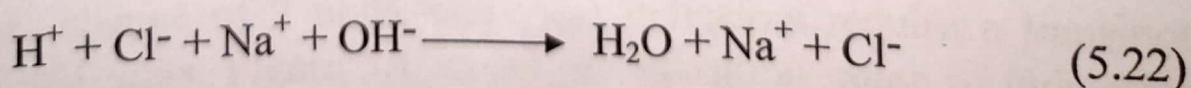


5.3.12 Acid–Base Titrations

An acid-base titration involves neutralization reaction in which an acid is reacted with an equivalent amount of base. If the process involves the determination of the strength of an acid solution by titrating it with a standard alkali it is termed as acidimetry, while the reverse process is termed as alkalimetry.

For such titrations a standard solution is essentially required. To prepare standard solutions it is necessary to have a primary standard (one which can be prepared from the substance directly with a sufficient degree of accuracy). It is prepared by weighing out the necessary amount of the substance and dissolving it in a measured volume of water to obtain the desired concentration. Some such as oxalic acid, sodium carbonate etc. can be obtained in a state of high purity and can be weighted accurately. For others such as HCl and H₂SO₄ a solution of approximate strength has to be prepared first (Box 5.7) and then its exact strength can be determined by titrating against a primary standard solution. Once the exact strength of this solution is known, it can be used as a primary standard solution.

In the case of a strong acid versus a strong base, both the titrant and the analyte are completely ionized. An example is the titration of hydrochloric acid with sodium hydroxide:



Initially, the pH is low, but as titration proceeds part of the H⁺ is removed from solution as H₂O. So the concentration of H⁺ gradually

decreases, on the contrary the pH increases. As the equivalence point is approached, the concentration reduces rapidly until at the equivalence point, when the neutralization is complete, a neutral solution of NaCl remains and the pH is 7. The point at which the reaction is complete is called the end point and is chosen such that it coincides with or is very close to the equivalence point.

The most obvious way of determining the end point is to measure the pH at different points of the titration. However, a more convenient way is to add an indicator to the solution and visually detect a colour change. An indicator for an acid-base titration is either a weak acid or a weak base that is highly coloured. These substances are usually composed of highly conjugated organic compounds that give rise to the colour. The colour of the ionized form is markedly different from the unionized form.

Assuming the indicator to be a weak acid, designated as HIn and that the unionized form is red while the ionized form is blue; then:



we can write a Henderson – Hasselbach equation for this as follows:

$$\text{pH} = \text{pka} + \log \frac{[\text{In}^-]}{[\text{HIn}]} \quad (5.24)$$

The indicator changes colour over a pH range. With indicators in which both forms are coloured, generally one colour is observed if the ratio of the concentration is 10:1 and only the colour of the more concentrated form is seen. We can calculate the pH transition range from one colour to the other. When the colour of the unionized form is seen, $[\text{In}^-] / [\text{HIn}] = 1/10$. Therefore,

$$\text{pH} = \text{pka} + \log 1 / 10 = \text{pka} - 1 \quad (5.25)$$

when only the colour of the ionized form is observed

$$\text{pH} = \text{pka} + \log 10 / 1 = \text{pka} + 1$$

So the pH in moving from one colour to the other has changed from to . This is a pH change of 2 units, and most indicators require a transition change of 2 pH units. Figure 5.1 shows the transition range of indicators, while Figure 5.2 shows some of the apparatus required for volumetric analysis.