Oxidation and reduction reactions II

Purpose of the exercise: the ability to determine the potential of a solution with a specific pH containing an oxidant and a reducer with specific concentrations.

1. Execution of the exercise

1.1. pH changes during the redox reactions

- For test-tube containing 1 – 2 cm^3 of **0.1 M KIO_3** add a drop of methyl orange and 2 – 3 drops of **0.1 M HCl** to obtain a red indicator color. Add to the second test-tube 1 - 2 cm^3 of **0.1 M $Na_2S_2O_3$**, then a drop of methyl orange and 2-3 drops of **0.1 M HCl.** The contents of both test-tubes combine and observe a change in the color that indicates a change in the concentration of H^+ ions.
- Mix a few drops of solution containing **SO_3^{2-}** and **0.01 M iodine** on a porcelain plate [bottle marked as: **"I_2"**]. Determine the pH of the mixture using a universal paper.

1.2. Predicting the direction of the reaction depending on the pH

1.2.1 For test-tube containing 1 – 2 cm^3 **6% H_2O_2** add 1 – 2 drops of **2 M H_2SO_4**. Next add a few drops of **0.01 M iodine** [I_2]. After shaking several times, record the color of the solution and basify the solution with **2 M $NH_{3\ aq}$**. Watch the change in color.

1.2.2 For 3 test-tubes add 1 cm^3 of **0.01 M $KMnO_4$** and respectively:

1) 1 cm^3 of **2 M H_2SO_4**,

2) 1 cm^3 of **2 M NaOH**,

3) keep the neutral environment - no additives.

Then add to each test-tube 1 cm^3 **0,1 M $Na_2S_2O_3$** – and observe a change in color.

For test-tube containing 1 – 2 cm^3 iodine ions [**I^-**], add 2 – 3 drops of concentrated **H_2SO_4** and a few drops of potassium iodate(V) solution [**KIO_3**]. Check the presence of iodine using starch solution.

1.3. Disproportionation under the influence of pH

1.3.1. For test-tube containing a few drops of **0.01 M** iodine [**I_2**], add 1 – 2 cm^3 of **2 M NaOH** and observe a change in color.

1.4. The influence of pH on chloride oxidation

1.4.1. To 1 - 2 cm^3 chloride solution [**1M NaCl**] add a few drops of **2 M H_2SO_4** and 1 - 2 cm^3 of **0.01 M $KMnO_4$**. Carefully heat the solution and check the presence of chlorine in the vapors above the test-tube with a moistened with water iodine-starch paper.

In order to demonstrate the role of the reaction of the environment, repeat the experiment in an inert environment [**without acidifying the sample**].

1.4.2. To 1 – 2 cm^3 of chloride solution [**1M NaCl**] add a few drops of **2 M HNO_3** and a little bit of **PbO_2**. Carefully heat the solution and check the presence of chlorine with a moistened with water iodine-starch paper.

In order to demonstrate the role of the reaction of the environment, repeat the experiment in an inert environment [**without acidifying the sample**].

Perform the experiment as in point 1.4.2. using manganese dioxide as an oxidant [**MnO_2**].

2. Development of results

- Write an electron balance for the redox reaction. Explain why there has been a change in the color of methyl orange to yellow.
- Justify the observed pH changes with redox equations.
- Write oxidation and reduction reactions for all experiments.
- Write the equation of the redox reaction in an inert, alkaline and acidic environment giving the approximate values of the apparent redox potentials for the oxidizing reagent reacting at equal pH.
- Write redox reactions, electron balance and normal potentials for oxidants and reducers involved in reactions. Which redox reaction does not take place? For the study, attach the response in the form of the E=f(pH) graph for individual halogens.
- Write a reaction of the disproportionation of chlorine and iodine under the influence of change pH the environment. Draw a graph E = f (pH) and justify the reaction using the apparent redox potentials in a function of pH.
- Write the chemical equation for redox reactions. Complete the table:

No.	Oxidizing agent	The potential of the oxidizer		The presence Cl_2 in the fumes	
		Neutral environment	Acidic environment	Neutral environment	Acidic environment
1.	$KMnO_4$				
2.	PbO_2				
3.	MnO_2				

3. Conclusion

Formulate the conclusion that a change in the concentration of hydrogen ions affects the equilibrium constant of the redox reaction.

4. The scope of material

- The effect of change oxidation - reducing properties and influence on pH change.
- Apparent potential.
- Ability to construct and read dependency diagrams apparent normal potential in pH function.

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- Polyoxidants, ampholytes - examples.
- Oxidized water as an oxidant and reducer in redox reactions.

5. Literature

- G. Charlot *Qualitative Inorganic Analysis,* Wiley 2007 (https://archive.org/details/in.ernet.dli.2015.151602)
- Ulrich Müller *Inorganic Structural Chemistry, 2nd Edition, Wiley 2006*