UNIT 9 APPLICATIONS OF EMF MEASUREMENTS

Structure

9.1 INTRODUCTION

In Unit 3 you were introduced to elementary concepts of potentiometry such as types of electrodes, electrode potential, galvanic cell, cell potential, emf, Nernst equation, etc. We also described a demonstration experiment for construction of a Daniell cell and measurement of its emf. In this unit we shall describe two experiments which are based on the applications of emf measurements. In these experiments, you will measure the emf of appropriate galvanic cells using low cost potentiometer and then determine the pH of solutions and the solubility product of a sparingly soluble salt such as silver chloride.

After studying this unit and having performed the experiments described in this unit, you should be able to:

● explain the terms electrodes, electrode potential, galvanic cell, cell potential, emf, Nernst equation, pH and solubility product,

● state the relationship between emf of a cell (or half-cell potential) and the concentration of species involved in cell reaction (or half-cell reaction),

● describe the relationship between emf of a cell and the pH of the solution in the indicator half-cell,

● construct copper, silver and quinhydrone electrodes,

● measure emf of a cell using low cost potentiometer,

● determine pH of an acid solution and solubility product of a sparingly soluble salt by measuring emf with a low cost potentiometer.

9.2 EXPERIMENT 8 : DETERMINATION OF pH BY EMF METHOD

In Chemistry Lab - I, CHE-03(L), course, you performed an experiment based on low cost pH meter. You performed a pH metric titration and determined the amount of acetic acid in a commercial sample of vinegar. The low cost pH meter can also be used for measuring pH of solutions directly. But in this experiment, you will use the low cost pH meter as a potentiometer. You will measure the emf of a galvanic cell and then, calculate the pH of hydrochloric acid and acetic acid solution by making use of Nernst equation.

9.2.1 Principle

In Unit 3 of this course, you studied that the quinhydrone electrode is reversible with respect to the H^+ ions. You have also studied that the half-cell reaction of this electrode is

$$C_6H_4O_2(s) + 2H^+(c) + 2e^- \longrightarrow C_6H_4(OH)_2(s) \qquad \qquad ..... (3.27)$$
$$\quad (Q) \qquad \qquad \qquad \qquad (H_2Q)$$
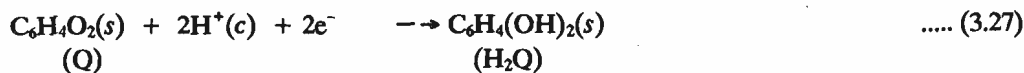
The expression for electrode potential of a quinhydrone electrode (E_{Q/H_2Q}) in a solution of H^+ ions can be written, by using Nernst equation (Eq.3.18), as follows :

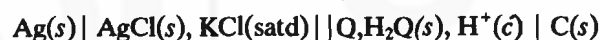$$E_{Q/H_2Q} = E^0_{Q/H_2Q} + \frac{2.303\,RT}{2F} \log [\,H^+]^2 \qquad \qquad .... (3.18)$$

$$= E^0_{Q/H_2Q} + \frac{2.303\,RT \times 2}{2F} \log [\,H^+]$$

$$= E^0_{Q/H_2Q} + \frac{2.303\,RT}{F} \cdot pH \qquad \qquad .... (3.30)$$

The standard reduction potential, E^0_{Q/H_2Q}, for quinhydrone electrode is 0.699 V. Therefore, from the above we get,

$$E_{Q/H_2Q} = 0.699 - \frac{2.303\,RT}{F}\, pH \qquad \qquad .... (9.1)$$

Thus, you can see that the electrode potential of a quinhydrone electrode is related to the pH of the solution with which it is in contact. Now if you set up a galvanic cell by connecting the quinhydrone indicator electrode to a saturated silver-silver chloride reference electrode, the quinhydrone electrode will act as a reduction electrode and the silver-silver chloride electrode will act as an oxidation electrode. The cell can be represented as follows:

$$Ag(s)\,|\,AgCl(s),\,KCl(satd)\,|\,|\,Q,H_2Q(s),\,H^+(c)\,|\,C(s)$$

The half-cell reactions are :

$$Ag(s) + Cl^-(aq) \longrightarrow AgCl(s) + e^-,\ and$$

$$Q(s) + 2H^+(c) + 2e^- \longrightarrow H_2Q(s)$$

Using Eq. 3.14 of Unit 3 of this course, the emf of the cell can be written as:

$$E_{cell} = E_{Q/H_2Q} - E_{AgCl/Cl^-}$$

But E_{AgCl/Cl^-} (as per literature value) = 0.222 V

Using this in Eq. 9.1, we get

$$E_{cell} = 0.699 - \frac{2.303\,RT}{F} pH - 0.222$$

$$\text{or}\quad E_{cell} = 0.477 - \frac{2.303\,RT}{F} pH$$

$$\text{or}\quad pH = \frac{(E_{cell} - 0.477)F}{2.303\,RT} \qquad \qquad .... (9.2)$$

Thus, by measuring the emf of the cell, the pH of the solution can be calculated.

9.2.2 Requirements

Apparatus			Chemicals
1.	Low cost potentiometer	1 No.	1. Quinhydrone
2.	Silver-silver chloride electrode	1 No.	2. Potassium chloride
3.	Carbon electrode	1 No.	

As per literature value, E_{AgCl/Cl^-} = 0.222 V

4.	Salt Bridge	1 No.
5.	Beakers 100 cm³	2 Nos.

Stock Solutions Provided

1. Hydrochloric acid (1.00×10^{-2} M)
2. Acetic acid - (1.00×10^{-2} M)

9.2.3 Procedure

Calibrate the low cost potentiometer by following the method given in Sub-section 3.4.6 of Block 1 of this course.

Set up a silver-silver chloride half-cell by taking 20-25 cm³ of saturated solution of KCl in one 100 cm³ beaker and by placing a silver-silver chloride electrode in it. Set up a quinhydrone half-cell by taking 20-25 cm³ of 1.00×10^{-2} M hydrochloric acid in another 100 cm³ beaker. Add a pinch of quinhydrone and insert a carbon electrode in this solution. Connect the two half-cells through a KCl salt bridge. If a salt bridge is not available, take a strip of filter paper, dip it in saturated KCl solution and use it in place of the salt bridge.

Connect the carbon electrode to the indicator terminal (through the red crocodile clip) and the silver-silver chloride electrode to the reference terminal (through the black crocodile clip) of the instrument.

Keep the range selector at 2 V and then measure the emf of the cell. Remember each division on the meter scale is equal to 0.04 V in the 2 V range. Take the measurements at least two more times for the same set-up, and record the values in Table 9.1.

Now set up the quinhydrone half-cell again, this time using acetic acid (1.00×10^{-2} M), and measure the emf value in the same manner. Repeat the measurement at least two more times and record the values in Table 9.1.

9.2.4 Observations

Room temperature, T = °C = K

Table 9.1 : Emf Values of Cells

Cell	Emf/V
Ag(s) | AgCl(s), KCl(satd) | | Q,H$_2$Q(s), HCl(1.00×10^{-2} M) | C(s)	--------
	--------

Average value of emf of the cell	--------
Ag(s)| AgCl(s), KCl(satd) | | Q, H$_2$Q(s), CH$_3$COOH(1.00×10^{-2}M)| C(s)	--------
	--------

Average value of emf of the cell	--------

9.2.5 Calculations

pH value of 1.00×10^{-2} M   HCl

Average value of emf of the cell containing 1.00×10^{-2} M HCl	=	 V
Temperature (T)	=	..... K
Gas constant (R)	=	8.314 J K^{-1} mol^{-1}
Faraday constant (F)	=	96500 C mol^{-1}

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Using Eq. 9.2, pH of 1.00×10^{-2} M HCl $= \dfrac{(E_{cell} - 0.477)\, 96500}{2.303 \times 8.314\, T}$

(You can compare this with the theoretical pH value of 2)

pH Value of 1.00×10^{-2} M CH$_3$COOH

Average value of emf of the
cell containing 1.00×10^{-2} M CH$_3$COOH = ... V

Again, using Eq. 9.2, pH of 1.00×10^{-2} M CH$_3$COOH $= \dfrac{(E_{cell} - 0.477)\, 96500}{2.303 \times 8.314\, T}$

9.2.6 Result

The pH values of 1.00×10^{-2} M HCl and 1.00×10^{-2} M CH$_3$COOH are found to be ---
and ---, respectively.

Give reason for the relative magnitudes of pH of 1.00×10^{-2} M HCl and 1.00×10^{-2} M
CH$_3$COOH in the space given below:

. . . .

. . . .

. . . .

. . . .

. . . .

9.3 EXPERIMENT 9 : DETERMINATION OF SOLUBILITY PRODUCT OF AgCl FROM EMF MEASUREMENTS

In Experiment 8 you have made use of emf of a cell in determining the pH of a solution. In
Unit 17 of the CHE-04 course, you had studied that the emf measurement can also be used
in determining the solubility product of a sparingly soluble salt. In this experiment, you will
determine the solubility product of a sparingly soluble salt, silver chloride, by measuring the
emf of an appropriate galvanic cell.

9.3.1 Principle

You know that the silver chloride with a solubility of 1.37×10^{-5} mol dm^{-3} in water is a sparingly
soluble salt. As a result, when silver chloride is placed in water, only a very small amount of
it dissolves in water to give a saturated solution. You also know that all the sparingly soluble
salts are strong electrolytes; therefore, AgCl that dissolves will be completely dissociated into
Ag$^+$ and Cl$^-$ ions. Thus, there will be an equilibrium between undissolved AgCl and its ions
in solution:

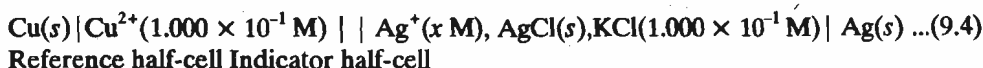$$AgCl(s) \quad \underset{\longleftarrow}{\overset{H_2O}{\longrightarrow}} \quad Ag^+(aq) + Cl^-(aq)$$
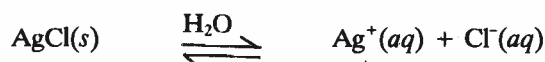
The equilibrium constant for this reaction, which is generally known as solubility product
constant, K_{sp}, is given by

$$K_{sp} = [Ag^+(aq)][Cl^-(aq)] \qquad\qquad ..... (9.3)$$

where $[Ag^+(aq)]$ and $[Cl^-(aq)]$ represent the molar concentrations of respective ions. The
potentiometric determination of K_{sp} is based on the following simple idea.

An aqueous solution of AgCl contains equal amounts of Ag$^+$ and Cl$^-$ ions. However, if to
this solution, a known excess of KCl is added, then [Cl$^-$] is essentially determined by the

concentration of KCl. The $[Ag^+]$ can be determined by setting up and then measuring the emf of the following galvanic cell:

$$Cu(s)|Cu^{2+}(1.000 \times 10^{-1}\,M)\,|\,|\,Ag^+(x\,M), AgCl(s), KCl(1.000 \times 10^{-1}\,M)|\,Ag(s)\,...(9.4)$$
Reference half-cell   Indicator half-cell

The emf of the cell is given by

$$
\begin{aligned}
E_{cell} &= E_{Ag^+/Ag} - E_{Cu^{2+}/Cu} \\
&= \left(E^{\circ}_{Ag^+/Ag} + \frac{2.303\,RT}{F}\log[Ag^+]\right) - \left(E^{\circ}_{Cu^{2+}/Cu} + \frac{2.303\,RT}{F}\log[Cu^{2+}]\right) \\
&= E^{\circ}_{Ag^+/Ag} - E^{\circ}_{Cu^{2+}/Cu} - \frac{2.303\,RT}{2F}\log(1.000 \times 10^{-1}) + \frac{2.303\,RT}{F}\log[Ag^+] \\
&= \left(E^{\circ}_{Ag^+/Ag} - E^{\circ}_{Cu^{2+}/Cu} + \frac{2.303\,RT}{2F}\right) + \frac{2.303\,RT}{F}\log[Ag^+] \qquad .....(9.5)
\end{aligned}
$$

$$
\begin{aligned}
&-\frac{2.303\,RT}{2F}\log(1.000 \times 10^{-1}) \\
&= -\frac{2.303\,RT}{2F}(-1) \\
&= \frac{2.303\,RT}{2F}
\end{aligned}
$$

At a given temperature, the value of the term within brackets is constant. Thus, by measuring the emf of the cell, the value of $[Ag^+]$ can be calculated using Eq. 9.5. A more accurate method of obtaining $[Ag^+]$ in solutions of unknown concentration from E_{cell} values is to plot a calibration curve between $[Ag^+]$ values of silver nitrate solutions of known concentration and the corresponding E_{cell} values. This is explained below.

Plotting a calibration curve

For the purpose of drawing calibration curve, you shall use the following cell:

$$Cu(s)|\,Cu^{2+}(1.000 \times 10^{-1}\,M)\,|\,|\,Ag^+(c)|Ag(s) \qquad .....(9.6)$$
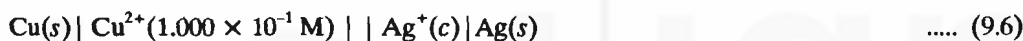
where $Ag^+(c)$ is the concentration of standard silver nitrate solution. You shall use $AgNO_3$ solutions of $5.000 \times 10^{-2}\,M$, $1.000 \times 10^{-2}\,M$, $5.000 \times 10^{-3}\,M$, $1.000 \times 10^{-3}\,M$, $5.000 \times 10^{-4}\,M$ and $1.000 \times 10^{-4}\,M$ concentration and measure E_{cell} every time and note down the values in Table 9.2.

The actual plot as per Eq. 9.5 should be between E_{cell} and $\log[Ag^+]$. Since $[Ag^+]$ in various $AgNO_3$ solutions is much less than 1, $\log[Ag^+]$ will be a negative quantity in each case. But it will be easier to plot a graph using positive quantities. Hence, Eq. 9.5 is transformed into Eq. 9.7 by multiplying the last term in right hand side of Eq. 9.5 by (–1) twice.

$$E_{cell} = \left(E^{\circ}_{Ag^+/Ag} - E^{\circ}_{Cu^{2+}/Cu} + \frac{2.303RT}{2F}\right) - \frac{2.303\,RT}{F}\left(-\log[Ag^+]\right) \qquad .....(9.7)$$

You should use this equation for the purpose of obtaining the calibration curve. Using E_{cell} and $-\log[Ag^+]$ values from Table 9.2, plot the calibration curve. Shape of the curve should be as shown in Fig. 9.1.
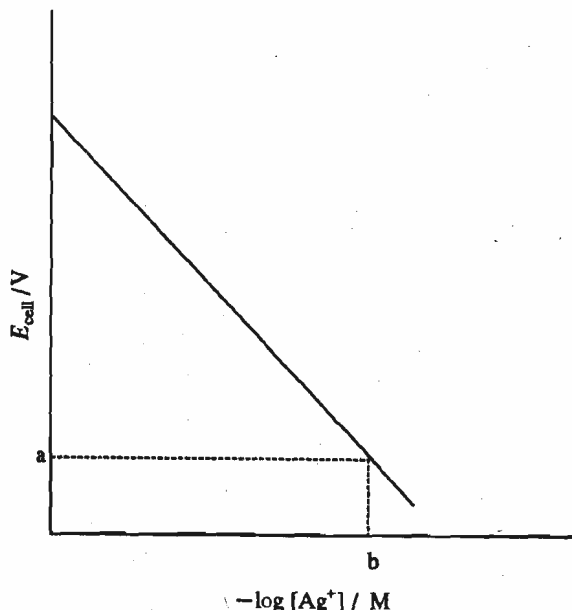
You can note that Eq. 9.5 is applicable to both the cells set up as per Eq. 9.4 and 9.6 since
i) the reference electrode is the same in each case, and
ii) the indicator electrode in each case is reversible with respect to Ag^+ ions.

To understand this clearly, you are advised to go through Section 17.8 of Unit 17 of the CHE-04 course.

Fig. 9.1 : Calibration curve and $[Ag^+]$ by interpolation

Obtaining unknown [Ag$^+$] from the calibration curve

The emf of the cell set up as per Eq. 9.4 is measured to obtain [Ag$^+$] in a saturated solution of AgCl by interpolation method. You have studied the details of interpolation method in Section 2.6 of Unit 2 of this course. Following this method, find out the x-coordinate value (b) from the curve corresponding to the E_{cell} value (a) obtained for the cell mentioned in Eq. 9.4. Once the x-coordinate value (b) is known, [Ag$^+$] in a saturated solution of AgCl can be calculated using the following relationship:

$$-\log [Ag^+] = b$$

Hence, $[Ag^+] = 10^{-b}$ M .....(9.8)

(as shown in Example 8 discussed in Sub-section 2.3.2 of Unit 2 of Block 1 of this course).

Calculation of K_{sp}

[Cl$^-$] = 3.333 × 10^{-2} M as shown in the calculation part of this experiment.

Using [Ag$^+$] value obtained above by interpolation method and [Cl$^-$] as 3.333×10^{-2} M, we can calculate K_{sp} using Eq. 9.3.

K_{sp} = ([Ag$^+$] as obtained by interpolation method) × ([Cl$^-$] in the indicator half-cell)

or $K_{sp} = (10^{-b} \times 3.333 \times 10^{-2})$ .....(9.9)

9.3.2 Requirements

Apparatus

|		
1.	Low cost potentiometer	1 No.
2.	Copper electrode	1 No.
3.	Silver electrode	1 No.
4.	Salt bridge/KNO$_3$	1 No.
5.	Beaker 100 cm^3	4 No.
6.	Graduated pipette 20 cm^3	1 No.
7.	Graduated pipette 10 cm^3	1 No.

Stock Solutions Provided

1. 1.000×10^{-4} M, 5.000×10^{-4} M, 1.000×10^{-3} M, 5.000×10^{-3} M, 1.000×10^{-2} M, 5.000×10^{-2} M and 1.000×10^{-1} M AgNO$_3$ solutions
2. 1.000×10^{-1} M CuSO$_4$ solution
3. 1.000×10^{-1} M KCl solution

9.3.3 Procedure

1. Calibrate the low cost potentiometer as you did in Experiment 8 by following the method given in Sub-section 3.4.6 of Block 1 of this course.

2. Set up the following galvanic cell (as per Eq. 9.6) using 1.000×10^{-1} M CuSO$_4$ and 1.000×10^{-4} M AgNO$_3$ solutions, copper and silver electrodes, and KNO$_3$ salt bridge:

   Cu(s)| Cu^{2+}(1.000 × 10^{-1} M) | | Ag$^+$(1.000 × 10^{-4} M) | Ag(s)
   Reference half-cell   Indicator half-cell
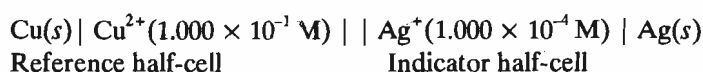
3. Connect the copper electrode to the reference terminal (through the black crocodile clip) and the silver electrode to the indicator terminal (through the red crocodile clip) of the instrument.

4. Keep the range selector at 2 V and then measure the emf of the cell. Remember each division on the meter scale is equal to 0.04 V in the 2 V range. Note down the measurement in Table 9.2 and remove the salt bridge as soon as possible.

5. Now take 5.000×10^{-4} M silver nitrate solution in the indicator half-cell. Connect the two half-cells through the salt bridge and note down the emf. Similarly measure the emf for the rest of the silver nitrate solutions (1.000×10^{-3} M, 5.000×10^{-3} M, 1.000×10^{-2}M and 5.000×10^{-2} M) and record the values in Table 9.2.

6. Prepare a system by mixing 10 cm^3 of 1.000×10^{-1} M AgNO$_3$ solution with 20 cm^3 of 1.000×10^{-1} M KCl solution. Shake the mixture well. Take this mixture in the indicator half- cell and set up a galvanic cell as shown in Eq. 9.4.

7. Measure the emf of the cell and record the value below Table 9.2.

8. Repeat the steps 6 and 7 once more and record the value of emf.

9.3.4 Observations

Room temperature =   °C

Table 9.2 : Emf Values of Cells constructed as per Eq. 9.8

S.No.	$[Ag^+]/M$	E_{cell}/V	$\log [Ag^+]/M$	$-\log [Ag^+]/M$
1.	1.000×10^{-4}			
2.	5.000×10^{-4}			
3.	1.000×10^{-3}			
4.	5.000×10^{-3}			
5.	1.000×10^{-2}			
6.	5.000×10^{-2}			

Emf of the cell constructed as per Eq. 9.4 is measured twice and the values are recorded below.

(i) V

(ii) V

The average of the two measurements = ... V

We shall use this as a, for the purpose of interpolation.

9.3.5 Calculations

(i) Plotting of calibration curve

Plot a calibration curve between E_{cell} and -log $[Ag^+]$ as explained in Sub-section 9.3.1. For this purpose, use the values from Table 9.2. You will get a calibration curve as shown in Fig. 9.1.

(ii) Obtaining $[Ag^+]$ from the calibration curve

As explained in Sub-section 9.3.1, obtain x- coordinate (b) corresponding to the average E_{cell} value (a) of the cell constructed as per Eq. 9.4.

b = ...

Then using Eq. 9.8, $[Ag^+]$ = 10^{-b}M

= .... M

(iii) Calculation of [Cl⁻]

Volume of 1.000×10^{-1} M KCl taken
in the indicator half-cell $= 20 \text{ cm}^3$

Amount of KCl present $= \dfrac{1.000 \times 10^{-1} \times 20}{1000} = 2.000 \times 10^{-3}$ mol

Volume of 1.000×10^{-1} M AgNO₃ taken $= 10 \text{ cm}^3$

Amount of AgNO₃ present $= 1.000 \times 10^{-1} \times \dfrac{10}{1000} = 1.000 \times 10^{-3}$ mol

Now, 1.000×10^{-3} mol of KCl will react with 1.000×10^{-3} mol of AgNO₃ to give 1.000×10^{-3} mol of AgCl and 1.000×10^{-3} mol of KNO₃.

Hence, amount of KCl left in the solution

$= 2.000 \times 10^{-3}$ mol $- 1.000 \times 10^{-3}$ mol

$= (2.000 - 1.000) \times 10^{-3}$ mol

$= 1.000 \times 10^{-3}$ mol

Volume of the solution $= (10 + 20) \text{ cm}^3 = 30 \text{ cm}^3$

Molar concentration of Cl⁻ $= \dfrac{1.000 \times 10^{-3}}{30} \times 1000$ M

or $\quad$ [Cl⁻] $= 3.333 \times 10^{-2}$ M

(iv) Calculation of K_{sp}

Now using Eq. 9.3, we can calculate the solubility product of AgCl as shown below:

$K_{sp} = [Ag^+][Cl^-] = (10^{-b} \times 3.333 \times 10^{-2})$

$= \dots\dots\dots\dots\dots\dots\dots\dots$

9.3.6 Result

The solubility product of silver chloride at .... °C = ...

Literature value at 25°C $= 1.5 \times 10^{-10}$

9.4 SUMMARY

In this unit, you had studied two experiments – one for determining the pH of solutions and the other for determining the solubility product of a sparingly soluble salt silver chloride. These experiments are based upon the applications of measurement of emf of galvanic cells.

Hydrochloric acid : 1.00×10^{-2} M : Using a burette or a measuring cylinder, transfer in small lots 9 cm^3 of concentrated hydrochloric acid to a 500 cm^3 volumetric flask which contains about 250 cm^3 of deionised water. Shake the flask after the addition of each lot. Make up the solution to the mark with deionised water. Let this solution, say solution (A),is approximately 0.1 M. Now titrate this solution against 1.000×10^{-1} M sodium carbonate solution using methyl orange indicator. Calculate the molarity of HCl solution (A) by using the following formula:

$$2\, M_{Na_2CO_3} \cdot V_{Na_2CO_3} = M_{HCl} \cdot V_{HCl}$$

Let the molarity of the HCl solution (A) be x mol dm^{-3}.

Prepare HCl solution of 1.00×10^{-2} M concentration by diluting appropriate volume of HCl solution (A) with deionised water. Volume of solution (A) required can be calculated by using the formula,

$$M_1V_1 = M_2V_2$$

or $\quad V_1 \quad = \dfrac{M_2 V_2}{M_1}$

Thus, volume of HCl solution (A) required to prepare 1 dm^3 of 1.00×10^{-2} M HCl

$$= \frac{1.00 \times 10^{-2}\ M \times 1000\ cm^3}{x\ M} = \frac{10}{x}\ cm^3$$

Therefore, take $10/x$ cm^3 of HCl solution (A) and dilute to 1 dm^3 with deionised water in a volumetric flask. The concentration of this HCl solution would be 1.00×10^{-2} M.

Acetic acid : 1.00×10^{-2} M : Using a burette, transfer 1.5 cm^3 of glacial acetic acid into a 250 cm^3 volumetric flask containing 100 cm^3 of deionised water and make up the solution to the mark. This solution say, acetic acid (A), is approximately 0.1 M. Find out the exact molarity of this solution by titrating against a standard solution of NaOH using phenolphthalein indicator. Let the exact molarity of acetic acid (A) be y mol dm^{-3}.

Now prepare 1.00×10^{-2} M acetic acid by diluting appropriate volume of acetic acid (A) with deionised water. Volume of acetic acid (A) required can be calculated by using molarity equation $M_1V_1 = M_2V_2$. Thus, volume of acetic acid (A) required to prepare 1 dm^3 of 1.00×10^{-2} M acetic acid $= \dfrac{1.00 \times 10^{-2}\,M \times 1000\ cm^3}{y\ M}$

$$= \frac{10\ cm^3}{y}$$

Therefore, take $10/y$ cm^3 of acetic acid (A) and dilute with deionised water to 1 dm^3 in a volumetric flask. This would give an acetic acid solution of concentration 1.00×10^{-2} M.

1.000×10^{-1} M sodium carbonate solution can be prepared by dissolving 1.325 g of anhydrous sodium carbonate in minimum quantity of deionised water and making up to the mark in a 250 cm^3 volumetric flask.

Details of preparation of standard NaOH solution are given in Appendix-II to Unit 12 of this Block.

Copper sulphate : 1.000×10^{-1} M : Dissolve 12.484 g of A.R. $CuSO_4.5H_2O$ in deionised water and make up the volume to 500 cm^3 in a volumetric flask. Before making up the volume, add 5cm^3 of acetic acid to prevent hydrolysis.

Potassium chloride : 1.000×10^{-1} M : Dissolve 3.728 g of A.R. KCl in deionised water and make up the volume to 500 cm^3 in a volumetric flask.

Silver nitrate : 1.000×10^{-1} M : Dissolve 16.987 g of A.R. $AgNO_3$ in deionised water and make up the volume to 1 dm^3 in a volumetric flask.

Silver nitrate : 5.000×10^{-2} M : Dilute 250 cm^3 of 1.000×10^{-1} M $AgNO_3$ to 500 cm^3 with deionised water in a volumetric flask.

Silver nitrate : 1.000×10^{-2} M : Dilute 50 cm^3 of 1.000×10^{-1} M $AgNO_3$ to 500 cm^3 with deionised water in a volumetric flask.

Silver nitrate : 5.000×10^{-3} M : Dilute 25 cm^3 of 1.000×10^{-1} M $AgNO_3$ to 500 cm^3 with deionised water in a volumetric flask.

Silver nitrate : 1.000×10^{-3} M : Dilute 10 cm^3 of 5.000×10^{-2} M $AgNO_3$ to 500 cm^3 with deionised water in a volumetric flask.

Silver nitrate : 5.000×10^{-4} M : Dilute 25 cm^3 of 1.000×10^{-2} M $AgNO_3$ to 500 cm^3 with deionised water in a volumetric flask.

Silver nitrate : 1.000×10^{-4} M : Dilute 10 cm^3 of 5.000×10^{-3} M $AgNO_3$ to 500 cm^3 with deionised water in a volumetric flask.