

2.1 General Principles

The Nernst equation tells us that a measurable quantity - voltage or potential - is related to the concentration of species in solution. In many cases, the measured potential is due to a number of species, and hence the concentration of one can be difficult or impossible to determine. Electrodes have been designed to respond only to one (or a very small number of) species, thus allowing the measurement of its concentration. This is known as **direct potentiometry**, and will be discussed in the next chapter.

In other cases, the analyte can be determined by means of a titration, and **the change in its concentration monitored by measurement of the solution potential**. This is known as a **potentiometric titration**.

In titration measurements, the errors (activity differences and junction potentials) mentioned in the previous chapter are not a problem because we are not interested in relating a single voltage measurement to the solution concentration: we are only interested in identifying the endpoint volume of the titration by the change in voltage from one volume addition to the next. Figure 2.1 shows the typical shape of a potentiometric titration curve.

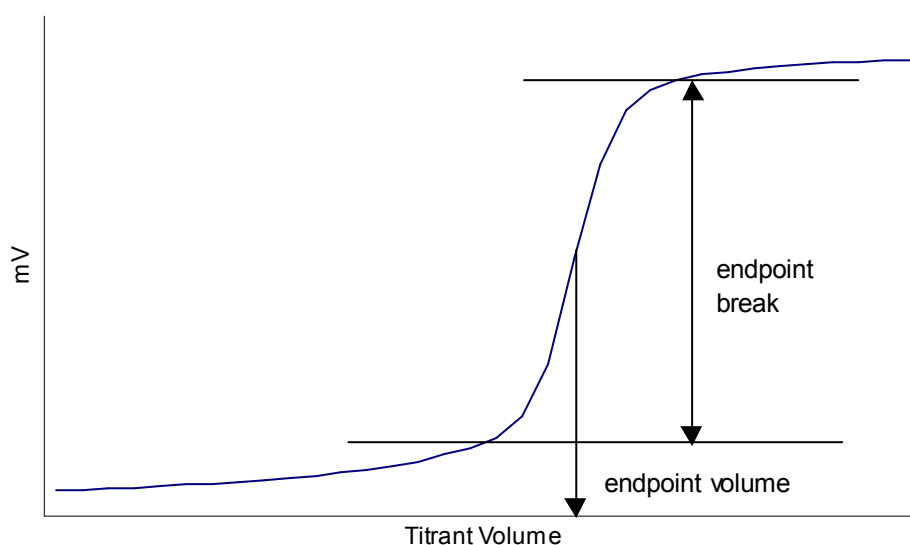


FIGURE 2.1 Typical potentiometric titration curve

Features of the titration curve

- the wave-like shape occurs because of the rapid change in voltage around the endpoint of the reaction
- the endpoint break is the large change in voltage around the endpoint
- the endpoint break should be as large as possible to improve accuracy of detection: this is done by choosing the titrant carefully (discussed in later sections of this chapter)
- the endpoint volume is defined as the volume half-way up the endpoint break
- the voltage values before the endpoint are due to the analyte
- the voltage values after the endpoint are due to the titrant

2.2 Why Choose a Potentiometric Titration?

Compared to titrations using an indicator to detect the endpoint, potentiometric titrations are:

- *more expensive* – you need a stirrer and bar, electrodes, and a voltmeter
- *slower to set up* – making sure the electrodes are correctly in place and functioning properly, getting the burette above the beaker and out of the way of the electrodes, making sure the stirrer bar isn't going to smash the end off the electrode
- *slower to perform* – having to record all the data (in most cases)
- *slower to get the endpoint volume* – you just don't read the volume off the burette at the end (in most cases), you have to process the data in some way

So given all this, why would you choose to use this method:

- *background colour* – the colour of the sample (eg red wine) is too great to be able to see an indicator change colour
- *correct indicator not known* – if the sample is of unknown composition, then the correct indicator to accurately determine the endpoint cannot be identified; however, after the potentiometric titration is done, the correct indicator can be chosen for further analyses of that sample type (discussed later)
- *solution is too dilute* – indicator titrations are not particularly sensitive, ie they can't accurately detect endpoints for analytes below about 0.01 M; potentiometric methods can do better than that

2.3 Data Collection and Processing

Collection

In an indicator-based titration, we titrate until we see the colour change. In a potentiometric titration, the endpoint is determined by data derived from the titration, and we generally continue to add titrant to at least 10 mL after the equivalence point. How do we know where the endpoint is? Simply, the potential readings will begin to change quite rapidly around the equivalence point as you have seen in Figure 2.1.

When performing a potentiometric titration, **it is crucial that as many data points as possible are obtained close to the endpoint**. Typically, the operator should aim at 0.1 mL intervals in the 1 mL around endpoint. The basic approach is to add the titrant in volumes that do not cause large changes in the potential (or pH). 5 or 10 mL additions to start are satisfactory, then the additions should be reduced to 1, 0.5 and 0.1 mL as endpoint approaches. After endpoint, the process is reversed until at least one 5 mL addition is made.

Endpoint detection

The very steep part of the graph is the region in which the equivalence point is found. The point at which the slope becomes greatest is defined as the equivalence point. Experimentally, this means the volume half way up the endpoint break.

To determine this **graphically**, plot the graph manually on graph paper or by computer, using gridlines to simulate graph paper. Extend lines from the two flat parts of the graph and then determine the half-way distance between them. Draw a line across to where it intersects with the graph to find the endpoint volume.

Problems with the graphical method include:

- buffer regions not being flat or parallel, making the measurement of the half-way point difficult
- manual plotting of graphs is very subjective and therefore inconsistent
- plotting of graphs using a spreadsheet and printing on plain paper (even with gridlines) requires knowledge of the program and is less accurate than on real graph paper

Therefore, a mathematical method, known as the **first derivative method** has been developed to avoid the need to use a graph. It is quicker, totally consistent and relatively simple. The first derivative method involves calculation of a sequence of values based on successive data points. The formula is shown in Equation 2.1

$$\text{1st deriv. value} = \left| \frac{mV_2 - mV_1}{V_2 - V_1} \right| \quad \text{Eqn 2.1}$$

where mV and V are the potential and volume data for successive data points, and the vertical lines mean that all values must be positive.

The **endpoint by the first derivative method is indicated by the maximum 1st derivative value**, and the **endpoint volume is the midpoint between the two volumes** that give the maximum value.

EXAMPLE 2.1

Determine the endpoint from the following data by the first derivative method.

Volume (mL)	mV	First deriv.
19.0	235	
$\frac{19.5 - 19.0}{2} = (19.25)$		$\frac{240 - 235}{19.5 - 19.0} = 10$
19.5	240	
(19.75)		12
20.0	246	
(20.25)		24
20.5	258	
(20.55)		40
20.6	262	
(20.65)		170
20.7	279	
(20.75)		190
20.8	298	
(20.85)		490
20.9	347	
(20.95)		1470
21.0	494	
(21.05)		2910
21.1	785	
(21.15)		1290
21.2	914	
(21.25)		450
21.3	959	
(21.35)		130
21.4	972	
(21.45)		90
21.5	983	
(21.75)		26
22.0	996	
(22.25)		18
22.5	1005	
(22.75)		12
23.0	1011	

The endpoint from the first derivative method is the midpoint between 21.0 and 21.1: **21.05 mL**.

CLASS EXERCISE 2.1

Determine the endpoint from the following potentiometric data.

Vol.	mV	1 st deriv. value	Vol.	mV	1 st deriv. value
0	56		13.6	480	
2	64		13.8	682	
4	73		14.2	751	
6	80		14.5	804	
8	90		15	842	
10	102		15.5	869	
11	112		16	893	
12	128		17	906	
12.5	139		18	911	
13	169		20	922	
13.2	188		22	929	
13.4	237		24	936	

Endpoint volume = _____

Therefore, the first derivative value is entirely satisfactory, and less work. In practice, you should look at the data to see where the large changes are occurring, and only calculate the first derivative values for a few points close to the endpoint. In the above example, this region is indicated by the double lines.

2.4 Apparatus

The basic apparatus required for all potentiometric titrations are the same and is shown in Figure 2.2:

- *indicator electrode* – responds to changes in the species of interest
- *reference electrode* – produces an unchanging voltage regardless of solution changes; most commonly a calomel electrode
- *voltmeter/pH meter* – measures the voltage/pH of the solution
- *stirrer and follower* – to avoid the need to manually swirl the solution between additions
- *burette* – to add the titrant of course

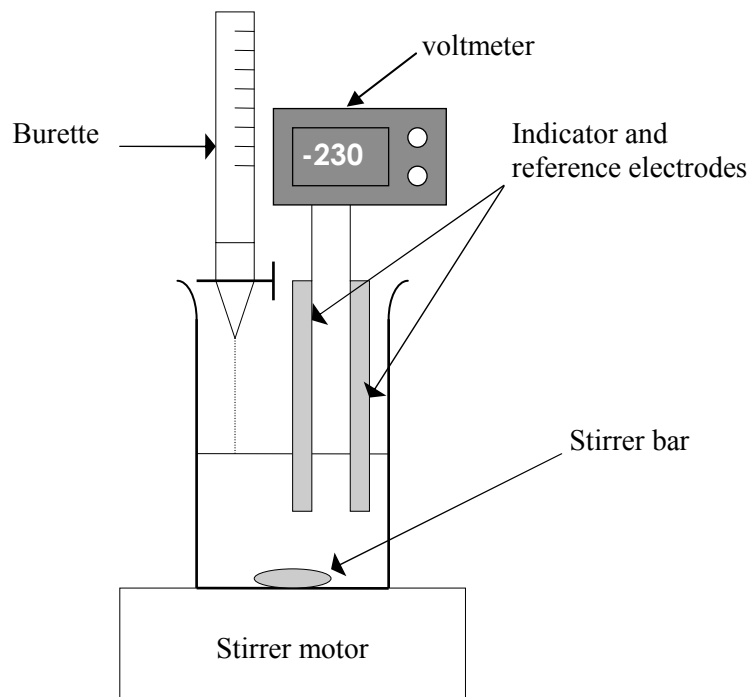


FIGURE 2.2 Typical apparatus for potentiometric titrations

Different titration reactions do require specific differences in not just the titrant in the burette, but also the two electrodes. Table 2.1 lists the standard electrode combinations for the normal titration types.

TABLE 2.1 Electrodes used in potentiometric titrations

<i>Reaction type</i>	<i>Indicator electrode</i>	<i>Reference electrode</i>	<i>Salt bridge needed?</i>
Acid-base	pH	contained in pH	No
Redox	Pt	calomel	No
Precipitation (with Ag^+)	Ag	calomel	Yes

CLASS EXERCISE 2.2

Select suitable electrodes for the following analyses.

Analysis	Electrode(s)
acid content of red wine	
salt content in seawater	
hydrogen peroxide in bleach	

2.5 Reaction Types

Titration reactions are of four basic classes:

- acid-base
- redox
- precipitation
- complexometric

Of these, only the last – complexometric – is unsuited to potentiometric measurements. This is because there are no suitable electrodes to monitor solution changes.

In any titration, a standardised titrant is used to react with the analyte in the sample. To make measurement of the endpoint as accurate as possible, **the size of the endpoint break should be as large as possible**. This rule applies to indicator-based titrations as well as potentiometric titrations. As mentioned earlier, the voltage values after the equivalence point are due to the titrant, so choice of the titrant can affect the size of the endpoint break, as shown in Figure 2.3: titrant B is the preferred choice.

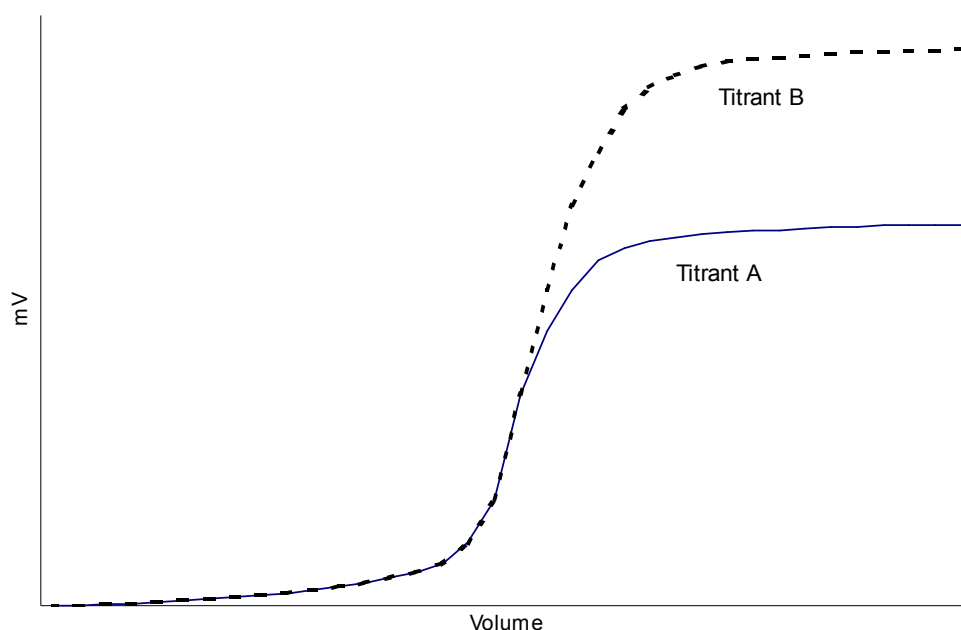


FIGURE 2.3 Effect of titrant on endpoint break

The following sections which examine the specific reaction types will focus on what chemical aspect of the titrant affects the size of the endpoint break.

Acid-base reactions

You should be familiar with the meaning and measurement of pH from other subjects. Figure 2.4 shows the pH curve for the titration of 0.1 M NaOH with 0.1 M HCl.

Because you are familiar with pH, we can use this titration to explain why titration curves are the shape they are. The calculations below in Table 2.2 that show how the pH changes in a titration are not examinable.

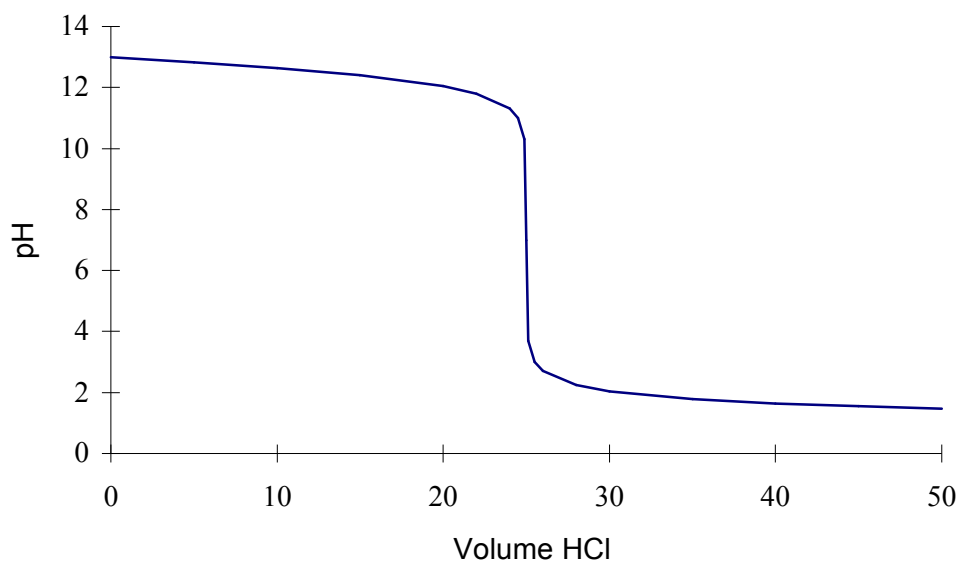


FIGURE 2.4 *pH* curve for 0.1 M NaOH/0.1 M HCl titration

TABLE 2.2 *pH* changes in the titration of 25 mL of 0.1 M NaOH with 0.1 M HCl

<i>Vol. HCl (mL)</i>	<i>mmoles HCl added</i>	<i>mmoles NaOH left</i>	<i>pH</i>
0	0	2.5	13.00
5	0.5	2	12.82
10	1	1.5	12.63
20	2	0.5	12.05
22	2.2	0.3	11.81
24	2.4	0.1	11.31
24.5	2.45	0.05	11.00
24.6	2.46	0.04	10.91
24.7	2.47	0.03	10.78
24.8	2.48	0.02	10.60
24.9	2.49	0.01	10.30
25	2.5	0	7.00
		<i>mmoles HCl in excess</i>	
25.1	2.51	0.01	3.70
25.2	2.52	0.02	3.40
25.3	2.53	0.03	3.22
25.4	2.54	0.04	3.10
25.5	2.55	0.05	3.00
26	2.6	0.1	2.71
30	3	0.5	2.04
35	3.5	1	1.78
40	4	1.5	1.64

As you can see the pH does not change consistently with volume. This is because pH is a logarithmic scale, so a **decrease of one pH unit requires a 10 x increase in $[H^+]$ (or in this case a 10 x decrease in $[OH^-]$** . When there is a lot of OH^- , like at the start of the titration, you need to add a relatively large volume of HCl to remove 90% (reduce by a factor of 10) of the NaOH. When there is a small amount of OH^- , like near the endpoint, only a small volume of HCl is required to reduce its level by 90%.

There are various factors that affect the shape of an acid-base titration curve, as discussed in the following sections. These allow us to make judgements and choose the best titrant for a particular analysis, and to select indicators where a potentiometric analysis is not necessary.

Before and after endpoint

The two principal factors that affect the shape of the titration curve before and after endpoint are:

- the strength of the acid and base – their level of dissociation (see Figure 2.5)
- the concentration of the acid and base – the concentration of analyte essentially dictates the concentration of the titrant (see Figure 2.6)

CLASS EXERCISE 2.3

Why is it necessary in all titrations that the concentration of the titrant should be similar to the concentration of the analyte?

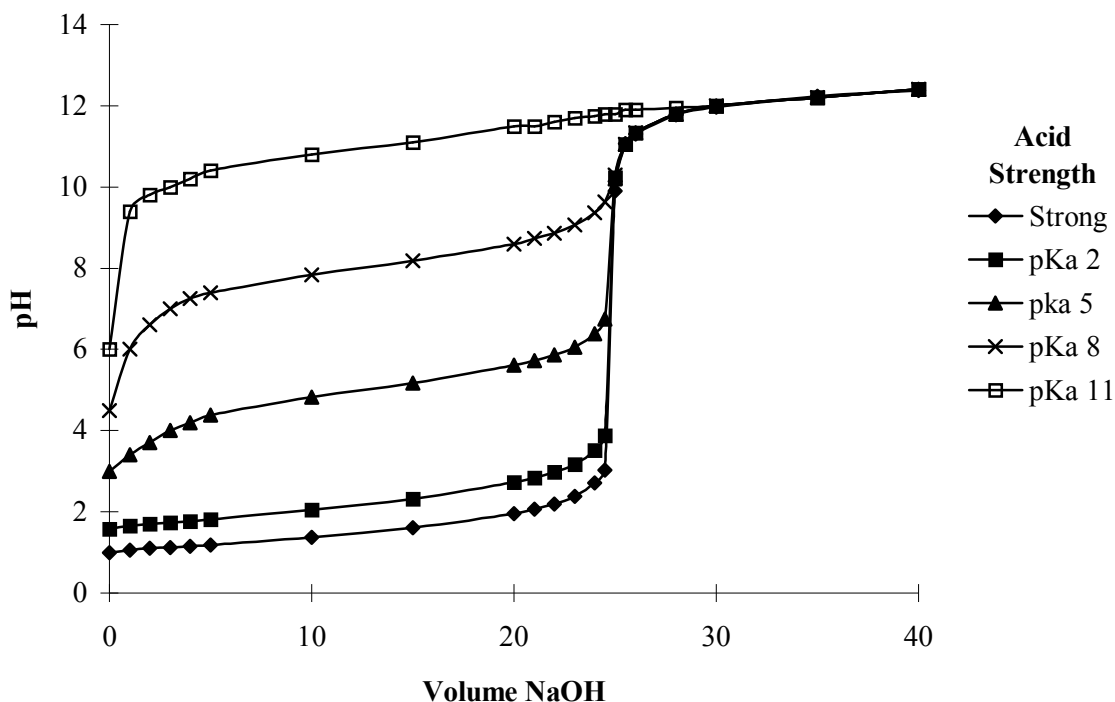


FIGURE 2.5 The effect of acidic strength on pH break

CLASS EXERCISE 2.4

What trend is apparent in Figure 2.5?

Given that the same trend is observed for bases, what titrant would be best for the analysis of the analysis for the weak base ammonia?

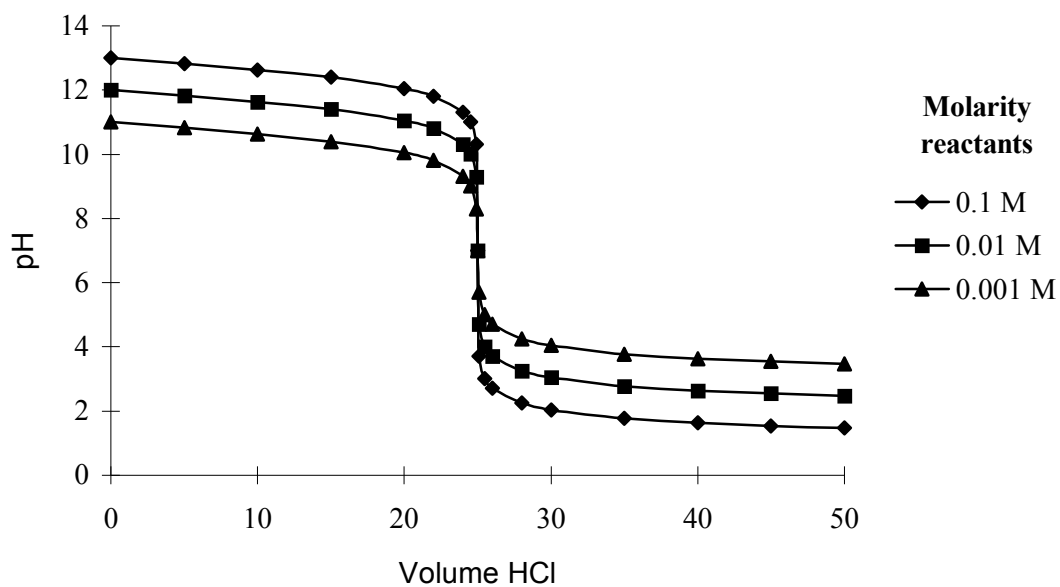


FIGURE 2.6 The effect of concentration on pH break for titrations of NaOH with HCl

CLASS EXERCISE 2.5

What trend is apparent in Figure 2.6?

GENERAL RULES

1. Strong acids and strong bases are always chosen as the titrants in potentiometric and indicator-based titrations because they maximise the endpoint break.
2. Lower concentrations decrease the endpoint break, making the detection more difficult, and less likely that an indicator will be accurate.

At endpoint

The pH of the equivalence point of an acid-base titration is not always 7. While we often describe these reactions as neutralisations, the endpoint is often not neutral pH 7. Only for a titration involving a strong acid and a strong base does the equivalence point occur at pH 7. Once weak acids or bases are involved, the equivalence points moves. Table 2.3 revises the general rules regarding the pH of equivalence points.

TABLE 2.3 *pH of equivalence points for different acid/base combinations*

<i>Acid strength</i>	<i>Base strength</i>	<i>pH at equivalence point</i>
Strong	Strong	Exactly 7
Strong	Weak	Less than 7 (5-7)
Weak	Strong	Greater than 7 (7-9)
Weak	Weak	Unpredictable

CLASS EXERCISE 2.6

Predict the endpoint pH (relative to 7) for the following titrations.

<i>Acid</i>	<i>Base</i>	<i>pH</i>
Hydrochloric	Sodium hydroxide	
Sulfuric	Ammonia	
Ethanoic	Potassium Hydroxide	
Ethanoic	Ammonia	

EXAMPLE 2.1

Draw a representation of the titration curve for 25 mL of 0.1 M ethanoic acid with 0.1 M NaOH.

Start pH: Ethanoic acid is a weak acid. While we can't calculate the exact pH of a 0.1 M solution of it, we know it must be greater than 1, since 0.1 M strong acid has a pH of 1.

End pH: If we titrate far enough past endpoint, we will end up with approximately 0.1 M NaOH, which has a pH of 13.

Endpoint pH: A weak acid and a strong base gives an endpoint pH greater than 7.

Endpoint volume: the reaction is 1:1, so 25 mL of NaOH will be needed.

An approximation of the titration curve is shown in Figure 2.7.

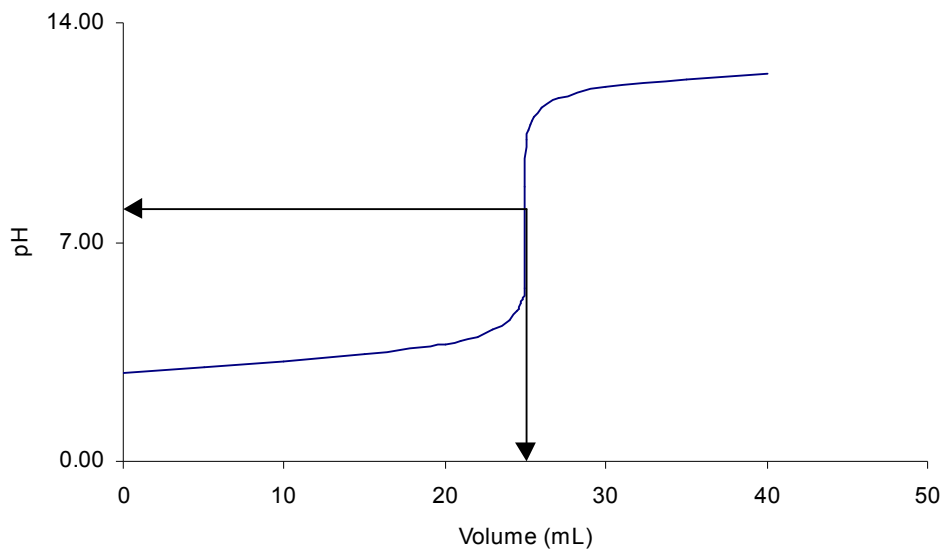


FIGURE 2.7 Approximate titration curve for Example 2.1

CLASS EXERCISE 2.7

Draw approximate titration curves for:

(a) the titration of 25 mL of 0.1 M ammonia with 0.1 M HCl



(b) the titration of 50 mL of 0.01 M HCl with 0.01 M NaOH



Multi-endpoint samples

Some samples will give multiple endpoints during a potentiometric titration, whether acid-base or otherwise. In acid-base titrations, there are two basic reasons for multiple endpoints:

- the compound has more than one reactive site
- the sample is a mixture of strong and weak acids

Compounds with more than one reactive site

This only applies to weak acids or bases, where the compound has or reacts with more than one hydrogen ion, and the two reactions have very different $K_{a/b}$ values. For example the weak base, carbonate, can react with two hydrogen ions, and phosphoric acid has three protons of different strengths (see Figure 2.8 – the reason that a third endpoint is not shown is because the third acid is so weak, it doesn't show an endpoint – remember Figure 2.5).

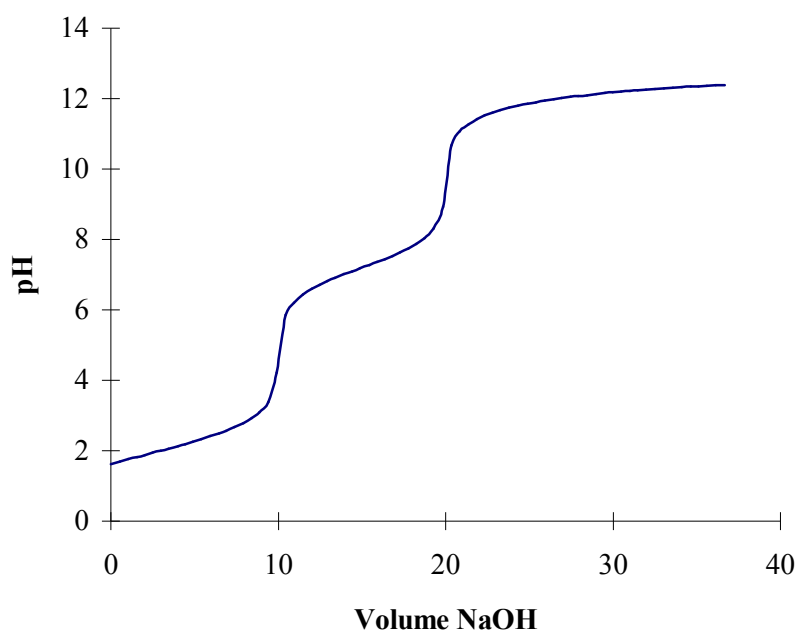


FIGURE 2.8 pH curve for phosphoric acid/NaOH titration

Each endpoint corresponds to the reaction of one proton from all molecules of the compound in the solution, and therefore, the volume of added titrant to the first endpoint will be exactly the same as the volume between the first and second endpoints. In general, **the first endpoint break will be smaller than the second, so the more accurate answer is obtained by using the latter**, and remembering the reaction ratio is not 1:1.

EXAMPLE 2.2

A 25.0 mL aliquot of carbonate solution is potentiometrically titrated with 0.0104 M HCl. The endpoints are determined to be 12.3 and 25.7 mL. Calculate the molarity of the carbonate.

Given the first endpoint would have a shorter pH break than the second, it is generally more accurate to use the second value, where the ratio is $2\text{HCl} + 1\text{CO}_3^{2-}$.

$$\text{Moles HCl} = (25.7/1000) \times 0.0104 = 2.673 \times 10^{-4}$$

$$\text{Moles CO}_3 = \frac{1}{2} \text{moles HCl} = 1.336 \times 10^{-4}$$

$$\text{Molarity CO}_3 = 1.336 \times 10^{-4} / 0.025 = 0.00535\text{M}$$

CLASS EXERCISE 2.8

A 50.0 mL aliquot of phosphoric acid is titrated with 0.121 M NaOH, and the process monitored using a pH meter. The data was analysed, and the volume required to reach the first endpoint was 23.1 mL. The TOTAL volume required to reach the second endpoint was 46.0 mL. Calculate the molarity of the phosphoric acid.

It should be made clear that di- or triprotic acids or bases will give multiple endpoints only if the pKa values are sufficiently different. In the case of citric acid, a triprotic acid, with pKa values of 3.1, 4.8 and 6.4, the dissociation strengths are close enough to mean that only one endpoint is visible.

Mixtures Of Strong And Weak Acids/Bases

A strong acid dissociates completely, and when mixed with a weak acid, stops the latter from dissociating at all. When NaOH is added, it will react with the free H⁺ only, meaning that **the strong acid reacts first**. Only when almost all of it has reacted, does the weak acid begin to dissociate and react. This means that two endpoints can be seen, but the first is often small and broad (see Figure 2.9). The concentration of each compound can be determined using the volume used for each endpoint, though the accuracy must be limited by the poorly defined first endpoint.

FIGURE 2.7 pH curve for titration of a mixture of a strong acid and a weak acid with NaOH

EXAMPLE 2.3

A mixture of nitric acid and benzoic acid in a 50.0 mL sample were titrated with 0.114 M NaOH, using a pH electrode to allow the plotting of a pH curve. The first endpoint occurred at a volume of 19.2 mL, and the second at a volume of 31.8 mL. Calculate the molarities of the two acids.

Nitric acid is a strong acid, while benzoic acid is a weak acid. Therefore, the volume of NaOH required to titrate the nitric acid is 19.2 mL, and the volume required to titrate the benzoic acid is 31.8 - 19.2 = 12.6 mL.

$$\text{Molarity nitric acid} = (0.0192 \times 0.114) \div 0.050 \text{ L} = 0.0438 \text{ M}$$

$$\text{Molarity benzoic acid} = (0.0126 \times 0.114) \div 0.050 \text{ L} = 0.0287 \text{ M}$$

CLASS EXERCISE 2.9

25.0 mL of each of an ethanoic acid solution and a sulfuric acid solution were pipetted into a beaker, and titrated with a 0.0953 M NaOH solution. A pH curve was plotted, and the volume required to reach the first endpoint was 35.2 mL, and a FURTHER 16.1 mL required to reach the second endpoint. Calculate the molarities of the two acids.

It should be obvious that the same comments can be made regarding mixtures of strong and weak bases, e.g. hydroxide and carbonate.

Total acidity/alkalinity

What is you don't know the composition of the mixture, so that any early endpoints in the titration could be for either reason. In that case, only a total number of moles of acid (or base) can be determined. In most samples of known origin, there will be a dominant species – for example citric acid is the major acid in wine – so that the moles of acid/base are reported as though the only species is that dominant one. This is known as the **total acidity** (or alkalinity).

EXAMPLE 2.4

A 10 mL aliquot of lemon juice is titrated potentiometrically with 0.102 M NaOH, and the final endpoint is 16.7 mL. What is the total acidity of the sample (as citric acid)?

$$\text{Moles of NaOH} = 0.0167 \times 0.102 = 1.703 \times 10^{-3} \text{ moles}$$

$$\text{Moles of citric acid} = \text{moles of NaOH} \div 3 = 5.678 \times 10^{-4} \text{ moles}$$

$$\text{Mass citric acid} = 5.678 \times 10^{-4} \times 192 = 0.109 \text{ g}$$

$$\text{Total acidity (g/L)} = 0.109 \text{ g} \div 0.01 \text{ L} = 10.9 \text{ g/L}$$

CLASS EXERCISE 2.10

A 25.0 mL aliquot of red wine is titrated to endpoint with 0.0752 M NaOH. Report the total acidity of the sample as gram of tartaric acid/100 mL, given the FW of tartaric acid is 150 and the reaction ratio is 1 acid: 2 NaOH.

Using potentiometric data to find endpoint pH

Because a pH electrode can be calibrated accurately (unlike the electrodes used for other titrations), the actual pH values can have some meaning. Because the endpoint pH for a particular combination of acid and base will be constant, we can use this to set the final pH for the titration, and not have to record all the data before, only the volume on the burette when the pH reaches that value.

For example, the endpoint pH for the titration of wine with NaOH is accepted generally to be 8.4 (small variations for different wines are not important). Therefore, analysis of wine is made easier by this approach: simply calibrate the pH electrode, titrate the sample to 8.4 and write down the burette volume.

What about a new sample whose endpoint pH is unknown? Simply do a full potentiometric titration with all the data on a few samples and determine the endpoint pH from the first-derivative method as **the average pH of the two data points which surround the endpoint**. This becomes the endpoint pH now for that sample type as 8.4 is for wine.

CLASS EXERCISE 2.11

Determine the endpoint pH for an alkaline waste sample given the titration data below.

Vol.	pH	First deriv.
21.2	7.25	
21.3	6.88	
21.4	6.07	
21.5	5.01	
21.6	4.23	
21.7	3.92	

For some sample where the usual reasons for using a potentiometric titration are not a problem (colour, concentration), a potentiometric titration might still be necessary because the sample is of unknown composition, so the shape of the titration curve is not known. In this case, while an indicator can be used, you don't know which one.

Acid-base indicators, such as phenolphthalein, change colour over a specific pH range, usually 1-2 pH units. This is due to their individual chemical structure. For example, phenolphthalein changes colour between 8.2 and 10. The rule for choosing an indicator which will accurately determine the endpoint volume is that **its colour change must fall completely within the endpoint break for that titration**. Table 2.4 lists the colour change ranges for some common indicators.

Method for determining a suitable indicator

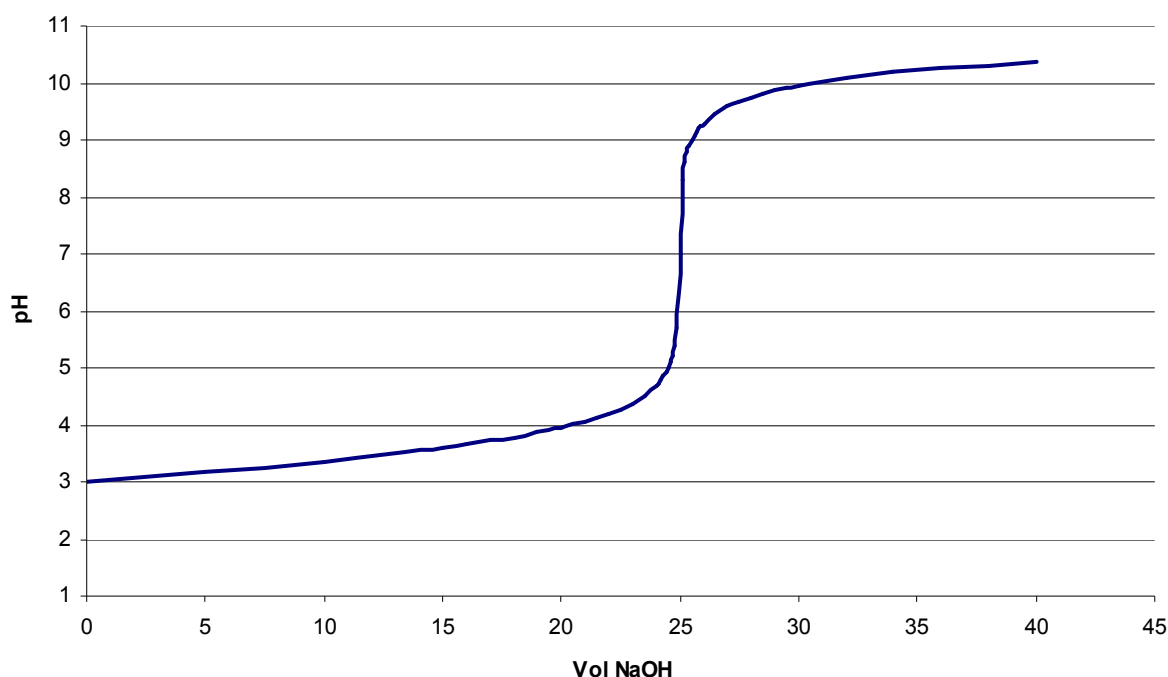
1. Determine whether the sample is acidic or basic (use pH paper)
2. Carry out a full potentiometric titration
3. Plot the data on a graph
4. Find an indicator whose colour change range is completely within the endpoint break

TABLE 2.4 Indicator colour changes

<i>Indicator</i>	<i>Colour Change pH</i>
Methyl orange	3.2 – 4.4
Bromocresol green	3.8 – 5.4
Methyl red	4.8 – 6.0
Bromothymol blue	6.0 – 7.6
Phenol red	6.6 – 8.0
Phenolphthalein	8.2 – 10.0

CLASS EXERCISE 2.12

Determine a suitable indicator from Table 2.4 for the titration graphed below.



CLASS EXERCISE 2.13

Why is phenolphthalein chosen for most titrations involving NaOH (except where the concentrations are $< 0.01\text{ M}$ – see Exercise 2.12) when bromothymol blue and phenol red have colour change ranges that include the endpoint pH of 7, and phenolphthalein doesn't?

Redox and precipitation reactions

The same basic principles discussed thoroughly for acid-base reactions also apply here. The factors that affect the shape of the titration curve are:

- *chemical nature of reactants* – here strong/weak is not appropriate, but the size of the endpoint break is still affected by particular chemical properties; for redox reactions, the greater the standard reduction/oxidation potential, the larger is the endpoint break, and for precipitation titrations, the smaller the value of the solubility product, K_{sp} ,
- *concentration* – the lower the concentrations, the smaller the endpoint breaks
- *multiple endpoints* – if the sample composition is known, then the individual species can be analysed separately, otherwise a total value must be given; no single species gives multiple endpoints

CLASS EXERCISE 2.14

Given the following data, choose the best titrant for each analysis.

Oxidant	E° (V)	Halide	K_{sp} (for Ag salt)
Permanganate	1.51	Chloride	1.6×10^{-10}
Dichromate	1.33	Bromide	7.7×10^{-13}
Iodine	0.54	Iodide	1.5×10^{-16}

(a) an oxidising titrant to analyse coloured iron (II) solutions

(b) a titrant to analyse silver solutions

In potentiometric titrations involving these types of reactions, millivolts are measured to detect the endpoint. Unlike the pH electrode, the Pt/SCE and Ag/SCE electrode combinations can't be accurately calibrated, so we can't preset millivolt values to titrate to, as we did with pH titrations (eg wine at 8.4). Nor is it as easy to select indicators – redox indicators such as ferroin change colour at a particular mV range – for the same reason. So full data-collection titrations have to be done every time unfortunately.

2.6 Automatic Titrators

A potentiometric titration, performed to collect all data, is tedious and time-consuming. When a set pH can be assigned as the endpoint, the process speeds up substantially. The nature of the potentiometric apparatus lends itself readily to automation. Automatic titrators are available which can perform either form of a potentiometric titration: collection of data and calculation of endpoint, and titration to a set pH.

The equipment is relatively simple, the key piece being the control of the titrant addition. In the simplest case where titration is performed to a set pH, a flexible piece of tubing is squeezed shut until the command is given to begin adding. The titrant is added at a fast rate until the pH reaches a predetermined value before endpoint. The control mechanism then closes the delivery tube to slow the addition of titrant. When the endpoint value is reached, the delivery is stopped. In a fully automated system, there will be some method of sensing how much titrant was added. Otherwise, the analyst can measure it.

Recording autotitrators deliver the titrant at a constant rate, and monitor the solution potential continually, producing data on a chart recorder or digitally into a computer. Other autotitrators are programmed to detect the endpoint by first- or second-derivative methods, meaning that there is no need for preset pH (or mV) values, an obvious advantage where different samples, and unknown samples, are being analysed.

Each type lends itself to multi-sample unsupervised analysis, which is a growing trend in instrumental design: the autosampler, which is a rotating turntable with many spaces for different samples, which analysed successively by the instrument.

What You Need To Be Able To Do

- distinguish between direct potentiometry and potentiometric titrations
- determine the endpoint of a titration from potentiometric data by the first derivative method
- draw a diagram showing the essential apparatus for a potentiometric titration
- list reasons for using potentiometric measurements instead of an indicator in a titration
- draw pH curves for acid-base titrations
- describe the effect on the pH curve of acid/base strength and concentration
- describe procedures where multiple endpoints are obtained in acid-base titrations
- determine the endpoint pH and use this to determine a suitable indicator
- outline variations in redox and precipitation titration curves
- explain how autotitrators work