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Course Code: 109 | **ID: 2021-3-60-016**
Course Title: CHE | **Section: 5**

Experiment-6: Determination of the acid dissociation constant (Ka) of an weak acid (acetic acid, CH3COOH).

Theory:

When a weak acid is dissolved in water, it breaks apart or dissociates to a slight extent. A proton from the acid is donated to a water molecule. The equations for the equilibrium and the equilibrium constant expression are as follows:

$$HA + H_2O \leftrightarrows H_3O^+ + A^-$$

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

Where A - represents the anion of the weak acid and the square brackets indicate molar concentrations of the species. The value of the equilibrium constant, Ka , indicates to what extent the reaction occurs. The greater the value of Ka , the stronger the acid, and the greater the amount of dissociation. Acetic acid and acetate ions are conjugate acid-base pairs. A conjugate acid is a substance that has one more proton in its structure than its corresponding conjugate base. This combination also results from a mixture of a weak acid, acetic acid, and its salt, sodium acetate. The equilibrium constant expression is:

$$K_a = \frac{[H_3O^+][C_2H_3O_2^-]}{[HC_2H_3O_2]}$$

If a solution contains equal concentrations of $HC_2H_3O_2$ and $C_2H_3O_2^-$ these concentration terms cancel out in the above equation so that $Ka = [H_3O^+]$.

Procedure:

1. Fill the burette with 0.1M standardized NaOH solution.

2. Take exactly 50.0 mL of acetic acid (unknown concentration, 0.1 M) in a beaker and record the pH of that acid solution (solution 1).

3. Pour precisely 25.0 mL of the weak acid solution into an Erlenmeyer flask and add 2 drops (0.2 mL) of phenolphthalein solution.

4. Titrate this solution with NaOH solution by continuously adding small volumes (~0.2 mL or less) until slight changing its color to a permanent pink. This process converts all the weak acid,

HC2H3O2 , in the flask into its conjugate base, C2H3O2 - , according to the neutralization reaction (OH- + HA = H2O + A). Record the pH of the solution (solution 2).

(Note: the beaker contains exactly one-half of the original acid (25 mL), essentially all of which is in the undissociated form, HC2H3O2 , and the flask contains an equal amount of the C2H3O2 - of the weak acid.)

5. To calculate the Ka of the weak acid, follow any one of the following calculations (a or b).

(a) Pour the contents of the flask into the beaker and mix the solution. Measure the pH of this solution using both a pH meter.

The pH is the pKa of the acid. Calculate the value of Ka of the acid. (note: pKa = -log K a).

(b) Calculate the final pH of the mixture of solution 1 and solution 2 mathematically to follow the equation:

$$C1\ V1 + C2\ V2 = Cf\ Vf$$

$$Or,\ Cf = (C1\ V1 + C2\ V2)/V\ f$$

Where,

$C1$: [H +] of solution 1; $C2$: [H +] of solution 2,

$V1$: volume of solution 1 (25 mL of acid);

$V2$: volume of solution 2 (25 mL of acid + added NaOH);

Vf : final volume of (solution 1 + solution 2)

Cf : final [H +] of (solution 1 + solution 2) =?

Therefore, the pH = -log [Cf]; The pH is the pKa of the acid.

Calculate the value of Ka of the acid. (note: pKa = -log K a).

<u>Data:</u>

Volume of NaOH, ml	pH of Acid
0	2.88
0.2	3.02
0.4	3.14
0.6	3.25
0.8	3.34
1	3.42
1.2	3.49
1.4	3.55
1.6	3.61
1.8	3.66
2	3.71
2.2	3.75
2.4	3.79
2.6	3.83
2.8	3.86
3	3.9
3.2	3.93
3.4	3.96
3.6	3.99
3.8	4.01
4	4.04
4.2	4.07
4.4	4.09
4.6	4.11
4.8	4.14
5	4.16
5.2	4.18
5.4	4.2
5.6	4.22
5.8	4.24
6	4.26
6.2	4.28
6.4	4.3
6.6	4.31
6.8	4.33

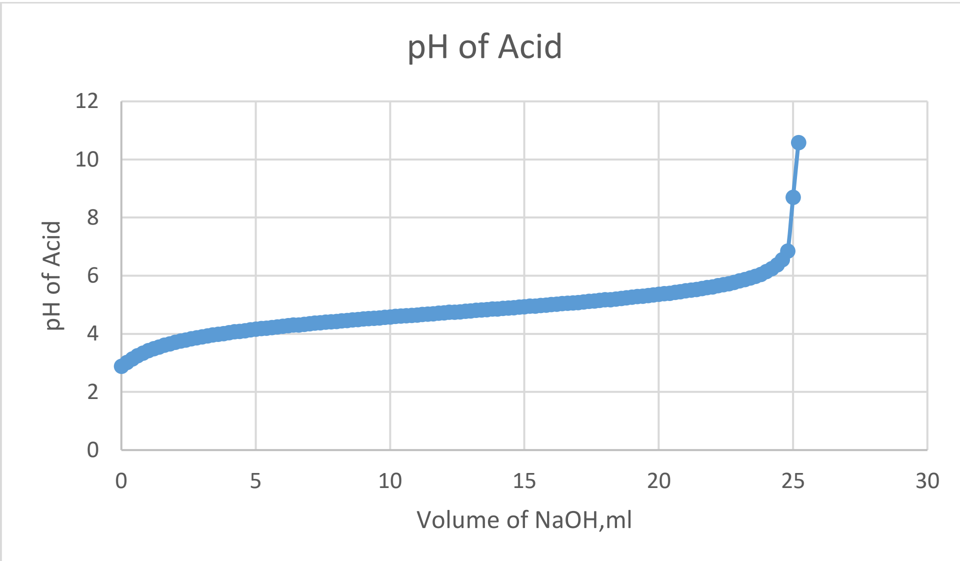
7	4.35
7.2	4.37
7.4	4.38
7.6	4.4
7.8	4.41
8	4.43
8.2	4.45
8.4	4.46
8.6	4.48
8.8	4.49
9	4.51
9.2	4.52
9.4	4.54
9.6	4.55
9.8	4.57
10	4.58
10.2	4.6
10.4	4.61
10.6	4.62
10.8	4.64
11	4.65
11.2	4.67
11.4	4.68
11.6	4.69
11.8	4.71
12	4.72
12.2	4.74
12.4	4.75
12.6	4.76
12.8	4.78
13	4.79
13.2	4.81
13.4	4.82
13.6	4.83
13.8	4.85
14	4.86
14.2	4.88
14.4	4.89
14.6	4.9
14.8	4.92
15	4.93
15.2	4.95
15.4	4.96
15.6	4.98
15.8	4.99
16	5.01
16.2	5.02
16.4	5.04
16.6	5.05
16.8	5.07

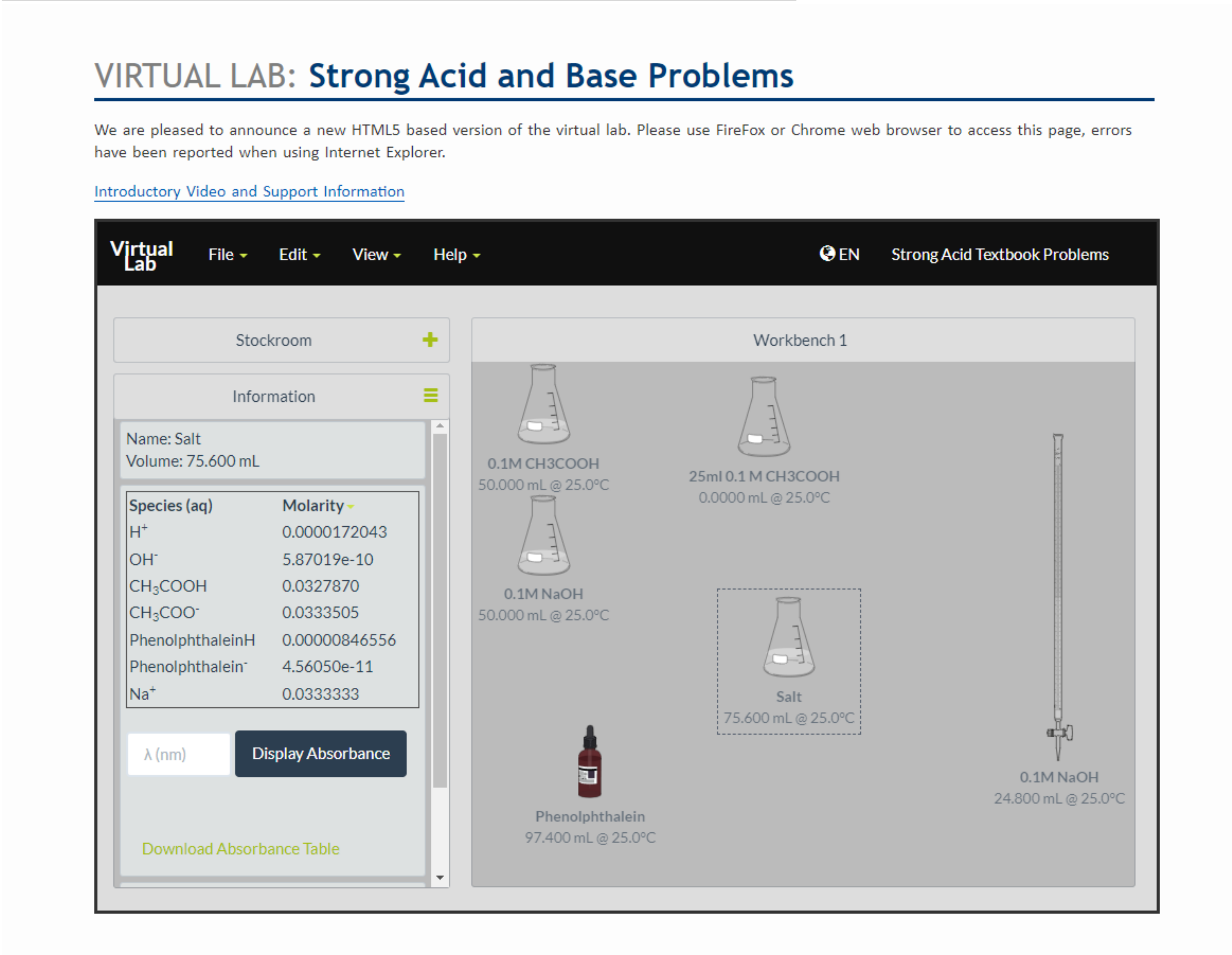
17	5.08
17.2	5.1
17.4	5.12
17.6	5.13
17.8	5.15
18	5.17
18.2	5.18
18.4	5.2
18.6	5.22
18.8	5.24
19	5.26
19.2	5.28
19.4	5.3
19.6	5.32
19.8	5.34
20	5.36
20.2	5.38
20.4	5.4
20.6	5.43
20.8	5.45
21	5.48
21.2	5.5
21.4	5.53
21.6	5.56
21.8	5.59
22	5.62
22.2	5.66
22.4	5.69
22.6	5.73
22.8	5.77
23	5.82
23.2	5.87
23.4	5.92
23.6	5.98
23.8	6.05
24	6.14
24.2	6.24
24.4	6.37
24.6	6.55
24.8	6.85
25	8.7
25.2	10.58

Calculation:

pH= pKa + log [salt]/ [acid
pH= pKa + log [50]/ [50]
pH= pKa + log [1]

pH= pKa + 0
pH= pKa= 4.76
Now, pKa= -logKa
Ka = 1.73*10^-5





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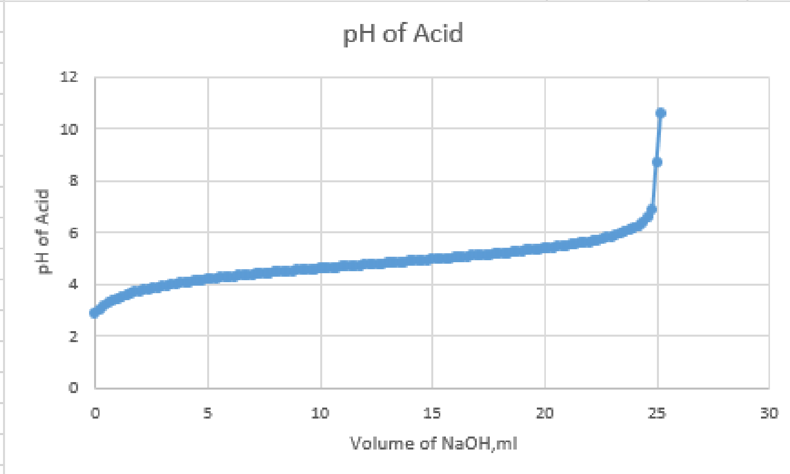

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pH= pKa + log [50]/ [50]
pH= pKa + log [1]
pH= pKa + 0
pH= pKa= 4.76

Now, pKa= -logKa
Ka = 1.73*10^-5

61	11.8	4.71
62	12	4.72
63	12.2	4.74
64	12.4	4.75
65	12.6	4.76
66	12.8	4.78
67	13	4.79
68	13.2	4.81
69	13.4	4.82
70	13.6	4.83
71	13.8	4.85
72	14	4.86
73	14.2	4.88
74	14.4	4.89
75	14.6	4.9
76	14.8	4.92
77	15	4.93
78	15.2	4.95
79	15.4	4.96
80	15.6	4.98
81	15.8	4.99
82	16	5.01
83	16.2	5.02
84	16.4	5.04
85	16.6	5.05
86	16.8	5.07
87	17	5.08
88	17.2	5.1
89	17.4	5.12
90	17.6	5.13
91	17.8	5.15
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