

## **States of matter and Kinetic Theory**

### **State the properties of matter (s, l and g)**

The three states of matter excluding plasma and Bose Einstein Condensate solids (s), liquids (l) and gases (g)

\*Refer to chapter 1 for the properties of each state of matter\*

### **Describe the kinetic theory in terms of the movement of particles whose average energy is proportional to temperature**

The kinetic theory describes the behavior of particles in matter, stating that particles are always in motion. According to this theory, the average energy of these particles is directly proportional to the temperature. As the temperature increases, the particles move faster because they have more energy. Conversely, as the temperature decreases, the particles move slower because they have less energy.

This relationship explains how temperature affects the movement and energy of particles in solids, liquids, and gases.

## **Collision Theory**

### **Describe the collision theory**

The collision theory states that in a gaseous phase, the chemical reaction rate is based on the kinetic energy of the particles. Furthermore, collision theory involves how the orientation of the collisions between particles also determines whether a reaction takes place or not. In order for a chemical reaction to occur, the reactant molecules have to collide with sufficient energy and the correct orientation.

### **Explain collision theory taking into account the qualitative effects of particles whose average energy is proportional to temperature**

Collision theory explains that chemical reactions occur when reacting particles collide with sufficient energy and proper orientation. The average energy of these particles is directly proportional to the temperature. As temperature increases, particles move faster and have higher energy. This higher energy increases the frequency and intensity of collisions between particles.

Qualitatively, when the temperature rises, more particles possess the required activation energy to undergo successful reactions. This results in a higher number of effective collisions, thereby increasing the reaction rate.

## **Equilibrium**

### **Define the thermal dissociation, reversible reaction and thermal decomposition**

Thermal dissociation: a thermal decomposition reaction which is easily reversible because usually only a single bond is broken when increasing the temperature

Reversible reactions: reactions where the reactants react together to form the products, and the product react together to form the reactants

Thermal decomposition: the chemical decomposition of a substance into several other substances at a high temperature. Thermal decomposition involves rearranging and breaking many bonds, making it not so easily reversible

### **Outline the characteristics of chemical and physical systems in a state of equilibrium**

Chemical system:

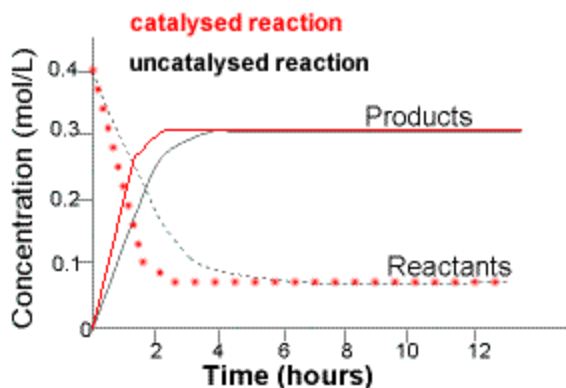
- In a chemical equilibrium, reactions still occur, but the forward and reverse reactions happen at the same rate, so the amounts of reactants and products stay constant.
- The concentrations of reactants and products remain stable because the rates of the forward and reverse reactions are equal.

Physical system:

- Physical equilibrium involves a balanced state where processes such as phase changes occur at equal rates, maintaining stability over time.
- Despite ongoing phase changes, properties like pressure, temperature, and phase proportions remain constant in the system.

### **State and explain the effect of a catalyst on an equilibrium reaction**

When a catalyst is used in a reaction, the position of equilibrium remains unchanged. However, it increases the rates of the backward and forward reactions, bringing a closed system to equilibrium more quickly



Apply Le Chatelier's principle to predict the qualitative effects of changes of temperature, pressure and concentration on the positions of equilibrium

## Le Chatelier's Principle

Disturbing a system at dynamic equilibrium shifts the equilibrium in the direction that counteracts the change.



### Concentration

↑ reactant concentration  
 ↑ favors product formation

↑ product concentration  
 ↑ favors reactant formation

### Temperature

↑ temperature  
 favors endothermic reaction

↓ temperature  
 favors exothermic reaction

### Pressure

↑ pressure  
 favors side with fewer molecules

↓ pressure  
 favors side with more molecules

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Factor	Effect on Equilibrium Position	Example
Temperature	Increase: Shifts equilibrium in the endothermic direction (heat-absorbing). Decrease: Shifts equilibrium in the exothermic direction (heat-releasing).	Example: $A + B \rightleftharpoons C + \text{heat}$ Increasing temperature shifts equilibrium towards the right ( $C + \text{heat}$ ). Decreasing temperature shifts equilibrium towards the left ( $A + B$ ).

Pressure	Increase: Shifts equilibrium towards the side with fewer gas moles. Decrease: Shifts equilibrium towards the side with more gas moles.	Example: $A(g) + B(g) \rightleftharpoons C(g)$ Increasing pressure shifts equilibrium towards the left ( $A(g) + B(g)$ ). Decreasing pressure shifts equilibrium towards the right ( $C(g)$ ).
Concentration	Increase: Shifts equilibrium away from the added component (opposite to reactant or product). Decrease: Shifts equilibrium towards the added component (same as reactant or product).	Example: $A + B \rightleftharpoons C$ Increasing concentration of A shifts equilibrium towards the right (C). Decreasing concentration of A shifts equilibrium towards the left (A + B).

**Apply the concepts of equilibrium to the Haber processes and use Le Chatelier to create the maximum yield of Ammonia**

The Haber process is a crucial industrial reaction for producing ammonia ( $NH_3$ ) from nitrogen ( $N_2$ ) and hydrogen ( $H_2$ ). It's represented by the equation:

$N_2(g) + 3H_2(g) \rightarrow 2NH_3$  To maximize the yield of ammonia, we can apply Le Chatelier's Principle. Factors affecting the yield of ammonia in the Haber process:

1. Concentration of Reactants: Increasing the concentration of nitrogen ( $N_2$ ) and hydrogen ( $H_2$ ) will shift the equilibrium towards the product side (ammonia,  $NH_3$ ). According to Le Chatelier's Principle, the equilibrium will shift to consume the added reactants, thus increasing the yield of ammonia.
2. Temperature: The forward reaction (formation of ammonia) is exothermic, meaning it releases heat. Therefore, lower temperatures favor the formation of ammonia. However, the reaction rate decreases at lower temperatures, so a compromise temperature is chosen in industry (around 400-500°C) to achieve a reasonable yield without sacrificing reaction rate excessively.
3. Pressure: Increasing the pressure favors the side with fewer moles of gas. In this case, the forward reaction (formation of ammonia) reduces the number of moles of gas (4 moles on the left side and 2 moles on the right side), so increasing pressure will favor the formation of ammonia. Therefore, higher pressure is used in the industrial Haber process to maximize the yield of ammonia.

**Reaction Rate**

### **Define the term rate of reaction**

The rate of reaction is the speed at which a reaction takes place. The formula for rate of reaction is:  $\text{ROR} = \text{Volume of product} / \text{Time}$

### **Describe suitable experimental procedures for measuring rates of reaction**

Surface area: React marble (calcium carbonate) stones of different sizes (powder, small pieces, large pieces) with hydrochloric acid (HCl). Measure the amount of carbon dioxide gas produced over a fixed period of time using a gas syringe. Compare the results to see how surface area affects the reaction rate.

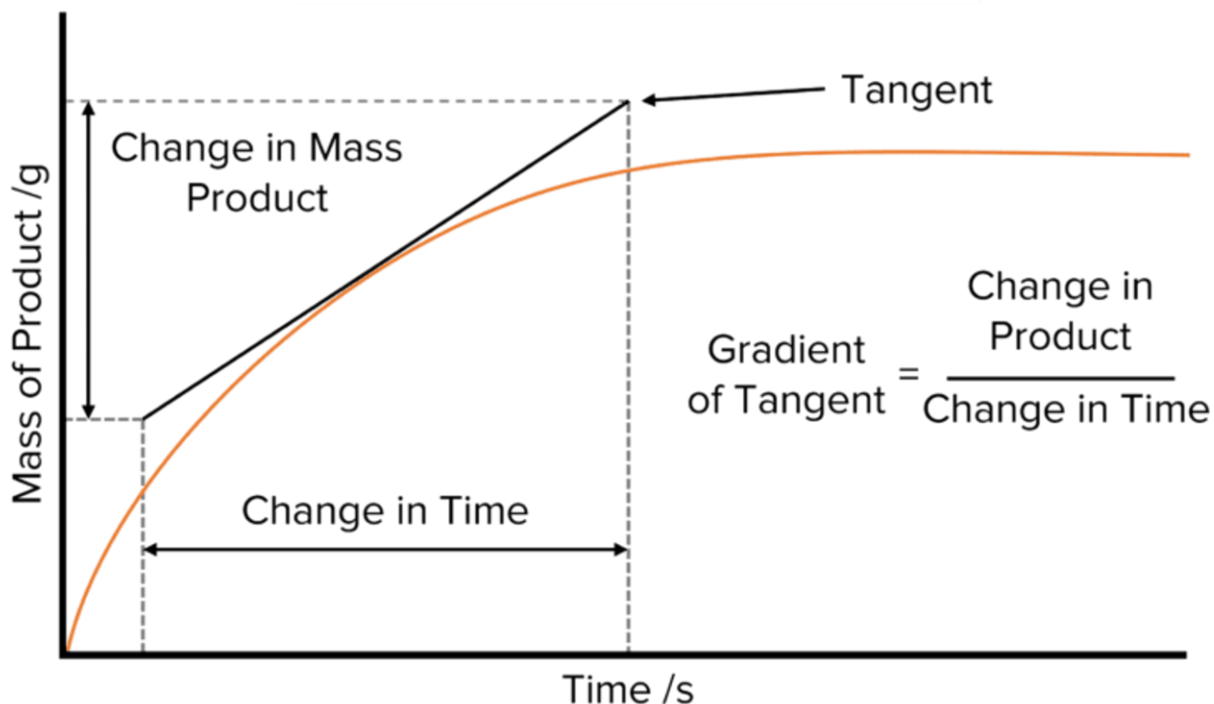
Temperature: React sodium thiosulphate solution with hydrochloric acid (HCl) at different temperatures. Observe the time taken for a thick, white precipitate to form, obscuring a cross drawn on a sheet of paper placed beneath the reaction vessel. Compare the times to see how temperature affects the reaction rate.

Concentration: React a fixed length of magnesium ribbon with hydrochloric acid (HCl) solutions of different concentrations. Measure the amount of hydrogen gas produced over a fixed period of time using a gas syringe. Compare the results to see how concentration affects the reaction rate.

### **Analyse data from rates experiments**

The average rate of reaction can be calculated by dividing the total volume of product collected and dividing it by the time taken. Alternatively, to calculate the reaction rate at a particular time, a tangent must be drawn from which the gradient will be the rate of reaction.

## Change in Product Mass Over Time



### **Explain factors which affect reaction rate: effects of particle size, temperature, concentration and pressure and presence of a catalyst**

1. Particle Size: Imagine reactants bumping into each other like billiard balls. The more surface area a reactant particle has, the more opportunities there are for collisions. Grinding a solid reactant into a powder increases its surface area, significantly speeding up the reaction.
2. Temperature: Think of temperature as the average kinetic energy (movement) of particles. As temperature rises, particles move faster and collide more frequently, leading to a faster reaction rate.
3. Concentration: This relates to how much reactant is packed into a given space. According to collision theory, a higher concentration of reactants means more frequent collisions and a faster reaction. Imagine a crowded room versus an empty one - collisions are bound to happen more often in the crowded space.
4. Pressure (for gases only): For gaseous reactions, pressure plays a role similar to concentration. Increased pressure squeezes the gas molecules closer together, leading to more frequent collisions and a faster reaction rate.

5. Presence of a Catalyst: A catalyst is like a matchmaker for reactions. It participates in the reaction without being consumed, lowering the activation energy (the minimum energy required for a reaction to occur) and speeding up the process. Imagine the catalyst weakening the bonds between the reactants, making it easier for them to break apart and form new products.

### **List and describe the limitations of each rate of reaction experiment**

Temperature:

Control Precision: Maintaining a constant temperature throughout the experiment can be challenging, leading to inconsistent results.

Thermal Decomposition: Some reactants may decompose at higher temperatures, affecting the reaction rate and product yield.

Heat Loss: Heat loss to the surroundings can alter the reaction temperature, leading to inaccurate measurements.

Concentration:

Measurement Accuracy: Precisely measuring and maintaining reactant concentrations can be difficult, affecting reproducibility.

Dilution Effects: Dilution errors can occur during the preparation of solutions, impacting the reaction rate.

Side Reactions: Higher concentrations may promote side reactions, complicating the interpretation of the results.

Surface Area:

Uniformity: Achieving and maintaining a consistent surface area of solid reactants can be difficult, leading to variable results.

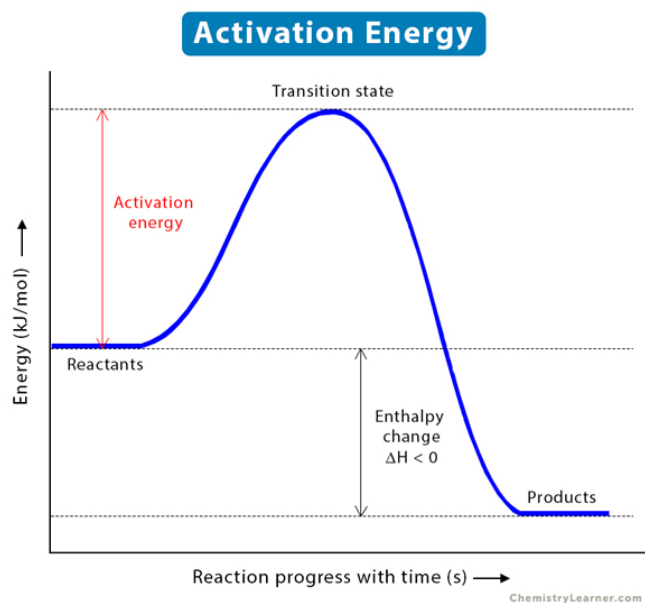
Measurement Difficulties: Quantifying the exact surface area exposed to the reaction can be challenging.

Physical Changes: Changes in particle size or shape during the reaction can alter the surface area, affecting the reaction rate.

### **Catalysts**

## Activation energy

Activation energy ( $E_a$ ) is the minimum energy required for a reaction to occur. It varies uniquely for each reaction based on the involved reactants. This energy difference is calculated as the gap between the highest point on the energy curve and the energy of the reactants.



## Effect of Catalysts

Catalysts are substances utilized to alter the reaction rate without being consumed. They provide an alternative pathway for the reaction to proceed, often with a lower activation energy compared to the uncatalyzed reaction. By employing catalysts, the activation energy is decreased, consequently reducing the reaction time. This occurs because the alternative reaction pathway provided by the catalyst has a lower activation energy compared to the original, uncatalyzed reaction.



