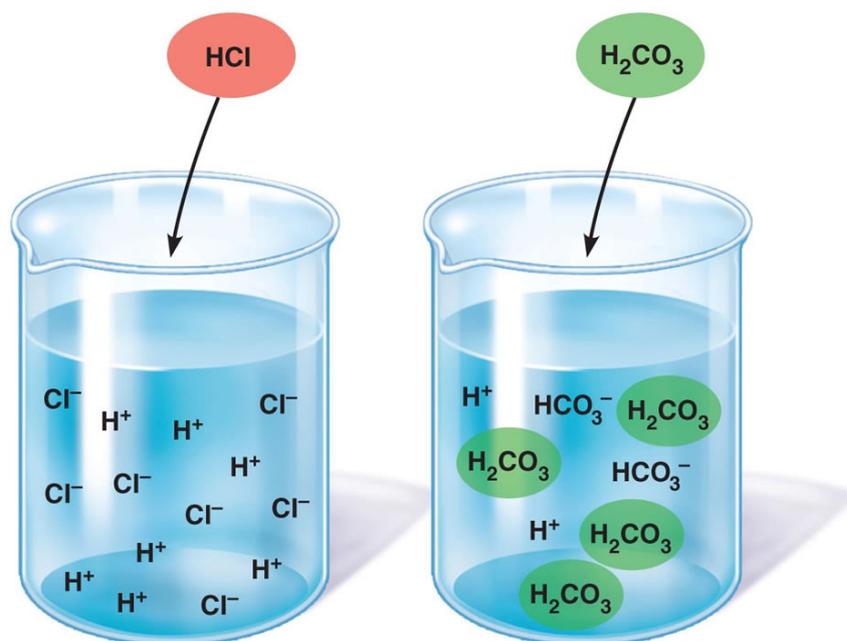


# Advanced Higher Chemistry (Revised)

## Unit 2 - Physical Chemistry

### *Weak Acids & Bases*



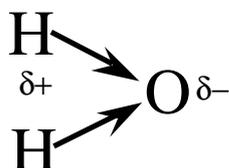
**(a) A strong acid such as HCl dissociates completely into its ions.**

**(b) A weak acid such as  $\text{H}_2\text{CO}_3$  does not dissociate completely.**

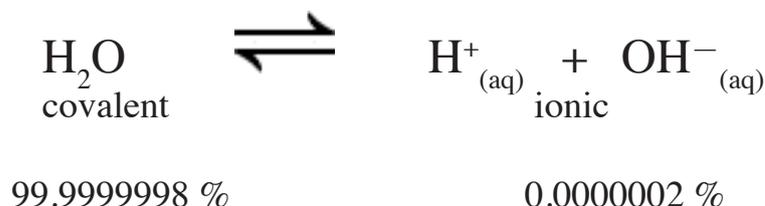
## Primer Notes (stuff that used to be covered in Higher Course)

## Water Equilibrium

This activity considers the equilibrium reactions that take place between the covalent and ionic forms of water.

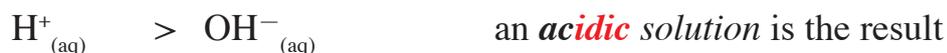


Water molecules have a tendency to 'react' to form  $\text{H}^+_{(\text{aq})}$  and  $\text{OH}^-_{(\text{aq})}$  ions. This reaction is, however, **reversible** so an **equilibrium mixture** exists.



The **equilibrium** lies well over to the **left**, though there are still enough **ions** present to make water a reasonably good **conductor** at high voltages.

In **pure water** the **number** of  $\text{H}^+_{(\text{aq})}$  and  $\text{OH}^-_{(\text{aq})}$  ions are always the **same**, and water is **neutral**. When other chemicals are added to water, this **equilibrium** can be **disturbed** resulting in



Much of this Topic will deal with **Equilibria in Solutions**.

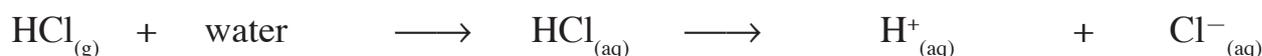
## Strong & Weak Acids

This lesson is about what strong and weak acids are about and why their properties differ.

### Definitions

This activity links the terms strong acid and weak acid with what happens when an acid solution is formed.

**Hydrochloric acid**, along with **sulphuric** and **nitric**, remain our 'main 3 acids'. From now on, they are our 'main 3 **strong acids**'. So what does the word "strong" tell you about an acid? All acids start off as (polar) **covalent** substances.



When **dissolved** in water, the **covalent** acid molecules **dissociate** (split up) into **ions**. With **strong acids** this **conversion** is 100% and a **single arrow**  $\longrightarrow$  is correct.

**A strong acid is one that is completely dissociated into ions in solution**

**Ethanoic acid**, along with other **carboxylic acids**, will be our ‘**main weak acids**’. So what does the word “**weak**” tell you about an acid? Again all acids start off as (**polar**) **covalent** substances.



When *dissolved* in water, very few of the **covalent** acid molecules *dissociate* (split up) into **ions**. With **weak acids** this *conversion* is usually less than 1% . In fact the **reverse reaction** dominates so an **equilibrium mixture** is formed that contains *very few ions* compared to **covalent molecules**. *The equilibrium lies well over to the left.*

*A weak acid is one that is only **partially dissociated** into ions in solution*

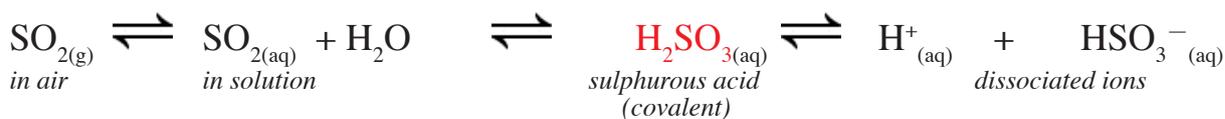
Other examples of **weak acids** are less ‘*obvious*’ and rely more on knowledge gained over the Standard grade and Higher courses. For example,

CO<sub>2</sub> is a **soluble** gas that (like other **non-metal oxides**) will *dissolve/react* with **water** to produce an **acidic solution**.

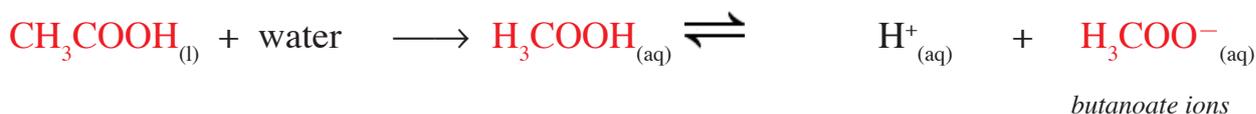
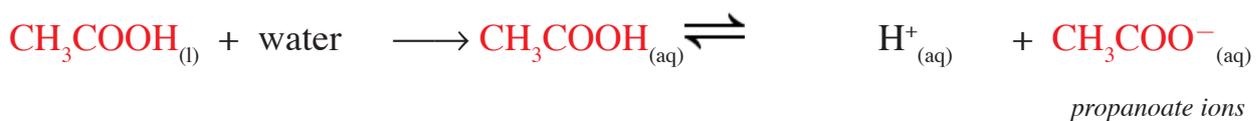
Solutions of CO<sub>2</sub> (think of fizzy drinks) lose CO<sub>2</sub> quickly when open to the air i.e. even the dissolving of CO<sub>2</sub> is a **reversible reaction**.



Similarly for SO<sub>2</sub> :

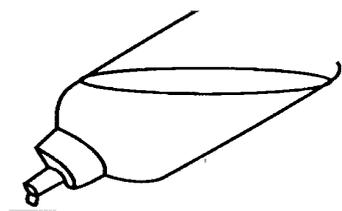


Most of the time, however, you can expect to deal mainly with **ethanoic**, **propanoic**, and **butanoic** acid.



## Comparing Strong & Weak Acids

This activity compares three properties of hydrochloric acid with the same three properties of ethanoic acid



Both acids are **equimolar** - they contain exactly the **same number** of moles (**same number** of molecules) per litre of solution.

Both acids are **monoprotic** (meaning that they are *capable* of releasing one **hydrogen ion** per molecule).

Comparison	0.1 M HCl	0.1 M CH <sub>3</sub> COOH
pH		
Conductivity		
Reaction Rate		

### Discussion of Results

**pH** there were **equal numbers** of molecules of acid dissolved in both solutions and 100% of the **HCl** molecules **dissociated** to produce many H<sup>+</sup> - **very low pH**

Less than 1% of the **CH<sub>3</sub>COOH** molecules **dissociated** so there were **much fewer** H<sup>+</sup> ions present - pH still < 7, but **higher** than HCl.

**Conductivity** there were **equal numbers** of molecules of acid dissolved in both solutions and 100% of the **HCl** molecules **dissociated** to produce many H<sup>+</sup> - **very high conductivity**

Less than 1% of the **CH<sub>3</sub>COOH** molecules **dissociated** so there were **much fewer** H<sup>+</sup> ions present - conducts, but **not as good as HCl**.

**Reaction Rate** there were **equal numbers** of molecules of acid dissolved in both solutions and 100% of the **HCl** molecules **dissociated** to produce many H<sup>+</sup> - **very fast reaction**

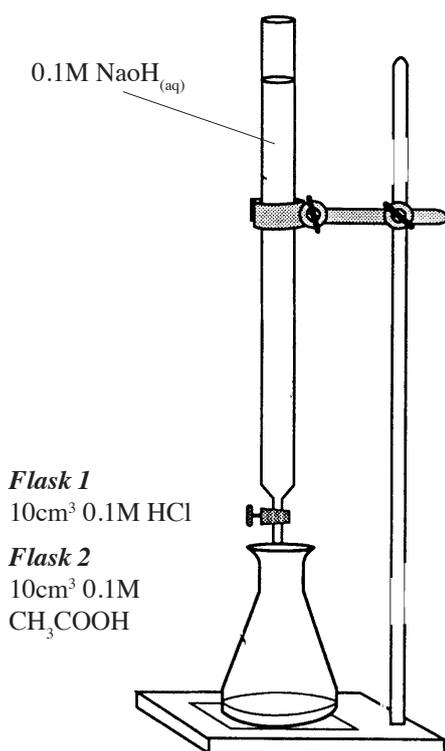
Less than 1% of the **CH<sub>3</sub>COOH** molecules **dissociated** so there were **much fewer** H<sup>+</sup> ions present - **fewer collisions** with magnesium ribbon so rate of **reaction is slower**.

### Conclusion

**Weak acids produce less H<sup>+</sup> ions than strong acids and, therefore, are less acidic than strong acids of equal concentration**

## Titrating Strong & Weak Acids

This activity compares the amounts of alkali required to neutralise equal quantities of a strong and weak acid



Vol. needed to **neutralise**  
 $10 \text{ cm}^3 \text{ 0.1M HCl} = \quad \text{cm}^3$

Vol. needed to **neutralise**  
 $10 \text{ cm}^3 \text{ 0.1M CH}_3\text{COOH} = \quad \text{cm}^3$

Perhaps surprisingly, it takes **exactly the same amount** of NaOH to **neutralise** a **strong** acid and a **weak** acid.

Both acids were the **same concentration** so they both contained the **same number of molecules** to begin with.

The strong acid will have **dissociated completely** meaning that **all** the  $\text{H}^+_{(\text{aq})}$  ions were available to react with the  $\text{OH}^-_{(\text{aq})}$  ions added from the **burette**.

The weak acid is only **partially dissociated** so less than 1% of the  $\text{H}^+_{(\text{aq})}$  ions are available at the beginning.

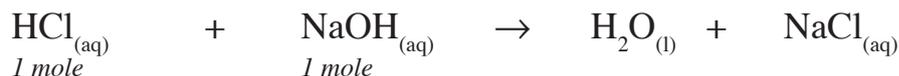
Weak acid means **reversible** reactions;  $\text{CH}_3\text{COOH}_{(\text{aq})} \rightleftharpoons \text{H}^+_{(\text{aq})} + \text{CH}_3\text{COO}^-_{(\text{aq})}$   
ethanoate ions

However, as the  $\text{H}^+_{(\text{aq})}$  ions react with  $\text{OH}^-_{(\text{aq})}$  they are **effectively removed** from the **equilibrium mixture**. This **slows down** or even **stops** the **reverse reaction**, but the **forward reaction continues**:



More  $\text{H}^+_{(\text{aq})}$  ions will be produced which will then react with the  $\text{OH}^-_{(\text{aq})}$  ions and so on until every single  $\text{CH}_3\text{COOH}_{(\text{aq})}$  molecule **dissociates** to form  $\text{H}^+_{(\text{aq})}$  ions.

Overall, a weak acid can supply **exactly the same number of  $\text{H}^+_{(\text{aq})}$**  as an equal quantity of a strong acid.



**Stoichiometry** is the word used to describe the numerical proportions (usually expressed in moles) of substances involved in reactions.

Strong and weak acids have **different properties**, but they have the **same stoichiometry**.

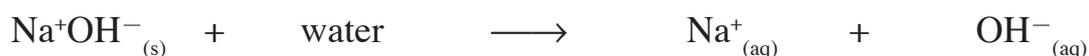
## Strong and Weak Bases

This lesson is about strong and weak bases and how their properties differ

### Two Examples of Bases

This activity considers whether a base is strong or weak in terms of whether or not it gives an equilibrium mixture in water.

**Sodium hydroxide**, along with **potassium hydroxide** and **lithium hydroxide**, were our ‘main 3 alkalis’. From now on, they are our ‘main 3 **strong** bases’. So what does the word “strong” tell you about a base? Most bases start off as **ionic** substances, (*oxides* or *hydroxides*)



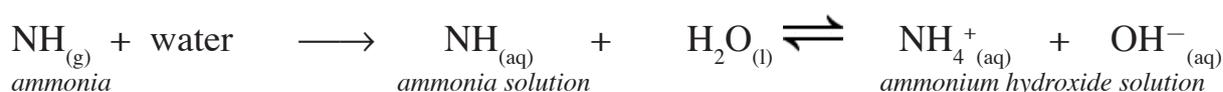
When *dissolved* in water, the **bases dissociate** (split up) into separate **ions**. With **strong** bases this *conversion* is 100% and a **single** arrow  $\longrightarrow$  is correct.

A **strong base** is **completely dissociated** to form  $\text{OH}^-$  ions in solution

Less **soluble** hydroxides (*Data Book*) are unlikely to **dissociate completely** (except at very low **concentrations**) and will be considered **weak bases**:

Group I	metal hydroxides	-	very soluble
Group II	metal hydroxides	-	slightly soluble
Transition	metal hydroxides	-	insoluble

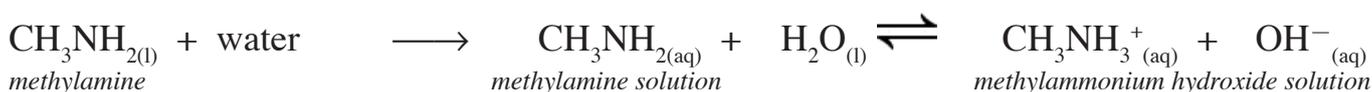
The **most important weak base** that you *must* .... repeat, **must** know about is **ammonia**,  $\text{NH}_3$ . Similar to **weak acids**, **ammonia** is a **polar covalent** molecule that *dissolves readily* in water (**fountain experiment**). Most of the **ammonia** (99.6%) *remains as molecules*, but a small amount *reacts with water* to produce **hydroxide** ions. This reaction is **reversible**.



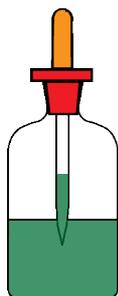
Bottles can be labelled ‘*ammonia solution*’ or ‘*ammonium hydroxide solution*’ - they are both the **same mixture**. Since the **equilibrium** lies well *over to the left*, the more correct label would be ‘*ammonia solution*’

A **weak base** is only **partially dissociated** to form  $\text{OH}^-$  ions in solution

Other examples of **weak bases** would be the *organic* equivalents of **ammonia** - the *amines*



## Comparing pH & Conductivity



This activity compares the pH and conductivity of equimolar solutions of a strong base and a weak base

Both bases are **equimolar** - they contain exactly the **same number** of moles (**same number** of molecules/units) per litre of solution.

Both bases are **monobasic** (meaning that they are *capable* of

Comparison	0.1 M NaOH	0.1 M NH <sub>3</sub>
pH		
Conductivity		

### Discussion of Results

**pH** there were **equal numbers** of molecules/units of base dissolved in both solutions and 100% of the **NaOH** units **dissociated** to produce **many** OH<sup>-</sup> - **very high pH**

Less than 1% of the **NH<sub>3</sub>** molecules **dissociated** so there were **much fewer** OH<sup>-</sup> ions present - pH still >7, but **lower** than NaOH

**Conductivity** there were **equal numbers** of molecules/units of base dissolved in both solutions and 100% of the **NaOH** units **dissociated** to produce **many** OH<sup>-</sup> - **high conductivity**

Less than 1% of the **NH<sub>3</sub>** molecules **dissociated** so there were **much fewer** OH<sup>-</sup> ions present - conducts, but **not as good as NaOH**.

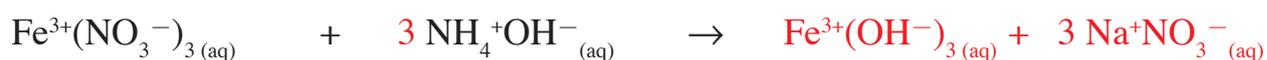
### Conclusion

*Weak bases produce less OH<sup>-</sup> ions than strong bases and, therefore, are less alkaline than strong bases of equal concentration*

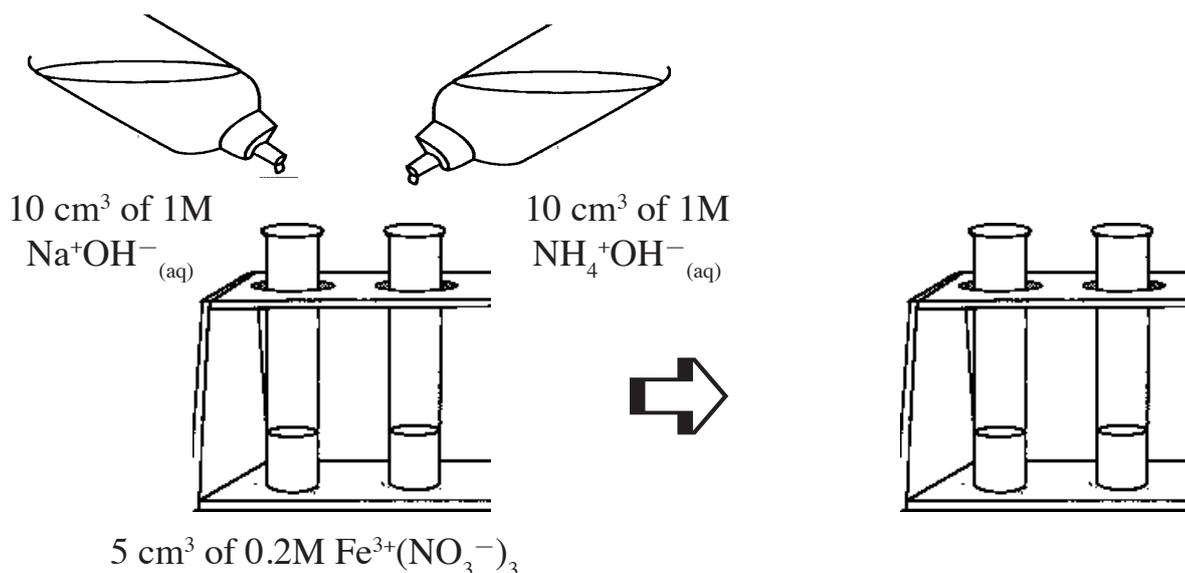
## Stoichiometry of Reactions

This activity compares the amounts of precipitate produced with iron (III) nitrate solution using equal amounts of a strong and a weak base.

The equations for the two reactions will look very similar:



The only difference between these two reactions is that the NaOH has 100% of its  $\text{OH}^-$  ions available to react, while the ammonia/ammonium hydroxide has less than 0.4%.



Perhaps surprisingly, *exactly the same amount* of precipitate is produced by a **strong** base and a **weak** base.

Both bases were the **same concentration** so they both contained the **same number of molecules / units** to begin with.

The strong base will have **dissociated completely** meaning that **all** the  $\text{OH}^-_{(\text{aq})}$  ions were available to react with the  $\text{Fe}^{3+}_{(\text{aq})}$  ions.

The weak base is only **partially dissociated** so less than 1% of the  $\text{OH}^-_{(\text{aq})}$  ions are available at the beginning.

However, as the  $\text{OH}^-_{(\text{aq})}$  ions react with  $\text{Fe}^{3+}_{(\text{aq})}$  they are *effectively removed* from the **equilibrium mixture**. This slows down or even stops the **reverse reaction**, but the **forward reaction continues**:



More  $\text{OH}^-_{(\text{aq})}$  ions will be produced which will then react with the  $\text{Fe}^{3+}_{(\text{aq})}$  ions and so on until every single  $\text{NH}_{(\text{aq})}$  molecule **dissociates** to form  $\text{OH}^-_{(\text{aq})}$  ions.

Overall, a weak base can supply *exactly the same number of*  $\text{OH}^-_{(\text{aq})}$  as an equal quantity of a strong base

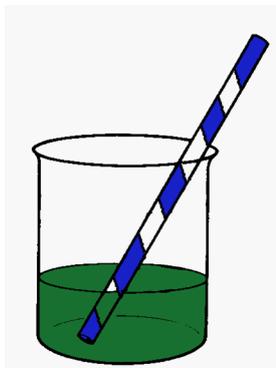
Strong and weak bases have **different properties**, but they have the **same stoichiometry**.

## Aqueous Solutions

This lesson looks at the effect on pH of dissolving oxides and salts in water

### Solutions of Covalent Oxides

This activity examines how it is possible for covalent oxides to affect the pH of water



When you **blow** CO<sub>2</sub> into **water** the pH will **drop** from pH = 7 to about pH = 5

This is because CO<sub>2</sub> **reacts** with **water** to produce the **weak acid** called **carbonic acid**. Most of the **carbonic acid** molecules produced remain as **covalent molecules** - only a few **dissociated ions** are produced - so the **position of equilibrium** is well over to the **left**.



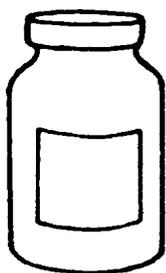
Similarly, SO<sub>2</sub> dissolves to produce the weak acid, **sulphurous acid**, (acid rain).



Other **covalent oxides** are either **insoluble** (CO, NO) or dissolve to form **strong acids** (NO<sub>2</sub> → *nitric acid*), (SO<sub>3</sub> → *sulphuric acid*), (P<sub>2</sub>O<sub>5</sub> → *phosphoric acid*)

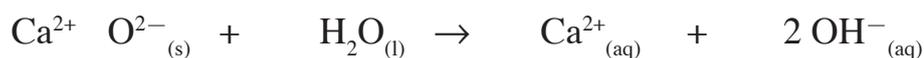
### Solutions of Ionic Oxides

This activity examines how soluble ionic oxides are able to affect the pH of water

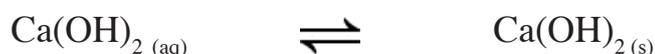


When you add CaO to **water** the pH will **rise** from pH = 7 to about pH = 10.

This is because CaO also **reacts** with **water** to form **calcium hydroxide**.



This reaction is **not reversible**, so **calcium hydroxide** is considered a **strong base**. However, Ca(OH)<sub>2</sub> is **not very soluble** so if too much is produced, some of it will **precipitate** out as **solid**:



Only the **Group I oxides/hydroxides** can be relied upon to be **soluble** at **high concentrations**. **Group II oxides/hydroxides** are only **soluble** at **low concentrations**. **Transition metal oxides/hydroxides** are **insoluble** (DATA BOOK)

## Solutions of Salts

This activity looks at the effect that dissolving salts can have on the pH of water



Salts are **ionic compounds** derived from an **acid** and a **base**. Though we describe the reaction between an **acid** and a **base** as **neutralisation**, it would be wrong to assume that the final **salt solution** is always **neutral**.

It depends on the **strength** of the parent **acid** and the parent **base**.

Salt	Formula	Parent Base & Strength	Parent Acid & Strength	pH and type of solution
sodium chloride	$\text{Na}^+\text{Cl}^-$	sodium hydroxide <b>STRONG</b>	hydrochloric acid <b>STRONG</b>	pH = 7 <b>NEUTRAL</b>
sodium sulphite	$(\text{Na}^+)_2\text{SO}_3^{2-}$	sodium hydroxide <b>STRONG</b>	sulphurous acid <b>WEAK</b>	pH = >7 <b>ALKALI</b>
ammonium nitrate	$\text{NH}_4^+\text{NO}_3^-$	ammonia <b>WEAK</b>	nitric acid <b>STRONG</b>	pH = <7 <b>ACID</b>
magnesium sulphate	$\text{Mg}^{2+}\text{SO}_4^{2-}$	magnesium hydroxide <b>WEAK</b>	sulphuric acid <b>STRONG</b>	pH = <7 <b>ACID</b>
sodium stearate	$\text{Na}^+\text{C}_{17}\text{H}_{35}\text{COO}^-$	sodium hydroxide <b>STRONG</b>	stearic acid <b>WEAK</b>	pH = >7 <b>ALKALI</b>
ammonium sulphate	$\text{NH}_4^+\text{SO}_4^{2-}$	ammonia <b>WEAK</b>	sulphuric acid <b>STRONG</b>	pH = <7 <b>ACID</b>

From the results above, 3 rules can be made:

- ① **ACIDIC** solutions when parent **ACID STRONG**, parent **BASE WEAK**
- ② **ALKALI** solutions when parent **ACID WEAK**, parent **BASE STRONG**
- ③ **NEUTRAL** solutions when parent **ACID STRONG**, parent **BASE STRONG**

These rules can be used to **PREDICT** the pH of a salt solution.

To **EXPLAIN** the pH of a salt solution is more complicated and these explanations come next.

## Equilibria in Salt Solutions

This topic examines how it is possible for salts to have an effect on the pH of water

### Sodium Ethanoate Solution

This activity explains why salts of strong bases and weak acids form alkaline solutions

Sodium ethanoate is the salt made from **sodium hydroxide** (**strong base**) and **ethanoic acid** (**weak acid**) and you would **predict** that it would produce an **alkaline** solution

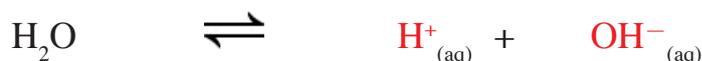
<b>STRONG</b>	<b>WEAK</b>	<b>predict</b>	pH > 7
<b>BASE</b>	<b>ACID</b>		

But how do we *explain* this?

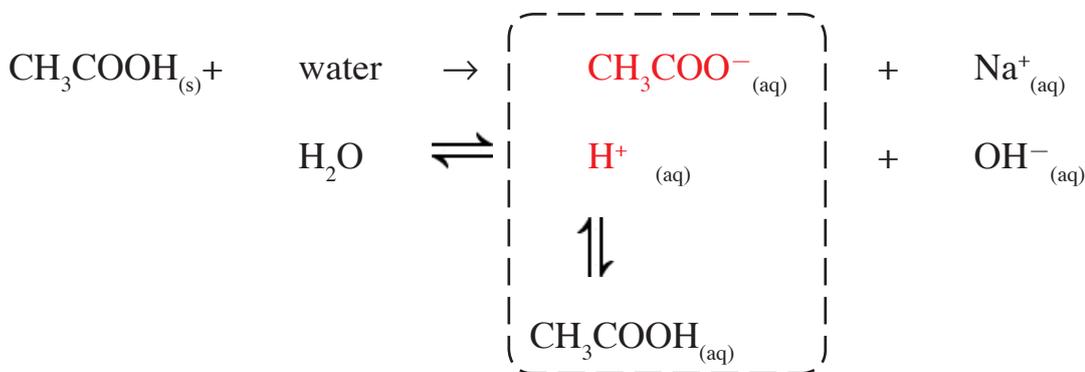
Sodium ethanoate is **soluble**, and **dissolves** to release **ions** into the **water**;



There are, however, already two **reactions** taking place in the water;



*Secondary reactions* between the salt ions and the water ions are possible;



**Ethanoate ions** ( $\text{CH}_3\text{COO}^-_{(aq)}$ ) will react with **hydrogen ions** ( $\text{H}^+_{(aq)}$ ) to produce **covalent molecules** of **ethanoic acid** ( $\text{CH}_3\text{COOH}_{(aq)}$ ). Being a **weak acid**, only a **small proportion** of these molecules will **dissociate** to reform  $\text{H}^+_{(aq)}$  ions.

As a result  $\text{H}^+_{(aq)}$  ion **concentrations** will **drop**. This will **slow down** the **backward** reaction so for a while;



The **extra**  $\text{H}^+_{(aq)}$  ions produced will be **mopped up** by the **ethanoate ions** ( $\text{CH}_3\text{COO}^-_{(aq)}$ ) but the **extra**  $\text{OH}^-_{(aq)}$  ions build up so,

$\text{OH}^- > \text{H}^+$  and pH > 7, alkaline solution

**Ammonium Chloride Solution**

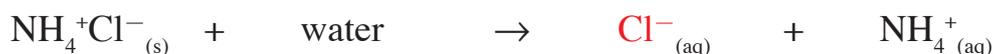
This activity explains why salts of weak bases and strong acids form acidic solutions

Ammonium chloride is the salt made from **ammonium hydroxide** (*weak base*) and **hydrochloric acid** (*strong acid*) and you would **predict** an **acidic** solution

**WEAK BASE**                      **STRONG ACID**                      *predict*                      pH < 7

But how do we *explain* this?

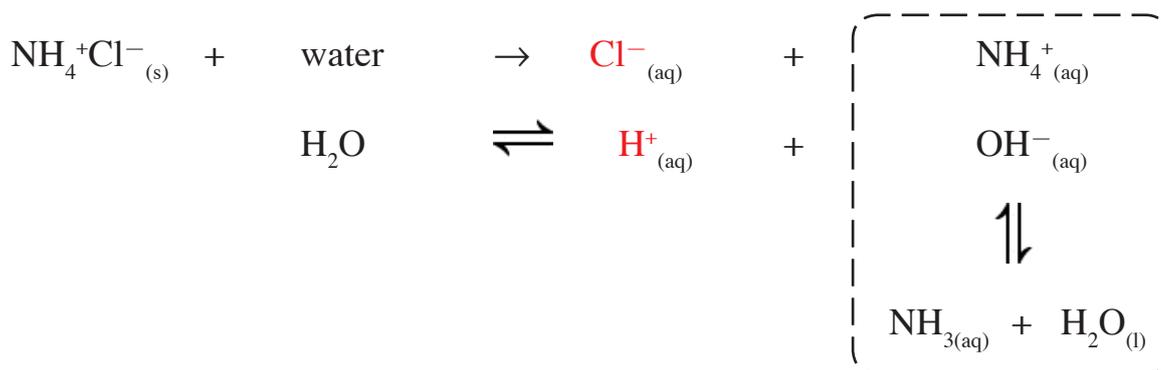
Ammonium chloride is **soluble**, and **dissolves** to release **ions** into the **water**;



There are, however, already two **reactions** taking place in the water;



*Secondary reactions* between the salt ions and the water ions are possible;



**Ammonium ions** ( $\text{NH}_4^+_{(aq)}$ ) will react with **hydroxide ions** ( $\text{OH}^-_{(aq)}$ ) to produce **covalent molecules** of **ammonia** ( $\text{NH}_3_{(aq)}$ ) and **water** ( $\text{H}_2\text{O}_{(l)}$ ). Being a **weak base**, only a **small proportion** of these molecules will **dissociate** to reform  $\text{OH}^-_{(aq)}$  ions.

As a result  $\text{OH}^-_{(aq)}$  ion **concentrations** will **drop**. This will **slow down** the **backward** reaction so for a while;



The **extra**  $\text{OH}^-_{(aq)}$  ions produced will be **mopped up** by the **ammonium ions** ( $\text{NH}_4^+_{(aq)}$ ) but the **extra**  $\text{H}^+_{(aq)}$  ions build up so,

$\text{H}^+ > \text{OH}^-$                       and pH < 7, acidic solution

## Sodium Chloride Solution

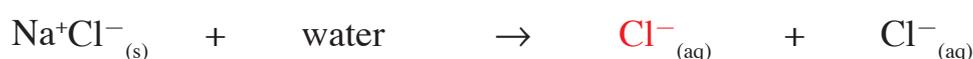
*This activity considers the changes taking place when the salt of a strong base and a strong acid is added to water*

Sodium chloride is the salt made from **sodium hydroxide** (**strong base**) and **hydrochloric acid** (**strong acid**) and you would **predict** that it would produce a **neutral** solution

**STRONG**                      **STRONG**                      *predict*                      pH = 7  
**BASE**                              **ACID**

But how do we *explain* this?

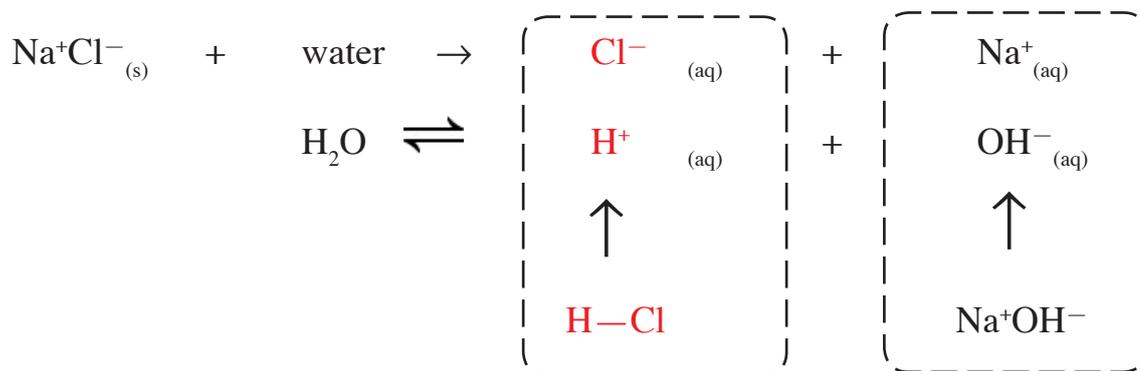
Sodium chloride is **soluble**, and **dissolves** to release **ions** into the **water**;



Again there are already two **reactions** taking place in the water;



*Secondary reactions*, however, between the salt ions and the water ions are **not** possible;



If  $\text{H}^+_{(aq)}$  ions were to react with  $\text{Cl}^-_{(aq)}$  they would form molecules of  $\text{HCl}_{(aq)}$  but, being a strong acid, these molecules would dissociate 100% so all the  $\text{H}^+_{(aq)}$  ions would return.

If  $\text{OH}^-_{(aq)}$  ions were to react with  $\text{Na}^+_{(aq)}$  they would form  $\text{NaOH}$  but, being a strong base, would dissociate 100% so all the  $\text{OH}^-_{(aq)}$  ions would return.

$\text{H}^+ = \text{OH}^-$                       and pH = 7, neutral solution

*When asked to **explain** the pH of salt solutions, you will be expected to produce the two equations involved and show which ions react.*

## Concentration & pH

This lesson considers the relationship between the concentration of ions in solution and the pH value

### Hydrogen Ions & pH

This activity investigates the numerical relationship between the hydrogen ion concentration and pH value



pH is a **measure** of the **concentration** of  $H^+_{(aq)}$  ions in a **solution**.

The 'p' in pH stands for 'power of' and 'H' is, of course, 'Hydrogen'

Square brackets, [ ], are used in Chemistry for 'concentration of', so  $[H^+_{(aq)}]$  = concentration of  $H^+_{(aq)}$  ions

Concentration of acid (in terms of molarity)	Concentration of $H^+_{(aq)}$ (as a power of 10)	pH (to nearest whole number)
0.1 M	$10^{-1} \text{ mol l}^{-1}$	1
0.01 M	$10^{-2} \text{ mol l}^{-1}$	2
0.001 M	$10^{-3} \text{ mol l}^{-1}$	3
0.0001 M	$10^{-4} \text{ mol l}^{-1}$	4
0.00001 M	$10^{-5} \text{ mol l}^{-1}$	5

From the table, it can be seen that the power of 10 index corresponds to the pH value

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

For example,

an acid of pH = 3	means $[H^+] = 10^{-3} \text{ mol l}^{-1}$
water of pH = 7	means $[H^+] = 10^{-7} \text{ mol l}^{-1}$
alkali of pH = 12	means $[H^+] = 10^{-12} \text{ mol l}^{-1}$

Notice that, even when we deal with alkalis, we continue to measure  $[H^+]$ .

acids	$[H^+] >$ in water	ie,	$[H^+] > 10^{-7} \text{ mol l}^{-1}$
alkalis	$[H^+] <$ in water	ie,	$[H^+] < 10^{-7} \text{ mol l}^{-1}$

As you've seen, there is a simple relationship between  $[H^+]$  and pH;

0.1 mol $l^{-1}$ HCl	=	$10^{-1}$ mol $l^{-1}$ $[H^+]$	=	pH of 1
1.0 mol $l^{-1}$ HCl	=	$10^0$ mol $l^{-1}$ $[H^+]$	=	pH of 0
10 mol $l^{-1}$ HCl	=	$10^1$ mol $l^{-1}$ $[H^+]$	=	pH of -1
100 mol $l^{-1}$ HCl	=	$10^2$ mol $l^{-1}$ $[H^+]$	=	pH of -2

In reality, there are no acids that are **soluble** enough to allow 100 moles of acid to dissolve in 1 litre of water. **Concentrated** HCl is about 12 mol  $l^{-1}$  while **concentrated**  $H_2SO_4$  is about 20 mol  $l^{-1}$ .

Similarly 1 mol  $l^{-1}$  and 10 mol  $l^{-1}$  **alkali** solutions are possible but **solubility** limits prevent **alkali** solutions of higher **concentrations**.

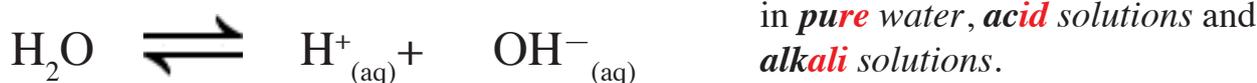
Therefore, our *pH scale* is really;

... -1 0 1 2 3 4 5 6 7 8 9 10 11 12 13 14 15 ...

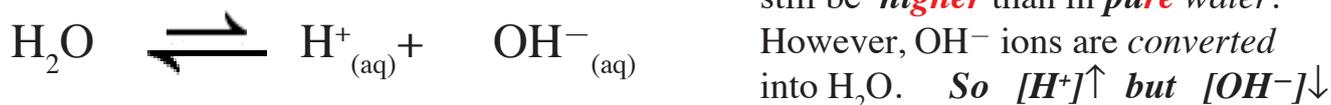
### Ionic Product for Water

*This activity considers the relationship between the concentrations of hydrogen ions and hydroxide ions in solutions of different pH values.*

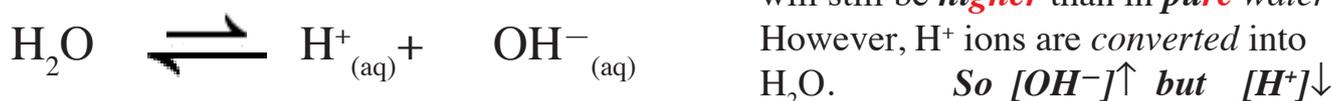
The water **equilibrium** is a **dynamic equilibrium**, both reactions taking place all the time;



If an **acid** is **dissolved** in water, the  $[H^+]$  will be **increased**. This will **speed up** the **backward** reaction. Some of the added  $H^+$  ions will be **converted** into  $H_2O$  but the  $[H^+]$  will still be **higher** than in **pure** water.



If a **base** is **dissolved** in water, the  $[OH^-]$  will be **increased**. This will **speed up** the **backward** reaction. Some of the added  $OH^-$  ions will be **converted** into  $H_2O$  but the  $[OH^-]$  will still be **higher** than in **pure** water.



Clearly, there is always going to be a **relationship** between  $[H^+]$  and  $[OH^-]$ .

In **pure** water the pH is 7, so  $[H^+] = 10^{-7}$  mol  $l^{-1}$  and  $[OH^-] = 10^{-7}$  mol  $l^{-1}$

Multiplying the two **concentrations** gives what is called the **Ionic Product** for water

$$\text{ionic product} = [H^+] \times [OH^-] = 10^{-7} \text{ mol } l^{-1} \times 10^{-7} \text{ mol } l^{-1} = 10^{-14} \text{ (units?)}$$

$[NaOH]$	$pH$ <i>measured</i>	$[H^+]$ <i>calculated</i>	$[OH^-]$	$[H^+] \times [OH^-]$ <i>ionic product</i>
0.0001 M	10	$10^{-10} \text{ mol l}^{-1}$	$10^{-4} \text{ mol l}^{-1}$	$10^{-14} \text{ mol}^2 \text{ l}^{-2}$
0.01 M	12	$10^{-12} \text{ mol l}^{-1}$	$10^{-2} \text{ mol l}^{-1}$	$10^{-14} \text{ mol}^2 \text{ l}^{-2}$
1.0 M	14	$10^{-14} \text{ mol l}^{-1}$	$10^0 \text{ mol l}^{-1}$	$10^{-14} \text{ mol}^2 \text{ l}^{-2}$

From the table above, it can be seen that the ionic product is a constant; its value remains unchanged in a solution of any pH.

$$\text{ionic product} = [H^+] \times [OH^-] = 10^{-14} \text{ mol}^2 \text{ l}^{-2}$$

### Calculations Involving the Ionic Product

*This activity uses the ionic product of water to calculate the concentrations of hydrogen or hydroxide ions in solutions*

**Example 1.** In a solution of pH 9 what is the concentration of  $OH^-_{(aq)}$ ?

**Example 2.** If  $[OH^-_{(aq)}]$  is  $0.01 \text{ mol l}^{-1}$  what is  
and

- the concentration of hydrogen ions
- the pH value?

**Example 3.** If a very concentrated solution of NaOH contains 40g in 100 cm<sup>3</sup> of solution, calculate

- a)  $[OH^-_{(aq)}]$
- b)  $[H^+_{(aq)}]$
- and c) the pH

**Example 4.** Very concentrated hydrochloric acid can contain 1 mol in 100 cm<sup>3</sup> solution.

- a) What is the pH value?
- b) Calculate the hydroxide ion concentration.

## Acids & Bases

### Acids and Bases

6. The **pH scale** is a continuous range from **below 0 to above 14**.

0.1 M HCl = pH 1, 1M = pH 0, 10M = pH -1  
0.1 M NaOH = pH 13, 1M = pH 14, 10M = pH 15

7. **Integral** (whole number) values from pH 0 to 14 can be related to concentrations of  $H^+_{(aq)}$  in mol / l.

0.1 M  $H^+_{(aq)}$  =  $10^{-1}$  = pH 1  
0.01 M  $H^+_{(aq)}$  =  $10^{-2}$  = pH 2  
0.001 M  $H^+_{(aq)}$  =  $10^{-3}$  = pH 3

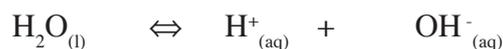
8. In water and aqueous solutions with a pH value of 7, the concentrations of  $H^+_{(aq)}$  and  $OH^-_{(aq)}$  are both  $10^{-7}$  mol / l at 25°C.

0.0000001 M  $H^+_{(aq)}$  =  $10^{-7}$  = pH 7

9. The concentration of  $H^+_{(aq)}$  and  $OH^-_{(aq)}$  can be calculated from the concentration of the other by using  $[H^+_{(aq)}][OH^-_{(aq)}] = 10^{-14}$  mol / l.

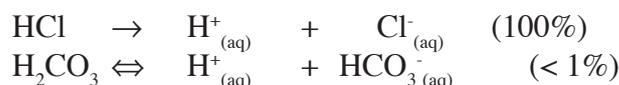
$[H^+_{(aq)}]$  = 'concentration of'  $H^+_{(aq)}$   
pH = 12; so  $[H^+_{(aq)}] = 10^{-12}$   
 $10^{-12} \times [OH^-_{(aq)}] = 10^{-14}$  so  $[OH^-_{(aq)}] = 10^{-2}$

10. In water and aqueous solutions there is an equilibrium between  $H^+_{(aq)}$  and  $OH^-_{(aq)}$  and water molecules.



### Strong and Weak

11. In aqueous solutions, **strong acids are completely dissociated** but **weak acids are only partially dissociated**.



12. Equimolar solutions of weak and strong acids differ in pH, conductivity and reaction rates but not in stoichiometry (amount) of reactions.

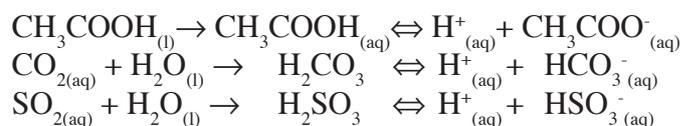
**pH** -  
0.1 M HCl, pH = 1 ( $10^{-1} H^+_{(aq)}$ )  
0.1 M  $CH_3COOH$ , pH = 4 ( $10^{-4} H^+_{(aq)}$ )

**conductivity** -  
0.1 M HCl, 100mA ( $10^{-1} H^+_{(aq)}$ )  
0.1 M  $CH_3COOH$ , 15mA ( $10^{-4} H^+_{(aq)}$ )

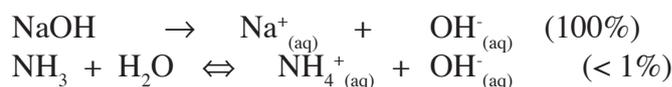
**rates** -  
0.1 M HCl + Mg, fast ( $10^{-1} H^+_{(aq)}$ )  
0.1 M  $CH_3COOH$  + Mg, slow ( $10^{-4} H^+_{(aq)}$ )

**stoichiometry** (amounts) -  
25cm<sup>3</sup> 0.1 M HCl + 25cm<sup>3</sup> 0.1 M NaOH  
25cm<sup>3</sup> 0.1 M  $CH_3COOH$  + 25cm<sup>3</sup> 0.1 M NaOH

13. The weakly acidic nature of **ethanoic acid, sulphur dioxide and carbon dioxide** can be explained by reference to equations showing the equilibrium.

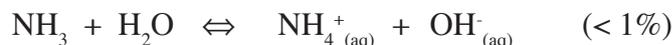
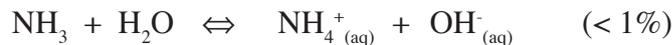


14. In aqueous solutions, strong bases are **completely dissociated** but **weak bases are only partially dissociated**.



15. Equimolar solutions of weak and strong bases differ in pH, conductivity but not in stoichiometry (amount) of reactions.

16. The weakly alkaline nature of **ammonia**, can be explained by reference to an equation showing the equilibrium.



### pH of salt solutions

17. A soluble salt of a **strong acid** and a **strong base** dissolves in water to produce a **neutral** solution.

18. A soluble salt of a **weak acid** and a **strong base** dissolves in water to produce an **alkaline** solution

19. A soluble salt of a **strong acid** and a **weak base** dissolves in water to produce an **acidic** solution

20. **Soaps** are salts of weak acids and strong bases.

21. The acidity, alkalinity or neutrality of the above kinds of salt can be **explained** by reference to the appropriate **equilibria** (plural !).

Salt	Acid	Base	pH
eg NaCl	strong	strong	= 7
NaCH <sub>3</sub> COO	weak	strong	> 7
NH <sub>4</sub> Cl	strong	weak	< 7
NH <sub>4</sub> CH <sub>3</sub> COO	weak	weak	?

