Titration Curves

SCH3U

Reactions in solution are used for analysis as well as synthesis. Titration is an important method for determining the amount of a substance present in solution. A solution of known concentration is called a **standard solution**.

The purpose of a titration is usually to standardize a solution, i.e. to determine the concentration of an acid or a base; and some times to determine the molar mass of and unknown acid. The standard solution is usually added from a burette to the solution being analyzed in the Erlenmeyer flask.

In order to determine these, the equivalence point of the titration must be determined.

The equivalence point may be determined by plotting a titration curve or by the use of an indicator.

The **equivalence point** is when stoichiometrically equivalent amounts of an acid and base required by the equation for the reaction have been combined.

End point is when the indicator changes colour.

A substance whose change in colour shows that the end point has been reached is called an **indicator.** The indicator is chosen so that the end point takes place as near the equivalence point as possible.

Titrant is the solution being slowly added from a burette to a solution in the receiving flask.

A **titration curve** is when the pH of a solution at different stages of a titration is plotted against the volume of titrant added; i.e. a titration curve for an acid– base titration is a graph of pH against the number of milliliters of acid or base added in the titration.

The titration curve is obtained by calculating the pH of the solution after a certain volume of the titrant has been added. A titration curve gives the clearest picture of what happens during a titration

Usually the pH is determined:

1. Initially, i.e. when no titrant has been added
2. Between the start and the equivalence point
3. At the equivalence point
4. After the equivalence point

Titration Curve for Strong Acid – Strong Base

Let us determine the titration curve when:

25.0 mL of 0.100 mol L^{-1} $HCl_{(aq)}$ is titrated using 0.100 mol L^{-1} $NaOH_{(aq)}$

Equation for the reaction: $HCl_{(aq)} + NaOH_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)}$

Net -ionic equation: $H_3O^{+1}{}_{(aq)} + OH^{-1}{}_{(aq)} \longrightarrow H_2O_{(l)}$

1. **Calculate the Initial pH**

25.0 mL of 0.100 mol L^{-1} $HCl_{(aq)}$

.: $[H_3O^{+1}] = 0.100$ mol L^{-1}

.: $pH = -\log [H_3O^{+1}]$

$= 1$

2. **Calculate the pH after the addition of 10.0 mL of 0.100 mol L^{-1} $NaOH_{(aq)}$**

mol (HCl) = c v = (0.100 mol L^{-1}) . (0.025 L) = 2.50×10^{-3} mol

mol (NaOH) = c v = (0.100 mol L^{-1}) . (0.010 L) = 1.00×10^{-3} mol

The base neutralizes some of the acid. One $H_3O^{+1}{}_{(aq)}$ removes one $OH^{-1}{}_{(aq)}$ added, (see net-ionic equation).

.: amount of HCl remaining after reaction = (2.50×10^{-3} mol) — (1.00×10^{-3} mol)
= 1.50×10^{-3} mol

concentration HCl remaining after reaction = $\dfrac{\text{number of mol of HCl}}{\text{Total volume of solution}} = \dfrac{1.50 \times 10^{-3}\ \text{mol}}{(25.0 + 10.0)/1000\ \text{L}}$

= 0.0429 mol L^{-1}

.: $[H_3O^{+1}]$ = 0.0429 mol L^{-1}

pH of solution after the addition of 10.0 mL of 0.100 mol L^{-1} $NaOH_{(aq)}$ = — log $[H_3O^{+1}]$
= — log [0.429] = 1.37

.: pH of solution after the addition of 10.0 mL of 0.100 mol L^{-1} $NaOH_{(aq)}$ = 1.37

3. **Calculate the pH at the equivalence point: after the addition of 25.0 mL of 0.100 mol L^{-1} $NaOH_{(aq)}$ to 25.0 mL of 0.100 mol L^{-1} $HCl_{(aq)}$**

All the acid has been neutralized, but no excess base has been added, the solution contains only salt and water.

Complete neutralization has occurred: $[H_3O^{+1}]$ = $[OH^{-1}]$.: pH = 7

.: **pH at the equivalence point: after the addition of 25.0 mL of 0.100 mol L^{-1} = 7**

4. **After the addition of 35.0 mL of 0.100 mol L^{-1} $NaOH_{(aq)}$ to 25.0 mL of 0.100 mol L^{-1} $HCl_{(aq)}$**

After the equivalence point, excess base is added to the neutral salt solution.

mol (NaOH) = c v = (0.100 mol L^{-1}) . (0.035 L) = 3.50×10^{-3} mol

mol (NaOH) in excess = (3.50×10^{-3} mol) – (2.50×10^{-3} mol) = 1.00×10^{-3} mol

$[OH^{-1}] = \dfrac{\text{number of mol of NaOH}}{\text{Total volume of solution}} = \dfrac{1.00 \times 10^{-3}\ \text{mol}}{0.060\ \text{L}}$ = 0.0167 mol L^{-1}

pOH = – log $[OH^{-1}]$ = – log (0.0167) = 1.78 pH = 14 – 1.78 = 12.22

.: pH after the addition of 35.0 mL of 0.100 mol L^{-1} $NaOH_{(aq)}$ = 12.22

Summary: Titration curve of a Strong Acid with a Strong Base

Initial pH = 1

pH of solution after the addition of 10.0 mL of 0.100 mol L^{-1} $NaOH_{(aq)}$ = 1.37

pH at the equivalence point: after the addition of 25.0 mL of 0.100 mol L^{-1} = 7

pH after the addition of 35.0 mL of 0.100 mol L^{-1} $NaOH_{(aq)}$ = 12.22

As can be seen during the first part of the titration, the pH increases slowly and gradually.

The concentration of the hydrogen ion is large at the beginning of the titration, and the addition of hydroxide ions produces only a small change in $[H^{+1}]$.

Near the equivalence point, the concentration of the hydrogen ion is very small; the addition of the OH^{-1} produces a large proportional change in $[H^{+1}]$.

There is a very rapid change in pH around the equivalence point.

At the equivalence point, addition of one drop (about 0.05 mL) of 0.100 mol L^{-1} $NaOH_{(aq)}$ results in a change in pH of more than 5.2 units, a change in $[H^{+1}]$ of about 100 000-fold.

Indicators usually change colour over a pH range of about 2 units.
If an indicator is to signal the equivalence point of a titration, an indicator must be chosen that will change colour when the stoichiometric quantity (indicated by the equation) of the standard solution has been added.
Any indicator that changes colour in the pH range of 3.5 – 10.5 is suitable for measuring the equivalence point of the titration of a strong acid with a strong base.
Bromocresol green, methyl red, litmus, bromothymol blue, and phenolphthalein are all suitable.

Analysis of a Titration of a Strong Acid with a Strong Base

Assume that you have 25.0 cm^3 of 0.100 mol dm^{-3} hydrochloric acid, $HCl_{(aq)}$, in an Erlenmeyer flask.
You are in the process of titrating this with 0.100 mol dm^{-3} sodium hydroxide, $NaOH_{(aq)}$.
Fill in the following table.
Plot a titration curve for the strong acid – strong base titration being performed, using the data in the table below.
Plot the "pH " along the y - axis of a graph and "volume of base added" along the x – axis.
Indicate the equivalence point on the graph. Indicate a suitable indicator for this titration.

# mol 0.100 mol L^{-1} $HCl_{(aq)}$	Volume of 0.100 mol L^{-1} $NaOH_{(aq)}$	# mol $NaOH_{(aq)}$ added	# mol H^{+1} in excess	# mol OH^{-1} in excess	Total flask volume (L)	[H^{+1}] in flask in excess	pH	[OH^{-1}] in flask in excess, pOH, .: pH
	0.00							
	5.00							
	10.00							
	15.00							
	20.00							
	25.00							
	30.00							
	35.00							
	40.00							