Friday Worksheet **Name: ……………….**

Acid Base equilibria worksheet 1

1) What is the pH of a 100.0 mL 0.325 M H_3BO_3 solution at 25 °C?

$Ka_{boric\ acid} = 5.8 \times 10^{-10}$

$pH = -\log[H_3O^+]$

Step 1 find the $[H_3O^+]$

⇨ $K_a = \frac{[H_2BO_3^-][H_3O^+]}{[H_3BO_3]} = 5.8 \times 10^{-10}$

Since according to the stoichiometry $[H_2BO_3^-] = [H_3O^+]$ we can write the expression below.

⇨ $K_a = \frac{[H_3O^+]^2}{[H_3BO_3]} = 5.8 \times 10^{-10}$

⇨ Since Ka is very small assume that negligible amount ionises. So we can write the expression below

⇨ $\frac{[H_3O^+]^2}{0.325} = 5.8 \times 10^{-10}$

⇨ $[H_3O^+]^2 = 5.8 \times 0.325 \times 10^{-10}$

⇨ $[H_3O^+]^2 = 1.9 \times 10^{-10}$

⇨ $[H_3O^+] = 1.4 \times 10^{-5} = 10^{0.15} \times 10^{-5} = 10^{-4.85}$

⇨ pH = 4.85

2) Ethanoic acid is a weak monoprotic acid.

a) Write the equation that represents the ionisation reaction of ethanoic acid.

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

b) Write the equilibrium expression for this reaction.

$K_e = \frac{[CH_3COO^-][H_3O^+]}{[CH_3COOH]\ [H_2O]}$

c) Write the expression for the Ka of ethanoic acid

$K_a = \frac{[CH_3COO^-][H_3O^+]}{[CH_3COOH]}$

d) Which has the highest pH and offer an explanation.

i) 10.0 mL 0.100 M HCOOH solution or 100.0 mL 0.100 M CH_3COOH solution

Since the Ka(HCOOH) = 1.8 X 10^{-4} and the Ka(CH_3COOH) = 1.7 X 10^{-5} there will be a greater degree of ionisation, at the same concentration, in the methanoic acid solution than the ethanoic acid solution. A greater degree of ionisation leads to a higher [H_3O^+] and hence a lower pH. So ethanoic acid has the higher pH.

ii) 10.0 mL of 0.01 M HCOOH solution or 10.0 mL 0.100 M HCOOH solution

Dilution increases the extent of ionisation however the [H_3O^+] decreases with dilution.

$$K_a = \frac{[H_3O^+]^2}{[HCOOH]} = 1.8 \text{ X } 10^{-4}$$

$$\Rightarrow \frac{[H_3O^+]^2}{[0.100]} = 1.8 \text{ X } 10^{-4}$$

$$\Rightarrow [H_3O^+]^2 = 1.8 \text{ X } 10^{-5}$$

=> $[H_3O^+] = 10^{-2.37}$ pH for 0.1 M HCOOH is 2.37

Now for thre 0,01M HCOOH

$$\Rightarrow \frac{[H_3O^+]^2}{[0.010]} = 1.8 \text{ X } 10^{-4}$$

$$\Rightarrow [H_3O^+]^2 = 1.8 \text{ X } 10^{-6}$$

=> $[H_3O^+] = 10^{-2.87}$
pH for 0.01 M HCOOH is 2.87

e) Explain why diluting a solution of 0.100M HCOOH to 0.001M HCOOH, at constant temperature, increases the percentage ionisation of HCOOH.

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

$$K_a = \frac{[HCOO^-]\,[H_3O^+]}{[HCOOH]} = 1.8 \text{ X } 10^{-4}$$

Diluting the solution lowers the concentration of the ions and drives the reaction forward hence ionising a greater percentage of the HCOOH present. Dilution, however, ultimately reduces the [H_3O^+] and increases pH.

3) A 20.00 mL aliquot of 0.200 M CH_3CH_2COOH (propanoic acid) is titrated with 0.250 M NaOH. The equation for the reaction between propanoic acid and NaOH solution is represented below.

$CH_3CH_2COOH(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CH_3CH_2COO^-(aq)$

a) Write the expression for the acidity constant.

$$K_a = \frac{[CH_3CH_2COO^-][H_3O^+]}{[CH_3CH_2COOH]}$$

b) What volume of NaOH is required to completely react with the acid.

$n_{propanoic\ aicd}$ = C X V = 0.02000 X 0.200 = 4 X 10^{-3}

n_{NaOH} = 4 X 10^{-3}/0.250 = 16.0 mL

c) Calculate the pH of the 0.200 M propanoic acid solution before any NaOH solution has been added.

$Ka_{propanoic\ acid}$ =1.3 X 10^{-5}

pH = -log[H_3O^+]

Step 1 find the [H_3O^+]

$$K_a = \frac{[CH_3CH_2COO^-][H_3O^+]}{[CH_3CH_2COOH]} = 1.3 \times 10^{-5}$$

- ⇨ Since according to the stoichiometry [$CH_3CH_2COO^-$] = [H_3O^+] we can write the expression below.
- ⇨ $K_a = \frac{[H_3O^+]^2}{[CH_3CH_2COOH]} = 1.3 \times 10^{-5}$

- ⇨ Since Ka is very small assume that negligible amount of propanoic acid ionises. So we can write the expression below

- ⇨ $\frac{[H_3O^+]^2}{0.200} = 1.3 \times 10^{-5}$

- ⇨ $[H_3O^+]^2 = 2.60 \times 10^{-6}$
- ⇨ $[H_3O^+]^2 = 1.61 \times 10^{-3}$
- ⇨ $[H_3O^+] = 10^{0.207} \times 10^{-3} = 10^{-2.79}$
- ⇨ pH = 2.79