Friday Worksheet **Name: ……………….**

Acid base equilibria worksheet 2

1) Consider the two solutions below at 25°C
   i. 100.0 mL of 0.100 M HCOOH
   ii. 100.0 mL of 0.100 M HCl

   a) What is the pH of each solution?
   The HCl solution will have a pH of 1.000 since it is a strong acid and full ionisation will occur.
   The pH of the weak methanoic acid will have to be calculated using the Ka at 25°C.
   Ka of methanoic acid = 1.8×10^{-4} at 25°C.
   => $1.8 X 10^{-4} = [H_3O^+] [HCOO^-] / [HCOOH]$ According to the stoichiometry
   $[H_3O^+] = [HCOO^-]$ hence we can write the expression below
   $1.8 X 10^{-4} = [H_3O^+]^2 / [HCOOH]$.
   => $1.8 X 10^{-4} = [H_3O^+]^2 / [0.100]$. Here we assume negligible ionisation hence the concentration of methanoic acid is still 0.100M.
   => $1.8 X 10^{-5} = [H_3O^+]^2$
   => $4.24 X 10^{-3} = [H_3O^+]$
   $pH = -\log_{10}[4.24 X 10^{-3}] = 2.37$

   b) The pH of which solution will undergo the greatest change when 900 mL of water is added to the solution. Explain
   The HCl is fully ionised. Hence a 1:10 dilution will drive the pH from 1 to 2 as the $[H_3O^+]$ changes from 0.100M to 0.0100M, 10 fold reduction.
   For the methanoic acid, however, being a weak acid, the reaction below

$$HCOOH(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + HCOO^-(aq)$$

   will shift to the right to increase the $[H_3O^+]$ as the Ka remains constant at the same temperature.
   The pH of a solution of 0.0100M HCOOH at 25°C is
   => $1.8 X 10^{-4} = [H_3O^+]^2 / [0.0100]$
   => $1.8 X 10^{-6} = [H_3O^+]^2$
   => $10^{-2.872} = [H_3O^+]$
   => $pH = -\log_{10} [10^{-2.872}]$
   => 2.872

2) The ionisation of ethanoic acid can be represented by the equation

$CH_3COOH(aq) + H_2O(l) \rightleftharpoons CH_3COO^-(aq) + H_3O^+(aq)$

Which of the following solutions has the highest percentage ionisation. Verify mathematically and show all working out.

A. 50 mL 1.0 M CH_3COOH solution..

B. 100 mL 0.01 M CH_3COOH solution.

Dilution causes the percentage ionisation of weak acids to increase.

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

The above equilibrium will shift to the right when diluted. On dilution all the concentrations are decreased and the system is pushed out of equilibrium.

For option A.

=> $1.7 \times 10^{-5} = [CH_3COO^-][H^+] / [CH_3COOH]$

=> $1.7 \times 10^{-5} = [CH_3COO^-]^2 / [1.0]$

=> $0.0041M = [CH_3COO^-]$

=> % ionisation = (0.0041 / 1.0) X 100 = 0.41%

For option **B**.

=> $1.7 \times 10^{-5} = [CH_3COO^-][H^+] / [CH_3COOH]$

=> $1.7 \times 10^{-5} = [CH_3COO^-]^2 / [0.01]$

=> $0.00041M = [CH_3COO^-]$

=> % ionisation = (0.00041 / 0.01) X 100 = 4.1%

4) A 20.00 mL aliquot of a 0.200 M CH_3COOH (ethanoic acid) is titrated with 0.150 M NaOH. The equation for the reaction between the ethanoic acid and NaOH solution is represented below.

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

What volume of the NaOH solution is required to completely react with the ethanoic acid?

Step 1 find the mol of ethanoic acid

=> n = C X V = 0.200 X 0.02000 = 0.00400

Step 2 find the volume of NaOH

=> V = C / n = 0.150 / 0.00400 = 6.00×10^{-3} L

5) A weak monoprotic acid has a K_a of $10^{-4.994}$ at 25°C and the solution has a pH of 4.523. What percentage of the acid is ionised?

Step 1 find the initial concentration of the acid

$XH(aq) + H_2O(l) \Rightarrow H_3O^+(aq) + X^-(aq)$

=> $[H_3O^+][X^-] / 10^{-4.994} = [XH]$

=> $[H_3O^+]^2 / 10^{-4.994} = [XH]$

=> $10^{-9.46} / 10^{-4.994} = [XH] = 10^{-4.466}$

⇨ $(10^{-4.523}/10^{-4.466}) \times 100 = 87.7\%$