Ministry of Higher Education and Scientific Research
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Department of Biochemistry

Physical Chemistry

Lecture 5

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First semester

Collision theory & Reaction Rates

By

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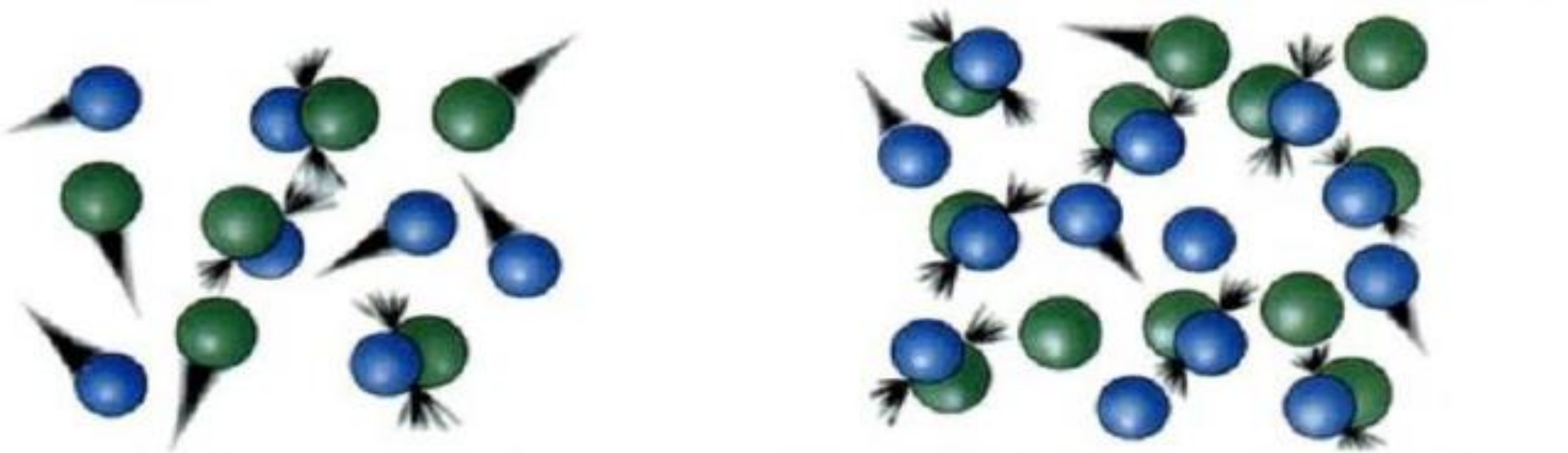
Collision Theory & Reaction Rates

Chemistry 30

Unit 3: Reaction Rates and Equilibrium

What is the collision theory?

- **Collision Theory** – the reacting particles must collide with one another for a reaction to take place.
  - It helps to explain why some reactions occur at different rates than others.
  - It also can be used to find out why a reaction is occurring slowly or quickly and what to change.



Low concentration = Few collisions

High concentration = More collisions

What makes a reaction happen?

- However, just because two particles collide does not mean that they will react.

- There are certain requirements that need to be met in order for a reaction to take place.
- The particles must collide:
  - With sufficient energy
  - With the correct orientation

What makes a reaction happen?

- Factors that affect the rate of a reaction:
  - Speed of the particles
    - which will give them more energy
  - Orientation of the particles
    - which will line them up properly to react
  - Frequency of collisions
    - The more chances of hitting each other, the more chances of achieving the other two factors

- The **rate of reaction** is determined by the frequency of the collisions

What is a Reaction Rate?

- Reaction rates refer to the speed at which a reaction occurs.
- Since our bodies and daily activities rely on chemical reactions, it is essential that we know why some reactions occur quickly and others very slowly.

Determining Reaction Rate

- **How do we measure the reaction rate?**
- One way would be to measure the rate at which the reactant disappears, or how fast the products are produced.
- Another way would be to measure how the concentration of a participant in the reaction changes
  - We will be focusing on <u>concentration of reactants and products</u> in this unit.

How is rate calculated?

- **Rate:** an average measure of how something changes over time

- Rate of reaction: change in moles/litre

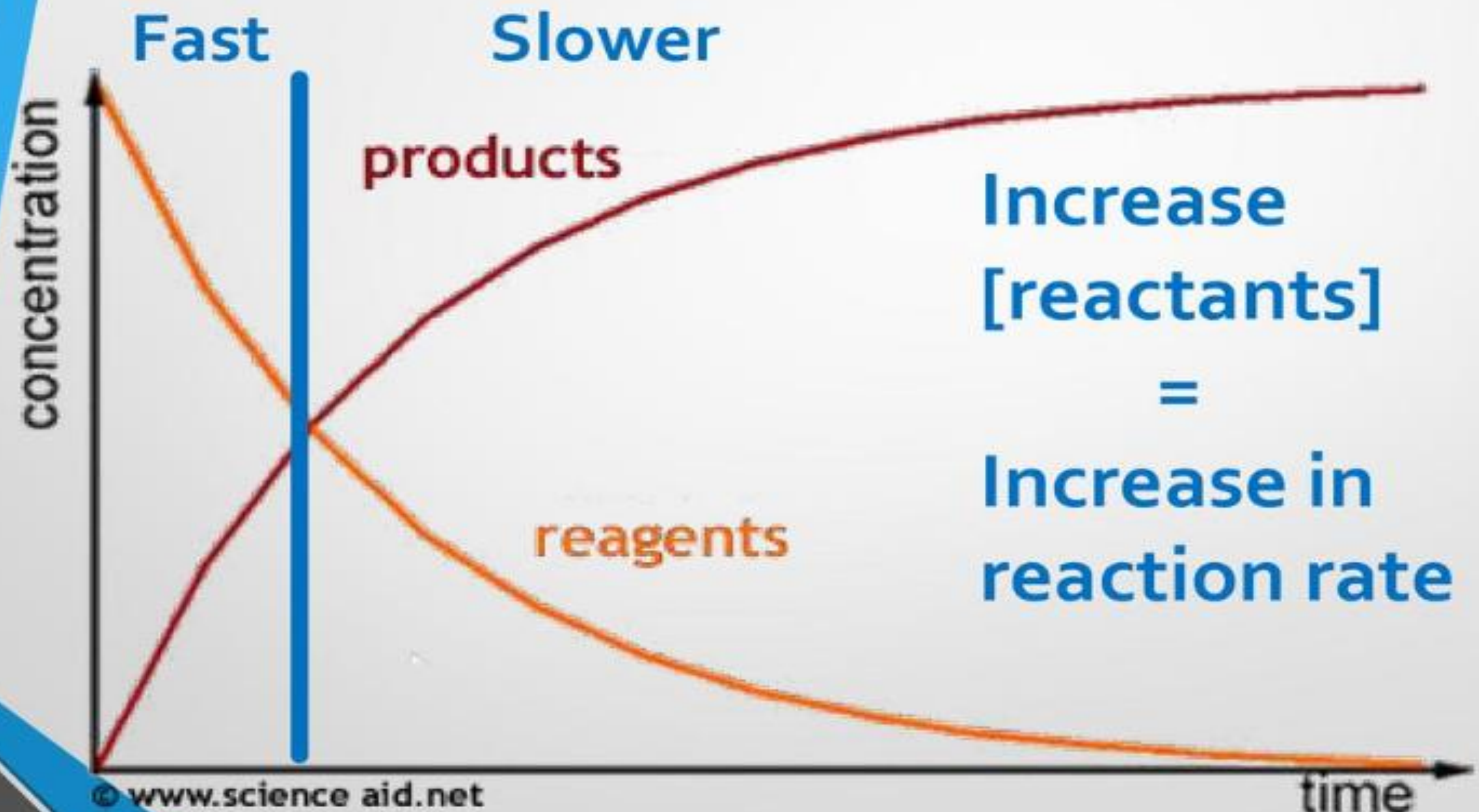
  change in time (sec)

- Rate = Δ M / Δ sec

When during a reaction will the rate be the quickest?

- When there are <u>more reactants</u>, the forward reaction can take place more easily and so the <u>rate is quicker</u>.
- Likewise, when the reaction has already started to take place and the <u>concentration of reactants has decreased</u>, the <u>rate will be slower</u>.
- So, what does that look like in a graph?

Reaction Rates Graphed



Increase [reactants]

=

Increase in reaction rate

Why is this so?

- The more particles present, the quicker they will collide and react.
- If you have fewer particles, it will take longer for each reactant to find each other and react.
- So....the rate at the start of a reaction is going to be different from the rate at the end.
- That is why we typically talk about the average rate of reaction.

Time and Rate

- Since rate is a product of time, you need to look at rates occurring at the same time in the reaction.
- For example, look at the rates in A → B:

Time (sec)	A (mol/L)	B (mol/L)
0.00	1.000	0.000
3.00	0.400	0.600
6.00	0.250	0.750

- What is the average rate of reaction for substance A?

Example 1

- **What is the average rate of reaction for <u>substance A</u>?**

Time (sec)	A (mol/L)	B (mol/L)
0.00	1.000	0.000
3.00	0.400	0.600
6.00	0.250	0.750

Use the equation:

rate = Δ M / Δ sec

$$\text{Rate of A} = \frac{(0.250\ \text{M} - 1.000\ \text{M})}{(6.00\ \text{sec} - 0.00\ \text{sec})} = \frac{-0.750\ \text{M}}{6.00\ \text{sec}}$$

Rate of A = **- 0.125 M/s**

(so A is decreasing at a rate of 0.125 M/s)

Example 1

- **Compare this value to the rate of reaction for the first 3 seconds:**

Time (sec)	A (mol/L)	B (mol/L)
0.00	1.000	0.000
3.00	0.400	0.600
6.00	0.250	0.750

Use the equation:

rate = Δ M / Δ sec

$$\text{Rate of A} = \frac{(0.400\text{ M} - 1.000\text{ M})}{(3.00\text{ sec} - 0.00\text{ sec})} = \frac{-0.600\text{ M}}{3.00\text{ sec}}$$

Rate of A = **- 0.200 M/s**

(so A is decreasing at a rate of 0.200 M/s)

Example 1

Time (sec)	A (mol/L)	B (mol/L)
0.00	1.000	0.000
3.00	0.400	0.600
6.00	0.250	0.750

Averages of both rates of reactions?

Rate of A = **- 0.125 M/s**

Rate of B = **0.125 M/s**

Notice they are the <u>same</u> because 1 A $\rightarrow$ 1 B

This is only true because our balanced equation shows us a 1:1 molar ratio between A & B

Example 2

- We would not find the same rates if we did not have a 1:1 relationship between reaction participants.

- For example if we examine the following reaction:

$$2\,H_2O_2 \rightarrow 2\,H_2O + O_2$$

- We find that only one mole of oxygen forms for every two moles of hydrogen peroxide that decomposes.
- Therefore, we could make the following relationships:

Example 2

$2\ H_2O_2$	→	$2\ H_2O$	+	O_2
If 2 moles of H_2O_2 decompose				1 mole of O_2 is formed
if the rate of decomposition of H_2O_2 is $4.00\ mol \cdot L^{-1} \cdot min^{-1}$				the rate of formation of O_2 is ½ × 4.00 or **$2.00\ mol \cdot L^{-1} \cdot min^{-1}$**

Example 3

- $N_2(g) + 3\ H_2(g) \longrightarrow 2\ NH_3(g)$

If the rate of loss of hydrogen gas is 0.03 M/s, what is the rate of production of ammonia?

$$\text{Rate of } NH_3 = 0.03\ \text{M/s}\ H_2 * \frac{2\ \text{mol}\ NH_3}{3\ \text{mol}\ H_2} = \mathbf{0.02M/s}$$

Example 4

- In the following decomposition reaction,

$$2\,N_2O_5 \rightarrow 4\,NO_2 + O_2$$

- Oxygen gas is produced at the average rate of 9.1×10^{-4} mol · L^{-1}· s^{-1}. Over the same period, what is the average rate of the following:
  - A) the production of nitrogen dioxide
  - B) the loss of nitrogen pentoxide


Thank you