

Acid-Base Titrations:

Introduction to Acid-Base Titrations

A **titration** is a procedure used in analytical chemistry to determine the amount or concentration of a substance. In a titration one reagent, the **titrant**, is added to another slowly. As it is added a chemical stoichiometric reaction occurs until one of the reagents is exhausted, and some process or device signals that this has occurred. The purpose of a titration is generally to determine the quantity or concentration of one of the reagents, that of the other being known beforehand. In any titration there must be a rapid quantitative reaction taking place as the titrant is added, and in acid-base titrations this is a stoichiometric neutralization. The type of titration is simply the type of chemical reaction taking place, and so in this section we consider **acid-base titrations**.

Acid-Base Titration Reactions

All acid-base titration reactions, as all acid-base reactions, are simply exchanges of protons. The reaction could be strong acid + strong base \rightarrow (neutral) salt, as in the case of $\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O}$, although the reaction would be correctly written as $\text{H}_3\text{O}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$ since strong acids and strong bases are totally dissociated to protons and hydroxide ions in water. For reactions which are strong acid + weak base \rightarrow (acidic) salt, such as the example $\text{HCl} + \text{CH}_3\text{NH}_2 \rightarrow \text{CH}_3\text{NH}_3^+\text{Cl}^-$, or strong base + weak acid \rightarrow (basic) salt, such as the example $\text{NaOH} + \text{CH}_3\text{COOH} \rightarrow \text{Na}^+\text{CH}_3\text{COO}^- + \text{H}_2\text{O}$, the cations and anions could be omitted as they do not actually participate in the reaction. (Some chemists call these **bystander ions**.)

Virtually all acid-base titrations are carried out using a strong acid or strong base. In most cases the strong acid or strong base is used as the titrant. It is less common, but equally feasible, to place the strong acid or strong base in the titration vessel and use the weak acid or weak base as the titrant. A weak acid-weak base titration would have only a small pH change at the equivalence point. This small change is difficult to detect, and for this reason weak acid-weak base titrations are uncommon.

Standards in Acid-Base Titrations

One of the substances involved in a titration must be used as a standard for which the amount of substance present is accurately known. The standard can be present either in the form of a pure substance or as a standard solution, which is a solution whose composition is accurately known. A standard can be prepared in only two ways: use a primary standard or standardize by titration against some previously standardized solution. A **primary standard** is some substance such as oxalic acid which can be precisely weighed out in pure form, so that the number of moles present can be accurately determined from the measured weight and the known molar mass. For example, we might prepare a 0.1000 molar solution of primary standard oxalic acid by weighing out exactly 0.1 moles of oxalic acid and diluting to one litre in a volumetric flask.

The standard solutions used in an acid-base titration need not always be primary standards. A standard solution which has been prepared by quantitative dilution of a

primary standard is an excellent secondary standard solution. Secondary standards can also be prepared by titration against a primary standard solution.

Stoichiometry of Acid-Base Titrations

In an acid-base titration, we slowly add the titrant strong acid or strong base until the equivalence point is reached as indicated in the previous section. The equivalence point is that point at which the number of moles of acid or base added as titrant is exactly equivalent to the number of moles of acid or base present originally in the other solution in accordance with the stoichiometric reaction.

Example. The equivalence point for the titration of 50.00 mL of 0.100 molar HCl with 0.200 molar NaOH could be calculated as follows: $50 \text{ mL} \times 0.1 \text{ mol/L} = 5.0 \text{ mmol HCl}$. The titration uses the stoichiometric reaction $\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O}$, which could just as accurately be written as $\text{H}_3\text{O}^+ + \text{OH}^- \rightarrow 2\text{H}_2\text{O}$. Since the reaction is a 1:1 reaction, 5.00 mmol of HCl are equivalent to 5.00 mmol NaOH. The volume of NaOH required can be calculated: $5.0 \text{ mmol NaOH} = 0.2 \text{ mol/L} \times V \text{ mL}$, $V = 5.00/0.200 = 25.00 \text{ mL NaOH}$

Example. In the titration of H_2SO_4 , sulfuric acid, the reaction requires 2 moles of NaOH per mole of H_2SO_4 . A complete titration of 50.00 mL of 0.100 molar H_2SO_4 would therefore require 50.00 mL of 0.200 molar NaOH rather than the 25.00 mL needed for the monoprotic acid HCl in the preceding example.

Detecting the Equivalence Point

In acid-base titrations, there is a sharp change in pH at the equivalence point. The titration of 2.5 mmol HCl (solid curve) and then that of 2.5 mmol CH_3COOH with 0.1 molar NaOH (dashed curve) are shown in the Figure below.

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The stoichiometric chemical reactions are $\text{Na}^+ + \text{OH}^- + \text{H}_3\text{O}^+ + \text{Cl}^- \rightarrow \text{H}_2\text{O} + \text{Na}^+ + \text{Cl}^-$ and $\text{Na}^+ + \text{OH}^- + \text{H}_3\text{O}^+ + \text{CH}_3\text{COO}^- \rightarrow \text{H}_2\text{O} + \text{Na}^+ + \text{CH}_3\text{COO}^-$. The reaction stoichiometry is 1:1 in both cases. The pH change can be detected by a pH meter, an electrochemical device whose discussion we will defer to later sections, or by a chemical indicator. **Chemical indicators are acid-base conjugate pairs whose acid form and base form are different in color.** A table of useful chemical indicators is given below.

Table: Properties of Aqueous Acid-Base Indicators at 25°C

Indicator	pH range	pKa	Acid Form	Base Form
methyl violet	0.0 - 1.6	0.8	yellow	blue
thymol blue	1.2 - 2.8	1.6	red	yellow
methyl yellow	2.9 - 4.0	3.3	red	yellow
methyl orange	3.1 - 4.4	4.2	red	yellow
bromocresol green	3.8 - 5.4	4.7	yellow	blue

methyl red	4.2 - 6.2	5.0	red	yellow
chlorophenol red	4.8 - 6.4	6.0	yellow	red
bromothymol blue	6.0 - 7.6	7.1	yellow	blue
phenol red	6.4 - 8.0	7.4	yellow	red
cresol purple	7.4 - 9.0	8.3	yellow	purple
thymol blue	8.0 - 9.6	8.9	yellow	blue
phenolphthalein	8.0 - 9.8	9.7	colorless	red
thymolphthalein	9.3 - 10.5	9.9	colorless	blue
alizarin yellow R	10.1 - 12.0	11.0	yellow	red
indigo carmine	11.4 - 13.0	12.2	blue	yellow

A chemical indicator is a compound which can change color to indicate that the endpoint of a titration has been reached. The color of an indicator changes because it is affected by the concentrations of ions in the solution. An acid-base indicator is an acid-base conjugate pair, a weak acid-weak base system in which the two forms have different colors. This is added, in low concentration so that it exerts essentially no control on the pH of the system. For an indicator, the acid ionization constant $K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$ is usually written as $K_a = \frac{[\text{H}_3\text{O}^+][\text{In}^-]}{[\text{HIn}]}$, where HIn is the acid form of the indicator and In^- is the base form of the indicator. Where the ratio $[\text{In}^-]/[\text{HIn}]$ is one, $[\text{H}_3\text{O}^+] = K_a$ and $\text{pH} = \text{p}K_a$. This point is in the color-change region of the indicator, usually at its center. The color-change region of the indicator is usually ± 1 pH unit around $\text{p}K_a$. To select a proper indicator, then, requires determination of the pH of the solution at the equivalence point of the titration and selection of an indicator whose $\text{p}K_a$ is as close to that pH as possible.

Indicators are usually organic structures of some complexity or natural products and, like dyeing materials, are generally known by their trivial names as in the above table. A few indicators, of which thymol blue is an example, are polyprotic acids which can change color more than once as pH is continuously increased. Such indicators can be used at a pH equal to either of their $\text{p}K_a$ values.

Equilibria and Acid-Base Titration Curves

The **titration curve** for an acid-base titration is a plot of the solution pH, normally on the vertical axis, against the volume of titrant added. We will investigate these curves by following a particular example in detail.

Example. Let us calculate the pH values along the course of a particular titration, the titration of exactly 100 mL of 0.100 molar acetic acid with 0.100 molar sodium hydroxide. In practice, we would probably titrate 25.00 mL instead for convenience, but these volumes are selected to simplify some of the calculations. In the course of the calculations and discussion, we will answer such questions as the titrant to use, the indicator to select, and what manner of solution is present at various points along the titration curve.

First, consider the titrant solution. Acetic acid is a weak acid, so a base is required to obtain any acid-base reaction. To get a pH break which is as large as possible at the equivalence point we should choose a strong base. The common strong bases are NaOH and KOH, so we choose one of these. It is normal to choose a concentration of base that will give us a reasonable volume to add, so that the volume can be accurately measured and yet not be too large. Here we choose 0.100 molar NaOH and

will therefore require about 100 mL of it, since the stoichiometric reaction is $\text{CH}_3\text{COOH} + \text{Na}^+ + \text{OH}^- \rightarrow \text{CH}_3\text{COO}^- + \text{Na}^+ + \text{H}_2\text{O}$.

The pH of the Starting Solution

At the start of the titration the solution contains only the weak acid CH_3COOH , for which

$$K_a = 1.75 \times 10^{-5} = [\text{H}_3\text{O}^+][\text{CH}_3\text{COO}^-]/[\text{CH}_3\text{COOH}]; K_a = [\text{H}_3\text{O}^+]^2/[\text{CH}_3\text{COOH}]$$

We know that $[\text{H}_3\text{O}^+]$ is approximately equal to $[\text{CH}_3\text{COO}^-]$, from the stoichiometric dissociation of the weak acetic acid and the fact that it is the major source of hydrated protons in the solution. Solving:

$$[\text{H}_3\text{O}^+]^2 = 1.75 \times 10^{-5} \times 0.1, [\text{H}_3\text{O}^+] = 1.32 \times 10^{-3}, \text{pH} = 2.88$$

Before the Equivalence Point

The pH slowly rises as the NaOH added reacts with the acetic acid. The acetic acid has been **one-quarter** titrated when 25 mL of the NaOH solution have been added. At this point 1/4 of the original moles of CH_3COOH have been titrated to CH_3COO^- , while 3/4 of the original moles of CH_3COOH remain, so $3[\text{CH}_3\text{COO}^-] = [\text{CH}_3\text{COOH}]$ and

$$K_a = 1.75 \times 10^{-5} = [\text{H}_3\text{O}^+][\text{CH}_3\text{COO}^-]/[\text{CH}_3\text{COOH}] = [\text{H}_3\text{O}^+]/3$$

$$[\text{H}_3\text{O}^+] = 3 \times 1.75 \times 10^{-5} = 5.25 \times 10^{-5}, \text{pH} = 4.28$$

Now consider the situation when **half** of the acetic acid has been titrated. On the volume axis, that would be at $1/2 \times 100 \text{ mL} = 50 \text{ mL}$ of added NaOH. In the solution at that point, half of the original acetic acid has been titrated and half has not, so $\text{mol CH}_3\text{COOH} = \text{mol CH}_3\text{COO}^-$ and, since both are present in the same solution, $[\text{CH}_3\text{COOH}] = [\text{CH}_3\text{COO}^-]$. As a consequence, $K_a = 1.75 \times 10^{-5} = [\text{H}_3\text{O}^+][\text{CH}_3\text{COO}^-]/[\text{CH}_3\text{COOH}]$ and $[\text{H}_3\text{O}^+] = 1.75 \times 10^{-5}$, $\text{pH} = 4.76$.

The acetic acid is **three-quarters** titrated when 75 mL of the NaOH solution have been added. At this point 3/4 of the original moles of CH_3COOH have been titrated to CH_3COO^- , while 1/4 of the original moles of CH_3COOH remain, so $[\text{CH}_3\text{COO}^-] = 3[\text{CH}_3\text{COOH}]$ and

$$K_a = 1.75 \times 10^{-5} = [\text{H}_3\text{O}^+][\text{CH}_3\text{COO}^-]/[\text{CH}_3\text{COOH}] = 3[\text{H}_3\text{O}^+]/1$$

$$[\text{H}_3\text{O}^+] = (1/3)(1.75 \times 10^{-5}) = 5.83 \times 10^{-6}, \text{pH} = 5.23$$

At the Equivalence Point

The equivalence point is reached after 100 mL of the NaOH solution have been added, just enough to react with all of the acetic acid present. Since the reaction is that of a weak acid plus a strong base to yield a weak base and water, the solution at the

equivalence point is simply a solution of the weak base CH_3COONa , sodium acetate. **By definition of the equivalence point no excess of either weak acid or strong base can be present.** The situation is therefore just as it would be for a solution of the weak base sodium acetate in water. Since there were originally 100 mL of 0.1 molar CH_3COOH , or 10 mmol CH_3COOH , there are now 10 mmol of CH_3COONa . These are contained in 200 mL of solution because we started with 100 mL and added another 100 mL, so the formal concentration of acetate is $10 \text{ mmol}/200 \text{ mL} = 0.05 \text{ molar}$.

The equilibrium constants are $K_a = 1.75 \times 10^{-5} = [\text{H}_3\text{O}^+][\text{CH}_3\text{COO}^-]/[\text{CH}_3\text{COOH}]$ and $K_b = 5.77 \times 10^{-10} = [\text{OH}^-][\text{CH}_3\text{COOH}]/[\text{CH}_3\text{COO}^-]$. It can no longer be assumed that $[\text{H}_3\text{O}^+]$ is approximately equal to $[\text{CH}_3\text{COO}^-]$, because the solution is now basic, containing the weak base acetate ion. The major source of the $[\text{H}_3\text{O}^+]$ is not the dissociation of CH_3COOH since there is virtually no CH_3COOH left to dissociate. However, it can now be assumed that $[\text{OH}^-]$ is approximately equal to $[\text{CH}_3\text{COOH}]$, because the major source of hydroxide ion is the weak base CH_3COO^- which hydrolyzes, giving CH_3COOH and OH^- in a stoichiometric 1:1 ratio, following the reaction equilibrium $\text{CH}_3\text{COO}^- + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{COOH} + \text{OH}^-$. To a good approximation, $K_b = [\text{OH}^-]^2/[\text{CH}_3\text{COO}^-]$. Since $[\text{CH}_3\text{COO}^-]$ is about 0.05 molar, $[\text{OH}^-]^2 = K_b[\text{CH}_3\text{COO}^-]$, $[\text{OH}^-] = 5.37 \times 10^{-6}$.

$$\text{pOH} = -\log(5.37 \times 10^{-6}) = 5.27, \text{pH} = 14.00 - 5.27 = 8.73$$

The equivalence point pH having been determined as 8.73, an indicator can now be chosen for this titration. Those indicators given in a [table of acid-base indicators](#) for which $\text{p}K_a$ is approximately equal to the pH at the equivalence point, 8.73, are thymol blue, whose $\text{p}K_a$ is 8.9, and cresol purple, whose $\text{p}K_a$ is 8.3. Either would be satisfactory for this titration. For thymol blue, the color change would be from the yellow color of the acid form to the blue color of the base form.

Notice that the pH slowly rises throughout that part of the titration curve prior to the equivalence point. If an indicator such as bromocresol green ($\text{p}K_a = 4.7$) had been chosen it would have changed from its acid color, yellow, to green well before a chemist was 3/4 of the way to the equivalence point. The titration results would be meaningless, because the endpoint at which color change occurs would not be easily relatable to the equivalence point.

Beyond the Equivalence Point

As NaOH addition continues beyond the equivalence point, the pH continues to rise. When twice as much NaOH has been added as is required, in other words when 200 mL of the NaOH solution have been added, the resulting solution is then a mixture of the strong base NaOH and the weak base CH_3COONa . In any mixture of a strong base and a weak base of comparable concentration, the strong base controls the pH, so that $[\text{OH}^-]$ is approximately equal to the excess concentration of added NaOH. There were added $200 \text{ mL} \times 0.1 \text{ molar} = 20 \text{ mmol}$ NaOH, but of this 10 mmol was used to react with the 10 mmol of acetic acid originally present. As a consequence, there are now $20 - 10 = 10 \text{ mmol}$ NaOH present in $100 + 200 = 300 \text{ mL}$ solution. Then $[\text{OH}^-] = 10 \text{ mmol}/300 \text{ mL} = 0.033 \text{ molar}$,

$$[\text{OH}^-] = 3.3 \times 10^{-2}, \text{pOH} = 1.48, \text{pH} = 14.00 - 1.48 = 12.52$$

The points calculated so far along this titration curve are shown in the Figure below.

The titration curve of an acid with a base will vary not only as the amount of the acid present varies, but also with the strength of the acid, as the Figure below shows. The plateau region of the titration curve shifts to higher pH values, and the magnitude of the pH break at the equivalence point therefore decreases, as the acid becomes weaker.

We now consider the comparable case of the titration of a weak base with a strong acid.

Example. A total of 50 mL of aqueous ammonia, NH_3 , solution which is approximately 0.01 molar is to be titrated. Let us consider what it should be titrated with, what indicator should be used, and what color change would then be seen.

To titrate this weak base, we could choose any strong acid, such as HCl , with a convenient concentration of about 0.01 molar. For ammonium ion, $K_a = 5.60 \times 10^{-10} = [\text{H}_3\text{O}^+][\text{NH}_3]/[\text{NH}_4^+]$.

At the equivalence point, the solution is essentially a solution of NH_4Cl . On hydrolysis one ammonium ion yields one hydrogen ion and one ammonia; $\text{NH}_4^+ \rightarrow \text{NH}_3 + \text{H}_3\text{O}^+$. To a good approximation, $[\text{H}_3\text{O}^+] = [\text{NH}_3]$, so $K_a = [\text{H}_3\text{O}^+]^2/[\text{NH}_4^+]$; $[\text{H}_3\text{O}^+] = \sqrt{K_a[\text{NH}_4^+]}$. There were $0.01 \text{ mmol/L} \times 50 \text{ mL} = 0.5 \text{ mmol}$ NH_4^+ ; this is in $50 + 50 = 100 \text{ mL}$ solution so $[\text{NH}_4^+] = 0.5 \text{ mmol}/100 \text{ mL} = 0.005 \text{ mmol/L}$ NH_4^+ .

$$[\text{H}_3\text{O}^+]^2 = (5.60 \times 10^{-10})(5 \times 10^{-3}) = 2.80 \times 10^{-12}$$

$$[\text{H}_3\text{O}^+] = 1.67 \times 10^{-6}, \text{pH} = 5.78$$

Since the pH should be approximately equal to $\text{p}K_a$ of the indicator, a reasonable choice would be either chlorphenol red or methyl red. The change is from base color to acid color, which is from red to yellow for chlorphenol red and yellow to red for methyl red.

Considerations in Acid-Base Titrimetry

Acid-base titrations are routine procedures in any chemical laboratory. They can be used to determine the molar concentrations of acids and bases in aqueous and nonaqueous solutions, although we shall here concentrate on aqueous solutions. When used in this way the precision of their concentration measurements can be a few parts per thousand using the burettes and pipettes which are normally available in modern chemical laboratories.

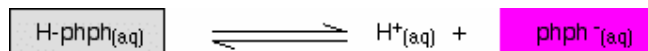
Acid-base titrations are also used to determine the values of acid-base ionization constants. To do this, a chemist takes advantage of the fact that when equal quantities of a weak acid and its conjugate base are present in an aqueous solution, the solution pH must be equal to the pK_a of the weak acid. This point in a titration curve is called the **half-titrated point** of an acid-base titration, and it occurs when the volume of the added titrant is exactly one-half of the volume of titrant required to reach the **equivalence point** of the titration. The value of an acid ionization constant can therefore be determined by simple inspection of an acid-base titration curve.

When an acid-base titration is carried out, a chemist normally tries to adjust the experimental conditions so that the sharp change in pH near the equivalence point is as large and as sharp as possible. The advantage in doing so is a more easily detected and more precisely defined endpoint for the titration. If acid-base indicators are used, a larger pH break at the equivalence point also permits a greater selection of useful indicators. For this reason, **a solution of a strong acid is usually used to titrate a weak base and a solution of a strong base is usually used to titrate a weak acid.** The most common strong acid used is hydrochloric acid and the most common strong base used is sodium hydroxide.

Acid-base titrations require the use of standard solutions, either prepared from primary standards or from secondary standards, as is discussed in [another section](#). The titration can be carried out equally well, in most cases, with the **titrant** solution in the burette being either a standard solution or a solution of an acid or base whose concentration is unknown; the other solution, or a solid sample, is placed in the titration flask.

Phenolphthalein

Phenolphthalein is another commonly used indicator for titrations, and is another weak acid.



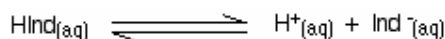
In this case, the weak acid is colourless and its ion is bright pink. Adding extra hydrogen ions shifts the position of equilibrium to the left, and turns the indicator colourless. Adding hydroxide ions removes the hydrogen ions from the equilibrium which tips to the right to replace them - turning the indicator pink.

The half-way stage happens at pH 9.3. Since a mixture of pink and colourless is simply a paler pink, this is difficult to detect with any accuracy!

The pH range of indicators

The importance of pK_{ind}

Think about a general indicator, HInd - where "Ind" is all the rest of the indicator apart from the hydrogen ion which is given away:



Because this is just like any other weak acid, you can write an expression for

$$K_{\text{ind}} = \frac{[\text{H}^+] [\text{Ind}^-]}{[\text{HInd}]}$$

K_a for it. We will call it K_{ind} to stress that we are talking about the indicator.

Think of what happens half-way through the colour change. At this point the concentrations of the acid and its ion are equal. In that case, they will cancel out of the K_{ind} expression.

$$K_{\text{ind}} = \frac{[\text{H}^+] \cancel{[\text{Ind}^-]}}{\cancel{[\text{HInd}]}}$$
$$K_{\text{ind}} = [\text{H}^+]$$

You can use this to work out what the pH is at this half-way point. If you re-arrange the last equation so that the hydrogen ion concentration is on the left -

$$[\text{H}^+] = K_{\text{ind}}$$

$$\text{pH} = \text{p}K_{\text{ind}}$$

hand side, and then convert to pH and $\text{p}K_{\text{ind}}$, you get:

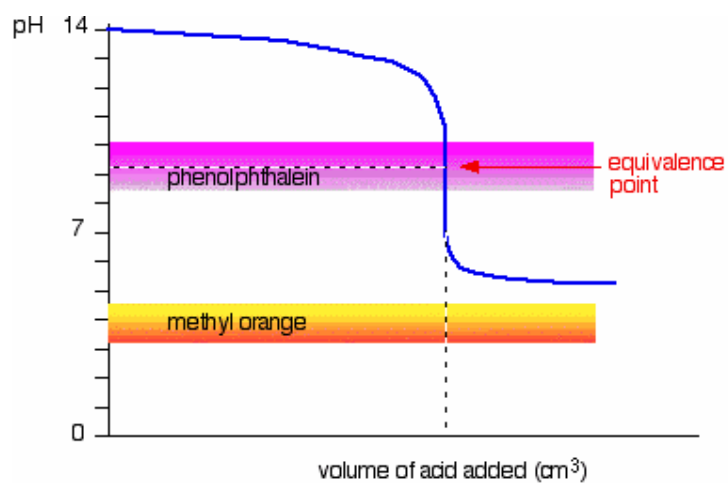
That means that the end point for the indicator depends entirely on what its $\text{p}K_{\text{ind}}$ value is. For the indicators we've looked at above, these are:

indicator	$\text{p}K_{\text{ind}}$
litmus	6.5
methyl orange	3.7
phenolphthalein	9.3

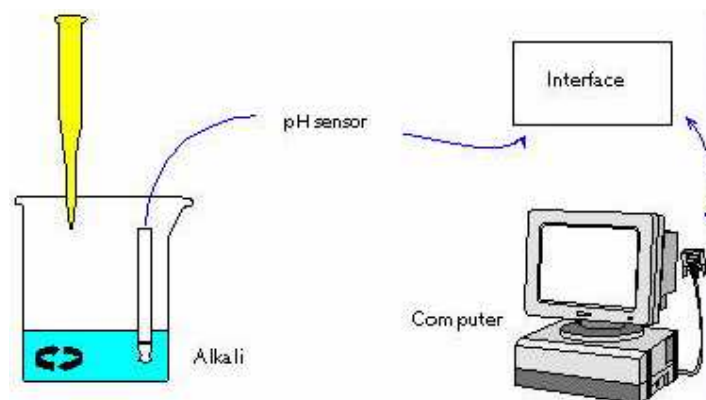
Choosing indicators for titrations

Remember that the equivalence point of a titration is where you have mixed the two substances in exactly equal proportions. You obviously need to choose an indicator which changes colour as close as possible to that equivalence point. That varies from titration to titration.

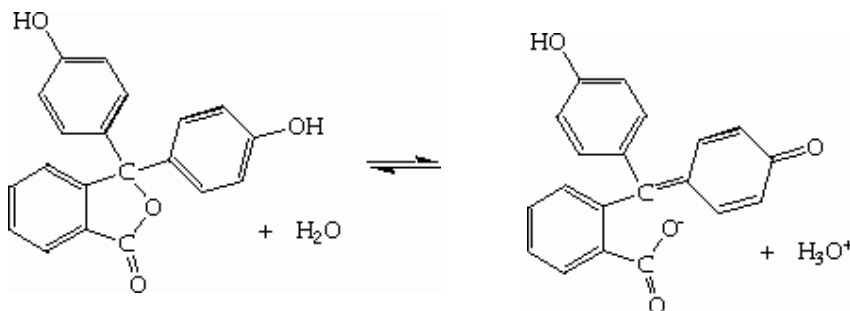
Weak acid v strong base



Evaluation – more accurate readings of pH



Phenolphthalein



Phenolphthalein is a colourless, weak acid which dissociates in water forming pink anions. Under acidic conditions, the equilibrium is to the left, and the concentration of the anions is too low for the pink colour to be observed. However, under alkaline conditions, the equilibrium is to the right, and the concentration of the anion becomes sufficient for the pink colour to be observed.

Acid-Base Titrations

Acid-base titrations are frequently used to determine the concentration of an unknown acid or alkali. The most common method involves using an **indicator** to show when the **end-point** (or **equivalence point**) has been reached. At this point the acid and base have reacted together in the correct stoichiometric ratio.



The above equation tells us that exactly 2 moles of sodium hydroxide must be added for each mole of sulphuric acid present in solution in order to reach equivalence. An end-point can be detected by other means, such as conductivity measurements or by thermometric titration (as neutralisation reactions are exothermic).

Titration Curves

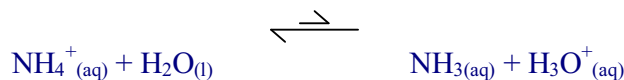
A graph showing the changes in pH during a titration is called a **titration curve**. (*You should have already produced these curves during your practical work.*) The exact shape of the curve depends upon the acid-base combination involved. Titration curves for the four possible combinations are shown below (anticlockwise, starting top left):



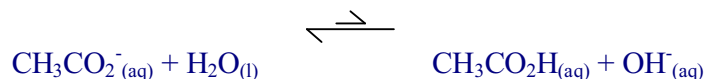
Note the rapid change in pH as the end-point is reached, allowing the easy detection of the end-point by an indicator (see below). At the exact equivalence point, an aqueous solution of sodium chloride is present, so the pH is 7.



Once again there is a rapid change of pH as the end-point is reached, but this time the pH at the equivalence point is less than 7. This is due to the nature of the salt formed - an aqueous solution of ammonium ions is slightly acidic:

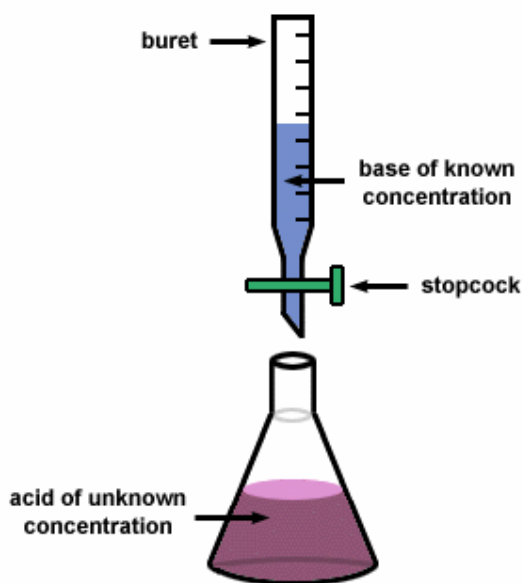


In this case the pH at the equivalence point is above 7. An aqueous solution of ethanoate ions is slightly alkaline:



Here there is a more gradual change in pH as the end-point is reached, and this means there is no satisfactory indicator for determining the exact equivalence point. The pH

at equivalence will be approximately 7, although the exact value will depend upon the acid-base character of the ions in the salt solution formed.

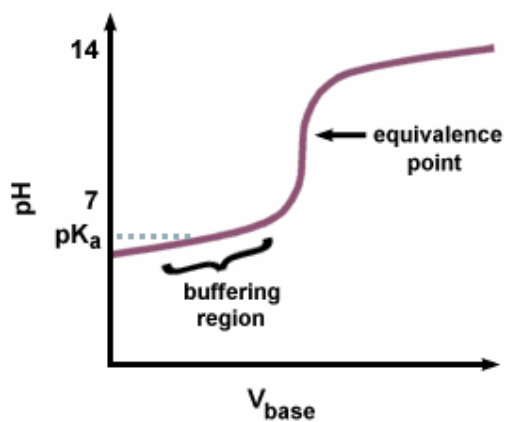


We use this instrumentation to calculate the amount of unknown acid in the receiving flask by measuring the amount of base, or **titrant**, it takes to neutralize the acid. There are two major ways to know when the solution has been neutralized. The first uses a pH meter in the receiving flask adding base slowly until the pH reads exactly 7. The second method uses an **indicator**. An indicator is an acid or base whose **conjugate acid** or **conjugate base** has a color different from that of the original compound. The color changes when the solution contains a 1:1 mixture of the differently colored forms of the indicator. As you know from the [Henderson-Hasselbalch equation](#), the pH equals the pK_a of the indicator at the **endpoint** of the indicator. Since we know the pH of the solution and the volume of titrant added, we can then deduce how much base was needed to neutralize the unknown sample.

1.2 Titration Curves

A **titration curve** is drawn by plotting data attained during a titration, titrant volume on the x -axis and pH on the y -axis. The titration curve serves to profile the unknown solution. In the shape of the curve lies much chemistry and an interesting summary of what we have learned so far about acids and bases.

However, if a strong base is used to titrate a **weak acid**, the pH at the equivalence point will not be 7. There is a lag in reaching the equivalence point, as some of the weak acid is converted to its conjugate base. You should recognize the pair of a weak acid and its conjugate base as a **buffer**. In [Figure 1.3](#), we see the resultant lag that precedes the equivalence point, called the buffering region. In the buffering region, it takes a large amount of NaOH to produce a small change in the pH of the receiving solution.

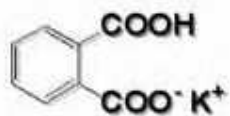


Standard titrants

Standardization of NaOH.

Secondary standard solutions are prepared by titration of a primary standard.

Potassium acid phthalate - KHP



Most commonly used

High equivalent weight - 204.2

High purity

Thermally stable when drying

Reacts rapidly with NaOH / KOH