# Edexcel IAL Chemistry A-Level

## Topic 14 - Acid-Base Equilibria

Flashcards





# What is a Brønsted-Lowry acid?

# What is a Brønsted-Lowry acid?

A proton donor.

# What is a Brønsted-Lowry base?

# What is a Brønsted-Lowry base?

A proton acceptor.

# What do acid-base reactions involve the transfer of?

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Protons ($H^+$ ions).

What is the proton donor and proton acceptor in the reaction below?

$$HNO_{2(s)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + NO_2^-{}_{(aq)}$$

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$HNO_{2(s)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + NO_2^-{}_{(aq)}$

Proton donor: $HNO_{2(aq)}$

Proton acceptor: $H_2O_{(l)}$

# What is a conjugate acid-base pair?

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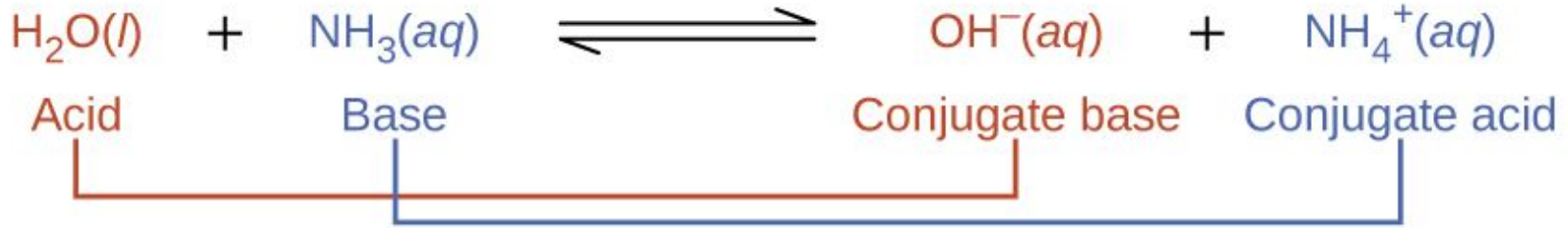
A conjugate acid-base pair contains two species that can be easily converted by transferring a proton.



# What is pH?

# What is pH?

A value which expresses the acidity or alkalinity of a substance, on a scale where 7 is neutral.

It is a way of expressing hydrogen ion concentration.

# How do you calculate the pH of a strong acid?

# How do you calculate the pH of a strong acid?

$HA \rightarrow H^+ + A^-$

For a strong acid the concentration of acid = concentration of $H^+$ ions, $[HA] = [H^+]$.

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

What expression can be used to calculate the hydrogen ion concentration of a strong acid, given its pH?

What expression can be used to calculate the hydrogen ion concentration of a strong acid, given its pH?

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

# What is a strong acid?

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A strong acid is an acid that completely dissociates in solution to form hydrogen ions:

$$HCl_{(aq)} \rightarrow H^{+}_{(aq)} + Cl^{-}_{(aq)}$$

# What is a strong base?

# What is a strong base?

A strong base is a base that completely dissociates in solution to form hydroxide ions:

$$NaOH_{(aq)} \rightarrow Na^{+}_{(aq)} + OH^{-}_{(aq)}$$

# What is a weak acid?

# What is a weak acid?

A weak acid is an acid that only partially dissociates in solution:

$$CH_3COOH_{(aq)} \rightleftharpoons CH_3COO^-_{(aq)} + H^+_{(aq)}$$

# What is a weak base?

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A weak base is a base that only partially dissociates in solution:

$$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$$

What is the acid dissociation constant, $K_a$?

# What is the acid dissociation constant, $K_a$?

A quantitative measure of the strength of an acid.

$$K_a = \frac{[H^+][A^-]}{[HA]}$$

# How do you calculate the pH of a weak acid?

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$HA \rightleftharpoons H^+ + A^-$

Write the $K_a$ expression: $$K_a = \frac{[H^+][A^-]}{[HA]}$$

For a weak acid we can assume $[H^+] = [A^-]$ so: $$K_a = \frac{[H^+]^2}{[HA]}$$

Rearrange to make $[H^+]$ the subject: $$[H^+] = \sqrt{K_a \times [HA]}$$

Calculate $[H^+]$, then use $pH = -\log[H^+]$ to find pH.

# What assumption is made when calculating the pH of a weak acid?

What assumption is made when calculating the pH of a weak acid?

The concentration of weak acid at equilibrium is equal to the original concentration of acid.

# What is $pK_a$?

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$pK_a = -\log(K_a)$

$pK_a$ values are much more manageable than $K_a$ values and make it easier to compare relative acidic strengths.

# What is $K_w$?

# What is $K_w$?

$K_w$ is the ionic product of water. Water molecules can function as acids and bases:

$$H_2O \rightleftharpoons H^+ + OH^-$$

$K_w = [H^+][OH^-] = 1 \times 10^{-14}\ mol^2\ dm^{-6}$ at RTP

Only a very small amount of water is ionised at any moment so the concentration remains effectively constant. Therefore, $H_2O$ doesn't appear in the $K_w$ expression.

# What is $pK_w$?

# What is $pK_w$?

$$pK_w = -\log[K_w]$$

At RTP:

$$pK_w = -\log[1 \times 10^{-14}]$$

$$= 14$$

# How do you calculate the pH of a strong base?

# How do you calculate the pH of a strong base?

$$XOH \rightarrow X^+ + OH^-$$

For a strong base: concentration of base = concentration of $OH^-$ ions, $[XOH] = [OH^-]$.

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

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

# How can you experimentally generate a pH curve?

# How can you experimentally generate a pH curve?

- Add a volume of acid to a conical flask, using a pipette.
- Measure the initial pH with a pH probe.
- Add small volumes of base from a burette into the conical flask, measuring the pH each time.
- Continue adding the base until the pH no longer changes.
- Plot a graph of pH (y-axis) against the volume of base added (x-axis).

# What does a pH curve look like for a strong acid-strong base reaction?

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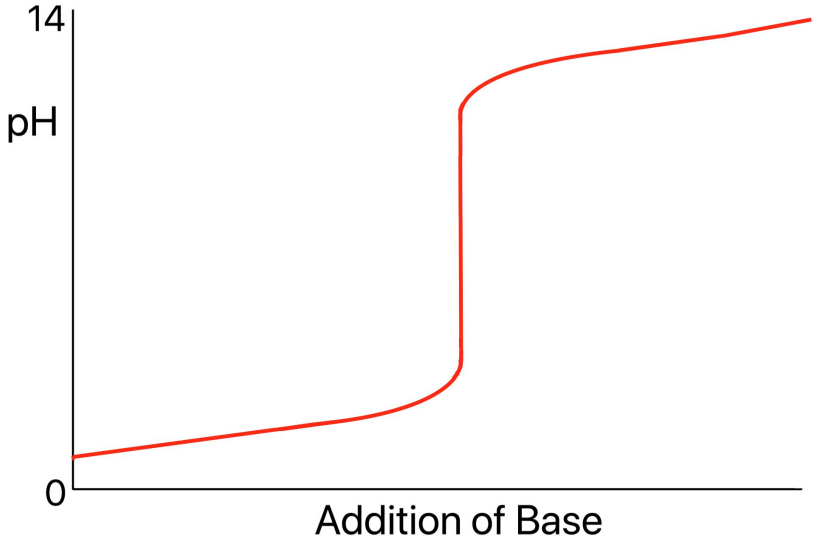

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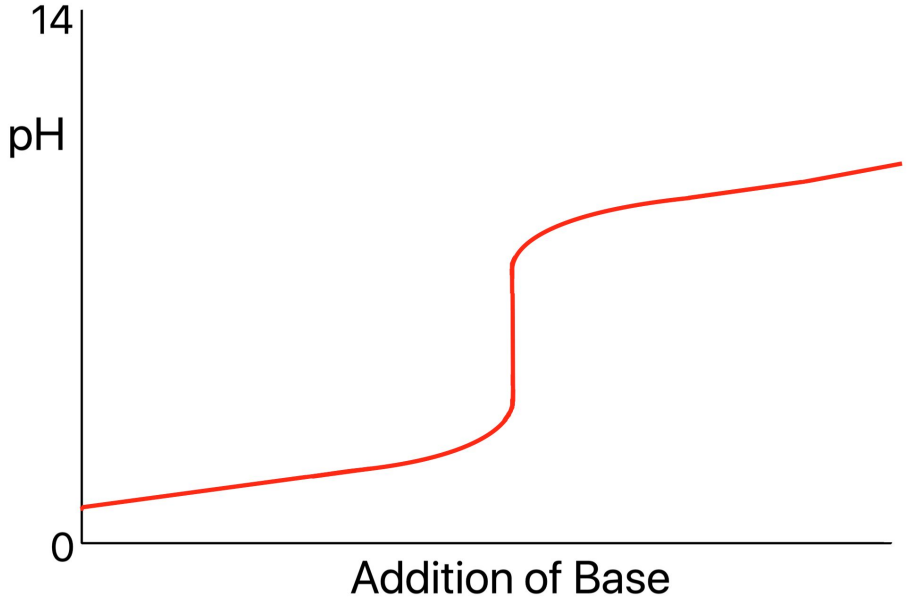

# What does a pH curve look like for a weak acid-strong base reaction?

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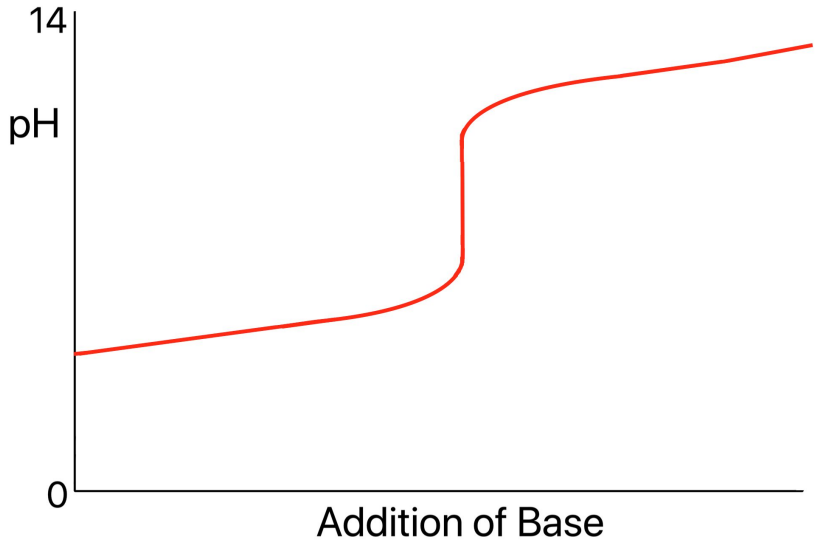

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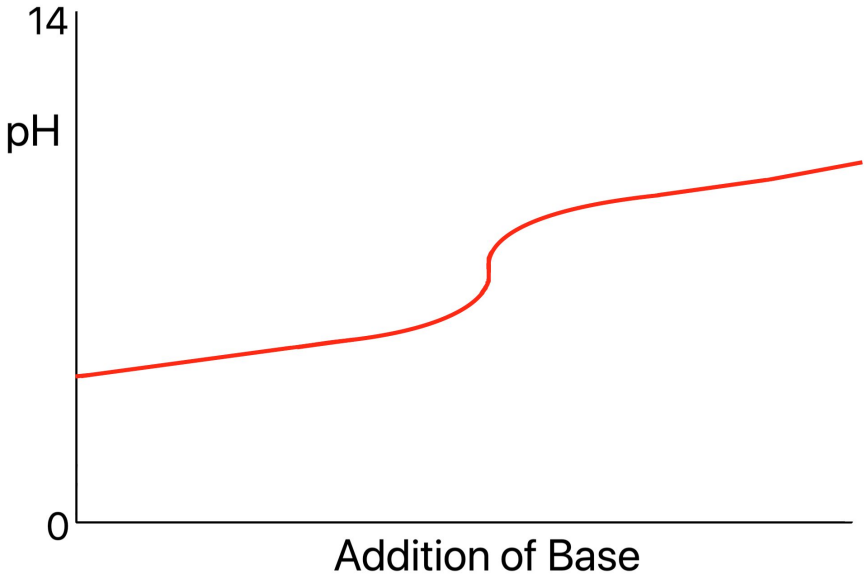

# How would you use a pH curve to determine a suitable indicator for a reaction?

# How would you use a pH curve to determine a suitable indicator for a reaction?

You want the pH range of the indicator to be in the vertical region of the pH curve for the reaction.

Weak acid-weak base reactions have no suitable indicator because they have no vertical region.

# What colour is methyl orange in acid and alkali?

# What colour is methyl orange in acid and alkali?

Acid - Red

Alkali - Yellow

# What colour is blue litmus paper in acid and alkali?

# What colour is blue litmus paper in acid and alkali?

Acid - Turns red

Alkali - Stays blue

# What colour is red litmus paper in acid and alkali?

# What colour is red litmus paper in acid and alkali?

Acid - Stays red

Alkali - Turns blue

# What colour is phenolphthalein in acid and alkali?

# What colour is phenolphthalein in acid and alkali?

Acid - Colourless

Alkali - Pink

# What is a buffer?

# What is a buffer?

A system that minimises changes in pH upon addition of small amounts of an acid or base.

# How do buffers work (upon small additions of acid)?

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E.g. $CH_3COOH_{(aq)} \rightleftharpoons CH_3COO^-_{(aq)} + H^+_{(aq)}$

Upon addition of acid:

More $H^+$ ions are present in the solution and so they combine with $CH_3COO^-_{(aq)}$ to form $CH_3COOH$. The reverse reaction is favoured (Le Chatelier's principle) and the position of equilibrium shifts to the left.

# How do buffers work (upon small additions of base)?

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E.g. $CH_3COOH_{(aq)} \rightleftharpoons CH_3COO^-_{(aq)} + H^+_{(aq)}$

Upon addition of base:

More $OH^-$ ions are present in the solution and so combine with $H^+$ to form $H_2O$. The forward reaction is favoured because $[H^+]$ is decreasing (Le Chatelier's principle) and the position of equilibrium shifts to the right.

How do you calculate the pH of a weak acid buffer solution like the following?

$$CH_3COOH_{(aq)} \rightleftharpoons CH_3COO^-_{(aq)} + H^+_{(aq)}$$

# How do you calculate the pH of a weak acid buffer solution?

$$CH_3COOH_{(aq)} \rightleftharpoons CH_3COO^-_{(aq)} + H^+_{(aq)}$$

- Write $K_a$ expression: $K_a = \frac{[H^+][CH_3COO^-]}{[CH_3COOH]}$
- Make $[H^+]$ the subject: $[H^+] = \frac{K_a \times [CH_3COOH]}{[CH_3COO^-]}$
- Calculate $[H^+]$ and then substitute into $pH = -\log[H^+]$

How can you determine $K_a$ if you know the volume of alkali needed to neutralise half of the acid?

# How can you determine $K_a$ if you know the volume of alkali needed to neutralise half of the acid?

$$HA_{(aq)} \rightleftharpoons H^+_{(aq)} + A^-_{(aq)}$$

- As half of the acid has been titrated/neutralised: $[H^+] = [HA] = [A^-]$.
- $[A^-]$ and $[HA]$ can be cancelled in the $K_a$ equation as they equal the same thing- and so $K_a = [H^+]$.
- The pH value of the mixed solutions can be converted to $[H^+]$ ($[H^+] = 10^{-pH}$) which can be used to give a $K_a$ value.

# What is the importance of buffers in biological environments?

# What is the importance of buffers in biological environments?

- In the blood: $H_2CO_3 \rightleftharpoons HCO_3^- + H^+$

  Blood has an optimum pH that is important to maintain the structure of the proteins in blood cells, among other things. This buffer helps to prevent blood pH change, keeping us alive.

- Buffers are also present in foods, preventing deterioration due to pH change.