

EXPERIMENT 7: CONDUCTOMETRIC AND POTENTIOMETRIC TITRATIONS

Before the experiment: *Read the booklet carefully. Be aware of the safety issues.*

a. CONDUCTOMETRIC TITRATION

Object

To investigate neutralization and precipitation titrations by observing the change in electrical conductivity.

Theory

A reaction that causes change in the number or type of ions present in a solution can be observed by monitoring the electrical conductivity of the solution [1]. The titration or precipitation reactions involve the addition of a reagent to a solution that contains a substance with unknown molarity. Just before and after the equivalence point, there is a marked difference in the rate of change of conductance with incremental addition of reagent. It is then possible to find the amount of reagent equivalent to the solution titrated by measuring the conductance. The use of an excess of the standard reagent establishes the correct location of the equivalence point. [2]

Conductance of a solution, denoted as G , equals to the inverse of its resistance (R). The unit of conductance is Ω^{-1} , or officially denoted as S (siemens) [1]. There are some advantages of conductometric titrations [3];

- i. No need of indicator.
- ii. End point can be identified graphically with minimum error.
- iii. Can be used with colored liquids for which ordinary indicators cannot work.
- iv. Can be used for dilute solutions as well as for very weak acids.

Consider first the titration of a strong acid like HCl with a strong base like NaOH. Before the addition of the base, the acid solution has a high conductance. As alkali is added, the hydrogen ions are removed to form water, and they are replaced by the much slower cations of the base. Consequently, the conductance of the solution decreases, and keeps on falling with addition of the

base until the equivalent point is reached. Further addition of alkali introduces now an excess of the fast hydroxyl ions, and they causes the conductance to rise again. When this variation of the conductance of the solution is plotted against the volume of alkali added the result looks like the curve ABC shown in Figure 1. Here the ordinate represents the conductance $1/R$. At the minimum there is no excess present of either acid or base, and hence it is the equivalence point.

When the acid is weak, say acetic acid, and the base is strong, the titration curve has the general form of A'B'C' in Figure 1. Since the acid is weak, its conductance is relatively low. As the base is added, the poorly conducting acid is converted to highly ionized salt, and consequently the conductance goes up along A'B'. Once the acid is neutralized, addition of excess base causes another sharp increase in conductance, and curve rises along B'C'[2].

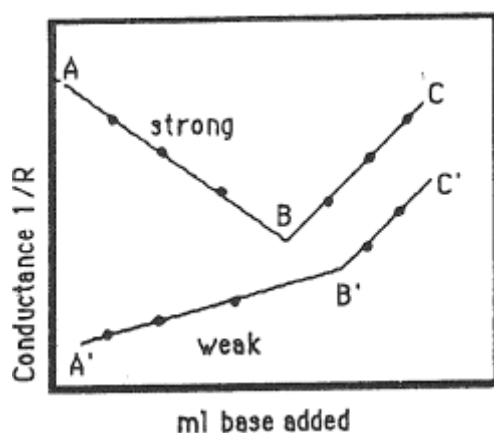


Figure 1. Conductometric titrations of acids with strong base.

In order to find the concentration of a component used during the titration, mole balance can be done at the neutralization point. At the neutralization point, number of moles of hydrogen ions in the medium is equal to the number of moles of hydroxyl ions. Hence, mole balance at the neutralization points yields:

$$n_{H^+} = n_{OH^-} \quad (1)$$

$$M_{acid} \times V_{acid} = M_{base} \times V_{base} \quad (2)$$

where M_i is the molarity of the solution, and V_i is the volume of the solution used [4].

Experimental Work

Apparatus and Chemicals: Conductivity meter, 1 ml and 10 ml pipettes, 200 ml beakers, 100 ml graduated cylinder, 0.1 M CH_3COOH , 0.1 M HCl , 0.1 M NaOH , distilled water.

Procedure

a) Titration of CH_3COOH with NaOH

150 ml of distilled water is added on 5 ml CH_3COOH solution and its conductivity is measured. Then it is measured after addition of 1 ml 0.1 M NaOH . This is repeated until total volume of 10 ml NaOH is added.

b) Titration of HCl with NaOH

150 ml of distilled water is added on 5ml of HCl solution and its conductivity is measured. Then it is measured after addition of 1 ml 0.1 M NaOH solution. This is repeated until 10 ml of NaOH is added.

Safety Issues: 0.1 M hydrochloric acid (HCl), 0.1 M acetic acid (CH_3COOH), and 0.1 M sodium hydroxide (NaOH) solutions are used in this experiment. 0.1 M hydrochloric acid solution is hazardous in case of eye contact and slightly hazardous in case of skin contact, ingestion and inhalation. 0.1 M acetic acid solution is slightly hazardous in case of skin and eye contact, ingestion, and inhalation. 0.1 M sodium hydroxide solution is hazardous in case of skin and eye contact, and ingestion and slightly hazardous in case of inhalation. Be aware of the risks when handling the chemicals. In case of eye contact, remove any contact lenses and flush eyes with plenty of water. In case of skin contact, flush skin with plenty of water, remove contaminated clothes, and cover the irritated skin with an emollient [5, 6, 7]. Use splash goggles, and wear lab coats and gloves during the experiment.

Calculations

1. For each experiment, draw the conductance ($1/R$) vs amount of reacting agent added (ml).
2. Find the concentration of CH_3COOH and HCl solutions.

b. POTENTIOMETRIC TITRATION

Object

To study the titrations of solutions of acids and bases by observing the change in their pH values.

Theory

A reaction, in which hydrogen ions are produced or consumed, can be followed by observing the pH of the reaction. Negative logarithm of the concentration of the H^+ cations gives the pH of the solution [4]:

$$pH = -\log[H^+] \quad (3)$$

As in the first part of the experiment, potentiometric titration can serve to determine the equivalence point or end point of titration. As the titration is carried out, as long as the acid is in excess, the pH will vary slightly with addition of base. However, near the equivalence point, concentration of the H^+ cations fall rapidly. In Figure 2 at curve (a), pH change is shown, during the titration of a strong acid with a strong base. Curve (b) presents the differential pH change during the titration, and the peak point is the end point of the titration. [2, 4]

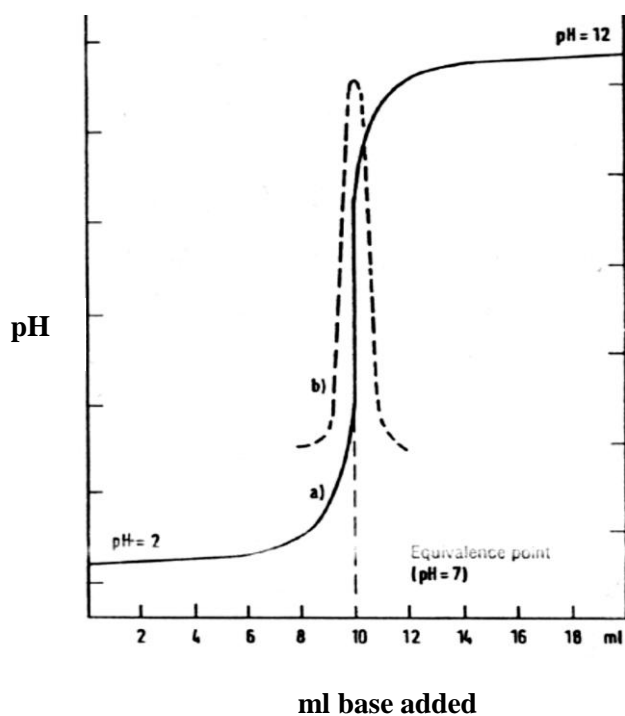


Figure 2. Potentiometric acid-base titration [2].

Experimental Work

Apparatus and Chemicals: pH meter, 1 ml and 5 ml pipettes, 200 ml beakers, 100 ml graduated cylinders, 0.1 M H₂SO₄, 1 M NaOH, distilled water.

Procedure

Titration of H₂SO₄ with NaOH

50 ml of distilled water is added into 5 ml of H₂SO₄ solution. Initial pH is read by pH meter. 0.2 ml of 1 M NaOH is added into the solution and the new value of pH is recorded. This is repeated until a total of 2 ml of NaOH is added.

Safety Issues: 0.1 M sulfuric acid (H₂SO₄), and 1 M sodium hydroxide (NaOH) solutions are used in this experiment. 0.1 M sulfuric acid solution is extremely hazardous in case of eye and skin contact, ingestion and inhalation. Skin contact may cause burns. 1 M sodium hydroxide solution is very hazardous in case of skin and eye contact, and ingestion and slightly hazardous in case of inhalation. Be aware of the risks when handling the chemicals. In case of eye contact, remove any contact lenses and flush eyes with plenty of water. In case of skin contact, flush skin with plenty of water, remove contaminated clothes, and cover the irritated skin with an emollient [7, 8]. Use splash goggles, and wear lab coats and gloves during the experiment.

Calculations

1. Draw the graphs of:
 - a) pH vs amount V of reagent added (ml)
 - b) [$\Delta(\text{pH})/\Delta V$] vs V of reagent added (ml)
2. Find the concentration of H₂SO₄ solution.

References

- [1] Atkins, P. W. , Physical Chemistry, 9th Ed., Oxford University Press, Oxford, 2009, Chapters 20 and 21.
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