Reactivity 3.1

IB CHEMISTRY SL

Reactivity 3.1.1
Understandings:

- Brønsted–Lowry acid is a proton donor and a Brønsted–Lowry base is a proton acceptor.

Learning outcomes:

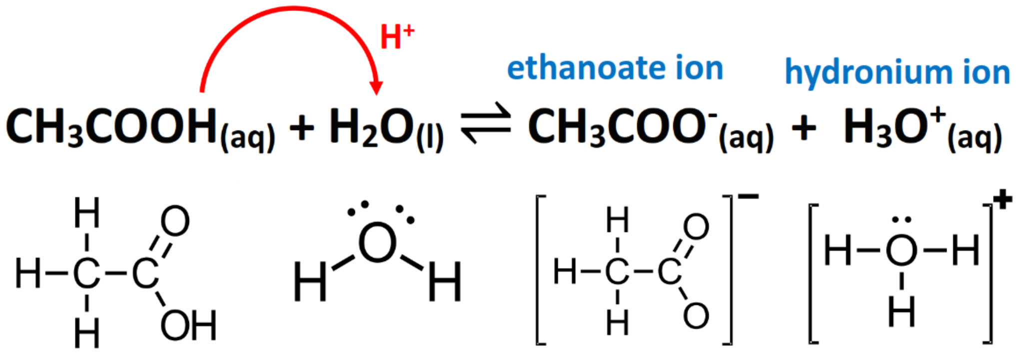
- Deduce the Brønsted–Lowry acid and base in a reaction.

Additional notes:

- A proton in aqueous solution can be represented as both $H^+(aq)$ and $H_3O^+(aq)$.
- The distinction between the terms "base" and "alkali" should be understood.

Brønsted–Lowry theory of acid and bases

- A Brønsted–Lowry acid is a proton (H^+) donor.
- A Brønsted–Lowry base is a proton (H^+) acceptor.
- A proton is $H^+(aq)$ but exists in solution as $H_3O^+(aq)$.
- The reaction below shows ethanoic acid reacting with water to form the ethanoate ion and the hydronium ion.

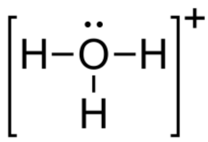


- Ethanoic acid is behaving as a Brønsted-Lowry acid (a proton donor) and water is behaving as a Brønsted-Lowry base (a proton acceptor).
- The reaction is shown in equation form below; note that both equations show the same reaction.

$$CH_3COOH(aq) + H_2O(l) \rightleftharpoons CH_3COO^-(aq) + H_3O^+(aq)$$

$$CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$$

- The H^+ ion exists in solution as $H_3O^+(aq)$; its Lewis structure is shown below.



$H_3O^+_{(aq)}$ / $H^+_{(aq)}$

Alkalis and bases

- Bases such as ZnO and MgO are insoluble in water.
- An alkali is a base that is soluble in water.
- Examples of alkalis include NaOH and LiOH.
- All alkalis are bases but not all bases are alkalis.

Exercises:

1) Define an acid and base according to the Brønsted–Lowry theory.

2) Outline what is meant by a proton in the Brønsted–Lowry theory of acids and bases.

3) Identify the Brønsted–Lowry acid and base in the forward reaction and in the reverse reaction in the reaction below.

$$NH_{3(aq)} + H_2O_{(l)} \rightleftharpoons \underset{\text{ammonium ion}}{NH_4^+{}_{(aq)}} + \underset{\text{hydroxide ion}}{OH^-{}_{(aq)}}$$

4) Outline the difference between a base and an alkali.

Reactivity 3.1.2

Understandings:

- A pair of species differing by a single proton is called a conjugate acid–base pair.

Learning outcomes:

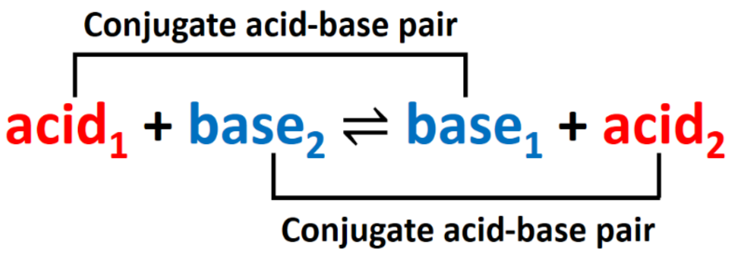
- Deduce the Brønsted–Lowry acid and base in a reaction.

Linking question(s):

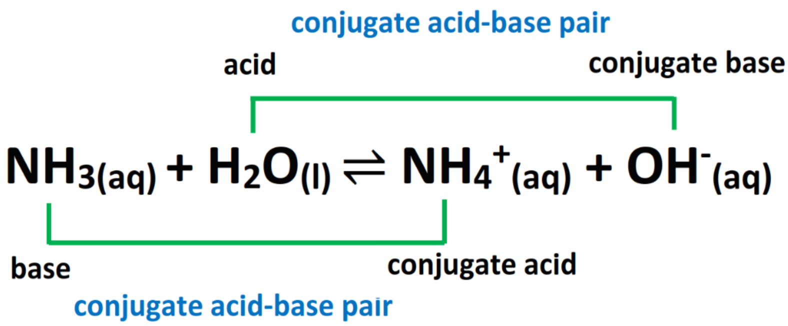
- Structure 2.1 What are the conjugate acids of the polyatomic anions listed in Structure 2.1?

Conjugate acid-base pairs

- Conjugate acid-base pairs differ by a single proton (H^+).
- A conjugate acid is formed when a base accepts a proton.
- A conjugate base is formed when an acid donates a proton.



- In the reaction below, H_2O and OH^- are conjugate acid-base pairs, as are NH_4^+ and NH_3.



Example: Identify the conjugate acid-base pairs in the following reaction:

$$CH_3COOH(aq) + H_2O(l) \rightleftharpoons CH_3COO^-(aq) + H_3O^+(aq)$$

Exercise: Identify the conjugate acid-base pairs in the following reactions.

1) $HClO_4(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + ClO_4^-(aq)$

2) $H_2SO_3(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + HSO_3^-(aq)$

3) $H_2S(g) + H_2O(l) \rightleftharpoons H_3O^+(aq) + HS^-(aq)$

Reactivity 3.1.3

Understandings:

- Some species can act as both Brønsted–Lowry acids and bases.

Learning outcomes:

- Interpret and formulate equations to show acid–base reactions of these species.

Linking question(s):

- Structure 3.1 What is the periodic trend in the acid–base properties of metal and non-metal oxides?
- Structure 3.1 Why does the release of oxides of nitrogen and sulfur into the atmosphere cause acid rain?

Amphiprotic species

- An amphiprotic species is a substance that can act as a Brønsted-Lowry acid or Brønsted-Lowry base.
- An amphiprotic species must be able to donate a proton (H^+) to another species and be able to accept a proton (H^+) from another species.
- In the examples below, H_2O is acting as both a Brønsted-Lowry acid and a Brønsted-Lowry base; it is amphiprotic.

$$CH_3COOH_{(aq)} + H_2O_{(l)} \rightleftharpoons CH_3COO^-_{(aq)} + H_3O^+_{(aq)}$$

$$NH_{3(aq)} + H_2O_{(l)} \rightleftharpoons NH_4^+{}_{(aq)} + OH^-_{(aq)}$$

Amphiprotic and amphoteric

- Amphiprotic refers to any substance that can both accept and donate a proton (Brønsted-Lowry theory).
- Amphoteric refers to any species that can act like an acid or a base. All amphiprotic species are also amphoteric.
- The term amphoteric can be applied in different theories of acids and bases (Lewis theory).

Exercises:

1) Using H_2O as an example, explain what is meant by an amphiprotic species.

2) Write equations for the following amphiprotic species reacting with the hydronium ion (H_3O^+) and the hydroxide ion (OH^-).

a) Hydrogen sulfate ion $HSO_4^-(aq)$

(i) $HSO_4^-(aq)$ reacting with $H_3O^+(aq)$

(ii) $HSO_4^-(aq)$ reacting with $OH^-(aq)$

b) Dihydrogen phosphate ion $H_2PO_4^-(aq)$

(i) $H_2PO_4^-(aq)$ reacting with $H_3O^+(aq)$

(ii) $H_2PO_4^-(aq)$ reacting with $OH^-(aq)$

c) Explain the difference between the terms amphiprotic and amphoteric.

Reactivity 3.1.4

Understandings:

- The pH scale can be used to describe the $[H^+]$ of a solution: $pH = -\log_{10}[H^+]$; $[H^+] = 10^{-pH}$

Learning outcomes:

- Perform calculations involving the logarithmic relationship between pH and $[H^+]$.

Additional information:

- Include the estimation of pH using universal indicator, and the precise measurement of pH using a pH meter/probe.
- The equations for pH are given in the data booklet.

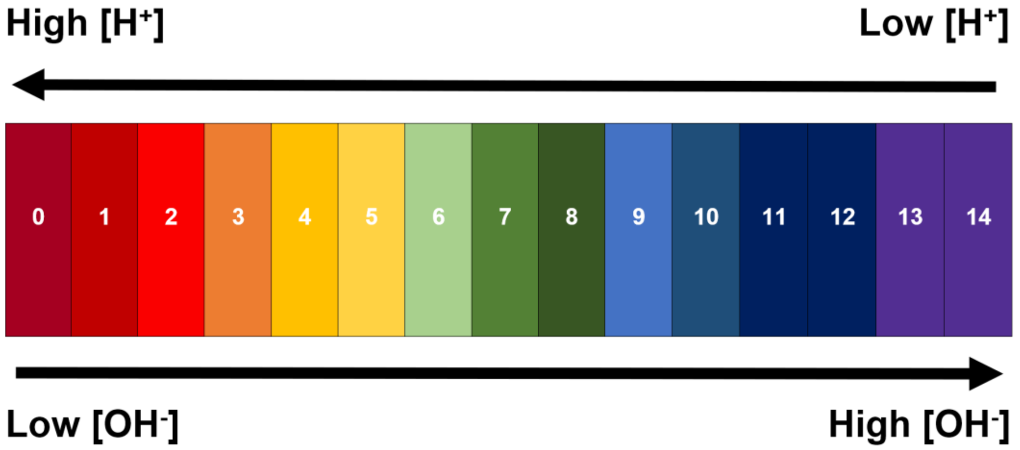
The pH scale

- The pH scale is a measure of the concentration of H^+ ions $[H^+]$ in solution.

$$pH = -\log_{10}[H^+]$$

$$[H^+] = 10^{-pH}$$

- The pH scale is inverse; a high concentration of hydrogen ions $[H^+]$ in solution results in a low pH value.
- Conversely a low concentration of hydrogen ions in solution $[H^+]$ results in a high pH value.



Exercises:

1) Complete the following table:

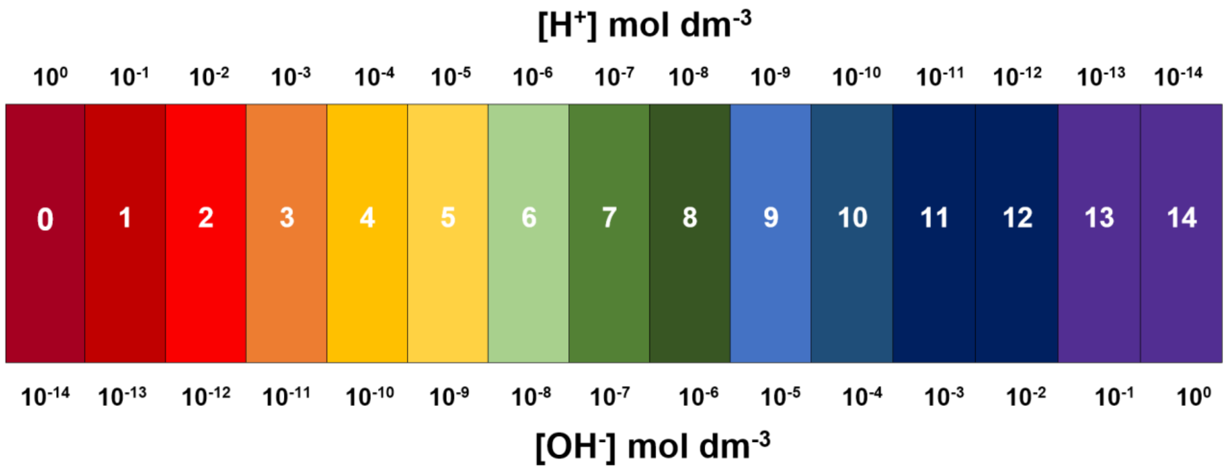
	Acidic, basic or neutral	**pH at 298 K**
$[H^+] = [OH^-]$		
$[H^+] > [OH^-]$		
$[OH^-] > [H^+]$		

2) Complete the sentences below.

If the $[H^+] = [OH^-]$, the pH is ______ at 298 K.
If the $[H^+] > [OH^-]$, the pH is ___________ than 7 at 298 K.
If the $[H^+] < [OH^-]$, the pH is ___________ than 7 at 298 K.

Changes in pH

- A change in one unit of pH represents a ten-fold change in $[H^+]$.



Example: Black coffee has a pH of 5 and toothpaste has a pH of 8. Identify which is more acidic and deduce how many times the $[H^+]$ is greater in the more acidic product.

Exercises:

1) Describe the relationship between the pH value and the $[H^+]$.

2) Calculate the following:

a) the pH of a solution that has $[H^+] = 3.2 \times 10^{-5}$ mol dm^{-3}.

b) the $[H^+]$ of a solution that has a pH of 4.6.

3) The pH of a solution is 2. If its pH is increased to 6, how many times greater is the $[H^+]$ of the original solution?

Reactivity 3.1.5

Understandings:

- The ion product constant of water, K_w, shows an inverse relationship between $[H^+]$ and $[OH^–]$. $K_w = [H^+][OH^–]$

Learning outcomes:

- Recognize solutions as acidic, neutral and basic from the relative values of $[H^+]$ and $[OH^–]$.

Additional information:

- The equation for K_w and its value at 298 K are given in the data booklet.

Linking question(s):

- Reactivity 2.3 Why does the extent of ionization of water increase as temperature increases?

Ionic product constant (K_w)

- Water ionises but only very slightly; the equilibrium position in the reaction below lies very much to the left.

$$H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq)$$

- K_w is the ionic product constant of water. It has a value of 1.00×10^{-14} at 298 K.

$$K_w = [H^+][OH^-]$$

$$K_w = 1.00 \times 10^{-14} \text{ at } 298 \text{ K}$$

Example: A solution has a pH of 3.72 at 298 K. Determine the $[H^+]$ and $[OH^-]$ of the solution.

Exercises:

1) Complete the following table (assume 298 K).

$[H^+]$ (mol dm^{-3})	$[OH^-]$ (mol dm^{-3})	pH	Acidic or basic
3.2×10^{-8}			
	1.6×10^{-10}		
3.2×10^{-9}			
	7.8×10^{-2}		

2) Lemon juice has a pH of 2.90 at 298 K. Calculate the $[H^+]$ and $[OH^-]$ in lemon juice.

Reactivity 3.1.6

Understandings:

- Strong and weak acids and bases differ in the extent of ionisation.

Learning outcomes:

- Recognise that acid–base equilibria lie in the direction of the weaker conjugate.

Additional information:

- HCl, HBr, HI, H_2SO_4 and HCl are strong acids, and group 1 hydroxides are strong bases.
- The distinction between strong and weak acids or bases and concentrated and dilute reagents should be covered.

Linking question(s):

- Reactivity 2.3 How would you expect the equilibrium constants of strong and weak acids to compare?
- Reactivity 1.1 Why does the acid strength of the hydrogen halides increase down group 17?

Names and formulae of common acid and bases

Name	Formula	Weak or strong acid or base
Sulfuric acid	H_2SO_4	Strong acid
Hydrochloric acid	HCl	Strong acid
Hydrobromic acid	HBr	Strong acid
Nitric acid	HNO_3	Strong acid
Methanoic acid	CH_3OH	Weak acid
Ethanoic acid	CH_3CH_2OH	Weak acid
Carbonic acid	H_2CO_3	Weak acid
Phosphoric acid	H_3PO_4	Weak acid
Sodium hydroxide	NaOH	Strong base
Potassium hydroxide	KOH	Strong base
Barium hydroxide	$Ba(OH)_2$	Strong base
Calcium hydroxide	$Ca(OH)_2$	Strong base
Ammonia	NH_3	Weak base
Phenylamine	$C_6H_5NH_2$	Weak base

Strong and weak acids and bases

- Strong acids completely ionise (or dissociate) in solution.
- For a strong acid the equilibrium position lies to the right.
- Strong acids are good proton donors and have weak conjugate bases.

$$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$$

- Strong bases completely ionise (or dissociate) in solution.
- For a strong base the equilibrium position lies to the right.
- Strong bases are good proton acceptors and have weak conjugate acids.

$$NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq)$$

- Weak acids partially ionise (or dissociate) in solution.
- For a weak acid the equilibrium position lies to the left.
- Weak acids are poor proton donors and have stronger conjugate bases.

$$CH_3COOH(aq) + H_2O(l) \rightleftharpoons CH_3COO^-(aq) + H_3O^+(aq)$$

- Weak bases partially ionise (or dissociate) in solution.
- For a weak base the equilibrium position lies to the left
- Weak bases are poor proton acceptors and have stronger conjugate acids.

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

Exercise: Complete the table.

	Degree of dissociation/ ionisation	**Strength of conjugate acid or base**
Strong acid		
Weak acid		
Strong base		
Weak base		

Distinguish between strong and weak acids and bases

Electrical conductivity

- Strong acids and bases have a higher concentration of mobile ions in solution, therefore they have higher electrical conductivity than weak acids and bases of equal concentration.

Reactions with active metals

- Strong acids have a higher rate of reaction with reactive metals than weak acids of equal concentration because they have a higher $[H^+]$ in solution.

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

$$Mg(s) + 2CH_3COOH(aq) \rightarrow Mg(CH_3COO)_2(aq) + H_2(g)$$

pH value

- Strong acids have a lower pH value than weak acids of equal concentration because they have a higher $[H^+]$.
- Strong bases have a higher pH value than weak bases of equal concentration because they have a higher [OH-].

Exercises:

1) Describe two different methods, one chemical and one physical, other than measuring the pH, that could be used to distinguish between ethanoic acid and hydrochloric acid solutions of the same concentration.

2) The nitrite ion is present in nitrous acid, HNO_2, which is a weak acid. The nitrate ion is present in nitric acid, HNO_3, which is a strong acid.

a) Distinguish between the terms strong and weak acid and state the equations used to show the dissociation of each acid in aqueous solution.

b) A small piece of magnesium ribbon is added to solutions of nitric and nitrous acid of the same concentration at the same temperature. Describe two observations that would allow you to distinguish between the two acids.

Calculating the pH of strong acids and bases

- Strong acids completely ionise in solution, so $[H^+]$ is assumed to be the same as the initial concentration.
- Strong bases also completely ionise in solution, so the assumption is the same as above (except that strong bases produce OH^- ions rather than H^+ ions).

Example 1: Calculate the pH of a 1.00 mol dm^{-3} solution of hydrochloric acid, HCl, at 298 K.

HCl dissociates as follows:

$$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$$

A 1.00 mol dm^{-3} solution of HCl will contain $[H^+]$ of 1.00 mol dm^{-3}.

pH = $-\log[H^+]$

pH = $-\log(1.00)$

pH = 0

Example 2: Calculate the pH of a 1.00 mol dm^{-3} solution of potassium hydroxide, KOH, at 298 K.

KOH dissociates as follows:

$$KOH(s) \rightarrow K^+(aq) + OH^-(aq)$$

A 1.00 mol dm^{-3} solution of KOH will contain $[OH^-]$ of 1.00 mol dm^{-3}.

At 298 K, $K_w = 1.00 \times 10^{-14}$

$K_w = [H^+][OH^-]$

$1.00 \times 10^{-14} = [H^+][1.00]$

$[H^+] = 1.00 \times 10^{-14} / 1.00$

$[H^+] = 1.00 \times 10^{-14}$

pH = $-\log[H^+]$

pH = $-\log(1.00 \times 10^{-14})$

pH = 14

Exercises:

1) Calculate the pH of a 0.100 mol dm^{-3} solution of nitric acid, HNO_3.

2) Calculate the pH of a 0.500 mol dm^{-3} solution of sodium hydroxide, NaOH.

Reactivity 3.1.7

Understandings:

- Acids react with bases in neutralisation reactions.

Learning outcomes:

- Formulate equations for the reactions between acids and metal oxides, metal hydroxides, hydrogencarbonates and carbonates.

Additional information:

- Identify the parent acid and base of different salts.
- Bases should include ammonia, amines, soluble carbonates and hydrogencarbonates; acids should include organic acids.

Linking question(s):

- Reactivity 1.1 Neutralisation reactions are exothermic. How can this be explained in terms of bond enthalpies?
- Reactivity 3.2 How could we classify the reaction that occurs when hydrogen gas is released from the reaction between an acid and a metal?

Properties of acids and bases

- Acids react with bases in neutralisation reactions to produce a salt and water.

metal hydroxide + acid → salt + water

$$NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$$

- Acids react with metal oxides to produce a salt and water:

metal oxide + acid → salt + water

$$MgO(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2O(l)$$

- Acids react with reactive metals (those above hydrogen on the activity series) to produce a salt and hydrogen gas.

metal + acid → salt + hydrogen

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

- Acids react with metal carbonates and hydrogen carbonates to produce a salt, carbon dioxide and water.

metal carbonate + acid → salt + water + carbon dioxide

$$Na_2CO_3(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2O(l) + CO_2(g)$$

metal hydrogen carbonate + acid → salt + water + carbon dioxide

$$NaHCO_3(s) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l) + CO_2(g)$$

Exercises:

1) Complete and balance the following equations.

a) $Zn(s) + HCl(aq) \rightarrow$

b) $Mg(s) + H_2SO_4(aq) \rightarrow$

c) $Ca(s) + HNO_3(aq) \rightarrow$

d) $Ni(s) + HCl(aq) \rightarrow$

2) Complete and balance the following equations.

a) $CaCO_3(s) + HCl(aq) \rightarrow$

b) $CaCO_3(s) + H_2SO_4(aq) \rightarrow$

c) $MgCO_3(s) + HNO_3(aq) \rightarrow$

3) Complete and balance the following equations.

a) $NaOH(aq) + HCl(aq) \rightarrow$

b) $KOH(aq) + HNO_3(aq) \rightarrow$

c) $LiOH(aq) + H_2SO_4(aq) \rightarrow$

d) $NH_3(aq) + HCl(aq) \rightarrow$

e) $NH_3(aq) + H_2SO_4(aq) \rightarrow$

f) $CH_3COOH(aq) + NaOH(aq) \rightarrow$

g) $CH_3COOH(aq) + NH_3(aq) \rightarrow$

4) Complete and balance the following equations.

a) $CuO(s) + HCl(aq) \rightarrow$

b) $MgO(s) + HNO_3(aq) \rightarrow$

c) $CaO(s) + H_2SO_4(aq) \rightarrow$

5) Complete and balance the following equations.

a) $NaHCO_3(s) + H_2SO_4(aq) \rightarrow$

b) $Mg(HCO_3)_2(s) + HNO_3(aq) \rightarrow$

Reactivity 3.1.8

Understandings:

- pH curves for neutralization reactions involving strong acids and bases have characteristic shapes and features.

Learning outcomes:

- Sketch and interpret the general shape of the pH curve.

Additional information:

- Interpretation should include the intercept with the pH axis and equivalence point.
- Only monoprotic neutralization reactions will be assessed.

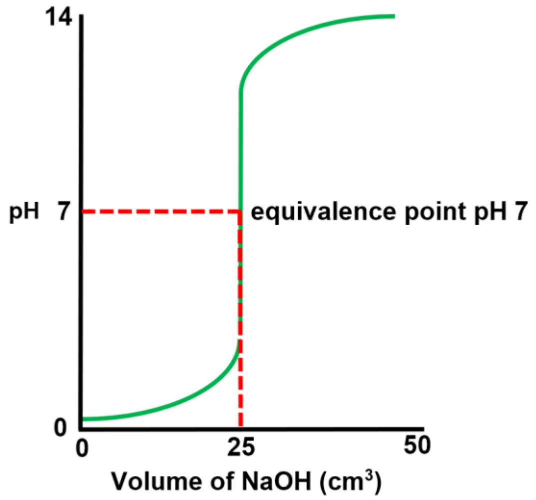
Linking question(s):

- Structure 1.4 Why is the equivalence point sometimes referred to as the stoichiometric point?

pH curves

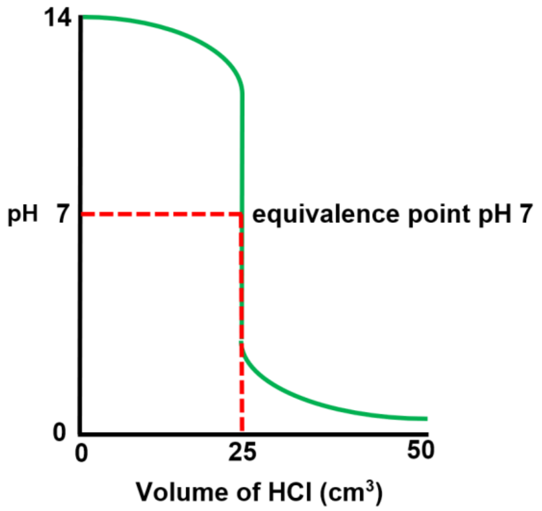
- A pH curve is produced in an acid-base titration when an acid and base react together to produce a salt and water.
- The equivalence point in an acid-base titration occurs when stoichiometrically equivalent amounts of acid and base have reacted and the solution contains only salt and water.

pH curve for the addition of a strong base to a strong acid



The pH curve is for the addition of a strong base to a strong acid.
The curve starts at a low pH – the pH of a strong acid.
The equivalence point is at pH 7 – the point at which the solution contains only salt and water.
The curve ends at a high pH – the pH of a strong base.

pH curve for the addition of a strong acid to a strong base



The pH curve is for the addition of a strong acid to a strong base.
The curve starts at a high pH – the pH of a strong base.
The equivalence point is at pH 7 – the point at which the solution contains only salt and water.
The curve ends at a low pH – the pH of a strong acid.