МІНІСТЕРСТВО ОХОРОНИ ЗДОРОВ’Я УКРАЇНИ
ХАРКІВСЬКИЙ НАЦІОНАЛЬНИЙ МЕДИЧНИЙ УНІВЕРСИТЕТ

# OXIDATION-REDUCTION REACTIONS. ELECTRODE POTENTIALS AND MECHANISM OF THEIR ORIGIN. ROLE OF ELECTROCHEMICAL PHENOMENA IN BIOLOGICAL PROCESSES. POTENTIOMETRY.

Methodical instructions for 1st year students’ self-work
in Medical Chemistry

# ОКИСНО-ВІДНОВНІ РЕАКЦІЇ. ЕЛЕКТРОДНІ ПОТЕНЦІАЛИ ТА МЕХАНІЗМИ ЇХ ВИНИКНЕННЯ. РОЛЬ ЕЛЕКТРОХІМІЧНИХ ЯВИЩ У БІОЛОГІЧНИХ ПРОЦЕСАХ. ПОТЕНЦІОМЕТРІЯ.

Методичні вказівки для самостійної роботи студентів 1-го курсу
з медичної хімії

Затверджено
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Oxidation-reduction reactions. Electrode potentials and mechanism of their origin. Role of electrochemical phenomena in biological processes. Potentiometry: methodical instructions for 1st year students' self-work in Medical Chemistry / compiled by A.O. Syrovaya, S.N. Kozub, V.O. Makarov et al. – Kharkiv: KhNMU, 2017. – 28 p.

Compiled by:
A.O. Syrovaya,
S.N. Kozub,
V.O. Makarov,
V.N. Petyunina,
S.V. Andreeva,
L.V. Lukyanova,
T.S. Tishakova,
O.L. Levashova,
N.N. Chalenko,
E.V. Savelieva,
N.V. Kopoteva
O.S. Kalinenko,
O.O. Zavada,
M.O. Vodolagenko

Окиснювально-відновні реакції. Електродні потенціали та механізми їх виникнення. Роль електрохімічних явищ у біологічних процесах. Потенціометрія: метод. вказ. для самостійної роботи студентів 1-го курсу з мед. хімії /уклад. Г.О. Сирова, С.М. Козуб, В.О. Макаров та ін. − Харків: ХНМУ, 2017.− 28 с.

Укладачі:
Г.О. Сирова
С.М. Козуб
В.О. Макаров
В.М. Петюніна
С.В. Андрєєва
Л.В. Лук'янова
Т.С. Тішакова
О.Л. Левашова
Н.М. Чаленко
О.В. Савельєва
Н.В. Копотєва
О.С. Каліненко
О.О. Завада
М.О. Водолаженко

## Subject «**Oxidation-reduction reactions. Electrode potentials and mechanism of their origin. Role of electrochemical phenomena in biological processes. Potentiometry**»

1. **Number of hours** 4
2. **Material and methodological support.**

Multimedia support (presentation, scientific films).

Posters:

1. ORR. Ion-electron method: basic medium
2. ORR. Ion-electron method: acid medium
3. ORR. Ion-electron method: neutral medium
4. Redox properties of hydrogen peroxide
5. The direction of the redox process
6. ORR in biological systems
7. Nernst's equation for electrode potential determination.
8. The scheme of galvanic cell (Daniell cell)
9. The standard electrode potentials
10. The standard redox potentials
11. The electrochemical circuits
12. The schemes of the indicating and reference electrodes
13. pH-meters

### Educational literature

1. Medical chemistry / V.A. Kalibabchuk, V.I. Halynska, V.I. Hryshchenko et al.; – Kyiv:AUS Medicine Publishing, 2010, - 224 p.

2. Fundamentals of medical chemistry: manual for students' self-work / A.O. Syrovaya, E.R. Grabovetskaya, L.G. Shapoval. – Kharkiv: KhNMU, 2015.–196 p.

3. Medical chemistry. Adapted concise course: manual for students' self-work / A.O. Syrovaya, E.R. Grabovetskaya, L.G. Shapoval. - Kharkiv: KhNMU, 2013. – 160 p.

4. Medical chemistry: workbook for self-work of first-year students of medical and dentistry faculties / compiled by A. O. Syrovaya, V. N. Petunina, V. A. Makarov et al. – Kharkiv : KhNMU, 2017. – 72 p.

5. Oxidation-reduction reactions. Electrode potentials and mechanism of their origin. Role of electrochemical phenomena in biological processes. Potentiometry: methodical instructions for 1st year students' self-work in Medical Chemistry / compiled by A.O. Syrovaya, S.N. Kozub, V.O. Makarov et al. – Kharkiv: KhNMU, 2017. – 28 p.

6. Individual tasks for students' self-control of knowledge in Medical Chemistry / A.O. Syrovaya, L.G. Shapoval, V.N. Petiunina, et al. – Kharkiv: KhNMU, 2014. –50 p.

7. Text of lecture.

3. **Substantiation for the subject.** Oxidation-reduction reactions take place in the human organism on different levels including a cell. $CO_2$ and $H_2O$ are products of biological oxidation. They form the basis of some volumetric analysis methods (oxidimetry), which enable to determine the content of oxidizing and reducing agents in solutions including the biological fluids.

Knowledge of this subject is necessary for a future physician for the successful studying of such disciplines as bioorganic and biological chemistry, normal and pathological physiology, pharmacology and others.

Electrochemical processes are widely used in scientific research. Potentiometric determination of pH is widely used in medical practice and medico-biological investigations. Diffuse and membrane potentials play a great role in biological objects. Biopotentials are very sensitive to the physiological changes in cells and organs and it is used in electrocardiography and encephalography.

4. **The purpose of the subject:**

- general: to clarify the role of oxidation-reduction reactions in the vital activity of the organism as well as their application in medico-biological investigations; to determine the values of electrodes potentials, to make up the patterns of electrodes and galvanic elements;
- specific: to be able to make up ORR equations using ion-electron method (the half-reaction method); to determine the equivalent, the equivalence factor and the molar weight of the equivalent of oxidizing and reducing agents in ORR; to know the nature of formation of electrode potentials, to analyze the principles of potentiometry and to make the conclusions about its use in medico-biological investigations, to measure redox potentials and predict the direction of redox reactions;

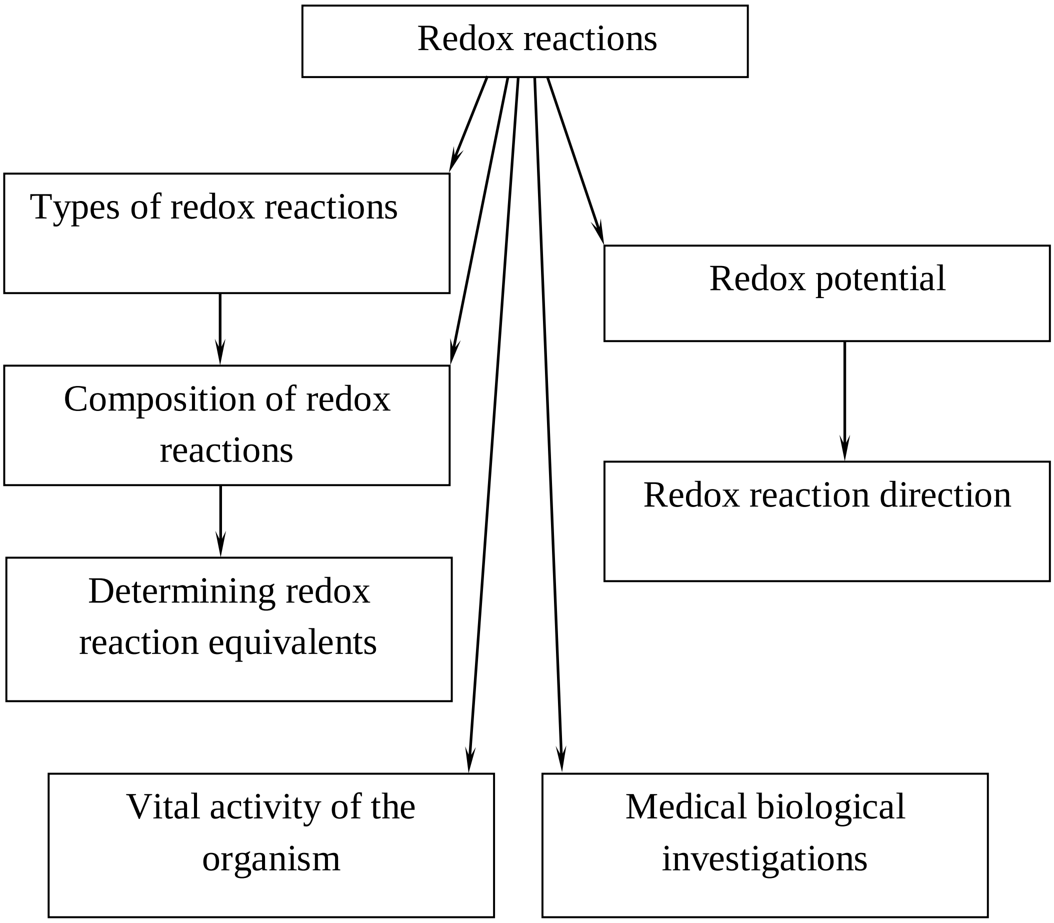
a) **to know:** the role of oxidation - reduction processes in the body; the role of redox potentials;

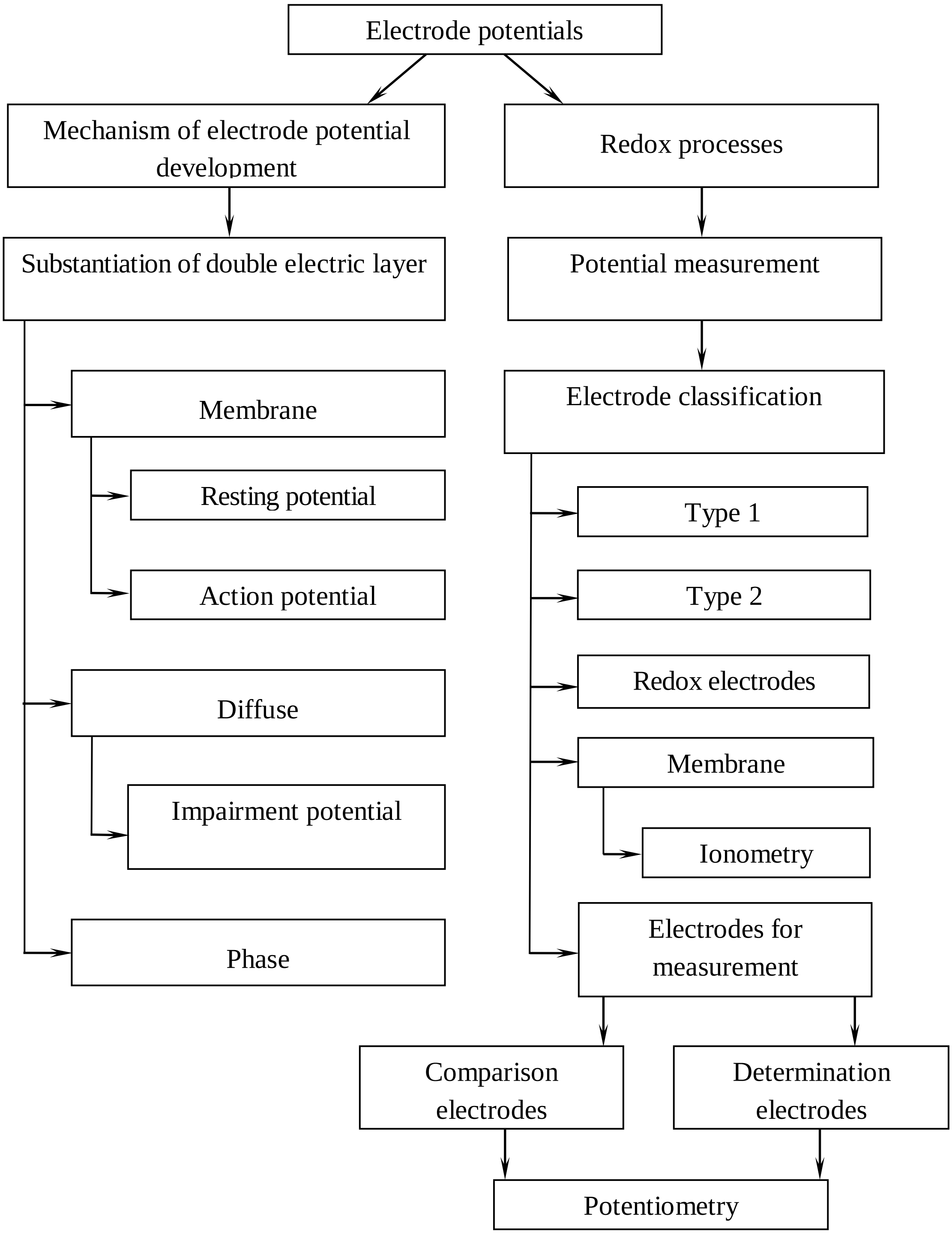
b) **to be able to:** make up ORR equations using ion-electron method (the half-reaction method), determine the equivalent, the equivalence factor and the molar weight of the equivalent of oxidizing and reducing agents in ORR, determine the direction of ORR by the values of the standard oxidation-reduction potentials; determine the values of electrode potentials, to make up the patterns of electrodes and galvanic elements.

c) **practical skills**:

- to make up ORR equations using ion-electron method,
- to determine the equivalent, the equivalence factor and the molar weight of the equivalent of oxidizing and reducing agents in ORR,
- to determine the direction of ORR by the values of the standard oxidation-reduction potentials,
- to determine the values of electrodes potentials and electromotive force,
- to make up the patterns of electrodes and galvanic elements;
- to determine pH of biological liquids by potentiometry.

## 5. Graph structures of the subject.




Electrode potentials
Mechanism of electrode potential development
Redox processes
Substantiation of double electric layer
Potential measurement
Membrane
Resting potential
Action potential
Diffuse
Impairment potential
Phase
Electrode classification
Type 1
Type 2
Redox electrodes
Membrane
Ionometry
Electrodes for measurement
Comparison electrodes
Determination electrodes
Potentiometry

## 6. Plan of students' work

| № | Stages | Time, minutes | Training and visual aids | Location |
|---|---|---|---|---|
| 1. | Motivational characteristics and plan of the subject. Answers to the students' questions | 25 | Text book (work book) | |
| 2. | Initial control | 20 | | Class room |
| 3. | Students' self-work with methodical literature, the solution of educational problems | 55 | Methodical instruction for students, lecture notes, text book for students' self-work, reference data, posters. | |
| 4. | Discussion of procedure of laboratory work | 5 | | |
| 5. | Performing of laboratory work and recording | 40 | | |
| 6. | Defence of laboratory work | 5 | | |
| 7. | Control of knowledge | 20 | | |
| 8. | Analysis and conclusions<br>Home work | 10 | | |

## 7. Tasks for self work:

**- list of questions to be studied:**

1. Oxidation-reduction reactions (ORR). Types of oxidation-reduction reactions.
2. Balancing redox equations: ion-electron method.
3. Determination the equivalent, the equivalence factor and the molar weight of the equivalent of oxidizing and reducing agents in ORR.
4. Redox potential, as a measure of the oxidazing and reducing capacity of the system. The direction of ORR.
5. The role of oxidation-reduction reactions in the vital activity of the organism.
6. The use of redox reactions in medico-biological investigations.
7. Electrode potentials and mechanism of their origin.
8. Measurement of electrode potentials. Standard hydrogen electrode. Nernst equation.
9. Classification of electrodes:

a) electrodes of I type;
b) electrodes of II type;
c) electrodes of III type;
d) membrane electrodes.

10. Galvanic cells.
11. Potentiometry. Determination of pH and active concentration by means of potentiometry.
12. Diffuse and membrane potentials. Biopotentials.

## 1. Oxidation-reduction reactions (ORR). Types of oxidation-reduction reactions.

Reactions as a result of which the oxidation states of elements change are called oxidation-redaction reactions. (ORR).

To understand chemistry of oxidation-reduction phenomena let's remember rule for assigning oxidation states of elements in molecule:

1. The oxidation state of an element in its free state is zero. Thus atoms in the molecules of elementary substance ($H_2$, $O_2$, Fe, Cl, Na) have zero oxidation state.

2. The oxidation state of a monatomic ion is the same as the charge on that ion, for example, $Na^+$ +1, $Ca^{2+}$ +2, a $Cl^-$ -1.

3. The oxidation state of hydrogen in most ionic compounds is +1. This rule is valid for hydrogen compounds such as $H_2O$, $NH_3$ or $CH_4$ except for metal hydrides, such as NaH, LiH, etc., in which the oxidation state for H is -1.

4. The oxidation state of oxygen in its compounds is -2 (for example, $H_2O$, $H_2SO_4$, NO, $CO_2$ and $CH_3OH$) except for peroxides like $H_2O_2$, and $Na_2O_2$, in which the oxidation state for O is -1. Another exception from the rule is $OF_2$, where oxidation state of oxygen is +2 and fluorine is -1.

5. In the compounds of nonmetals, which don't contain hydrogen and oxygen, nonmetal with higher electronegativity is considered negatively charged. The oxidation state of this non-metal is equal to the charge of its negative ion. For example, oxidation state of chlorine in $CCl_4$ is -1 and oxidation state of carbon is +4. In $CH_4$ the oxidation state of hydrogen is +1, and the oxidation state of carbon is -4. The oxidation state of florine in $SF_2$ is -1 and the oxidation state of sulfur is +2, but the oxidation state of sulfur in $CS_2$ is -2 and the oxidation state of carbon is +4. Concept of the oxidation state loses its significance in the molecules like $N_4S_4$ with covalent bonds (where atoms are joined similar in electronegativity).

6. The algebraic sum of oxidation states for all atoms in a molecule of the compound should be zero. Thus, in the molecule of $NH_4Cl$ sum of oxidation states for hydrogen atoms is 4·(+1), and oxidation state of chlorine is -1, that's why the oxidation state of nitrogen is -3.

7. The algebraic sum of oxidation states for all atoms in the complex ion must be equal to its total charge. Thus, the oxidation state of nitrogen must be -3 in the $NH_4^+$ ion and, therefore for hydrogen oxidation state: -3+4 =+1. In the $SO_4^{-2}$ ion sum of oxidation states for four oxygen atoms is equal to -8, that's why sulfur has an oxidation state +6.

8. In the chemical reactions conservation rule of the algebraic sum of oxidation states for all atoms must be complied. The increase in oxidation state of an atom through a chemical reaction is known as an oxidation; a decrease in oxidation state is known as a reduction. Oxidation and reduction processes must compensate each other in the complete equation.

The oxidation state for atoms in the molecules and complex ions is calculated using the above rules. Each element has definite oxidation state associated with its position in the periodic system.

Oxidation and reduction properly refer to a change in oxidation state. Changing of the oxidation states means that electrons are transferred. *Oxidation* is a loss of electrons. *Reduction* is a gain of electrons.

A substance that loses electrons in a chemical reaction (as indicated by an increase in oxidation state) is said to be **oxidized.**

A substance that gains electrons in a chemical reaction (as indicated by a decrease in oxidation state) is said to be **reduced.**

For example:

$$H_2S^{-2} + I_2^0 \rightarrow 2HI^{-1} + S^0$$

the oxidation state of sulfur increases, it is oxidized. The oxidation state of iodine decreases, it is reduced. The oxidizing agent is an iodine and reducing agent is the hydrogen sulfide.

Compounds containing an element in its highest oxidation state can be only oxidizing agent. Compounds containing an element in its lowest oxidation state can only be reducing agent. If element in compound is in an intermediate oxidation state it can serve both oxidizing and reducing agent.

For example: $H_2S^{+6}O_4$ is only oxidizing agent, $H_2S^{+4}O_3$ can be as oxidizing or reducing agent depending on the conditions, $H_2S^{-2}$ is a reducing agent.

**Major oxidizing agents are:**

**1)** Electronegative elements (fluorine, oxygen, chlorine, etc);
**2)** Oxygen-containing acids having non-metal in its highest oxidation state ($H_2SO_4$, $HNO_3$, $HMnO_4$) and salts of these acids ($KMnO_4$, $K_2Cr_2O_7$, etc);
**3)** Metal ions in highest oxidation states ($Fe^{3+}$, $Cu^{2+}$, etc);
**4)** Oxides of non-metals and transition elements in highest oxidation state ($SO_3$, $MnO_2$, etc), and all peroxides ($H_2O_2$, $Na_2O_2$, etc).

**Major reducing agents are:**

1) Metals;
2) Hydrogen, carbon, CO;
3) Hydracids and their salts ($H_2S$, HI, etc).

All redox reactions can be divided into three types:

1) Intramolecular oxidation-reduction reaction. In this process, one constituent part of the substance is an oxidizing agent, and the other is a reducing agent: $2KClO_3 = 2KCl + 3O_2$.
2) Intermolecular oxidation-reduction. These are ordinary redox reactions in which element change oxidation states being in composition of different molecules.
3) Disproportionation or autoxidation-autoreduction reactions. In these reactions, within the same substance, one part of the element is oxidized and the other part is reduced:

$$Cl_2 + 2Ca(OH)_2 = CaOCl_2 + CaCl_2 + 2H_2O$$

(chlorine undergoes both oxidation and reduction).

## 2. Balancing redox equations: ion-electron method.

To make up the equations for redox reactions there are two methods: electronic (method of electronic balance) and ion-electron method (also known as the half-reaction method). First method is used to balance intramolecular reactions and disproportionation reactions.

The electron-ion method used for setting up equations intermolecular reactions in solutions. The method is based on the preparation of half-reactions for oxidation and reduction processes and summing them in the general equation.

Let’s consider the ion-electron scheme by the example of sodium sulfite oxidation with potassium permanganate in different medium.

***Acidic medium:***

$$2KMnO_4 + 5Na_2SO_3 + 3H_2SO_4 \rightarrow 2MnSO_4 + 5Na_2SO_4 + K_2SO_4 + 3H_2O$$

| | |
|---|---|
| $SO_3^{2-} + H_2O - 2\bar{e} \rightarrow SO_4^{2-} + 2H^+$ | 5 |
| $MnO_4^- + 8H^+ + 5\bar{e} \rightarrow Mn^{2+} + 4H_2O$ | 2 |
| $2MnO_4^- + \cancel{16}H^+ + 5SO_3^{2-} + \cancel{5H_2O} \rightarrow 2Mn^{2+} + \cancel{8}H_2O + 5SO_4^{2-} + \cancel{10H^+}$ | |

$$2MnO_4^- + 6H^+ + 5SO_3^{2-} \rightarrow 2Mn^{2+} + 5SO_4^{2-} + 3H_2O$$

**Basic medium:**

$$2KMnO_4 + Na_2SO_3 + 2KOH \rightarrow 2K_2MnO_4 + Na_2SO_4 + H_2O$$

| | |
|---|---|
| $MnO_4^- + \bar{e} \rightarrow MnO_4^{2-}$ | 2 |
| $SO_3^{2-} + 20H^- - 2\bar{e} \rightarrow SO_4^{2-} + H_2O$ | 1 |
| $2MnO_4^- + SO_3^{2-} + 2OH^- \rightarrow 2MnO_4^{2-} + SO_4^{2-} + H_2O$ | |

**Neutral medium:**

$$2KMnO_4 + 3Na_2SO_3 + H_2O \rightarrow MnO_2 + 3Na_2SO_4 + 2KOH$$

| | |
|---|---|
| $MnO_4^- + 2H_2O + 3\bar{e} \rightarrow MnO_2 + 4OH^-$ | 2 |
| $SO_3^{2-} + 2H_2O - 2\bar{e} \rightarrow SO_4^{2-} + 2H^+$ | 3 |
| $2MnO_4^- + \cancel{7}H_2O + 3SO_3^{2-} \rightarrow 2MnO_2 + 3SO_4^{2-} + \cancel{8}OH^- + \cancel{6H^+}$ | |

$$2MnO_4^- + H_2O + 3SO_3^{2-} \rightarrow 2MnO_2 + 3SO_4^{2-} + 2OH^-$$

Making up the redox reactions by the above-mentioned method ions and molecules of oxidizing agent, reducing agent and products of the reaction are considered in the form in which they exist in solution. Weak electrolytes and poorly soluble substances is written in molecular form. Take into account the following:

1. Redox reaction can be balanced by emphasizing the equal exchange of electrons.

2. The total number of atoms of the same name in the left side of the equation must be equal to their number from the right.

3. In neutral or acidic solution, $H_2O$ and $H^+$ may be used for balancing oxygen and hydrogen atoms. The oxygen atoms are balanced first. For each excess oxygen atom on one side of an equation, balance is obtained by adding one $H_2O$ to the other side. Then $H^+$ is used to balance hydrogens. Note that $O_2$ and $H_2$ are not used to balance the oxygen and hydrogen atoms unless they are actually reactants or products.

In acidic solution: $O^{-2} + 2H^+ = H_2O$.

In basic and neutral solutions: $HOH + O^{-2} = 2OH^-$.

If the solution is alkaline, OH and $H_2O$ may be used. For each excess oxygen on one side of an equation, balanced is obtained by adding one $H_2O$ to the same side and 2 $OH^-$ to the other side. If hydrogen is still unbalanced after this is done, balance is obtained by addition of one OH for each excessed hydrogen on the same side and one $H_2O$ to the other side. If both oxygen and hydrogen are in excess on the same side of the half-reaction, an OH can be written on the other side for each paired excess of H and O.

In acidic and neutral solutions: $H_2O = O^{-2} + 2H^+$.

In basic solution: $2OH^- = O^{-2} + H_2O$.

In reactions involving hydrogen peroxide: as already indicated, $H_2O_2$ may be as both oxidizing and reducing agent. Depending on the medium, hydrogen peroxide undergoes transformation:

$H_2O_2$ – oxidizing agent:

$H_2O_2 + 2\bar{e} \rightarrow 2OH^-$ in basic medium;

$H_2O_2 + 2H^+ + 2\bar{e} \rightarrow 2H_2O$ in acidic medium.

$H_2O_2$ – reducing agent:

$H_2O_2 - 2\bar{e} \rightarrow 2H^+ + O_2$ in acidic medium;

$H_2O_2 + 2OH^- - 2\bar{e} \rightarrow 2H_2O + O_2$ in basic medium.

## 3. Determination the equivalent, the equivalence factor and the molar weight of the equivalent of oxidizing and reducing agents in ORR.

Equivalent of oxidizing and reducing agent in the ORR is not constant and may vary depending on the reaction conditions. To determine the equivalent of the oxidizing agent is equal to its molar mass divided by the number of electrons gaining by one molecule of the oxidizing agent.

For example:

| |
|---|
| $SO_3^{2-} + H_2O - 2\bar{e} \rightarrow SO_4^{2-} + 2H^+$ |
| $MnO_4^- + 8H^+ + 5\bar{e} \rightarrow Mn^{2+} + 4H_2O$ |

The equivalent of a reducing agent is called its amount, which corresponds with one electron in the ORR. To determine the equivalent of the reducing agent is equal to its molar mass divided by the number of electrons losing by one molecule of the reducing agent.

$$E(FeSO_{4(red.)}) = 152/1 = 152 \text{ g/mol}$$

Value of standard redox potential can be used to determine tendency for electron transfer and hence, the direction of the reaction. The E.M.F. which is difference between the redox potential of oxidizing and reducing agent is used for this.

$$E.M.F. = e_{Ox} - e_{Red}$$

If E.M.F. > 0, reaction occurs, but if E.M.F. <0 reaction does not occur.

For example, whether the reaction proceeds $Fe_2(SO_4)_3 + NaI$ ?

$$E^0_{Fe^{2+}/Fe^{3+}} = 0,77 \text{ V}, \; e^0_{2I^-/I_2} = 0,53 \text{ V}.$$

$$E.M.F. = e_{ox} – e_{red} = 0,77 – 0,53 = 0,24 \text{ V}.$$

This reaction is possible because the value of E.M.F. is positive.

$$Fe_2(SO_4)_3 + 2NaI \rightarrow 2FeSO_4 + Na_2SO_4 + I_2 .$$

Standard redox potential in biochemistry is different from the electrochemical. It is determined for the conditions:

$$a = 1, \; T = 298 \text{ K and } pH = 7.$$

Standard oxidation - reduction biochemical potentials $e^{0'}$ are measured in volts and are connected to a standard electrochemical potential:

$$e^{0'} = e^0 – 0,059 \cdot pH = e^0 – 0,059 \cdot 7 = e^0 – 0,41 \text{ V}.$$

In biological oxidation hydrogen transfer from the substrate molecule to the oxygen is carried out in several stages. All way of biological oxidation is characterized by a change of potential from - 0.42 to + 0.81 V.

### 4. Redox potential, as a measure of the oxidazing and reducing capacity of the system. The direction of ORR.

Quantitative measure of oxidizing and reducing agents is normal redox potential $e^0$, which refers to the two states of the element. A greater oxidation state determines the oxidized form, but a smaller oxidation state - a reducing form.

For example:

$$e^0_{Fe^{3+}/Fe^{2+}} = + 0,77 \text{ V}$$

$Fe^{3+}$ – oxidizing form; $Fe^{2+}$ – reducing form.

Normal redox potentials are the energy released (+) or absorbed (-) under electron attachment to the oxidizing form for transition to the reducing form if the active concentration of these forms is equal to 1 mol/L (or 1 g-ion/L). All normal oxidation - reduction potentials are measured by the the value of hydrogen electrode taken as 0, $e^0_{H2|2H+} = 0$, that is, these values are relative, and can be found in chemical handbooks.

In many cases, the transition of element from oxidizing form to reducing form requires a specific medium (acidic, neutral, basic). Often medium is involved in the process. In this case, corresponding ions and molecules determining medium are added to the oxidizing form of oxidant. The transition can be achieved only with their participation. Thus $KMnO_4$ is an oxidizing agent, reduced in different manner, depending on the medium. The medium determines the strength of the oxidant. In acidic medium $MnO_4^-$ ions can convert into $Mn^{2+}$ ion (in the presence of a reducing agent), if oxygen binds with hydrogen ions to form water (4 atoms of oxygen need 8 hydrogen ions).

$$MnO_4^- + 8H^+ + 5\bar{e} \rightarrow Mn^{2+} + 4\,H_2O$$

potential of this process $e^0$= 1,56 V.

In neutral medium: $e^0$ = 0,6 V.

$$MnO_4^- + 2H_2O + 3\bar{e} \rightarrow MnO_2 + 4OH^-$$

In basic medium: $e^0$ = 0,56 V.

$$MnO_4^- + \bar{e} \rightarrow MnO_4^{2-}.$$

In this case, the basic medium is not involved in the formation of the final ion, but determines the convertion to $MnO_4^{2-}$ ion. If the value $e^0$ is more the system has high oxidizing properties. If this value is less, the system has great reduction properties.

Taking into account the value of normal oxidation - reduction potentials allows to make up ORR using half-reactions that can be practically implemented. The reaction between two oxidation - reduction systems can take place if the potential difference $\Delta e^0$ ($e^0_1$-$e^0_2$) is larger than 0.2 V. The more difference the quicker and more complete the reaction is.

The Nernst's equation is used for calculation of the redox potential:

$$e = e^0_{ox} + (0.059/n)\lg(C_{ox}/C_{red.})$$

The use of $\lg^0_{ox/red}$ is possible for evaluation of different systems:

a) redox properties of the systems on the example of:

| $F_2 + 2\bar{e} \rightarrow 2F^-$ | $H_2 + 2\bar{e} \rightarrow 2H^+$ |
|---|---|
| $e^0_{ox/red}$ = +2,85 V | $e^0_{ox/red}$ =-2,23 V |
| oxidizing agent | reducing agent |

b) direction of oxidation-reduction process:

$$MnO_4^- + 5Fe^{+2} + 8H^+ \rightarrow Mn^{2+} + 5Fe^{+3} + 4H_2O$$

oxidizing form:

| | |
|---|---|
| $MnO_4^- + 8H^+ + 5\bar{e} \rightarrow Mn^{2+} + 4H_2O$ | $e^0_{ox/red} = +1{,}63$ V |
| $Fe^{+3} + \bar{e} \rightarrow Fe^{+2}$ | $e^0_{ox/red} = +0{,}77$ V |

as $lg^0_{ox/red}$ of the first system is higher than for second, that's why first system will act as the oxidizing agent, but second one - as a reducing agent.

## 5. The role of oxidation-reduction reactions in the vital activity of the organism.

Oxidation-reduction reactions are of great importance in biological systems. Photosynthesis, breathing, digestion – all these are chains of oxidation-reduction reactions.

The essence of the biological oxidation is in the oxygen addition to the substrate or hydrogen proton and electron detachment and transfer from one molecule (which is being oxidized) to another (being reduced). The oxidation-reduction processes are connected with each other in an organism. Specific enzymes catalyze them: oxidases, dehydrogenases and others.

Energy necessary for vital activity is given off, low-molecular organic substances used for the biosynthesis of specific complex substances of an organism are formed, and the processes of destruction and neutralization of metabolites takes place due to the oxidation-reduction process.

It is necessary to admit that biological redox processes have a number of specific features. First, they proceed through a number of intermediate stages; the energy is released in small portions, which allows the cell to use it most completely. Second, the released energy is stored in the bonds of ATP and other macroergic compounds (process of oxidative phosphorylation).

The redox reactions in the organism are:

- tricarboxylic acid cycle reactions (Krebs cycle);
- tissue respiration;
- photosynthesis reaction;
- fermentation and glycolysis reaction;
- oxidation and fatty acid synthesis reactions.

Impairment of the redox processes is a cause of disease. In particular, when oxygen blood saturation is law, this results in hypoxia. Severe hypoxia can arrest oxygen transport to the tissues (anoxia). This results in loss of consciousness with a possible lethal outcome. Such pathological changes are frequently observed in poisoning with carbon monoxide, industrial gases containing aniline, nitrobenzene and the so-called blood poisons.

## 6. The use of redox reactions in medico-biological investigations.

Oxidation-reduction reactions are the basis of some quantitative analysis methods that are called oxidimetry (redoximetry). Oxidimetry methods are widely used in biochemistry and clinical analysis. Main methods of oximetry depending on the type of the reaction and the operating solutions are permanganatometry, iodometry, chromatometry, etc. The method of permanganatometry (a solution of $KMnO_4$ is the operating solution) is used to determine the content of uric acid in urine, $Ca^{2+}$ in plasma and catalase enzyme content.

In sanitary-hygienic practice these methods are used to investigate drinking water and sewage, hydrochloride lime analysis. In pharmaceutics these reactions are used to analyze drugs.

## 7. Electrode potentials and mechanism of their origin.

When a strip of metal is placed in a solution of the same metal, according to Pysargevsky-Izgaryshev solvation theory of electrode potentials there are two processes in the solution of its salt:

The first one (for active metals such as, Zn):

metal ions as a result of the effect of polar water molecules (solvation), break away from the metal, moving into the solution, if $\Delta H_{\text{solv.}} > \Delta H_{\text{cryst. lat.}}$

$Me^0 - n\bar{e} \rightarrow Me^{n+}$, $Zn^0 - 2e \rightarrow Zn^{2+}$

The second one (for inactive metals such as, Cu):

metal ions coat the strip

$\Delta H_{\text{solv.}} < \Delta H_{\text{cryst. lat.}}$

$Me^{n+} + n\bar{e} \rightarrow Me^0$, $Cu^{2+} + 2e \rightarrow Cu^0$

It is also known that the dissolution process is characterized by the equation:

$\Delta H_{\text{solution}} = \Delta H_{\text{cryst. lat.}} + \Delta H_{\text{solv.}}$

If the energy of the crystal lattice is less than solvation energy, then the first process is possible; if the energy of the crystal lattice is more than solvation energy, the second process occurs. Charged strip of metal attracts oppositely charged ions from the solution. Chemically active metals (e.g., zinc, aluminium) when immersed in a solution of its salt at any concentration partially pass into the solution, and strips become negatively charge.

The opposite phenomenon is observed for the less active metals (like copper, silver, gold): at any concentration of the solution metal ions precipitate on the strip. Then this strip becomes positively charged. Thus, double electrical layer forms on the metal-solution interface. Absolute electrode potential arises on the metal-solution interface near a surface of metal.

A system in which the metal is in contact with the electrolyte solution, characterized by the occurrence of the electrical double layer with some potential e is called *electrode*.

## 8. Measurement of electrode potentials. Standard hydrogen electrode. Nernst equation.

The absolute value of electrode potential cannot be measured, that's why the potential difference between given electrode and a reference electrode is measured. A hydrogen electrode is used as a reference electrode. Its potential equals zero. Electrode potential (relative potential) is the value of E.M.F. of galvanic cell, consisting of the indicating electrode and reference electrode (standard hydrogen electrode):

$Pt(H_2)|H^+ \,||Me^+|Me$.

$\alpha_{H+} = 1 \quad \alpha_{Me}{}^{+} = 1$

E.M.F. of this chain $E = e^0_{Me+|Me} – e^0_{H+|1/2\,H2}$ .

If $e^0_{H+/1/2H2} = 0$, than $E = e^0_{Me+|Me0}$ .

The values of standard electrode potentials ($e^0$) are tabular values, which give the opportunity to determine which of the element in galvanic cell is cathode and anode. Cathode has higher potential but anode has lower potential.

For example, in the scheme consisting of gold and iron electrodes, from the left must be iron electrode because its $e^0{}_{Fe^{2+}|Fe} = -0{,}44$ V, and from the right must be gold electrode because $e^0{}_{Au^{+}|Au} = +1{,}498$ V.

$– \; Fe|Fe^{2+}||Au^+|Au \; +$

Current appears in this circuit. This example shows that metals and alloys of different electrochemical activity can not be applied together for restoration in stomatology.

The value of the electrode potential (e) depends on the standard electrode potential ($e^0$) and the concentration of the ions. This is illustrated by the Nernst equations (at 298K):

$$e = e^0 + \frac{0{,}059}{n} \lg C_{cat} \text{ – if ions are cations}$$

$$e = e^0 - \frac{0{,}059}{n} \lg C_{an} \text{ – if ions are anions}$$

where T = 298 К; n – is the number of electrons; $e^0$ – is the standard electrode potential; C – is the molar concentration of the ions in the solution (mol/L).

## 9. Classification of electrodes

The electrodes used in the electrochemistry have been classified depending on their composition and electrode reaction.

1. *I type*, potential of which depends on the active concentration of metal ions;
2. *II type*, potential of which depends on the concentration of anions and cations;
3. *III type or oxidation-reduction electrodes.*
4. *Membrane electrodes.*

a) The electrodes of the first type *are electrodes consisting of metal, which is in the solution of its salt.*

The scheme of the electrode of the first type: $Me \,|\, Me^{n+}$;

the electrode reaction is $Me – ne^- \rightleftarrows Me^{n+}$.

The electrode potential value depends on the active concentration of metal ions, and, therefore such an electrode can be used for their determination.

*Hydrogen electrode* belongs to the electrodes of I type. It is made by adding platinum black to platinum wire or a platinum plate. Such plate is able to absorb high volumes of molecular hydrogen, which after decomposition convert to the solution in the form of ions:

$$H_2(gas) \leftrightarrow 2H^+ + 2e$$

The scheme of the hydrogen electrode is:

$$Pt(H_2) \mid H^+$$

at $P_{H2}$ = 101,3 kPa; $C_{H+}$ = 1 mol/l; $e_{H+/1/2H2}$ = - 0,059pH

$e^0$ =0, then the Nernst equation is as follows:

$$e=e^0 + 0.059/n\cdot lg[H^+],\ e_{H2/H+}=-0.059pH$$

**b) The electrodes of the second type** *are the electrodes consisting of metal covered with a hardly soluble compound of this metal (salt, oxide, hydroxide) and plunged into a solution of the readily soluble compound with the same anion*. Particularly calomel or silver-chloride ones are used as reference electrodes:

calomel: Hg, $Hg_2Cl_2$ |KCl (sat)
silver-chloride: Ag, AgCl|KCl (sat)
antimony: Sb, $Sb_2O_3$ $|H^+$

The reaction takes place on the electrodes (respectively):

$Hg_2Cl_2 + 2e \rightarrow 2Hg^0 + 2Cl^-$
$AgCl + e \rightarrow Ag^0 + Cl^-$
$Sb_2O_3 + 6H^+ \rightarrow 2Sb + 3H_2O$

Therefore, they can be written:

$$e_{cal}=e^0_{cal} - 0.059/n\cdot lg[Cl^-]$$
$$e_{s.c.}=e^0_{s.c.} - 0.059/n\cdot lg[Cl^-]$$

Given that electrodes are immersed in a saturated solution of KCl, that is $[Cl^-]$ = const, than $e_{cal.}$ = const = 0,248 V, $e_{s.c.}$=0,222 V.

The antimony electrode's potential depends on the concentration of $H^+$ ions in a solution ($Sb_2O_3$, Sb is in a crystal state).

**c) The electrodes of the third type (Oxidation-reduction electrodes)** *are the half-element consisting of metal (platinum or gold) in a solution containing oxidized and reduced forms of the same substance.*

Here is the scheme of the oxidation-reduction electrode:

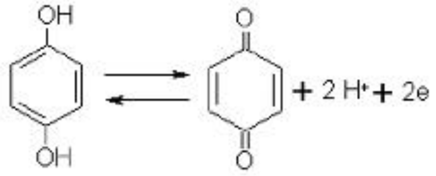
$Pt\left|\frac{[Ox]}{[Red]}\right.$; The electrode reaction is $OX + ne^- \rightleftarrows RED$.

$$Pt\left|\frac{ox}{red}\right. \quad \text{for example,} \quad Pt\left|\frac{Fe^{3+}(\text{oxidation form})}{Fe^{2+}(\text{reduction form})}\right.$$

Platinum doesn't take part in the electrode reaction in the oxidation-reduction electrodes but it acts as carrier of electrons. These electrodes' potential depends on the value of the standard oxidation-reduction potential as well as of the activities ratio of the oxidized and reduced forms of a substance:

$$e = e^{o} + \frac{0,059}{n}\lg\frac{[Ox]}{[Red]}$$

***Quinhydrone electrode***, which **constitutes** the platinum wire soldered in the glass tube and immersed into the investigated solution to which some quinhydrone crystals have been added, is referred to the oxidation-reduction electrodes. Here is the scheme of such an electrode:



The scheme of quinhydrone electrode: $^{Pt\,|\,H+,}{}_{Quinhydrone}$.

The potential of a quinhydrone electrode depends on the $H^{+}$ ions concentration:

$$e_{quinhydrone} = e^{o} + 0,059\lg C_{H}^{+} = e^{o} - 0,059pH;\ e^{o}=0,7V,$$

so $\underline{\mathbf{e_{quinhydrone}=0,7 - 0,059pH}}$

It should be noted that quinone and hydroquinone structures are often found in the living cells (vitamin E, K, coenzyme Q).

The electrode reactions in some oxidation-reduction electrodes take place in the presence of $H^{+}$ ions.

The oxidation-reduction electrode potential for such systems depends not only on the activity ratio of the oxidized and reduced forms of a substance but on the concentration of $H^{+}$ ions as well:

$$MnO_4^{-} + 8H^{+} + 5e \leftrightarrow Mn^{2+} + 4H_2O$$

$$e_{MnO^{-}_4/Mn^{2+}}=e^{0}_{MnO^{-}_4/Mn^{2+}} + (0.059/5)\cdot\lg([MnO^{-}_4][H^{+}]^{8})/[Mn^{2+}]$$

### d) Membrane electrodes

When a semipermeable membrane is placed between two solutions of different concentrations, then from the one side of the membrane excess of negative ions appears and from the other - positive. In this case appeared electrochemical potential difference is called membrane potential.

The ion-selective electrodes belong to the electrodes with membrane potential. Depending on the type of the electrode material the ion-selective electrodes can be divided into three groups: solid, liquid and membrane.

Nowadays 30 types of ionoselective electrodes are produced, using which more than 50 cations and anions and molecular compounds can be identified. The base of the ionoselective electrodes is a semi-permeable membrane that has selective ion conductivity. $Na^{+}$, $K^{+}$, $Mg^{2+}$, $Cl^{-}$, $Cu^{2+}$ and other ions can be determined using ion selective electrodes . Ionometry has advantages over the other methods:

1) Specific feature of ionometry is that this method allows to determine the active ions concentration on the background of its total concentration.

2) To carry out measurements in opaque, colored and turbid environments, even in viscous pastes. This method excludes long, time-consuming operation of filtration, distillation and extraction.

3) Have a wide range of measurements. Limit of qualification (LOQ) for ions in various natural and industrial objects can range from a few mol/L to $10^{-6}$ mol/l. The volume needed for analysis is from 0.05 ml to 0.1 ml.

The above-mentioned advantages of this method allow using its widely in medico-biological investigations.

The electrodes selective to the surface - active substances (determination of impurities by the detergents) and to acetylcholine (determination of impurities by the organophosphate pesticides) are the most important for solving of ecological problems.

The most common type of electrodes of this type is the glass electrode. The potential of this electrode is dependent on the concentration (activity) of $H^+$ ions in the solution.

The membrane is made of sodium ($Si_2$ - $Na_2O$ - Ca) or lithium ($Si_2$ - $Li_2O$ - Ca) glass, it has a cation exchange properties because in an aqueous solution ions of alkali metal (Na or Li) during the hydrogenation can exchange only with hydrogen ions of internal and external solution:

$$H^+(\text{solution}) + Na^+(\text{glass}) \leftrightarrow H^+(\text{glass}) + Na^+(\text{solution})$$

The most widespread membrane electrode is a glass electrode, which consists of a glass tube with a ball made of a special electrode glass on its end. A buffer solution is inserted inside of the system and a silver-chloride electrode is placed there to provide it with a current-conductor. The potential that occurs on the silver-chloride electrode remains constant and doesn’t influence the potential occurring between the glass surface and the investigated solution. When we place the glass electrode into a solution, the hydrogen ions from the solution penetrate into the surface layer, forcing out the alkali metals ions (sodium or lithium) contained in the glass. It results in such $H^+$ ions distribution between the glass and the solution that these phases’ surfaces obtain the opposite charges and the potential which value depends on pH occurs between the glass and the surface.

$$e_{glass} = e^{0'}_{glass} + 0{,}059 \lg \alpha_{H+} = e^{0'} – 0{,}059 \text{ pH}.$$

Glass electrodes have several advantages by changing the pH of biological liquids:

- are not sensitive to the ORR.
- not adsorb proteins.
- indifferent to SAS.

The membrane of solid electrodes is created from a mono or polycrystalline salts poorly soluble in water. Therefore electrodes with a solid membrane are used for identification of those ions, which are part of the membrane.

Electrodes with a liquid membrane are the diaphragm, which pores are filled with the solution of electrode-active material in an organic solvent. The ion exchangers (cationites or anionites) or neutral molecules – membrane-active complexes capable of chelate formation – are used as the electrode-active agents.

## 10. Galvanic cells

Galvanic cell is a system consisting of two electrodes, in which the energy of a chemical reaction is converted into electrical energy. At the heart of any electrochemical element the ORR are taking place separately at the anode - oxidation at the cathode - the reduction.

An example of a galvanic cell is Jacobi – Daniel cell.

Electricity produced in chemical reactions:

$$Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$$

The reaction consists of oxidation and reduction process. The electrode at which oxidation takes place is called the anode; reduction takes place at the cathode:

(–) anode: $Zn^0 - 2e \rightarrow Zn^{2+}$ – oxidation;

(+) cathode: $Cu^{2+} + 2e \rightarrow Cu^0$ – reduction.

The scheme of the galvanic cell is written as per the rule "right plus". It means that from the right must be always written positive electrode (Fig. 1).

(–) Zn | $ZnSO_4$ || $CuSO_4$ | Cu (+) or (–) Zn | $Zn^{2+}$ || $Cu^{2+}$ | Cu (+).

| – the borderline between the metal and the solution in the electrode.

|| – interface between the electrodes.

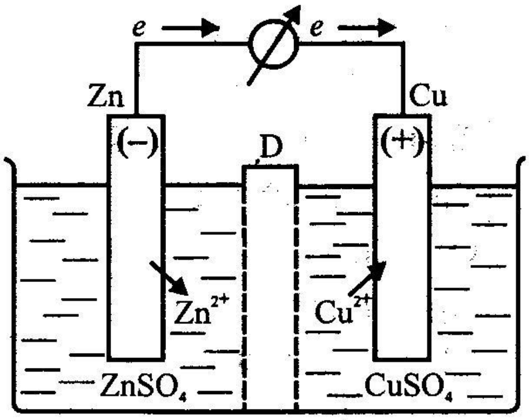


Fig. 1. The scheme of the galvanic cell (Jacobi – Daniel cell).

The electromotive force (E.M.F.) appears in the galvanic cell between two electrodes because of the difference in the electrode potentials.

$$E.M.F. = e^+ - e^-$$

The concentrating galvanic cells are the cells where the chemical reaction energy turns into the electric energy. The similar electrodes in such chains are plunged into the solutions of the same electrolytes having different concentrations:

(–) Ag | $AgNO_3$ || $AgNO_3$ | Ag (+)

The electrode in the solution with a higher concentration ($C_2$) will serve as a cathode, the electrode in the solution with a lower concentration ($C_1$) will serve as an anode.

$$E_k = \frac{0{,}059}{n} \lg \frac{C_1}{C_2}$$

*The oxidation-reduction chains* consist of two oxidation-reduction electrodes:

$$Pt(Cu^{2+}) \,|\, Cu^{+} \,\|\, Fe^{3+} \,|\, (Fe^{2+})Pt$$

## 11. Potentiometry. Determination of pH and active concentration by means of potentiometry.

Potentiometry is a method of analysis based on the measurement of electrode potential and electromotive force of galvanic cells. Potentiometric methods have several advantages: in comparison with the calorimetry. This method is more precise: up to 0,02-0,05 mol/L, which is important for biological researches. As the equilibrium potential value establishes quickly then potentiometric measurements are time-effective. This method can be used for analysis of multicomponent systems, turbid and colored solutions, viscous media, etc.

To measure pH by potentiometric method it is necessary to compose a galvanic cell consisting of two electrodes. One of them is the electrode potential of which depends on pH. Such electrode is called indicating electrode.

*Hydrogen, quinhydrone, antimony* and *glass electrode* are used as indicating electrodes:

1) $Pt(H_2)|H^{+}$, $e = -0{,}059pH$;
2) $Pt|H^{+}_{quin.}$, $e = 0{,}7 - 0{,}059pH$;
3) $Sb,Sb_2O_3|H^{+}$, $e = e^{0'} - 0{,}059pH$;
4) $glass|H^{+}$, $e = e^{0'}_{glass} - 0{,}059pH$.

Second electrode is called reference electrode, potential of it is constant.

Calomel and silver-chloride electrodes are as follows:

$Hg,Hg_2Cl_2|KCl$ sat. $e^0_{cal} = 0{,}248$ V at 298 K, $e_{cal} = e^0_{cal} - 0{,}059\lg\alpha_{Cl^-}$

$Ag,AgCl|KCl$ sat. $e^0_{s.c.} = 0{,}222$ V at 298 K, $e_{s.c.} = e^0_{s.c.} - 0{,}059\lg\alpha_{Cl^-}$

Silver-chloride electrode is more often used as reference electrode (it is more comfortable in design).

For the pH measuring can be used the following galvanic cells:

1) A chain consisting of hydrogen and reference electrodes.

$$Pt(H_2)|H^{+}||KCl|AgCl,Ag$$

$$E.M.F. = e^{+} - e^{-} = e_{s.c.} - (-0{,}059pH) = e_{s.c.} + 0{,}059pH$$

$$pH = (E.M.F. - e_{s.c.})/0{,}059$$

2) A chain consisting of quinhydrone and reference electrodes.

$$Ag,AgCl|KCl_{sat.}||H^{+}|Pt_{quin.}$$

$$E.M.F. = e^{+} - e^{-} = e_{quin.} - 0{,}059pH - e_{s.c.}$$

$$pH=(e^0_{quin.} – e_{s.c.} – E.M.F.)/0{,}059$$

3) A cell composed of an antimony electrode and a reference electrode. This kind of cell is used for pH measuring of a stomach (Linar probe):

$$Sb,Sb_2O_3|H^+||KCl|Hg_2Cl_2,Hg$$

sat.

$$pH=\frac{e_{cal}-e_{ant}-E}{0{,}059}$$

Last chain is used in medicine for intragastric pH determination using two-channel Linar probe. Using this probe pH in the stomach can be determined, where the acidic medium, and in the astral section where glands secrete basic secretion.

Potentiometric titration is carried out in cases where chemical indicators can not be use or do not have a specific indicator.

In potentiometric titration the potentiometer electrodes immersed in titrated solution are used as indicator. During titration changing the concentration of ions occurs. After recording potentiometer readings in units of pH or mV, plot volume of titrant as a function of pH (titration curve), determine the equivalence point and the volume of titrant consumed on titration.

## 12. Diffuse membrane potential. Biopotentials.

Diffuse potential is the potential difference on the borderline of two electrolyte solutions with different concentration or different composition due to different ion mobility.

The value of diffuse potential can be calculated using Henderson equation:

$$e=\frac{U-V}{U+V}\cdot\frac{RT}{nF}\cdot\ln\frac{C_1}{C_2}$$

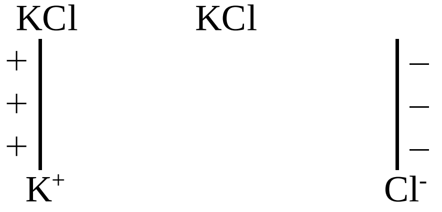
where U – cation mobility; V – anion mobility.

Diffuse potentials can appear in biological objects in the result of impairment, e.g. cellular membranes.

The electrolytes travel from the place of impairment to the intact areas.

The damaged tissue is charged negatively when compared with the intact one, it means that diffuse potential of the impairment occurs. It equals about 30–40 mV.

Membrane potential is formed on the borderline of two solutions if there is a semi-permeable membrane, which passes cations and captures anions. Therefore, one side of the membrane is charged positively, the other – negatively.



$C_1 < C_2$

where $C_1$ and $C_2$ — are cation concentration on the both sides of the membrane, and $C_2 > C_1$, that is membrane potential is calculated using Nernst equation for concentration chain electromotive force.

Semi-permeable membrane transmits cations $K^+$, and $Cl^-$ anions are delayed because the outer surface of the membrane positively charged, and the inner surface of the membrane negatively charged, which results in membrane potential.

Value of $K^+$ concentration in the cell and outside the cell is 20: 1, so you can calculate the membrane potential:

$$e_{memb} = 0{,}059 \cdot \lg 20 = 0{,}059 \cdot 1{,}3 = 0{,}077\ V$$

The changes in membrane potential, which accompany transmission of nerve impulses or muscular contraction, are due to the flow of potassium cations from the cell and sodium cations to it. This results in potential reduction, which can be registered using macroelecrodes placed inside and outside the cell.

## BIOPOTENTIALS

Potential difference measured in the state of physiological rest of the cell is called rest potential.

Rest potential in different cells equals 50–100 mV.

The cause of biopotential appearance is uneven distribution of $K^+$ and $Na^+$ ions. The number of $K^+$ ions in the intracellular fluid is 20–40 times higher than in the extracellular fluid, while $Na^+$ is 10–20 times higher in the extracellular one.

The ions of organic acids pass the membrane with difficulty. At rest $K^+$ ions pass from the intracellular fluid to extracellular.

Thus, inner surface of the cell is charged negatively, while external is charged positively.

The value for rest potential can be determined using the Nerst equation:

$$e_r=(RT/nF)\ln([K^+]_{in}/[K^+]_{ex}),$$

where «in» is inner, «ex» external.

Action potential is explained by increase of cellular membrane permeability for $Na^+$ at the moment of excitation while $K^+$ ions rate remains the same.

In the initial moment of action potential appearance, termed *depolarization*, resting potential decreases, that is $Na^+$ ion flow into the cell results in reduction of the membrane negative charge on the internal surface and later in recharge.

The value of action potential is calculated using the formula:

$$E_{action} = \frac{RT}{F}\left[\ln\frac{[K^+]_{ex.}}{[K^+]_{in.}} + \ln\frac{[Na^+]_{in.}}{[Na^+]_{ex.}}\right]$$

The period of membrane depolarization lasts less than a second, then the flow of $K^+$ ions from the cell increases and the contrary flow of $Na^+$ ions decreases. This goes on until the potential restoration to the value of resting potential. After that $K^+$ ion diffusion decreases up to the initial value. This phase is called *repolarization*.

Biocurrent registration is used in medical diagnosis, electrocardiography in particular. Electrocardiography is the study of the heart physiology by means of registration of the electrical potentials, which develop during the work of the heart muscle.

Electroencephalography is the method of investigation of the bran activity based on registration of the potentials produced by the nerve cells of the brain.

## Laboratory work «Determination of hydrochloric acid concentration by potentiometry».

**Experimental procedure.**

Measure 10 mL of hydrochloric acid solution into a beaker. Set up the burette, filled with the 0.1M NaOH, over the beaker. Potentiometric titration is performed using the electrode system consisting in two electrodes: indicating (glass) electrode and reference (silver chloride) electrode. Place system of electrodes in the beaker with investigated solution, turn on the pH-meter and measure pH value before the addition of titrant. Add 0.5 mL of NaOH standard solution to every 10 mL of HCl solution. Note down results in the table.

| № | Volume of NaOH, mL | pH |
|---|---|---|
| 1 | 0.5 | |
| 2 | 1.0 | |
| 3 | 1.5 | |
| 4 | 2.0 | |
| 5 | 2.5 | |
| 6 | 3.0 | |
| 7 | 3.5 | |

Plot a titration curve according to the measurement results. Place the volume of titrant (NaOH) added to the investigated solution on the x-axis, and place the pH of the investigated solution on the y-axis. After determination the equivalence point calculate volume of NaOH consumed on the titration of 10 mL of hydrochloric acid. The end of the titration can be defined by the changing of potential of indicator electrode near equivalence point.

Using the data about the number of moles of acid ($C_aV_a$) and base ($C_bV_b$) it is possible to calculate concentration of acid and make conclusion.

Conclusions.

- list of practical skills to be mastered.

After studying the subject you must be able to make the equation of the ORR by ion-electron method, to calculate the equivalent, the equivalence factor, molar mass of equivalent of oxidizing and reducing agent in the ORR, to determine the direction the ORR using the values of the standard redox potentials; to be able to determine the value of electrode potentials, to make the scheme of the electrodes and galvanic cells, to determine the pH of biological liquids by potentiometric method.

## 8. Tasks for knowledge control.

1. Redox reaction can occur if $e^0_{ox}$-$e^0_{red}$

A. is 0 V
B. less than 0 V
C. more than 0 V

2. Oxidation is the process of:

A. gain of electrons
B. loss of electrons
C. gain of protons $H^+$
D. loss of $OH^-$ group

3. Does the medium (acidic, neutral or basic) influence the element conversion from oxidizing to reducing form?

A. No;
B. Yes;
C. In some cases;
D. Rarely

4. Calculate electrode potential of magnesium electrode if active concentration of magnesium ions is 1 mol/l, T=298 K.

A. e = – 2,67 V
B. e = + 2,67 V
C. e = 0 V

5. Calculate electrode potential of a zinc electrode if active concentration of zinc ions is 1 mol/l, T=298 K.

A. 0,45 V
B. 0,76 V
C. 0,85 V.

6. Calculate electromotive force of a galvanic element consisting of two silver electrodes immersed in a solution of silver nitrate with active concentration of silver ions 0.1 mol/l and 0.01 mol/l, T=298 K.

A. E.M.F. = 0,049 V
B. E.M.F. = 0,059 V
C. E.M.F. = 0,069 V

7. What is the direction of redox reaction: $Cu^+ + Fe^{3+} \rightarrow Cu^{2+} + Fe^{2+}$ ?

A. left
B. right
C. the reaction does not proceed

Answers: 1 – B; 2 – A; 3 – B; 4 – A; 5 – B; 6 – B; 7 – B.

## 9. Recommendations for the work results design

Algorithms for solving educational problems of class work and self-work, and conclusions of laboratory work should be recorded in the workbook.

## 10. Suggested readings

1. Inorganic Chemistry: manual / V.O. Kalibabchuk, V.V. Ogurtsov, V.B.Emelianov, V.I.Galinska et all. - Kiev: BCB “Medicine”, 2017. - 300 p.

## Навчальне видання

# Окислительно-восстановительные реакции. Электродные потенциалы и механизмы их возникновения. Роль электрохимических явлений в биологических процессах. Потенциометрия.

## Методичні вказівки для самостійної роботи студентів 1-го курсу з медичної хімії

Укладачі:
Сирова Ганна Олегівна

Козуб Світлана Миколаївна

Макаров Володимир Олександрович

Петюніна Валентина Миколаївна

Андрєєва Світлана Віікторівна

Лук’янова Лариса Володимирівна

Тішакова Тетяна Станіславівна

Левашова Ольга Леонидівна

Чаленко Наталія Миколаївна

Савельєва Олена Валеріївна

Копотєва Наталія Василівна

Каліненко Ольга Сергіївна

Завада Оксана Олександрівна

Водолаженко Марія Олександрівна

Відповідальний за випуск Козуб С.М.