# CAIE Chemistry A-level

## Topic 7 - Equilibria

### Flashcards

# What is a reversible reaction?

# What is a reversible reaction?

A reaction in which the products can react together to form the original reactants.

# What is dynamic equilibrium?

# What is dynamic equilibrium?

The rate of the forward reaction equals the rate of the reverse reaction, hence the concentrations of the reactants and products remains constant.

# What is Le Chatelier's principle?

# What is Le Chatelier's principle?

If a system in dynamic equilibrium is subjected to a change, then the position of equilibrium will shift to minimise this change.

# How does increasing temperature affect the position of equilibrium?

# How does increasing the temperature affect the position of equilibrium?

For an equilibrium where the forward reaction is exothermic, increasing the temperature will shift the equilibrium left (so more endothermic reactions occur) to take in more heat energy and reduce the temperature.

The reverse is true when the forward reaction is endothermic.

# How does decreasing the temperature affect the position of equilibrium?

# How does decreasing the temperature affect the position of equilibrium?

For an equilibrium where the forward reaction is exothermic, decreasing the temperature will shift the position of equilibrium to the right (so more exothermic reactions occur) to release more heat energy and increase the temperature.

The reverse is true when the forward reaction is endothermic.

# How does the concentration of reactants affect the position of equilibrium?

# How does the concentration of reactants affect the position of equilibrium?

Increasing the concentration of reactants causes the position of equilibrium to shift right in order to reduce the concentration of reactants and form more products.

The reverse occurs if the concentration of reactants is decreased.

# How does pressure affect the position of equilibrium?

# How does pressure affect the position of equilibrium?

Increasing the pressure will cause the position of equilibrium to shift to the side with the fewest gaseous molecules in order to decrease the pressure.

The opposite occurs if pressure is decreased.

If there is an equal number of gaseous molecules on both sides of the equation, changing the pressure will have no effect on the position of equilibrium.

How does the presence of a catalyst affect the position of equilibrium and the magnitude of the equilibrium constant?

# How does the presence of a catalyst affect the position of equilibrium and the magnitude of the equilibrium constant?

The presence of a catalyst doesn't affect the position of equilibrium so the magnitude of the equilibrium constant is unaffected.

The catalyst increases the rate of the forward and reverse reactions equally so equilibrium is established sooner.

# How does temperature affect the magnitude of the equilibrium constant?

# How does temperature affect the magnitude of the equilibrium constant?

If the forward reaction is **exothermic**, increasing the temperature shifts the position of equilibrium to the **left** so $K_c$ decreases.

If the forward reaction is **endothermic**, increasing the temperature shifts the equilibrium to the **right** so $K_c$ increases.

The reverse is true if temperature is decreased.

# How does pressure affect the magnitude of the equilibrium gas constant?

# How does pressure affect the magnitude of the equilibrium gas constant?

$K_p$ remains the same:

- Doubling the pressure will double both the partial pressures and concentrations of the species on both sides of the equation.
- The system is no longer in equilibrium so partial pressures of reactants and products must change to keep $K_p$ the same.
- New equilibrium position will be reached whereby $K_p$ is restored (the ratio of the $K_p$ expression is the same as before).

# How does concentration affect the magnitude of the equilibrium constant?

# How does concentration affect the magnitude of the equilibrium constant?

Changing the concentration of a reactant or product means that the system is no longer in equilibrium.

The concentrations of the reactants and products now must change so that the ratio, and hence $K_c$, is restored.

$K_c$ is therefore unaffected by concentration changes.

# What does the equilibrium constant tell you?

# What does the equilibrium constant tell you?

The position of equilibrium of a reaction.

The magnitude indicates whether there are more reactants or products in an equilibrium system.

How would you calculate the equilibrium constant ($K_c$) for the following reaction?

$$3H_2 + N_2 \rightleftharpoons 2NH_3$$

How would you calculate the equilibrium constant ($K_c$) for the following reaction? $3H_2 + N_2 \rightleftharpoons 2NH_3$

$$K_c = \frac{[NH_3]^2}{[N_2][H_2]^3}$$

Multiply all the concentrations of the products to the power of their balancing number. Divide this by the product of all the concentrations of the reactants to the power of their balancing number. This gives $K_c$, the equilibrium constant.

Use the $K_c$ expression to work out the units for the equilibrium constant for the equation below:

$$3H_2 + N_2 \rightleftharpoons 2NH_3$$

Use the $K_c$ expression to work out the units for the equilibrium constant for the equation below:

$3H_2 + N_2 \rightleftharpoons 2NH_3$

$$K_c = \frac{[NH_3]^2}{[N_2][H_2]^3}$$

Substitute in the units and cancel down:

$$K_c = \frac{(mol\ dm^{-3})^2}{mol\ dm^{-3} \times (mol\ dm^{-3})^3} = \frac{mol^2\ dm^{-6}}{mol^4\ dm^{-12}} = dm^6\ mol^{-2}$$

# What is $K_p$?

# What is $K_p$?

The equilibrium constant for reactions in the gaseous phase.

It is similar to $K_c$ but it uses partial pressures instead of concentrations.

# How do you calculate the mole fraction of a gas?

# How do you calculate the mole fraction of a gas?

## For gas A:

$$Mole\ fraction,\ X_a = \frac{No.\ of\ moles\ of\ gas\ A}{Total\ no.\ of\ moles\ of\ all\ present\ species}$$

# How do you calculate the partial pressure of a gas?

# How do you calculate the partial pressure of a gas?

## For gas A:

$$Partial\ pressure\ of\ A,\ p(A) = Mole\ fraction,\ X_A \times Total\ pressure$$

Write the $K_p$ expression for the following reaction: $H_{2(g)} + I_{2(g)} \rightarrow 2HI_{(g)}$

Write the $K_p$ expression for the following reaction:

$H_{2(g)} + I_{2(g)} \rightarrow 2HI_{(g)}$

$$K_p = \frac{p(HI)^2}{p(H) \times p(I)}$$

where *p(X)* is the partial pressure of gas X and the power is the balancing number in the equation.

Use the $K_p$ expression to work out the units for the equilibrium constant for the equation below:

$$H_{2(g)} + I_{2(g)} \rightarrow 2HI_{(g)}$$

Use the $K_p$ expression to work out the units for the equilibrium constant for the equation below:

$H_{2(g)} + I_{2(g)} \rightarrow 2HI_{(g)}$

Substitute the units into the $K_p$ expression and cancel any common units. In this case all units cancel out:

$$K_p = \frac{p(HI)^2}{p(H) \times p(I)}$$

$$K_p = \frac{(kPa)^2}{kPa \times kPa} = \frac{kPa^2}{kPa^2} = no\ units$$

$$CH_4 + 2H_2O \rightleftharpoons CO_2 + 4H_2$$

Initially, a mixture contains 3.00 mol of $H_2O$ and 1.50 mol of $CH_4$. At equilibrium, this mixture contains 0.250 mol of $CO_2$. Calculate the quantity of each compound at equilibrium.

$CH_4 + 2H_2O \rightleftharpoons CO_2 + 4H_2$

Initially, a mixture contains 3.00 mol of $H_2O$ and 1.50 mol of $CH_4$. At equilibrium, this mixture contains 0.250 mol of $CO_2$. Calculate the quantity of each compound at equilibrium.

**Black**- from question
**Red** - use the stoichiometric ratio in the equation.
**Blue** - calculated quantities (subtract reacted from initial)

| | $CH_4$ + | $2H_2O$ ⇌ | $CO_2$ + | $4H_2$ |
|---|---|---|---|---|
| Initial amount / mol | **1.50** | **3.00** | 0.00 | 0.00 |
| Amount reacted / mol | 0.250 | 0.500 | 0.00 | 0.00 |
| Amount at equilibrium / mol | 1.50 - 0.250 = 1.25 | 3.00 - 0.500 = 2.50 | **0.250** | 1.00 |

No products in initial mixture

# Calculate $K_c$ using the quantities at equilibrium given below. The total volume is 4 $dm^3$.

| | $CH_4$ + | $2H_2O$ ⇌ | $CO_2$ + | $4H_2$ |
|---|---|---|---|---|
| **Amount at equilibrium / mol** | 1.25 | 2.50 | 0.250 | 1.00 |

# Calculate $K_c$ using the quantities at equilibrium below. The volume is 4 $dm^3$.

Calculate concentration using 'concentration = moles ÷ volume':

| | $CH_4$ + | $2H_2O$ ⇌ | $CO_2$ + | $4H_2$ |
|---|---|---|---|---|
| **Amount at equilibrium / mol** | 1.25 | 2.50 | 0.250 | 1.00 |
| **Concentration / mol $dm^{-3}$** | 1.25 ÷ 4 = 0.313 mol $dm^{-3}$ | 2.50 ÷ 4 =0.625 mol $dm^{-3}$ | 0.250 ÷ 4 0.0625 mol $dm^{-3}$ | 1.00 ÷ 4 = 0.250 mol $dm^{-3}$ |

$$K_c = \frac{[H_2]^4[CO_2]}{[CH_4][H_2O]^2} = \frac{(0.250)^4(0.0625)}{(0.313)(0.625)^2} = 2\ x\ 10^{-3}\ mol^2 dm^{-6}$$

What is the Haber process? Write an equation for this reaction and state the source of the reactants

# What is the Haber process? Write an equation for this reaction and state the source of the reactants

It is a process that produces ammonia:

$$N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$$

- Nitrogen - fractional distillation of air
- Hydrogen - natural gas

# What conditions are used for the Haber process?

# What conditions are used for the Haber process?

- 450℃
- 200 atm
- Iron catalyst

What temperature would you expect to be used for the Haber process and why? Why is the temperature used in industry a compromise?

What temperature would you expect to be used for the Haber process and why? Why is the temperature used in industry a compromise?

The forward reaction is exothermic so a low temperature would give the greatest yield.

However, a low temperature gives a slow rate of reaction and so a higher temperature is normally used to strike a balance between yield and rate.

What pressure would you expect to be used for the Haber process and why? Why is the pressure used in industry a compromise?

What pressure would you expect to be used for the Haber process and why? Why is the pressure used in industry a compromise?

According to Le Chatelier's principle, a high pressure would give the greatest yield of ammonia.

However, high pressures can be dangerous and expensive to maintain so a slightly lower, safer pressure is used.

What is the Contact process? Write an equation for this reaction and state the source of the reactants

# What is the Contact process? Write an equation for this reaction and state the source of the reactants

The process for creating sulfur trioxide:

$$2SO_{2(g)} + O_{2(g)} \rightleftharpoons 2SO_{3(g)}$$

- Sulfur dioxide - heat sulfur in air
- Oxygen - air

# What conditions are used for the Contact process?

# What conditions are used for the Contact process?

- 400 - 450℃
- 1 - 2 atm
- $V_2O_5$ catalyst

What temperature would you expect to be used for the Contact process and why? Why is the temperature used in industry a compromise?

What temperature would you expect to be used for the Contact process and why? Why is the temperature used in industry a compromise?

The forward reaction is exothermic so a low temperature would give the greatest yield.

However, a low temperature results in a slow rate of reaction and so a higher temperature may be used to strike a balance between yield and rate.

What pressure would you expect to be used for the Contact process and why? Why is the pressure used in industry a compromise?

# What pressure would you expect to be used for the Contact process and why? Why is the pressure used in industry a compromise?

According to Le Chatelier's principle, a high pressure would give the greatest yield of sulfur trioxide.

However, even at pressures close to atmospheric pressure, 99.5% of $SO_2$ is converted into $SO_3$ so increasing the pressure would only see a minute improvement in yield that wouldn't be economically worthwhile.

# 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 is the proton donor and acceptor in this reaction:

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

What is the proton donor and acceptor in this reaction: $HNO_{2(s)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + NO_2^-{}_{(aq)}$

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

Proton acceptor: $H_2O_{(l)}$

What is a conjugate acid-base pair?

Label the conjugate acid-base pairs in the equation below:

$$HCl_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + Cl^-_{(aq)}$$

What is a conjugate acid-base pair? Label the conjugate acid-base pairs in the equation below:

$HCl_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + Cl^-_{(aq)}$

A conjugate acid-base pair is two species that differ from each other by a proton ($H^+$ ion)

- HCl and $Cl^-$ are a conjugate acid-base pair.
- $H_2O$ and $H_3O^+$ are a conjugate acid-base pair.

What is a strong acid? Write an equation to show that HCl is a strong acid

# What is a strong acid? Write an equation to show that HCl is a strong acid

A strong acid is an acid that completely dissociates in solution:

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

Typical pH is 0 - 1.

What is a strong base? Write an equation to show that NaOH is a strong base

# What is a strong base? Write an equation to show that NaOH is a strong base

A strong base is a base that completely dissociates in solution:

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

Typical pH is close to 14.

What is a weak acid? Write an equation to show that ethanoic acid is a weak acid

# What is a weak acid? Write an equation to show that ethanoic acid is a weak acid

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

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

Typical pH is 2 - 6.

What is a weak base? Write an equation to show that ammonia is a weak base

# What is a weak base? Write an equation to show that ammonia is a weak base

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

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

Typical pH is above 7 but lower than a strong base.

What type of reaction is the following?

$$HCl + NaOH \rightarrow H_2O + NaCl$$

What type of reaction is the following?

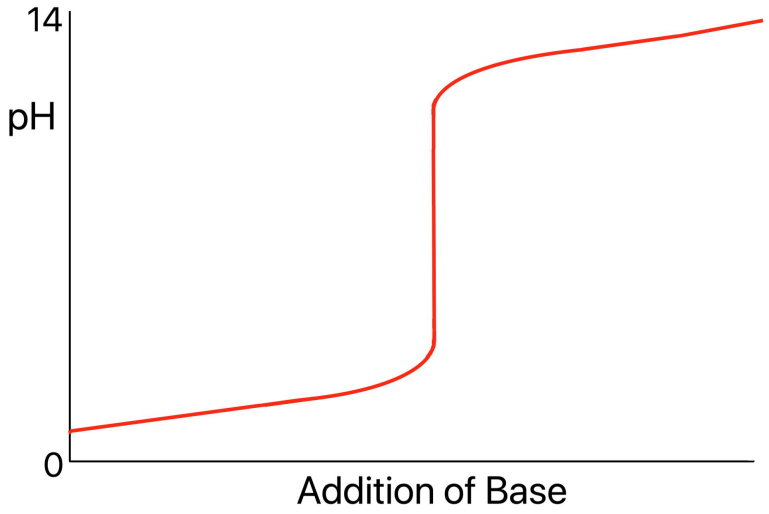
$HCl + NaOH \rightarrow H_2O + NaCl$

Neutralisation

An acid and a base react to form a salt and water.

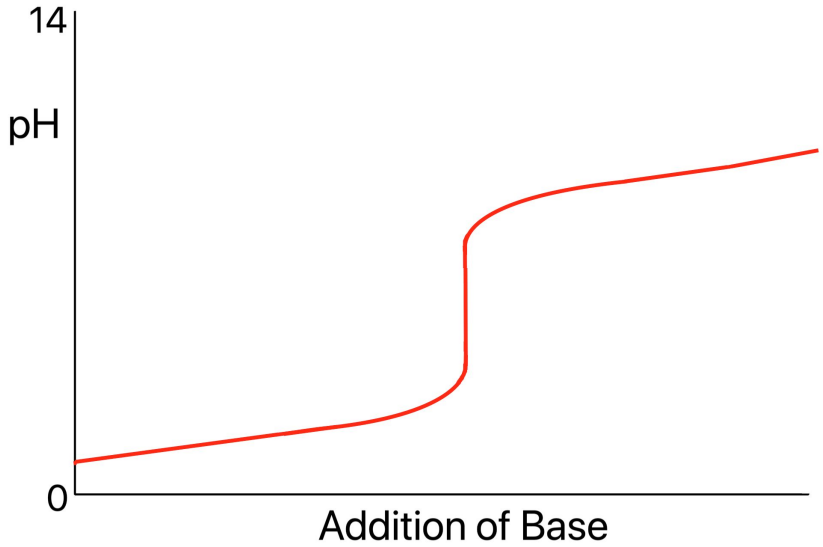
# Draw a pH titration curve for a strong acid-strong base reaction

# Draw a pH titration curve for a strong acid-strong base reaction



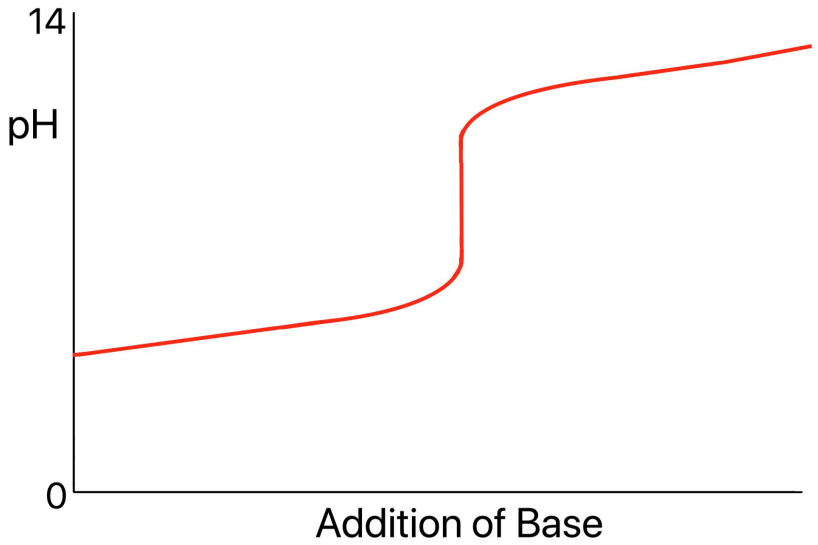
# Draw a pH titration curve for a strong acid-weak base reaction

# Draw a pH titration curve for a strong acid-weak base reaction



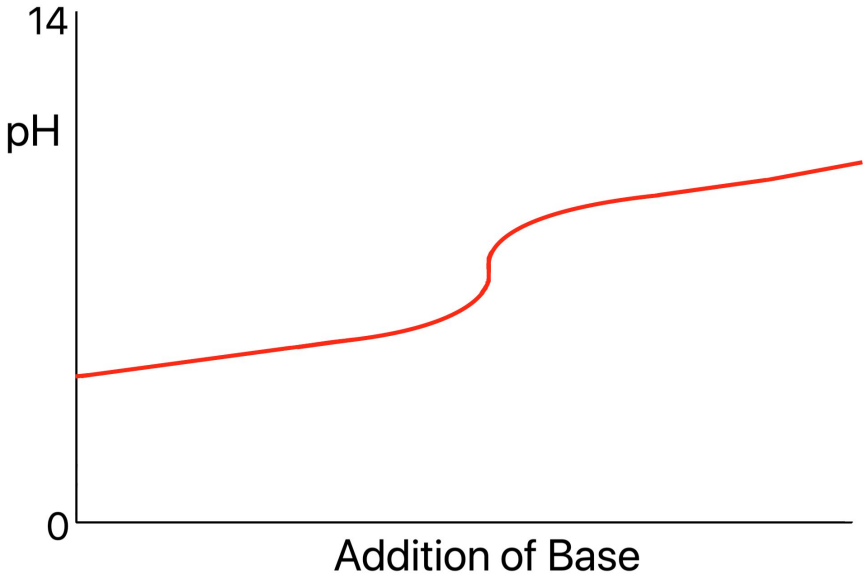
# Draw a pH titration curve for a weak acid-strong base reaction

# Draw a pH titration curve for a weak acid-strong base reaction



# Draw a pH titration curve for a weak acid-weak base reaction

# Draw a pH titration curve for a weak acid-weak base reaction



# How would you use a pH titration curve to determine a suitable indicator for an acid-base titration?

# How would you use a pH titration curve to determine a suitable indicator for an acid-base titration?

The pH range of the colour change of the indicator must be within the vertical section of the pH curve.

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

# Why must an indicator change colour within the vertical section of pH titration curve for an acid-base titration?

# Why must an indicator change colour within the vertical section of pH titration curve for an acid-base titration?

Within the vertical section, the volume of acid (or alkali) added changes very little. When the indicator changes colour, this is the end point of the titration. It shows the point at which the correct volumes of acid and alkali have been combined so that neither reactant is in excess. If the indicator changes colour outside the vertical sections, the volumes of acid and alkali won't be close to the actual volumes required for the neutralisation reaction.