MARITIME UNIVERSITY OF SZCZECIN

Institute of Mathematics, Physics and Chemistry
Department of Chemistry

EXERCISE INSTRUCTION

Laboratory Exercise 6

The reaction rate and catalysis

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EXERCISE SHEET

1	**Relation to subjects**: ESO/26, DiRMiUO/26, EOUNIE/26		
	Specialty/Subject	**Learning outcomes for the subject**	**Detailed learning outcomes for the subject**
	ESO/25 Technical chemistry	EKP1 K_W01,K_W02, K_U05 EKP2 K_U08, K_U09	SEKP7 – Mastering the experimental knowledge in studying the rate of chemical reactions and drawing conclusions from the conducted experiments. Mastering the knowledge of the mechanism and type of catalysis as well as identifying catalysts and inhibitors. Expanding knowledge on the use of catalysts in technology and environmental protection.
	ESO/26 Chemistry of water, fuels and lubricants	EKP3 K_U014, K_U015, K_U016	SEKP6 – Determination of oxygen and ammoniacal nitrogen content in technical water. SEKP6 – Determination of corrosion inhibitors in technical water.
	DiRMiUO/26 Chemistry of water, fuels and lubricants	EKP3 K_U014, K_U015, K_U016	SEKP6 – Determination of oxygen and ammoniacal nitrogen content in technical water SEKP6 – Determination of corrosion inhibitors in technical water.
	EOUNIE/26 Chemistry of water, fuels and lubricants	EKP3 K_U014, K_U015, K_U016	SEKP6 – Determination of oxygen and ammoniacal nitrogen content in technical water SEKP6 – Determination of corrosion inhibitors in technical water
2	**Purpose of the exercise:** 1. Learning and consolidation of the basic concepts related to the rate of chemical reactions and catalysis. 2. Understanding the practical impact of selected factors on the rate of chemical reactions in laboratory conditions. 3. Analysing the theoretical way of influencing the acceleration or delay of chemical processes.		
3	**Prerequisites:** general chemical knowledge of the rate of chemical reactions, catalysis and catalysts		
4	**Description of the laboratory workplace:** Basic laboratory equipment – test tube rack, water baths, micro-spatula, stopwatch, chemical reagents – 0.1 M sulphuric acid(VI) H_2SO_4, 0.1 M potassium manganate(VII), $KMnO_4$; 0.1 M oxalic acid $H_2C_2O_4$, 0.1 M acetic acid; 3% hydrogen peroxide solution, H_2O_2, manganese(IV) oxide MnO_2, lead(IV) oxide PbO_2**, manganese(II) sulphate(VI) $MnSO_4$, sodium sulphate(IV), Na_2SO_3;**		

5	**Risk assessment:** Chemical burns resulting from contact with 0.2 M sulphuric acid and caustic soda are very unlikely, the possible effects are minor, Final assessment – **VERY SMALL THREAT** **Security measures required:** 1. Lab coats, gloves and safety glasses. 2. Health and safety cleaning products, paper towels.
6	**The course of the exercise:** 1. Getting to know the workplace manual (appendix 1), 2. Performing individual exercises according to the instructions.
7	**Exercise report:** 1. Develop an exercise in accordance with the instructions contained in the workplace manual. 2. Solve the given task and/or answer the questions included in the set of tasks and questions to be completed by the student.
8	**Archiving of research results:** Submit the report on the exercise in the applicable form at the beginning of the next laboratory exercises.
9	**Assessment method and criteria:** a. EKP1, EKP2 – control of the knowledge of basic chemical concepts regarding the rate of chemical reactions, catalysis, catalysts and inhibitors during the classes, b. SEKP7 – the detailed learning outcome for an individual student will be assessed on the basis of the solutions to tasks and problems presented in the report, given for independent solution/development: – mark 2,0 – the student has too little knowledge of the rate of chemical reactions, factors influencing the rate of reaction, the catalysis process, catalysts and inhibitors, or is unable to solve simple tasks related to the above-mentioned concepts; – mark 3,0 – has a basic chemical knowledge of the rate of chemical reactions, catalysis, catalysts and inhibitors and is able to solve simple problems in his profession related to the above-mentioned concepts; – mark 3,5 – 4,0 – has extensive knowledge of chemistry and the rate of chemical reactions and the catalysis process, and has the ability to solve complex tasks in his specialty regarding the above-mentioned concepts; – mark 4,5 – 5,0 – has the ability to apply complex knowledge regarding the rate of chemical reactions and the catalysis process, and is able to solve problem tasks in his specialty regarding the above-mentioned concepts.
10	**References:** 1. https://assets.openstax.org/oscms-prodcms/media/documents/Chemistry2e-WEB.pdf (accessed 15.07.22). 2. A. Kozłowski, A. Kalbarczyk-Jedynak, M. Ślączka-Wilk, K. Ćwirko, C. Wiznerowicz, G. Gorzycka, Instrukcje stanowiskowe do ćwiczeń laboratoryjnych: Szybkość reakcji chemicznych. Kataliza, AM Szczecin 2022 (in Polish). 3. J. E. McMurry, R. C. Fay, J. K. Robinson, Chemistry, 7th edition, global edition, publisher: Pearson, 2016. 4. A. Blackman, S. Bottle, S. Schmid, M. Mocerino, U. Wille, Chemistry, 2nd edition, publisher: John Wiley&Sons, 2012. 5. G. Curran, Chemistry, publisher: The Career Press, 2011. 6. J. T. Moore, Chemistry for Dummies, publisher: Wiley Publishing, 2015.

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11	Notes

1. THEORY

KEYWORDS:

- reaction rate,
- factors that affect the rate (the speed) of reaction,
- chemical equilibrium and Le Chatelier's Principle,
- catalyst/inhibitor.

Reaction rate (reaction speed; speed of reaction; rate of reaction) – is the rate of change in concentration in reactants or products over time. In other words the reaction rate is defined as how quickly or slowly the reaction takes place.

For this reaction:

$C \longrightarrow D$ Rate = $\triangle[D] / \triangle$ time [mol/dm^3 · s] the rate of appearance of product
$C \longrightarrow D$ Rate = $-\triangle[C] / \triangle$ time [mol/dm^3 · s] the rate of loss of reactant
[] – molar concentration; mol/dm^3; mol/l

The rate law – explanation based on reaction: $A + B \longrightarrow AB$; the rate usually depends on [A] and [B]. This is an equation that states the instantaneous rate of reaction as a function of reactant concentrations.

First-order reaction: rate law: Rate = $k_1[A]$ or Rate = $k_1[B]$; k = rate constant; [A], [B] – molar concentrations of A and B.

Second –order reaction: rate law: Rate = $k_2[A]^2$ or Rate = $k_2[B]^2$ or Rate = $k_2[A][B]$.

Zero-order reaction: rate law: Rate = k

The main factors that affect the rate (the speed) of reaction:

- activation energy (E_a) – is a minimum energy required to initiate a chemical reaction [kJ/mol], in other words, this is the energy that we have to put in to get the reaction going,
- concentration – the higher concentration of one or more reactants the faster the rate of reaction; if we increase the concentration of one or more reactants, the speed of reaction (the rate of reaction) will also increase,
- temperature – an increase in temperature increases the speed of reaction,

Arrhenius equation: $k = Ae^{-Ea/RT}$ Ea – activation energy; R – ideal gas constant; T – temperature [K]; A – a frequency factor (frequency of collisions) – it depends on reaction:

- pressure of reacting gases – the higher the pressure the faster the speed of reaction; an increase in pressure increases the rate of reaction,
- surface area of solid reactants – an increase in surface area increases the speed of reaction,
- presence of a catalyst – a catalyst is a substance that increase the speed of reaction without being affected chemically or being used up itself during the reaction of interest. A catalyst reduces an activation energy (Ea) that is why it increases the rate of reaction.

Chemical equilibrium – typical for reversible reaction – two opposite reactions (forward and reverse) are taking place at the same time in the same place and with equal rates. Represented in the chemical balanced equation by:

a A + b B $\rightleftarrows$ c C + d D at equilibrium, the process is described by constant (Kc; equilibrium constant; K_{eq}):

$K_c = [C]_{eq}^c \times [D]_{eq}^d / [A]_{eq}^a \times [B]_{eq}^b$ [ ] refers to molarity (M = mol/l; mol/dm^3)

(Guldberg and Waage – Law of Mass Action – the law states that equilibrium condition is expressed by equation of equilibrium constant).

It is said that chemical equilibrium is dynamic, it means that reactants and products are continually being formed and re-formed.

For all other conditions (except equilibrium: eq) the process is expressed by the reaction quotient, Q_c:

$Q_c = [C]^c \cdot [D]^d / [A]^a \cdot [B]^b$ [ ] refers to molarity (M=mol/l; mol/dm^3)

when $Q_c = K_c$ – the reaction is at equilibrium

when $Q_c > K_c$ – the reaction will go to the left

when $Q_c < K_c$ – the reaction will go to the right

Le Chatelier's principle:

When a system at dynamic equilibrium is perturbed (disturbed, changed), the system will adjust itself in order to minimize the effects of the perturbation (the effect of changes).

Summary of Le Chatelier's principle:

- CONCENTRATION: adding reactant/removing product – equilibrium shifts in the forward direction; removing reactant/adding product – equilibrium shifts in the reverse direction; K_c does not change,
- TEMPERATURE: increasing temperature – equilibrium shifts in the forward direction for endothermic reactions, in the reverse direction for exothermic reactions; decreasing temperature - equilibrium shifts in the reverse direction for endothermic reactions, in the forward direction for exothermic reactions; K_c changes,
- CATALYSTS: no shift; K_c does not change,
- PRESSURE: increasing pressure – equilibrium shifts toward the side of the reaction with fewer moles of gas; decreasing pressure – equilibrium shifts toward the side of the reaction with more moles of gas; K_c does not change.

Enthalpy (H) is the measure of energy that can be released as heat; ΔH – reaction enthalpy (the enthalpy change).

When **$\Delta H < 0$**, the reaction is exothermic (heat is released; gives off heat to its surroundings); when **$\Delta H > 0$**, the reaction is endothermic (heat is consumed; absorbs energy from its surroundings).

Catalyst – is a substance that increase the speed of reaction without being affected chemically or being used up itself during the reaction of interest. A catalyst reduces an activation energy (Ea) that is why it increases the rate of reaction.

Heterogeneous catalyst – is a substance that acts in a different phase than the reactants.

Homogeneous catalyst – is a substance that acts in the same phase as the reactants.

Biocatalyst – mainly enzymes – protein based catalysts in metabolism and catabolism.

Autocatalytic reaction – the product is itself the catalyst for this reaction.

Inhibitor – is a substance that increases an activation energy that is why it slows down a reaction (decreases the speed of reaction).

Example 1 – The reaction: $2NO_2 \rightleftharpoons 2NO + O_2$ is at equilibrium at some temperature, T and the following equilibrium concentrations are measured:
$[NO_2] = 0.05$ mol/dm^3; $[NO] = 0.22$ mol/dm^3; $[O_2] = 0.14$ mol/dm^3. Calculate the equilibrium constant K_c.
$K_c = [NO]^2 \cdot [O_2] / [NO_2]^2 = (0.22)^2 \cdot 0.14 / (0.05)^2 = 2.71$

Example 2 – Write the equilibrium constant (K_c) expression for the following reaction:
$H_2 + I_2 \rightleftharpoons 2HI$ $\qquad K_c = [HI]^2 / [H_2] \cdot [I_2]$

Example 3 – How will the rate of reaction (Rate) change according to the kinetic equation:
Rate = $k \cdot [A]^2 \cdot [B]$ if the molar concentrations of A and B increase 3 times?
$Rate_x = k \cdot 3A^2 \cdot 3B = k \cdot 9A^2 \cdot 3B = k27A^2 \cdot B$
$Rate_x$ / Rate = $k\ 27A^2 \cdot B / k \cdot A^2 \cdot B = 27$ – the rate of reaction will increase 27 times.

Additional tasks and questions to be performed by the student:

1. Write the equilibrium constant (K_c) expression for the following reactions:
   a. $N_2 + 3H_2 \rightleftharpoons 2NH_3$
   b. $2H_2S \rightleftharpoons 2H_2 + S_2$
   c. $2N_2 + O_2 \rightleftharpoons 2N_2O$
2. The reaction: $2SO_2 + O_2 \rightleftharpoons 2SO_3$ is at equilibrium at some temperature, T and the following equilibrium concentrations are measured:
   $[SO_2] = 0.70$ mol/dm^3; $[SO_3] = 0.9$ mol/dm^3; $[O_2] = 0.15$ mol/dm^3. Calculate the equilibrium constant Kc. Answer: K=11,02.
3. The equilibrium constant (Kc) for reaction: $3H_2 + N_2 \rightleftharpoons 2NH_3$ equals 0.1. Given: equilibrium concentrations of: $[H_2] = 0.2$ mol/dm^3; $[NH_3] = 0.08$ mol/dm^3. Calculate the concentration of $[N_2]$ at equilibrium. Answer: $[N_2] = 8$ mol/dm^3.
4. Find several examples of the application of catalysts/inhibitors in environmental protection and in daily life.

2. INSTRUCTION 6 – LABORATORY EXERCISE 6

Experiment 1 – The effect of concentration of reagents on the reaction rate

Materials and reagents:

Glass test tube set, solutions of sulfuric acid (0.1M H_2SO_4), sodium thiosulfate (0.5M $Na_2S_2O_3$), stopwatch, measuring cylinder.

Experimental procedure:

Prepare three identical test tubes: 1, 2, 3. Pour reagents to each of three test tubes according to the table: first make solutions of sodium thiosulfate and distilled water (test tubes 1, 2, 3) and then add sulfuric acid to the first test tube and start measuring the time, stop the stopwatch when the solution in the first test tube is turbid. Record the elapsed time – t[s], do the same with test tubes 2 and 3.

Test tube	Reagents $Na_2S_2O_3$ [cm^3] / H_2O [cm^3] +H_2SO_4 [cm^3]	Molar concentration (molarity) C [mol/dm^3]	Chemical reaction (balanced chemical equation) – ionic form	Elapsed time t [s]	Conclusions
1.	3 / 6 + 1				
2.	6 / 3 + 1				
3.	9 / 0 + 1				

Data analysis (after the experiment):

1. Write the balanced chemical equation (reaction from the experiment) – ionic form.
2. Fill in the table.
3. Draw a graph of reactant ($Na_2S_2O_3$) concentration versus time.
4. Explain how the reaction rate depends on the concentration of reagents?

Experiment 2 – The effect of temperature on the reaction rate

Materials and reagents:

Glass test tube set, laboratory water bath, solutions of potassium permanganate (0.1M $KMnO_4$), sulfuric acid (0.1M H_2SO_4), oxalic acid (0.1M $H_2C_2O_4$), stopwatch, measuring cylinder.

Experimental procedure:

Pour 4 cm^3 of potassium permanganate (0.1M $KMnO_4$) to each of three test tubes, add 5 drops of sulfuric acid (0.1M H_2SO_4) solution (test tubes 1, 2, 3) and add 2cm^3 of oxalic acid (0.1M $H_2C_2O_4$) solution (test tubes 1, 2, 3). Put the first test tube into the laboratory water bath and start your stopwatch. Heat the first test tube in the laboratory water bath at 80ºC (353K) until a complete discoloration. Record the elapsed time. Put the second test tube into the laboratory water bath and start your stopwatch. Heat the second test tube in the laboratory water bath at 90ºC (363K) until a complete discoloration. Record the elapsed time. Treat the third test tube as a control sample.

Data analysis (after the experiment):

Fill in the table below:

Test tube	Temperature [K]	Elapsed time [s]	Chemical reaction (balanced chemical equation)-ionic form	Calculated temperature coefficient	Conclusions
1.	353				
2.	363				
3.	WITHOUT HEATING				

Experiment 3 – The effect of hydronium ions (oxonium ions; hydrated hydrogen ion) upon reaction rate

Materials and reagents:

Glass test tube set, micro spatula, measuring cylinder, solutions of potassium permanganate (0.1M $KMnO_4$), sulfuric acid (0.1M H_2SO_4), acetic acid (0.1M CH_3COOH), solid sodium sulfite (Na_2SO_3).

Experimental procedure:

Pour 4 cm^3 of potassium permanganate (0.1M $KMnO_4$) to each of two test tubes, add 10 drops of sulfuric acid (0.1M H_2SO_4) solution into the first test tube and 10 drops of acetic acid (0.1M CH_3COOH) solution into the second test tube. Add a pinch of solid sodium sulfite (Na_2SO_3) into both test tubes. Record the color in both test tubes.

Data analysis (after the experiment):

1. Write the balanced chemical equation (reactions from the experiment) – ionic form.
2. Explain how the reaction rate depends on the hydronium ions?

Experiment 4 – The effect of a catalyst on the reaction rate

Materials and reagents:

Glass test tube set, micro spatula, measuring cylinder, solutions of potassium permanganate (0.1M $KMnO_4$), sulfuric acid (0.1M H_2SO_4), oxalic acid (0.1M $H_2C_2O_4$), solid manganese(II) sulfate ($MnSO_4$).

Experimental procedure:

Pour 4 cm^3 of potassium permanganate (0.1M $KMnO_4$) and 4 cm^3 of oxalic acid (0.1M $H_2C_2O_4$) into the first test tube. Do the same with the second test tube: pour 4 cm^3 of potassium permanganate (0.1M $KMnO_4$) and 4 cm^3 of oxalic acid (0.1M $H_2C_2O_4$). Add 2 – 3 drops of sulfuric acid solution (0.1M H_2SO_4) to each of two test tubes of interest. Add a pinch of solid manganese(II) sulfate ($MnSO_4$) into the second test tube. Observe the color of obtained solutions in both test tubes.

Data analysis (after the experiment):

1. Write the balanced chemical equation (reaction from the experiment) – ionic form.
2. Explain the effect of a catalyst on the reaction rate.

Experiment 5 – The effect of a catalyst/inhibitor on the reaction rate

Materials and reagents:

Glass test tube set, micro spatula, measuring cylinder, solutions of hydrogen peroxide (3% H_2O_2), phosphoric acid (0.1M H_3PO_4), solids: manganese(IV) oxide (MnO_2), lead(IV) oxide (PbO_2).

Experimental procedure:

Pour 4 cm^3 of hydrogen peroxide (3% H_2O_2) to each of two test tubes. Treat the first test tube as a control sample. Add solid manganese (IV) oxide (MnO_2) (half of the micro spatula!) into the second test tube. Compare the changes in the rate of hydrogen peroxide decomposition in both test tubes. Light a wooden splint, blow out the flame, allowing the splint to continue glowing and put it into the second test tube. The glowing splint will relight (test for oxygen).

Pour 4 cm^3 of hydrogen peroxide (3% H_2O_2) to each of two test tubes. Treat the first test tube as a control sample. Add solid lead (IV) oxide (PbO_2) (half of the micro spatula!) into the second test tube. Compare the changes in the rate of hydrogen peroxide decomposition in both test tubes. Light a wooden splint, blow out the flame, allowing the splint to continue glowing and put it into the second test tube. The glowing splint will relight (test for oxygen).

Pour 4 cm^3 of hydrogen peroxide (3% H_2O_2) into one test tube. Add few drops of phosphoric acid (0.1M H_3PO_4). Light a wooden splint, blow out the flame, allowing the splint to continue glowing and put it into the test tube.

Data analysis (after the experiment):

1. Write the balanced chemical equation of the decomposition reaction of hydrogen peroxide.
2. Write the definition of a catalyst/inhibitor. Determine the nature of added substances (catalyst or inhibitor) because of their effect on the rate of the hydrogen peroxide decomposition reaction.

3. Guidelines for writing the final laboratory report

1. First page of the report – The Laboratory Report Cover Sheet found on our website: https://www.am.szczecin.pl/en/facilities/institute-of-mathematics-physics-and-chemistry/department-of-chemistry/chemistry-lab-manuals/
2. Second page of the report – „The Theoretical Part" – on a maximum of one page including brief description of keywords.
3. Third page of the report – „The Experimental Part" – including all performed experiments with titles, raw data, reactions, calculations, tables, graphs, etc. It should be written in accordance with „Data analysis (after the experiment)".
4. Additional task/tasks given by the academic teacher.
5. References.

4. In order to pass the laboratory exercise students must pass „the entry test" and submit the final laboratory report at the next laboratory meeting. The lab report must be accepted by the academic teacher.