# Edexcel GCSE Chemistry

Your notes

## Dynamic Equilibria

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## Reaching Equilibrium

# Predicting the Conditions

## Predicting the Ideal Conditions

- We have previously seen that ammonia is manufactured using the **Haber Process**
- The hydrogen and nitrogen react to form ammonia in the following **reversible reaction:**

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \qquad \Delta H = -92 \text{ kJ mol}^{-1}$$

- The formation of ammonia is exothermic, so using Le Chatelier's Principle we would predict that
  - The reaction will produce a higher yield at low temperatures
  - Using a high pressure would increase the yield as there are fewer moles of gas on the right than the left of the equation

## Reaching Equilibrium

- Equilibrium occurs when during the course of a reversible reaction, the rate of the forward reaction **equals** the rate of the reverse reaction
- This means that products are being formed in the forward reaction as fast as reactants are being formed in the reverse reaction
- It is reached at a faster rate when:
  - A **higher pressure** is used as there are more **successful collisions**
  - A **higher temperature** is used as the particles have greater **kinetic energy**
  - A **higher concentration** is used as there are **more particles** per given volume, hence there are more collisions
  - A **catalyst** is used as it speeds up the **rate** of reaction, allowing it to reach equilibrium faster

### Examiner Tips and Tricks

Remember that Le Chatelier's Principle tell us that any change to a system at equilibrium results in the equilibrium responding by opposing the effect of that change. Cooling an exothermic reaction results in the equilibrium shifting in the exothermic direction (to produce more heat and raise the temperature). Increasing pressure results in the equilibrium shifting to the side with the fewer gas molecules (to decrease the gas pressure).

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# Choosing the Conditions

## Economic Considerations

- Like all industries, companies that manufacture and sell chemical goods do so to make a profit
- Part of the industrial process is the economic decision on how and where to design and implement a manufacturing site
- The availability and cost of **raw materials** is a major consideration which must be studied well before any decisions are taken
- In the Haber Process the raw materials are readily available and inexpensive to purify:
  - **Nitrogen** - from the **air**
  - **Hydrogen**- from **natural gas**
- If the cost of extraction of raw materials is too high or they are unavailable then the process is no longer **economically viable**
- Many industrial processes require huge amounts of **heat** and **pressure** which is very expensive to maintain
- Production energy costs are also a factor to be considered carefully and alongside the raw materials issue

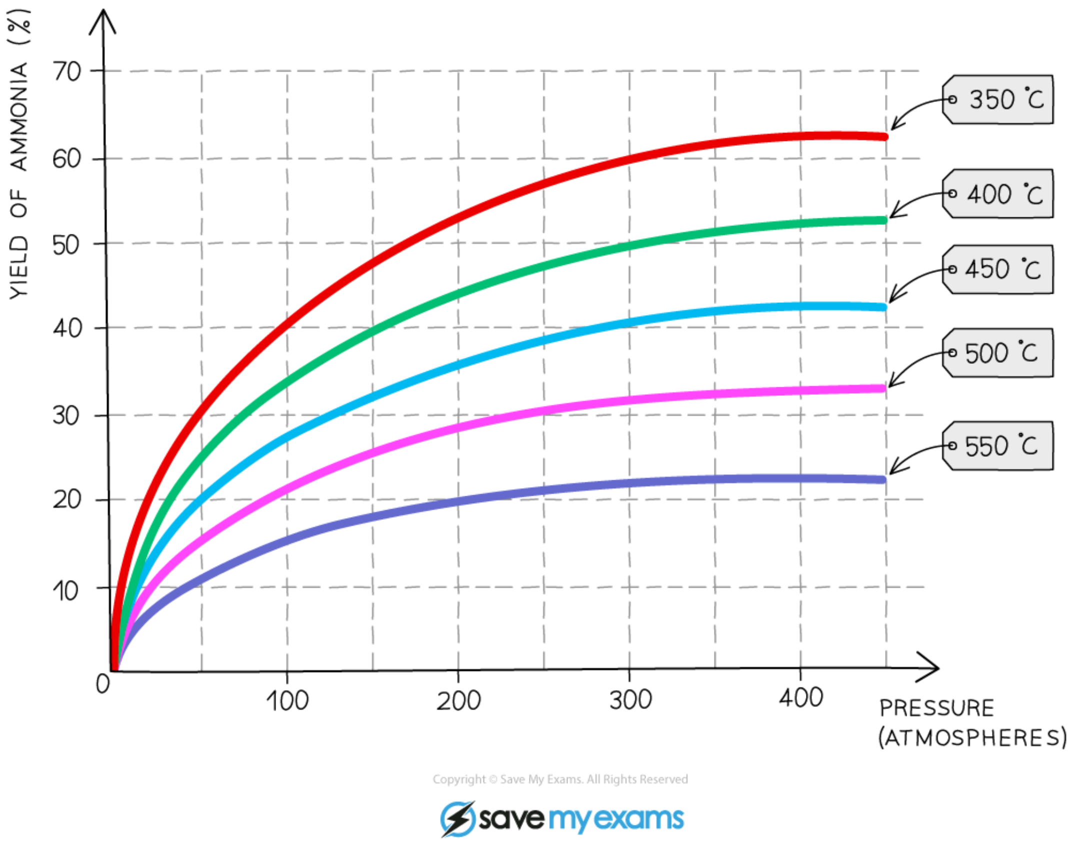
## Temperature: 450°C

- A **higher** temperature would favour the reverse reaction as it is endothermic (takes in heat) so a higher yield of **reactants** would be made
- If a **lower** temperature is used it favours the forward reaction as it is exothermic (releases heat) so a higher yield of **products** will be made
- However at a lower temperature the rate of reaction is very **slow**
- So 450°C is a compromise temperature between having a **lower yield** of products but being made more quickly

## Pressure: 200 atm

- A **lower** pressure would favour the reverse reaction as the system will try to increase the pressure by creating more molecules (4 molecules of gaseous reactants) so a higher yield of **reactants** will be made

- A **higher** pressure would favour the forward reaction as it will try to decrease the pressure by creating fewer molecules (2 molecules of gaseous products) so a higher yield of **products** will be made
- However, high pressures can be dangerous and very expensive equipment is needed
- So 200 atm is a **compromise** pressure between a lower yield of products being made **safely** and **economically**




***Choosing the conditions for the Haber Process***

## Catalyst

- The presence of a catalyst does **not** affect the position of equilibrium but it does increase the **rate** at which equilibrium is reached
- This is because the catalyst increases the rate of **both** the forward and backward reactions by the same amount (by providing an alternative pathway requiring **lower activation energy**)
- As a result, the **concentration** of reactants and products is nevertheless the **same** at equilibrium as it would be without the catalyst.

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So a catalyst is used as it helps the reaction reach equilibrium quicker

- It allows for an acceptable yield to be achieved at a **lower temperature** by lowering the activation energy required
- Without it the process would have to be carried out at an even **higher temperature**, increasing **costs** and decreasing **yield** as the higher temperature **decomposes** more of the $NH_3$ molecules

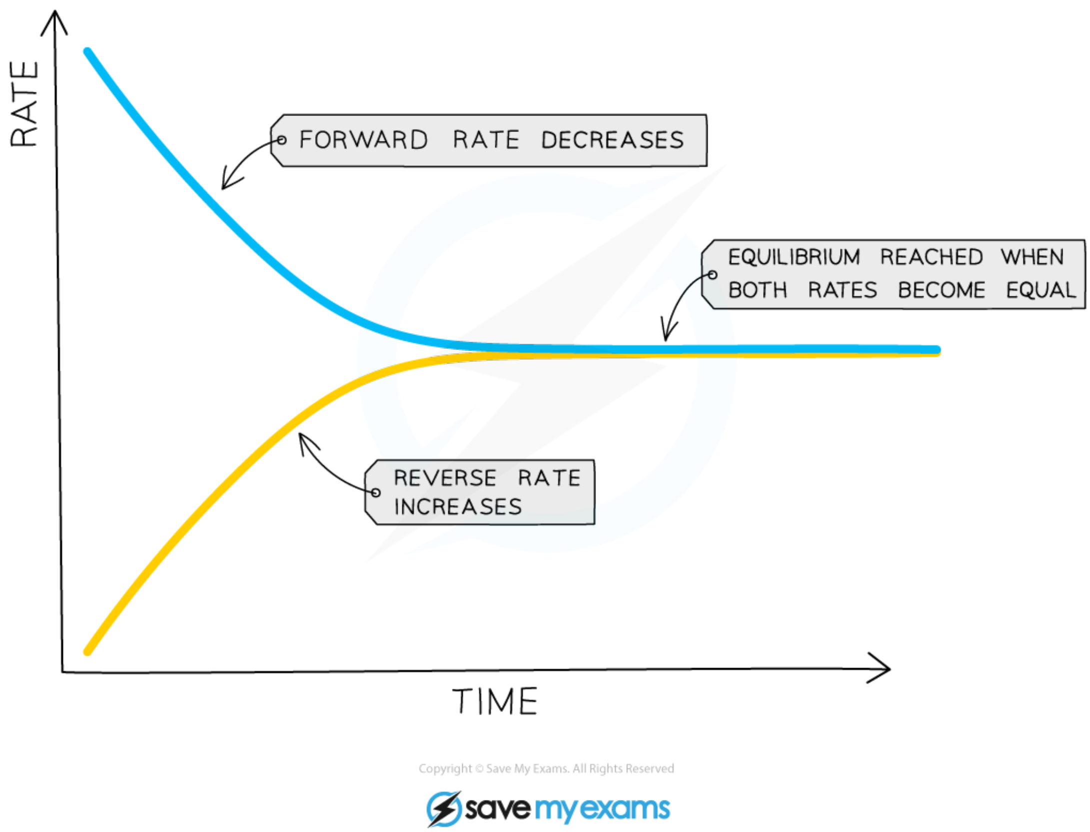


***Diagram showing the effect of catalyst on equilibrium position***

### Examiner Tips and Tricks

The reaction conditions chosen for the Haber process are not ideal in terms of the yield but do provide balance between product yield, reaction rate and production cost. These are called **compromise conditions** as they are chosen to give a good compromise between the yield, rate and cost.

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## Fertilisers

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# Fertilisers

- Compounds containing **nitrogen, potassium** and **phosphorus** are used as fertilisers to increase crop yields
- **NPK** fertilisers are formulations containing appropriate ratios of all three elements
- From these three essential elements:
  - Nitrogen promotes healthy **leaves,**
  - Potassium promotes **growth,** healthy **fruit** and **flowers**
  - Phosphorus promotes healthy **roots**
- A distinct advantage of artificial fertilisers is that they can be **designed** for specific needs whereas in natural fertilizers, such seaweed or manure, the proportions of elements cannot be controlled
- Fertiliser compounds contain the following **water soluble** ions:
  - Ammonium ions, $NH_4^+$ and nitrate ions, $NO_3^-$, which are sources of soluble nitrogen
  - Phosphate ions, $PO_4^{3-}$, which are a source of soluble phosphorus
- Most common potassium compounds dissolve in water to produce potassium ions, $K^+$

| Fertiliser | Formula | Essential Element |
|---|---|---|
| Ammonium nitrate | $NH_4NO_3$ | Nitrogen |
| Calcium phosphate | $Ca_3(PO_4)_2$ | Phosphorus |
| Potassium nitrate | $KNO_3$ | Potassium, nitrogen |
| Potassium sulfate | $K_2SO_4$ | Potassium |



- Ammonia is an **alkaline** substance and neutralises acids producing a salt and water
- The salt it produces contains the **ammonium** ion, $NH_4^+$, which is a component of several fertilisers

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- Ammonia also undergoes oxidation to produce nitric acid, $HNO_3$
- Nitric acid is used as the source of the **nitrate** ion, $NO_3^-$, which is another important ion found in fertilisers
- Ammonium nitrate, a fertiliser and one of the most important ammonium salts, is made by reacting ammonia with nitric acid:

$$NH_3\,(aq) + HNO_3\,(aq) \rightarrow NH_4NO_3\,(aq)$$

### Examiner Tips and Tricks

Fertilisers must be water soluble so the nutrients they provide can be effectively absorbed and transported by the plant.

# Making Fertilisers

## Preparation of Ammonium Sulfate in the Laboratory

**Aim:**

- To prepare ammonium sulfate by titration:

$$2NH_3 + H_2SO_4 \rightarrow (NH_4)_2SO_4$$

**Materials:**

- Dilute ammonia solution, dilute solution of sulfuric acid, methyl orange indicator
- Clamp and stand, burette and volumetric pipette, conical flask, white tile

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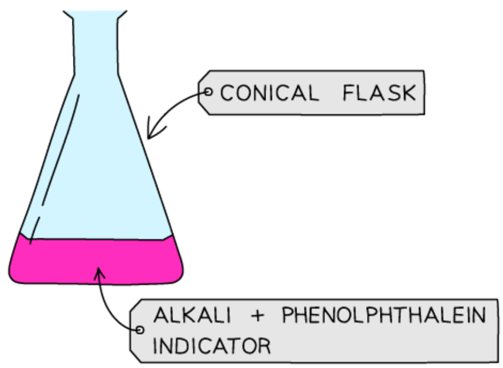
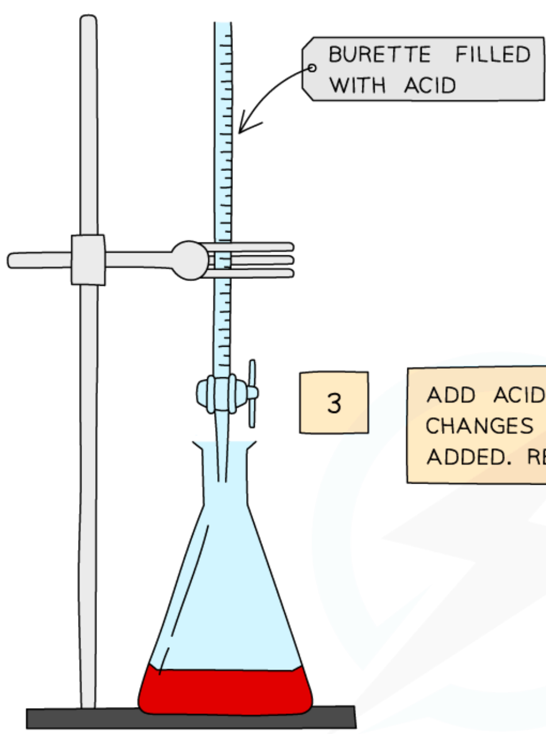
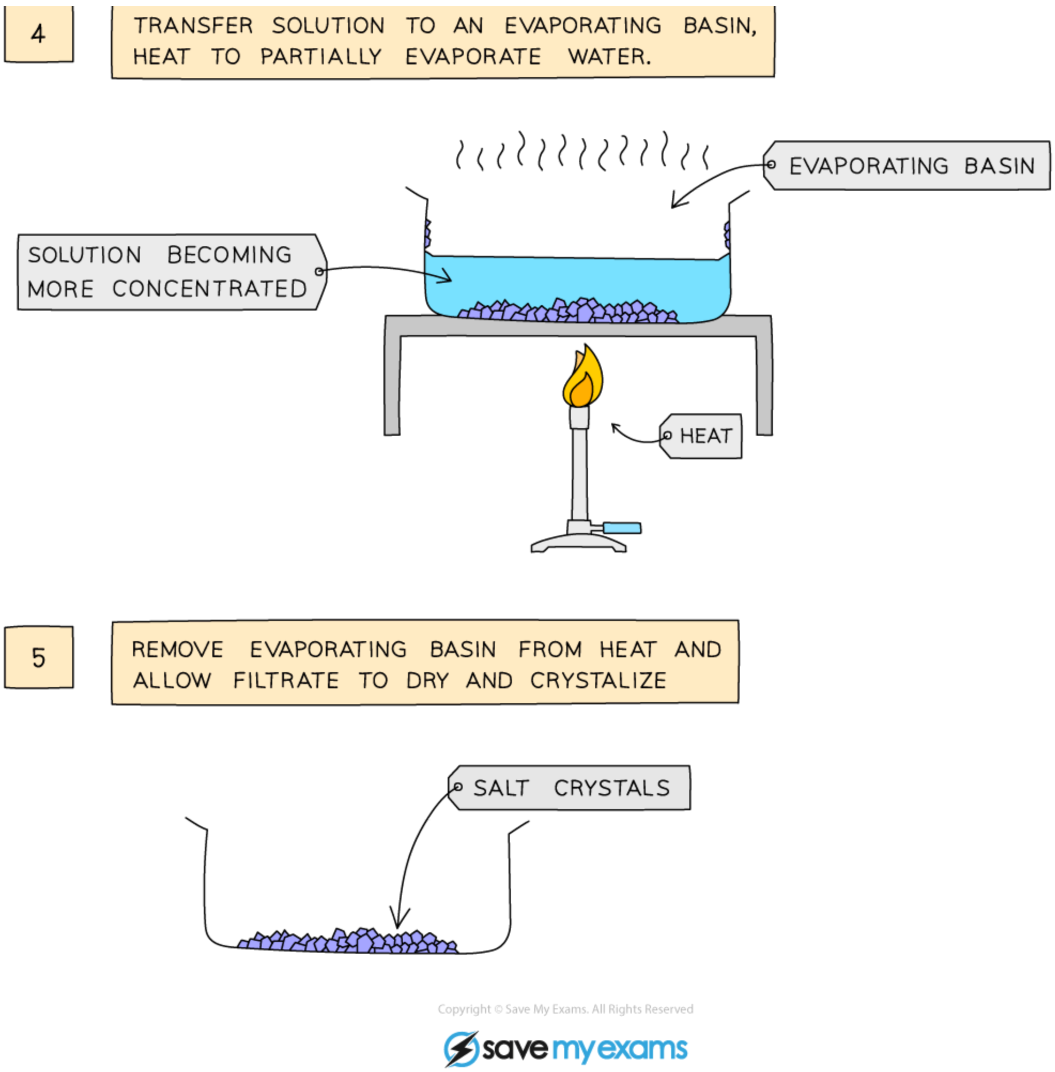
1 ADD ALKALI + INDICATOR TO CONICAL FLASK USING A PIPETTE




2 ADD ACID TO BURETTE, NOTING THE STARTING VOLUME




3 ADD ACID TO ALKALI SLOWLY UNTIL INDICATOR CHANGES COLOUR. CALCULATE VOLUME OF ACID ADDED. REPEAT STEPS 1–3 WITHOUT INDICATOR




***Steps in the laboratory preparation of ammonium sulfate***

**Method:**

- Add an **exact volume** of ammonia to the conical flask and place on the white tile
- Add a few drops of indicator and swirl, it should turn **yellow**
- Add the acid to the flask solution **drop by drop**, swirling the flask in between
- Continue until the colour turns red sharply and record the **titre**
- Repeat by adding **exactly** the same amount of acid but this time without the indicator which is an **impurity**

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- Pour the reaction mixture in an **evaporating dish** and **gently** heat in a water bath to remove some of the water
- Stop heating when the volume has been reduced to roughly **one third** of its volume
- Leave in a dry place so the remaining water evaporates, allowing **crystallisation** to occur
- This may take a few days depending on ambient conditions

**Analysis of results:**

- After a few days ammonium sulfate crystals should appear
- Filter to remove any remaining water

## Industrial Preparation

- The industrial preparation of ammonium sulfate is a **large scale** operation consisting of several **stages**
- Ammonia is prepared by the **Haber** process and sulfuric acid by the **Contact** process
- Both processes require their own supplies of **raw materials**, **energy** and **equipment**
- The most common industrial process of manufacturing ammonium sulfate involves filling a large reactor chamber with **ammonia** gas.
- **Sulfuric acid** is **sprayed** into the chamber from above and ammonium sulfate powder is produced
- Another method involves pumping a mixture of ammonia gas and steam in a reactor which contains some sulfuric acid and a concentrated solution of ammonium sulfate
- The reaction is carried out at **60 °C** and concentrated sulfuric is added gradually

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## Comparing the laboratory preparation & industrial production of ammonium sulfate

- Comparing the two processes highlights the challenges of scaling up laboratory preparations to industrial levels

**Different Ways to Make Ammonium Sulfate**

| | Laboratory | Industrial |
|---|---|---|
| **Equipment** | **Simple** equipment needed, prepared using a titration apparatus | Hugely **expensive** and **complex** |
| **Reactant concentration** | **Low** concentrations, less heat given off | **High** concentrations, very exothermic reaction |
| **Separation of product** | **Crystallisation** is used which is a slow process | The heat produced is used to evaporate water from the reaction mixture to make **very concentrated** ammonium nitrate product |



### Examiner Tips and Tricks

Notice that when writing ammonia solution as $NH_3$ (aq), water does not appear to be a product of the neutralisation reaction. However, ammonia solution may also be written as, $NH_4OH$ (aq), ammonium hydroxide, in which case water is produced:

$$NH_4OH\ (aq) + HNO_3\ (aq) \rightarrow NH_4NO_3\ (aq) + H_2O\ (l)$$

Either formula may be used to show the reactions.