

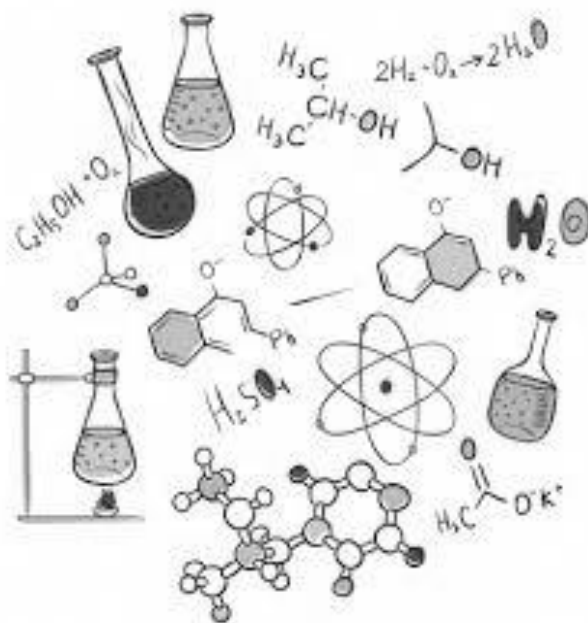
NATIONAL UNIVERSITY OF LIFE AND ENVIRONMENTAL
SCIENCES OF UKRAINE

Department of Analytical and Bioinorganic chemistry &
Water Quality

Kopilevich V.A., Lavryk R.V, Prokopchuk N.M., Voytenko L.V.

Part # 1: GENERAL CHEMISTRY

Manual for Bachelor students
Specialty “Biotechnology”



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У навчально-методичному виданні викладено основні теоретичні положення дисципліни «Загальна та неорганічна хімія». У нього включено тестові завдання та лабораторний практикум.

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Рецензенти: д.х.н., проф. Максін В.І., НУБіП України
к.пед. н., доц. Вощевська О.В., НУБіП України

Автори: Копілевич В.А., Лаврик Р.В., Прокопчук Н.М., Войтенко Л.В.

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Manual expounded the main theoretical notions of General and Inorganic Chemistry. It includes test questions for students' self-control and laboratory training.

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Reviewers: Doctor of Chemical Sci., Prof. Maksin V.I., NULES of Ukraine
Candidate of Pedagogical Sci., Associate Prof. Voshchevska O.V., NULES of Ukraine

Authors: Kopilevich V.A., Lavryk R.V, Prokopchuk N.M., Voytenko L.V.

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1. Solubility of inorganic compounds and method of their transference in solution.

What is necessary to know:

What systems are called solutions?

Why is the down difference between and chemical compounds?

What kind of factors does depend the solubility of compounds?

Thermic effect caused by dissolution of substances.

Methods of determination of concentration of solutions.

How to determine equivalent of oxide, acids, bases, salts?

What one should be able to do:

To calculate the quantity of dissolve substance and quantity of dissolvent for preparation of solution of different concentration.

What one should be concerned in:

Aqueous solutions.

Theory solution of D.I.Mendeleev. Solvate, hydrate, crystalohydrates.

Solution – is homogeneous similar thermodynamic durable mixture of substance – dissolvent and matter. Dissolvent is a substance which (to create solution no to change own aggregate state and usually present in big quantity as for matter. The type of dependence of aggregate state solutions are: hard (alloy of metals), aqueous, gaseous (air).

Homogeneity of solutions tcan be made. Part of solid matterSub time formation of solution in interaction with solvent substance separated in up to molecular or atomic level (10^{-9} M).

Solubility of inorganic compounds aree depend from their nature and described by the rule – the same is ,, similar is dissolved in similar,, (fig. 1).

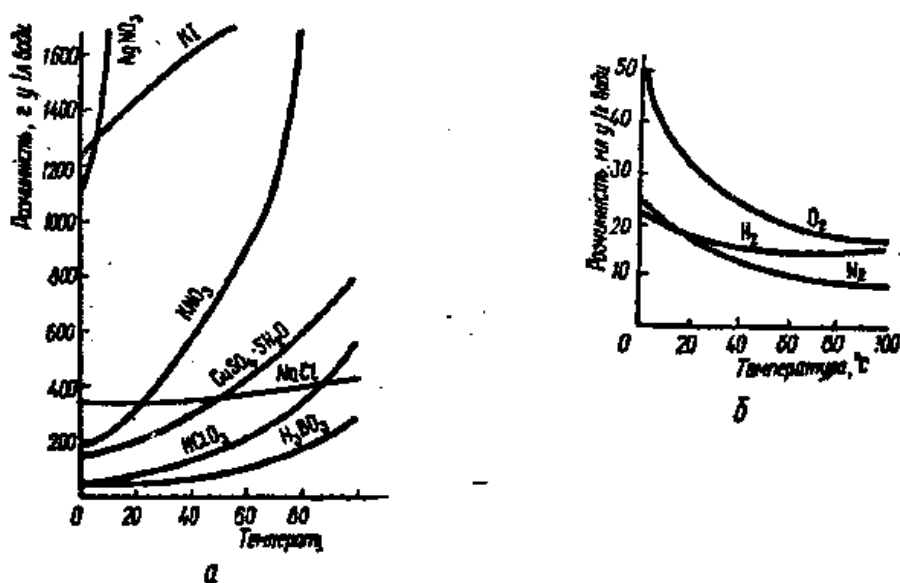


Fig. 1. Dependence of solubility of compounds from temperature:
 a) hard substances; б) gasses.

The concentration of solution is the measurement of the amount of solute in a given amount of solvent or solution. Some medications are solutions of drugs - one-teaspoonful dose at correct concentration might

cure the condition; the same dose in the wrong concentration might kill the patient.

Sometimes solutions are referred to as “concentrated” or “diluted”. These are relative terms. “Diluted” means that there is relatively small amount of solute in solvent. “Concentrated”, on the other hand, means that these terms are unrelated to the degree in which the solution is saturated. Saturated solution of substance can not be considered to be diluted.

Percent of Mass (mass concentration, percent concentration)

Percent of Mass (P) of a solute in solution is the number of grams of solute dissolved in 100 g of solution. For example, the percent of mass of a solution made from 10 g of Sodium Hydroxide (NaOH) dissolved in 90 g of water is found as follows:

$$\begin{aligned} P &= \frac{\text{Mass of solute } m_s}{\text{mass of solute } m_s + \text{mass of solvent } m_{\text{solv}}} \times 100\% = \\ &= \frac{10 \text{ g NaOH}}{10 \text{ g NaOH} + 90 \text{ g H}_2\text{O}} \times 100\% = 10\% \text{ NaOH (by mass)}. \end{aligned}$$

Note that the mass of solution is the sum of the masses of the solute and solvent. Percent concentration is common for solutions used for practical purposes - household or industrial cleaning, killing pests in the garden, medical applications. Percent-of-mass concentration is based only on the mass of solute and is unrelated to its chemical formula of molar mass.

$$P = \frac{m_{\text{solute}}}{m_{\text{solution}}} \times 100\%$$

Molarity (molar concentration)

Molarity (M) is the number of moles of solute in one litre of solution. To find the molarity of a solution one must know the molar mass of the solute. For example, “one-molar” solution of Sodium hydroxide

(NaOH) contains one mole of NaOH in each litre of solution. The concentration of the proceeding solution is written as 1 M NaOH.

One mole of Sodium hydroxide (NaOH) has a mass of 40 g/mol. This quantity of NaOH dissolved in enough water to make exactly 1,00 L of solution gives a 1 M solution. If 20,0 g of NaOH, which is 0,500 mol, is dissolved in 1,00 L of solution, 0,500 M NaOH solution is produced:

$$\text{Molarity} = \frac{\text{Number of moles of solute}}{\text{Number of Litres of Solution}} = \frac{v}{V} = \frac{0,500 \text{ mol NaOH}}{1,00 \text{ L}} = 0,500 \text{ M NaOH.}$$

Note that one molar solution is not made by adding 1 mol of solute to 1 L of solvent. In such case, the final total volume would be slightly different from 1 L. Instead of this 1 mol of salute is firstly dissolved in less than 1 L of solvent. Then, the resulting solution is carefully diluted with more solvent to bring the total volume to 1 L, using volumetric flask.

Molarity is useful taking into consideration the quantity of solute participating in a chemical reaction in solution. Any required molar quantity of the solute can be selected by measuring out the appropriate volume of the solution of known molarity.

$$M = \frac{v}{V} = \frac{\text{mass of solute}}{\text{Molar mass of solute} \times \text{Volume of solution in Litres}} \quad \text{or}$$

$$M = \frac{m_{\text{solute}}}{M_{\text{solute}} \times V}$$

Normality (normal concentration)

The concentration of solutions can be expressed by stating the quantity of the solute in equivalents rather than in moles. This

measurement of concentration is called Normality (N). The normality of solution is the number of equivalents (v_E) of solute per litre (V) of solution:

$$\text{Normality} = \frac{\text{Number of equivalents of solute}}{\text{Number of Litres of Solution}} = \frac{v_E}{V}.$$

The normality of an acid or base solution is commonly expressed using the number of H^+ or OH^- ions available for a complete neutralization.

Therefore:

$$E_{\text{acid}} = \frac{M_{\text{acid}}}{\text{number of hydrogen atoms in acid}},$$

$$E_{\text{base}} = \frac{M_{\text{base}}}{\text{number of OH groups in base}}.$$

For neutral salts equivalent mass may be expressed as:

$$E_{\text{neutral salt}} = \frac{M_{\text{neutral salt}}}{\text{number of atoms of metals in salt} \times \text{Valency of metals}},$$

where E_{acid} , E_{base} , $E_{\text{neutral salt}}$ - equivalent mass of acid, base or neutral salt respectively, g/g-eq;

M_{acid} , M_{base} , $M_{\text{neutral salt}}$ - molar mass of acid, base or neutral salt respectively, g/mol.

$$\text{So, } v_E = \frac{m_{\text{solute}}}{E_{\text{solute}}} \text{ and}$$

$$N = \frac{m_{\text{solute}}}{E_{\text{solute}} \times V}$$

Titre

Titre (T) is a special unit for measuring of concentration connected with chemical quantitative analysis. Titre is the number of grams of solute in 1 millilitre of solution. For example, the titre of a

solution made from 0,3 g of Silver Nitrate (AgNO_3) dissolved in 150 mL of solution is found as follows:

$$T = \frac{\text{Mass of solute } m_s \text{ in grams}}{\text{Volume of solution in milliliters}} =$$

$$= \frac{0,3 \text{ g AgNO}_3}{1150 \text{ mL of solution}} = 0,002 \text{ g/mL.}$$

So,

$$T = \frac{m_{\text{solute in g}}}{V_{\text{solution in mL}}}$$

Table 1. Interconnection of concentration units

Measuring unit		Formulas for re-calculation			
Name	Symbol	M	N	T	P
Molarity (molar concentration)	M	-	$\frac{N \cdot E}{M_r}$	$\frac{T \cdot 1000}{M_r}$	$\frac{P \cdot 10d}{M_r}$
Normality (normal concentration)	N	$\frac{M \cdot M_r}{E}$	-	$\frac{T \cdot 1000}{E}$	$\frac{P \cdot 10d}{E}$
Titre	T	$\frac{M \cdot M_r}{1000}$	$\frac{N \cdot E}{1000}$	-	$\frac{P \cdot d}{100}$
Percent by Mass (mass concentration)	P	$\frac{M \cdot M_r}{10d}$	$\frac{N \cdot E}{10d}$	$\frac{T \cdot 100}{d}$	-

M_r - relative molar mass of solute, g/mol;

E - equivalent mass of solute, g/g-eq;

d - density of solution, g/L.

The fundamental methods of transformation of inorganic compounds in solution are:

1. dissolution in the solvent by heating - NaNO_3 in water.
2. separation in the solvent (H_2SO_4 in water).
3. warming (by heating to 650°C) – 5% NiO in NaPO_3 .

2. Ionic product of water. Notion of pH.

What is necessary to know:

What is normal standard electrode potential?

The Fundamental equation of electro-chemistry.

What is red-ox systems?

Which compounds are called complex?

The Fundamental conception and determination of complex in chemistry.

What one should be able to do:

To prognosticate the direction of chemical reactions to find the meaning of red-ox potential of red-ox systems. To determine and to calculate down the constants of stability and instability of complex ions.

What one should be concerned in:

Red-ox systems and direction of electrochemical processes in such systems.

The part of complex compounds and their application in analytical chemistry.

When the law of mass action is applied to the dissociation of water, we have:

$$\frac{[H^+] \cdot [OH^-]}{[H_2O]} = K \text{ or } [H^+] \cdot [OH^-] = K \cdot [H_2O].$$

Denoting $K \cdot [H_2O]$ by K_w , we obtain $[H^+] \cdot [OH^-] = K_w$. The quantity K_w is equal to product of concentration of the hydrogen and hydroxyl ions in water, and, is, therefore, called the ionic product of water. Its numerical value can easily be found, since K and $[H_2O]$ are known: the dissociation constant of water, K , equals $1,8 \cdot 10^{-16}$, while the concentration of undissociated water molecules is, owing to insignificant dissociation, practically equal to the total number of moles of water per litre, i.e. $100/18=55,56$ moles/L. Consequently, $K_w = 1,8 \cdot 10^{-16} \times 55,56 = 1 \cdot 10^{-14}$.

The ion product of water is an extremely important value, since it enables the concentration of OH^- ions to be found for any aqueous solution with a known concentration of H^+ ions and vice versa. For example, for pure water

$$[H^+] = [OH^-] = \sqrt{K_w} = 10^{-7} \text{ g-ions/L.}$$

By using the ion product of water, the acidity or alkalinity of any solution can be expressed in terms of hydrogen ion concentration.

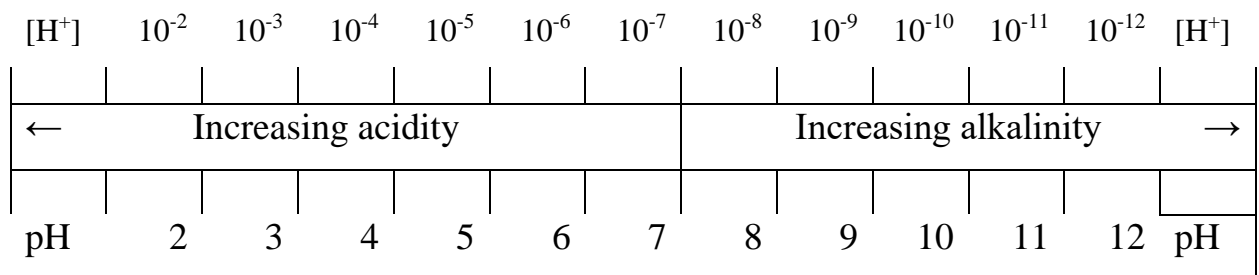


Fig. 2. Acidity and alkalinity scale.

In neutral solution, $[H^+]$ (often denoted by C_H) is, according to the foregoing, equal to 10^{-7} . Obviously, it is higher in an acid solution and lower in an alkaline. Thus, as a solution becomes more and more acid, C_H

will pass through the values 10^{-6} , 10^{-5} , 10^{-4} , etc. And, conversely, as the solution becomes more and more alkaline, we shall have: 10^{-8} , 10^{-9} , 10^{-10} , etc.

The quantitative designation of the acidity or alkalinity of a solution may be still further simplified by using the hydrogen ion concentration index (pH). This index is defined by the relation: $\text{pH} = -\log C_{\text{H}}$. Thus a neutral solution will have $\text{pH}=7$, an acid solution will have pH less than 7, and an alkaline solution will have pH more than 7.

In practice, the acidity or alkalinity of a solution is conveniently determined by means of indicators – substances which change colour depending on the relative concentrations of H^+ and OH^- ions. The best known indicator is litmus, which turns red when there is a surplus of H^+ ions, i.e., in acid solution, blue when there is a surplus of OH^- ions, i.e., in alkaline solution, and violet in neutral solution. If a filter paper impregnated with litmus (“litmus paper”) is immersed in the solution to be tested, its colour shows at once whether the solution is acid, alkaline or neutral.

The colour change of different indicators occurs at different hydrogen ion concentrations, which is important for chemical analysis. For example, litmus changes its colour (red to blue) at approximately $\text{pH}=7$, methyl orange (red to yellow) at $\text{pH}=4$, and phenolphthalein (colourless to pink) at $\text{pH}=9$.

For more precise measuring of pH it is widely used the special tools – pH -meters, which provides assurance of measuring within the limits of $\pm 0,01$.

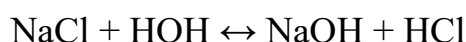
3. *General notion of Hydrolysis*

When various substances are dissolved in water a chemical reaction of an exchange nature often occurs. Such processes are grouped under the heading of **hydrolysis**.

In general, hydrolysis is an exchange reaction of substances with water. In this case, the equilibrium of the dissociated or a sparingly soluble product.

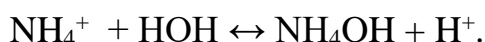
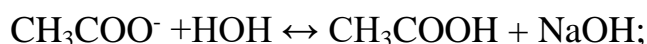
The following discussion will be confined to the hydrolysis of salts, which, in practice, is by far the most common case of hydrolysis.

When a salt formed by a strong base and a strong acid, such as NaCl, dissolves, the water dissociation equilibrium does not shift, since the ions of such a salt do not form slightly dissociated products with H⁺ and OH⁻ ions. Therefore, in the system

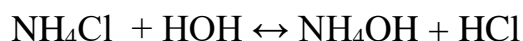
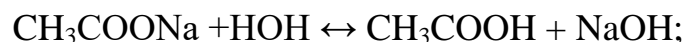


the only remaining slightly dissociated compound is water itself. As a result, the equilibrium of the reaction is almost wholly shifted to the left, i.e., practically no hydrolysis of NaCl occurs, and there is no appreciable surplus of either H⁺ or OH⁻ ions in the solution.

It is different when a salt of a strong base and a weak acid (CH₃COONa), or vice versa, (NH₄Cl) dissolves. In the former case, the H⁺ will be partially bound, while in the latter, the OH⁻ ion will be partially bound due to the equations:



Since both CH₃COOH and NH₄OH are much more dissociated than water, both these equilibria are greatly displaced to the left. Therefore, hydrolysis of the salts in accordance with the equations



Takes place only to an insignificant extent. Nevertheless, the first solution contains a certain surplus of OH^- ions, while the second solution contains a certain surplus of H^+ ions.

Obviously, if both substances forming the salt – the base and the acid – are slightly dissociated, the equilibrium of the hydrolysis.

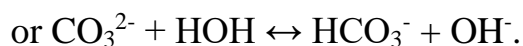
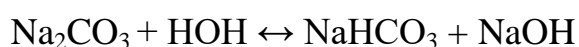
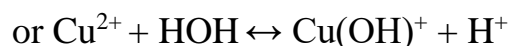
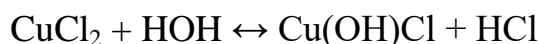
Since both CH_3COOH and NH_4OH are much more dissociated than water, both these equilibria are greatly displaced to the left. Therefore, hydrolysis of the salts in accordance with the equations



should be shifted further to the right. In other words, the hydrolysis of a salt forms by a weak base and a weak acid will be, generally speaking, more complete than if only one of these substances is slightly dissociated.

The acidity or alkalinity of solutions of such salts depends on the relative strength of the acid and the base. If their strengths are equal, the solution may be neutral, which is the case, for example, in the hydrolysis, of $\text{CH}_3\text{COONH}_4$. Thus, neutrality of a solution does not, in itself, prove that the salt is not hydrolysed.

In practice, one mostly has to deal with the hydrolysis of salts with a multiply charged ions of the weak constituent (acid or base) and singly charged ions of the strong constituent. Hydrolysis of such compounds, e.g., CuCl_2 and Na_2CO_3 results, as a rule, in basic or acidic salts, respectively:



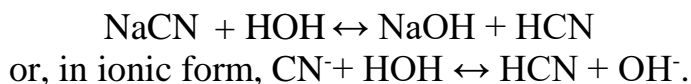
Hydrolysis of such salts does not usually proceed any further, i.e., to formation of the free weak base or acid, owing to the accumulation of H^+ or OH^- ions, respectively, in solution. Exceptions are those cases where the basic or acidic properties of the multivalent constituent are extremely weak. In such hydrolysis often proceeds practically to completion.

Different types of Hydrolysis

Individual possible cases of hydrolysis of salts are compared below (from the point of view of the acidity or alkalinity of the solution and the nature of the product obtained). The basis for the comparison is, on the one hand, the strength of the corresponding acids and bases, and, on the other hand, the type of salt with regard to valency of the cation and the anion.

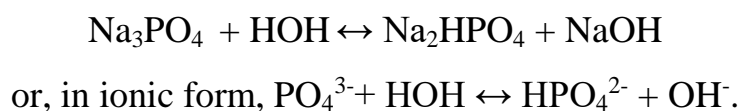
I. Strong base, weak acid. The solution is alkaline. The following cases are possible, depending on the valency of the cation or the anion.

a) Cation and anion monovalent. The hydrolysis products are a free acid and a free base, e.g.:

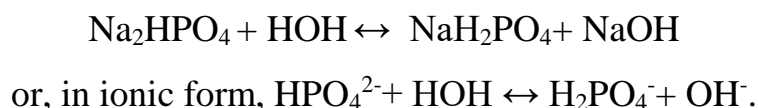


b) Cation monovalent, anion multivalent. This is a most typical case.

Hydrolysis gives acid salts and a free alkali, e.g.:



In a large volume of water hydrolysis proceeds further:

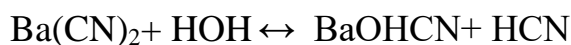


However, it does not reach the stage of formation of a free weak acid owing to the accumulation of free alkali (OH^- ions) in the solution.

c) Cation multivalent, anion monovalent. This case is rather rare.

Hydrolysis results in a basic salt and a free acid. But the solution is alkaline

since it contains more OH^- ions from the basic salt (formed by the strong base) than H^+ ions from the weak acid, e.g.:

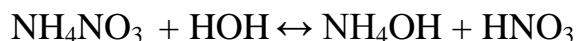


or, in ionic form, $\text{CN}^- + \text{HOH} \leftrightarrow \text{HCN} + \text{OH}^-$.

d) Cation and anion multivalent. This case is never encountered in practice, as the compounds of this group are all insoluble in water.

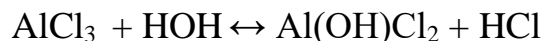
II. Weak base, strong acid. The solution is acid. The following cases are possible, depending on the valency of the cation or the anion.

a) Cation and anion monovalent. The hydrolysis gives a free acid and a free base, e.g.:



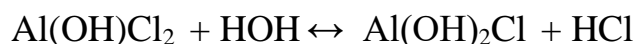
or, in ionic form, $\text{NH}_4^+ + \text{HOH} \leftrightarrow \text{NH}_4\text{OH} + \text{H}^+$.

b) Cation multivalent, anion monovalent. This is a most typical case. Hydrolysis results in basic salts and a free acid, e.g.:



or, in ionic form, $\text{Al}^{3+} + \text{HOH} \leftrightarrow \text{Al}(\text{OH})^{2+} + \text{H}^+$.

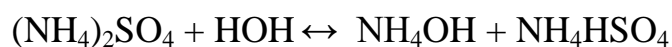
In a large volume of water hydrolysis proceeds further:



or, in ionic form, $\text{Al}(\text{OH})^{2+} + \text{HOH} \leftrightarrow \text{Al}(\text{OH})_2^{2+} + \text{H}^+$.

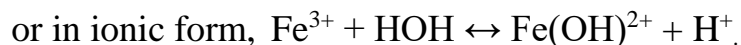
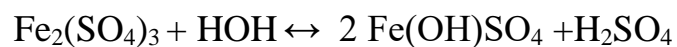
However, it does not reach the stage of formation of a free weak base owing to the accumulation of free acid (H^+ ions) in the solution.

c) Cation monovalent, anion multivalent. This is a very rare case. The result of hydrolysis is an acid salt and a free base. In this case, the solution contains more H^+ ions from the acid salt (formed by the strong acid) than OH^- ions from the weak base, e.g.:

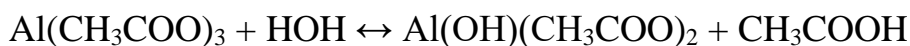


or, in ionic form, $\text{NH}_4^+ + \text{HOH} \leftrightarrow \text{NH}_4\text{OH} + \text{H}^+$.

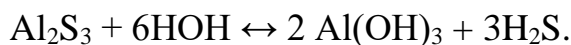
d) Cation and anion multivalent. Hydrolysis results in a basic salt and a free acid, e.g.:



III. Weak base, weak acid. In this case the alkalinity or acidity of the solution depends on the relative strengths of the acid and the base. The nature of the hydrolysis products depends on the strengths of the acid and the base as well as on the valencies of the cation and the anion. For example, aluminum acetate hydrolyses to form basic salts, according to the equations:



while aluminium sulphide, a salt of a much weaker acid, is hydrolysed to the free base and the free acid:

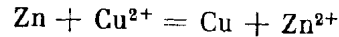


The last case concerns to the salts marked as symbol - (dash) in Table of salts and bases Solubility in water.

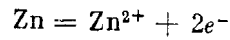
IV. Strong base, strong acid. Solutions of such salts are neutral to litmus, since they are practically unhydrolysed. This group includes NaCl, Na₂SO₄, BaCl₂ etc.

4. Chemical Sources of Electrical Energy

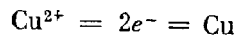
We know that in any redox reaction, electrons pass from The reducing agent to the oxidising one. For example, when a zinc plate is lowered into a copper sulphate solution, the following reaction occurs:



Here the reducing agent—zinc—loses electrons. This half-reaction is



The oxidising agent —the copper ion—gains electrons. The equation of this half-reaction is



In the above example, both half-reactions occur at the place of contact of the zinc with the solution so that the electrons pass directly from the zinc atoms to the copper ions. We can conduct this reaction in such a way, however, that the oxidation and reduction half-reactions will be separated in space, and the electrons will pass from the reducing agent to the oxidising one not directly, but via a conductor of an electric current—through an external circuit. This directed flow of electrons is an electric current. When a redox reaction is conducted in such a way, its energy will be converted into electrical energy that can be used by connecting a consumer of electrical energy (for example, an electrical heating appliance or an electrical lamp) to the external circuit.

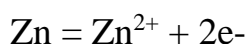
Devices used for the direct conversion of the energy of a chemical reaction into electrical energy are called galvanic (or voltaic) cells. They are also known as chemical sources of electrical energy or chemical sources of current.

It is customary practice in engineering to apply the name galvanic

cell only to a. chemical source of current in which virtually irreversible reactions proceed. Such current sources cannot usually be recharged: they are intended for use only once (in one or more stages). Chemical sources of current in which virtually reversible reactions occur are called accumulators: they can be recharged and used repeatedly.

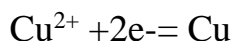
The functioning of any galvanic cell is based on the proceeding of a redox reaction in it. The simplest galvanic cell consists of two plates or rods made from different metals and immersed in a solution of an electrolyte. This system makes possible the separation of the redox reaction in space: oxidation occurs on one metal, and reduction on the other. Thus, the electrons are transferred from the reducing agent to the oxidising one via the external circuit. Let us consider as an example a copper-zinc galvanic cell operating at the expense of the energy of the reaction between zinc and copper sulphate described above. This cell (a Jacobi-Daniell cell) consists of a copper plate immersed in a solution of copper sulphate solution (a copper electrode) and a zinc plate immersed in a zinc sulphate solution (a zinc electrode). The two solutions are in contact with each other, but to prevent mixing, they are separated by a partition made from a porous material.

In operation of the cell, i.e. when the circuit is closed, the zinc becomes oxidised: on the surface of its contact with the solution, the zinc atoms transform into ions and after becoming hydrated pass into the solution. The liberated electrons travel through the external circuit to the copper electrode. The entire collection of these processes is depicted schematically by a half-reaction equation, or an electro-chemical equation:



At the copper electrode, reduction of the copper ions takes place. The

electrons arriving here from the zinc electrode combine with the copper ions becoming dehydrated and leaving the solution; copper atoms are produced that separate in the form of metal. The corresponding electrochemical equation has the form



The net equation of the reaction proceeding in the cell is obtained when the equations of the two half-reactions are summated. Thus, in the operation of a galvanic cell, the electrons from the reducing agent pass to the oxidising agent through the external circuit, electrochemical processes occur at the electrodes, and the directed motion of the ions is observed in the solution.

The direction of motion of the ions in the solution is due to the electrochemical processes occurring at the electrodes. We have already indicated that at the zinc electrode the cations emerge into the solution, creating an excess positive charge in it, while at the copper electrode the solution, conversely, constantly becomes leaner in cations so that here the solution is charged negatively. The result is the setting up of an electric field in which the cations in the solution (Cu^{2+} and Zn^{2+}) move from the zinc electrode to the copper one, while the anions— SO_4^{2-} — move in the opposite direction. As a result, the liquid at both electrodes remains electrically neutral.

The motion of the electrons and ions in the operation of a copper zinc cell is shown schematically (fig. 3).

The electrode at which oxidation occurs is called the **anode**. The one at which reduction occurs is called the **cathode**. In the copper zinc cell, the zinc electrode is the anode, and the copper one—the cathode.

The redox reaction proceeding in a galvanic cell is a complicated process. It includes the electrochemical steps proper (the transformations

of atoms, ions, or molecules at the electrodes), the transfer of electrons, and the transfer of ions. All these steps are interlinked and proceed at the same rate. The number of electrons given up by the zinc in unit time equals the number of electrons received during this time by the copper ions. Hence, the rate of a reaction proceeding in a galvanic cell is proportional to the quantity of electricity transferred through the circuit in unit time, i.e. to the current in the circuit.

The electric current flowing in the external circuit of a galvanic cell can do useful work. But the work that can be done at the expense of the energy of a chemical reaction depends on its rate— it is the greatest when the reaction is conducted infinitely slowly, reversibly. Consequently, the work that can be done at the expense of the reaction proceeding in a galvanic cell depends on the magnitude of the current taken from it. If we lower the current in the external circuit to an infinitely small value by increasing the resistance of the circuit, the rate of the reaction in the cell will also be infinitely small, and the work will be maximum. The heat evolved in the internal circuit of the cell, on the contrary, will be minimum.

The work of an electric current is expressed by the product of the quantity of electricity flowing through, the circuit and the voltage V .

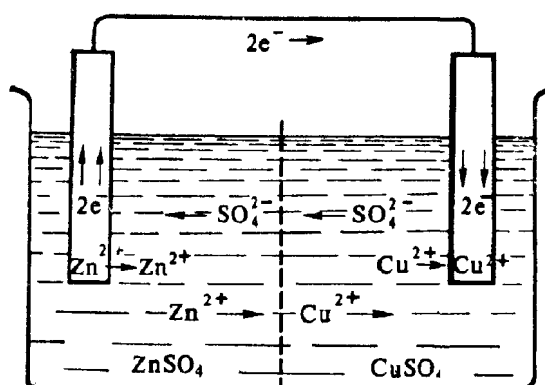


Fig. 3. Motion and electrons in the operation of a copper-zinc

In a copper-zinc cell, when one equivalent of zinc is oxidised and simultaneously one equivalent of copper ions is reduced, a quantity of electricity equal to one faraday ($1 F = eN_A = 96\,485$ coulombs*) where e is the elementary electric charge and N_A is the Avogadro constant) will pass through the circuit. Hence, the useful work W which a current can do will be

$$W' = FV$$

where F = Faraday's constant, numerically equal to one faraday V = voltage between the poles of the cell.

But since this work depends on the current, the voltage between the poles of the cell also depends on the current (F is a constant). In the limiting case corresponding to the reversible proceeding of the reaction, the voltage will be maximum. The maximum value of the voltage of a galvanic cell corresponding to the reversible proceeding of the reaction is called the electromotive force (e.m.f.) of the given cell.

For this limiting case, the useful work done by the current in a copper-zinc cell when one equivalent of zinc reacts with one equivalent of copper ions is expressed by the equation

$$W'_{\max} = FV_{\max} = FE$$

where $E = V_{\max}$ is the e.m.f. of the cell.

It is clear that when one mole of zinc atoms reacts with one mole of copper ions, the equation becomes

$$W_{\max} = 2W'_{\max} = 2FE$$

In the general case upon the dissolving (or separation) of one mole of a substance whose ions have a charge of z , the maximum useful work is related to the e.m.f. by the equation

$$W_{\max} = zFE$$

At a constant temperature and pressure, the maximum useful work of a reaction equals the change in the Gibbs energy ΔG taken with the reverse sign. Hence:

$$\Delta G = -zFE$$

If the concentrations (more exactly, the activities) of the substances participating in a reaction equal unity, i.e. if standard conditions are observed, the e.m.f. of a cell is called its standard electromotive force and is designated by the symbol E° . The last equation now becomes:

$$\Delta G^\circ = -zFE^\circ$$

* In calculations, we shall use the value of this quantity approximated to three significant digits (96500 C/mol).

With a view to the standard change in the Gibbs energy of a reaction being associated with its equilibrium constant (see Sec. 68) by the expression

$$\Delta G^\circ = -RT \ln K$$

we get an equation relating the standard e.m.f. to the equilibrium constant of a reaction proceeding in a galvanic cell- E.m.f. can be measured with a high accuracy.

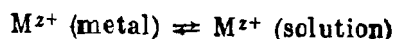
$$RT \ln K = zFE^\circ$$

These measurements are one of the most accurate ways of finding the standard Gibbs energies and, consequently, the equilibrium constants for oxidation-reduction reactions in solutions.

Redox reactions proceed in a galvanic cell notwithstanding the fact that the oxidising and reducing agents are not in direct contact with each other. To gain an understanding of how this occurs, of how an e.m.f. appears when the processes of oxidation and reduction are separated in space, let us consider in greater detail the phenomena occurring at the phase interfaces

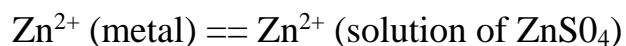
in a galvanic cell.

Direct experiments with the use of radioactive indicators show that if we bring a metal (M) into contact with a solution of its salt, then the metal ions (M^{z+}) pass from the metal phase into the solution and from the solution into the metal. Since the energy state of the ions in these phases is not the same, then at the first moment after contact has been established, the metal ions pass from the metal into the solution and in the reverse direction at different rates. If transition of the ions from the metal phase into the solution predominates, the solution acquires a positive charge, while the metal electrode is charged negatively. As these charges increase, the transition of the cations into the solution having a like charge is hindered so that the rate of this process diminishes. On the other hand, the rate of transition of the cations from the solution onto the negatively charged electrode grows. As a result, the rates of the two processes level out, and equilibrium sets in between the metal and the solution:

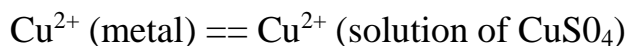


The metal electrode is charged negatively, and the solution positively. If, when contact is established between a metal and a solution, the rate of transition of the cations from the metal into the solution was lower than the rate of their transition in the reverse direction, equilibrium also sets in between the electrode and the solution. Here, however, the electrode is charged positively, and the solution negatively.

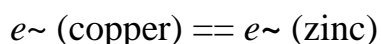
In the Jacobi-Daniel cell, the corresponding equilibria set in between the zinc electrode and the zinc sulphate solution:



and also between the copper electrode and the copper sulphate solution:



This cell has two other phase interfaces: between the solutions of zinc and copper sulphates, and also between the copper and the zinc. The interface between the solutions does not appreciably affect either the magnitude of the e.m.f. or the proceeding of the reaction when the cell is operating. As regards the interface between the metals, electrons can pass through it instead of ions as at a metal-solution interface*. Here too, owing to the different energy state of the electrons in the copper and in the zinc, the initial rates of electron transition from one metal into the other and in the reverse direction are different. In this case, equilibrium also sets in rapidly, however, and the metals also acquire charges of the opposite sign:



Consequently, when the circuit is open, equilibria set in on the three phase interfaces in the Jacobi-Daniell cell, the phases becoming charged. As a result, the energy state of the electrons at the ends of the open circuit becomes different: on the copper conductor in contact with the zinc electrode the Gibbs energy of the electrons is higher, and on the end connected to the copper electrode it is lower. It is exactly the difference, between the Gibbs energies of the electrons at the ends of the circuit that determines the e.m.f. of a given cell.

When the external circuit is closed, the electrons move from the zinc electrode to the copper one. Therefore, equilibria at the phase interfaces are violated. A directed transition of the zinc ions takes place from the metal into the solution, of the copper ions—from the solution into the metal, and of the electrons—from the zinc to the copper. A redox reaction occurs.

In principle, any redox reaction can yield electrical energy. The number of reactions practically used in chemical sources of electrical energy,

however, is not great. This is associated with the circumstance that not any redox reaction makes it possible to create a galvanic cell having technically valuable properties (a high and virtually constant e.m.f., the possibility of obtaining high currents, a long lifetime, etc.). In addition, many redox reactions require the use of costly substances.

Unlike the copper-zinc cell, all modern galvanic cells and accumulators use not two, but one electrolyte. Such current sources are considerably more convenient in operation. In virtually all the galvanic cells produced at present, the anode is made from zinc, while oxides of less active metals are used as the substance for the cathode.

For a description of the most important galvanic cells and of accumulators see.

In means of communication (radio, telephone, telegraph) and in electrical measuring apparatus, they are sources of electrical power, on motor vehicles and aeroplanes they are used for actuating starters and other equipment, on transport and in many other fields they are used in portable lanterns for illumination. With the growing shortage of petroleum products in many countries, greater and greater attention is being given to the development of electric vehicles powered by accumulators.

All conventional chemical current sources are not free of two shortcomings. First, the cost of the substances needed for their operation (for example, lead and cadmium) is high. Second, the ratio of the amount of energy that a cell can give up to its mass is low. The last few decades have seen the conducting of investigations aimed at creating cells whose operation would require the consumption of cheap substances having a low density, similar to liquid or gaseous fuel (natural gas, kerosene, hydrogen, etc.). Such galvanic cells are known as fuel cells. Much attention is being given to the fuel cell problem at present, and it can be assumed that in the

nearest future fuel cells will find broad application.

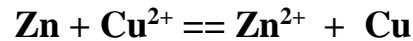
5. *Electrode Potentials*

Every oxidation-reduction reaction consists of oxidation and reduction half-reactions. When a reaction proceeds in a galvanic cell or is conducted by electrolysis, each half-reaction occurs at the corresponding electrode. This is why half-reactions are also called electrode processes.

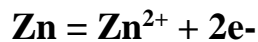
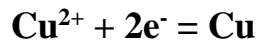
We showed earlier that the e.m.f. E of a galvanic cell corresponds to the redox reaction proceeding in this cell. The e.m.f. is related to the change in the Gibbs energy of the reaction ΔG by the equation

$$\Delta G = -zFE$$

In accordance with the division of a redox reaction into two half-reactions, it is customary practice to represent the e.m.f.'s too in the form of the difference between two quantities, each of which corresponds to the given half-reaction. These quantities are known as the electrode potentials. For a copper-zinc cell, the reaction proceeding during its operation:



s divided into two half-reactions:



Accordingly, the e.m.f. of this element (E) can be represented as the difference between the electrode potentials (φ), one of which (φ_1) corresponds to the first, and the other (φ_2) to the second of the above half-reactions:

$$E = \varphi_1 - \varphi_2$$

The change in the Gibbs energy ΔG_1 that corresponds to the thermodynamically reversible reduction of one mole of copper ions is

$$\Delta G_1 = -2F\varphi_1$$

while the change in the Gibbs energy ΔG_2 corresponding to the thermodynamically reversible oxidation of one mole of zinc atoms is

$$\Delta G_2 = -2F\varphi_2$$

In the general case, the electrode potential φ and a change in the Gibbs energy ΔG equal to

$$\Delta G = -zF\varphi$$

correspond to any electrode process $\text{Ox} + ze^- = \text{Red}$

where Ox and Red stand for the oxidised and reduced forms, respectively, of the substances participating in the electrode process.

In the following, when dealing with electrode processes, we shall write their equations in the direction of reduction (except, naturally, when we are speaking exactly about oxidation).

Investigation of the potentials of various electrode processes has shown that their magnitudes depend on the following three factors:

- (1) the nature of the substances participating in the electrode process;
- (2) the ratio between the concentrations* of these substances, and
- (3) the temperature of the system. This relationship is expressed by the equation:

$$\varphi = \varphi^\circ + \frac{2.3RT}{zF} \log \frac{[\text{Ox}]}{[\text{Red}]}$$

Strictly speaking, the magnitude of an electrode potential depends on the ratio of the activities and not of the concentrations of substances. In all the following equations, the activity should be substituted for the concentration. But at low concentrations of solutions, the error introduced by substituting the concentration for the activity is not great.

Here φ° is the standard electrode potential of a given process— a

constant whose physical meaning is considered below; R is the molar gas constant; T is the absolute temperature; z is the number of electrons participating in the process; F is the Faraday constant; $[Ox]$ and $[Red]$ are the products of the concentrations of the substances participating in the process in the oxidised (Ox) and reduced (Red) forms.

The physical meaning of the quantity φ° becomes clear if we consider the case when the concentrations (activities) of all the substances participating in a given electrode process are unity. For this condition, the second addend in the right-hand side of the equation vanishes ($\log 1=0$), and the equation becomes

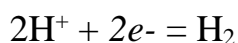
$$\varphi = \varphi^\circ$$

Concentrations (activities) equal to unity are called standard concentrations (activities). Therefore, the potential corresponding to this case is also called the standard potential. Thus, the standard electrode potential is the potential of a given electrode process at concentrations (more accurately, at activities) of all the substances participating in it equal to unity.

Hence, the first addend in the equation of the electrode potential (φ°) takes into account the influence of the nature of the substances on its magnitude, and the second addend $2.3RT/zF \log[Ox]/[Red]$ the influence of their concentrations. In addition, both terms vary with the temperature.

For the customary standard temperature used in electrochemical measurements ($25\text{ }^\circ\text{C} = 298\text{ K}$), when the values of the constant quantities are introduced [$R = 8.31\text{ J}/(\text{mol}\cdot\text{K})$, $F = 96\,500\text{ C}/\text{mol}$], the equation becomes. To construct a numerical scale of the electrode potentials, it is necessary to assume that the potential of a definite electrode process is

zero. The following electrode process has been taken as the standard for constructing such a scale:



The change in the Gibbs energy associated with the proceeding of this half-reaction in standard conditions is adopted equal to zero. Accordingly, the standard potential of this electrode process is also taken equal to zero. All electrode potentials indicated in this book, and also in the majority of other modern publications, are expressed according to this hydrogen scale.

The above electrode process is carried out on a hydrogen electrode. The latter is a platinum plate electrolytically coated with spongy platinum and immersed in a solution of an acid through which hydrogen is passed. The hydrogen dissolves well in the platinum; the hydrogen molecules partly decompose into atoms (the platinum catalyses this decomposition). Oxidation of the hydrogen atoms or reduction of the hydrogen ions can proceed on the surface of contact of the platinum with the acid solution. The platinum does not virtually participate in the electrode reactions and plays, as it were, the role of a sponge impregnated with atomic hydrogen.

The potential of the hydrogen electrode is reproduced with a very high accuracy. This is exactly why the hydrogen electrode has been taken as the standard in creating a scale of electrode potentials.

Let us establish the form of the general equation of the electrode potential for the hydrogen electrode. For this electrode, $z = 2$, $[\text{Ox}] = [\text{H}^+]^2$, $[\text{Red}] = [\text{H}_2]$. The concentration of the hydrogen dissolved in the platinum is proportional to its partial pressure $P(\text{H}_2)$:

$$[\text{H}_2] = kp(\text{H}_2)$$

where k is a quantity that is constant at the given temperature. Using the

equation of an electrode process and introducing the constant k into the value of φ° , we get

$$\varphi = \varphi^\circ + 0.059 \log [\text{H}^+] - 0.030 \log p(\text{H}_2)$$

The partial pressure of hydrogen $P(\text{H}_2)$ is usually kept equal to standard atmospheric pressure, which is conventionally taken as unity. In this case, the last term of the equation obtained vanishes ($\log 1 = 0$). Hence:

$$\varphi = \varphi^\circ + 0.059 \log [\text{H}^+]$$

Since the standard potential of the process being considered is assumed to equal zero, then

$$\varphi = 0.059 \log [\text{H}^+]$$

or since $\log [\text{H}^+] = -\text{pH}$, we finally get

$$\varphi = -0.059 \text{ pH}$$

To find the potential of an electrode process, it is necessary to form a galvanic cell from the electrode being tested and a standard hydrogen electrode and measure its e.m.f. Seeing that the potential of a standard hydrogen electrode is zero, the measured e.m.f. will be the potential of the given electrode process.

A standard hydrogen electrode is not generally used in practice as a reference electrode because this involves considerable complications. More convenient electrodes are employed whose potentials in comparison with the standard hydrogen electrode are known. The e.m.f. of the cell should be calculated by the equation

$$E = |\varphi_{\text{ref}} - \varphi_x|$$

Here E is the e.m.f. of the cell, φ_{ref} is the known potential of the reference electrode, and φ_x is the potential of the electrode being tested.

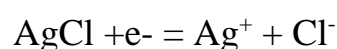
Solving the equation with respect to φ_x , we get:
when $\varphi_{\text{ref}} > \varphi_x$

$$\varphi_x = \varphi_{\text{ref}} - E$$

when $\varphi_{\text{ref}} < \varphi_x$

$$\varphi_x = \varphi_{\text{ref}} + E$$

The reference electrodes in greatest favour are the silver-silver chloride and the calomel electrodes. The former is a silver wire coated with a layer of AgCl and immersed in a solution of hydrochloric acid or its salt. The following reaction proceeds in it when the circuit is closed:



The calomel electrode consists of mercury coated with a suspension of calomel (Hg_2Cl_2) in a solution of KCl. The potentials of these electrodes are reproduced with a high accuracy.

To find the value of the electrode potential, it is necessary to measure the voltage of the operating cell, but its e.m.f. When measuring the latter, the resistance of the external circuit (i.e. of the measuring device) is very high. Virtually no reaction proceeds in the cell. Thus, the electrode potentials correspond to the reversible occurring of processes or, which is the same, to the state of electrochemical equilibrium at the electrodes.

Consequently, electrode potentials are often called equilibrium electrode potentials or simply equilibrium potentials.

Let us now consider the form which the general equation of the electrode potential will acquire in the most important cases.

1. The electrode process is expressed by the equation



where M stands for the atoms of a metal, and M^{z+} are its z-charged ions.

This case includes both electrodes of a copper-zinc cell and in general any metal electrode in a solution of a salt of the same metal. Here the metal

ions are the oxidised form of the metal, and its atoms are the reduced form. Hence, $[Ox] = [M^{z+}]$, and $[Red] = \text{const}$ because the concentration of the atoms in a metal is a constant quantity at a constant temperature. Including the value of this constant in the quantity φ° , we get:

$$\varphi = \varphi^\circ + \frac{0.059}{z} \log [M^{z+}]$$

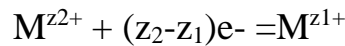
For instance, for the process $Ag^+ + e^- = Ag$, we have:

$$\varphi = 0.799 + 0.059 \log [Ag^+]$$

and for the process $Zn^{2+} + 2e^- = Zn$, we get:

$$\varphi = -0.763 + 0.030 \log [Zn^{2+}]$$

2. The electrode process is expressed by the equations



Here both the oxidised (M^{z_2+}) and reduced (M^{z_1+}) forms of the metal are in the solution, and their concentrations are variable quantities. Consequently,

$$\varphi = \varphi^\circ + \frac{0.059}{z_2 - z_1} \log \frac{[M^{z_2+}]}{[M^{z_1+}]}$$

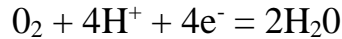
$$\varphi = 0.771 + 0.059 \log \frac{[Fe^{3+}]}{[Fe^{2+}]}$$

In this and in the cases treated below, the electrode at which the electrode process occurs is made from an inert material. Most often, this material is platinum.

We have considered examples when only ions consisting of one element participated in the electrode processes. Quite often, however, the substance becoming oxidised or reduced consists not of one, but of two or more elements. The oxidising agent most frequently contains oxygen; here water and the products of its dissociation—hydrogen ions (in an acid solution) or

hydroxide ions (in a basic solution) generally participate in the electrode process. Let us see what the equations of the electrode process potentials will be like in such cases.

3. The electrode process is expressed by the equation:



This half-reaction (when it proceeds in the direction of reduction) plays a very great part in the corrosion of metals. Oxygen is the most widespread oxidising agent causing metals to corrode in aqueous solutions.

In the above electrode process, oxygen is reduced with the participation of hydrogen ions to form water. Hence, $[Red] = [H_2O]^2$, and $[Ox] = [O_2][H^+]^4$. The concentration of the water in dilute solutions may be considered constant. The concentration of oxygen in a solution is proportional to its partial pressure over the solution $\{[O_2] = kp(O_2)\}$. After performing the required transformations and designating the sum of the constant quantities by φ° , we get

$$\varphi = \varphi^\circ + 0.059 \log [H^+] + \frac{0.059}{4} \log p(O_2)$$

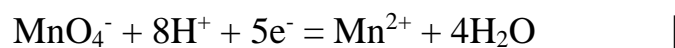
For the process being considered, $\varphi^\circ = 1.228$ V; hence,

$$\varphi = 1.228 - 0.059 \text{ pH} + 0.015 \log p(O_2)$$

At a partial pressure of the oxygen equal to standard atmospheric pressure (which is conventionally taken equal to unity), $\log p(O_2) = 0$, and the last equation becomes

$$\varphi = 1.228 - 0.059 \text{ pH}$$

4. For electrode processes represented by more complicated equations, the expressions for the potentials contain a greater number of variable concentrations. Let us consider, for example, the electrode process



This half-reaction proceeds (in the direction of reduction) when potassium permanganate reacts with most reducing agents in an acid solution.

The concentrations of all the substances participating in this electrode process (except water) are variable quantities. For this process, $\varphi^0 = 1.507$ V. The equation of the electrode potential has the form

$$\begin{aligned}\varphi &= \varphi^0 + \frac{0.059}{5} \log \frac{[\text{MnO}_4^-]}{[\text{Mn}^{2+}]} + \frac{8 \times 0.059}{5} \log [\text{H}^+] = \\ &= 1.507 + 0.012 \log \frac{[\text{MnO}_4^-]}{[\text{Mn}^{2+}]} - 0.095 \text{ pH}\end{aligned}$$

Examples 3 and 4 show that for electrochemical processes occurring with the participation of water, the hydrogen ion concentration is in the numerator of the logarithmic term of the potential equation. For this reason, the electrode potentials of such processes depend on the pH of the solution and their values grow with an increasing acidity of the solution.

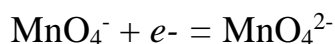
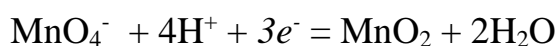
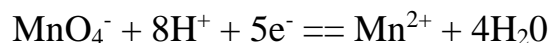
We have already mentioned that the dependence of the electrode potential on the nature of the substances participating in an electrode process is described by the quantity φ^0 . In this connection, it is customary practice to arrange all electrode processes in a series according to the magnitude of their standard potentials. Table. 2 gives the equations of the most important electrode processes and the relevant electrode potentials in the order of increasing values of φ^0 .

The position of an electrochemical system in this series characterises its oxidation-reduction ability. By an electrochemical system here is meant the collection of all the substances participating in the given electrode process.

The oxidation-reduction ability characterises an electrochemical system, but the oxidation-reduction ability of a substance (or ion) is also often spoken about. It must be borne in mind, however, that many substances can be oxidised or reduced to different products. For instance, potassium

permanganate (the ion MnO_4^-) can, depending on the conditions and first of all on the pH of the solution, be reduced either to the ion Mn^{2+} , or to MnO_2 , or to the ion MnO_4^{2-} .

The corresponding electrode processes are expressed by the equations:



Since the standard potentials of these three electrode processes are different (see Table. 2), the positions of these three systems in the series of φ^0 are also different. Hence, the same oxidising agent (MnO_4^-) can occupy several positions in the series of standard potentials.

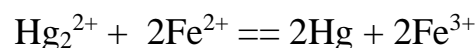
Elements exhibiting only one oxidation number in their compounds have simple oxidation-reduction characteristics and occupy a small number of positions in the series of standard potentials. They chiefly include the metals of the main subgroups of Groups I-III of the periodic table. On the other hand, many positions in the series of φ^0 are occupied by the elements that form compounds with different oxidation numbers—non-metals and many metals of the auxiliary subgroups of the periodic table.

The series of standard electrode potentials makes it possible to solve the problem about the direction of spontaneous redox reactions.

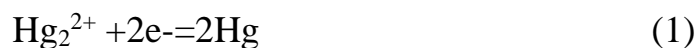
In the general case of any chemical reaction, the determining factor here is the sign of the change in the Gibbs energy of the reaction. If we form a galvanic cell from two electrochemical systems, then in its operation the electrons will pass spontaneously from the negative pole of the cell to the positive one, i.e. from an electrochemical system with a lower value of the electrode potential to a system with a higher value of it. But this signifies that the first system will play the role of a reducing agent, and the second

that of an oxidising agent. Consequently, in a galvanic cell, *an oxidation-reduction reaction can proceed spontaneously in a direction in which the electrochemical system with a higher value of the electrode potential plays the role of an oxidising agent, i.e. becomes reduced.* Upon the direct reaction of the substances, the possible direction of the reaction will naturally be the same as when it is conducted in a galvanic cell.

If the oxidising and reducing agents are far from each other in the series of φ° , the direction of the reaction is practically completely determined by their mutual position in the series. For example, zinc ($\varphi^\circ = -0.763$ V) will displace copper ($\varphi^\circ = +0.337$ V) from an aqueous solution of its salt at any practically achievable concentration of the solution. If, on the other hand, the values of φ° for the oxidising and reducing agents are close to each other, then when solving the problem of the direction in which the reaction will proceed spontaneously, it is also necessary to take account of how the concentrations of the relevant substances affect the electrode potentials. For example, the reaction can proceed spontaneously either from the left to the right or from the right to the left.



Its direction is determined by the concentrations of the iron and mercury ions. Two electrochemical systems participate in this reaction:



The following potentials correspond to the relevant electrode processes:

$$\varphi_1 = 0.788 + 0.030 \log [\text{Hg}_2^{2+}]$$

$$\varphi_2 = 0.771 + 0.059 \log \frac{[\text{Fe}^{3+}]}{[\text{Fe}^{2+}]}$$

Let us calculate the values of φ_1 and φ_2 at $[\text{Hg}_2^{2+}] = [\text{Fe}^{2+}] = 10^{-1}$ and $[\text{Fe}^{3+}] = 10^{-4}$ mol/1000 g of H_2O :

$$\varphi_1 = 0.788 + 0.030 \log 10^{-1} = 0.788 - 0.030 \approx 0.76 \text{ V}$$

$$\varphi_2 = 0.771 + 0.059 \log \frac{10^{-4}}{10^{-1}} = 0.771 - 3 \times 0.059 \approx 0.59 \text{ V}$$

Table 2.

Electrode Potentials in Aqueous Solutions at 25 °C and at a Partial Pressure of the Gases Equal to Standard Atmospheric Pressure

Electrode process	Equation of electrode potential
$\text{Li}^+ + e^- = \text{Li}$	$\varphi = -3.045 + 0.059 \log [\text{Li}^+]$
$\text{Rb}^+ + e^- = \text{Rb}$	$\varphi = -2.925 + 0.059 \log [\text{Rb}^+]$
$\text{K}^+ + e^- = \text{K}$	$\varphi = -2.924 + 0.059 \log [\text{K}^+]$
$\text{Cs}^+ + e^- = \text{Cs}$	$\varphi = -2.923 + 0.059 \log [\text{Cs}^+]$
$\text{Ca}^{2+} + 2e^- = \text{Ca}$	$\varphi = -2.866 + 0.030 \log [\text{Ca}^{2+}]$
$\text{Na}^+ + e^- = \text{Na}$	$\varphi = -2.714 + 0.059 \log [\text{Na}^+]$
$\text{Mg}^{2+} + 2e^- = \text{Mg}$	$\varphi = -2.363 + 0.030 \log [\text{Mg}^{2+}]$
$\text{H}_2 + 2e^- = 2\text{H}^-$	$\varphi = -2.251 - 0.059 \log [\text{H}^-]$
$\text{Al}^{3+} + 3e^- = \text{Al}$	$\varphi = -1.663 + 0.020 \log [\text{Al}^{3+}]$
$\text{Ti}^{2+} + 2e^- = \text{Ti}$	$\varphi = -1.630 + 0.030 \log [\text{Ti}^{2+}]$
$\text{Mn}^{2+} + 2e^- = \text{Mn}$	$\varphi = -1.179 + 0.030 \log [\text{Mn}^{2+}]$
$\text{Cr}^{2+} + 2e^- = \text{Cr}$	$\varphi = -0.913 + 0.030 \log [\text{Cr}^{2+}]$
$\text{Zn}^{2+} + 2e^- = \text{Zn}$	$\varphi = -0.763 + 0.030 \log [\text{Zn}^{2+}]$
$\text{Cr}^{3+} + 3e^- = \text{Cr}$	$\varphi = -0.744 + 0.020 \log [\text{Cr}^{3+}]$
$[\text{Au}(\text{CN})_2]^- + e^- = \text{Au} + 2\text{CN}^-$	$\varphi = -0.61 + 0.059 \log \frac{[\text{Au}(\text{CN})_2]^-}{[\text{CN}^-]^2}$
$2\text{H}^+ + 2e^- = \text{H}_2$	$\varphi = -0.059 \text{ pH}$
$\text{I}_2(\text{c}) + 2e^- = 2\text{I}^-$	$\varphi = -0.536 - 0.059 \log [\text{I}^-]$
$\text{MnO}_4^- + e^- = \text{MnO}_4^{2-}$	$\varphi = 0.564 + 0.059 \log \frac{[\text{MnO}_4^-]}{[\text{MnO}_4^{2-}]}$
$\text{Fe}^{3+} + e^- = \text{Fe}^{2+}$	$\varphi = 0.771 + 0.059 \log \frac{[\text{Fe}^{3+}]}{[\text{Fe}^{2+}]}$
$\text{NO}_3^- + 2\text{H}^+ + e^- = \text{NO}_2 + \text{H}_2\text{O}$	$\varphi = 0.78 + 0.059 \log [\text{NO}_3^-] - 0.118 \text{ pH}$
$\text{Hg}_2^{2+} + 2e^- = 2\text{Hg}$	$\varphi = 0.788 + 0.030 \log [\text{Hg}_2^{2+}]$
$\text{Ag}^+ + e^- = \text{Ag}$	$\varphi = 0.799 + 0.059 \log [\text{Ag}^+]$
$\text{Hg}^{2+} + 2e^- = \text{Hg}$	$\varphi = 0.850 + 0.030 \log [\text{Hg}^{2+}]$
$2\text{Hg}^{2+} + 2e^- = \text{Hg}_2^{2+}$	$\varphi = 0.920 + 0.030 \log \frac{[\text{Hg}_2^{2+}]}{[\text{Hg}^{2+}]^2}$
$\text{Br}_2(\text{lq}) + 2e^- = 2\text{Br}^-$	$\varphi = 1.065 - 0.059 \log [\text{Br}^-]$
$\text{Pt}^{2+} + 2e^- = \text{Pt}$	$\varphi = 1.188 + 0.030 \log [\text{Pt}^{2+}]$
$\text{O}_2 + 4\text{H}^+ + 4e^- = 2\text{H}_2\text{O}$	$\varphi = 1.228 - 0.059 \text{ pH}$
$\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 6e^- = 2\text{Cr}^{3+} + 7\text{H}_2\text{O}$	$\varphi = 1.333 + 0.010 \log \frac{[\text{Cr}_2\text{O}_7^{2-}]}{[\text{Cr}^{3+}]^2} - 0.138 \text{ pH}$
$\text{Cl}_2(\text{g}) + 2e^- = 2\text{Cl}^-$	$\varphi = 1.359 - 0.059 \log [\text{Cl}^-]$
$\text{PbO}_2 + 4\text{H}^+ + 2e^- = \text{Pb}^{2+} + 2\text{H}_2\text{O}$	$\varphi = 1.449 - 0.030 \log [\text{Pb}^{2+}] - 0.118 \text{ pH}$
$\text{Au}^{3+} + 3e^- = \text{Au}$	$\varphi = 1.498 + 0.020 \log [\text{Au}^{3+}]$
$\text{MnO}_4^- + 8\text{H}^+ + 5e^- = \text{Mn}^{2+} + 4\text{H}_2\text{O}$	$\varphi = 1.507 + 0.012 \log \frac{[\text{MnO}_4^-]}{[\text{Mn}^{2+}]} - 0.095 \text{ pH}$
$\text{PbO}_2 + \text{SO}_4^{2-} + 4\text{H}^+ + 2e^- = \text{PbSO}_4 + 2\text{H}_2\text{O}$	$\varphi = 1.685 + 0.030 \log [\text{SO}_4^{2-}] - 0.118 \text{ pH}$
$\text{Au}^+ + e^- = \text{Au}$	$\varphi = 1.692 + 0.059 \log [\text{Au}^+]$
$\text{MnO}_4^- + 4\text{H}^+ + 3e^- = \text{MnO}_2 + 2\text{H}_2\text{O}$	$\varphi = 1.692 + 0.020 \log [\text{MnO}_4^-] - 0.079 \text{ pH}$
$\text{H}_2\text{O}_2 + 2\text{H}^+ + 2e^- = 2\text{H}_2\text{O}$	$\varphi = 1.776 + 0.030 \log [\text{H}_2\text{O}_2] - 0.059 \text{ pH}$
$\text{S}_2\text{O}_8^{2-} + 2e^- = 2\text{SO}_4^{2-}$	$\varphi = 2.010 + 0.030 \log \frac{[\text{S}_2\text{O}_8^{2-}]}{[\text{SO}_4^{2-}]^2}$
$\text{F}_2 + 2e^- = 2\text{F}^-$	$\varphi = 2.87 - 0.059 \log [\text{F}^-]$

6. Oxidation-Reduction Reactions. Electrochemistry

Thus, with the ratio of the concentrations we have taken, $\varphi_1 > \varphi_2$ and the reaction proceeds from left to right.

Now let us calculate φ_1 and φ_2 for the reverse ratio of the concentrations. Assume that $[\text{Hg}_2^{2+}] - [\text{Fe}^{2+}] = 10^{-4}$, and $[\text{Fe}^{3+}] = 10^{-1}$ mol/1000 g of H_2O :

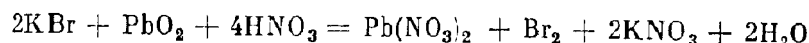
$$\begin{aligned}\varphi_1 &= 0.788 + 0.030 \log 10^{-4} = 0.788 - 4 \times 0.030 \approx 0.67 \text{ V} \\ \varphi_2 &= 0.771 + 0.059 \log \frac{10^{-1}}{10^{-4}} = 0.771 + 3 \times 0.059 \approx 0.95 \text{ V}\end{aligned}$$

Consequently, at these concentrations, $\varphi_1 > \varphi_2$, and the reaction proceeds from right to left.

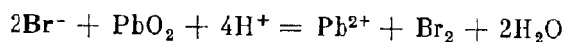
If a redox reaction proceeds with the participation of water and hydrogen or hydroxide ions, then the pH of the solution must also be taken into consideration.

Table. 2 includes 38 half-reactions. Combining them with **one** another, we can solve the problem of the directions of the spontaneous proceeding of $(38 \times 37)/2 = 703$ reactions.

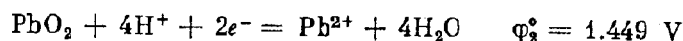
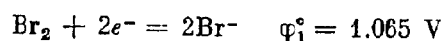
Example. Establish the direction in which the following reaction can proceed:



We shall write the equation of the reaction in the net ionic form.



In Table. 2, we find the standard electrode potentials for the electrochemical systems participating in the reaction:



The oxidising agent is always the electrochemical system with a higher value of the electrode potential. Since here φ_2° is considerably greater than φ_1° , then at virtually any concentrations of the reacting substances the bromide ion will be the reducing agent and will be oxidised by the lead

dioxide: the reaction will proceed spontaneously from left to right.

The farther a system is in the series of standard potentials, i.e. the higher its standard potential, the stronger an oxidising agent is its oxidised form. And, conversely, the earlier a system appears in the series, i.e. the smaller the value of p° , the stronger a reducing agent is its reduced form. Indeed, among the oxidised forms of systems at the bottom of the series, we find such strong oxidising agents as F_2 , H_2O_2 , and MnO_4^- . The strongest reducing agents, on the other hand, are the reduced forms of the systems at the top of the series; the alkali and alkaline-earth metals.

When redox reactions proceed, the concentrations of the reactants diminish, and of the products grow. The result is a change in the magnitudes of the potentials of both half-reactions: the electrode potential of the oxidising agent drops, and that of the reducing agent grows. When the potentials of both processes become equal to each other, the reaction terminates—a state of chemical equilibrium sets in.

Electromotive Series of Metals

If from the entire series of standard electrode potentials we separate only the electrode processes that correspond to the general equation

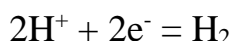


we get an electromotive (or activity) series of metals. In addition to metals, hydrogen is included in this series, which permits us to see what metals are capable of displacing hydrogen from aqueous solutions of acids. The electromotive series for the most important metals is given in Table. 3.

Table. 3. Electromotive Series of Metals

Equation of electrode process	Standard potential φ° at 25 °C, V	Equation of electrode process	Standard potential φ° at 25 °C, V
Li ⁺ + e ⁻ = Li	-3.045	Co ²⁺ + 2e ⁻ = Co	-0.277
Rb ⁺ + e ⁻ = Rb	-2.925	Ni ²⁺ + 2e ⁻ = Ni	-0.250
K ⁺ + e ⁻ = K	-2.924	Sn ²⁺ + 2e ⁻ = Sn	-0.136
Cs ⁺ + e ⁻ = Cs	-2.923	Pb ²⁺ + 2e ⁻ = Pb	-0.126
Ca ²⁺ + 2e ⁻ = Ca	-2.866	Fe ³⁺ + 3e ⁻ = Fe	-0.037
Na ⁺ + e ⁻ = Na	-2.714	2H ⁺ + 2e ⁻ = H ₂	0
Mg ²⁺ + 2e ⁻ = Mg	-2.363	Bi ³⁺ + 3e ⁻ = Bi	0.215
Al ³⁺ + 3e ⁻ = Al	-1.663	Cu ²⁺ + 2e ⁻ = Cu	0.337
Ti ²⁺ + 2e ⁻ = Ti	-1.630	Cu ⁺ + e ⁻ = Cu	0.520
Mn ²⁺ + 2e ⁻ = Mn	-1.179	Hg ₂ ²⁺ + 2e ⁻ = 2Hg	0.788
Cr ²⁺ + 2e ⁻ = Cr	-0.913	Ag ⁺ + e ⁻ = Ag	0.799
Zn ²⁺ + 2e ⁻ = Zn	-0.763	Hg ²⁺ + 2e ⁻ = Hg	0.850
Cr ³⁺ + 3e ⁻ = Cr	-0.744	Pt ²⁺ + 2e ⁻ = Pt	1.188
Fe ²⁺ + 2e ⁻ = Fe	-0.440	Au ³⁺ + 3e ⁻ = Au	1.498
Cd ²⁺ + 2e ⁻ = Cd	-0.403	Au ⁺ + e ⁻ = Au	1.692

The position of a metal in the series characterises its ability to participate in oxidation-reduction reactions in aqueous solutions in standard conditions. Ions of the metals are oxidising agents, and the metals in the form of elementary substances are reducing agents. The farther a metal is in the electromotive series, the stronger an oxidising agent in an aqueous solution are its ions. Conversely, the nearer a metal is to the top of the series the stronger are the reducing properties exhibited by the elementary substance—metal. The potential of the electrode process



in a neutral medium ($\text{pH} = 7$) is $-0.059 \times 7 = -0.41 \text{ V}$. The active metals at the top of the series having a potential that is considerably more negative than -0.41 V displace hydrogen from water. Magnesium displaces hydrogen only from hot water. The metals between magnesium and cadmium do not usually displace hydrogen from water. The surfaces of these metals become covered with oxide films having a protective action.

The metals between magnesium and hydrogen displace hydrogen from solutions of acids. The surfaces of some metals also become covered with protective films that inhibit the reaction. For example, the oxide film on aluminium makes this metal stable not only in water, but also in solutions of some acids. Lead does not dissolve in sulphuric acid when its concentration is below 80% because the salt PbSO_4 formed when lead reacts with sulphuric acid is insoluble and produces a protective film on the surface of the metal. The phenomenon of the deep inhibition of the oxidation of a metal due to the presence of protective oxide or salt films on its surface is called passivation, and the state of the metal is called the passive state.

Metals are capable of displacing one another from solutions of salts. The direction of the reaction is determined by their relative position in the electromotive series. When considering specific cases of such reactions, one must remember that active metals displace hydrogen not only from water, but also from any aqueous solution. Consequently, the mutual displacement of metals from solutions of their salts occurs in practice only with metals below magnesium in the series.

The relative position of some metals in the electromotive series and in the periodic table, at first sight, is not the same. For example, according to its position in the periodic table, the chemical activity of potassium should be greater than that of sodium, and of sodium greater than that of lithium.

In the electromotive series, however, lithium is the most active, while potassium occupies a position between lithium and sodium. Zinc and copper as regards their position in the periodic table should have an approximately identical chemical activity, but in the electromotive series, zinc is considerably ahead of copper. The reason for such discrepancies is as follows.

In comparing metals occupying definite positions in the periodic table, the magnitude of the ionisation energy of free atoms is taken as a measure of their chemical activity—their reducing power. Indeed, as we pass, for example, down the main subgroup of group I of the periodic table, the ionisation energy of atoms diminishes, which is associated with an increase in their radii (i.e., with a greater distance to the outer electrons from the nucleus) and with an increasing screening of the positive charge of the nucleus by intermediate electron layers. Consequently, potassium atoms display a greater chemical activity—have stronger reducing properties—than sodium atoms, and sodium atoms display a greater activity than lithium atoms.

When comparing metals in the electromotive series, the measure of their chemical activity is the work of transforming a metal in the solid state into hydrated ions in an aqueous solution. This work can be represented as the sum of three addends: the energy of atomisation—the transformation of a metal crystal into isolated atoms, the energy of ionisation of the free atoms of the metal, and the energy of hydration of the ions formed. The atomisation energy characterises the strength of the crystal lattice of a given metal. The ionisation energy of atoms—the energy needed to detach valence electrons from them—is directly determined by the position of the metal in the periodic table. The energy evolved in hydration depends on the electronic structure of an ion, its charge and radius. Ions of lithium and

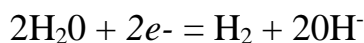
potassium, having an identical charge, but different radii, will set up different electric fields around themselves. The field set up near small lithium ions will be stronger than that near big potassium ions. It is thus clear that lithium ions will become hydrated with the liberation of more energy than potassium ions.

Thus, in the course of the transformation being considered, energy is spent for atomisation and ionisation, and energy is evolved in hydration. When the summary expenditure of energy is smaller, the some electrochemical processes. Such cases will be given special mention below.

In considering cathode processes occurring in the electrolysis of aqueous solutions, we shall limit ourselves to the most important case—cathode reduction leading to separation of the elements in the free state. Here account must be taken of the magnitude of the potential in the process of hydrogen ion reduction. This potential depends on the concentration of the hydrogen ions and for neutral solutions (pH =7) it is $\phi = -0.059 \times 7 = -0.41$ V. Therefore, if the cation of an electrolyte is a metal whose electrode potential is considerably more positive than -0.41 V, the metal will separate from a neutral solution of such an electrolyte at the cathode. In the electromotive series, such metals are near hydrogen (beginning approximately with tin) and after it. Conversely, if a metal having a potential considerably more negative than -0.41 V is the cation of an electrolyte, the metal will not be reduced, but hydrogen will be liberated. Such metals include the ones at the top of the electromotive series, approximately up to titanium. Finally, if the potential of a metal is close to the value -0.41 V (metals in the middle part of the series—Zn, Cr, Fe, Cd, Ni), then depending on the concentration of the solution and the conditions of electrolysis both reduction of the metal and the liberation of

hydrogen are possible. Often the joint deposition of the metal and liberation of hydrogen are observed.

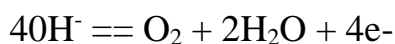
The electrochemical liberation of hydrogen from acid solutions occurs owing to the discharge of the hydrogen ions. For neutral or alkaline media, it is the result of the electrochemical reduction of water:



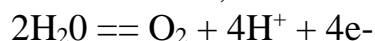
Thus, the nature of the cathode process in the electrolysis of aqueous solutions is determined first, of all by the position of the relevant metal in the electromotive series. In a number of cases, the pH of the solution, the concentration of the metal ions, and other conditions of electrolysis are of great importance.

In considering anode processes, we must remember that the material of the anode may become oxidised in the course of electrolysis. In this connection, electrolysis with an inert anode and electrolysis with an active anode are distinguished. An anode is called inert if its material does not become oxidised in electrolysis. An anode is called active if its material may become oxidised in the course of electrolysis. Most frequently, graphite, carbon, and platinum are used as the materials for inert anodes.

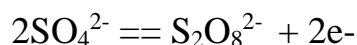
The most important conditions of electrolysis include the current density, the temperature, and the composition of the solution. The current density is defined as the ratio of the current to the working surface area of an electrode. The electrochemical oxidation of water with the liberation of oxygen occurs at an inert anode in the electrolysis of aqueous solutions of alkalies, oxyacids and their salts, and also hydrogen fluoride and fluorides. Depending on the pH of the solution, this process goes on differently and can be expressed by different equations. For an alkaline medium, the equation is



and for an acid or a neutral one, it is

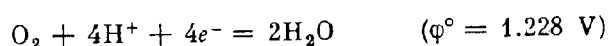


In the above cases, the electrochemical oxidation of water is the most advantageous process from the energy viewpoint. Oxygen-containing anions are either not capable of becoming oxidised, or their oxidation occurs at very high potentials. For instance, the standard potential of oxidation of the ion SO_4^{2-}



is 2.010 V, which considerably exceeds the standard potential of oxidation of water (1.228 V). The standard potential of oxidation of the ion F^- has a still higher value (2.87 V).

In the electrolysis of aqueous solutions of acids not containing oxygen and their salts (except for HF and fluorides), the anions are discharged at the anode. Particularly, in the electrolysis of solutions of HI, HBr, HCl and their salts, the corresponding halogen is liberated at the anode. We must note that the liberation of chlorine in the electrolysis of HCl and its salts contradicts the mutual position of the systems



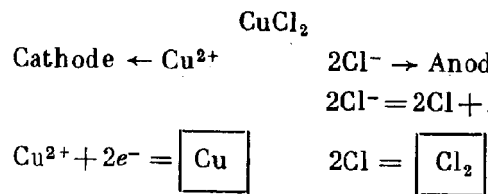
in the series of standard electrode potentials. This anomaly is connected with the considerable overvoltage of the second of these two electrode processes—the material of the anode has an inhibiting action on the process of oxygen liberation.

With an active anode, the number of competing oxidising processes grows to three: the electrochemical oxidation of water with the liberation of oxygen, discharge of the anion (i.e. its oxidation), and electrochemical oxidation of the anode metal (the anode dissolving of the metal). Of these possible processes, the one will occur that is more profitable from the

energy viewpoint. If the metal of the anode is above the other two electrochemical systems in the series of standard potentials, then anode dissolving of the metal will be observed. Otherwise, oxygen will be liberated or the anion will be discharged. Let us consider some typical cases of electrolysis of aqueous solutions.

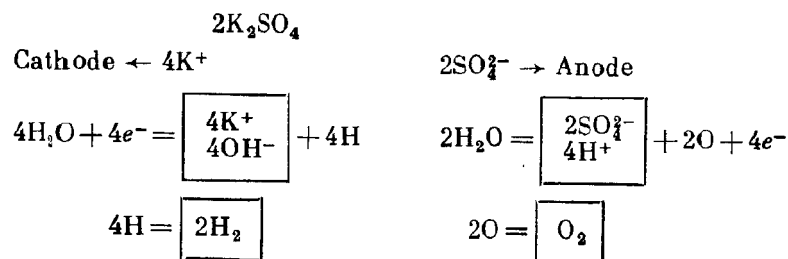
The Electrolysis of a CuCl_2 Solution with an Inert Anode. Copper is below hydrogen in the electromotive series. Consequently, the discharge of the Cu^{2+} ions and the deposition of metallic copper will occur at the cathode. The chloride ions will discharge at the anode.

The scheme of electrolysis* of a solution of copper(II) chloride is as follows:



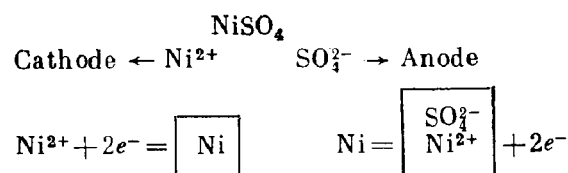
Electrolysis of a K_2SO_4 Solution with an Inert Anode. Since potassium is much higher than hydrogen in the electromotive series, the liberation of hydrogen and the accumulation of OH^- ions will occur at the cathode. The liberation of oxygen and the accumulation of H^+ ions will occur at the anode. At the same time, K^+ ions will arrive at the cathode space, and SO_4^{2-} ions at the anode space. Thus, the solution will remain electrically neutral in all its parts. An alkali will accumulate in the cathode space, however, and an acid in the anode space.

The scheme of electrolysis of a potassium sulphate solution is:



Electrolysis of an NiSO₄ Solution with a Nickel Anode. The standard potential of nickel (—0.250 V) is somewhat higher than —0.41 V. Hence, in the electrolysis of a neutral solution of NiSO₄, discharge of the Ni²⁺ ions and deposition of the metal will mainly occur at the cathode. The opposite process—oxidation of the metal—occurs at the anode because the potential of nickel is much lower than the potential of oxidation of water and, moreover, than the potential of oxidation of the SO₄²⁻ ion. Hence, in the given case, electrolysis consists in dissolving of the anode metal and its deposition at the cathode.

In this and the following schemes, the boxes contain the formulas of the substances that are the products of electrolysis. The scheme of electrolysis of a nickel sulphate solution is



This process is used for the electrical purification of nickel.

Laws of Electrolysis

Electrolysis was first studied quantitatively in the thirties of the 19th century by the outstanding British scientist M. Faraday who established the following laws of electrolysis:

1. The mass of a substance formed in electrolysis is proportional to the amount of electricity that has passed through the solution.

This law follows from the essence of electrolysis. We have already mentioned that an electrochemical process—the reaction of the ions or molecules of the electrolyte with the metal electrons—occurs at the place of contact of the metal with the solution, so that the electrolytic formation of a substance is a result of this process. It is quite obvious that the amount

of a substance obtained at an electrode will always be proportional to the number of electrons passing through the circuit, i.e. to the amount of electricity.

2. In the electrolysis of different chemical compounds, equal amounts of electricity result in the electrochemical transformation of equivalent amounts of substances.

Assume, for example, that an electric current consecutively passes through solutions of hydrochloric acid, silver nitrate, copper (II) chloride, and tin (IV) chloride.

Michael Faraday (1791-1867) was one of the most eminent British physicists and chemists. The major part of his works relate to the field of electricity. He established the laws of electrolysis, and discovered the phenomenon of electromagnetic induction. He was the first to obtain a number of gases (chlorine, ammonia, etc.) in the liquid state. He discovered benzene and isobutylene.

After a certain time, the amounts of the electrolysis products are determined. It is found that in the time during which one gram of hydrogen, i.e. one mole of atoms is liberated from a solution of hydrochloric acid, the masses of the metals indicated below will be deposited from the other solutions:

	Electrolyte		
Mass of metal deposited at cathode, g	AgNO ₃ 107.9	CuCl ₂ 31.8	SnCl ₄ 29.7
Atomic mass of metal	107.87	63.55	118.69

Comparing the deposited masses of the metals with their atomic masses, we find that one mole of silver atoms, one-half mole of copper atoms, and one-fourth mole of tin atoms are deposited. In other words, the amounts of substances formed at the cathode equal their equivalents. The same result is obtained by measuring the amounts of the substances liberated at the

anode. Thus, in each of the first, third and fourth cells, 35.5 g of chlorine are liberated, and in the second cell 8 g of oxygen. It is not difficult to see that here too the substances are liberated in amounts equal to their equivalents.

In considering the second law of electrolysis from the viewpoint of the electron theory, it is not difficult to understand why substances are liberated in chemically equivalent amounts in electrolysis. Let us take as an example the electrolysis of copper (II) chloride. When copper is deposited from the solution, each of its ions receives two electrons from the cathode, and at the same time two chloride ions give up electrons to the anode, transforming into chlorine atoms. Consequently, the number of deposited copper atoms will always be half the number of liberated chlorine atoms, i.e. the ratio of the masses of the copper and chlorine will equal the ratio of their equivalent masses.

Measurements have established that the amount of electricity causing the electrochemical transformation of one equivalent of a substance is 96 485 (approximately 96 500) -coulombs. This amount of electricity is called the *faraday or Faraday's constant and is* designated by the symbol *F*.

The second law of electrolysis gives a direct way of determining the equivalents of various elements. Calculations associated with electrochemical production processes are based on the same law.

The laws of electrolysis relate to the electrolysis of solutions, melts, and solid electrolytes with purely ionic conductivity.

Electrolysis in Industry

Electrolysis finds a very important application in the metallurgical and chemical industries and in electrodeposition.

In the metallurgical industry, electrolysis of molten compounds and aqueous solutions is used to produce metals, and also for electrolytic refining—the purification of metals from harmful impurities and the extraction of valuable components.

The **electrolysis of melts** is used to produce metals having strongly negative electrode potentials, and some of their alloys.

At a high temperature, the electrolyte and the electrolysis products may enter into a reaction with one another, with air, and also with the materials of the electrodes and the electrolyser. As a result, this scheme of electrolysis, which is simple in principle (for instance, the electrolysis of MgCl_2 in the production of magnesium) becomes more intricate.

Usually, not individual molten compounds, but their mixtures are used as the electrolyte. A very important advantage of mixtures is their relatively low melting point allowing electrolysis at a lower temperature.

At present, electrolysis of melts is used to produce aluminium, magnesium, sodium, lithium, beryllium, and calcium. It is virtually not used to produce potassium, barium, rubidium, and cesium owing to the high chemical activity of these metals and their high solubility in the molten salts. In recent years, the electrolysis of molten media has been acquiring some significance for the production of certain refractory metals.

The electrolytic separation of a metal from a solution is called electrical extraction. The ore or concentrate is treated with definite reagents to transfer the metal into a solution. After purification, the solution is subjected to electrolysis. The metal is deposited at the cathode and in the majority of cases is highly pure. This method is used mainly to produce

zinc, copper, and cadmium.

Metals are subjected to electrolytic refining for removing impurities from them and for transferring the components contained in them into products convenient to process. The metal to be purified is cast into plates, and they are placed as anodes in an electrolyser. The passage of a current through the circuit causes the anode metal to dissolve—it passes into solution in the form of cations. The metal cations then become discharged at the cathode and form a compact deposit of pure metal. The impurities in the anode either remain undissolved, settling in the form of anode mud, or pass into the electrolyte, whence they are periodically or continuously withdrawn.

Let us consider as an example the electrolytic refining of copper. The main component of the solution is copper sulphate—the most widespread and cheapest salt of this metal. But a solution of CuSO_4 has a low electrical conductivity. To increase it, sulphuric acid is added to the electrolyte. Small amounts of additions facilitating the production of a compact deposit of the metal are also introduced into the solution.

The metallic impurities contained in unrefined ("blister") copper can be divided into two groups:

(1) Fe, Zn, Ni, Co. These metals have considerably more negative electrode potentials than copper. Therefore, they are dissolved from the anode together with the copper, but are not deposited at the cathode. They accumulate in the electrolyte, and in this connection the latter is periodically purified.

(2) Au, Ag, Pb, Sn. The noble metals (Au, Ag) do not undergo anode dissolving, and in the course of the process settle near the anode, forming together with other impurities an anode mud that is periodically extracted. Tin and lead, on the other hand, dissolve together with the copper, but form

poorly soluble compounds in the electrolyte that precipitate and are also removed.

Copper, nickel, lead, tin, silver, and gold are subjected to electrolytic refining.

Electrodeposition includes electroplating and galvanoplastics. Electroplating processes involve the application of other metals on the surface of metal articles by electrolysis to protect the articles from corrosion, to impart hardness to their surfaces, and also for decorative purposes. Among the numerous electroplating processes used in engineering, the most important are chromium plating, zinc plating (galvanisation), and nickel plating.

The essence of electroplating is as follows. The well cleaned and degreased component to be protected is immersed in a solution containing a salt of the metal which it is to be coated with, and is connected as the cathode to a direct-current circuit. When a current flows through the circuit, a coat of the protecting metal is deposited onto the component. The best protection is provided by finely crystalline dense deposits. Such deposits also have better mechanical properties.

Galvanoplastics (or galvanoplasty) is the name given to processes of obtaining accurate metal copies of relief (embossed) objects by the electrodeposition of metal. Galvanoplastics is used to produce moulds and dies for pressing various articles (phonograph records, buttons, etc.), matrices for stamping leather and paper, printed radio engineering circuits, and stereotype plates. Galvanoplastics was discovered by the Russian academician B. Jacobi (1801-1874) in the thirties of the 19th century.

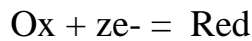
There are also other kinds of electrochemical treatment of metal surfaces such as the electrical polishing of steel and the oxidation of aluminium and magnesium. The latter consists in anode treatment of the metal in the

course of which the structure of the oxide film on its surface is changed in a definite way. The result is improvement of the corrosion resistance of the metal. In addition, the metal acquires an attractive appearance.

In the chemical industry, electrolysis is used to prepare various products. Among them are fluorine, chlorine, sodium hydroxide, many oxidising agents, in particular highly pure hydrogen, and hydrogen peroxide.

Electrochemical Polarisation. Overvoltage

When an electrode is at a potential equal to the equilibrium one, electrochemical equilibrium sets in on it:



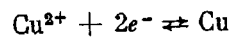
When the potential of the electrode is displaced in the positive or negative direction, oxidation or reduction processes begin to proceed on it. The deviation of the potential of an electrode from its equilibrium value is called *electrochemical polarisation* or simply *polarisation*.

An electrode can be polarised by connecting it to a direct-current circuit. For this purpose, an electrolytic cell must be formed from an electrolyte and two electrodes—the one being studied and an auxiliary one. By connecting it to a direct-current circuit, we can make the electrode being studied the cathode or (with reverse connection of the cell) the anode. This method of polarisation is called polarisation from an external source of electrical energy.

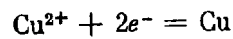
Let us consider a simple example of polarisation. Assume that a copper electrode is in a 0.1 *m* solution of CuSO₄ containing no impurities and no dissolved oxygen. As long as the circuit is not closed, the electrode potential at 25 °C will have an equilibrium value equal to

$$\varphi = 0.337 + 0.030 \log 10^{-1} \approx 0.31 \text{ V}$$

and electrochemical equilibrium will set in at the metal-solution interface:



Let us connect the electrode to the negative pole of the current source—we shall make it the cathode. The surplus of electrons that now appears at the electrode will shift the potential of the electrode in the negative direction and simultaneously violate equilibrium. The electrons will attract copper cations from the solution—a reduction process will occur:



If we connect the electrode to the positive pole of the current source instead of to the negative one, i.e. make it the anode, then owing to the withdrawal of part of the electrons, the potential of the electrode will shift in the positive direction and equilibrium will also be violated.

6. Nomenclature of Coordination Compounds

A **complex** is a substance in which a metal atom or ion is associated with a group of neutral molecules or anions called **ligands**. **Coordination compounds** are neutral substances (i.e. uncharged) in which **at least** one ion is present as a complex. You will learn more about coordination compounds in the lab lectures of experiment 4 in this course. The coordination compounds are named in the following way. (At the end of this tutorial we have some examples to show you how coordination compounds are named.)

A. To name a coordination compound, no matter whether the complex ion is the cation or the anion, always name the cation before the anion. (This is just like naming an ionic compound.)

B. In naming the complex ion:

1. **Name the ligands first, in alphabetical order, then the metal atom or ion. Note:** The **metal atom or ion** is written **before** the **ligands** in the **chemical formula**.

2. The names of some common ligands are listed in Table 1.

□□ *For anionic ligands end in "-o"; for anions that end in "-ide" (e.g. chloride), "-ate" (e.g. sulfate, nitrate), and "-ite" (e.g. nitrite), change the endings as follows: -ide → -o; -ate → -ato; -ite → -ito*

□□ *For neutral ligands, the common name of the molecule is used e.g. $H_2NCH_2CH_2NH_2$ (ethylenediamine). Important*

exceptions: water is called ‘aqua’, ammonia is called ‘ammine’, carbon monoxide is called ‘carbonyl’, and the N₂ and O₂ are called ‘dinitrogen’ and ‘dioxygen’.

Table. 4. Names of Some Common Ligands.

Anionic Ligands	Names	Neutral Ligands	Names
Br ⁻	bromo	NH ₃	ammine
F ⁻	fluoro	H ₂ O	aqua
O ²⁻	oxo	NO	Nitrosyl
OH ⁻	Hydroxo	CO	Carbonyl
CN ⁻	cyano	O ₂	dioxygen
C ₂ O ₄ ²⁻	oxalato	N ₂	dinitrogen
CO ₃ ²⁻	carbonato	C ₅ H ₅ N	pyridine
CH ₃ COO ⁻	acetato	H ₂ NCH ₂ CH ₂ NH ₂	ethylenediamine

3. Greek prefixes are used to designate the number of each type of ligand in the complex ion, e.g. di-, tri- and tetra-. If the ligand already contains a Greek prefix (e.g. ethylenediamine) or if it is polydentate ligands (ie. can attach at more than one binding site) the prefixes bis-, tris-, tetrakis-, pentakis-, are used instead. The numerical prefixes are listed in Table. 5.

Table. 5. Numerical Prefixes.

Number	Prefix	Number	Prefix	Number	Prefix
1	mono	5	penta (pentakis)	9	nona (ennea)
2	di (bis)	6	hexa (hexakis)	10	deca
3	tri (tris)	7	hepta	11	undeca
4	tetra (tetrakis)	8	octa	12	dodeca

4. After naming the ligands, name the central metal. If the complex ion is a cation, the metal is named same as the element. For example, Co in a complex cation is call cobalt and Pt is called platinum. If the complex ion is an anion, the name of the metal ends with the suffix –ate. For example, Co in a complex anion is called cobaltate and Pt is called platinate. For some metals, the Latin names are used in the complex anions e.g. Fe is called ferrate (not ironate).

5. Following the name of the metal, the oxidation state of the metal in the complex is given as a Roman numeral in parentheses.

C. To name a neutral complex molecule, follow the rules of naming a complex cation. **Remember:** Name the (possibly complex) cation **BEFORE** the (possibly complex) anion.

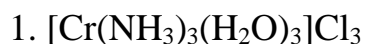
Table. 6: Name of Metals in Cationic and Anionic Complexes

Transition Metal	Name if in Cationic Complex	Name if in Anionic Complex
Sc	scandium	Scandate
Ti	titanium	titanate
V	vanadium	vanadate
Cr	chromium	chromate
Mn	manganese	manganate
Fe	iron	ferrate
Co	cobalt	cobaltate
Ni	nickel	nickelate
Cu	copper	cuprate
Zn	Zinc	zincate

For historic reasons, some coordination compounds are called by their common names. For example, $\text{Fe}(\text{CN})_6^{3-}$ and $\text{Fe}(\text{CN})_6^{4-}$ are named ferricyanide and ferrocyanide respectively, and $\text{Fe}(\text{CO})_5$ is called iron carbonyl.

Examples

Give the systematic names for the following coordination compounds:



Answer: triamminetriaquachromium(III) chloride

Solution: The complex ion is inside the parentheses, which is a cation.

The ammine ligands are named before the aqua ligands according to alphabetical order.

Since there are three chlorides binding with the complex ion, the charge on the complex ion must be +3 (since the compound is electrically neutral).

From the charge on the complex ion and the charge on the ligands, we can calculate the oxidation number of the metal. In this example, all the ligands are neutral molecules. Therefore, the oxidation number of chromium must be same as the charge of the complex ion, +3.



Answer: pentaamminechloroplatinum(IV) bromide

Solution: The complex ion is a cation, the counter anion is the 3 bromides.

The charge of the complex ion must be +3 since it bonds with 3 bromides.

The NH_3 are neutral molecules while the chloride carries \square 1 charge. Therefore, the oxidation number of platinum must be +4.



Answer: dichlorobis(ethylenediamine)platinum(IV) chloride

Solution: ethylenediamine is a bidentate ligand, the bis- prefix is used instead.



Answer: tris(ethylenediamine)cobalt(III) sulfate

Solution: The sulfate is the counter anion in this molecule. Since it takes 3 sulfates to bond with two complex cations, the charge on each complex cation must be +3.

Since ethylenediamine is a neutral molecule, the oxidation number of cobalt in the complex ion must be +3.

Again, remember that you never have to indicate the number of cations and anions in the name of an ionic compound.



Answer: potassium hexacyanoferrate(II)

Solution: potassium is the cation and the complex ion is the anion.

Since there are 4 K^+ binding with a complex ion, the charge on the complex ion must be - 4.

Since each ligand carries -1 charge, the oxidation number of Fe must be +2.

The common name of this compound is potassium ferrocyanide.

6. $\text{Na}_2[\text{NiCl}_4]$

Answer: sodium tetrachloronickelate(II)

Solution: The complex ion is the anion so we have to add the suffix –ate in the name of the metal.

7. $\text{Pt}(\text{NH}_3)_2\text{Cl}_4$

Answer: diamminetetrachloroplatinum(IV)

Solution: This is a neutral molecule because the charge on Pt^{+4} equals the negative charges on the four chloro ligands.

If the compound is $[\text{Pt}(\text{NH}_3)_2\text{Cl}_2]\text{Cl}_2$, even though the number of ions and atoms in the molecule are identical to the example, it should be named: diamminedichloroplatinum(II) chloride, a big difference.

8. $\text{Fe}(\text{CO})_5$

Answer: pentacarbonyliron(0)

Solution: Since it is a neutral complex, it is named in the same way as a complex cation. The common name of this compound, iron carbonyl, is used more often.

9. $(\text{NH}_4)_2[\text{Ni}(\text{C}_2\text{O}_4)_2(\text{H}_2\text{O})_2]$

Answer: ammonium diaquabis(oxalato)nickelate(II)

Solution: The oxalate ion is a bidentate ligand.

10. $[\text{Ag}(\text{NH}_3)_2][\text{Ag}(\text{CN})_2]$

Answer: diamminesilver(I) dicyanoargentate(I)

You can have a compound where both the cation and the anion are complex ions. Notice how the name of the metal differs even though they are the same metal ions.

THE MAIN CLASSES OF INORGANIC COMPOUNDS *LABORATORY TRAINING*

Experiment 1. Interaction between oxides and water

To add a few grains of solid CaO, CuO, ZnO (total volume of each chemicals - near one match-head) into three tubes. To add 10-20 drops of distilled water into each tube. To mix a tube content. What oxides do directly react with water? To write a corresponding reaction.

REACTION	
OBSERVATION	
REACTION	
OBSERVATION	
REACTION	
OBSERVATION	

Experiment 2. Preparation of insoluble hydroxides

To add separately 1-4 drops of each soluble Iron (III) salts (FeCl_3 or $\text{Fe}_2(\text{SO}_4)_3$), Copper (CuSO_4 or CuCl_2 or $\text{Cu}(\text{NO}_3)_2$), Manganese (II) (MnSO_4 or MnCl_2), and Nickel (NiSO_4), Cobalt (II) (CoSO_4 or CoCl_2) into tubes #1, 2, 3, 4, 5. To add 1-2 drops of any alkali into each tube. What do you observe? To write corresponding reactions. To point out color of precipitate in each tube. To explain why color of Manganese (II) Hydroxide is changed.

REACTION	
OBSERVATION	
REACTION	

OBSERVATION	
REACTION	
OBSERVATION	
REACTION	
OBSERVATION	
REACTION	
OBSERVATION	

Experiment 3. Preparation of amphoteric hydroxides

To add 1 drop of any alkali (NaOH or KOH) into three tubes. To add by drops the solution of Aluminium salt (for example, $\text{Al}_2(\text{SO}_4)_3$) into tube # 1 up to the precipitate appearance; into tube # 2 - Zinc salt (for example, ZnSO_4) analogously to previous case; into tube # 3 - Chromium (III) salt (for example, $\text{Cr}_2(\text{SO}_4)_3$). To divide each prepared precipitate into two tubes. To add solution of any acid (for example, HNO_3) to the first part of precipitate, alkali solution up to the whole dissolving of precipitate - to the second one. What do you observe? To write corresponding reactions.

REACTION	
OBSERVATION	
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OBSERVATION	

Experiment 4. Interaction of salts solution (metathesis)

To add 2-3 drops of BaCl₂ solution into tube # 1 and to drop the same quantity of Sodium Sulfate; 2-3 drops of Lead (II) Nitrate - into tube # 2 and to add 1-2 drops of Potassium Iodide; 2-3 drops of Cobalt (II)

Sulfate into tube # 3 and to add 1-2 drops of Sodium HydroPhosphate. What do you observe? To write corresponding reactions.

REACTION	
OBSERVATION	
REACTION	
OBSERVATION	
REACTION	
OBSERVATION	

TEST QUESTIONS

1. Point the correspondence of inorganic substances and their chemical nature:

A.	N_2O_5	1.	Basic oxide
B.	SnO	2.	Acidic oxide
C.	HCl	3.	Amphoteric oxide
D.	Ca(OH)_2	4.	Base
E.	K_2O	5.	Acid

A. _____, B. _____, C. _____, D. _____, E. _____.

2. Point the correspondence of oxides and their chemical nature::

A.	$\text{ZnO}, \text{Cr}_2\text{O}_3$	1.	Basic oxide
B.	SO_2, CO_2	2.	Acidic oxide
C.	CaO, BaO	3.	Amphoteric oxide
D.	$\text{CrO}_3, \text{Mn}_2\text{O}_7$		
E.	$\text{K}_2\text{O}, \text{MgO}$		

A. _____, B. _____, C. _____, D. _____, E. _____.

3. What base is formed by the addition of water to potassium oxide (K_2O)?

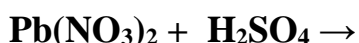
A.	K_2OH	C.	K(OH)_2
B.	KOH	D.	HKO_2

4. Calculate sum of coefficients in the molecular equation:



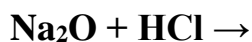
A.	4	C.	6
B.	5	D.	8

5. Calculate sum of coefficients in the molecular equation:



A.	4	C.	6
B.	5	D.	8

6. Calculate sum of coefficients in the molecular equation:



A.	4	C.	6
B.	5	D.	7

7. Calculate sum of coefficients in the molecular equation:



A.	2	C.	5
B.	4	D.	6

8. Choose metals.

A.	K	D.	H
B.	C	E.	Na
C.	O	F.	As

9. Which reaction is with correct products $\text{Mg(OH)NO}_3 + \text{HNO}_3 =$

A.	$\text{Mg(OH)NO}_3 + \text{HNO}_3 = \text{Mg(NO}_3)_2 + \text{H}_2\text{O}$
B.	$\text{Mg(OH)NO}_3 + \text{HNO}_3 = \text{Mg(NO}_2)_2 + \text{H}_2\text{O}$
C.	$\text{Mg(OH)NO}_3 + \text{HNO}_3 = \text{MgOH(NO}_2)_2 + \text{H}_2\text{O}$
D.	$\text{Mg(OH)NO}_3 + \text{HNO}_3 = \text{MgOH(NO}_2)_2$
E.	$\text{Mg(OH)NO}_3 + \text{HNO}_3 = \text{Mg(NO}_2)(\text{OH}) + \text{H}_2\text{O}$

10. Choose amphoteric compounds:

A.	$\text{Zn(OH)}_2, \text{Al(OH)}_3$	D.	$\text{Fe(OH)}_2, \text{Mg(OH)}_2$
B.	$\text{Ca(OH)}_2, \text{Be(OH)}_2$	E.	$\text{Zn(OH)}_2, \text{Ba(OH)}_2$
C.	$\text{Cr(OH)}_3, \text{KOH}$		

11. Determine the Chromium valence in K_2CrO_4 :

A.	9	D.	6
B.	8	E.	5
C.	7		

12. Give the formula for the base formed by the addition of water to barium oxide (BaO):

A.	BaOH	C.	HBaO_3
B.	Ba(OH)_2	D.	Ba_2O_3

13. How many possible reactions are between Ba(OH)_2 and H_2SO_4 (taking into account the possibility of neutral, acidic and basic salts forming).

A.	1	C.	3
B.	2	D.	4

14. How many possible reactions are between Al(OH)_3 and HCl (taking into account the possibility of neutral, acidic and basic salts forming).

A.	1	C.	3
B.	2	D.	4

15. What is it necessary to add to Na_3PO_4 , so that Na_2HPO_4 can be formed:

A.	NaOH	C.	H_2SO_4
B.	NaCl	D.	H_3PO_4

16. Which formula contains error?

A.	CaSO_4	C.	HSO_4
B.	H_2SO_4	D.	BaCl_2

17. Which of the following is a basic oxide?

A.	CO_2	C.	N_2O_3
B.	Na_2O	D.	CrO_3

18. Which one of the following is the acid anhydride for the acid, HClO₄?

A.	ClO	D.	ClO ₄
B.	ClO ₂	E.	Cl ₂ O ₇
C.	ClO ₃		

19. Which of the following is a soluble base?

A.	Aluminium oxide	C.	Sodium hydroxide
B.	Copper(II) oxide	D.	Iron(II) hydroxide

20. Put in sentence a missing word:

An oxide is a chemical compound that contains at least one atom and one other element in its chemical formula.

21. Al₂O₃ is an amphoteric oxide. When reacted with HCl, it acts as a base to form the salt AlCl₃. When reacted with NaOH, it acts as an acid to form NaAlO₂.

A.	true
B.	false

22. Salts are composed of related numbers of cations (positively charged ions) and anions (negative ions).

A.	true
B.	false

23. Put in sentence missing words:

.....is used in the manufacture of fertilizers.

A.	Sodium nitrate	C.	Copper nitrate
B.	Ammonium nitrate	D.	Lead nitrate

24. Choose the acid salt from the following:

A.	NaNO ₃	C.	NaHSO ₄
B.	Na ₂ SO ₄	D.	Na ₂ CO ₃

25. Identify the products in the following reaction:
Na₂CO₃ + HCl =

A.	sodium chloride and water	C.	sodium carbonate, carbon dioxide and water
B.	water and carbon dioxide	D.	sodium chloride, carbon dioxide and water

26. Identify the salt that has formed in the following reaction:
NH₄OH + H₂SO₄ =

A.	NH ₄ S	C.	NH ₄ SO ₄
B.	(NH ₄) ₂ SO ₃	D.	(NH ₄) ₂ SO ₄

UNITS OF CONCENTRATION

TEST QUESTIONS

1. How many grams of NaOH are in 500.0 mL of a 0.175 N NaOH solution?

A.	$2.19 \cdot 10^{-3}$ g	C.	14.0 g
B.	3.50 g	D.	114 g

2. How many grams of H₃PO₄ are in 175 mL of a 3.5 M solution of H₃PO₄?

A.	0.61 g	C.	20 g
B.	4.9 g	D.	60 g

3. What volume of a concentrated solution of potassium hydroxide (6 N) must be diluted to 200 ml to make a 0.880 N solution of potassium hydroxide?

A.	26.4 mL	C.	50.0 mL
B.	29.3 mL	D.	176 mL

4. What is the molarity of a solution prepared by diluting 43.72 mL of 5.005 M aqueous K₂Cr₂O₇ to 500 ml?

A.	0.0044 M	C.	0.438 M
B.	0.0879 M	D.	0.870 M

5. A 28% solution of phosphoric acid. This means that:

A.	1 mL of this solution contains 28 g of phosphoric acid
B.	1 L of this solution has a mass of 28 g
C.	100 g of this solution contains 28 g of phosphoric acid
D.	1 L of this solution contains 28 ml of phosphoric acid

6. What is the equivalent concentration of a solution prepared by dissolving 7,58 grams of potassium nitrate in enough water to prepare 250 mL of solution?

A.	0,0937 g-eq /L	D.	3,065 g-eq /L
B.	0,300 g-eq /L	E.	3,34 g-eq /L
C.	1,895 g-eq /L		

7. 66,7 mL of 18 molar sulfuric acid solution was dissolved in enough water to make 500 mL of solution. The molarity of diluted mixture is:

A.	2,40 molar	D.	9,00 molar
B.	0,135 molar	E.	0,00741 molar
C.	36,0 molar		

8. Choose the correct name of the following concentration unit:

$$\frac{\text{mass of solute}}{\text{mass of solution}} \cdot 100$$

A.	Molarity	D.	Molality
----	----------	----	----------

B.	Percent by mass	E.	Titre
C.	Normality		

9. Define the following concentration units:

A. Molarity -

B. Normality -

C. Titre -

D. Percent by mass -

10. What is the volume of a 0.2 M AgNO₃ solution containing 8.5 grams of AgNO₃?

A.	25 mL	D.	500 mL
B.	100 mL	E.	250 mL
C.	1000 mL		

11. What is the percent by mass of solution when 65.3 g of the compound KNO₃ is placed in 163 g of water?

A.	0.286 %	D.	6.53%
B.	65.3%	E.	15.6%
C.	28.6%		

12 A solution that contains all the solute it can hold at a given temperature is

A.	Unsaturated	D.	Insoluble
B.	Soluble	E.	Superunsaturated
C.	Saturated		

13. How many grams of Na₂SO₄ are in 1000 mL of a 0.0053 T solution?

A.	53 g	C.	5 g
B.	3 g	D.	5.3 g

14. What is the titer of a 0.1 M $\text{Na}_2\text{B}_4\text{O}_7$ solution?

A.	0.0381 g/mL	C.	0.3814 g/mL
B.	0.0191 g/mL	D.	1.9071 g/mL

15. How many mL of 2.50 % salt solution would contain 1.80 g of salt?

A.	105 mL	C.	45 mL
B.	104 mL	D.	72 mL

16. Which statement about solutions is correct?

A.	When vitamin D dissolves in fat, vitamin D is the solvent and fat is the solute.
B.	In a solution of NaCl in water, NaCl is the solute and water is the solvent.
C.	An aqueous solution consists of water dissolved in a solute.
D.	The concentration of a solution is the amount of solvent dissolved in 1 mL of solution.

17. Hydrogen peroxide that is 3% by mass concentration has 3 grams of pure hydrogen peroxide mixed with:

A.	97 grams of solution	C.	97 grams of water
B.	100 grams of solution	D.	100 grams of water

18. In a solution of ammonia gas in water:

A.	the water is the solute and ammonia is the solvent	D.	both the ammonia and the water are solutes
B.	the ammonia is the solute and water is the solvent	E.	both the ammonia and the water are solvents

THEORY OF ELECTROLYTIC DISSOCIATION

LABORATORY TRAINING

Experiment 1. Slightly soluble substances production

- A) Put separately 5-6 drops of Na_2SO_4 , $(\text{NH}_4)_2\text{CO}_3$, Na_3PO_4 solutions into three test-tubes and add 3-4 drops of BaCl_2 solution into each of them. What does happen? Draw molecular and net ionic equations.

REACTION	
OBSERVATION	
REACTION	
OBSERVATION	
REACTION	
OBSERVATION	

- B) Put 5-6 drops of CuSO_4 and $\text{Cu}(\text{NO}_3)_2$ solutions into two test-tubes. Add 4-5 drops of NaOH solution into one test-tube and 4-5 drops of $\text{Ba}(\text{OH})_2$ into another. What does happen? Draw molecular and net ionic equations.

REACTION	
OBSERVATION	
REACTION	

OBSERVATION	

Experiment 2. Weak electrolytes production

Put 5-6 drops of CH_3COONa and $\text{Pb}(\text{CH}_3\text{COO})_2$ solutions into two test-tubes, add 5-6 drops of 2 N HCl or 2 N H_2SO_4 into each of them. Heat the mixture and identify by odor the acetic acid vapor isolation. Draw molecular and net ionic equations.

REACTION	
OBSERVATION	
REACTION	
OBSERVATION	

Experiment 3. Gases production

A) Put 5-6 drops of NH_4Cl into the test-tube; add the same volume of NaOH . Heat the mixture on water bath. Identify by odor the kind of gas. Draw molecular and net ionic equations.

REACTION	
OBSERVATION	

B) Put 0,01 - 0,02 g of CaCO_3 or BaCO_3 solid salt into the test-tube and add 4-5 drops of 2 N HCl solution. What does happen? Draw molecular and net ionic equations.

REACTION	
OBSERVATION	

Experiment 4. Properties of amphoteric hydroxides and their production

Produce Zn(OH)_2 precipitation in two test-tubes by adding drops of ZnSO_4 solution to 1-2 drops of NaOH solution to forming of sediment. Add 5-6 drops of 2 N HCl into one test-tube and 5-6 drops of 2 N NaOH into another. What does happen with Zn(OH)_2 precipitation? Draw molecular and net ionic reactions of Zn(OH)_2 production and its dilution in acid and alkali.

REACTION	
OBSERVATION	
REACTION	
OBSERVATION	
REACTION	
OBSERVATION	

TEST QUESTIONS

1. Each of the following compounds classify as a strong or a weak electrolyte:

A.	KOH	1.	strong electrolyte
B.	CH ₃ COOH	2.	weak electrolyte
C.	NaCl		
D.	HNO ₂		
E.	NaOH		
F.	H ₂ SO ₄		
G.	Na ₃ PO ₄		
H.	H ₂ S		

A._____, B._____, C._____, D._____, E._____, F._____, G._____,
H._____.

2. How many ions will form after dissociation of Chromium(III) Sulphate molecular?

A.	2	D.	5
B.	3	E.	6
C.	4		

3. Chose substances that dissociate to ions in one stage:

A.	HNO ₃	C.	Ca(NO ₃) ₂
B.	H ₂ CO ₃	D.	NaClO ₃

4. All of the following compounds would produce an electrolyte solution when dissolved in water EXCEPT:

A.	Glucose (C ₆ H ₁₂ O ₆)	C.	Potassium iodide (KI)
B.	Ammonium chloride (NH ₄ Cl)	D.	Magnesium sulfate (MgSO ₄)

5. Which one of the following compounds is a nonelectrolyte when dissolved in water?

A.	CuSO ₄	C.	CaCl ₂
B.	C ₆ H ₁₂ O ₆	D.	NaCl

6. Which of the following net-ionic equation represents the reaction which takes place when HNO₃ is added to Mg(OH)₂?

A.	$\text{HNO}_3 + \text{Mg}(\text{OH})_2 \rightarrow \text{Mg}(\text{NO}_3)_2 + \text{H}_2\text{O}$
B.	$\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$
C.	$\text{HNO}_3 + \text{OH}^- \rightarrow \text{NO}_3^- + \text{H}_2\text{O}$
D.	$\text{H}^+ + \text{Mg}(\text{OH})_2 \rightarrow \text{Mg}^{2+} + \text{H}_2\text{O}$

7. Calculate sum of coefficients in net-ionic equation that represents the reaction: BaCl₂+H₂SO₄=

A.	3	C.	5
----	---	----	---

B.	4	D.	6
----	---	----	---

8. In the reaction: $K_2SO_4 + Ba(NO_3)_2 \rightarrow BaSO_4 + 2KNO_3$ which ions are the spectator ions?

A.	Ba^{2+} and SO_4^{2-}	D.	K^+ and SO_4^{2-}
B.	Ba^{2+} and K^+	E.	K^+ and NO_3^-
C.	Ba^{2+} and NO_3^-		

9. Which one of the following species is a weak electrolyte?

A.	HCl	C.	NH_4OH
B.	NaOH	D.	H_2SO_4

10. Which substances are strong electrolytes?

A.	$Zn(OH)_2$	D.	HF
B.	HNO_3	E.	CH_3COOH
C.	HClO	F.	$CaCl_2$

11. Which one of the following compounds produces 3 ions per formula unit by dissociation when dissolved in water?

A.	$NaBrO_3$	D.	KClO
B.	$Al(NO_3)_3$	E.	NH_4Cl

12. Which one of the following compounds produces 3 ions per formula unit by dissociation when dissolved in water?

A.	$NaClO_2$	D.	NaClO
B.	$LiClO_4$	E.	$(NH_4)_2SO_4$

13. Put in following sentence the missing word:

When electrolytes are dissolved in water (or other polar solvents) they ionize into positive ions - cations and negative ions -

14. Nonelectrolytes are nonpolar (covalent) compounds, that do not dissociate into ions.

A.	true
B.	false

15. Which statement below is correctly worded and states a fact?

A.	Acids are strong electrolytes and ionize completely when dissolved in water.
B.	Bases are weak electrolytes and ionize completely when dissolved in water.
C.	Bases are strong electrolytes and ionize completely when dissolved in water.
D.	Soluble salts are strong electrolytes and dissociate completely when dissolved in water.
E.	Soluble salts are weak electrolytes and dissociate partially when dissolved in water.

16. The equation for the reaction: $2AgNO_3 + K_2CrO_4 \rightarrow Ag_2CrO_4 +$

2KNO_3 , can be written as an ionic equation. In this ionic equation the spectator ions are:

A.	Ag^+ and K^+	D.	K^+ and NO_3^-
B.	Ag^+ and CrO_4^{2-}	E.	CrO_4^{2-} and NO_3^-
C.	K^+ and CrO_4^{2-}		

17. Which set of compounds below is a set in which all members are considered soluble in water:

A.	BaCO_3 , NaBrO_3 , $\text{Ca}(\text{OH})_2$, PbCl_2	D.	NaCl , AgBr , Na_2CO_3 , $\text{Hg}(\text{NO}_3)_2$
B.	NaCl , BaCl_2 , NH_4NO_3 , LiClO_4	E.	PbCl_2 , $\text{Pb}(\text{NO}_3)_2$, AgClO_4 , HgCl_2
C.	NiCO_3 , PbSO_4 , AgCl , $\text{Mg}(\text{OH})_2$		

18. Which of the following choices represents the net-ionic equation which takes place in solution when CH_3COOH is added to $\text{Ba}(\text{OH})_2$?

A.	$\text{CH}_3\text{COOH} + \text{Ba}(\text{OH})_2 \rightarrow \text{Ba}(\text{CH}_3\text{COO})_2 + \text{H}_2\text{O}$
B.	$\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$
C.	$\text{CH}_3\text{COOH} + \text{OH}^- \rightarrow \text{CH}_3\text{COO}^- + \text{H}_2\text{O}$
D.	$\text{H}^+ + \text{Ba}(\text{OH})_2 \rightarrow \text{Ba}^{2+} + \text{H}_2\text{O}$
E.	$\text{CH}_3\text{COOH} + \text{Ba}^{2+} \rightarrow \text{Ba}(\text{CH}_3\text{COO})_2 + \text{H}^+$

19. Write and balance molecular, ionic, and net-ionic equations for the following reaction:

molecular	$(\text{NH}_4)_2\text{C}_2\text{O}_4 + \text{Al}(\text{ClO}_4)_3 \rightarrow$
ionic	
net-ionic	

20. Which ions are present in a solution of sodium carbonate, Na_2CO_3 ?

A.	Na^+ , CO_3^{2-}	D.	Na_2^+ , CO_3^{2-}
B.	Na^+ , C^{4+} , O_3^{2-}	E.	Ca^{2+} , C^+ , O^{2-}

IONIC PRODUCT OF WATER. HYDROLYSIS OF SALTS

LABORATORY TRAINING

Experiment 1. Identification of reaction medium in solutions of salts

Draw the reactions of hydrolysis for salt solutions of Na_2CO_3 , $(\text{NH}_4)_2\text{SO}_4$, ZnSO_4 , $\text{Cu}(\text{NO}_3)_2$, $\text{Al}_2(\text{SO}_4)_3$, Na_3PO_4 and theoretically provide for pH of these salt solutions. Then identify reaction medium by putting the drop of solution on universal litmus paper. Draw molecular and ionic equations of hydrolysis in form of the next table:

REACTION	
OBSERVATION	
REACTION	
OBSERVATION	
REACTION	
OBSERVATION	
REACTION	
OBSERVATION	
REACTION	

OBSERVATION	

Experiment 2. Influence of temperature to hydrolysis

A) Fill each of two test tubes with 5-6 drops of 2N CH₃COONa solution and one drop of phenolphthalein solution. Leave one test-tube for comparison and heat another on water bath and note color change. Draw molecular and ionic hydrolysis equations.

REACTION	
OBSERVATION	

B) Fill test-tube with 5-6 drops of FeCl₃ and 4-5 drops of CH₃COONa. Note solution color change as the result Iron (III) Acetate production. Add 10 - 20 mL of distilled water and heat the mixture on water bath up to brown precipitation formation. Draw molecular and ionic equations of hydrolysis (1st and 2^d steps).

REACTION	
OBSERVATION	
REACTION	
OBSERVATION	

Experiment 3. Hydrolysis strengthening in case of solution dilution

Put 4-5 drops of Bithmus (III) Chloride BiCl_3 and by drops (15-20 drops) distilled water up to basic salt precipitation formation. Save the solution with precipitation for the next experiment. Draw hydrolysis equation taking into account that $\text{Bi(OH)}_2\text{Cl}$ is produced as the result of hydrolysis.

REACTION	
OBSERVATION	

Experiment 4. Reversibility of hydrolysis process

Add by drops hydrochloric acid solution up to precipitation produced in previous experiment. Then add distilled water up to basic salt precipitation production. Explain the phenomenon, which occurs. Draw molecular and ionic equations.

REACTION	
OBSERVATION	
REACTION	
OBSERVATION	

TEST QUESTIONS

1. What group of salts are hydrolyzed?

A. NH_4Br , KCl , MgSO_4	C. Fe_2S_3 , AlCl_3 , ZnSO_4
B. $\text{Cr}(\text{NO}_3)_3$, BaCl_2 , CuCl_2	D. NaNO_3 , NH_4NO_3 , NiSO_4

2. What salt have acidic medium in water solution?

A. ZnSO_4	C. NaNO_3
B. LiNO_3	D. Na_2SO_4

3. What is the action of Na_2CO_3 solution on litmus?

A. litmus turns red	C. litmus turns yellow
B. litmus turns blue	D. litmus turns green

4. Predict the pH of the following eques salt solutions?

A. LiCl	1.	pH \approx 7
B. CaSO_3	2.	pH $<$ 7
C. Na_2CO_3	3.	pH $>$ 7
D. $\text{Fe}(\text{NO}_3)_3$		
E. KClO_4		

A.____, B____, C____, D.____, E.____.

5. Write molecular, complete and net-ionic equations for the 1 st step hydrolysis of Na_2SiO_3 .

6. Write molecular, complete and net-ionic equations for the 1 st step hydrolysis of ZnSO_4 .

7. Write molecular, complete and net-ionic equations for the 1 st step hydrolysis of NH_4NO_2 .

8. Calculate sum of coefficients in net-ionic equation for the 1 st step hydrolysis of $\text{Ba}(\text{NO}_2)_2$

A. 4	C. 6
B. 5	D. 2

9. Which of the following net-ionic equations is represent the 1 st step hydrolysis of Na_2CO_3 :

A. $\text{CO}_3^{2-} + 2\text{HOH} = 2\text{OH}^- + \text{H}_2\text{CO}_3$	
--	--

B.	$\text{HOH} = \text{OH}^- + \text{H}^+$
C.	$\text{CO}_3^{2-} + \text{HOH} = \text{OH}^- + \text{H}_2\text{CO}_3$
D.	$\text{CO}_3^{2-} + \text{HOH} = \text{HCO}_3^- + \text{OH}^-$

10. Which of the following molecular equations represents the 1 st step hydrolysis of $(\text{NH}_4)_3\text{PO}_4$:

A.	$(\text{NH}_4)_3\text{PO}_4 + \text{HOH} = (\text{NH}_4)_2\text{HPO}_4 + \text{NH}_4\text{OH}$
B.	$(\text{NH}_4)_3\text{PO}_4 + 2\text{HOH} = \text{NH}_4\text{H}_2\text{PO}_4 + 2\text{NH}_4\text{OH}$
C.	$(\text{NH}_4)_3\text{PO}_4 + 2\text{HOH} = \text{NH}_4\text{HPO}_4 + 2\text{NH}_4\text{OH}$
D.	$(\text{NH}_4)_3\text{PO}_4 + 3\text{HOH} = \text{H}_3\text{PO}_4 + 3\text{NH}_4\text{OH}$

11. Which of the following molecular equations represents the 1 st step hydrolysis of CuCl_2 :

A.	$\text{CuCl}_2 + 2\text{HOH} = \text{CuCl}_2 + 2\text{HCl}$
B.	$\text{CuCl}_2 + \text{HOH} = \text{CuOHCl} + \text{HCl}$
C.	$\text{CuCl}_2 + 2\text{HOH} = \text{Cu}(\text{OH})_2 + 2\text{HCl}$
D.	$\text{CuCl}_2 + \text{HOH} = \text{OHCl} + \text{CuHCl}$

12. Which of the following molecular equations represents the 1 st step hydrolysis of BaS :

A.	$3\text{BaS} + 2\text{HOH} = (\text{BaOH})_2\text{S} + \text{Ba}(\text{HS})_2$
B.	$2\text{BaS} + 2\text{HOH} = \text{Ba}(\text{HS})_2 + \text{Ba}(\text{OH})_2$
C.	$\text{BaS} + 2\text{HOH} = \text{H}_2\text{S} + \text{Ba}(\text{OH})_2$
D.	$2\text{BaS} + 2\text{HOH} = \text{H}_2\text{S} + (\text{BaOH})_2\text{S}$

13. Which of the following net-ionic equations represents the 1 st step hydrolysis of LiNO_2 :

A.	$\text{LiNO}_2 + \text{HOH} = \text{OH}^- + \text{H}^+$
B.	$\text{NO}_2^- + \text{HOH} = \text{OH}^- + \text{H}^+$
C.	$\text{NO}_2^- + \text{HOH} = \text{OH}^- + \text{HNO}_2$
D.	$\text{Li}^+ + \text{HOH} = \text{LiOH} + \text{H}^+$

14. Which of the following substances are acidic when mixed with water?

A.	KBr, NaCl, BaSO_4	C.	NH_4Cl , AlCl_3 , ZnSO_4
B.	$\text{Ca}(\text{NO}_3)_2$, BaF_2 , CuCl_2	D.	RbNO_3 , NH_4Br , NH_4NO_3

15. Which of the following solutions is likely to have a pH greater than 7?

A.	sodium chloride	C.	potassium carbonate
B.	ammonium chloride	D.	pure water

ATOMIC STRUCTURE AND CHEMICAL BONDING

TEST QUESTIONS

1. The element with electron configuration $1s^2 2s^2 2p^6 3s^2 3p^2$ is:

A.	Mg (Z = 12)	C.	S (Z = 16)
B.	C (Z = 6)	D.	Si (Z = 14)

2. What is the electron configuration for nitrogen, atomic number 7?

A.	$1s^2 2s^2 2p^3$	C.	$1s^2 2s^3 2p^2$
B.	$1s^2 2s^3 2p^1$	D.	$1s^2 2s^2 2p^2 3s^1$

3. The possible values of an electron's spin quantum number are:

A.	-1, 0, or 1	C.	+1 or -1
B.	+1/2 or -1/2	D.	0 or 1

4. How many electrons can occupy the *s* orbital?

A.	two, if they have opposite spins	C.	one
B.	two, if they have the same spin	D.	no more than eight

5. If *n* is the principal quantum number of a main energy level, the number of electrons in that energy level is:

A.	n	C.	n^2
B.	2n	D.	$2n^2$

6. The nucleus of an atom consists of:

A.	electrons	C.	protons and neutrons
B.	neutrons	D.	protons, neutrons, and electrons

7. The three basic components of an atom are:

A.	protons, neutrons, and ions	C.	protium, deuterium, and tritium
B.	neutrons, protons, neutrons, and electrons	D.	protium, deuterium, and neutrons

8. Which two particles would be attracted to each other?

A.	electrons and neutrons	C.	protons and neutrons
B.	electrons and protons	D.	ions and neutrons

9. Put in a sentence a missing word:

An	consists of protons, neutrons and electrons.
----------	--

10. A typical atom consists of parts, which include protons, neutrons, and electrons. Each electron has a positive electrical charge. Each proton has a negative electrical charge.

A.	yes
B.	no

11. A nuclear particle that has about the same mass as a proton, but with no electrical charge, is called a(n)

A.	nuclide	C.	neutron.
B.	electron	D.	isotope.

12. The nucleus of an atom has all of the following characteristics EXCEPT that it

A.	is positively charged	C.	is very dense
B.	contains nearly all of the atom's mass	D.	contains nearly all of the atom's volume

13. Isotopes of an element contain different numbers of ...

A.	electrons	C.	protons
B.	neutrons	D.	nuclides

14. How many quantum numbers are needed to describe the energy state of an electron in an atom?

A.	1	C.	3
B.	2	D.	4

15. How many electrons are needed to completely fill the fourth energy level?

A.	8	C.	32
B.	18	D.	40

16. Point the correspondence between formula of compound and type of a chemical bond:

A.	BaCl ₂	1.	An ionic bond
B.	Mg	2.	A non-polar covalent bond
C.	Cl ₂	3.	A polar covalent bond
D.	NH ₃	4.	A metallic bond

A. _____, B. _____, C. _____, D. _____.

17. Point the correspondence between symbol of chemical element and its electron formula:

A.	[Ar]4s ² 4p ³	1.	As
B.	[Ar]4s ² 3d ¹	2.	Sc
C.	[He]2s ² 2p ⁵	3.	Cr
D.	[Ar]4s ¹ 3d ⁵	4.	F

A. _____, B. _____, C. _____, D. _____.

18. What electrical charge has a single proton?

A.	no charge	C.	negative charge
B.	positive charge	D.	either a positive or negative charge

19. Silicon-30 contains 14 protons. It also contains:

A.	16 electrons	C.	30 neutrons
B.	16 neutrons	D.	44 neutrons

20. Chemical bond in I₂ molecule is...

A.	covalent non-polar	C.	ionic
B.	covalent polar	D.	zero-order

21. Chemical bond in NaF molecule is...

A.	covalent non-polar	C.	ionic
B.	covalent polar	D.	zero-order

22. Chemical bond in PH₃ molecule is...

A.	covalent non-polar	C.	ionic
B.	covalent polar	D.	zero-order

23. What is electron configuration for the Mg²⁺ ion?

A.	1s ² 2s ² 2p ⁶ 3s ² 3p ²	D.	1s ² 2s ² 2p ⁶ 3s ² 3p ⁶ 4s ² 3d ²
B.	1s ² 2s ² 2p ⁶ 3s ¹	E.	1s ² 2s ² 2p ⁶
C.	1s ² 2s ² 2p ⁶ 3s ² 3p ⁶		

24. Which of the following is an ionic compound?

A.	NF ₃	D.	N ₂
B.	Na ₂ O	E.	CH ₄
C.	CO ₂		

25. Use the following information to identify the atom or ion: 8 protons, 8 neutrons, and 10 electrons.

A.	S ²⁺	D.	S ²⁻
B.	O ²⁻	E.	Ne
C.	O ²⁺		

26. The ground state electron configuration of Nickel (₂₈Ni) is given by

A.	1s ² 2s ² 2p ⁶ 3s ² 3p ⁶ 4s ²	D.	1s ² 2s ² 3s ² 3p ⁶ 3d ⁸
B.	1s ² 2s ² 2p ⁶ 3s ² 3p ⁴ 4s ² 3d ⁸	E.	1s ² 2s ² 2p ⁶ 3s ² 3p ⁶ 4d ⁵ 4s ¹ 4p ³
C.	[Kr]5s ² 4d ⁶		

27. Which one of the following is the correct orbital diagram for ground state of nitrogen (₇N)?

A.	$\boxed{\uparrow\downarrow}$	$\boxed{\uparrow\downarrow}$	$\boxed{\uparrow}$	$\boxed{\uparrow}$	$\boxed{\uparrow}$	C.	$\boxed{\uparrow\downarrow}$	$\boxed{\uparrow}$	$\boxed{\uparrow\downarrow}$	$\boxed{\uparrow}$	$\boxed{\uparrow}$
	1s	2s	2p				1s	2s	2p		
B.	$\boxed{\uparrow\downarrow}$	$\boxed{\uparrow\downarrow}$	$\boxed{\uparrow\downarrow}$	$\boxed{\uparrow}$	$\boxed{}$	D.	$\boxed{\uparrow\downarrow}$	$\boxed{\uparrow\downarrow}$	$\boxed{\uparrow\downarrow}$	$\boxed{\uparrow}$	$\boxed{\uparrow}$
	1s	2s	2p				1s	2s	2p		

28. Which of the following has 2 unpaired electrons?

A.	As	D.	Ge ⁺²
B.	P	E.	Si
C.	S ⁻²		

29. Which of the following elements has the largest atomic radius?

A.	Si	D.	C
B.	O	E.	Mg
C.	Be		

30. The ability of an atom in a molecule to attract electron density to itself is termed ...

A.	paramagnetism	D.	electron affinity
B.	diamagnetism	E.	ionization potential
C.	electronegativity		

OXIDATION-REDUCTION REACTIONS

LABORATORY TRAINING

Experiment 1. Reducing properties of metal ions of lower oxidation number

Add 1-2 drops of Iron (III) Sulfate $\text{Fe}_2(\text{SO}_4)_3$ to 2-3 drops of Ammonium Thiocyanate NH_4SCN . Register color of solution. Further add to the same tube the solution of Tin (II) Chloride SnCl_2 (by drops) up to whole decoloration of mixture. Draw up equations of reactions. For Redox reaction write the scheme of electron balance, and mark oxidizing and reducing agents.

REACTION	
OBSERVATION	
REACTION	
OBSERVATION	

Experiment 2. Influence of medium to redox reactions

Put 2-3 drops of Potassium Permanganate solution KMnO_4 in the 3 tubes. Add 1-2 drops of diluted Sulfuric acid to the first tube; - 5-10 drops of distilled water to the second one; 2-3 drops of alkali solution (NaOH or KOH) in the third one. Is the color of KMnO_4 solution changed? Then add in each tube few crystals of Sodium Sulfite or Potassium Sulfite. What do you observe? Draw up equations of reactions, write the scheme of electron balance, mark oxidizing and reducing agents.

REACTION	
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OBSERVATION	
REACTION	
OBSERVATION	
REACTION	
OBSERVATION	

Experiment 3. Redox properties of Chromium compounds

A) Put 2-3 drops of Potassium Dichromate $K_2Cr_2O_7$ solution in to the tube, add 2-3 drops of diluted Sulfuric acid and few crystals of solid Sodium Sulfite or Potassium Sulfite. What do you observe? Draw up equations of reactions, write the scheme of electron balance, mark oxidizing and reducing agents.

REACTION	
OBSERVATION	

B) Put 1-2 drops of alkali solution (NaOH or KOH), add by drops the solution of Chromium (III) Sulfate $Cr_2(SO_4)_3$ up to the formation of precipitate. Register the color of mixture. Further add alkali solution (NaOH or KOH) to the same tube up to the dissolving of precipitate (register the color of solution), then - 3-4 drops of 10% solution of Hydrogen Peroxide H_2O_2 . Heat a content of tube during 2-3 minutes in water bath. How is the a color of solution changed? Draw up equations of

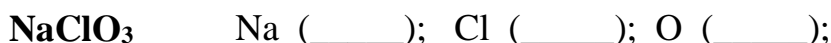
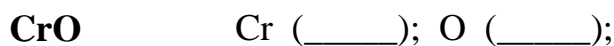
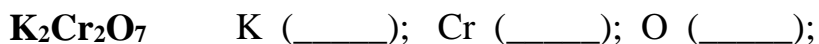
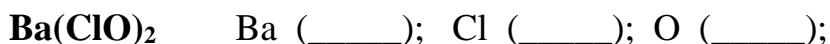
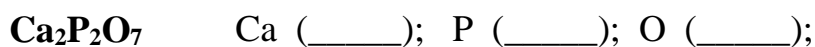
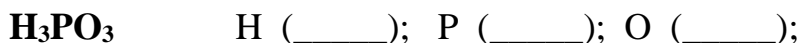
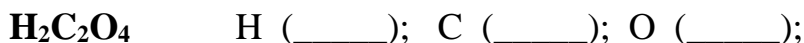
reactions of corresponded transformations: $\text{Cr}^{3+} \rightarrow \text{Cr}(\text{OH})_3 \downarrow \rightarrow \text{CrO}_2^- \rightarrow \text{CrO}_4^{2-}$. For Redox reaction write the scheme of electron balance, mark oxidizing and reducing agents.

REACTION	
OBSERVATION	
REACTION	
OBSERVATION	
REACTION	
OBSERVATION	

SKILL-DEVELOPING EXERCISES

Write answer:

1. Determine oxidation number of each element on the following compounds:

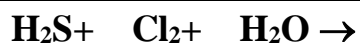


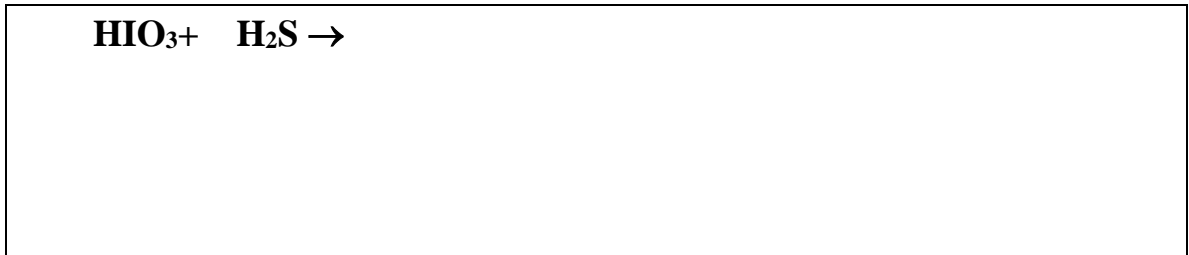
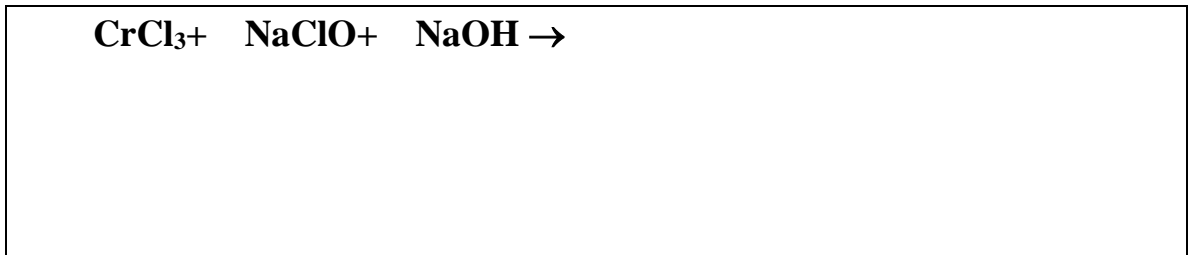
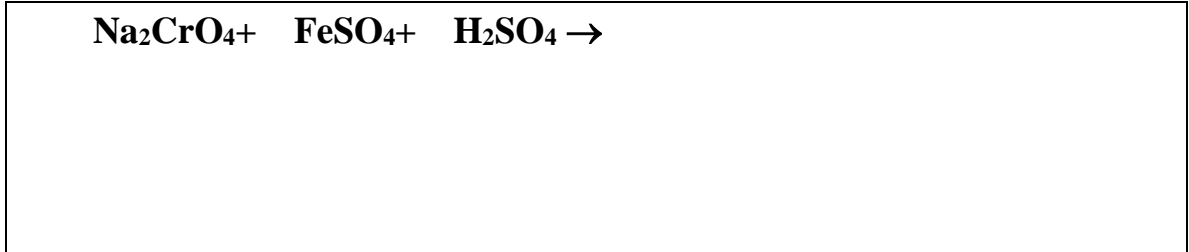
2. Balance Redox reactions using method of electron balance. Determine reducing and oxidizing agents.





3. Complete and balance Redox reactions, determine reducing and oxidizing agents:





TEST QUESTIONS

1. Define the following terms:

E. Oxidation -

F. Reduction -

G. Oxidizing agent -

H. Reducing agent -

2. In a particular Redox reaction? The oxidation number of Phosphorus changed from - 3 to 0. From this it may be concluded that Phosphorus:

A.	Lost 3 electrons and was reduced	C.	Gained 3 electrons and was reduced
B.	Lost 3 electrons and was oxidized	D.	Gained 3 electrons and was oxidized

3. Which species is losing electrons in the following Redox reaction:



A.	H	C.	O
B.	Cl	D.	Sn

4. Which of the following reaction represents a Redox reaction?

A.	$2\text{HCl} + \text{Na}_2\text{SO}_3 \rightarrow 2\text{NaCl} + \text{H}_2\text{O} + \text{SO}_2$
B.	$\text{CuS} + \text{H}_2 \rightarrow \text{H}_2\text{S} + \text{Cu}$
C.	$\text{AgNO}_3 + \text{NaCl} \rightarrow \text{AgCl} + \text{NaNO}_3$
D.	$\text{H}_2\text{CO}_3 \rightarrow \text{CO}_2 + \text{H}_2\text{O}$

5. What is the oxidation number of Nitrogen in NH_4Cl ?

A.	-3	C.	0
B.	-2	D.	+1

6. Electrons are lost by the:

A.	Reducing agent as it	C.	Oxidizing agent as it
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	undergoes oxidation		undergoes oxidation
B.	Reducing agent as it undergoes reduction	D.	Oxidizing agent as it undergoes reduction

7. Balance the following Redox reaction.

What element is a reducing agent?

$\text{NaNO}_2 + \text{KMnO}_4 + \text{H}_2\text{SO}_4 \rightarrow \text{MnSO}_4 + \text{NaNO}_3 + \text{K}_2\text{SO}_4 + \text{H}_2\text{O}$			
A.	N	C.	S
B.	Mn	D.	O

8. In which of the following half-reactions Chromium is oxidation agent?

A.	$\text{Cr}^{+3} \rightarrow \text{Cr}^0$	C.	$\text{Cr}^{+3} \rightarrow \text{Cr}^{+2}$
B.	$2\text{Cr}^{+3} \rightarrow \text{Cr}_2\text{O}_7^{2-}$	D.	$2\text{CrO}_4^{2-} \rightarrow \text{Cr}_2\text{O}_7^{2-}$

9. Balance the following Redox reaction using method of electron balance. Determine reducing and oxidizing agents:

$\text{FeS}_2 + \text{O}_2 \rightarrow \text{Fe}_2\text{O}_3 + \text{SO}_2$			
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10. Copper has an oxidation number of +1 in:

A.	$\text{Cu}(\text{CH}_3\text{COO})_2$	C.	CuC_2O_4
B.	CuBr	D.	CuO

11. Oxidation is defined as:

A.	gain of a proton	C.	gain of an electron
B.	loss of a proton	D.	loss of an electron

12. Oxidation is defined as:

A.	gain of a proton	C.	gain of an electron
B.	loss of a proton	D.	loss of an electron

13. What is the oxidation number of the Chlorine atom in the HClO_4 molecule?

A.	-1	D.	+7
B.	+3	E.	+9
C.	+5		

14. The reaction, $\text{AgNO}_3 + \text{NH}_4\text{Br} \rightarrow \text{AgBr} + \text{NH}_4\text{NO}_3$, involves changes in oxidation numbers and is therefore classified as a redox reaction.

A.	True	D.	False
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15. Is the process, $\text{MnO}_4^- \rightarrow \text{Mn}^{+4}$ an oxidation or a reduction?

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16. Which of the following processes represents a reduction?

A.	$\text{Ba}^{2+} + \text{CrO}_4^{2-} \rightarrow \text{BaCrO}_4$	D.	$\text{MnO}_2 \rightarrow \text{MnO}_4^-$
B.	$2\text{H}^+ + \text{CO}_3^{2-} \rightarrow \text{CO}_2 + \text{H}_2\text{O}$	E.	$2\text{CrO}_4^{2-} + 2\text{H}^+ \rightarrow \text{Cr}_2\text{O}_7^{2-} + \text{H}_2\text{O}$
C.	$\text{CrO}_4^{2-} \rightarrow \text{Cr}^{+3}$		

17. What is the oxidation number of the Vanadium atom in the compound $(\text{NH}_4)_3\text{VO}_4$?

A.	+2	D.	+6
B.	+3	E.	+7
C.	+5		

COMPLEX (COORDINATION) COMPOUNDS

LABORATORY TRAINING

Experiment 1. Compound formation with complex cation

Add by drops ammonia solution firstly up to the formation of HydroxoSulfate and then up to its complete dilution and formation of complex compound $[\text{Cu}(\text{NH}_3)_4]\text{SO}_4$ into the test-tube with 4-5 drops of cuprous Sulfate or Chloride (II). How does the initial color of the solution change? Equate the reaction according to the process stages of reagent interactions.

REACTION	
OBSERVATION	
REACTION	
OBSERVATION	

Experiment 2. Compound formation with complex anion

A) Put 4-5 drops of $\text{Hg}(\text{NO}_3)_2$ solution into the test-tube and add by one drop of KI solution up to the formation of red precipitate of Mercury (II) iodide. Continue to add KI solution up to complete dilution of precipitate and formation of soluble complex compound - Potassium Tetraiodomercurate (II). Equate the reaction taking into account the stages.

REACTION	
OBSERVATION	

REACTION	
OBSERVATION	

B) Put 5-7 drops of Cobalt (II) Sulfate or Chloride solution into the test-tube. Add little by little the granules of KNO_2 up to solution saturation (there are some undissolved salt at the bottom of test-tube). Add by drops concentrated acetic acid to the mixture up to the formation of yellow precipitate of $\text{K}_3[\text{Co}(\text{NO}_2)_6]$. Equate the reaction according to stages taking into account that nitrous acid, which is isolated under acidation, oxidizes Co^{2+} to Co^{3+} :

REACTION	$\text{CoCl}_2 + \text{KNO}_2 \rightarrow$
OBSERVATION	
REACTION	$\text{KNO}_2 + \text{CH}_3\text{COOH} \rightarrow \text{HNO}_2 + \dots$
OBSERVATION	
REACTION	$\text{Co}(\text{NO}_2)_2 + \text{HNO}_2 \rightarrow \text{NO} + \dots$
OBSERVATION	
REACTION	$\text{Co}(\text{NO}_2)_3 + \text{KNO}_2 \rightarrow$

OBSERVATION	

Experiment 3. Complex compounds in metathesis

A) Add 2-3 drops of $K_3[Fe(CN)_6]$ solution to 4-5 drops of any Iron (II) salt solution. What is the color of precipitate? Equate molecular and ionic reactions.

REACTION	
OBSERVATION	

B) Add 2-3 drops of $K_4[Fe(CN)_6]$ solution to 4-5 drops of any Cooper (II) salt solution. Identify the color of precipitate and equate molecular and ionic reactions.

REACTION	
OBSERVATION	

C) Add 2-3 drops of $K_4[Fe(CN)_6]$ solution to 4-5 drops of Iron (III) salt solution. Identify the color of precipitate and equate molecular and ionic reactions.

REACTION	
OBSERVATION	

D) Add 2-3 drops of KSCN or NH₄SCN solution to 4-5 drops of Iron (III) salt solution. Identify the color of precipitate and equate molecular and ionic reactions.

REACTION	
OBSERVATION	

SKILL-DEVELOPING EXERCISES

Write answer:

1. Write the molecular formulas of the following coordination compounds:

1. hexaammineiron(III) nitrate_____

2. ammonium tetrachlorocuprate(II)_____

3. sodium monochloropentacyanoferrate(III)_____

4. potassium hexafluorocobaltate(III)_____

2. Give the name of the following coordination compounds:

1. $[\text{Co}(\text{NH}_3)_5\text{Br}]\text{SO}_4$ _____

2. $[\text{Fe}(\text{NH}_3)_6][\text{Cr}(\text{CN})_6]$ _____

3. $[\text{Ni}(\text{H}_2\text{O})_6](\text{NO}_3)_2$ _____

4. $[\text{Fe}(\text{OH})(\text{H}_2\text{O})_5]\text{Cl}_2$ _____

3. In the compound $[\text{Cu}(\text{NH}_3)_4]\text{SO}_4$, what are:

inner sphere_____

outer sphere_____

central atom_____

ligands_____

coordination number_____

4. In the compound $\text{Na}_3[\text{Ni}(\text{H}_2\text{O})(\text{CN})_5]$, what are:

inner sphere_____

outer sphere_____

central atom_____

ligands_____

coordination number_____

5. In the compound $[\text{Cr}(\text{H}_2\text{O})_4\text{Cl}_2]$, what are:

inner sphere_____

outer sphere_____

central atom_____

ligands_____

coordination number _____

6. In the compound $K[AuCl_4]$, what are:

inner sphere _____

outer sphere _____

central atom _____

ligands _____

coordination number _____

7. Complete complexation reactions :

a) $Cu(NO_3)_2 + NH_3 \rightarrow$ _____

b) $Zn(OH)_2 + KOH_{(excess)} \rightarrow$ _____

c) $CdI_2 + KI_{(excess)} \rightarrow$ _____

d) $CoCl_3 + NH_4OH_{(excess)} \rightarrow$ _____

e) $Fe(CN)_3 + KCN \rightarrow$ _____

8. Write chemical reactions between the following substances:

a) $ZnSO_4 + K_3[Fe(CN)_6] \rightarrow$ _____

b) $KCl + Na_3[Co(NO)_2] \rightarrow$ _____

c) $FeCl_3 + K_3[Fe(CN)_6] \rightarrow$ _____

TEST QUESTIONS

1. What is the charge of the complex ion in $\text{Mg}_2[\text{Fe}(\text{NO}_3)_6]$?

A.	2-	D.	3+
B.	2+	E.	4-
C.	3-		

2. What is the oxidation number of chromium in $[\text{Cr}(\text{NH}_3)_4\text{Cl}_2]\text{Cl}$?

A.	-3	D.	-2
B.	+3	E.	0
C.	+2		

3. What is the ligand in $\text{Ca}_3[\text{Fe}(\text{CN})_6]_2$?

A.	Ca^{2+}	D.	$\text{Fe}(\text{CN})_6^{3-}$
B.	Fe^{+3}	E.	Fe^{+2}
C.	CN^-		

4. What is the charge of the central metal ion in $\text{K}_4[\text{Fe}(\text{CN})_6]$?

A.	0	D.	3+
B.	1+	E.	6+
C.	2+		

6. The charge of the complex ion in $[\text{Zn}(\text{H}_2\text{O})_3\text{Cl}]\text{Cl}$ is:

A.	0	D.	1+
B.	1-	E.	2-
C.	2+		

6. The coordination number for $[\text{Zn}(\text{NH}_3)\text{H}_2\text{O}]\text{SO}_4$ is:

A.	5	D.	1
B.	4	E.	6
C.	2		

7. What is the oxidation state of iron in $\text{CaNa}[\text{Fe}(\text{CN})_6]$?

A.	0	D.	+4
B.	+2	E.	+6
C.	+3		

8. What is the coordination number of cobalt in $[\text{Co}(\text{NH}_3)_5\text{I}](\text{NO}_3)_2$?

A.	12	D.	2
B.	8	E.	6
C.	4		

9. What is the metal ion in the porphine?

A.	iron	D.	magnesium
B.	calcium	E.	chlorophyll
C.	molybdenum		

10. EDTA is _____-dantate ligand.

A.	mono	D.	tetra
B.	bi	E.	hexa

C.	tri
----	-----

11. The correct name for $[\text{Ni}(\text{NH}_3)_6](\text{NO}_3)_3$ is:

A.	dinitrohexaamminenickel (II)	D.	hexaamminenickel (II) nitrate
B.	hexaamminenickel (III) trinitrate	E.	hexaamminenickel (III) nitrate
C.	dinitrohexaamminenickelate (III)		

12. The correct name for $\text{Na}_3[\text{CoF}_6]$ is:

A.	trisodium hexakisfluorocobalt(III)	D.	sodium hexafluorocobaltate(III)
B.	trisodium hexakisfluorocobalt(II)	E.	sodium hexafluorocobaltate(IV)
C.	trisodium hexakisfluorocobalt(IV)		

13. What is the oxidation number of Ni in $[\text{Ni}(\text{CN})_5]^{3-}$?

A.	1+	D.	4+
B.	2+	E.	5+
C.	3+		

14. Which of the following complexes has a coordination number of 6?

A.	$[\text{Co}(\text{en})_2\text{Cl}_2]^+$	D.	$[\text{Ag}(\text{NH}_3)_2]^+$
B.	$[\text{Pt}(\text{NH}_3)_2\text{Cl}_2]$	E.	None of these complexes has coordination number 6
C.	$[\text{Cu}(\text{NH}_3)_4]^{2+}$		

15. What metal is complexed in chlorophyll?

A.	iron	D.	vanadium
B.	chromium	E.	magnesium
C.	manganese		

16. Which one of the following is the correct formula for potassium diaqua-tetrachloromolybdate (III)?

A.	$\text{K}_2[\text{Mo}(\text{H}_2\text{O})_2\text{Cl}_4]$	D.	$\text{Mo}[\text{K}(\text{H}_2\text{O})_2]\text{Cl}_4$
B.	$\text{K}[\text{Mo}(\text{H}_2\text{O})_2\text{Cl}_2]\text{Cl}_2$	E.	$\text{K}_3[\text{Mo}(\text{H}_2\text{O})_2\text{Cl}_4]$
C.	$\text{K}[\text{Mo}(\text{H}_2\text{O})_2\text{Cl}_4]$		

17. The coordination sphere of a complex consists of:

A.	the central metal ion only	D.	the primary and secondary valencies
B.	the ligands	E.	coordination and steric numbers
C.	the central metal ion and the ligands bonded to it		

18. What is the oxidation state of chromium in $[\text{Cr}(\text{H}_2\text{O})_4\text{Cl}_2]^+$?

A.	0	D.	-3
B.	+2	E.	+3

Literature

- В.А.Копілевич, В.Є.Косматий, Л.В.Войтенко, Л.М.Абарбарчук, Т.К.Панчук, Л.В.Гаєвська, А.П.Попель, К.О.Чеботько, В.В.Трачевський. Аналітична хімія. (для аграрних спеціальностей).- м.Київ - 2003-300с.
- В.В.Сухан, Т.В.Табенська, А.Й.Капустян, В.Ф.Горлач. Хімія (посібник для вступників до вищих навчальних закладів).- м.Київ-1996-440с.
- В.Є.Косматий, В.А.Копілевич, Л.В.Войтенко, Л.М.Абарбарчук, Т.К.Панчук, Л.В.Гаєвська. Лабораторний практикум з неорганічної хімії.-м.Київ-2001-154с.
- О.І.Карнаухов, Д.О.Мельничук, К.О.Чеботько, В.А.Копілевич. Загальна та біонеорганічна хімія.- м.Київ-2001-578с.
- С.А.Неділько, П.П.Попель. Загальна й неорганічна хімія. (Задачі і вправи).-м.Київ-2001-328с.
- П.Г.Нагорний, О.В.Петренко. Хімія (Посібник для студентів факультетів нехімічних спеціальностей). - м.Київ-2004-128с.

Appendixes

Appendix 1. Ionization constants of some acids and bases 25⁰ C

Compounds 1	Expression for ion. const. 2	K _{ion} 3	pK = -lg K _{dis} 4
HF	$K = \frac{[H^+][F^-]}{[HF]}$	$6,8 \cdot 10^{-4}$	3,20
HCl	$K = \frac{[H^+][Cl^-]}{[HCl]}$	$1,0 \cdot 10^7$	-7,00
HBr	$K = \frac{[H^+][Br^-]}{[HBr]}$	$1,0 \cdot 10^9$	-9,00
HI	$K = \frac{[H^+][I^-]}{[HI]}$	$1,0 \cdot 10^{11}$	-11,00
HClO	$K = \frac{[H^+][ClO^-]}{[HClO]}$	$3,0 \cdot 10^{-8}$	7,53
HClO ₂	$K = \frac{[H^+][ClO_2^-]}{[HClO_2]}$	$1,1 \cdot 10^{-2}$	1,97
HBrO	$K = \frac{[H^+][BrO^-]}{[HBrO]}$	$2,5 \cdot 10^{-9}$	8,66
HIO	$K = \frac{[H^+][IO^-]}{[HIO]}$	$2,3 \cdot 10^{-11}$	10,64
HIO ₃	$K = \frac{[H^+][IO_3^-]}{[HIO_3]}$	$1,6 \cdot 10^{-11}$	10,80
H ₂ S	$K_1 = \frac{[H^+][HS^-]}{[H_2S]}$ $K_2 = \frac{[H^+][S^{2-}]}{[HS^-]}$	$1,0 \cdot 10^{-7}$ $2,5 \cdot 10^{-18}$	6,99 12,60
H ₂ SO ₃	$K_1 = \frac{[H^+][HSO_3^-]}{[H_2SO_3]}$ $K_2 = \frac{[H^+][SO_3^{2-}]}{[HSO_3^-]}$	$1,4 \cdot 10^{-2}$ $6,2 \cdot 10^{-8}$	1,85 7,20
H ₂ SO ₄	$K_1 = \frac{[H^+][HSO_4^-]}{[H_2SO_4]}$	$1,0 \cdot 10^3$ $1,2 \cdot 10^{-2}$	-3,00 1,94

	$K_2 = \frac{[H^+][SO_4^{2-}]}{[HSO_4^-]}$		
1	2	3	4
HNO ₂	$K = \frac{[H^+][NO_2^-]}{[HNO_2]}$	6,9·10 ⁻⁴	3,16
H ₃ PO ₄	$K_1 = \frac{[H^+][H_2PO_4^-]}{[H_3PO_4]}$ $K_2 = \frac{[H^+][HPO_4^{2-}]}{[H_2PO_4^-]}$ $K_3 = \frac{[H^+][PO_4^{3-}]}{[HPO_4^{2-}]}$	7,1·10 ⁻³ 6,2·10 ⁻⁸ 5,0·10 ⁻¹³	2,15 7,21 12,00
H ₄ P ₂ O ₇	$K_1 = \frac{[H^+][H_3P_2O_7^-]}{[H_4P_2O_7]}$ $K_2 = \frac{[H^+][H_2P_2O_7^{2-}]}{[H_3P_2O_7^-]}$ $K_3 = \frac{[H^+][HP_2O_7^{3-}]}{[H_2P_2O_7^{2-}]}$ $K_4 = \frac{[H^+][P_2O_7^{4-}]}{[HP_2O_7^{3-}]}$	1,2·10 ⁻¹ 7,9·10 ⁻³ 2,0·10 ⁻⁷ 4,8·10 ⁻¹⁰	0,91 2,10 6,70 9,32
H ₃ AsO ₄	$K_1 = \frac{[H^+][H_2AsO_4^-]}{[H_3AsO_4]}$ $K_2 = \frac{[H^+][HAsO_4^{2-}]}{[H_2AsO_4^-]}$ $K_3 = \frac{[H^+][AsO_4^{3-}]}{[HAsO_4^{2-}]}$	5,6·10 ⁻³ 1,7·10 ⁻⁷ 3,0·10 ⁻¹²	2,25 6,77 11,53
HAsO ₂	$K = \frac{[H^+][AsO_2^-]}{[HAsO_2]}$	6·10 ⁻¹⁰	9,22
HCOOH	$K = \frac{[H^+][HCOO^-]}{[HCOOH]}$	1,6·10 ⁻¹²	11,80
CH ₃ COOH	$K = \frac{[H^+][CH_3COO^-]}{[CH_3COOH]}$	1,75·10 ⁻⁵	4,75
HCN	$K = \frac{[H^+][CN^-]}{[HCN]}$	5,0·10 ⁻¹⁰	9,30

H_2CO_3 (“удавана” константа)	$K_1 = \frac{[\text{H}^+][\text{HCO}_3^-]}{[\text{H}_2\text{CO}_3]}$ $K_2 = \frac{[\text{H}^+][\text{CO}_3^{2-}]}{[\text{HCO}_3^-]}$	$4,45 \cdot 10^{-7}$ $4,69 \cdot 10^{-11}$	6,35 10,33
1	2	3	4
$\text{H}_2\text{C}_2\text{O}_4$	$K_1 = \frac{[\text{H}^+][\text{HC}_2\text{O}_4^-]}{[\text{H}_2\text{C}_2\text{O}_4]}$ $K_2 = \frac{[\text{H}^+][\text{C}_2\text{O}_4^{2-}]}{[\text{HC}_2\text{O}_4^-]}$	$5,6 \cdot 10^{-2}$ $5,4 \cdot 10^{-5}$	1,25 4,27
H_2SiO_3	$K_1 = \frac{[\text{H}^+][\text{HSiO}_3^-]}{[\text{H}_2\text{SiO}_3]}$ $K_2 = \frac{[\text{H}^+][\text{SiO}_3^{2-}]}{[\text{HSiO}_3^-]}$	$2,2 \cdot 10^{-10}$ $1,6 \cdot 10^{-12}$	9,66 11,80
H_2SnO_2	$K_1 = \frac{[\text{H}^+][\text{HSnO}_2^-]}{[\text{H}_2\text{SnO}_2]}$	10^{-15}	15,00
H_2SnO_3	$K_1 = \frac{[\text{H}^+][\text{HSnO}_3^-]}{[\text{H}_2\text{SnO}_3]}$	10^{-10}	10,00
H_2PbO_2	$K_1 = \frac{[\text{H}^+][\text{HPbO}_2^-]}{[\text{H}_2\text{PbO}_3]}$	10^{-11}	11,00
H_3BO_3	$K_1 = \frac{[\text{H}^+][\text{H}_2\text{BO}_3^-]}{[\text{H}_3\text{BO}_3]}$ $K_2 = \frac{[\text{H}^+][\text{HBO}_3^{2-}]}{[\text{H}_2\text{BO}_3^-]}$ $K_3 = \frac{[\text{H}^+][\text{BO}_3^{3-}]}{[\text{HBO}_3^{2-}]}$	$5,8 \cdot 10^{-10}$ $1,8 \cdot 10^{-13}$ $1,6 \cdot 10^{-14}$	9,22 12,74 13,80
$\text{H}_2\text{B}_4\text{O}_7$	$K_1 = \frac{[\text{H}^+][\text{HB}_4\text{O}_7^-]}{[\text{H}_2\text{B}_4\text{O}_7]}$ $K_2 = \frac{[\text{H}^+][\text{B}_4\text{O}_7^{2-}]}{[\text{HB}_4\text{O}_7^-]}$	$1,0 \cdot 10^{-4}$ $1,0 \cdot 10^{-9}$	4,00 9,00
HAlO_2	$K = \frac{[\text{H}^+][\text{AlO}_2^-]}{[\text{HAlO}_2]}$	$6,0 \cdot 10^{-15}$	14,22

H ₂ MnO ₄	$K_1 = \frac{[H^+][HMnO_4^-]}{[H_2MnO_4]}$ $K_2 = \frac{[H^+][MnO_4^{2-}]}{[HMnO_4^-]}$	$\sim 10^{-1}$ $7,1 \cdot 10^{-11}$	1 10,15
HMnO ₄	$K = \frac{[H^+][MnO_4^-]}{[HMnO_4]}$	$2,0 \cdot 10^2$	- 2,3
1	2	3	4
H ₂ CrO ₄	$K_1 = \frac{[H^+][HCrO_4^-]}{[H_2CrO_4]}$ $K_2 = \frac{[H^+][CrO_4^{2-}]}{[HCrO_4^-]}$	$1,6 \cdot 10^{-1}$ $3,2 \cdot 10^{-7}$	0,80 6,50
H ₂ MoO ₄	$K_1 = \frac{[H^+][HMoO_4^-]}{[H_2MoO_4]}$ $K_2 = \frac{[H^+][MoO_4^{2-}]}{[HMoO_4^-]}$	$2,9 \cdot 10^{-3}$ $1,4 \cdot 10^{-4}$	2,54 3,86
H ₂ PbO ₂	$K_1 = \frac{[H^+][HPbO_2^-]}{[H_2PbO_2]}$	$1 \cdot 10^{-12}$	12,00
AgOH	$K = \frac{[Ag^+][OH^-]}{[AgOH]}$	$1,1 \cdot 10^{-4}$	3,96
Al(OH) ₃	$K_3 = \frac{[Al^{3+}][OH^-]}{[AlOH^{2+}]}$	$1,38 \cdot 10^{-9}$	8,86
Ba(OH) ₂	$K_2 = \frac{[Ba^{2+}][OH^-]}{[Ba(OH)^-]}$	$2,3 \cdot 10^{-1}$	0,64
Ca(OH) ₂	$K_2 = \frac{[Ca^{2+}][OH^-]}{[Ca(OH)^-]}$	$4,0 \cdot 10^{-3}$	1,40
Cd(OH) ₂	$K_2 = \frac{[Cd^{2+}][OH^-]}{[CdOH^+]}$	$5,0 \cdot 10^{-3}$	2,80
Co(OH) ₂	$K_2 = \frac{[Co^{2+}][OH^-]}{[CoOH^+]}$	$4,0 \cdot 10^{-5}$	4,4
Cr(OH) ₃	$K_3 = \frac{[Cr^{3+}][OH^-]}{[CrOH^{2+}]}$	$1,02 \cdot 10^{-10}$	9,99

Cu(OH) ₂	$K_2 = \frac{[Cu^{2+}][OH^-]}{[CuOH^+]}$	$3,4 \cdot 10^{-7}$	6,47
Fe(OH) ₂	$K_2 = \frac{[Fe^{2+}][OH^-]}{[FeOH^+]}$	$1,3 \cdot 10^{-4}$	3,89
Fe(OH) ₃	$K_2 = \frac{[FeOH^{2+}][OH^-]}{[Fe(OH)_2^+]}$	$1,8 \cdot 10^{-11}$	10,74
	$K_3 = \frac{[Fe^{3+}][OH^-]}{[FeOH^{2+}]}$	$1,4 \cdot 10^{-12}$	11,87
1	2	3	4
Hg(OH) ₂	$K_1 = \frac{[Hg(OH)^+][OH^-]}{[Hg(OH)_2]}$	$7 \cdot 10^{-12}$	11,15
	$K_2 = \frac{[Hg^{2+}][OH^-]}{[HgOH^+]}$	$2 \cdot 10^{-23}$	22,70
LiOH	$K = \frac{[Li^+][OH^-]}{[LiOH]}$	$6,8 \cdot 10^{-1}$	0,17
Mn(OH) ₂	$K_2 = \frac{[Mn^{2+}][OH^-]}{[MnOH^+]}$	$5,0 \cdot 10^{-4}$	3,30
Mg(OH) ₂	$K_2 = \frac{[Mg^{2+}][OH^-]}{[MgOH^+]}$	$2,5 \cdot 10^{-3}$	2,60
NaOH	$K = \frac{[Na^+][OH^-]}{[NaOH]}$	5,9	- 0,77
Ni(OH) ₂	$K_2 = \frac{[Ni^{2+}][OH^-]}{[NiOH^+]}$	$2,5 \cdot 10^{-5}$	4,60
NH ₄ OH (“удавана” константа)	$K = \frac{[NH_4^+][OH^-]}{[NH_4OH]}$	$1,76 \cdot 10^{-5}$	4,755
Pb(OH) ₂	$K_1 = \frac{[Pb(OH)^+][OH^-]}{[Pb(OH)_2]}$	$9,6 \cdot 10^{-4}$	3,02
	$K_2 = \frac{[Pb^{2+}][OH^-]}{[Pb(OH)^+]}$	$3 \cdot 10^{-8}$	7,52

$\text{Sr}(\text{OH})_2$	$K_2 = \frac{[\text{Sr}^{2+}][\text{OH}^-]}{[\text{SrOH}^+]}$	$1,5 \cdot 10^{-1}$	0,82
$\text{Zn}(\text{OH})_2$	$K_2 = \frac{[\text{Zn}^{2+}][\text{OH}^-]}{[\text{Zn}(\text{OH})^+]}$	$4,0 \cdot 10^{-5}$	4,40

Appendix 2

Density of water solutions of the bases 18 °C, g/cm³

%	KOH	NaOH	NH ₃	%	KOH	NaOH	NH ₃
4	1,033	1,046	0,983	34	1,334	1,374	0,889
6	1,048	1,069	0,973	36	1,358	1,395	0,884
8	1,065	1,092	0,967	38	1,384	1,416	-
10	1,082	1,115	0,960	40	1,411	1,437	-
12	1,100	1,137	0,958	42	1,437	1,458	-
14	1,118	1,159	0,946	44	1,460	1,478	-
16	1,137	1,181	0,939	46	1,485	1,499	-
18	1,156	1,203	0,932	48	1,511	1,519	-
20	1,176	1,225	0,926	50	1,538	1,540	-
22	1,196	1,247	0,919	52	1,564	1,560	-
24	1,217	1,268	0,913	54	1,590	1,580	-
26	1,240	1,289	0,908	56	1,616	1,601	-
28	1,263	1,310	0,903	58	-	1,622	-
30	1,286	1,332	0,898	60	-	1,643	-
32	1,310	1,352	0,893				

Appendix 3

Density of water solutions of the acids 18 °C, g/cm³

%	H ₂ SO ₄	HNO ₃	HCl	%	H ₂ SO ₄	HNO ₃	HCl
4	1,027	1,020	1,018	52	1,415	1,322	-
6	1,038	1,031	1,028	54	1,435	1,334	-
8	1,052	1,043	1,038	56	1,456	1,345	-
10	1,066	1,054	1,047	58	1,477	1,356	-
12	1,080	1,066	1,057	60	1,498	1,367	-
14	1,095	1,078	1,068	62	1,520	1,377	-
16	1,109	1,090	1,078	64	1,542	1,387	-
18	1,124	1,103	1,088	66	1,565	1,396	-
20	1,139	1,115	1,098	68	1,587	1,405	-
22	1,155	1,128	1,108	70	1,611	1,413	-
24	1,170	1,140	1,119	72	1,634	1,422	-
26	1,176	1,153	1,129	74	1,657	1,430	-
28	1,202	1,167	1,139	76	1,681	1,438	-
30	1,219	1,180	1,149	78	1,704	1,445	-
32	1,235	1,193	1,159	80	1,727	1,452	-
34	1,252	1,207	1,169	82	1,749	1,459	-
36	1,268	1,221	1,179	84	1,769	1,466	-
38	1,286	1,234	1,189	86	1,787	1,372	-
40	1,303	1,246	1,198	88	1,802	1,477	-
42	1,321	1,259	-	90	1,814	1,483	-
44	1,338	1,272	-	92	1,824	1,487	-
46	1,357	1,285	-	94	1,8312	1,491	-
48	1,376	1,298	-	96	1,8355	1,495	-
50	1,395	1,310	-	98	1,8395	1,501	-
				100	1,8305	1,513	-

Solubility of salts and bases in water

Аніони	Катіони											
	K ⁺	Na ⁺	NH ₄ ⁺	Ba ²⁺	Sr ²⁺	Ca ²⁺	Mg ²⁺	Al ³⁺	Cr ³⁺	Fe ³⁺	Fe ²⁺	Mn ²⁺
OH ⁻	P	P	P	P	P	BP	P	H	H	H	H	H
F ⁻	P	P	P	BP	BP	BP	BP	BP	H	BP	BP	BP
Cl ⁻	P	P	P	P	P	P	P	P	BP	P	P	P
Br ⁻	P	P	P	P	P	P	P	P	BP	P	P	P
I ⁻	P	P	P	P	P	P	P	P	P	P	P	P
S ²⁻	P	P	P	P	P	P	BP	-	-	H	H	H
SO ₃ ²⁻	P	P	P	H	H	H	BP	-	-	-	H	-
SO ₄ ²⁻	P	P	P	H	BP	BP	P	P	P	P	P	P
PO ₄ ³⁻	P	P	P	H	H	H	H	H	BP	H	H	H
CrO ₄ ²⁻	P	P	P	H	BP	P	P	-	-	-	-	H
CO ₃ ²⁻	P	P	P	H	H	H	H	-	-	-	H	H
NO ₃ ⁻	P	P	P	P	P	P	P	P	P	P	P	P
NO ₂ ⁻	P	P	P	P	P	P	P	P	P	P	P	P
CH ₃ COO ⁻	P	P	P	P	P	P	P	P	P	P	P	P

Примітка: P - розчинні у воді; BP - важкорозчинні (малорозчинні); H - практично нерозчинні; риска означає, що речовина не існує або розкладається водою.

Аніони	Катіони									
	Zn ²⁺	Ni ²⁺	Co ²⁺	Ag ⁺	Hg ²⁺	Cu ²⁺	Cd ²⁺	Pb ²⁺	Bi ³⁺	Sn ²⁺
OH ⁻	H	H	H	BP	H	H	H	H	H	H
F ⁻	BP	P	P	P	-	BP	BP	H	H	P
Cl ⁻	P	P	P	H	P	P	P	BP	-	P
Br ⁻	P	P	P	H	P	P	P	BP	-	P
I ⁻	P	P	P	H	H	H	P	BP	H	P
S ²⁻	H	H	H	H	H	H	H	H	H	H
SO ₃ ²⁻	H	H	-	H	-	-	H	H	-	-
SO ₄ ²⁻	P	P	P	BP	-	P	P	BP	P	P
PO ₄ ³⁻	H	H	H	H	H	H	H	H	H	-
CrO ₄ ²⁻	H	H	H	H	H	H	H	H	H	H
CO ₃ ²⁻	H	H	H	H	H	H	H	H	H	-
NO ₃ ⁻	P	P	P	P	P	P	P	P	P	-
NO ₂ ⁻	P	P	P	BP	-	P	P	P	H	-
CH ₃ COO ⁻	P	P	P	BP	P	P	P	P	-	-

Standard electrode potentials (as reduction potentials)

ЕЛЕМЕНТ	ВИЩА СТУПІНЬ ОКИСНЕННЯ	КІЛЬКІСТЬ ЕЛЕКТРОНІВ	НИЖЧА СТУПІНЬ ОКИСНЕННЯ	E^0 , В
1	2	3	4	5
Ag	Ag^+	$+ e^-$	$\text{Ag} \downarrow$	+ 0,7994
	$\text{Ag}(\text{NH}_3)_2^+$	$+ e^-$	$\text{Ag} \downarrow + 2 \text{NH}_3$	+ 0,373
Al	Al^{3+}	$+ 3 e^-$	$\text{Al} \downarrow$	- 0,66
Ba	Ba^{2+}	$+ 2 e^-$	$\text{Ba} \downarrow$	- 2,91
Be	Be^{2+}	$+ 2 e^-$	$\text{Be} \downarrow$	- 1,97
Br	Br_2	$+ 2 e^-$	2Br^-	+ 1,087
	$2\text{HBrO} + 2 \text{H}^+$	$+ 2 e^-$	$\text{Br}_2 \uparrow + 2 \text{H}_2\text{O}$	+ 1,6
	$2\text{BrO}^- + 2 \text{H}_2\text{O}$	$+ 2 e^-$	$\text{Br}_2 \uparrow + 4 \text{OH}^-$	+ 0,45
	$\text{HBrO} + \text{H}^+$	$+ 2 e^-$	$\text{Br}^- + \text{H}_2\text{O}$	+ 1,34
	$\text{BrO}^- + \text{H}_2\text{O}$	$+ 2 e^-$	$\text{Br}^- + 2 \text{OH}^-$	+ 0,76
	$\text{BrO}_3^- + 5 \text{H}^+$	$+ 4 e^-$	$\text{HBrO} + 2 \text{H}_2\text{O}$	+ 1,45
	$\text{BrO}_3^- + 2 \text{H}_2\text{O}$	$+ 4 e^-$	$\text{BrO}^- + 4 \text{OH}^-$	+ 0,54
	$2\text{BrO}_3^- + 12 \text{H}^+$	$+ 10 e^-$	$\text{Br}_2 \uparrow + 6 \text{H}_2\text{O}$	+ 1,52
	$\text{BrO}_3^- + 6 \text{H}^+$	$+ 6 e^-$	$\text{Br}^- + 6 \text{H}_2\text{O}$	+ 1,45
Ca	Ca^{2+}	$+ 2 e^-$	$\text{Ca} \downarrow$	- 2,79
Cd	Cd^{2+}	$+ 2 e^-$	$\text{Cd} \downarrow$	- 0,403
Cl	Cl_2	$+ 2 e^-$	2Cl^-	+ 1,359
	$2\text{HClO} + 2\text{H}^+$	$+ 2 e^-$	$\text{Cl}_2 \uparrow + 4 \text{H}_2\text{O}$	+ 1,63
	$\text{HClO} + \text{H}^+$	$+ 2 e^-$	$\text{Cl}^- + \text{H}_2\text{O}$	+ 1,50
	$2\text{HClO}_2 + 6 \text{H}^+$	$+ 6 e^-$	$\text{Cl}_2 \uparrow + 4 \text{H}_2\text{O}$	+ 1,63
1	2	3	4	5
	$\text{HClO}_2 + 3 \text{H}^+$	$+ 4 e^-$	$\text{Cl}^- + 4 \text{H}_2\text{O}$	+ 1,63
	$\text{ClO}_2^- + 2 \text{H}_2\text{O}$	$+ 4 e^-$	$\text{Cl}^- + 4 \text{OH}^-$	+ 0,77
	$\text{ClO}_3^- + 6 \text{H}^+$	$+ 6 e^-$	$\text{Cl}^- + 6 \text{OH}^-$	+ 0,63
	$\text{ClO}_4^- + 2 \text{H}^+$	$+ 2 e^-$	$\text{ClO}_3^- + \text{H}_2\text{O}$	+ 1,19
	$2\text{ClO}_4^- + 16 \text{H}^+$	$+ 14 e^-$	$\text{Cl}_2 \uparrow + 8 \text{H}_2\text{O}$	+ 1,39

	$\text{ClO}_4^- + 4 \text{H}_2\text{O}$	$+ 8 e^-$	$\text{Cl}^- + 8 \text{OH}^-$	+ 0,56
Co	Co^{3+}	$+ e^-$	Co^{2+}	+ 1,95
	Co^{3+}	$+ 3 e^-$	$\text{Co} \downarrow$	+ 0,46
	Co^{2+}	$+ 2 e^-$	$\text{Co} \downarrow$	- 0,29
Cr	Cr^{3+}	$+ e^-$	Cr^{2+}	- 0,41
	Cr^{3+}	$+ 3 e^-$	$\text{Cr} \downarrow$	- 0,74
	Cr^{2+}	$+ 2 e^-$	$\text{Cr} \downarrow$	- 0,91
	$\text{CrO}_2^- + 2 \text{H}_2\text{O}$	$+ 3 e^-$	$\text{Cr} \downarrow + 4 \text{OH}^-$	- 1,2
	$\text{Cr}_2\text{O}_7^{2-} + 14 \text{H}^+$	$+ 6 e^-$	$2 \text{Cr}^{3+} + 7 \text{H}_2\text{O}$	+ 1,33
	$\text{CrO}_4^{2-} + 4 \text{H}_2\text{O}$	$+ 3 e^-$	$\text{Cr}(\text{OH})_3 \downarrow + 5 \text{OH}^-$	- 0,13
Cs	Cs^+	$+ e^-$	$\text{Cs} \downarrow$	- 2,923
Cu	Cu^{2+}	$+ e^-$	Cu^+	+ 0,159
	Cu^{2+}	$+ 2 e^-$	$\text{Cu} \downarrow$	+ 0,345
	Cu^+	$+ e^-$	$\text{Cu} \downarrow$	+ 0,531
	$\text{Cu}^{2+} + \text{Cl}^-$	$+ e^-$	$\text{CuCl} \downarrow$	+ 0,54
	$\text{Cu}^{2+} + \text{Br}^-$	$+ e^-$	$\text{CuBr} \downarrow$	+ 0,64
	$\text{Cu}^{2+} + \text{I}^-$	$+ e^-$	$\text{CuI} \downarrow$	+ 0,86
F	$\text{F}_2 \uparrow$	$+ 2 e^-$	2F^-	+ 2,77
Fe	Fe^{3+}	$+ e^-$	Fe^{2+}	+ 0,771
	Fe^{3+}	$+ 3 e^-$	$\text{Fe} \downarrow$	- 0,058
	Fe^{2+}	$+ 2 e^-$	$\text{Fe} \downarrow$	- 0,473
1	2	3	4	5
	$\text{FeO}_4^{2-} + 8 \text{H}^+$	$+ 3 e^-$	$\text{Fe}^{3+} + 4 \text{H}_2\text{O}$	+ 1,9
Ga	Ga^{3+}	$+ 3 e^-$	$\text{Ga} \downarrow$	- 0,56
Ge	Ge^{2+}	$+ 2 e^-$	$\text{Ge} \downarrow$	0,0
H	2H^+	$+ 2 e^-$	$\text{H}_2 \uparrow$	0,0000
	$2\text{H}_2\text{O}$	$+ 2 e^-$	$\text{H}_2 \uparrow + 2 \text{OH}^-$	- 0,828
Hf	$\text{HfO}^{2+} + 2 \text{H}^+$	$+ 4 e^-$	$\text{Hf} \downarrow + \text{H}_2\text{O}$	- 1,70
Hg	2Hg^{2+}	$+ 2 e^-$	Hg_2^{2+}	+ 0,907
	Hg^{2+}	$+ 2 e^-$	$\text{Hg} \downarrow$	+ 0,850
	2Hg^{2+}	$+ 2 e^-$	2Hg	+ 0,792
I	I_2	$+ 2 e^-$	2I^-	+ 0,621

	I_3^-	$+ 2 e^-$	$3I^-$	+0,546
	$2I\text{Br}$	$+ 2 e^-$	$I_2 \downarrow + 2\text{Br}^-$	+ 1,02
	2ICl	$+ 2 e^-$	$I_2 \downarrow + 2 \text{Cl}^-$	+ 1,19
	$2\text{HIO} + 2 \text{H}^+$	$+ 2 e^-$	$I_2 \downarrow + 2 \text{H}_2\text{O}$	+ 1,45
	$2\text{IO}^- + 2 \text{H}_2\text{O}$	$+ 2 e^-$	$I_2 \downarrow + 4 \text{OH}^-$	+ 0,45
	$\text{HIO} + \text{H}^+$	$+ 2 e^-$	$I^- + \text{H}_2\text{O}$	+ 0,99
	$\text{IO}^- + \text{H}_2\text{O}$	$+ 2 e^-$	$I^- + 2 \text{OH}^-$	+ 0,49
	$\text{IO}_3^- + 5 \text{H}^+$	$+ 4 e^-$	$\text{HIO} + 2 \text{H}_2\text{O}$	+ 1,14
	$\text{IO}_3^- + 2 \text{H}_2\text{O}$	$+ 4 e^-$	$\text{IO}^- + 4 \text{OH}^-$	+ 0,14
	$2\text{IO}_3^- + 12 \text{H}^+$	$+ 10 e^-$	$I_2 \downarrow + 6 \text{H}_2\text{O}$	+ 1,19
	$2\text{IO}_3^- + 6 \text{H}_2\text{O}$	$+ 10 e^-$	$I_2 \downarrow + 12 \text{OH}^-$	+ 0,21
	$\text{IO}_3^- + 6 \text{H}^+$	$+ 6 e^-$	$I^- + 3 \text{H}_2\text{O}$	+ 1,08
	$\text{IO}_3^- + 3 \text{H}_2\text{O}$	$+ 6 e^-$	$I^- + 6 \text{OH}^-$	+ 0,26
	$\text{H}_5\text{IO}_6 + \text{H}^+$	$+ 2 e^-$	$\text{IO}_3^- + 3 \text{H}_2\text{O}$	+ 1,6
	$\text{H}_3\text{IO}_6^{2-}$	$+ 2 e^-$	$\text{IO}_3^- + 3 \text{OH}^-$	+ 0,7
	$\text{H}_5\text{IO}_6 + 7\text{H}^+$	$+ 8 e^-$	$I^- + 6 \text{H}_2\text{O}$	+ 1,24
1	2	3	4	5
	$\text{H}_3\text{JO}_6^{2-} + 3 \text{H}_2\text{O}$	$+ 8 e^-$	$I^- + 9 \text{OH}^-$	- 0,37
In	In^{3+}	$+ 2 e^-$	In^+	- 0,45
	In^{3+}	$+ 3 e^-$	$\text{In} \downarrow$	- 0,34
Ir	Ir^{3+}	$+ 3 e^-$	$\text{Ir} \downarrow$	+ 1,15
K	K^+	$+ e^-$	$\text{K} \downarrow$	- 2,923
Li	Li^+	$+ e^-$	$\text{Li} \downarrow$	- 3,04
Mg	Mg^{2+}	$+ 2 e^-$	$\text{Mg} \downarrow$	- 2,37
Mn	Mn^{3+}	$+ e^-$	Mn^{2+}	+ 1,51
	Mn^{2+}	$+ 2 e^-$	$\text{Mn} \downarrow$	- 1,17
	$\text{MnO}_4^{2-} + 4 \text{H}^+$	$+ 2 e^-$	$\text{MnO}_2 \downarrow + 2 \text{H}_2\text{O}$	+ 2,26
	$\text{MnO}_4^{2-} + 2 \text{H}_2\text{O}$	$+ 2 e^-$	$\text{MnO}_2 \downarrow + 4 \text{OH}^-$	+ 0,6
	MnO_4^-	$+ e^-$	MnO_4^{2-}	+ 0,588
	$\text{MnO}_4^- + 4 \text{H}^+$	$+ 3 e^-$	$\text{MnO}_2 \downarrow + 2 \text{H}_2\text{O}$	+ 1,69
	$\text{MnO}_4^- + 8 \text{H}^+$	$+ 5 e^-$	$\text{Mn}^{2+} + 4 \text{H}_2\text{O}$	= 1,51
Mo	Mo^{3+}	$+ 3 e^-$	$\text{Mo} \downarrow$	- 0,2

	$\text{MoO}_2^+ + 4 \text{H}^+$	$+ 2 e^-$	$\text{Mo}^{3+} + 2 \text{H}_2\text{O}$	0,0
	$\text{H}_2\text{MoO}_4 + 6 \text{H}^+$	$+ 6 e^-$	$\text{Mo} \downarrow + 4 \text{H}_2\text{O}$	+ 0,48
	$\text{MoO}_4^{2-} + 4 \text{H}_2\text{O}$	$+ 6 e^-$	$\text{Mo} \downarrow + 8 \text{OH}^-$	-1,05
N	$\text{NH}_3 + 11 \text{H}^+$	$+ 8 e^-$	3NH_4^+	+ 0,69
	$\text{N}_3^- + 7 \text{H}_2\text{O}$	$+ 6 e^-$	$\text{N}_2\text{H}_4 + \text{NH}_3 + 7 \text{OH}^-$	- 0,62
	$3\text{N}_2 \uparrow + 2 \text{H}^+$	$+ 2 e^-$	2NH_3	- 3,1
	$3\text{N}_2 \uparrow$	$+ 2 e^-$	2N_3^-	- 3,4
	$\text{N}_2 \uparrow + 8 \text{H}^+$	$+ 6 e^-$	2NH_4^+	+ 0,26
	$\text{H}_2\text{N}_2\text{O}_2 + 2\text{H}^+$	$+ 2 e^-$	$\text{N}_2 \uparrow + 2 \text{H}_2\text{O}$	+ 2,65
	$2\text{HNO}_2 + 4\text{H}^+$	$+ 4 e^-$	$\text{H}_2\text{N}_2\text{O}_2 + 2\text{H}_2\text{O}$	+ 0,85
	$\text{HNO}_2 + \text{H}^+$	$+ e^-$	$\text{NO} \uparrow + \text{H}_2\text{O}$	+ 0,98
	1	2	3	4
	$\text{NO}_2^- + \text{H}_2\text{O}$	$+ e^-$	$\text{NO} \uparrow + 2 \text{OH}^-$	- 0,46
	$2 \text{HNO}_2 + 4 \text{H}^+$	$+ 4 e^-$	$\text{N}_2\text{O} \uparrow + 3 \text{H}_2\text{O}$	+ 1,29
	$2 \text{HNO}_2 + 6 \text{H}^+$	$+ 6 e^-$	$\text{N}_2 \uparrow + 4 \text{H}_2\text{O}$	+ 1,44
	$2 \text{NO}_2^- + 4 \text{H}_2\text{O}$	$+ 6 e^-$	$\text{N}_2 \uparrow + 8 \text{OH}^-$	+ 0,41
	$\text{HNO}_2 + 7 \text{H}^+$	$+ 6 e^-$	$\text{NH}_4^+ + 2 \text{H}_2\text{O}$	+ 0,86
	$\text{NO}_2^- + 6 \text{H}_2\text{O}$	$+ 6 e^-$	$\text{NH}_4\text{OH} + 7 \text{OH}^-$	- 0,15
	$\text{N}_2\text{O} \uparrow + 2\text{H}^+$	$+ 2 e^-$	$\text{N}_2 \uparrow + \text{H}_2\text{O}$	+ 1,77
	$\text{N}_2\text{O} \uparrow + \text{H}_2\text{O}$	$+ 2 e^-$	$\text{N}_2 \uparrow + 2\text{OH}^-$	+ 0,94
	$2\text{NO} \uparrow + 4\text{H}^+$	$+ 4 e^-$	$\text{N}_2 \uparrow + 2\text{H}_2\text{O}$	+ 1,68
	$2\text{NO} \uparrow + 2\text{H}_2\text{O}$	$+ 4 e^-$	$\text{N}_2 \uparrow + 4\text{OH}^-$	+ 0,85
	$\text{N}_2\text{O}_4 \uparrow + 2\text{H}^+$	$+ 2 e^-$	2HNO_2	+ 1,07
	$\text{N}_2\text{O}_4 \uparrow$	$+ 2 e^-$	2NO_2^-	+ 0,88
	$\text{N}_2 \uparrow + 2 \text{H}_2\text{O}$	$+ 8 e^-$	$\text{N}_2 \uparrow + 4\text{H}_2\text{O}$	+ 1,35
	$\text{N}_2\text{O}_4 \uparrow + 4\text{H}_2\text{O}$	$+ 8 e^-$	$\text{N}_2 \uparrow + 8\text{OH}^-$	+ 0,53
	$\text{NO}_3^- + 3 \text{H}^+$	$+ 2 e^-$	$\text{HNO}_2 + \text{H}_2\text{O}$	+ 0,94
	$\text{NO}_3^- + \text{H}_2\text{O}$	$+ 2 e^-$	$\text{NO}_2^- + 2 \text{OH}^-$	+ 0,01
	$\text{NO}_3^- + 2 \text{H}^+$	$+ e^-$	$\text{NO}_2 \uparrow + \text{H}_2\text{O}$	+ 0,80
	$\text{NO}_3^- + \text{H}_2\text{O}$	$+ e^-$	$\text{NO}_2 \uparrow + 2\text{OH}^-$	- 0,86
	$\text{NO}_3^- + 4\text{H}^+$	$+ 3 e^-$	$\text{NO} \uparrow + 2\text{H}_2\text{O}$	+ 0,96
	$\text{NO}_3^- + 2\text{H}_2\text{O}$	$+ 3 e^-$	$\text{NO} \uparrow + 4\text{OH}^-$	- 0,14

	$\text{NO}_3^- + 12\text{H}^+$	$+10 e^-$	$\text{N}_2\uparrow + 6\text{H}_2\text{O}$	+ 0,73
	$\text{NO}_3^- + 7\text{H}_2\text{O}$	$+ 8 e^-$	$\text{NH}_4\text{OH} + 9 \text{OH}^-$	- 0,12
	$\text{NO}_3^- + 10\text{H}^+$	$+ 8 e^-$	$\text{NH}_4^+ + 3 \text{H}_2\text{O}$	+ 0,87
Na	Na^+	$+ e^-$	$\text{Na} \downarrow$	- 2,713
Ni	Ni^{2+}	$+ 2 e^-$	$\text{Ni} \downarrow$	- 0,228
Rb	Rb^+	$+ e^-$	$\text{Rb} \downarrow$	- 2,924
O	$\text{O}_2\uparrow + 4\text{H}^+$	$+ 4 e^-$	$2\text{H}_2\text{O}$	+ 1,229
	$\text{O}_2\uparrow + 2\text{H}_2\text{O}$	$+ 4 e^-$	4OH^-	+ 0,401
	$\text{O}_2\uparrow + 2\text{H}^+$	$+ 2 e^-$	H_2O_2	+ 0,682
	$\text{O}_2\uparrow + \text{H}_2\text{O}$	$+ 2 e^-$	$\text{HO}_2^- + \text{OH}^-$	- 0,076
	$\text{H}_2\text{O}_2 + 2\text{H}^+$	$+ 2 e^-$	$2\text{H}_2\text{O}$	+ 1,77
	$\text{HO}_2^- + \text{H}_2\text{O}$	$+ 2 e^-$	3OH^-	+ 0,88
	$\text{O}_3\uparrow + 2\text{H}^+$	$+ 2 e^-$	$\text{O}_2\uparrow + \text{H}_2\text{O}$	+ 2,07
	$\text{O}_3\uparrow + \text{H}_2\text{O}$	$+ 2 e^-$	$\text{O}_2\uparrow + 2 \text{OH}^-$	+ 1,24
P	$\text{H}_3\text{PO}_2 + \text{H}^+$	$+ e^-$	$\text{P} \downarrow + 2 \text{H}_2\text{O}$	- 0,51
	$\text{H}_3\text{PO}_3 + 3 \text{H}^+$	$+ 3 e^-$	$\text{P} \downarrow + 3 \text{H}_2\text{O}$	- 0,50
	$\text{H}_3\text{PO}_4 + 5 \text{H}^+$	$+ 5 e^-$	$\text{P} \downarrow + 4 \text{H}_2\text{O}$	- 0,41
	$\text{H}_3\text{PO}_4 + 2 \text{H}^+$	$+ 2 e^-$	$\text{H}_3\text{PO}_3 + \text{H}_2\text{O}$	- 0,276
	$\text{PO}_4^{3-} + 2 \text{H}_2\text{O}$	$+ 2 e^-$	$\text{HPO}_3^{2-} + 3 \text{OH}^-$	- 1,12
Pb	Pb^{4+}	$+ 2 e^-$	Pb^{2+}	+ 1,66
	Pb^{4+}	$+ 4 e^-$	$\text{Pb} \downarrow$	+ 0,77
	Pb^{2+}	$+ 2 e^-$	$\text{Pb} \downarrow$	- 1,26
	$\text{HPbO}_2^- + \text{H}_2\text{O}$	$+ 2 e^-$	$\text{Pb} \downarrow + 3 \text{OH}^-$	- 0,54
	$\text{PbO}_3^{2-} + \text{H}_2\text{O}$	$+ 2 e^-$	$\text{PbO}_2^{2-} + 2 \text{OH}^-$	+ 0,2
Pt	Pt^{2+}	$+ 2 e^-$	$\text{Pt} \downarrow$	+ 1,2
S	$\text{S}_2\text{O}_3^{2-} + 6\text{H}^+$	$+ 4 e^-$	$2 \text{S} \downarrow + 3 \text{H}_2\text{O}$	+ 0,5
	$2\text{H}_2\text{SO}_3 + 2\text{H}^+$	$+ 4 e^-$	$\text{S}_2\text{O}_3^{2-} + 3 \text{H}_2\text{O}$	+ 0,40
	$\text{SO}_3^{2-} + 3 \text{H}_2\text{O}$	$+ 4 e^-$	$\text{S}_2\text{O}_3^{2-} + 6 \text{OH}^-$	- 0,58
	$2\text{H}_2\text{SO}_3 + \text{H}^+$	$+ 2 e^-$	$\text{HS}_2\text{O}_4^- + 2 \text{H}_2\text{O}$	- 0,08
	$2\text{SO}_3^{2-} + 2 \text{H}_2\text{O}$	$+ 2 e^-$	$\text{S}_2\text{O}_4^{2-} + 4 \text{OH}^-$	- 1,12
	$\text{SO}_4^{2-} + 4\text{H}^+$	$+ 2 e^-$	$\text{H}_2\text{SO}_3 + \text{H}_2\text{O}$	+ 0,17
	$\text{SO}_4^{2-} + \text{H}_2\text{O}$	$+ 2 e^-$	$\text{SO}_3^{2-} + 2 \text{OH}^-$	- 0,93

Solubility product constants

Сполука	ДР	Сполука	ДР	Сполука	ДР
AgBr	$5,3 \cdot 10^{-13}$	CsMnO ₄	$9,1 \cdot 10^{-5}$	Ni(OH) ₂	$2,0 \cdot 10^{-15}$
AgCl	$1,8 \cdot 10^{-10}$	Cs ₂ SiF ₆	$1,26 \cdot 10^{-5}$	Ni ₂ P ₂ O ₇	$1,7 \cdot 10^{-13}$
AgI	$8,3 \cdot 10^{-17}$	CuBr	$5,25 \cdot 10^{-9}$	NiS	$3,2 \cdot 10^{-19}$
Ag ₂ CrO ₄	$1,1 \cdot 10^{-12}$	CuI	$1,1 \cdot 10^{-12}$	PbCO ₃	$7,5 \cdot 10^{-14}$
Ag ₃ PO ₄	$1,3 \cdot 10^{-20}$	Cu ₂ (OH) ₂ CO ₃	$1,7 \cdot 10^{-34}$	PbC ₂ O ₄	$4,8 \cdot 10^{-19}$
Ag ₂ SO ₄	$1,6 \cdot 10^{-5}$	Cu(OH) ₂ (Cu ²⁺ , 2OH ⁻)	$2,2 \cdot 10^{-20}$	PbCrO ₄	$1,8 \cdot 10^{-14}$
Ag ₂ CO ₃	$8,2 \cdot 10^{-12}$	Cu ₃ (PO ₄) ₂	$1,7 \cdot 10^{-25}$	PbCl ₂	$1,6 \cdot 10^{-5}$
Al(OH) ₃ (Al ³⁺ , 3OH ⁻)	$1,0 \cdot 10^{-32}$	Cu ₂ P ₂ O ₇	$8,3 \cdot 10^{-16}$	PbI ₂	$1,1 \cdot 10^{-9}$
AlPO ₄	$5,8 \cdot 10^{-19}$	CuS	$6,3 \cdot 10^{-36}$	Pb ₃ (PO ₄) ₂	$7,9 \cdot 10^{-43}$
BaC ₂ O ₄	$1,1 \cdot 10^{-7}$	FeCO ₃	$3,5 \cdot 10^{-11}$	Pb(OH) ₂ (Pb ²⁺ , 2OH ⁻)	$5,0 \cdot 10^{-16}$
BaCO ₃	$5,1 \cdot 10^{-9}$	FeC ₂ O ₄	$2,0 \cdot 10^{-7}$	PbS	$2,5 \cdot 10^{-27}$
BaSO ₄	$1,1 \cdot 10^{-10}$	Fe(OH) ₂ (Fe ²⁺ , 2OH ⁻)	$1,0 \cdot 10^{-15}$	PbSO ₄	$1,6 \cdot 10^{-8}$
Ba ₃ (PO ₄) ₂	$6,0 \cdot 10^{-39}$	Fe(OH) ₃ (Fe ³⁺ , 3OH ⁻)	$3,2 \cdot 10^{-38}$	Pb(OH) ₄	$6,5 \cdot 10^{-71}$
Ba(OH) ₂	$5,0 \cdot 10^{-3}$	FePO ₄	$1,3 \cdot 10^{-22}$	Sn(OH) ₂ (Sn ²⁺ , 2OH ⁻)	$6,3 \cdot 10^{-23}$
Bi(OH) ₃	$3,2 \cdot 10^{-32}$	FeS	$5,0 \cdot 10^{-18}$	Sn(OH) ₄	$1 \cdot 10^{-57}$
Bi ₂ S ₃	$1,0 \cdot 10^{-97}$	HgS	$1,6 \cdot 10^{-52}$	SrCO ₃	$1,1 \cdot 10^{-10}$
CaCO ₃	$4,8 \cdot 10^{-9}$	Hg ₂ CO ₃ (Hg ₂ ²⁺ , CO ₃ ²⁻)	$8,9 \cdot 10^{-17}$	SrC ₂ O ₄	$5,6 \cdot 10^{-8}$
CaC ₂ O ₄	$2,3 \cdot 10^{-9}$	Li ₃ PO ₄	$3,2 \cdot 10^{-9}$	SrCrO ₄	$3,6 \cdot 10^{-5}$
CaCrO ₄	$2,3 \cdot 10^{-9}$	Li ₂ CO ₃	$4 \cdot 10^{-3}$	Sr(OH) ₂	$3,2 \cdot 10^{-4}$
Ca(OH) ₂ (Ca ²⁺ , 2OH ⁻)	$5,5 \cdot 10^{-6}$	MgCO ₃	$2,1 \cdot 10^{-5}$	Sr ₃ (PO ₄) ₂	$1,1 \cdot 10^{-31}$
Ca ₃ (PO ₄) ₂	$2,0 \cdot 10^{-29}$	MgC ₂ O ₄	$8,6 \cdot 10^{-5}$	SrMoO ₄	$2 \cdot 10^{-7}$
CaSO ₃	$1,3 \cdot 10^{-8}$	MgNH ₄ PO ₄	$2,5 \cdot 10^{-13}$	SrSO ₄	$3,2 \cdot 10^{-7}$
CaSO ₄	$9,1 \cdot 10^{-6}$	Mg(OH) ₂ (Mg ²⁺ , 2OH ⁻)	$6,0 \cdot 10^{-10}$	SrSO ₃	$4 \cdot 10^{-8}$
CdCO ₃	$5,2 \cdot 10^{-12}$	Mg ₃ (PO ₄) ₂	$1,0 \cdot 10^{-14}$	Tl ₂ SO ₄	$4 \cdot 10^{-3}$
Cd(OH) ₂	$2,0 \cdot 10^{-19}$	MgSO ₃	$3 \cdot 10^{-3}$	UO ₂ NH ₄ PO ₄	$4,4 \cdot 10^{-27}$
CoCO ₃	$1,4 \cdot 10^{-13}$	MnC ₂ O ₄	$5,0 \cdot 10^{-6}$	U(OH) ₄	$1 \cdot 10^{-45}$
Co(OH) ₂	$2,0 \cdot 10^{-15}$	MgNH ₄ PO ₄	$1 \cdot 10^{-12}$	ZnCO ₃	$1,45 \cdot 10^{-11}$
Co(OH) ₃	$4,0 \cdot 10^{-45}$	Mn(OH) ₂ (Mg ²⁺ , 2OH ⁻)	$1,9 \cdot 10^{-13}$	SrC ₂ O ₄	$1,5 \cdot 10^{-9}$
Cr(OH) ₃ (Cr ³⁺ , 3OH ⁻)	$6,3 \cdot 10^{-31}$	Mo(OH) ₄	$1 \cdot 10^{-50}$	Zn(OH) ₂	$1,2 \cdot 10^{-17}$
CrPO ₄	$2,4 \cdot 10^{-23}$	NiCO ₃	$1,3 \cdot 10^{-7}$	Zn ₃ (PO ₄) ₂	$9,1 \cdot 10^{-33}$

Appendix 7. Constants of ion. complex ions

Комплексо- утворювач	Дисоціація комплексного іона	Константа стійкості	ρ β_c
Ag⁺	$[\text{Ag}(\text{NH}_3)_2]^+ \Leftrightarrow \text{Ag}^+ + 2\text{NH}_3$	$1,1 \cdot 10^7$	7,03
	$[\text{Ag}(\text{CN})_2]^- \Leftrightarrow \text{Ag}^+ + 2\text{CN}^-$	$7,2 \cdot 10^{21}$	21,1
	$[\text{Ag}(\text{S}_2\text{O}_3)_2]^{3-} \Leftrightarrow \text{Ag}^+ + 2\text{S}_2\text{O}_3^{2-}$	$4,0 \cdot 10^{13}$	13,6
	$[\text{Ag}(\text{SCN})_2]^- \Leftrightarrow \text{Ag}^+ + 2\text{SCN}^-$	$3,7 \cdot 10^8$	7,57
	$[\text{AgEDTA}]^{3-} \Leftrightarrow \text{Ag}^+ + \text{EDTA}^{4-}$	$2,1 \cdot 10^7$	7,32
Al³⁺	$[\text{AlF}_6]^{3-} \Leftrightarrow \text{Al}^{3+} + 6\text{F}^-$	$6,9 \cdot 10^{19}$	19,84
	$[\text{Al}(\text{C}_2\text{O}_4)_3]^{3-} \Leftrightarrow \text{Al}^{3+} + 3\text{C}_2\text{O}_4^{2-}$	$6,2 \cdot 10^{16}$	16,8
	$[\text{AlEDTA}]^- \Leftrightarrow \text{Al}^{3+} + \text{EDTA}^{4-}$	$1,4 \cdot 10^{16}$	16,13
Cd²⁺	$[\text{Cd}(\text{NH}_3)_4]^{2+} \Leftrightarrow \text{Cd}^{2+} + 4\text{NH}_3$	$1,3 \cdot 10^7$	7,12
	$[\text{Cd}(\text{CN})_4]^{2-} \Leftrightarrow \text{Cd}^{2+} + 4\text{CN}^-$	$7,1 \cdot 10^{18}$	18,85
	$[\text{CdEDTA}]^{2-} \Leftrightarrow \text{Cd}^{2+} + \text{EDTA}^{4-}$	$3,1 \cdot 10^{16}$	16,48
Co²⁺	$[\text{Co}(\text{NH}_3)_4]^{2+} \Leftrightarrow \text{Co}^{2+} + 4\text{NH}_3$	$3,6 \cdot 10^5$	5,55
	$[\text{Co}(\text{NH}_3)_6]^{2+} \Leftrightarrow \text{Co}^{2+} + 6\text{NH}_3$	$1,3 \cdot 10^5$	5,11
	$[\text{Co}(\text{CN})_6]^{4-} \Leftrightarrow \text{Co}^{2+} + 6\text{CN}^-$	$1,3 \cdot 10^{17}$	17,11
	$[\text{Co}(\text{SCN})_4]^{2-} \Leftrightarrow \text{Co}^{2+} + 4\text{SCN}^-$	$6,0 \cdot 10^1$	1,78
	$[\text{CoEDTA}]^{2-} \Leftrightarrow \text{Co}^{2+} + \text{EDTA}^{4-}$	$1,3 \cdot 10^{16}$	16,1
Co³⁺	$[\text{Co}(\text{NH}_3)_6]^{3+} \Leftrightarrow \text{Co}^{3+} + 6\text{NH}_3$	$3,2 \cdot 10^{32}$	32,51
	$[\text{Co}(\text{CN})_6]^{3-} \Leftrightarrow \text{Co}^{3+} + 6\text{CN}^-$	$1,0 \cdot 10^{64}$	64,0
	$[\text{CoEDTA}]^- \Leftrightarrow \text{Co}^{3+} + \text{EDTA}^{4-}$	$1,0 \cdot 10^{36}$	36,0
Cu²⁺	$[\text{Cu}(\text{NH}_3)_4]^{2+} \Leftrightarrow \text{Cu}^{2+} + 4\text{NH}_3$	$4,7 \cdot 10^{12}$	12,67
	$[\text{Cu}(\text{C}_2\text{O}_4)_2]^{2-} \Leftrightarrow \text{Cu}^{2+} + 2\text{C}_2\text{O}_4^{2-}$	$1,1 \cdot 10^8$	8,04
	$[\text{Cu}(\text{CN})_4]^{2-} \Leftrightarrow \text{Cu}^{2+} + 4\text{CN}^-$	$2,0 \cdot 10^{27}$	27,3
	$[\text{CuEDTA}]^{2-} \Leftrightarrow \text{Cu}^{2+} + \text{EDTA}^{4-}$	$6,3 \cdot 10^{18}$	18,8
Hg²⁺	$[\text{Hg}(\text{NH}_3)_4]^{2+} \Leftrightarrow \text{Hg}^{2+} + 4\text{NH}_3$	$1,9 \cdot 10^{19}$	19,28
	$[\text{HgI}_4]^{2-} \Leftrightarrow \text{Hg}^{2+} + 4\text{I}^-$	$6,7 \cdot 10^{29}$	29,83
	$[\text{Hg}(\text{CN})_4]^{2-} \Leftrightarrow \text{Hg}^{2+} + 4\text{CN}^-$	$2,5 \cdot 10^{41}$	41,4
	$[\text{HgEDTA}]^{2-} \Leftrightarrow \text{Hg}^{2+} + \text{EDTA}^{4-}$	$1,4 \cdot 10^{22}$	22,15
Fe²⁺	$[\text{Fe}(\text{CN})_6]^{4-} \Leftrightarrow \text{Fe}^{2+} + 6\text{CN}^-$	$1 \cdot 10^{24}$	24,0
	$[\text{FeEDTA}]^{2-} \Leftrightarrow \text{Fe}^{2+} + \text{EDTA}^{4-}$	$9,8 \cdot 10^{14}$	14,45
Fe³⁺	$[\text{FeF}_6]^{3-} \Leftrightarrow \text{Fe}^{3+} + 6\text{F}^-$	$1,3 \cdot 10^{16}$	16,1
	$[\text{Fe}(\text{C}_2\text{O}_4)_3]^{3-} \Leftrightarrow \text{Fe}^{3+} + 3\text{C}_2\text{O}_4^{2-}$	$1,6 \cdot 10^{20}$	20,2
	$[\text{Fe}(\text{CN})_6]^{3-} \Leftrightarrow \text{Fe}^{3+} + 6\text{CN}^-$	$1,0 \cdot 10^{31}$	31,0
	$[\text{FeEDTA}]^- \Leftrightarrow \text{Fe}^{3+} + \text{EDTA}^{4-}$	$1,3 \cdot 10^{25}$	25,1
Ni²⁺	$[\text{Ni}(\text{NH}_3)_6]^{2+} \Leftrightarrow \text{Ni}^{2+} + 6\text{NH}_3$	$5,4 \cdot 10^8$	8,76
	$[\text{Ni}(\text{CN})_4]^{2-} \Leftrightarrow \text{Ni}^{2+} + 4\text{CN}^-$	$5,4 \cdot 10^{13}$	13,75
	$[\text{NiEDTA}]^{2-} \Leftrightarrow \text{Ni}^{2+} + \text{EDTA}^{4-}$	$2,8 \cdot 10^{18}$	18,45
Pb²⁺	$[\text{PbEDTA}]^{2-} \Leftrightarrow \text{Pb}^{2+} + \text{EDTA}^{4-}$	$1,6 \cdot 10^{18}$	18,2
Zn²⁺	$[\text{Zn}(\text{NH}_3)_4]^{2+} \Leftrightarrow \text{Zn}^{2+} + 4\text{NH}_3$	$2,9 \cdot 10^9$	9,46
	$[\text{ZnEDTA}]^{2-} \Leftrightarrow \text{Zn}^{2+} + \text{EDTA}^{4-}$	$3,1 \cdot 10^{16}$	16,5
	$[\text{Zn}(\text{CN})_4]^{2-} \Leftrightarrow \text{Zn}^{2+} + 4\text{CN}^-$	$7,7 \cdot 10^{16}$	16,89