

# Acids and bases

## SL (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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## 8.1 Theories of acids and bases

### Understandings:

- A Brønsted–Lowry acid is a proton/ $H^+$  donor and a Brønsted–Lowry base is a proton/ $H^+$  acceptor.
- Amphiprotic species can act as both Brønsted–Lowry acids and bases
- A pair of species differing by a single proton is called a conjugate acid-base pair.

### Applications and skills:

- Deduction of the Brønsted–Lowry acid and base in a chemical reaction.
- Deduction of the conjugate acid or conjugate base in a chemical reaction.

### Guidance:

- Lewis theory is not required here.
- The location of the proton transferred should be clearly indicated.
- Students should know the representation of a proton in aqueous solution as both  $H^+$  (aq) and  $H_3O^+$  (aq).
- The difference between the terms amphoteric and amphiprotic should be covered.

### Syllabus checklist

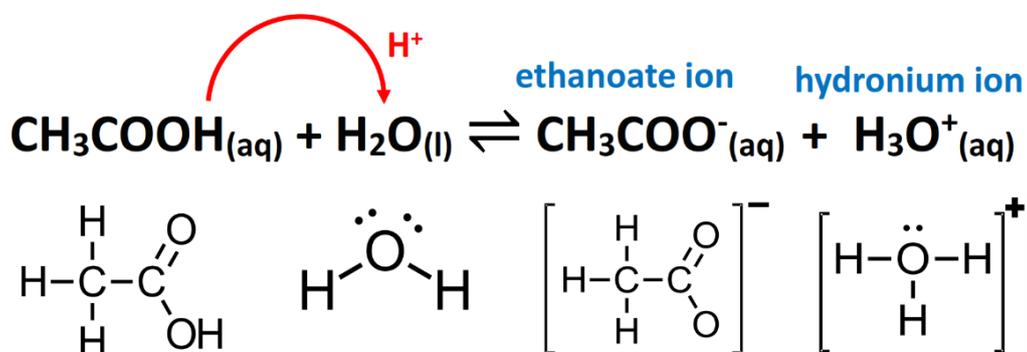
Objective	I am confident with this	I need to review this	I need help with this
Outline the Brønsted–Lowry theory of acids and bases			
Deduce the Brønsted–Lowry acid and base in a chemical reaction			
Identify the conjugate acid-base pair in a chemical reaction.			
Outline the difference between amphiprotic and amphoteric species			

### Names and formulae of common acid and bases

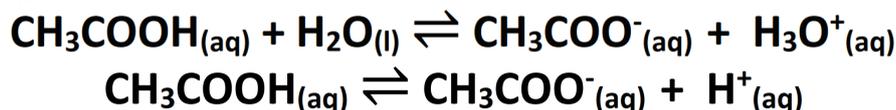
Name	Formulae	Weak or strong acid or base
Sulfuric acid	$H_2SO_4$	Strong acid
Hydrochloric acid	$HCl$	Strong acid
Nitric acid	$HNO_3$	Strong acid
Ethanoic acid	$CH_3COOH$	Weak acid
Methanoic acid	$HCOOH$	Weak acid
Carbonic acid	$H_2CO_3$	Weak acid
Phosphoric acid	$H_3PO_4$	Weak acid
Sodium hydroxide	$NaOH$	Strong base
Potassium hydroxide	$KOH$	Strong base
Barium hydroxide	$Ba(OH)_2$	Strong base
Ammonia	$NH_3$	Weak base

### Brønsted–Lowry theory of acid and bases

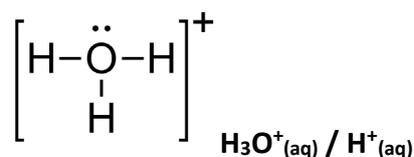
- A Brønsted–Lowry acid is a proton ( $\text{H}^+$ ) donor.
- A Brønsted–Lowry base is a proton ( $\text{H}^+$ ) acceptor.
- A proton is a  $\text{H}^+$ (aq) ion, but exists in solution as  $\text{H}_3\text{O}^+$ (aq)
- The reaction below shows ethanoic acid reacting with water to form the ethanoate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



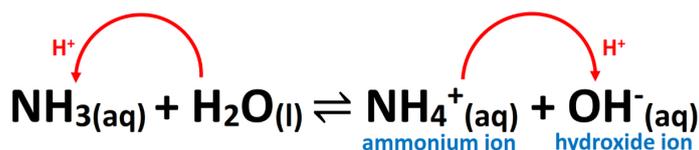
- Ethanoic acid is behaving as a Brønsted-Lowry acid and water is behaving as a Brønsted-Lowry base.
- The reaction is shown in equation form below; note that both equations show the same reaction.



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



**Exercise:** Identify the Brønsted–Lowry acid and base in the forward reaction and in the reverse reaction in the reaction below.



**Forward reaction:**

BL acid:  $\text{H}_2\text{O}$

BL base:  $\text{NH}_3$

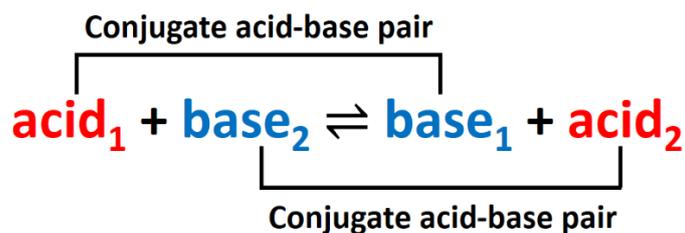
**Reverse reaction:**

BL acid:  $\text{NH}_4^+$

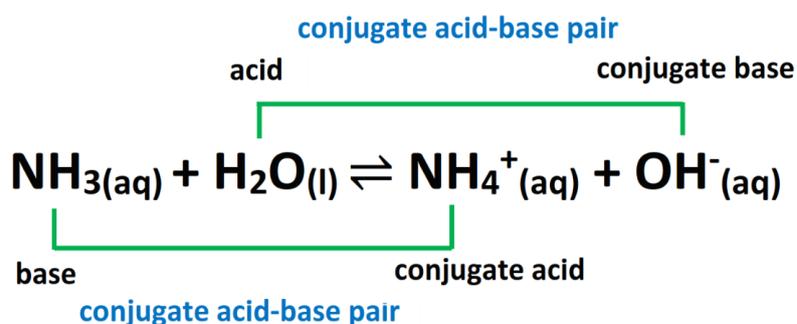
BL base:  $\text{OH}^-$

## Conjugate acid-base pairs

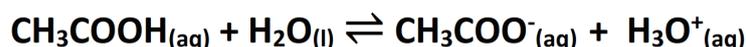
- Conjugate acid-base pairs differ by a proton.



- In the reaction below,  $\text{H}_2\text{O}$  and  $\text{OH}^-$  are conjugate acid-base pairs, as are  $\text{NH}_4^+$  and  $\text{NH}_3$ .



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



### Exercises:

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

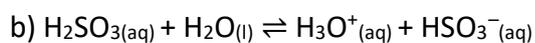
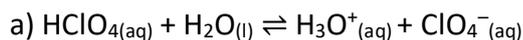
A Bronsted-Lowry acid is a proton ( $\text{H}^+$ ) donor.

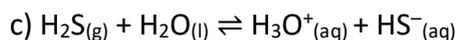
A Bronsted-Lowry base is a proton ( $\text{H}^+$ ) acceptor.

2) What is meant by a proton in the Brønsted–Lowry theory?

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

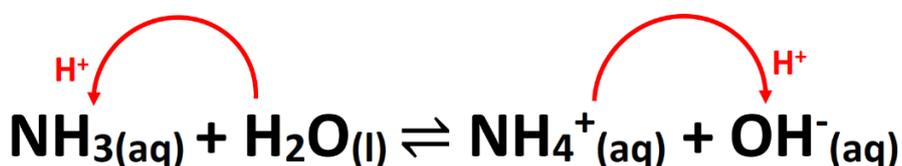
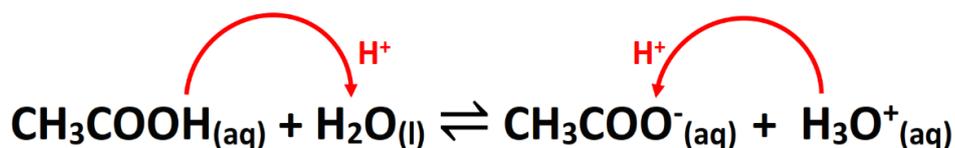
3) Identify the conjugate acid-base pairs in the following reactions:





### 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 two 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:

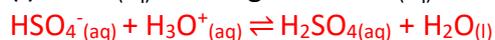
1) Using  $\text{H}_2\text{O}$  as an example, explain what is meant by an amphiprotic species.

An amphiprotic species is able to donate or accept a proton therefore 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 or a Brønsted-Lowry base.

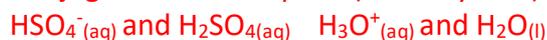
2) Write equations for the following amphiprotic species reacting with the hydronium ion ( $\text{H}_3\text{O}^+$ ) and the hydroxide ion ( $\text{OH}^-$ ).

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

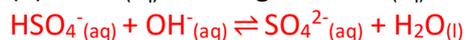
(i)  $\text{HSO}_4^-$ (aq) reacting with  $\text{H}_3\text{O}^+$ (aq)



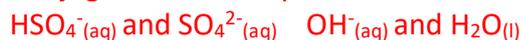
Conjugate acid-base pairs (differ by a  $\text{H}^+$ ):



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



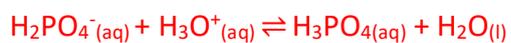
Conjugate acid-base pairs:



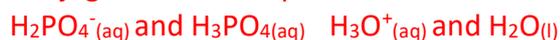
(iii) Identify the conjugate acid-base pairs in the above reactions.

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

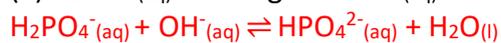
(i)  $\text{H}_2\text{PO}_4^-$ (aq) reacting with  $\text{H}_3\text{O}^+$ (aq)



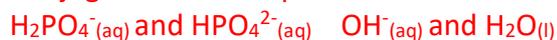
Conjugate acid-base pairs:



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



Conjugate acid-base pairs:



(iii) Identify the conjugate acid-base pairs in the above reactions.

3) 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.

## 8.2 Properties of acids and bases

### Understandings:

- Most acids have observable characteristic chemical reactions with reactive metals, metal oxides, metal hydroxides, hydrogencarbonates and carbonates.
- Salt and water are produced in exothermic neutralization reactions.

### Applications and skills:

- Balancing chemical equations for the reaction of acids.
- Identification of the acid and base needed to make different salts.
- Candidates should have experience of acid-base titrations with different indicators.

### Guidance:

- Bases which are not hydroxides, such as ammonia, soluble carbonates and hydrogen carbonates should be covered.
- The colour changes of different indicators are given in the data booklet in section 22.

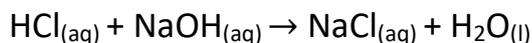
### Syllabus checklist

Objective	I am confident with this	I need to review this	I need help with this
Write balanced chemical equations for the reactions of: Acids with metal hydroxides, metal oxides, metal carbonates and hydrogencarbonates			
Identify the parent acid and base from the formula of a salt			

## Properties of acids and bases

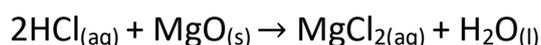
- Acids react with bases in neutralisation reactions to produce a salt and water.

**acid + metal hydroxide → salt + water**



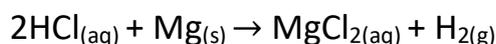
- Acids react with metal oxides to produce a salt and water:

**acid + metal oxide → salt + water**



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

**acid + metal → salt + hydrogen**



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

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

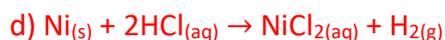
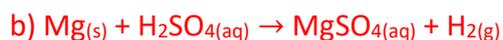
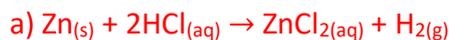


**acid + metal carbonate → salt + water + carbon dioxide**

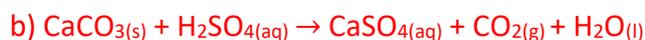
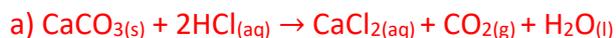


### Exercises:

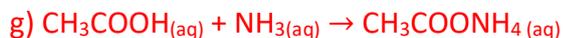
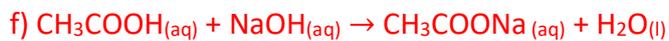
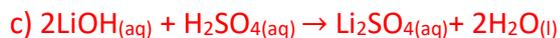
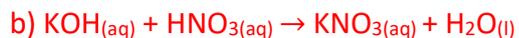
#### 1) Complete and balance the following equations:



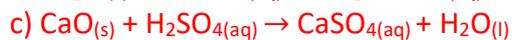
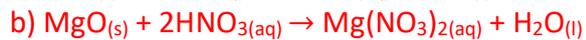
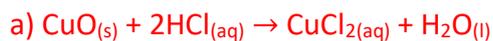
#### 2) Complete and balance the following equations:



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



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



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



### 8.3 The pH scale

#### Understandings:

- $\text{pH} = -\log[\text{H}^+]$  and  $[\text{H}^+] = 10^{-\text{pH}}$ .
- A change of one pH unit represents a 10-fold change in the hydrogen ion concentration
- Distinguish between acidic, neutral and alkaline solutions.
- The ionic product constant,  $K_w = [\text{H}^+][\text{OH}^-]$  at 298 K

#### Applications and skills:

- Solving problems including pH,  $[\text{H}^+]$  and  $[\text{OH}^-]$
- Students should be familiar with the use of a pH meter and universal indicator.

#### Guidance:

- Students will not be assessed on pOH values.
- Students should be concerned only with strong acids and bases in this sub-topic.
- Knowing the temperature dependence of  $K_w$  is not required.
- Equations involving  $\text{H}_3\text{O}^+$  instead of  $\text{H}^+$  may be applied.

#### Syllabus checklist

Objective	I am confident with this	I need to review this	I need help with this
Calculate the pH of an acid or base given the hydrogen ion concentration $[\text{H}^+]$			
Calculate the hydrogen ion concentration $[\text{H}^+]$ from the pH value			
Distinguish between acidic, basic and neutral solutions from $[\text{H}^+]$ and $[\text{OH}^-]$			
Deduce the change in hydrogen ion concentration from changes in pH and vice-versa.			
Solve problems using the ionic product constant of water, $K_w$			

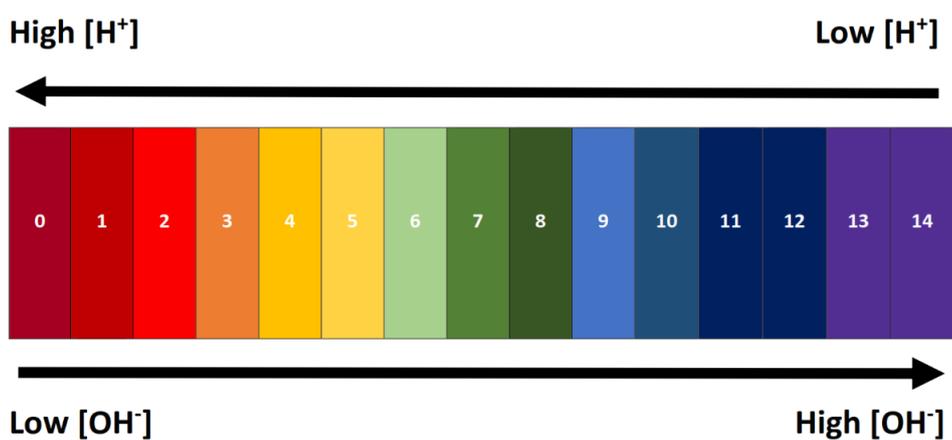
## The pH scale

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

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

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

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

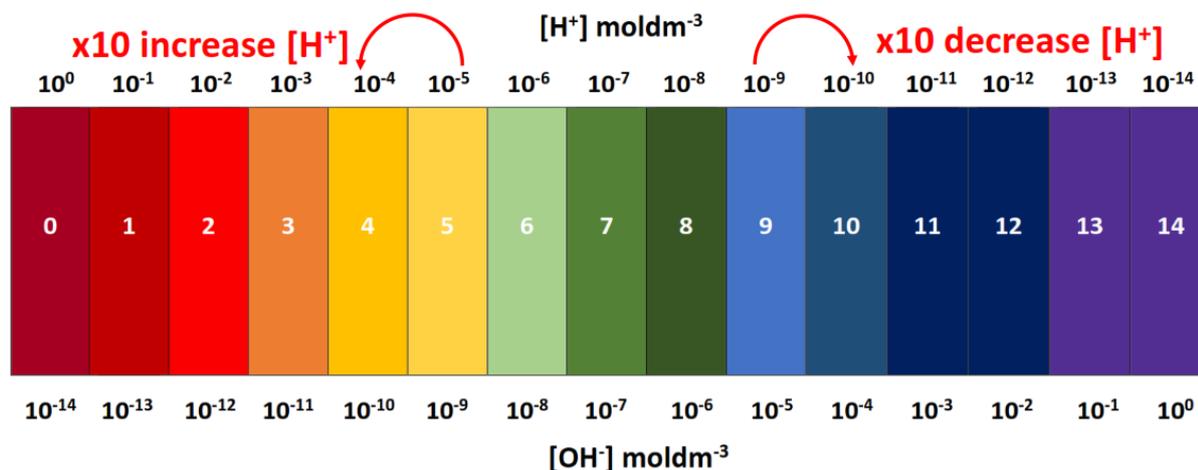


Complete the following table:

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

## 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 a 1000 times higher  $[H^+]$  than toothpaste.**

### Exercises:

1) Define the term pH.

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

2) Calculate the following:

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

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

$$\text{pH} = -\log(3.2 \times 10^{-5})$$

$$\text{pH} = 4.49$$

b) the  $[H^+]$  of a solution that has a pH of 4.600

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

$$[H^+] = 10^{-4.6}$$

$$[H^+] = 2.51 \times 10^{-5} \text{ mol dm}^{-3}$$

3) 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  $[H^+]$ .**

4) 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?

$$pH = 2$$

$$[H^+] = 10^{-2}$$

$$[H^+] = 1 \times 10^{-2} \text{ mol dm}^{-3}$$

$$pH = 6$$

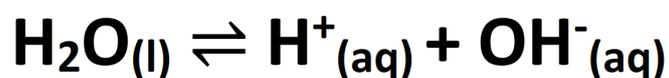
$$[H^+] = 10^{-6}$$

$$[H^+] = 1 \times 10^{-6} \text{ mol dm}^{-3}$$

The original solution is 10000 times more acidic.

#### Ionic product constant ( $K_w$ )

- Water ionises but only very slightly; the equilibrium in the reaction below lies to the left.
- The  $K_w$  is the ionic product constant of water. It has a value of  $1.00 \times 10^{-14}$  at 298 K.



$$K_c = \frac{[H^+][OH^-]}{[H_2O]}$$

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

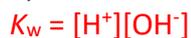
$$K_w = 1.00 \times 10^{-14} \text{ at } 298 \text{ K}$$

**Example:** Calculate the pH of pure water at 298 K.

See [video for solution](#).

**Exercises:**

1) Write the expression for the ionic product constant of water and state its value at 298 K.



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

2) 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

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

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

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

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

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

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

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

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

## 8.4 Strong and weak acids and bases

### Understandings:

- Strong and weak acids and bases differ in the extent of ionization.
- Strong acids and bases of equal concentrations have higher conductivities than weak acids and bases.
- A strong acid is a good proton donor and has a weak conjugate base.
- A strong base is a good proton acceptor and has a weak conjugate acid.

### Applications and skills:

- Distinction between strong and weak acids and bases in terms of the rates of their reactions with metals, metal oxides, metal hydroxides, metal hydrogen carbonates and metal carbonates and their electrical conductivities for solutions of equal concentrations.

### Guidance:

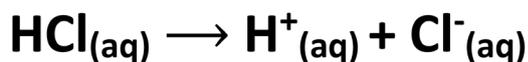
- The terms ionisation and dissociation can be used interchangeably.
- See section 21 in the data booklet for a list of weak acids and bases.

### Syllabus checklist

Objective	I am confident with this	I need to review this	I need help with this
State the degree of dissociation (or ionisation) for a weak and strong acid or base.			
Distinguish between weak and strong acids and bases based on their rates of reactions with metals, metal oxides, metal hydroxides, metal hydrogen carbonates and metal carbonates and their electrical conductivities for solutions of equal concentrations.			

### Strong and weak acids and bases

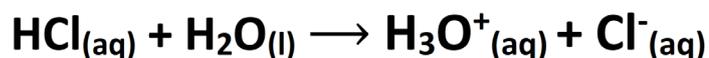
- Strong acids and bases completely ionize (or dissociate) in solution.
- For example, HCl is a strong acid and the equation for its dissociation is shown below.



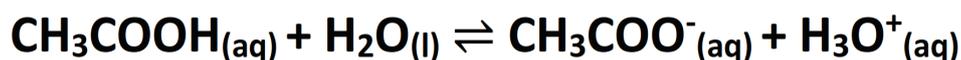
- Weak acids partially ionize (or dissociate) in solution.
- CH<sub>3</sub>COOH is a weak acid; the equilibrium sign is used to show its dissociation (the equilibrium lies to the left).



- Strong acids are good proton donors and have weak conjugate bases.
- The conjugate base of HCl, the Cl<sup>-</sup> ion, is a weak conjugate base.



- Weak acids are poor proton donors and have stronger conjugate bases.
- CH<sub>3</sub>COO<sup>-</sup>, the conjugate base of CH<sub>3</sub>COOH is a stronger conjugate base than Cl<sup>-</sup>



### Strong and weak bases

- Strong bases completely ionize (or dissociate) in solution.
- NaOH is a strong base that dissociates completely in solution, as shown below.



- Weak bases such as NH<sub>3</sub> partially ionize (or dissociate) in solution.
- The equilibrium sign is used to show its dissociation (the equilibrium lies to the left).



- Strong bases are good proton acceptors and have weak conjugate acids.
- Weak bases are poor proton acceptors and have stronger conjugate acids.

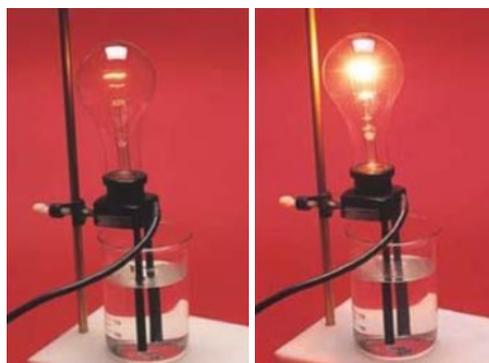
Complete the following table:

	Degree of dissociation/ ionization	Strength of conjugate acid or base
Strong acid	complete	weaker
Weak acid	partial	stronger
Strong base	complete	weaker
Weak base	partial	stronger

**Distinguish between strong and weak acids and bases**

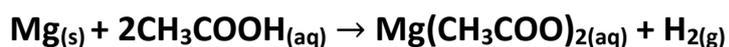
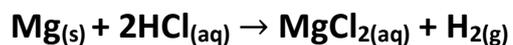
**Electrical conductivity**

- 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.



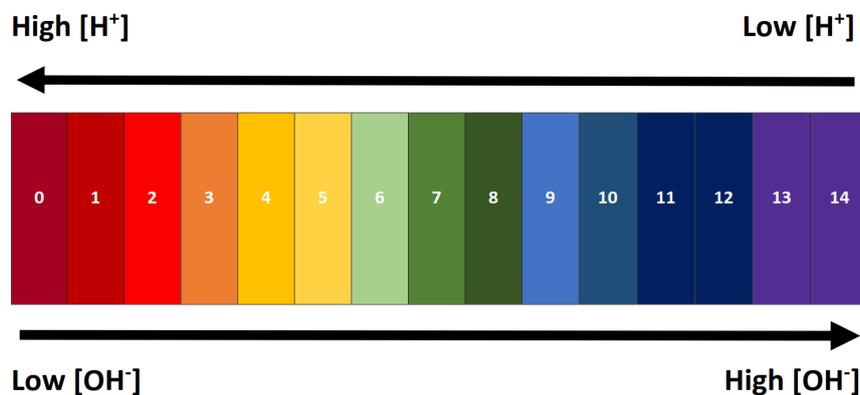
**Reactions with active metals**

- 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

- Strong acids have a lower pH than weak acids of equal concentration because they have a higher  $[H^+]$  in solution.
- Recall that a higher  $[H^+]$  equals a lower pH value.



### 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.

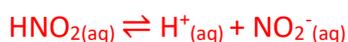
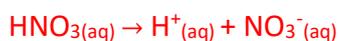
#### Chemical

reaction with reactive metal/Mg/Zn/carbonate/hydrogen carbonate  
hydrochloric acid would react faster/more vigorously / ethanoic acid would react slower/less vigorously

Physical: 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 acid completely dissociated/ionized and weak acid partially dissociated/ionized



- 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.

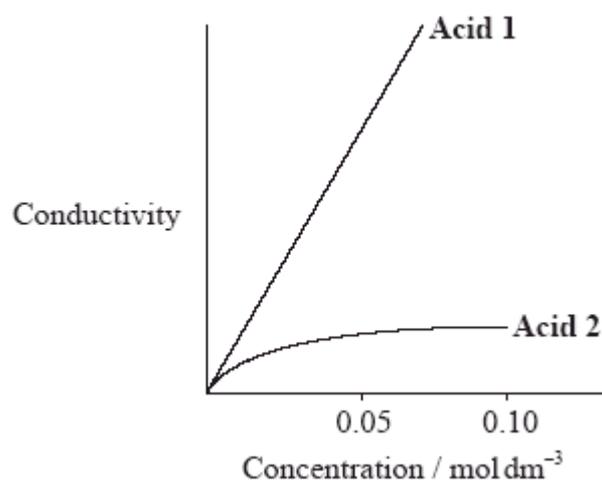
With  $\text{HNO}_3$ :

faster rate of bubble/gas/hydrogen production

faster rate of magnesium dissolving

higher temperature change

- c. The graph below shows how the conductivity of the two acids changes with concentration.



Identify Acid 1 and explain your choice.

Acid 1 is  $\text{HNO}_3$

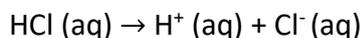
$\text{HNO}_3$  is a strong acid that is fully dissociated, therefore, it has higher conductivity because there is a higher concentration of mobile ions in solution.

## Calculating the pH of strong acids and bases

- Strong acids completely ionise in solution, so the  $[H^+]$  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.

HCl dissociates as follows:



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

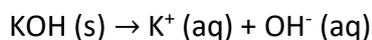
$$pH = -\log[H^+]$$

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

$$pH = 0$$

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

KOH dissociates as follows:



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

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

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

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

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

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

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

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

$$pH = 14$$

### Exercise:

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  $0.500 \text{ mol dm}^{-3}$  solution of sodium hydroxide, NaOH.

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

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

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

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

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

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

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

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

$$\text{pH} = 13.7$$

## 8.5 Acid deposition

### Understandings:

- Rain is naturally acidic because of dissolved  $\text{CO}_2$  and has a pH of 5.6. Acid deposition has a pH below 5.0.
- Acid deposition is formed when nitrogen or sulfur oxides dissolve in water to form  $\text{HNO}_3$ ,  $\text{HNO}_2$ ,  $\text{H}_2\text{SO}_4$  and  $\text{H}_2\text{SO}_3$ .
- Sources of the oxides of sulfur and nitrogen and the effects of acid deposition should be covered.

### Applications and skills:

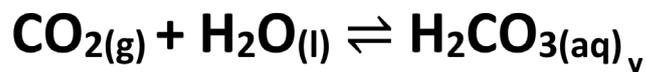
- Balancing the equations that describe the combustion of sulfur and nitrogen to their oxides and the subsequent formation of  $\text{HNO}_3$ ,  $\text{HNO}_2$ ,  $\text{H}_2\text{SO}_4$  and  $\text{H}_2\text{SO}_3$ .
- Distinction between the pre-combustion and post-combustion methods of reducing sulfur oxides emissions.
- Deduction of acid deposition equations for acid deposition with reactive metals and carbonates.

### Syllabus checklist

Objective	I am confident with this	I need to review this	I need help with this
Outline why rain is naturally acidic with a pH of 5.6			
Write balanced equations for the reactions for the formation of $\text{HNO}_3$ , $\text{HNO}_2$ , $\text{H}_2\text{SO}_4$ and $\text{H}_2\text{SO}_3$ in the atmosphere			
Outline pre- and post-combustion methods of reducing $\text{SO}_2$ emissions			
Deduce equations for the reactions of $\text{HNO}_3$ and $\text{H}_2\text{SO}_4$ with reactive metals and carbonates			

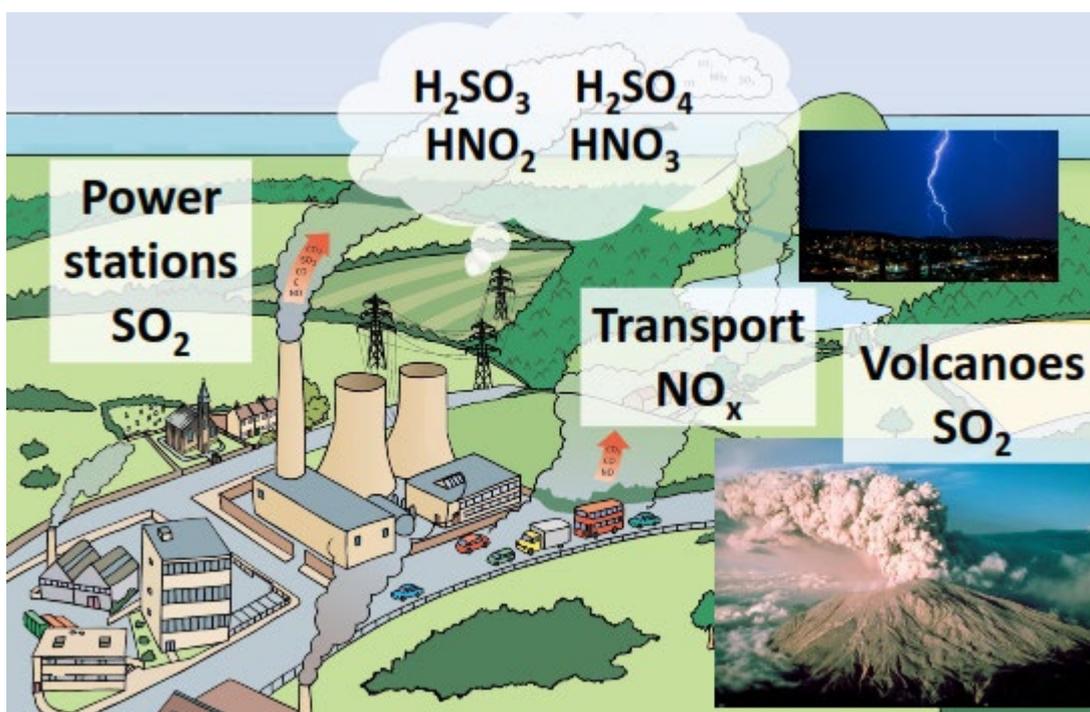
## Acid deposition

- Unpolluted rainwater is naturally acidic with a pH of 5.6
- The equation below shows that carbon dioxide dissolves in water to form carbonic acid  $\text{H}_2\text{CO}_3$ .



- Acid rain has a pH of less than 5.0
- Wet deposition includes acid rain, fog and snow.
- Dry deposition includes acidic gases and particles.

## Sources of acid deposition



**Exercise:** From the diagram above, identify some natural and anthropogenic sources (caused by human activity) of acid deposition.

### Anthropogenic:

NO<sub>x</sub> – from the combination of atmospheric nitrogen and oxygen at high temperatures in internal combustion engines.

SO<sub>2</sub> – combustion of coal that contains sulfur.

### Natural:

NO<sub>x</sub> – lightning

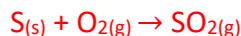
SO<sub>2</sub> – volcanoes

## Formation of acid rain

### Sulfuric and sulfurous acid

- Sulfuric acid ( $\text{H}_2\text{SO}_4$ ) and sulfurous acid ( $\text{H}_2\text{SO}_3$ ) are formed by the combustion of coal that contains high levels of sulfur.

- Sulfur (S) burns in oxygen ( $\text{O}_2$ ) to form sulfur dioxide ( $\text{SO}_2$ )



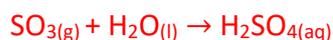
- $\text{SO}_2$  dissolves in water to form sulfurous acid ( $\text{H}_2\text{SO}_3$ )



- $\text{SO}_2$  can react with  $\text{O}_2$  to form sulfur trioxide ( $\text{SO}_3$ )



- $\text{SO}_3$  dissolves in water to form  $\text{H}_2\text{SO}_4$



### Nitric and nitrous acid

- Nitric ( $\text{HNO}_3$ ) and nitrous acid ( $\text{HNO}_2$ ) are formed from the reaction of nitrogen ( $\text{N}_2$ ) and oxygen ( $\text{O}_2$ ) at high temperatures in internal combustion engines.

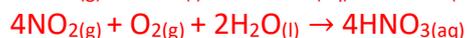
- Nitrogen ( $\text{N}_2$ ) and oxygen ( $\text{O}_2$ ) react at high temperatures in internal combustion engines to form nitrogen monoxide (NO)



- NO reacts with  $\text{O}_2$  to form nitrogen dioxide ( $\text{NO}_2$ )



- $\text{NO}_2$  dissolves in water to form  $\text{HNO}_3$  and  $\text{HNO}_2$



**Exercises:**

1. State an equation to show why rain water is naturally acidic.



2. State a natural and anthropogenic source of sulfur dioxide (SO<sub>2</sub>) and nitrogen monoxide (NO).

**SO<sub>2</sub>**

Anthropogenic source: combustion of coal containing sulfur in power stations

Natural source: volcanoes

**NO<sub>x</sub>**

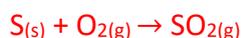
Anthropogenic source: reaction of nitrogen with oxygen at high temperatures in an internal combustion engine

Natural source: lightning

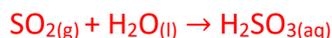
3. Acid rain has a pH of less than 5.0. Explain how the burning of coal can lead to the formation of acid rain.

Coal, containing sulfur, is burned in a power station to produce heat.

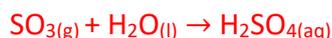
The sulfur reacts with oxygen to form sulfur dioxide



The sulfur dioxide then reacts with water to form sulfurous acid

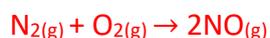


The sulfur dioxide can react with oxygen to form sulfur trioxide which then reacts with water to form sulfuric acid.



4. Outline the process responsible for the production of acid rain from the oxides of nitrogen.

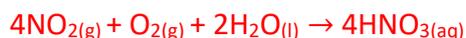
Nitrogen reacts with oxygen at high temperatures inside an internal combustion engine to produce nitrogen monoxide.



The nitrogen monoxide reacts with oxygen to form nitrogen dioxide.



The nitrogen dioxide then reacts with water (or oxygen and water) to form nitrous and nitric acids.



## Effects of acid deposition

- Acid rain reacts with calcium carbonate ( $\text{CaCO}_3$ ) in statues.



**Exercise:** Write equations for the reaction of calcium carbonate with sulfuric and nitric acid.

Reaction with sulfuric and nitric acid:



## Reduction of $\text{SO}_2$ emissions

**Pre-combustion methods** (remove or reduce sulfur in oil or coal before combustion).

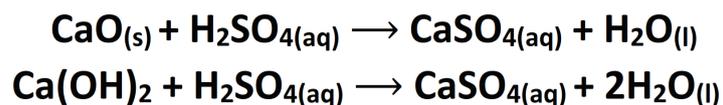
- Hydrodesulfurization is a chemical process that uses a catalyst to remove sulfur from natural gas and from refined petroleum products.
- The removal of sulfur reduces the  $\text{SO}_2$  emissions when fuels are burned.

**Post-combustion methods** (remove acidic gases after combustion by reacting with a base).

- Flu-gas desulfurization can remove up to 90% of  $\text{SO}_2$  emissions from power stations.

## Reducing acidity of water

- Powdered  $\text{CaCO}_3$  is added to lakes to reduce the acidity.
- $\text{CaO}$  and  $\text{Ca}(\text{OH})_2$  are also used to neutralise acidic water.



## Exercises:

1. Outline one pre-combustion and one post-combustion method of reducing sulfur emissions.

**Pre-combustion methods** (remove or reduce sulfur in oil or coal before combustion)

**Post-combustion methods** (remove acidic gases after combustion by reacting with a base)

2. Outline how acidic lake water can be neutralised.

The addition of CaO and Ca(OH)<sub>2</sub> neutralises acidic water.

