2=10 μg

V_1 (stock solution) = x

$$V_1 = \frac{10\mu g}{400\mu g/mL} = 0.025mL = 25\mu L$$

Weak acid: K_a, pH and molarity

PROBLEM 3: The degree of dissociation of a 0.1 M acetic acid (CH_3COOH) solution is α=0.132. What would be the pH and dissociation constant K_a of acetic acid?

Solution:

Acetic acid is a weak electrolyte and dissociates in acetate ion and hydrogen ion:

$$CH_3COOH_{(aq)} \leftrightarrows CH_3COO^-_{(aq)} + H^+_{(aq)}$$

If the initial concentration of acetic acid is taken as C_0 then initial and equilibrium concentration of all species would be:

	$CH_3COOH_{(aq)} \leftrightarrows$	$CH_3COO^-_{(aq)}$ +	$H^+_{(aq)}$
Initial concentration	C_0	0	0
Modification	$-\alpha C_0$	$+\alpha C_0$	$+\alpha C_0$
Concentration at the equilibrium	$C_0-\alpha C_0$	αC_0	αC_0

By computing the acid constant/dissociation constant formula:

$$K_a = \frac{\left[CH_3COO^-_{(aq)}\right]\cdot[H^+_{(aq)}]}{[CH_3COOH_{(aq)}]} = \frac{C_0^2\cdot\alpha^2}{C_0\cdot(1-\alpha)} = \frac{0.1\cdot 0.132^2}{(1-0.132)} = 2.007\cdot 10^{-3}$$

To calculate the pH we need to compute the $[H^+_{(aq)}]$. By definition we know that $\left[H^+_{(aq)}\right] = C_0 \cdot \alpha$, thus we have that:

$$[H^+_{(aq)}] = C_0\cdot\alpha = 0.1\cdot 0.132 = 0.0132M$$

thus the pH is

$$pH = -\log_{10}[H^+_{(aq)}] = 1.9$$

PROBLEM 4: A 0.228 M solution of uric acid ($HC_5H_3N_4O_3$) has a pH of 2.39. Calculate the K_a of uric acid.

Solution

Uric acid goes under dissociation

$$HC_5H_3N_4O_{3(aq)} \leftrightarrows C_5H_3N_4O^-_{3(aq)} + H^+_{(aq)}$$

$$K_a = \frac{\left[C_5H_3N_4O^-_{3(aq)}\right] \cdot [H^+_{(aq)}]}{[HC_5H_3N_4O_{3(aq)}]}$$

If we define 'x' as the unknown concentration (in mol/L):

	$C_5H_3N_4O_{3(aq)} \leftrightarrows$	$C_5H_3N_4O^-_{3(aq)}$ +	$H^+_{(aq)}$
Initial concentration	0.228	0	0
Modification	-x	+x	+x
Concentration at the equilibrium	0.228-x	x	x

From the pH:

$$x = [H^+_{(aq)}] = 10^{-pH} = 10^{-2.39} = 4.07 \cdot 10^{-3} M$$

So

$$K_a = \frac{\left[C_5H_3N_4O^-_{3(aq)}\right] \cdot [H^+_{(aq)}]}{[HC_5H_3N_4O_{3(aq)}]} = \frac{x^2}{0.228 - x} = 7.41 \cdot 10^{-5} M$$