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CHEMISTRY

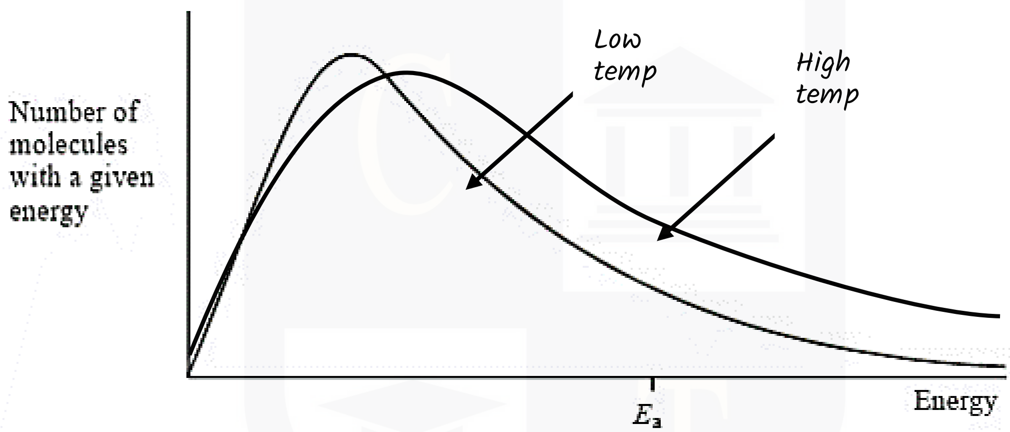
Physical Chemistry

Level & Board	AQA (A-LEVEL)
TOPIC:	KINETICS
PAPER TYPE:	SOLUTION - 2
TOTAL QUESTIONS	10
TOTAL MARKS	47

Kinetics - 2

1.

(a)



(3)

(b)

The rate of reaction increases at a higher temperature because at elevated temperatures, more molecules or particles possess the minimum energy needed to react. This minimum energy required for a successful reaction is termed as **activation energy.**

With increased temperature, a greater number of molecules or particles acquire sufficient kinetic energy, as per the Maxwell-Boltzmann distribution, allowing them to surpass the activation energy barrier. So, there's a higher frequency of successful collisions between reacting species, leading to an overall increase in the rate of the chemical reaction.

(3)

(c)

The rate of reaction increases because the catalyst lowers the activation energy required for the reaction to occur.

By lowering the activation energy, more molecules within the system are able to surpass the energy barrier necessary for the reaction to proceed. This reduction in activation energy enables a larger number of molecules to participate in successful collisions, thus increasing the rate of the chemical reaction.

(3)

2.

(a)

Activation energy is the minimum amount of energy required for a chemical reaction to occur. It represents the energy barrier that reactant molecules must overcome to transform into products during a chemical reaction.

(2)

(b)

The reaction between nitrogen dioxide and carbon monoxide is very slow at room temperature because **few molecules / particles have the required activation energy.**

(1)

(c)

An increase in pressure causes molecules to be closer together within the same volume.

As a result of this decreased space between molecules, they collide more frequently.

The closer distance of molecules due to higher pressure results in a higher collision frequency. This increased collision frequency is a key factor contributing to the accelerated rate of the reaction between nitrogen dioxide and carbon monoxide.

(2)

(d)

When the temperature is increased, more molecules within the reaction mixture gain kinetic energy, resulting in a greater proportion of molecules possessing energies equal to or greater than the activation energy (E_{Act}) necessary for the reaction to occur.

This higher population of molecules with sufficient energy to overcome the activation barrier greatly increases the likelihood of successful collisions between nitrogen dioxide (NO_2) and carbon monoxide (CO), leading to a substantial increase in the rate of the reaction.

(2)

(e)

A catalyst is a substance that accelerates a chemical reaction by providing an alternative reaction pathway, consequently lowering the activation energy required for the reaction to occur. Importantly, while facilitating the reaction, a catalyst remains chemically unchanged or unaltered at the end of the reaction.

It speeds up the reaction kinetics but does not get consumed or transformed into different substances during the reaction process. Therefore, a catalyst plays a pivotal role in expediting reactions without being consumed itself, allowing it to be reused in subsequent reactions.

(1)

(f)

A solid catalyst for a gas-phase reaction is often in powder form because it offers a larger surface area, enabling increased reactivity, better mixing, and improved kinetics due to greater accessibility of active sites for gas-phase reactants.

(1)

3.

Activation energy is the minimum amount of energy required for a chemical reaction to occur. It represents the energy barrier that reactant molecules must overcome to transform into products during a chemical reaction.
In a reaction, reactant molecules possess a certain amount of energy. However, for these molecules to undergo a chemical change and form products, they need to acquire an additional amount of energy to break the existing chemical bonds and initiate the reaction. ***This extra energy required to start the reaction and facilitate the conversion of reactants into products is referred to as activation energy.***

(2)

4.

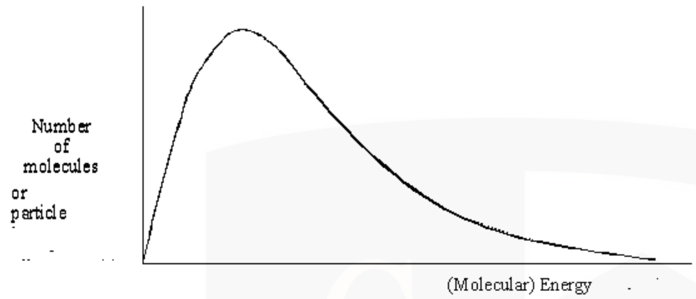
Measurement Procedure:
Volume of Gas or Mass Loss: *Measure the volume of gas produced or the mass loss due to the reaction at regular time intervals.*
Concentration Changes (if applicable): *Use colorimetry or spectrophotometry to track concentration changes.*

Time Intervals: *Record data points regularly and plot against time to determine the reaction rate.*

(2)

5.

(a) Labeled axes on the diagram are:



(1)

(b)
The area under the curve represents total number of particles (or molecules) present in the sample.

(1)

(c)
This curve starts at the origin because no molecules have no energy.

(2)

6.
(a)
In the context of a dynamic equilibrium, the term "dynamic" means that both the forward and reverse reactions occur simultaneously or at the same time. This means that both reactions conversion of reactants to products and the conversion of products back to reactants are ongoing concurrently, yet the overall concentrations of substances involved remain constant at equilibrium.

(1)

(b)
i.
Effect on Equilibrium Position:
No effect: When the total pressure is increased, there is no shift in the position of this equilibrium.
Explanation: The equilibrium position remains unchanged because there is an equal number of moles or molecules of gas on both sides of the reaction.

Therefore, the increase in total pressure does not cause the equilibrium to shift in a specific direction due to the equal number of gas molecules present on both sides of the reaction.

(2)

ii.

Effect on the Time Taken to Reach Equilibrium:

Less time / Equilibrium reached faster: *Increasing the total pressure results in a decrease in the time taken to reach equilibrium.*

Explanation: *The increase in total pressure causes more particles or molecules to occupy a given volume or space. As a result, the molecules are closer together, leading to an increased frequency of collisions between the reactants.*

These more frequent collisions occur with sufficient energy ($E>E_{Act}$) required to overcome the activation energy barrier, thereby resulting in a higher rate of successful or productive collisions. So, the increased collision frequency and more successful collisions lead to a faster attainment of equilibrium.

(3)

7. *(B)*

(Total 1 mark)

8.

Adding a catalyst decreases the time required for gas G to decompose compared to a similar sample without a catalyst.

This ***decrease*** *in time is due to the catalyst lowering the activation energy (E_a) needed for the decomposition reaction.*

The catalyst provides an alternative reaction pathway with a lower activation energy, allowing a larger number of gas G molecules to have sufficient energy to surpass the lowered activation energy barrier. Consequently, more molecules can effectively participate in the reaction, expediting the decomposition process.

(3)

9.

(a)

According to Le Chatelier's principle

Yield: The yield of hydrogen (H_2) is increased.

Reaction Endothermic: The reaction is endothermic.

Equilibrium Shift: The equilibrium shifts to the right or forward to counteract the increase in temperature by absorbing heat.

Increase in Pressure:

Yield: The yield of hydrogen (H_2) is decreased.

Increase in Moles of Gas: The increase in pressure will favor the side with fewer moles of gas, which is the left side with fewer gas molecules (2 moles - CH_4 and H_2O) compared to the right side (4 moles - CO and H_2), thus decreasing the yield of hydrogen.

Equilibrium Shift: The equilibrium shifts to the left or backward to reduce pressure by favoring the side with fewer gas molecules.

(6)

(b)

Equilibrium Yield of Hydrogen:

Unaffected Equilibrium Yield: The use of a catalyst does not change the equilibrium yield of hydrogen.

Explanation: The catalyst increases the rate or speed of both the forward and backward reactions equally. As a result, while the rate of reaching equilibrium is enhanced, the actual position of equilibrium remains the same. Therefore, the equilibrium yield of hydrogen is unaffected by the catalyst.

Amount of Hydrogen Produced in Time:

Increased Production of Hydrogen: The use of a catalyst results in more hydrogen being produced.

Explanation: The catalyst enhances the rate of the reaction, leading to a faster production of hydrogen. However, this increase in the rate of production only affects the speed at which hydrogen is formed but does not alter the total amount of hydrogen present at equilibrium. Thus, while more hydrogen is produced in a given time, the total amount of hydrogen at equilibrium remains unchanged.

(4)

10. (D)

(Total 1 mark)

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