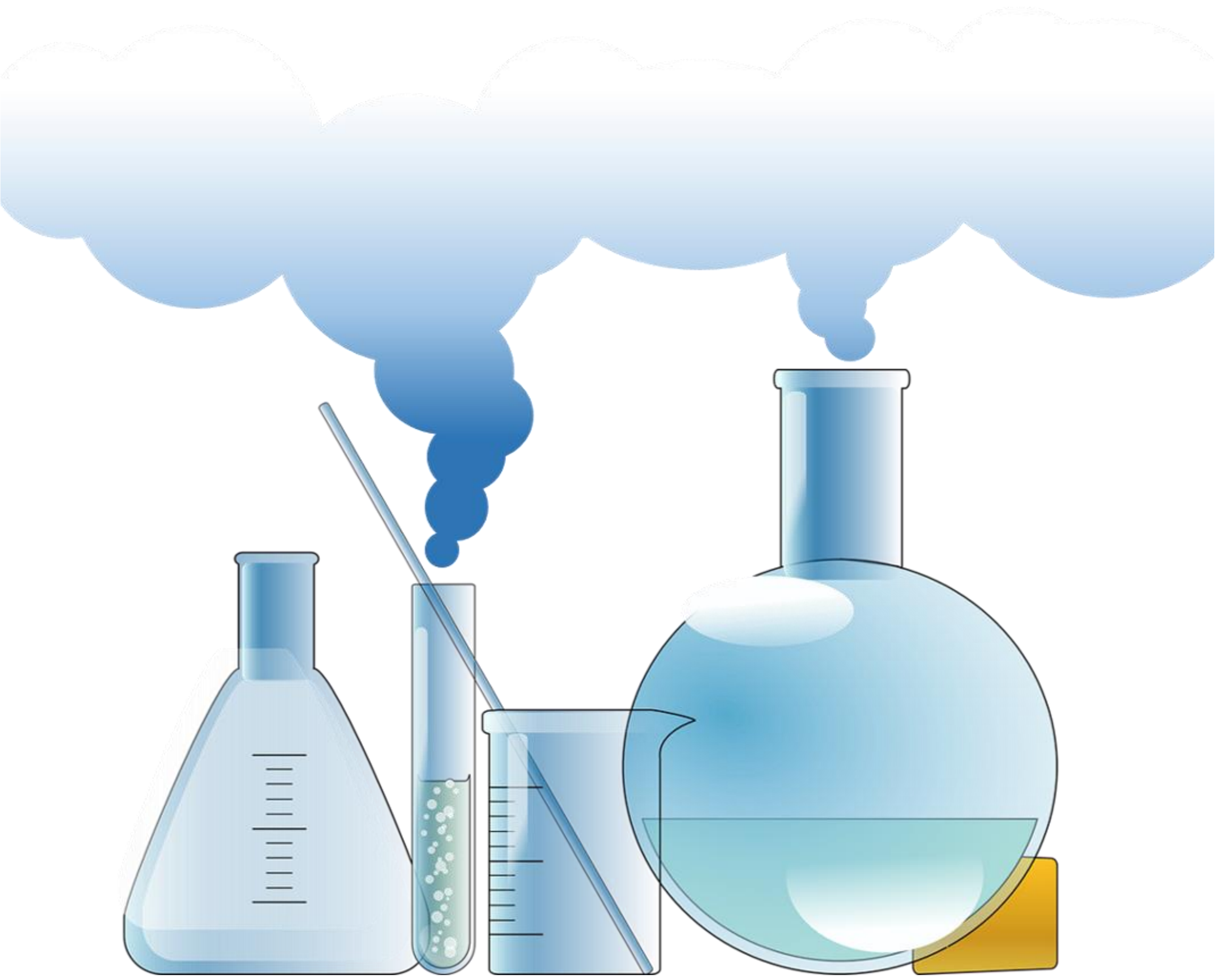
Chemical Equilibrium



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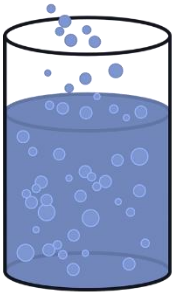
Understanding **Chemical Equilibrium**

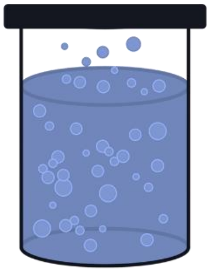
Open System



Matter and Energy
can leave the system

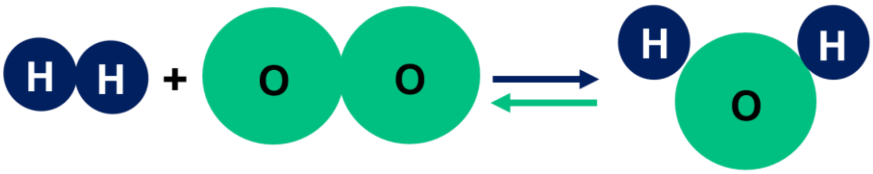
For a **closed system,** a **DYNAMIC CHEMICAL EQUILIBRIUM** will form when the rate of the **FORWARD** and **REVERSE** reactions are **equal**

Closed System



Matter and Energy
cannot leave the system

FORWARD reaction (reactants change into products)

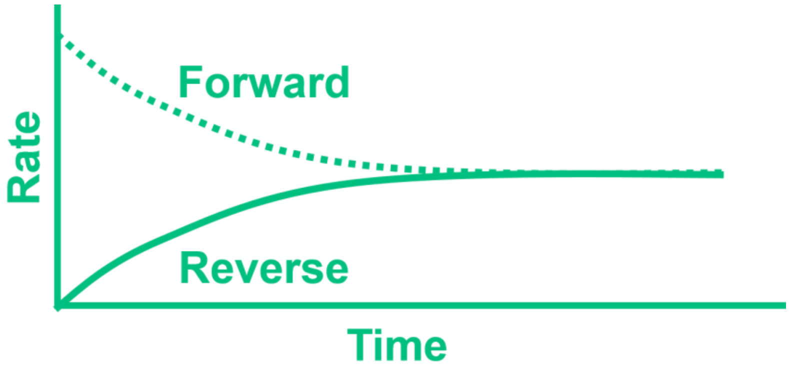


REVERSE reaction (products change into reactants)

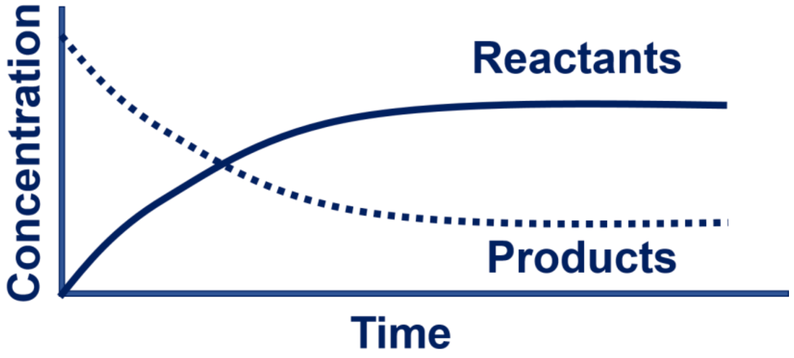
Many chemical reactions are REVERSIBLE

A system is in CHEMICAL EQUILIBRIUM when

The **RATE** of the forward reaction and the reverse reaction are **EQUAL**



The **CONCENTRATION** of the reactants and products are **CONSTANT**



Changes to the **Equilibrium Position**

Certain factors can **CHANGE** a system in equilibrium, affecting the **CONCENTRATION** of the reactants or products.

The change **DISRUPTS THE EQUILIBRIUM** causing the **RATE** of the forward or reverse reaction to either increase or decrease as the system comes to **A NEW EQUILIBRIUM**.

Factors Affecting Equilibrium

CHANGE THE CONCENTRATION
At constant volume and only applies to solutions (aq) and gases (g)

CHANGE THE TEMPERATURE

CHANGE IN PRESSURE
Only applies to gases (g)

Le Chatelier's Principle helps to predict how the system will respond to changes made to an equilibrium system.

Le Chatelier's Principle

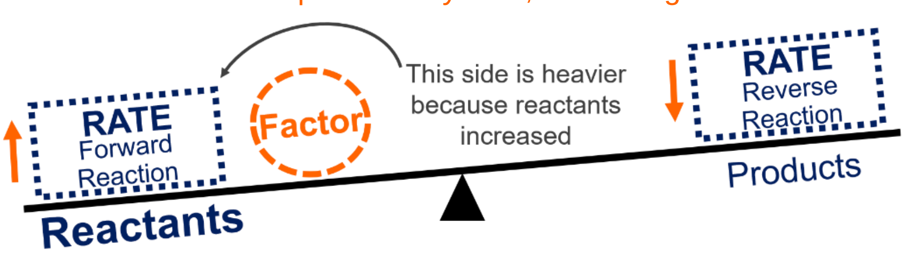
If a change is imposed on a system at equilibrium, the position of the equilibrium will shift in a direction that tends to reduce that change

Understanding **Le Chatelier's Principle**

Basically an equilibrium system will respond to reduce the effect of the change by favouring either the FORWARD or REVERSE reaction.

EXAMPLE

A factor causes a **CHANGE** to an equilibrium system, increasing the concentration of the reactants.



In **RESPONSE** the forward reaction is **FAVOURED,** decreasing the concentration of the reactants, restoring (a new) balance.

The Equilibrium Constant

The **EQUILIBRIUM CONSTANT (K_C)** is the ratio of products to reactants

It allows a chemist to determine the progress of a reaction, indicating whether the formation of the reactants or products is being favored.

For the chemical reaction at equilibrium

$$aA + bB \rightleftharpoons cC + dD$$

$$K_c = \frac{[\text{products}]}{[\text{reactants}]} = \frac{[C]^c.[D]^d}{[A]^a.[B]^b}$$

Phases of matter are IMPORTANT

Pure substances (liquid or solid) do not have a measurable concentration and are given a value of 1 in the K_c expression

HOMOGENOUS REACTIONS

Reactants and products are ALL in the SAME phase

$$aA(g) + bB(g) \rightleftharpoons cC(g) + dD(g)$$

HETEROGENEOUS REACTIONS

Reactants and products are NOT ALL in the SAME phase

$$aA(s) + bB(l) \rightleftharpoons cC(g) + dD(aq)$$

$$K_c = \frac{[\text{products}]}{[\text{reactants}]} = \frac{[C]^c.[D]^d}{[A]^a.[B]^b}$$

Only temperature can change K_c

The value of Kc does NOT change if a catalyst is added, concentration of reactants or products change or pressure of the system changes.

Understanding the VALUE of K_c

When **K_c = 1** : Concentration of reactants and products are about the same.

When **K_c > 1** : Concentration of products are higher than reactants.

When **K_c < 1** : Concentration of reactants are higher than products.

Ensure you know how to calculate the concentration and moles

$$c = \frac{n}{V} \quad \& \quad n = \frac{m}{M}$$

c – concentration ($mol \cdot dm^{-3}$) V – volume (dm^3)

n – number of moles (mol) M – molar mass ($g \cdot mol^{-1}$) m – mass (g)

Calculating K_C

IMPORTANT!

Ensure you are able to do the following THREE steps in a K_c calculation problem.

① Step-up your K_c expression

Check the phases of your reactants and products

Write your K_c expression

② Use an Equilibrium Table (R.I.C.E.E.)

Draw an Equilibrium Table.

	A (g)	B (g)	C (g)	D (aq)
Ratio	Mole ratio as per balanced chemical equation			
Initial	Starting amount (mol) of reagent present at the beginning			
Change	Change in reactants/products based on mole ratio			
Equilibrium	Amount of reactant/product present at equilibrium			
Equilibrium	Concentration of reactant/product at equilibrium			

③ Substitute into your K_c expression

Substitute into K_c equation from Step ①

WORKED Exam Question Paper 2, Oct/Nov 2019, Q.6

Initially 60,8 g pure carbon dioxide, $CO_2(g)$, is reacted with carbon, C(s), in a sealed container of volume 3 dm^3. The reaction reaches equilibrium at temperature **T** according to the following balanced equation:

$$C(s) + CO_2(g) \rightleftharpoons 2CO(g)$$

6.1 Define the term chemical equilibrium. (2)

The stage in a chemical reaction when the rate of forward reaction equals the rate of reverse reaction and the concentrations of reactants and products remain constant.

6.2 At equilibrium it is found that the concentration of the CO_2 is 0,054 $mol \cdot dm^{-3}$.

Calculate the:

6.2.1 Equilibrium constant, K_C, for this reaction at temperature **T** (7)

① $C(\mathbf{s}) + CO_2(\mathbf{g}) \rightleftharpoons 2CO(\mathbf{g})$ AND $\mathbf{K_c} = \frac{[CO]^2}{[CO_2]}$

② $n_{initial}(CO_2) = \frac{m}{M} = \frac{60.8}{44} = 1{,}382 \text{ mol}$

$n_{equil}(CO_2) = cV = 0.054 \times 3$
$= 0{,}162 \text{ mol}$

③ $\mathbf{K_c} = \frac{[CO]^2}{[CO_2]} = \frac{[\frac{2.44}{3}]^2}{[0.054]} = \frac{[0.813]^2}{[0.054]} = 12.24$

	$CO_2(\mathbf{g})$	$CO(\mathbf{g})$
Ratio	**1**	**2**
Initial (mol)	1.382	**0**
Change (mol)	1.22	2.44
Equilibrium (mol)	0.162	2.44

6.2.2 Minimum mass of C(s) that must be present in the container to obtain this equilibrium

The minimum mass would equal **the Change** (what reacted)

Mole ratio 1:1 for C(s) and CO_2 , Therefore $m = nM = n(C)_{reacted}M = 1{,}22(12) = 14{,}64 \text{ g}$

6.3 How will EACH of the following changes affect the AMOUNT of CO(g) at equilibrium?

Choose from INCREASES, DECREASES or REMAINS THE SAME.

6.3.1 More carbon is added to the container (1)

The carbon is in the solid phase. Only substances in the gaseous and aqueous phase can have a change in concentration and will be affected by a change in concentration, therefore the amount of CO(g) **REMAINS THE SAME**

6.3.2 The pressure is increased by reducing the volume of the container at constant temperature. Use Le Chatelier's principle to explain the answer. (3)

A change in pressure only applies to reactions with substances in the gaseous phase.

Pressure increases, therefore the system will try and reduce the pressure by favouring the reaction which has smaller volume of gas (lowest number of molecules).

The reverse reaction is favoured (1 gaseous molecule), therefore the amount of CO(g) **DECREASES**

6.4 The table below shows the percentages of CO_2(g) and CO(g) in the container at different temperatures.

TEMPERATURE (°C)	% CO_2(g)	% CO(g)
827	6,23	93,77
950	1,32	98,68
1 050	0,37	99,63
1 200	0,06	99,94

6.4.1 Is the reaction EXOTHERMIC or ENDOTHERMIC?
Refer to the data in the table and explain the answer. (3)

The reaction is ENDOTHERMIC since an increase in temperature favours the forward reaction (product yield increases).

An increase in temperature favours the endothermic reaction

6.4.2 Use the information in the table to determine temperature **T**.
Show clearly how you arrived at the answer. (3)

To calculate % of CO_2(g) and % of CO(g), we need to know the total volume first

$V_{TOTAL\ Equil} = 0{,}162 + 2{,}44 = 2{,}602\ dm^3$

% CO(g) = $\frac{2{,}44}{2{,}602} \times 100 = 93{,}77\%$

Temperature is 827 ^{0}C

Past Exam Question — Paper 2, May/June 2019, Q.5

Learners use the reaction of a sodium thiosulphate solution with dilute hydrochloric acid to investigate several factors that affect the rate of a chemical reaction.
The balanced equation for the reaction is:

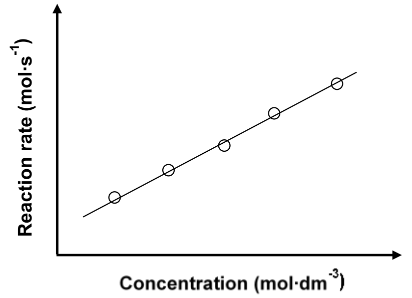
$$Na_2S_2O_3(aq) + 2HC\ell(aq) \rightarrow 2NaC\ell(aq) + SO_2(g) + S(s) + H_2O(\ell)$$

5.1 Define reaction rate. (2)
Three investigations (I, II and III) are carried out.

5.2 **INVESTIGATION I**
The results obtained in INVESTIGATION I are shown in the graph.

For this investigation, write down the:
5.2.1 Dependent variable (1)
5.2.2 Conclusion that can be drawn from the results (2)

GRAPH OF REACTION RATE VERSUS CONCENTRATION OF $Na_2S_2O_3$(aq)



The balanced equation below represents the reaction used in the Haber process to produce ammonia.

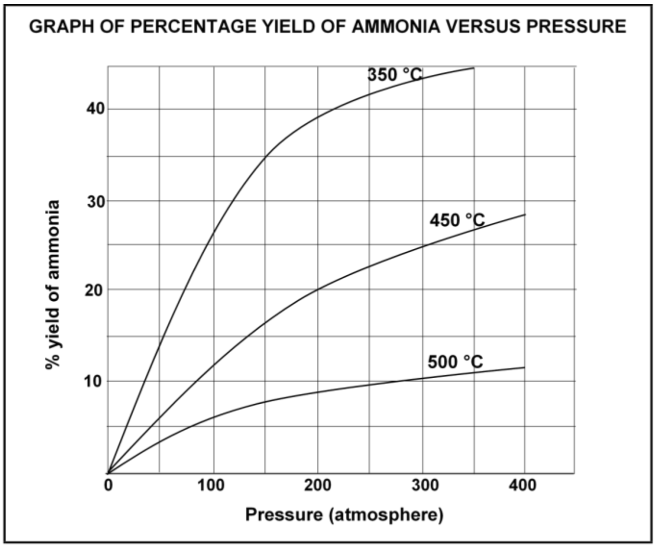
$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \qquad \Delta H < 0$$

In industry the product is removed as quickly as it forms.

6.1 Write down the meaning of the double arrow used in the equation above. (1)

6.2 Give ONE reason why ammonia is removed from the reaction vessel as quickly as it forms. (1)

The graph below shows the percentage yield of ammonia at different temperatures and pressures.



6.3 Write down the percentage yield of ammonia at 450 °C and 200 atmospheres. (1)

6.4 Refer to Le Chatelier's principle to explain EACH of the following deductions made from the graph:

6.4.1 For a given pressure, the yield of ammonia at 500 °C is much lower than that at 350 °C (3)

6.4.2 For a given temperature, the yield of ammonia at 350 atmospheres is much higher than that at 150 atmospheres (2)

6.5 A technician prepares $NH_3(g)$ by reacting 6 moles of $H_2(g)$ and 6 moles of $N_2(g)$.

6.5.1 Calculate the maximum number of moles of $NH_3(g)$ that can be obtained in this reaction. (2)

6.5.2 The above reaction now takes place in a 500 cm^3 container at a temperature of 350 °C and a pressure of 150 atmospheres. The system is allowed to reach equilibrium.

Use the graph above and calculate the equilibrium constant, K_c, for this reaction under these conditions. (7)

Past Exam Question — Paper 2, Oct/Nov 2018, Q.6

Dinitrogen tetraoxide, N_2O_4(g), decomposes to nitrogen dioxide, NO_2(g), in a sealed syringe of volume 2 dm^3.

The mixture reaches equilibrium at 325 °C according to the balanced equation:

$$N_2O_4(g) \rightleftharpoons 2NO_2(g)$$

Colourless — brown

When equilibrium is reached, it is observed that the colour of the gas in the syringe is brown.

6.1 State Le Chatelier's principle. (2)

6.2 The syringe is now dipped into a beaker of ice water. After a while the brown colour disappears. Is the forward reaction EXOTHERMIC or ENDOTHERMIC? Explain the answer using Le Chatelier's principle. (3)

6.3 The volume of the syringe is now decreased while the temperature is kept constant. How will EACH of the following be affected? Choose from: INCREASES, DECREASES or REMAINS THE SAME.

6.3.1 The number of moles of N_2O_4(g) (1)
6.3.2 The value of the equilibrium constant (1)
6.3.3 The rate of the forward and reverse reactions (1)

6.4 Initially **X** moles of N_2O_4(g) were placed in the syringe of volume 2 dm^3. At equilibrium, it was found that 20% of the N_2O_4(g) had decomposed.

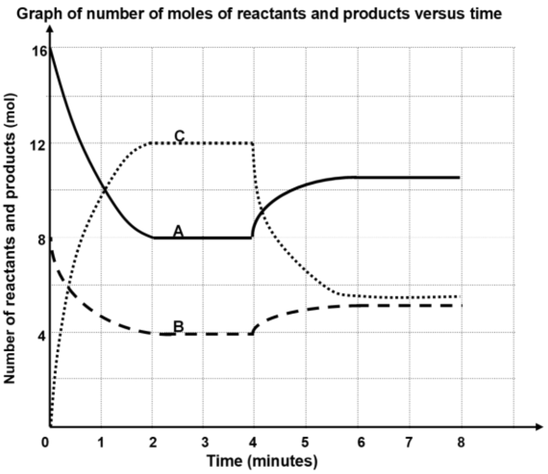
If the equilibrium constant, K_c, for the reaction is 0,16 at 325 °C, calculate the value of **X**. (8)

Past Exam Question — Paper 2, May/June 2018, Q.6

The equation below represents a hypothetical reaction that reaches equilibrium in a closed container after 2 minutes at room temperature. The letters x, y and z represent the number of moles in the balanced equation.

$$\mathbf{x}A(aq) + \mathbf{y}B(aq) \rightleftharpoons \mathbf{z}C(aq)$$

The graph below shows the change in the number of moles of reactants and products versus time during the reaction.



6.1 Define a dynamic equilibrium. (2)

6.2 Use the information in the graph and write down the value of:

6.2.1 x (1)
6.2.2 y (1)
6.2.3 z (1)

6.3 Calculate the equilibrium constant, K_c, for this hypothetical reaction at room temperature if the volume of the closed container is 3 dm^3. (7)

6.4 At t = 4 minutes, the temperature of the system was increased to 60°C. Is the REVERSE reaction EXOTHERMIC or ENDOTHERMIC? Explain how you arrived at the answer.

Hydrogen gas, $H_2(g)$, reacts with sulphur powder, $S(s)$, according to the following balanced equation:

$$H_2(g) + S(s) \rightleftharpoons H_2S(g) \qquad \Delta H < 0$$

The system reaches equilibrium at 90 °C.

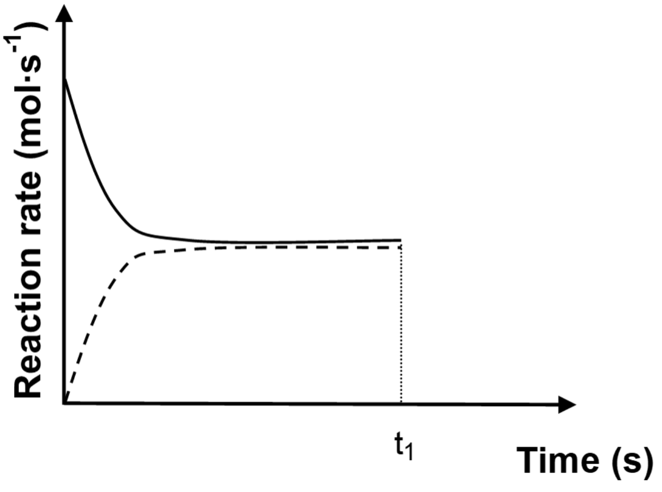
6.1 Define the term chemical equilibrium. (2)

6.2 How will EACH of the following changes affect the number of moles of $H_2S(g)$ at equilibrium? Choose from INCREASES, DECREASES or REMAINS THE SAME.

6.2.1 The addition of more sulphur (1)

6.2.2 An increase in temperature. Use Le Chatelier's principle to explain the answer. (4)

6.3 The sketch graph below was obtained for the equilibrium mixture.



A catalyst is added to the equilibrium mixture at time t_1.

Redraw the graph above in your ANSWER BOOK. On the same set of axes, complete the graph showing the effect of the catalyst on the reaction rates. (2)

Initially 0,16 mol $H_2(g)$ and excess $S(s)$ are sealed in a 2 dm^3 container and the system is allowed to reach equilibrium at 90 °C.

An exact amount of $Pb(NO_3)_2$ solution is now added to the container so that ALL the $H_2S(g)$ present in the container at EQUILIBRIUM is converted to $PbS(s)$ according to the following balanced equation:

$$Pb(NO_3)_2(aq) + H_2S(g) \rightarrow PbS(s) + 2HNO_3(aq)$$

The mass of the PbS precipitate is 2,39 g.

6.4 Calculate the equilibrium constant Kc for the reaction $H_2(g) + S(s) \rightleftharpoons H_2S(g)$ at 90 °C. (9)

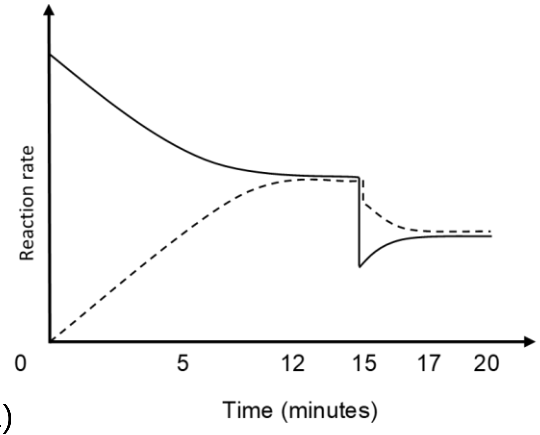
Past Exam Question

Paper 2, Feb/Mar 2015, Q.6

Pure hydrogen iodide, sealed in a 2 dm^3 container at 721 K, decomposes according to the following balanced equation:

$2HI(g) \rightleftharpoons H_2(g) + I_2(g) \qquad \Delta H = +26\ kJ\cdot mol^{-1}$

The graph below shows how reaction rate changes with time for this reversible reaction.



6.1 Write down the meaning of the term reversible reaction. (1)

6.2 How does the concentration of the reactant change between the 12th and the 15th minute? Write down only INCREASES, DECREASES or NO CHANGE. (1)

6.3 The rates of both the forward and the reverse reactions suddenly change at t = 15 minutes.

6.3.1 Give a reason for the sudden change in reaction rate. (1)

6.3.2 Fully explain how you arrived at the answer to QUESTION 6.3.1. (3)

The equilibrium constant (K_c) for the forward reaction is 0,02 at 721 K.

6.4 At equilibrium it is found that 0,04 mol HI(g) is present in the container. Calculate the concentration of $H_2(g)$ at equilibrium. (6)

6.5 Calculate the equilibrium constant for the reverse reaction. (1)

6.6 The temperature is now increased to 800 K. How will the value of the equilibrium constant (Kc) for the forward reaction change?
Write down only INCREASES, DECREASES or REMAINS THE SAME. (1)