FACTFILE: GCSE CHEMISTRY: UNIT 2.4

Rewarding Learning

Equilibrium

Learning outcomes

Student should be able to:

2.4.1 demonstrate knowledge and understanding that many chemical reactions are reversible and the direction of a reversible reaction can be changed by altering the reaction conditions;

2.4.2 demonstrate knowledge and understanding that dynamic equilibrium occurs in a closed system when the rates of forward and reverse reactions are equal and the amounts of reactants and products remain constant;

2.4.3 **state Le Châtelier's Principle and use it to predict the qualitative effects of changes of temperature, pressure and concentration on the position of equilibrium for a closed homogeneous system;**

2.4.4 **describe the Haber process as a reversible reaction between nitrogen and hydrogen to form ammonia and describe and explain the conditions used and the trade-off between rate of production and position of equilibrium.**

The majority of the reactions met at GCSE are irreversible. For example - when magnesium reacts with sulfuric acid a metal salt and hydrogen gas are produced.

$$\text{magnesium} + \text{sulfuric acid} \longrightarrow \text{magnesium sulfate} + \text{hydrogen}$$

$$Mg + H_2SO_4 \longrightarrow MgSO_4 + H_2$$

This reaction will stop when all of the magnesium metal has reacted, essentially the reaction has 'gone to completion' and this is indicated by an arrow $\longrightarrow$.

If a chemical reaction happens in a container where one or more of the reactants or products can escape this is termed an open system. If a chemical reaction happens in a container where none of the reactants or products can escape this is termed a closed system.

A reversible reaction is a chemical change in which the products can be converted back to the original reactants under suitable conditions. This means a reaction can go in either direction. *reversible reactions* that happen in a closed system eventually reach equilibrium.

A reversible reaction is shown by the sign $\rightleftharpoons$.

$$A + B \rightleftharpoons C + D$$

- a half-arrow to the right (direction of forward reaction),
- a half-arrow to the left (direction of backward reaction).

For this equation the forward reaction is:

$$A + B \longrightarrow C + D$$

And the backward reaction is:

$$C + D \longrightarrow A + B$$

A reversible reaction does not go to completion in either direction. All components, original reactants or ensuing products co-exist in the reaction mixture. In a reversible reaction, changing the reaction conditions will change the net direction in which the reaction goes. It will either go more to the right (forward direction) or more to left (backward direction).

The following conditions will change the equilibrium position in a reversible reaction:
- changing the concentration of a reactant;
- changing the concentration of a product;
- changing the temperature;
- changing the pressure in a reaction involving gases.

When a system is in a state of equilibrium nothing appears to be happening. However, the system is *dynamic* meaning it is in constant motion. A chemical system is in dynamic equilibrium when:
- the rate of the forward reaction is the same as the rate of the reverse reaction;
- the concentrations of the reactants and products remain the same.

Dynamic equilibrium – *this is the equilibrium when the rate of the forward reaction is equal to the rate of the reverse reaction and the amount of reactants and products remain constant.*

Le Châtelier's principle *states that if a change is made to the conditions of a system at equilibrium, then the position of equilibrium moves to oppose the change in conditions.*

This principle can be used to predict how the position of equilibrium will change when the reaction conditions are changed. The reactions are *homogeneous* which means the reactants and products are all in the same physical state.

In each example the general equation $A + B \rightleftharpoons C + D$ will be used.

It is important to remember that for this equation the forward reaction is:

$$A + B \longrightarrow C + D$$

And the backward reaction is:

$$C + D \longrightarrow A + B$$

Changing the concentration of a reactant

$$A + B \rightleftharpoons C + D$$

- Increasing the concentration of reactant (increasing concentration of A or B)

The position of equilibrium will shift to **decrease** the concentration of the reactant to minimise the change. The forward reaction uses up reactants A and B and therefore the equilibrium position shifts to favour this reaction – the position of equilibrium shifts to the RIGHT.

- Decreasing the concentration of reactant (decreasing concentration of A or B)

The position of equilibrium will shift to **increase** the concentration of the reactant to minimise the change. The backward reaction produces reactants A and B and therefore the equilibrium position shifts to favour this reaction – the position of equilibrium shifts to the LEFT.

Changing the concentration of a product

$$A + B \rightleftharpoons C + D$$

- Increasing the concentration of product (increasing concentration of C or D)

The position of equilibrium will shift to **decrease** the concentration of the product to minimise the change. The backward reaction uses up the products C and D and therefore the equilibrium position shifts to favour this reaction – the position of equilibrium shifts to the LEFT.

- Decreasing the concentration of product (decreasing concentration of C or D)

The position of equilibrium will shift to **increase** the concentration of the product to minimise the change. The forward reaction produces the products C and D and therefore the equilibrium position shifts to favour this reaction – the position of equilibrium shifts to the RIGHT.

Changing the temperature of the system

In these reactions it is important to know the energy change associated with the forward and backward reactions. A positive energy change is an *endothermic* process and a negative energy change is an *exothermic* process. Endothermic reactions will decrease the temperature in a system and exothermic reactions will increase the temperature in a system.

If you know the energy change for one of the directions in the reaction, the other direction is the opposite energy change.

$$\boxed{\text{EXOTHERMIC}}$$

$$\xrightarrow{\hspace{2cm}}$$

$$A + B \rightleftharpoons C + D \qquad \Delta H = \text{NEGATIVE}$$

$$\xleftarrow{\hspace{2cm}}$$

$$\boxed{\text{ENDOTHERMIC}}$$

- Increasing the temperature of the system.

The position of equilibrium will shift to **decrease** the temperature of the system to minimise the change. To decrease the temperature the endothermic reaction must be favoured. The backward reaction is endothermic so the equilibrium position shifts to favour this reaction – the position of equilibrium shifts to the LEFT.

- Decreasing the temperature of the system.

The position of equilibrium will shift to **increase** the temperature of the system to minimise the change. To increase the temperature the exothermic reaction must be favoured. The forward reaction is exothermic so the equilibrium position shifts to favour this reaction – the position of equilibrium shifts to the RIGHT.

Changing the pressure of the system

Changing the pressure will have an effect in those reactions containing gaseous species. The total number of gas moles on each side of the reaction must be calculated before predicting changes to the position of equilibrium.

$$2A(g) + B(g) \rightleftharpoons 3C(g) + 2D(g)$$

In this reaction there are 2 moles of A and one mole of B so a total of 3 moles of reactants would be produced in the backward reaction. There are 3 moles of C and 2 moles of D so a total of 5 moles of products would be produced in the forward reaction.

$$\underbrace{2A(g) + B(g)}_{\boxed{\text{3 moles}}} \rightleftharpoons \underbrace{3C(g) + 2D(g)}_{\boxed{\text{5 moles}}}$$

- Increasing the pressure of the system.

The position of equilibrium will shift to **decrease** the pressure of the system to minimise the change. To decrease the pressure the reaction which produces **fewer** moles of gas must be favoured. The backward reaction produces 3 moles compared to 5 moles in the forward reaction so the equilibrium position shifts to favour the backward reaction – the position of equilibrium shifts to the LEFT.

- Decreasing the pressure of the system.

The position of equilibrium will shift to **increase** the pressure of the system to minimise the change. To increase the pressure the reaction which produces **more** moles of gas must be favoured. The backward reaction produces 3 moles compared to 5 moles in the forward reaction so the equilibrium position shifts to favour the forward reaction – the position of equilibrium shifts to the RIGHT.

THE HABER PROCESS

Many important chemical processes exist as equilibrium systems including the preparation of ammonia from nitrogen and hydrogen in the Haber process. Ammonia is a very important material in the chemical industry – it is used to make, e.g. nitric acid, fertilisers, medicines, explosives and dyestuffs.

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

Using Le Châtelier's principle the theoretical conditions that would favour the production of ammonia are a high pressure and a low temperature. However there are drawbacks in using these theoretical conditions:

- A high pressure will produce a good yield of ammonia but it requires large amounts of energy to compress the gases which has significant cost. There are also safety implications to consider as a leak could cause harm to the workforce and the environment.
- A low temperature despite producing a good yield of ammonia would have a low rate of reaction.

Compromise Temperature

For the process to be viable the ammonia plant needs to produce a sufficient yield of ammonia at a reasonable cost and in as short a time as possible. Therefore a compromise between yield and rate is made.

Temperature – 450°C

- This is high enough to allow the reaction to proceed at a good rate and still produces and acceptable yield of ammonia.

Pressure – 200 atmospheres

- This is still quite high but safe enough to limit danger to workforce and environment.

Catalyst – iron

- This speeds up the rate of the reaction and allows the equilibrium to be established faster meaning lower temperatures can be used. Using a lower temperature will reduce costs.

REVISION QUESTIONS

1 (a) Each of the equations in the table below, labelled A to E represents a dynamic equilibrium.

A	$N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$	ΔH = +180 kJ/mol
B	$N_2O_4(g) \rightleftharpoons 2NO(g)$	ΔH = +58 kJ/mol
C	$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$	ΔH = –196 kJ/mol
D	$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$	ΔH = –92 kJ/mol
E	$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$	ΔH = –11 kJ/mol

(i) Explain what is meant by the term dynamic equilibrium.

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(ii) State what effect, if any, the addition of a oxygen would have on the position of the equilibrium in reaction C.

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(iii) State what effect, if any, the removal of NH_3 would have on the position of the equilibrium in reaction D.

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(iv) State what effect, if any, increasing the temperature in reaction B would have on the position of the equilibrium.

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(v) State what effect, if any, increasing the pressure in reaction E would have on the position of the equilibrium.

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2 The equilibrium reaction below is carried out at 325 °C in the presence of a nickel catalyst.

$$CO(g) + 3H_2(g) \rightleftharpoons CH_4(g) + H_2O(g) \qquad \Delta H = -206 \text{ kJ/mol}$$

(a) State and explain how the equilibrium position would change if:
- The temperature of the reaction was increased
- The total pressure was increased

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(b) State the effect on the equilibrium position of removing water from the reaction mixture. Explain your answer.

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(c) Give one reason why, in practice, the reaction is carried out at a high temperature.

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