

# Reactivity 3.1

## Answers

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IB CHEMISTRY SL

25 <b>Mn</b> Manganese 54.938045	16 <b>S</b> Sulfur 32.065	<b>J</b>	6 <b>C</b> Carbon 12.0107	2 <b>He</b> Helium 4.002602	25 <b>Mn</b> Manganese 54.938045
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### Reactivity 3.1.1

#### Understandings:

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

#### Learning outcomes:

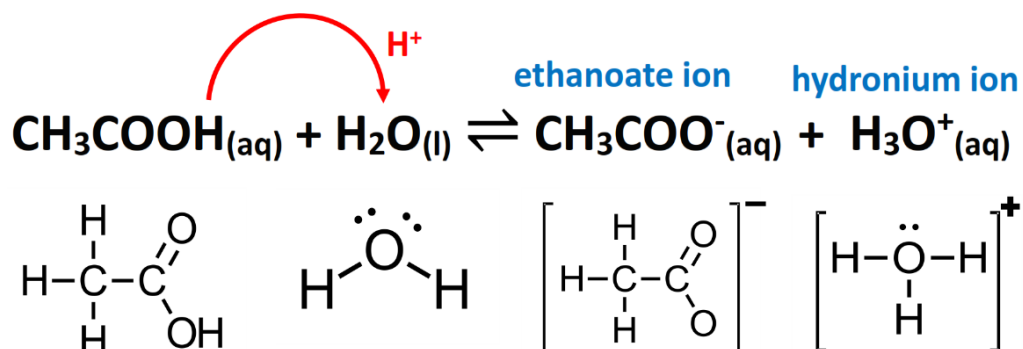
- Deduce the Brønsted–Lowry acid and base in a reaction.

#### Additional notes:

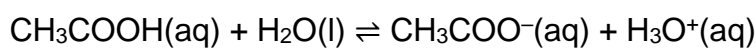
- A proton in aqueous solution can be represented as both  $\text{H}^+(\text{aq})$  and  $\text{H}_3\text{O}^+(\text{aq})$ .
- The distinction between the terms “base” and “alkali” should be understood.

### Brønsted–Lowry theory of acid and bases

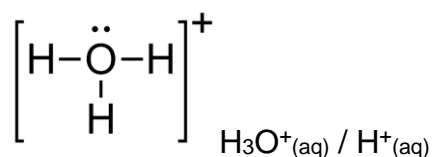
- A Brønsted–Lowry acid is a proton ( $\text{H}^+$ ) donor.
- A Brønsted–Lowry base is a proton ( $\text{H}^+$ ) acceptor.
- A proton is  $\text{H}^+(\text{aq})$  but exists in solution as  $\text{H}_3\text{O}^+(\text{aq})$ .
- The reaction below shows ethanoic acid reacting with water to form the ethanoate 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.



- The  $\text{H}^+$  ion exists in solution as  $\text{H}_3\text{O}^+(\text{aq})$ ; its Lewis structure is shown below.



### Alkalis and bases

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

## Exercises:

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

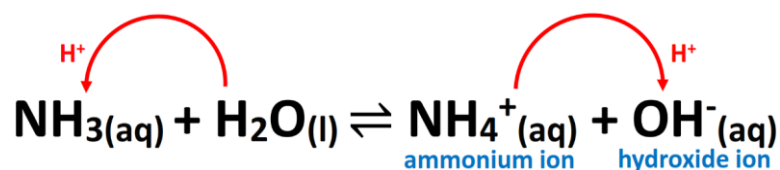
A Brønsted–Lowry acid is a proton ( $\text{H}^+$ ) donor.

A Brønsted–Lowry base is a proton ( $\text{H}^+$ ) acceptor.

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

A proton is a  $\text{H}^+$  ion. A hydrogen ion in solution can be represented as  $\text{H}^+(\text{aq})$  or  $\text{H}_3\text{O}^+(\text{aq})$ .

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



Forward reaction: Brønsted–Lowry acid is  $\text{H}_2\text{O}(\text{l})$  and Brønsted–Lowry base is  $\text{NH}_3(\text{aq})$ .

Reverse reaction: Brønsted–Lowry acid is  $\text{NH}_4^+(\text{aq})$  and Brønsted–Lowry base is  $\text{OH}^-(\text{aq})$ .

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

A base is insoluble in water whereas an alkali is soluble.

### 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?



### Reactivity 3.1.3

#### Understandings:

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

#### Learning outcomes:

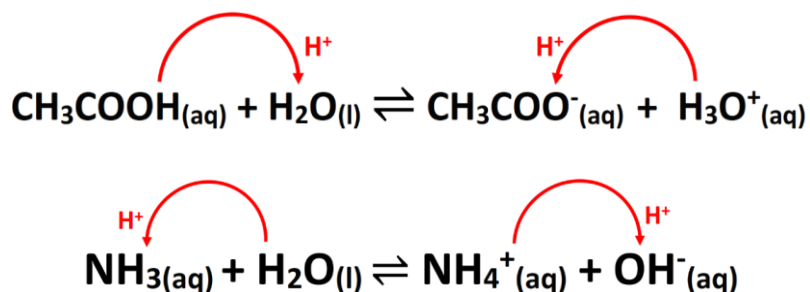
- Interpret and formulate equations to show acid–base reactions of these species.

#### 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 ( $\text{H}^+$ ) to another species and be able to accept a proton ( $\text{H}^+$ ) from another species.
- In the examples below,  $\text{H}_2\text{O}$  is acting as both a Brønsted-Lowry acid and a Brønsted-Lowry base; it is amphiprotic.



## 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:

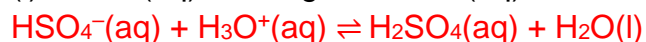
- 1) Using  $\text{H}_2\text{O}$  as an example, explain what is meant by an amphiprotic species.

An amphiprotic species is able to donate or accept a proton acting as a Brønsted-Lowry acid or a Brønsted-Lowry base.  $\text{H}_2\text{O}$  is able to donate a proton to form  $\text{OH}^-_{(\text{aq})}$  or accept a proton to form  $\text{H}_3\text{O}^+_{(\text{aq})}$ , therefore acting as a Brønsted Lowry acid and a Brønsted-Lowry base.

- 2) Write equations for the following amphiprotic species reacting with the hydronium ion ( $\text{H}_3\text{O}^+$ ) and the hydroxide ion ( $\text{OH}^-$ ).

- a) Hydrogen sulfate ion  $\text{HSO}_4^-_{(\text{aq})}$

(i)  $\text{HSO}_4^-_{(\text{aq})}$  reacting with  $\text{H}_3\text{O}^+_{(\text{aq})}$

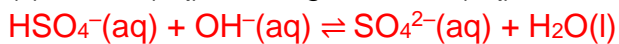


Conjugate acid-base pairs (differ by a  $\text{H}^+$ ):





(ii)  $\text{HSO}_4^-(\text{aq})$  reacting with  $\text{OH}^-(\text{aq})$

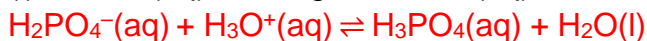


Conjugate acid-base pairs:



**b)** Dihydrogen phosphate ion  $\text{H}_2\text{PO}_4^-(\text{aq})$

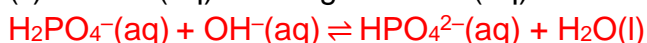
(i)  $\text{H}_2\text{PO}_4^-(\text{aq})$  reacting with  $\text{H}_3\text{O}^+(\text{aq})$



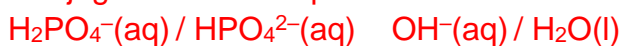
Conjugate acid-base pairs:



(ii)  $\text{H}_2\text{PO}_4^-(\text{aq})$  reacting with  $\text{OH}^-(\text{aq})$



Conjugate acid-base pairs:



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

Amphiprotic is used to describe a species that can donate or accept a proton and is specific to the Brønsted-Lowry theory of acids and bases. All amphiprotic species are also amphoteric.

Amphoteric refers to substances that can act as acids and bases and can be used in other theories of acids and bases where there is no proton transfer.

### 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.

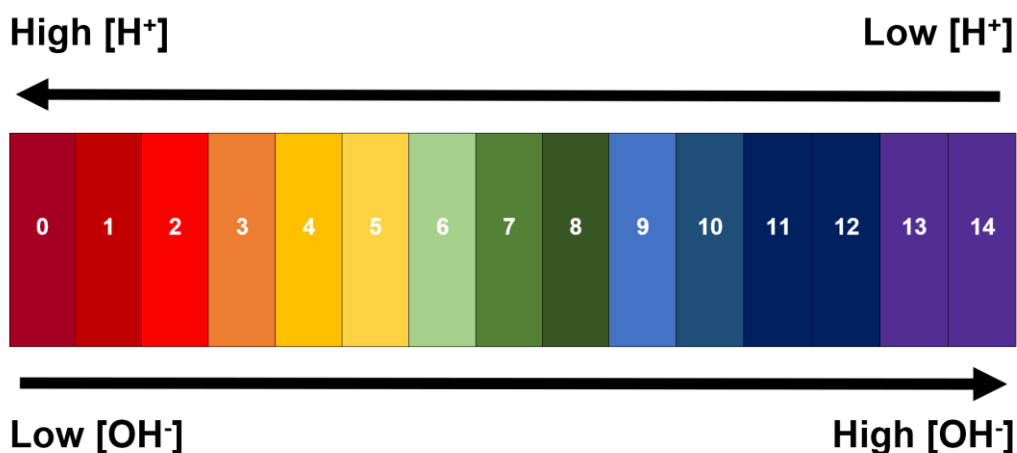
## The pH scale

- 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:

- Complete the following table:

	Acidic, basic or neutral	pH at 298 K
$[H^+] = [OH^-]$	Neutral	7
$[H^+] > [OH^-]$	Acidic	<7
$[OH^-] > [H^+]$	Basic	>7

- Complete the sentences below.

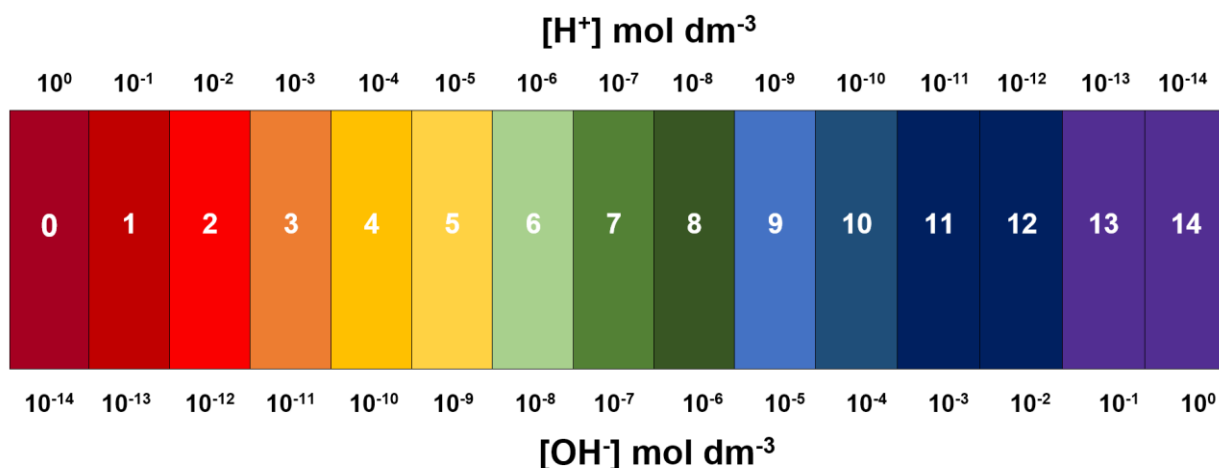
If the  $[H^+] = [OH^-]$ , the pH is **7** at 298 K.

If the  $[H^+] > [OH^-]$ , the pH is **greater** than 7 at 298 K.

If the  $[H^+] < [OH^-]$ , the pH is **less** 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.  
**Black coffee is more acidic, with 1000 times higher  $[H^+]$  than toothpaste.**

### Exercises:

- Describe the relationship between the pH value and the  $[H^+]$ .  
Inverse relationship - the lower the pH value, the higher the  $[H^+]$ . The higher the pH value, the lower the  $[H^+]$ .
- Calculate the following:
  - the pH of a solution that has  $[H^+] = 3.2 \times 10^{-5} \text{ mol dm}^{-3}$ .  
 $pH = -\log[H^+]$   
 $pH = -\log(3.2 \times 10^{-5})$   
 $pH = 4.49$
  - the  $[H^+]$  of a solution that has a pH of 4.6.  
 $[H^+] = 10^{-pH}$   
 $[H^+] = 10^{-4.6}$   
 $[H^+] = 2.51 \times 10^{-5} \text{ mol dm}^{-3}$
- 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?  
 $pH = 2$   
 $[H^+] = 10^{-2}$   
 $[H^+] = 1 \times 10^{-2} \text{ mol dm}^{-3}$   
 $pH = 6$   
 $[H^+] = 10^{-6} \text{ mol dm}^{-3}$   
 $[H^+] = 1 \times 10^{-6} \text{ mol dm}^{-3}$   
The original solution is 10000 times more acidic.

### 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:

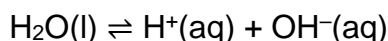
- The equation for  $K_w$  and its value at 298 K are given in the data booklet.

#### 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.



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

$$K_w = [\text{H}^+][\text{OH}^-]$$

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

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

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

$$[\text{H}^+] = 10^{-3.72}$$

$$[\text{H}^+] = 1.91 \times 10^{-4} \text{ mol dm}^{-3}$$

$$K_w = [\text{H}^+][\text{OH}^-] = 1.00 \times 10^{-14} \text{ at } 298 \text{ K}$$

$$1.00 \times 10^{-14} = 1.91 \times 10^{-4} [\text{OH}^-]$$

$$[\text{OH}^-] = 1.00 \times 10^{-14} / 1.91 \times 10^{-4}$$

$$[\text{OH}^-] = 5.24 \times 10^{-11} \text{ mol dm}^{-3}$$

### Exercises:

- Complete the following table (assume 298 K).

$[\text{H}^+]$ (mol dm <sup>-3</sup> )	$[\text{OH}^-]$ (mol dm <sup>-3</sup> )	pH	Acidic or basic
$3.2 \times 10^{-8}$	$3.1 \times 10^{-7}$	7.49	Basic
$6.3 \times 10^{-5}$	$1.6 \times 10^{-10}$	4.20	Acidic
$3.2 \times 10^{-9}$	$3.2 \times 10^{-6}$	8.50	Basic
$1.3 \times 10^{-13}$	$7.8 \times 10^{-2}$	12.89	Basic

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

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

$$[\text{H}^+] = 10^{-2.90}$$

$$[\text{H}^+] = 1.26 \times 10^{-3} \text{ mol dm}^{-3}$$

$$K_w = [\text{H}^+][\text{OH}^-] = 1.00 \times 10^{-14} \text{ at } 298 \text{ K}$$

$$1.00 \times 10^{-14} = 1.26 \times 10^{-3} [\text{OH}^-]$$

$$[\text{OH}^-] = 1.00 \times 10^{-14} / 1.26 \times 10^{-3}$$

$$[\text{OH}^-] = 7.94 \times 10^{-12} \text{ mol dm}^{-3}$$

### 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<sub>2</sub>SO<sub>4</sub> 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):

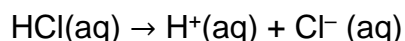
- Reactivity 2.3 How would you expect the equilibrium constants of strong and weak acids to compare?
- Reactivity 1.1 Why does the acid strength of the hydrogen halides increase down group 17?

#### Names and formulae of common acid and bases

Name	Formula	Weak or strong acid or base
Sulfuric acid	H <sub>2</sub> SO <sub>4</sub>	Strong acid
Hydrochloric acid	HCl	Strong acid
Hydrobromic acid	HBr	Strong acid
Nitric acid	HNO <sub>3</sub>	Strong acid
Methanoic acid	CH <sub>3</sub> OH	Weak acid
Ethanoic acid	CH <sub>3</sub> CH <sub>2</sub> OH	Weak acid
Carbonic acid	H <sub>2</sub> CO <sub>3</sub>	Weak acid
Phosphoric acid	H <sub>3</sub> PO <sub>4</sub>	Weak acid
Sodium hydroxide	NaOH	Strong base
Potassium hydroxide	KOH	Strong base
Barium hydroxide	Ba(OH) <sub>2</sub>	Strong base
Calcium hydroxide	Ca(OH) <sub>2</sub>	Strong base
Ammonia	NH <sub>3</sub>	Weak base
Phenylamine	C <sub>6</sub> H <sub>5</sub> NH <sub>2</sub>	Weak base

## Strong and weak acids and bases

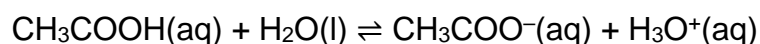
- Strong acids completely ionise (or dissociate) in solution.
- For a strong acid the equilibrium position lies to the right.
- Strong acids are good proton donors and have weak conjugate bases.



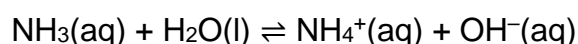
- Strong bases completely ionise (or dissociate) in solution.
- For a strong base the equilibrium position lies to the right.
- Strong bases are good proton acceptors and have weak conjugate acids.



- 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.



- 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.



**Exercise:** Complete the table.

	Degree of dissociation/ ionisation	Strength of conjugate acid or base
<b>Strong acid</b>	Completely	Weak
<b>Weak acid</b>	Partially	Strong
<b>Strong base</b>	Completely	Weak
<b>Weak base</b>	Partially	Strong



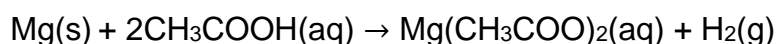
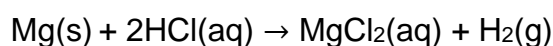
## Distinguish between strong and weak acids and bases

### 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.



### 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.

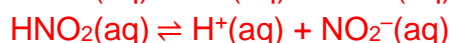
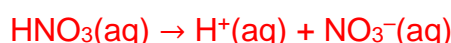
Chemical: reaction with reactive metal / Mg / Zn  
hydrochloric acid will react faster / more vigorously / ethanoic acid will react slower / less vigorously

Physical: electrical conductivity - hydrochloric acid will conduct more / higher / ethanoic acid will conduct less.

- 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.

Strong acids completely dissociated / ionised and weak acid partially dissociated / ionised



- 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.

With  $\text{HNO}_3$ :

faster rate of bubble/gas/hydrogen production

faster rate of magnesium dissolving

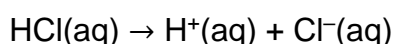
higher temperature change

## 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  $1.00 \text{ mol dm}^{-3}$  solution of hydrochloric acid, HCl, at 298 K.

HCl dissociates as follows:



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

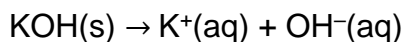
$$pH = -\log[H^+]$$

$$pH = -\log(1.00)$$

$$pH = 0$$

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

KOH dissociates as follows:



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

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  $0.100 \text{ mol dm}^{-3}$  solution of nitric acid,  $HNO_3$ .

Nitric acid is a strong acid, so  $[H^+] = 0.100 \text{ mol dm}^{-3}$ .

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

$$pH = -\log(0.100)$$

$$pH = 1$$

2) Calculate the pH of a 0.500 mol dm<sup>-3</sup> solution of sodium hydroxide, NaOH.

NaOH is a strong base, so  $[\text{OH}^-] = 0.500 \text{ mol dm}^{-3}$

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

$$K_w = [\text{H}^+][\text{OH}^-]$$

$$1.00 \times 10^{-14} = [\text{H}^+][0.500]$$

$$[\text{H}^+] = 1.00 \times 10^{-14} / 0.500$$

$$[\text{H}^+] = 2.00 \times 10^{-14}$$

$$\text{pH} = -\log [\text{H}^+]$$

$$\text{pH} = -\log(2.00 \times 10^{-14})$$

$$\text{pH} = 13.7$$

### 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:

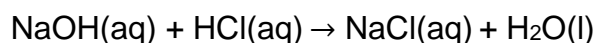
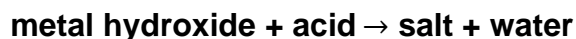
- Identify the parent acid and base of different salts.
- Bases should include ammonia, amines, soluble carbonates and hydrogencarbonates; acids should include organic acids.

#### Linking question(s):

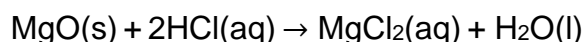
- Reactivity 1.1 Neutralisation reactions are exothermic. How can this be explained in terms of bond enthalpies?
- Reactivity 3.2 How could we classify the reaction that occurs when hydrogen gas is released from the reaction between an acid and a metal?

## Properties of acids and bases

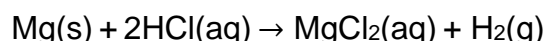
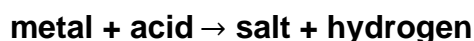
- Acids react with bases in neutralisation reactions to produce a salt and water.



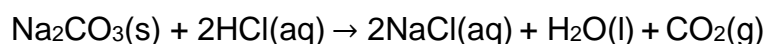
- Acids react with metal oxides to produce a salt and water:



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

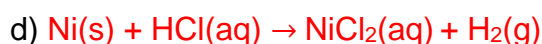
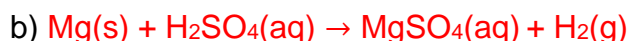
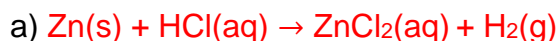


- Acids react with metal carbonates and hydrogen carbonates to produce a salt, carbon dioxide and water.

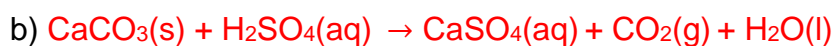


## Exercises:

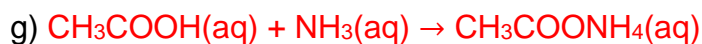
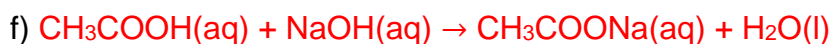
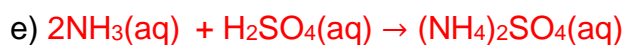
1) Complete and balance the following equations.



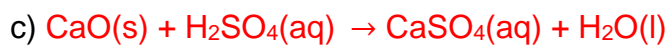
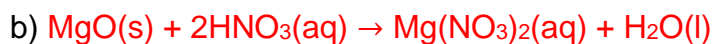
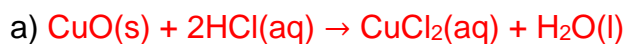
**2) Complete and balance the following equations.**



**3) Complete and balance the following equations.**



**4) Complete and balance the following equations.**



**5) Complete and balance the following equations.**



### **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):**

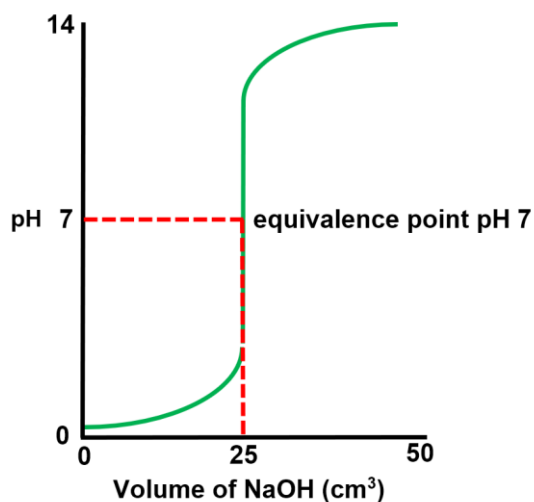
- Structure 1.4 Why is the equivalence point sometimes referred to as the stoichiometric point?



## pH curves

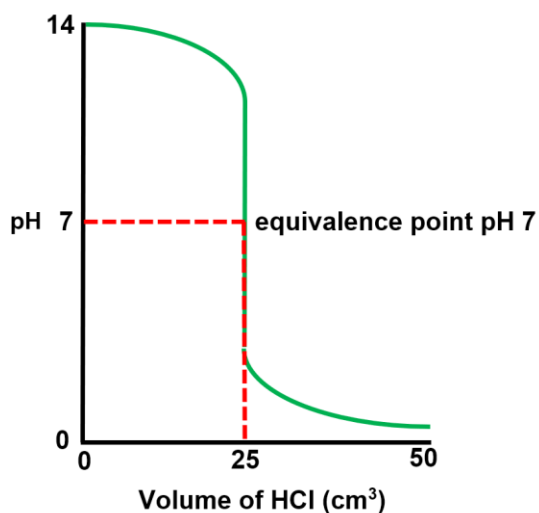
- A pH curve is produced in an acid-base titration when an acid and base react together to produce a salt and water.
- The equivalence point in an acid-base titration occurs when stoichiometrically equivalent amounts of acid and base have reacted and the solution contains only salt and water.

### 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.