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 H₂CO₃ 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 \( H^+_{(aq)} \) and \( OH^-_{(aq)} \) ions. This reaction is, however, reversible so an equilibrium mixture exists.

\[
\begin{align*}
\text{H}_2\text{O} & \rightleftharpoons H^+_{(aq)} + OH^-_{(aq)} \\
\text{covalent} & \quad \text{ionic}
\end{align*}
\]

99.9999998 % 0.0000002 %

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

\[
\begin{align*}
H^+_{(aq)} & > OH^-_{(aq)} \quad \text{an acidic solution is the result} \\
OH^-_{(aq)} & > H^+_{(aq)} \quad \text{an alkaline solution is the result}
\end{align*}
\]

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.

\[
\text{HCl}_{(g)} + \text{water} \rightarrow \text{HCl}_{(aq)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}
\]

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

A strong acid is one that is completely dissociated into ions in solution.
Weak Acids & Bases Primer

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.

\[
\text{CH}_3\text{COOH}(l) + \text{water} \rightarrow \text{CH}_3\text{COOH}(aq) \rightleftharpoons \text{H}^+(aq) + \text{CH}_3\text{COO}^-(aq)
\]

ethanoate ions

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,

CO₂ is a soluble gas that (like other non-metal oxides) will dissolve/react with water to produce an acidic solution.

Solutions of CO₂ (think of fizzy drinks) lose CO₂ quickly when open to the air i.e. even the dissolving of CO₂ is a reversible reaction.

\[
\text{CO}_2(g) \rightleftharpoons \text{CO}_2(aq) + \text{H}_2\text{O} \rightleftharpoons \text{H}_2\text{CO}_3(aq) \rightleftharpoons \text{H}^+(aq) + \text{HCO}_3^-(aq)
\]

carbonic acid (covalent)
dissociated ions

Similarly for SO₂:

\[
\text{SO}_2(g) \rightleftharpoons \text{SO}_2(aq) + \text{H}_2\text{O} \rightleftharpoons \text{H}_2\text{SO}_3(aq) \rightleftharpoons \text{H}^+(aq) + \text{HSO}_3^-(aq)
\]
sulphurous acid (covalent)
dissociated ions

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

\[
\text{CH}_3\text{COOH}(l) + \text{water} \rightarrow \text{CH}_3\text{COOH}(aq) \rightleftharpoons \text{H}^+(aq) + \text{CH}_3\text{COO}^-(aq)
\]
propanoate ions

\[
\text{CH}_3\text{COOH}(l) + \text{water} \rightarrow \text{H}_3\text{COOH}(aq) \rightleftharpoons \text{H}^+(aq) + \text{H}_3\text{COO}^-(aq)
\]
butanoate ions
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).

<table>
<thead>
<tr>
<th>Comparison</th>
<th>0.1 M HCl</th>
<th>0.1 M CH₃COOH</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>pH</strong></td>
<td></td>
<td></td>
</tr>
<tr>
<td><strong>Conductivity</strong></td>
<td></td>
<td></td>
</tr>
<tr>
<td><strong>Reaction Rate</strong></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**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⁺** - **very low pH**

Less than 1% of the CH₃COOH molecules **dissociated** so there were **much fewer** H⁺ 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⁺** - **very high conductivity**

Less than 1% of the CH₃COOH molecules **dissociated** so there were **much fewer** H⁺ 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⁺** - **very fast reaction**

Less than 1% of the CH₃COOH molecules **dissociated** so there were **much fewer** H⁺ ions present - **fewer collisions** with magnesium ribbon so rate of **reaction is slower**.

**Conclusion**
Weak acids produce less H⁺ 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 cm$^3$ 0.1M HCl = cm$^3$

Vol. needed to neutralise
10 cm$^3$ 0.1M CH$_3$COOH = 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 H$^+$ ions were available to react with the OH$^-$ ions added from the burette.

The weak acid is only partially dissociated so less than 1% of the H$^+$ ions are available at the beginning.

Weak acid means reversible reactions; CH$_3$COOH$\rightleftharpoons$ H$^+$(aq) + CH$_3$COO$^-$ eq

However, as the H$^+$ ions react with OH$^-$ they are effectively removed from the equilibrium mixture. This slows down or even stops the reverse reaction, but the forward reaction continues:

CH$_3$COOH$\rightarrow$ H$^+$(aq) + CH$_3$COO$^-$ eq

More H$^+$ ions will be produced which will then react with the OH$^-$ ions and so on until every single CH$_3$COOH molecule dissociates to form H$^+$ ions.

Overall, a weak acid can supply exactly the same number of H$^+$ as an equal quantity of a strong acid.

HCl$\rightleftharpoons$ NaOH $\rightarrow$ H$_2$O(l) + NaCl

CH$_3$COOH + NaOH $\rightarrow$ H$_2$O(l) + NaCH$_3$COO

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)

\[ \text{Na}^+ \text{OH}^- (\text{s}) + \text{water} \rightarrow \text{Na}^+(\text{aq}) + \text{OH}^- (\text{aq}) \]

When dissolved in water, the bases dissociate (split up) into separate ions. With strong bases this conversion is 100% and a single arrow \( \rightarrow \) 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 1** 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.

\[ \text{NH}_3 (\text{g}) + \text{water} \rightarrow \text{NH}_3(\text{aq}) + \text{H}_2\text{O}(\text{l}) \Leftrightarrow \text{NH}_4^+(\text{aq}) + \text{OH}^- (\text{aq}) \]

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

\[ \text{CH}_3\text{NH}_2(l) + \text{water} \rightarrow \text{CH}_3\text{NH}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) \Leftrightarrow \text{CH}_3\text{NH}_3^+(\text{aq}) + \text{OH}^- (\text{aq}) \]
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

<table>
<thead>
<tr>
<th>Comparison</th>
<th>0.1 M NaOH</th>
<th>0.1 M NH₃</th>
</tr>
</thead>
<tbody>
<tr>
<td>pH</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Conductivity</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**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⁻ - *very high* pH

Less than 1% of the NH₃ molecules *dissociated* so there were much fewer OH⁻ 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⁻ - *high conductivity*

Less than 1% of the NH₃ molecules *dissociated* so there were much fewer OH⁻ ions present - conducts, but *not as good as NaOH*.

**Conclusion**

Weak bases produce less OH⁻ 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:

\[
\text{Fe}^{3+}(\text{NO}_3^-)_3 \text{ (aq)} + 3 \text{Na}^+\text{OH}^- \text{ (aq)} \rightarrow \text{Fe}^{3+}(\text{OH}^-)_3 \text{ (aq)} + 3 \text{Na}^+\text{NO}_3^- \text{ (aq)}
\]

\[
\text{Fe}^{3+}(\text{NO}_3^-)_3 \text{ (aq)} + 3 \text{NH}_4^+\text{OH}^- \text{ (aq)} \rightarrow \text{Fe}^{3+}(\text{OH}^-)_3 \text{ (aq)} + 3 \text{Na}^+\text{NO}_3^- \text{ (aq)}
\]
The only difference between these two reactions is that the NaOH has 100% of its 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 OH$^{\text{aq}}$ ions were available to react with the Fe$^{3+\text{aq}}$ ions.

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

However, as the OH$^{\text{aq}}$ ions react with 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:

$$\text{NH}^{\text{aq}} + \text{H}_2\text{O} (l) \rightarrow \text{NH}_4^{\text{aq}} + \text{OH}^{-\text{aq}}$$

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

Overall, a weak base can supply exactly the same number of 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** CO₂ into **water** the pH will **drop** from pH = 7 to about pH = 5

This is because CO₂ **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**.

\[
\begin{align*}
\text{CO}_2(\text{g}) & \rightleftharpoons \text{CO}_2(\text{aq}) + \text{H}_2\text{O} \\
& \rightleftharpoons \text{H}_2\text{CO}_3(\text{aq}) \rightleftharpoons 2 \text{H}^+ (\text{aq}) + \text{CO}_3^{2-} (\text{aq})
\end{align*}
\]

Similarly, SO₂ dissolves to produce the weak acid, **sulphurous acid**, (acid rain).

\[
\begin{align*}
\text{SO}_2(\text{g}) & \rightleftharpoons \text{SO}_2(\text{aq}) + \text{H}_2\text{O} \\
& \rightleftharpoons \text{H}_2\text{SO}_3(\text{aq}) \rightleftharpoons 2 \text{H}^+ (\text{aq}) + \text{SO}_3^{2-} (\text{aq})
\end{align*}
\]

Other **covalent oxides** are either **insoluble** (CO, NO) or dissolve to form **strong acids** (NO₂ → nitric acid), (SO₃ → sulphuric acid), (P₂O₅ → 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 from **calcium hydroxide**.

\[
\text{Ca}^{2+} + \text{O}^{2-} (\text{s}) + \text{H}_2\text{O} (\text{l}) \rightarrow \text{Ca}^{2+} (\text{aq}) + 2 \text{OH}^- (\text{aq})
\]

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

\[
\text{Ca(OH)}_2 (\text{aq}) \rightleftharpoons \text{Ca(OH)}_2 (\text{s})
\]

Only the **Group 1 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.

<table>
<thead>
<tr>
<th>Salt</th>
<th>Formula</th>
<th>Parent Base &amp; Strength</th>
<th>Parent Acid &amp; Strength</th>
<th>pH and type of solution</th>
</tr>
</thead>
<tbody>
<tr>
<td>sodium chloride</td>
<td>Na⁺Cl⁻</td>
<td>sodium hydroxide</td>
<td>hydrochloric acid</td>
<td>pH = 7 NEUTRAL</td>
</tr>
<tr>
<td>sodium sulphite</td>
<td>(Na⁺)₂SO₃²⁻</td>
<td>sodium hydroxide</td>
<td>sulphurous acid</td>
<td>pH = &gt;7 ALKALI</td>
</tr>
<tr>
<td>ammonium nitrate</td>
<td>NH⁴⁺NO₃⁻</td>
<td>ammonia</td>
<td>nitric acid</td>
<td>pH = &lt;7 ACID</td>
</tr>
<tr>
<td>magnesium sulphate</td>
<td>Mg²⁺SO₄²⁻</td>
<td>magnesium hydroxide</td>
<td>sulphuric acid</td>
<td>pH = &lt;7 ACID</td>
</tr>
<tr>
<td>sodium stearate</td>
<td>Na⁺C₁₇H₃₅COO⁻</td>
<td>sodium hydroxide</td>
<td>stearic acid</td>
<td>pH = &gt;7 ALKALI</td>
</tr>
<tr>
<td>ammonium sulphate</td>
<td>NH⁴⁺SO₄²⁻</td>
<td>ammonia</td>
<td>sulphuric acid</td>
<td>pH = &lt;7 ACID</td>
</tr>
</tbody>
</table>

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.

\[
\begin{array}{c|c|c}
\text{STRONG } & \text{WEAK } & \text{predict} \\
\text{BASE } & \text{ACID } & \text{pH } > 7
\end{array}
\]

But how do we explain this?

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

\[
\text{CH}_3\text{COONa}_{(s)} + \text{water} \rightarrow \text{CH}_3\text{COO}^-_{(aq)} + \text{Na}^+_{(aq)}
\]

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

\[
\text{H}_2\text{O} \rightleftharpoons \text{H}^+_{(aq)} + \text{OH}^-_{(aq)}
\]

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

\[
\begin{align*}
\text{CH}_3\text{COOH} \_{(s)} & \text{water} \rightarrow \{ \text{CH}_3\text{COO}^-_{(aq)} + \text{Na}^+_{(aq)} \\
\text{H}_2\text{O} & \rightleftharpoons \text{H}^+_{(aq)} + \text{OH}^-_{(aq)}
\end{align*}
\]

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

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

\[
\text{H}_2\text{O} \rightleftharpoons \text{H}^+_{(aq)} + \text{OH}^-_{(aq)}
\]

The extra H\(^+\)\_(aq) ions produced will be mopped up by the ethanoate ions (CH\(_3\)COO\(^-\)\_(aq)) but the extra 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:

<table>
<thead>
<tr>
<th><strong>WEAK BASE</strong></th>
<th><strong>STRONG ACID</strong></th>
<th><strong>predict</strong></th>
<th><strong>pH</strong> $&lt; 7$</th>
</tr>
</thead>
</table>

But how do we explain this?

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

$$\text{NH}_4^+\text{Cl}^- (s) + \text{water} \rightarrow \text{Cl}^- (aq) + \text{NH}_4^+ (aq)$$

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

$$\text{H}_2\text{O} \quad \rightleftharpoons \quad \text{H}^+ (aq) + \text{OH}^- (aq)$$

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

$$\text{NH}_4^+\text{Cl}^- (s) + \text{water} \rightarrow \text{Cl}^- (aq) + \text{NH}_4^+ (aq) + \text{OH}^- (aq)$$

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;

$$\text{H}_2\text{O} \quad \rightleftharpoons \quad \text{H}^+ (aq) + \text{OH}^- (aq)$$

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,

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

\[
\text{STRONG BASE} \quad \text{STRONG ACID} \quad \text{predict} \quad \text{pH} = 7
\]

But how do we explain this?

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

\[
\text{Na}^+\text{Cl}^- (s) + \text{water} \rightarrow \text{Cl}^- (aq) + \text{Cl}^- (aq)
\]

Again there are already two reactions taking place in the water:

\[
\text{H}_2\text{O} \quad \leftrightarrow \quad \text{H}^+ (aq) + \text{OH}^- (aq)
\]

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

\[
\text{Na}^+\text{Cl}^- (s) + \text{water} \rightarrow \begin{cases} \text{Cl}^- (aq) + \text{Cl}^- (aq) \\ \text{H}^+ (aq) \uparrow \\downarrow \text{H}^-\text{Cl} \\ \text{Na}^+\text{OH}^- (aq) \end{cases}
\]

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.

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

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 $H^{+\text{(aq)}}$ 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 $[H^{+\text{(aq)}}] = \text{concentration of } H^{+\text{(aq)}}$ ions*

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

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

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

- acids $[H^+] > \text{in water}$ ie, $[H^+] > 10^{-7}$ mol l$^{-1}$
- alkalis $[H^+] < \text{in water}$ ie, $[H^+] < 10^{-7}$ mol l$^{-1}$
As you’ve seen, there is a simple relationship between \([H^+]\) and pH:

\[
\begin{align*}
0.1 \text{ mol } \ell^{-1} \text{ HCl} &= 10^{-1} \text{ mol } \ell^{-1} [H^+] = \text{pH of 1} \\
1.0 \text{ mol } \ell^{-1} \text{ HCl} &= 10^0 \text{ mol } \ell^{-1} [H^+] = \text{pH of 0} \\
10 \text{ mol } \ell^{-1} \text{ HCl} &= 10^1 \text{ mol } \ell^{-1} [H^+] = \text{pH of -1} \\
100 \text{ mol } \ell^{-1} \text{ HCl} &= 10^2 \text{ mol } \ell^{-1} [H^+] = \text{pH of -2}
\end{align*}
\]

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 \(\ell^{-1}\) while concentrated H\(_2\)SO\(_4\) is about 20 mol \(\ell^{-1}\).

Similarly 1 mol \(\ell^{-1}\) and 10 mol \(\ell^{-1}\) alkali solutions are possible but solubility limits prevent alkali solutions of higher concentrations.

Therefore, our pH scale is really:

\[
\ldots -1 \; 0 \; 1 \; 2 \; 3 \; 4 \; 5 \; 6 \; 7 \; 8 \; 9 \; 10 \; 11 \; 12 \; 13 \; 14 \; 15 \; \ldots
\]

**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;

\[
\text{H}_2\text{O} \rightleftharpoons [H^+] + [OH^-]
\]

In pure water, acid solutions and alkali solutions.

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 \(\text{H}_2\text{O}\) but the [H\(^+\)] will still be higher than in pure water. However, OH\(^-\) ions are converted into \(\text{H}_2\text{O}\). So [H\(^+\)]↑ but [OH\(^-\)]↓

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 \(\text{H}_2\text{O}\) but the [OH\(^-\)] will still be higher than in pure water. However, H\(^+\) ions are converted into \(\text{H}_2\text{O}\). So [OH\(^-\)]↑ but [H\(^+\)]↓

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

In pure water the pH is 7, so \([H^+] = 10^{-7} \text{ mol } \ell^{-1}\) and \([OH^-] = 10^{-7} \text{ mol } \ell^{-1}\)

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

\[
\text{ionic product} = [H^+] \times [OH^-] = 10^{-7} \text{ mol } \ell^{-1} \times 10^{-7} \text{ mol } \ell^{-1} = 10^{-14} \text{ (units?)}
\]
<table>
<thead>
<tr>
<th></th>
<th></th>
<th></th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>0.0001 M</td>
<td>10</td>
<td>10⁻¹⁰ mol l⁻¹</td>
<td>10⁻⁴ mol l⁻¹</td>
<td>10⁻¹⁴ mol² l⁻²</td>
</tr>
<tr>
<td>0.01 M</td>
<td>12</td>
<td>10⁻¹² mol l⁻¹</td>
<td>10⁻² mol l⁻¹</td>
<td>10⁻¹⁴ mol² l⁻²</td>
</tr>
<tr>
<td>1.0 M</td>
<td>14</td>
<td>10⁻¹⁴ mol l⁻¹</td>
<td>10⁰ mol l⁻¹</td>
<td>10⁻¹⁴ mol² l⁻²</td>
</tr>
</tbody>
</table>

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.

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

**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 mol l⁻¹ what is a) the concentration of hydrogen ions and b) the pH value?
Example 3. If a very concentrated solution of NaOH contains 40g in 100 cm$^3$ of solution, calculate:

a) $[\text{OH}^-_{(aq)}]$

b) $[\text{H}^+_{(aq)}]$

and c) the pH

Example 4. Very concentrated hydrochloric acid can contain 1 mol in 100 cm$^3$ solution.

a) What is the pH value?

b) Calculate the hydroxide ion concentration.
Acids and Bases

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

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

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

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

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

0.1 M HCl = pH 1, 1M = pH 0, 10M = pH -1
0.1 M NaOH = pH 13, 1M = pH 14, 10M = pH 15
0.1 M $H^+_{\text{(aq)}}$ = $10^{-1}$ = pH 1
0.01 M $H^+_{\text{(aq)}}$ = $10^{-2}$ = pH 2
0.001 M $H^+_{\text{(aq)}}$ = $10^{-3}$ = pH 3

$[H^+_{\text{(aq)}}] = \text{‘ concentration of ’ } [H^+_{\text{(aq)}}]$

$\text{pH} = 12; \text{ so } [H^+_{\text{(aq)}}] = 10^{-12}$

$10^{-12} \times [OH^-_{\text{(aq)}}] = 10^{-14}$ so $[OH^-_{\text{(aq)}}] = 10^{-2}$

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.

\[
\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ (\text{aq}) + \text{OH}^- (\text{aq}) \quad (<1\%) \\
\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ (\text{aq}) + \text{OH}^- (\text{aq}) \quad (<1\%)
\]

pH of salt solutions

17. A soluble salt of a strong acid and a strong base differs in pH, conductivity but not in stoichiometry (amount) of reactions.

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

<table>
<thead>
<tr>
<th>Salt</th>
<th>Acid</th>
<th>Base</th>
<th>pH</th>
</tr>
</thead>
<tbody>
<tr>
<td>eg NaCl</td>
<td>strong</td>
<td>strong</td>
<td>7</td>
</tr>
<tr>
<td>NaCH₃COO</td>
<td>weak</td>
<td>strong</td>
<td>&gt; 7</td>
</tr>
<tr>
<td>NH₄Cl</td>
<td>strong</td>
<td>weak</td>
<td>&lt; 7</td>
</tr>
<tr>
<td>NH₄CH₃COO</td>
<td>weak</td>
<td>weak</td>
<td>?</td>
</tr>
</tbody>
</table>

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!).

\[
\text{H}_2\text{O} (\text{l}) \rightleftharpoons \text{H}^+ (\text{aq}) + \text{OH}^- (\text{aq}) \\
\text{Na}_2\text{SO}_3 \rightarrow \text{Na}^+ (\text{aq}) + \text{SO}_3^{2-} (\text{aq}) \\
\text{Na}_2\text{SO}_3 \rightarrow \text{Na}^+ (\text{aq}) + \text{SO}_3^{2-} (\text{aq}) \\
\text{HSO}_3^- (\text{aq}) \rightarrow \text{H}^+ (\text{aq}) + \text{SO}_3^{2-} (\text{aq}) \\
\text{H}_2\text{SO}_3 \rightarrow \text{H}^+ (\text{aq}) + \text{SO}_3^{2-} (\text{aq})
\]

a weak acid so will not fully ionise; H⁺ ions removed