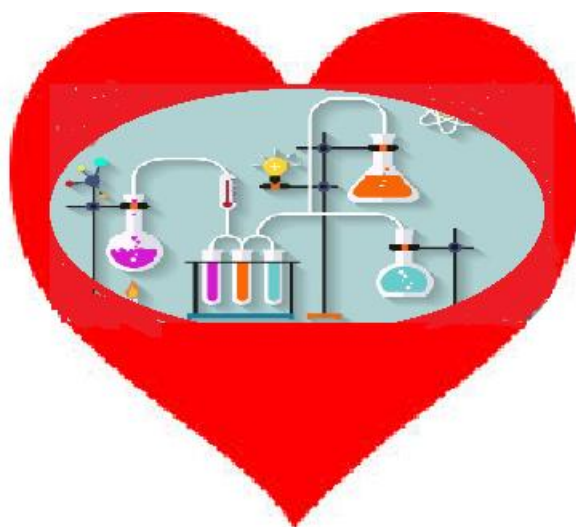


Ministry of Health of Ukraine
Poltava State Medical University



MEDICAL CHEMISTRY

Workbook

for English-speaking students who study at institutions
of higher medical education of Ministry of Health of Ukraine
in specialty 221 “Dentistry”

Student of group _____ *course* _____

Student's name _____

Poltava – 2021

*Recommended by Central Methodological Commission of Poltava State Medical University,
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SAFETY RULES FOR WORK IN EDUCATIONAL LABORATORIES OF MEDICAL CHEMISTRY:

1. Students must wear lab coats and medical caps to work in a chemistry laboratory.
2. Each student must keep his or her place at the desk orderly and clean. There should not be extra things on the table (bags etc.).
3. A student should begin to perform an experiment only after carefully reading the instruction and listening to the teacher's explanations. The work should be done neatly, exactly, without hurrying.
4. It is strictly forbidden to perform experiments that are not in the instructions for the subject.
5. It is strictly forbidden to take chemicals with bare hands or taste them. When determining a substance by its smell, the glass should be kept at a distance. Direct the air from the opening of the glass to the nose with the movement of your hand. To protect face and clothing when mixing reagents, do not lean over the opening of the reaction vessel.
6. Experiments should be performed only in clean glassware, with such quantity and concentration of reactants, and in the exact sequence and conditions, as indicated in the instruction.
7. Immediately cork or cap all glasses with reagents and solutions after use.
8. All experiments with concentrated acids and alkalis should be only performed under the extraction hood.
9. When diluting concentrated acids, especially sulfuric acid, you should carefully pour the acid into water, and not vice versa.
10. Do not lean over a boiling liquid because it can be thrown out of the test tube. The opening of the test tube should not be directed towards the person performing the experiment or towards those next to him.
11. It is strictly forbidden to switch on and off electrical devices without the permission of the teacher, as well as light gas burners without need.
12. At the end of the class, each student should wash the glassware he or she has worked with and clean the workplace. The person on duty should ensure that the entire laboratory is put in order.

Basic First Aid Rules:

1. In the case of an injury caused by glass shards, stop the bleeding with 3% solution of hydrogen peroxide, disinfect the wounds with an iodine solution and bandage.
2. In the case of chemical burns of hands or face, rinse off the reagent with plenty of water and then treat with either 2% solution of boric acid (if the burn is caused by an alkali), or 2% sodium hydrogen carbonate solution (if the burn is caused by an acid), and then rinse with water again.
3. In the case of burns with a hot object or hot liquid, treat the injury with a freshly prepared solution of potassium permanganate, and then apply a burn ointment.
4. In the case of chemical burns of eyes, rinse eyes with plenty of water and then seek medical advice.

When performing laboratory work students must observe order and the safety rules. Each student signs the safety register after familiarizing himself or herself with the safety rules.

Informed _____ " ____ " _____ 20 _____.
(student signature)

MODULE 1. "ACID-BASE EQUILIBRIUM AND FORMATION OF COMPLEXES IN BIOLOGICAL FLUIDS »

PRACTICAL CLASS № 1.

SUBJECT: Biogenic s-block elements: biological role, applications in medicine

IMPORTANCE: More than 70 chemical elements are found in the human body, and more than 40 elements are components of medicines, therefore study of these elements is very important.

AIM: to be able to explain properties and practically determine presence of biogenic s-block elements using qualitative reactions.

I. Theoretical part:

1. Biogenic elements: organogens; macroelements; microelements
2. Electronic structure of biogenic s-block elements.
3. Typical chemical properties of s-block elements and their compounds (reactions without changing degrees of oxidation).
4. Biological roles of biogenic s-block elements.
5. Relationship between the location of s-block elements in the Periodic System and their content in the organism.
6. Applications in medicine.

II. Practical part

1. Laboratory work

Qualitative reactions of biogenic s-block elements

Reactions of potassium cation (K⁺)

Experiment 1. *Interaction with sodium hexanitrocobaltate (III) Na₃[Co(NO₂)₆].*

Sodium hexanitrocobaltate (III) Na₃[Co(NO₂)₆] forms a yellow crystalline precipitate of potassium-sodium hexanitrocobaltate (III) K₂Na[Co(NO₂)₆] in reactions with salts of potassium:



The reaction must be performed in a neutral or slightly acidic medium.

Add two drops of a 0.5 N potassium salt solution to a test tube and then add an equal volume of a freshly prepared solution of sodium hexanitrocobaltate (III). A yellow crystalline precipitate forms.

Conclusions:

Reactions of sodium cation (Na⁺)

Experiment 2. *Interaction with potassium dihydrogen antimonate.*

Potassium dihydrogen antimonate forms a white crystalline precipitate of sodium dihydrogen antimonate in reactions with Na⁺:



Add 2-3 drops of a sodium salt solution to a test tube and then add 2-3 drops of potassium dihydrogen antimonate solution. If a precipitate is not obtained immediately, rub the inner walls of the test tube with a glass stick.

Conclusions:

Reactions of calcium cation (Ca²⁺)

Experiment 3. *Interaction with mineral acids*

Sulfuric acid forms white crystalline calcium (II) sulfate precipitate in reactions with Ca²⁺ ions upon heating and rubbing the inner walls of the test tube with a glass stick:

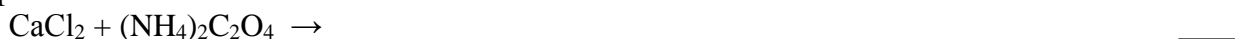


Add 3-5 drops of CaCl_2 solution into the test tube and add an equal volume of dilute H_2SO_4 . A white precipitate forms. When 2-3 drops of acetone or ethyl alcohol are added, precipitation is more complete, because the solubility of CaSO_4 in the new solvent decreases.

Conclusions:

Experiment 4. Interaction with ammonium oxalate

Ammonium oxalate $(\text{NH}_4)_2\text{C}_2\text{O}_4$ with Ca^{2+} ion forms a white crystalline calcium oxalate precipitate:



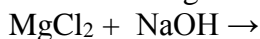
Add 3-5 drops of a CaCl_2 solution to a test tube, and add an equal volume of an ammonium oxalate solution. The precipitated CaC_2O_4 is readily soluble in mineral acids but not soluble in acetic acid.

Conclusions:

Reactions of magnesium cation Mg^{2+}

Experiment 5. Interaction with hydroxides

Hydroxides NaOH and KOH form a white amorphous magnesium hydroxide precipitate in reactions with cation Mg^{2+} :



Add four drops of a magnesium salt solution to a test tube, then add four drops of a saturated solution of sodium or potassium hydroxide NaOH or KOH .

Divide the obtained $\text{Mg}(\text{OH})_2$ precipitate into two test tubes. Add a solution of sulfuric acid dropwise to a test tube until complete dissolution of the precipitate:



Add a solution of ammonium chloride dropwise to the other test tube until the precipitate is dissolved:



Conclusions:

2. Control tasks

1. How to formulate the D.I.Mendeleev Periodic Law and why is it the basis for studying the properties of bioelements?

2. How do charges of nuclei, radii of atoms and ions, ionization energies within the periods and groups of the Periodic System of chemical elements change?

3. Write electronic configurations of sodium and magnesium atoms. Compare them. How does the electronic structure of these elements effect their properties?

4. Write electronic configurations of potassium and calcium atoms. Compare them. How does the electronic structure of these atoms effect their properties?

5. What are elements that are organogens, macroelements and microelements?

6. Biological roles of ions K^+ , Na^+ in the body.

7. Biological role of Mg^{2+} ion. _____

8. Compounds of *s*-block elements as medical preparations. Normal saline.

9. How to determine the presence of *s*-block elements in a sample by a change of the color of the flame? _____

15. Natural compounds of *s*-block elements, their role in the human body and applications in medicine.

Literature: P. 192-196

https://saylordotorg.github.io/text_general-chemistry-principles-patterns-and-applications-v1.0/s25-periodic-trends-and-the-s-bloc.html

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PRACTICAL CLASS № 2.

SUBJECT: Biogenic *p*-block and *d*-block elements: biological role, applications in medicine

IMPORTANCE: organogenic *p*-block elements carbon, nitrogen, oxygen, phosphorus, sulfur are the principal components of the main organic compounds of human organism: proteins, fats, carbohydrates, polynucleotides.

Biological roles of *d*-block elements are stipulated by their ability to participate in formation of coordination compounds, heterogeneous and protolytic reactions. Knowledge of atomic structure and properties of *d*-block elements and their compounds is necessary for medical students to master many sections of biochemistry, pharmacology, physiology and special disciplines. Understanding roles of their compounds in metabolic reactions is impossible without a study of properties of the elements and their simpler compounds.

AIM: to be able to explain properties and practically determine presence of biogenic *p*-block and *d*-block elements using qualitative reactions.

I. Theoretical part:

1. The electronic structure of *p*-block elements. Typical chemical properties of *p*-block elements and their compounds (reactions without changing the degree of oxidation). Biological role of biogenic *p*-block elements.

2. Relationship between the content of biogenic *p*-block elements in the human body and their content in the environment. Endemic diseases, their relationship to properties of biogeochemical provinces (areas with natural deficiency or excess of certain chemical elements in the lithosphere). Applications in medicine.

3. Problems of pollution of the biosphere and removal of toxic chemical compounds of technogenic origin from it.

4. Metals of life. Electron structure and electronegativity of biogenic *d*-block elements.

5. Typical chemical properties of *d*-block elements and their compounds:

a) reactions with changes of the degrees of oxidation;

b) formation of coordination compounds.

6. Biological roles of biogenic *d*-block elements. Toxicity of *d*-block elements and their compounds. Applications in medicine.

II. Practical part

1. Laboratory work

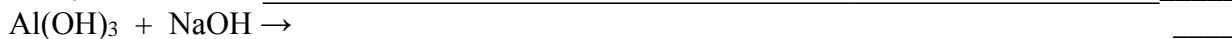
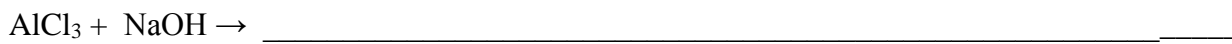
Qualitative reactions of biogenic *p*-block elements

Reactions of aluminum cation (Al^{3+})

Experiment 1. Interaction with alkalis

Alkalis form an amphoteric precipitate of aluminum hydroxide $\text{Al}(\text{OH})_3$ in a reaction with the

Al³⁺ cation. The precipitate dissolves in an excess of alkalis and acids, i.e. it exhibits amphoteric properties:



Add 2-3 drops of a solution of an aluminum salt to two test tubes and add dropwise a solution of sodium hydroxide until a white gelatinous precipitate is obtained. Add excess solution of sodium hydroxide to one test tube until the precipitate is dissolved. Add a solution of HCl to the other test tube until the precipitate is dissolved.

Conclusions: _____

Reactions of halide ions Cl⁻, Br⁻, I⁻

Experiment 2. Interaction of the silver (I) nitrate with halide ions

Silver nitrate forms a white precipitate in a reaction with **Cl⁻ ion**. The precipitate is not soluble in HNO₃, but it reacts with NH₄OH to form a silver ammonia complex compound.



The silver ammonia complex decomposes in nitric acid and silver chloride precipitates again



Add 2-3 drops of a NaCl solution to a test tube, then add 2-3 drops of an AgNO₃ solution. A white precipitate AgCl is obtained.

Silver nitrate forms a yellowish precipitate in a reaction with **Br⁻ ion**. The precipitate is not soluble in HNO₃, and is poorly soluble in a NH₄OH solution.



Silver nitrate forms a yellow precipitate in a reaction with **I⁻ ion**. The precipitate is not soluble in HNO₃ or NH₄OH solutions.



Add 2-3 drops of NaCl, NaBr and KI solutions to three test tubes. A white precipitate AgCl, yellowish precipitate AgBr, yellow precipitate AgI are formed.

Add several drops of NH₄OH solution to the resulting AgCl until it dissolves. Then add several drops of a HNO₃ solution to the obtained solution. AgCl precipitates again.

Conclusions: _____

Qualitative reactions of biogenic d-block elements

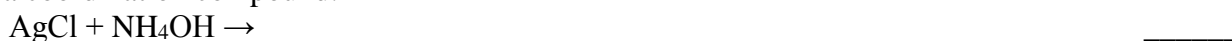
Reactions of silver cation (Ag⁺)

Experiment 3. Interaction with hydrochloric acid

Hydrochloric acid and solutions of its salts form a white precipitate of silver chloride in reactions with Ag⁺ cation:



Silver chloride is not soluble in acids and alkalis. It readily reacts with a solution of ammonia to form a coordination compound:



The complex [Ag(NH₃)₂]Cl is easily decomposed in reactions with acids.

Add 1-2 drops of a solution of a salt of silver to a test tube. Then add 1-2 drops of hydrochloric acid until a white precipitate is obtained. Add an ammonia solution dropwise to the precipitate until the it dissolves.

Conclusions: _____

If the nitric acid is added to the obtained solution, a precipitate of silver chloride is formed again.



Experiment 4. Interaction with alkalis

Alkalis form a black precipitate of silver oxide in reactions with Ag^+ cation that is insoluble in an excess of alkali.



Add 1-2 drops of a solution of a silver salt and 1-2 drops of alkali to a test tube. A black precipitate of silver oxide is obtained. The precipitate can be dissolved in ammonia, nitric acid and acetic acid.

Conclusions: _____

Experiment 5. Interaction with potassium iodide.

Potassium iodide forms a yellow precipitate of silver iodide in a reaction with Ag^+ cation:



Add 2-3 drops of a solution of a silver salt and 1-2 drops of potassium iodide to a test tube. A yellow precipitate of silver iodide is formed. The precipitate does not dissolve in ammonia.

Conclusions: _____

Experiment 6. Interaction with potassium chromate

Potassium chromate forms brick-red precipitate of silver chromate in reactions with Ag^+ cation. The precipitate can be dissolved in nitric acid and ammonia solutions.

Add 2-3 drops of a solution of a silver salt, 3-4 drops of distilled water and 1-2 drops of a potassium chromate solution to a test tube. A red precipitate of silver chromate is formed. The precipitate does not dissolve in acetic acid.



Conclusions: _____

Reactions of ferrous cation (Fe^{2+})

Experiment 7. Interaction with alkalis

Alkalis form a greenish precipitate of ferrous hydroxide in reactions with Ag^+ cation. In the presence of oxygen, the precipitate oxidizes to brownish ferric hydroxide.



Add 2-3 drops of a solution of Fe^{2+} salt and 2-3 drops of an alkali solution to a test tube. A pale green precipitate forms. It does not dissolve in alkalis but dissolves in mineral acids.

Conclusions: _____

Reactions of ferric cation (Fe^{3+})

Experiment 8. Interaction with alkalis

Alkalis and ammonia form a reddish-brown precipitate of ferric hydroxide in reactions with Fe^{3+} cation. The precipitate readily dissolves in acids and does not dissolve in excess of alkalis.



Add 3-4 drops of a FeCl_3 solution and 3-4 drops of an alkali solution to a test tube. A reddish-brown precipitate forms. Add hydrochloric acid to dissolve the precipitate.

Conclusions: _____

Experiment 9. Interaction with potassium rhodanide

Potassium rhodanide or ammonium rhodanide (NH_4CNS) forms a solution of a blood-red complex salt in reactions with Fe^{3+} cation.



Add 3-4 drops of a FeCl_3 solution, 1-2 drops of hydrochloric acid and 5-6 drops of a potassium rhodanide or ammonium rhodanide solution to a test tube.

Conclusions: _____

2. Control tasks:

1. What are the properties of the structure of atoms of *p*-block elements, possible degrees of oxidation, nature of their oxides and hydroxides? Amphoteric hydroxides.

2. Compounds of carbon and aluminum in human metabolism, applications of compounds of elements of IV(A) group in medicine.

3. Nitrogen and phosphorus as organogenic elements. Toxicity of compounds of As, Sb, Bi, Pb, Se, Te, Tl.

4. Biological roles of elements of the V(A) group, ozone. Sulfur and its compounds, their applications in medicine.

5. Biological role of halogens, applications of their salts as medicines.

6. Relationship between the content of biogenic *p*-block elements in the Periodic System and their content in the human body.

7. Concept of biogeochemical provinces, endemic diseases.

8. Various content of iodine in soils is recorded in different parts of Ukraine. What diseases are caused by deficiency and excess of this element?

9. Biological roles of carbon, nitrogen, phosphorus in the human body. Write down the electron structure of the elements.

10. Biological roles of oxygen, sulfur, fluorine in the human body. Write down the electron structure of the elements.

11. Biological roles of chlorine and iodine in the human body. Write down the electron structure of the elements.

12. What qualitative reactions are used to find presence of halogens? Write chemical equations.

13. Explain problems of pollution of the biosphere and removal of toxic chemical compounds of technogenic origin from it.

14. Biological roles of iron and copper in the human body. Write the electron structure of the elements.

15. Biological roles of zinc and manganese in the human body. Write the electron structure of the elements.

16. Biological roles of cobalt, chromium and nickel in the human body.

17. Toxicity of *d*-block elements and their compounds.

18. Write down the electron formula of the element with the number 30.

19. Write down the electron formula of the element with the number 26.

Literature: P. 196-215

https://saylordotorg.github.io/text_general-chemistry-principles-patterns-and-applications-v1.0/s26-the-p-block-elements.html

https://saylordotorg.github.io/text_general-chemistry-principles-patterns-and-applications-v1.0/s27-the-d-block-elements.html

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PRACTICAL CLASS № 3.

SUBJECT: Units of quantitative composition of solutions

IMPORTANCE OF THE SUBJECT: Blood plasma, saliva, gastric juice and other biological fluids in the organism are solutions. Processes of food assimilation and excretion are possible only in solutions. A great number of medical preparations are also solutions. The principal characteristic of solutions is their quantitative composition. When preparing, analyzing or using a drug, you need to know the amount of the active substance in the preparation, and also often have to convert the concentrations from one type of units to another. Therefore, a physician needs knowledge of units of quantitative composition of the solutions.

AIM: to master basic theoretical principles of the theory of solutions; types of units of solution concentration; solving calculation problems.

I. Theoretical part:

1. Composition of solutions.
2. Classification of solutions.
3. Units of quantitative composition of solutions:
 - 1) mass, volume and mole fraction;
 - 2) molarity (also deca-, centi-, milli- and micromolar);
 - 3) normality;
 - 4) molality;
 - 5) titer.

II. Practical part

Calculation tasks for the lesson

Task № 1

0.040 kg salt was dissolved in 0.200 L water. Calculate the mass fraction of the salt in the solution, if the density of water is 1 kg/L. _____

Task № 2

Calculate the mass of 10% (by mass) CuSO_4 solution and the mass of water, which are needed to prepare 0.5 kg 2% (by mass) CuSO_4 solution _____

Task № 3

Calculate the molarity of a sodium hydroxide solution with the mass fraction 0.2. The density of the solution is 1.29 kg/L. _____

Task № 4

Calculate the normality of a solution obtained by dissolving 0.0426 kg sodium sulfate in 0.3 kg water. The density of the solution is 1.12 kg/L. _____

Task № 5

A water solution obtained by dissolving 5 g glucose ($M_r = 180$ g/mol) in 95 g water is isotonic to blood plasma. Calculate the mass fraction, mole fraction, molarity and molality of glucose in the solution (density of the solution 1.018 g/cm³).

1. Control tasks:

1. Calculate the equivalency factors and equivalent molar masses for H_2SO_4 , $NaOH$, Na_2CO_3 , H_3PO_4 , H_3BO_3 , $Ca_3(PO_4)_2$, H_2SiO_3 , $Al_2(SO_4)_3$ in reactions that proceed to completion.

2. How many grams potassium chloride are needed to add to 450 g of a 8% KCl solution to obtain a 12% solution?

3. 400 g salt precipitated from 10 kg of a 20% solution at cooling. Calculate the mass percent of the resulting solution.

4. 5 L of a 2% HNO_3 solution (density 1.009 g/cm³) was added to 3 L of a 10% solution of the acid (density 1.054 g/cm³) Calculate the mass fraction and molarity of the solution (volume of the solution 8 L).

5. Now many moles of water are needed to be added to 1.6 kg of a 25% sodium hydroxide solution to obtain a 16% $NaOH$ solution?

6. One liter (dm³) of a solution contains 10.8 g of sulfuric acid. Calculate the molarity and normality of the solution.

7. One liter of a solution contains 192.6 g KNO_3 (density 1.14 g/cm^3). Calculate the molarity and mass fraction of the solution.

8. Calculate the mass fraction of sulfuric acid in a 5M solution of H_2SO_4 (density 1.29 g/cm^3).

9. 35 mL of a 1.025N HCl solution was used to dissolve a sample of chalk (calcium carbonate). Calculate the mass of calcium carbonate.

10. Calculate normality of a 18% NaOH solution (density 1.203 g/cm^3).

11. 20 g of iodine I_2 was dissolved in 500g tetrachloromethane CCl_4 . Calculate the mass fraction and mole fraction of iodine in the solution.

12. A solution of 5.0g toluene C_7H_8 in 225g benzene C_6H_6 has density 0.876 g/cm^3 . Calculate the mass fraction and molarity of the solution.

13. Calculate the volume of 0.2N HCl solution needed to neutralize 50 cm^3 of a solution that contains 0.40 g NaOH. Use the rule of equivalents for the calculation.

14. Calculate the volume of a 10% sodium hydrogencarbonate solution (density 1.105 g/cm^3) needed to prepare 5 L of a 2% solution (density 1.020 g/cm^3)?

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PRACTICAL CLASS № 4

SUBJECT: Preparation of solutions of required quantitative composition

IMPORTANCE OF THE SUBJECT: Solutions with molecular and ionic solutes are most important biological fluids. Aqueous solutions of electrolytes and low molecular weight substances provide constant osmotic pressure, acidity, buffering properties of bodily fluids, regulate values of membrane potentials, activity of enzymes, etc. The use of solutions with different content of the dissolved substance in the medical practice requires an ability to prepare solutions of a required concentration for practical use.

AIM: experimental preparation of a solution with a given concentration.

I. Theoretical part:

1. Solutions in life.
2. Enthalpy and entropy dissolution factors and their relationip with the mechanism of solubility.
3. Solubility of gases in liquids and its dependence on various factors. The Henry and Dalton laws.
4. Effect of electrolytes on solubility of gases (the Sechenov’s law). Solubility of gases in blood. The bends (caisson disease).
5. Solubility of liquids and solids. Distribution of substances between two immiscible liquids. The Nernst distribution law, its importance for permeability of biological membranes.
6. Preparation of solutions of a needed composition.

II Practical part

1. Laboratory work

Experiment 1. *Preparation of a solution with a given mass fraction of a solute.*

Prepare 200 g of 5% solution of copper (II) sulfate from crystalline salt and water. For this:

a) make the necessary calculations:

b) Weigh out the calculated weight on technical scales with accuracy up to 0.01 g and dissolve it in a suitable vessel with the required volume of water. Stir the prepared solution until complete dissolution of the solute.

Experiment 2. *Preparation of a solution of needed molarity.*

Prepare 100 ml of a 0.1 M solution of copper (II) sulfate from crystalline salt and water. For this:

a) make the necessary calculations:

b) weigh out the calculated weight with scales to the nearest 0.01 g and transfer it through a funnel to a 100 ml volumetric flask (pour a small volume of distilled water into the flask before it), wash it thoroughly from the funnel with distilled water. Dissolve the solute in the flask in a small amount of distilled water, then add water to the mark, cork the test tube and mix well.

II. Control tasks:

1. Solutions in life.

2. What is enthalpy dissolution factor?

3. What is entropy dissolution factor?

4. Gibbs' free energy. Write down the Gibbs' equation and explain the enthalpy dissolution factor, the entropy dissolution factor and their relationship to the mechanism of dissolution.

5. Explain the effect of temperature on solubility of gases. Provide examples.

6. Effect of pressure on solubility of gases. The Henry's Law.

7. The first Dalton law for gas mixtures.

8. The second Dalton law for gas mixtures.

9. Solubility of gases in blood. Caisson disease.

10. Effect of electrolytes on solubility of gases (the Sechenov's law).

11. Classification and characteristics of solubility of liquids in liquids.

12. The Nernst law of distribution. Practical applications.

13. Effect of the nature of the dissolved substance on dissolution of solids in water.

14. Effect of the nature of solvent and temperature on dissolution of solids in water.

15. Solubility of organic substances in nonpolar solvents.

Literature: P. 60-62

https://saylordotorg.github.io/text_general-chemistry-principles-patterns-and-applications-v1.0/s17-01-factors-affecting-solution-for.html

https://saylordotorg.github.io/text_general-chemistry-principles-patterns-and-applications-v1.0/s17-02-solubility-and-molecular-struct.html

https://saylordotorg.github.io/text_general-chemistry-principles-patterns-and-applications-v1.0/s17-04-effects-of-temperature-and-pre.html

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PRACTICAL CLASS № 5.

SUBJECT: Acid-base equilibrium and pH of biological fluids

IMPORTANCE OF THE SUBJECT: Acids, bases and salts are the principal classes of chemical substances. In the human organism, biochemical processes proceed in water solutions. In this, a water solution can be neutral, acidic or basic. The organism maintains acid-base equilibrium, and pH of biological fluids remains constant. Knowing pH values of biological fluids allows to find out pathologic alterations in the organism and prevent diseases.

AIM: Know basics of acid-base balance, learn to evaluate and predict processes that depend on the acidity of a solution.

I. Theoretical part:

1. Electrolytes in the human body. Dissociation degree and dissociation constant of weak electrolytes. Properties of solutions of strong electrolytes.
2. Types of protolytic reactions. Neutralization, hydrolysis and ionization reactions.
3. Hydrolysis of salts.
4. Degree of hydrolysis, its dependence on concentration and temperature.
5. Constant of hydrolysis.
6. Water-electrolyte balance in homeostasis.
7. Dissociation of water. Ionic product of water.
8. Power of hydrogen pH.
9. Normal and pathologic pH values of fluids of the human body.

II. Practical part

1. Laboratory work

Experiment 1. Reactions with formation of poorly dissociable compound

1) Add 3 - 4 mL of a sodium hydroxide solution to a test tube, add 2 - 3 drops of phenolphthalein. Then add a solution of sulfuric acid until discoloration.

2) Add 2 - 3 mL of a copper (II) sulfate solution to a test tube, add a NaOH solution dropwise until formation of a precipitate. Explain the phenomena that occur and write appropriate reaction equations.

Conclusions: _____

Experiment 2. Effect of identity of a salt on acidity of a solution

Add 2 - 3 mL of Na_2CO_3 , ZnSO_4 , NaCl and $\text{CH}_3\text{COONH}_4$ solutions to four test tubes. Check the acidity of these solutions using universal indicator paper, solutions of indicators of methyl orange and phenolphthalein. Record the results in the table.

Table

Solution of a salt	Color of the indicator			Solution medium	pH of the solution
	universal indicator paper	methyl orange	phenolphthalein		
Na_2CO_3					
ZnSO_4					
NaCl					
$\text{CH}_3\text{COONH}_4$					

Are all the salts subject to hydrolysis? Formulate equations of the salt hydrolysis reactions in molecular and ionic forms.

Conclusions: _____

Experiment 3. *Effect of temperature on the degree of hydrolysis*

Add 2 - 3 mL of a sodium acetate solution to two test tubes and add 2 - 3 drops of phenolphthalein. Heat one of the test tubes until boiling. Compare the colors of the cold and hot solutions. Cool the test tube in water and observe the color change of the solution. Explain the observations and write the reaction equation.

Conclusions: _____

Experiment 4. *Determination of pH value of biological fluids with indicators*

Determine pH values in biological fluids and make conclusions about acidity of the fluids.

Conclusions: _____

II. Control tasks:

1. Which of the salts – sodium chloride, ammonium sulfate, potassium nitrate, calcium acetate, potassium cyanide, sodium carbonate, ammonium acetate, sodium sulfate - undergo hydrolysis? Determine acidity of solutions of the salts.

2. Explain the effect of concentration of a salt and temperature on the degree of hydrolysis of the salt.

3. Calculate the pH value of a 1% HCl solution. Calculate the pH value of a 0.5 M ammonium hydroxide solution, $K_d = 1.85 \times 10^{-5}$.

4. How does the pH value of a 0.1 M nitric acid solution ($\alpha=0.86$) change, when 50 mL of water are added to 10 mL of the solution.

5. Indicate the relationship of values of pH and pOH in a solution: _____

6. Explain, how a change of pH effects processes in the organism.

7. Explain why pH in the inflammation zone is decreased: _____

8. Calculate the pH value of a solution obtained upon mixing of equal volumes of a 0.2 N sulfuric acid solution and 0.5 N sodium hydroxide solution. What is acidity of a medium is obtained in hydrolysis of CuCl_2 ?

9. Which salt, KCl or K_3PO_4 , is subject to hydrolysis and why?

10. What is active acidity?

11. What is alkalosis?

12. Calculate the pH value of a 0.01 M NaOH solution ($\alpha=1$).

13. Calculate $[\text{H}^+]$, if $\text{pH} = 5.3$.

14. The pH value of the pancreatic secretion is 8.5. What is the molar concentration of hydroxide ions and the pOH value?

15. Molar concentration of hydroxide ions in the gastric juice is equal to 10^{-12} mol/L. What is the pH value?

Literature: P. 70-80

https://saylordotorg.github.io/text_general-chemistry-principles-patterns-and-applications-v1.0/s20-aqueous-acid-base-equilibriums.html

“Accepted” “ _____ ” _____ 20 ____ . _____

PRACTICAL CLASS № 6.

SUBJECT: Titrimetric analysis. Methods of acid-base titration

IMPORTANCE OF THE SUBJECT: Diagnostics of diseases is often based on the data of clinical, biochemical and physico-chemical methods of analysis. Titrimetric analysis is one of the most important methods of chemical analysis. Permanganatometry is used in clinical laboratories to measure the amount of uric acid in blood plasma. In sanitary analysis the method is used to evaluate quality of drinking water and wastewater. Iodimetry is used in clinical laboratories to measure concentrations of glucose and the oxidative enzyme peroxidase in blood plasma; it is used in sanitary analysis to measure amount of active chlorine in chlorinated lime or residual chlorine in drinking water. Argentometry is used in clinical laboratories to measure concentrations of chlorides in biological fluids (blood, urine, gastric juice) and for analysis of drinking water. Complexometry is used to measure concentrations of cations of many metals in medicines, biological fluids and tissues of the organism, Mg^{2+} and Ca^{2+} cations in drinking water, cations of heavy metals in various objects.

AIM: On the basis of knowledge of physical and chemical properties of substances and solutions learn principal methods and techniques of titrimetric analysis. Develop practical skills of working with analytic utensils, get acquainted with the technique of performing volumetric analysis with the example of acid-base titration (neutralization method).

I. Theoretical part:

1. Basics of titrimetric analysis.
2. Classification of titration methods.
3. Application of the method in medical practice.
4. The method of acid-base titration.
5. Acid-base indicators.

II. Practical part

1. Laboratory work

Experiment 1. Titration of a strong acid with an alkali.

Fill a burette with a 0.1M KOH solution. Add 5 mL of a 0.1M HCl solution and 1-2 drops of methyl orange to a titration flask and titrate until the color of the indicator changes from pink to orange after adding one drop of an alkali solution. Repeat the titration three times (the difference between titrations must not exceed 0.2 cm³) and calculate the average volume. Fill in the table with the data.

Repeat the titration using phenolphthalein instead of methyl orange until the color of the indicator changes from colorless to light pink after adding one drop of an alkali solution. Draw titration curves, use the curves to make conclusions about the possibility to apply the indicators in the titration.

Table

Titration	Volume of the solution of a strong acid, mL	Volume of the solution of an alkali, mL	Average value of the volume of the alkali, mL
1st titration, methyl orange			
2nd titration, methyl orange			
3rd titration, methyl orange			

2. Calculate the mass of acetic acid in one liter of the solution:

$$m(\text{CH}_3\text{COOH}) = C_f(\text{CH}_3\text{COOH}) \cdot M_f(\text{CH}_3\text{COOH}) \cdot V,$$

where $M_f(\text{CH}_3\text{COOH})$ is the molar mass of equivalent CH_3COOH , g/mol;
 $V = 1\text{L}$ – the volume of the acetic acid solution.

3. Calculate the mass fraction of acetic acid in the weighed portion of concentrated acetic acid that had been used for preparation of 1L of the analyzed solution.

$$\omega(\text{CH}_3\text{COOH}) = \frac{m(\text{CH}_3\text{COOH}) \cdot 100\%}{m(\text{weighed portion})}.$$

Conclusion: _____

2. Control tasks

1. What is the process of titration? Basic requirements for reactions in titrimetric analysis.

2. What are titrants (work solutions)? _____

3. What are the equivalence point and neutrality point? When do they coincide? When don't they coincide? How to determine the equivalence point?

4. What reactions are methods of volumetric analysis based on? What are the requirements for reactions in titrimetric analysis? _____

5. What law are calculations of results of titrimetric analysis based on? Write down the formula. _____

6. What are acid-base indicators according to the Ostwald's theory? _____

7. What is the transition interval of the color of an indicator? _____

8. What are titration curves and how are they used to select an indicator?

9. Concentrations of what substances can be measured with the neutralization methods (alkalimetry, acidimetry)? _____

10. What work solutions (titrants) are used in the neutralization method? How to prepare them?

11. What are the standard (reference) substances in the neutralization method? What are the requirements for the substances? _____

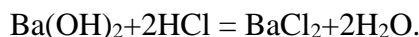
12. What indicators can be used in titrations of hydrochloric acid, orthophosphoric acid, ammonium chloride?

13. What are the types of acidity of gastric juice and in which units it is expressed? Determine the total acidity of gastric juice, if 13.6 mL of an alkali solution with a molar concentration of 0.0485 mol/L was used to titrate 10 mL of the gastric juice.

14. 25.0 mL ammonia solution was titrated against 25.05 mL 0.1244 N HCl. Calculate the mass of ammonia in 1 L of the solution.

15. 10.0 mL potassium hydroxide solution was titrated against 10.63 mL 0.02 M sulfuric acid. Calculate the mass of potassium hydroxide in 1 L of the solution.

16. What formula is used to calculate normality of barium hydroxide from the titration in the reaction:



Literature:

https://saylordotorg.github.io/text_general-chemistry-principles-patterns-and-applications-v1.0/s20-05-acid-base-titrations.html

“Accepted” “ _____ ” _____ 20 ____ . _____

PRACTICAL CLASS № 7.

SUBJECT: Properties of buffer solutions

IMPORTANCE OF THE SUBJECT: One of the most important aspects of homeostasis in the human body is the maintenance of a constant pH value carried out by buffer systems and physiological mechanisms. Buffer systems regulate concentrations of hydrogen and hydroxide ions and pH-dependent reactions. It is important to know properties and composition of buffer systems in studying biochemistry, physiology and clinical sciences.

AIM: learn to assess effect of buffer systems in maintaining a certain pH value; prepare buffer systems with a needed pH value and use knowledge of mechanism of action of buffer systems for prediction of biochemical processes in biological fluids in a change of acidity of a medium.

I. Theoretical part:

1. Classification of buffer solutions.
2. Mechanisms of action of buffer systems.
3. pH of buffer solutions (Henderson-Hasselbach equation).

II. Practical part

1. Laboratory work

Experiment 1. *Effect of acids and alkalis on pH of a buffer solution*

In a test tube, mix 5 mL of a 0.1M CH₃COOH solution and 5 mL of 0.1M CH₃COONa solution. Divide the obtained buffer system equally into three test tubes. Add 3 drops of a 0.1M HCl solution to the first test tube, add 3 drops of a 0.1M NaOH solution to the second test tube. Add 2 drops of methyl red indicator to each test tube. Compare the color of the solutions, write reaction equations, make conclusions.

Conclusions

Experiment 2. Effect of dilution on pH of a buffer solution.

Prepare a buffer solution (see Experiment 1) and divide it equally into two test tubes. Add 1 mL of water to the first test tube. Add 2 drops of methyl red indicator to each test tube. Compare the color of the solutions, write reaction equations, make conclusions.

Conclusions

2. Control tasks

1. Write a definition of buffer systems

2. Write classification of buffer systems, their formulas.

3. Explain why pH of hydrogencarbonate buffer system practically does not change upon addition of a small quantity of a strong acid.

4. Indicate the factors that effect pH of a buffer system.

5. Mechanism of action of buffer systems. Write reaction equations.

6. Buffer systems of the human organism. Write formulas and names of components.

7. Explain the action mechanism of the phosphate buffer system.

8. Indicate which factors effect buffer capacity.

9. Main equation of buffer systems. The Henderson-Hasselbach formula.

10. Explain the shift of acid-base metabolism that occurs at prolonged hypoventilation.

11. Calculate pH of a buffer solution that is obtained in mixing of 0.04 L of 0.15M ammonia solution and 0.02 L of 0.25M ammonium chloride solution. $pK_{\text{NH}_4\text{OH}} = 4,74$.

12. What is the alkali buffer capacity of a phosphate buffer system if a pH change of 1 unit occurred upon titration of 10.0 mL of the buffer solution against 5.3 mL of 0.01M potassium hydroxide solution.

13. Calculate the acid buffer capacity of the hydrogencarbonate buffer system of blood plasma, if pH of the buffer solution changed from 7.4 to 6.7 upon addition of 14.3 mL of 0.05M hydrochloric acid to 25 mL of the buffer solution.

14. Calculate the volumes of a 0.1M solution of CH_3COOH and 0.1M solution of CH_3COONa needed to prepare 200 mL of a buffer solution with $\text{pH} = 5.24$. pK_a of acetic acid is equal to 4.8.

Literature: P. 80-86

https://saylordotorg.github.io/text_general-chemistry-principles-patterns-and-applications-v1.0/s20-06-buffers.html

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PRACTICAL CLASS № 8.
SUBJECT: Role of buffer solutions in biosystems

IMPORTANCE OF THE SUBJECT: Blood, urine, saliva, extracellular fluid and other fluids in the human organism contain buffer systems to maintain stable pH. Buffer solutions play a vital role in maintaining almost constant pH in many chemical reactions occurring in biological and other systems. Certain disorders cause pH shift to acidic (acidosis) or alkaline (alkalosis). It is known that a prolonged shift in the pH of the blood by 0.3-0.4 pH units can lead to the death of the patient. Buffer capacity is an important parameter of a buffer system because it shows how much acid or base can be neutralized by the buffer system. Potentiometric measurement of pH change is used to calculate buffer capacity.

AIM: Explain the mechanism of action of buffer systems and their role in maintaining acid-base equilibrium in biosystems. Determine acid buffer capacity and alkali buffer capacity of buffer solutions.

I. Theoretical part:

1. Buffer systems of blood:
 - a) bicarbonate (hydrogen carbonate) buffer;
 - b) phosphate buffer;
 - c) protein buffer systems.
2. The concept of the acid-base state of blood.
3. Buffer capacity and factors on which it depends.

II. Practical part

1. Laboratory work

Experiment 1. Studying properties of buffer solutions with potentiometry

1. Calculate the volume of _____ and the volume of _____ needed to prepare 20 mL of _____ buffer solution with pH = _____.

The concentrations of acid and salt are both equal to 0.2 mol/L. $pK_a(\text{CH}_3\text{COOH}) = 4.75$.

2. Using a burette, measure out appropriate volumes of the acid and salt solutions and mix thoroughly.

3. Measure the pH of the resulting buffer solution with a pH meter using the manual for the instrument.

4. Using a pipette, take 5 mL of the buffer solution, transfer it to a 50 mL volumetric flask, make it to the mark with distilled water, mix thoroughly.

Divide the remaining buffer solution into two equal parts, and transfer them to two beakers. Using pipettes, add 1 mL of a 0.01M HCl solution to one of the beakers and 1 mL of a 0.01M NaOH solution to the other one.

5. Measure the pH of the solutions after the dilution and addition of acid and alkali with the same pH meter. Fill in the table with the data:

Solution	pH value	
	calculated	measured
Initial buffer solution		
After addition of HCl		
After addition of NaOH		
Diluted solution		

Processing of the experimental data:

1. Calculation of pH of the buffer solution after the addition of a strong acid.

2. Calculation of pH of the buffer solution after the addition of an alkali.
3. Calculation of pH of the buffer solution after dilution.
4. Make conclusions.

2. Control tasks

1. Buffer systems of the human body. Write formulas and names of components.

2. Explain what the mechanism of the phosphate buffer system's buffer action is based on.

3. Indicate what factors effect buffer capacity?

4. Explain the shift of acid-base metabolism that occurs at prolonged hypoventilation.

5. What is the alkali buffer capacity of a phosphate buffer system if a titration of 10.0 mL of the buffer system with 5.3 mL of 0.01M potassium hydroxide solution changed its pH value in 1 unit.

6. Calculate the acid buffer capacity of the hydrogencarbonate buffer system of blood plasma, if pH of the buffer solution changed from 7.4 to 6.7 upon addition of 14.3 mL of 0.05M hydrochloric acid to 25 mL of the buffer solution.

7. Calculate the acid buffer capacity and alkali buffer capacity of blood plasma, if:

- 1) 10 mL of 0.2M HCl solution should be added to 100 mL of the blood plasma to change its pH from 7.4 to 3.4;
- 2) 0.8 mL of 0.1M NaOH solution should be added to 100 mL of the blood plasma to change its pH

from 7.4 to 9.4.

Literature: P. 80-86

https://saylordotorg.github.io/text_general-chemistry-principles-patterns-and-applications-v1.0/s20-06-buffers.html

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PRACTICAL CLASS № 9.
SUBJECT: Colligative properties of solutions

IMPORTANCE OF THE SUBJECT: Knowledge of colligative properties of dilute solutions renders it possible to analyze such phenomena as diffusion, osmosis, turgor, transport of substances across biological membranes, ion exchange in the organism, hemolysis, membrane equilibrium.

AIM: learn to analyze association between colligative properties and concentrations of solutions; apply theory of osmotic phenomena to explain processes in the human organism.

I. Theoretical part:

1. Colligative properties of diluted solutions of non-electrolytes:
 - a) relative depression of the pressure of saturated vapor of the solvent over the solution. The Raoult's law;
 - b) boiling point elevation and freezing point depression of a solution compared to the solvent.
 - c) use of osmometry and cryometry in medical, biological and laboratory-diagnostic research.
 - e) osmosis; osmotic pressure. Van't Hoff's law. Hemolysis and plasmolysis.
2. Colligative properties of dilute solutions of electrolytes:
 - a) isotonic coefficient.
 - b) hypo-, hyper- and isotonic solutions in medical practice.
 - c) role of osmosis in biological systems.
3. Properties of semipermeable membranes.
4. Oncotic pressure of blood plasma.

II Practical part

Experiment 1. *Observing osmosis*

Fill an osmometer with a fuchsin-colored sugar solution. Immerse the osmometer in a beaker filled with water. Mark the initial level of the solution in the osmometer, and then mark it again after half an hour. Explain the phenomenon you are observing. Draw a scheme of the osmometer and record the results of the experiment.

Experiment 2. *Preparation of isotonic solution*

1. Prepare 500 g of a solution isotonic to blood with a mass fraction of NaCl 0.9%.

1. Calculations:

Calculate of the mass of NaCl needed to prepare the isotonic solution with the formula for mass fraction.

$$W(\text{solute}) = \frac{m_{\text{solute}}}{m_{\text{solution}}} \times 100\%,$$

Calculate the mass of water needed to prepare the solution:

2. Prepare a solution with the needed mass fraction:

Weigh out the calculated amount of the solute on a technical scale, transfer the solute into a convenient non-measuring vessel and add the calculated amount of water. Mix the solution until complete dissolution of the solute.

a) *Calculate the mass of NaCl needed to prepare 200 g of a solution isotonic to blood.*

b) *Calculate the mass of glucose needed to prepare 100 g of a 4% solution isotonic to blood.*

c) *Calculate the mass of water needed to prepare 50 mL of a 0.9% NaCl solution.*

2. Control tasks:

1. What properties of solutions belong to colligative properties? Give definitions and name the properties: _____

2. Semipermeable membranes (definitions, examples). _____

3. The van't-Hoff's law for osmotic pressure (definition), equation for non-electrolytes and electrolytes. _____

4. The van't-Hoff's isotonic coefficient, its relationship with the degree of dissociation.

13. A solution containing 18 g of a solute in 500 mL water has an osmotic pressure of 0.0456 MPa at 0° C. Calculate the molar mass of the solute.

14. Osmotic pressure of blood plasma at 37° C was 0.77 MPa. Calculate the mass of sucrose needed to prepare 0.5 L of a solution isotonic to the blood plasma.

15. Calculate the osmotic pressure of a 0.2 M urea solution at 0° C.

16. Calculate the osmotic pressure of a 5% (by mass) glucose solution, $t=37^\circ\text{C}$, $\rho=1\text{g/mL}$.

17. Calculate the osmotic pressure of a 5.85% (by mass) sodium chloride solution at 0° C. The degree of dissociation of sodium chloride is 0.96, and $\rho = 1.04\text{ g/mL}$.

Literature: P. 62-70

https://saylordotorg.github.io/text_general-chemistry-principles-patterns-and-applications-v1.0/s17-05-colligative-properties-of-solu.html

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PRACTICAL CLASS № 10.

SUBJECT: Principles of thermodynamics of biochemical processes

IMPORTANCE OF THE SUBJECT: knowledge of basics of chemical thermodynamics is necessary for understanding energy transformations in biochemical processes. Calculations of heat effects are used in nutrition science to determine caloric content of foods.

AIM: be able to perform thermochemical calculations for estimation of caloric content of foods; learn to calculate and experimentally determine heat effects of chemical reactions and processes.

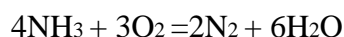
I. Theoretical part

1. Basic concepts of chemical thermodynamics: systems, system parameters, system state functions, processes, heat, work, internal energy, enthalpy.
2. The First Law of Thermodynamics, its biological importance.
3. Thermochemistry. Heat effects of chemical reactions. The Hess law and its consequences.
4. The Second Law of Thermodynamics, its biological importance.
5. Entropy, factors that effect it. Role of entropy in systems and processes.
6. Gibbs energy, its importance for thermodynamic calculations. Effect of entropy and enthalpy factors on spontaneity of processes.
7. Energy metabolism in living organisms as open systems. Macroergic compounds.
8. Conjugation of energy in living systems: exergonic and endergonic processes in the body.

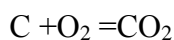
II. Practical part

1. Calculation problems

1. Determine Gibbs energy for the reaction and decide whether the reaction can occur spontaneously under standard conditions:



2. Calculate enthalpy of the reaction and make a conclusion about the heat effect of the reaction:



3. A diet at diabetes limits carbohydrate content in food to 50% while maintaining the physiological normal values of fats and proteins. Calculate the calorie content of the diet:

2. Control tasks

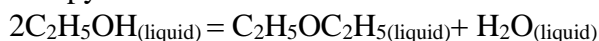
1. Calculate the standard enthalpy of the chemosynthesis reaction in bacteria *Thiobacillus denitrificans*:

$6\text{KNO}_3(\text{solid}) + 5\text{S}(\text{solid}) + 2\text{CaCO}_3(\text{solid}) = 3\text{K}_2\text{SO}_4(\text{solid}) + 2\text{CaSO}_4(\text{solid}) + 2\text{CO}_2(\text{gas}) + 3\text{N}_2(\text{gas})$,
from the values of the standard enthalpies of formation of the substances:

	K_2SO_4	CaSO_4	CO_2	KNO_3	CaCO_3
$\Delta H_{\text{formation}}^0$, kJ/mol	-1438	-1432	-393,5	493	-1207

Is the reaction endothermic or exothermic?

2. Calculate the standard enthalpy of the reaction:



from the values of the standard enthalpies of combustion of the substances:

$$\Delta H_{\text{combustion}}^0 \text{C}_2\text{H}_5\text{OH} = -1368 \text{ kJ/mol};$$

$$\Delta H_{\text{combustion}}^0 \text{C}_2\text{H}_5\text{OC}_2\text{H}_5 = -2727 \text{ kJ/mol.}$$

3. Using the data in the table, determine the role of the enthalpy and entropy factor in the reaction:

	$4\text{HCl}_{(\text{gas})} + \text{O}_{2(\text{gas})} = 2\text{Cl}_{2(\text{gas})} + 2\text{H}_2\text{O}_{(\text{liquid})}$			
	HCl	H ₂ O	O ₂	Cl ₂
$\Delta H_{\text{formation}}^0$, kJ/mol	-92,3	-286	0	0
S^0 , J/(mol·K)	187	70	205	233

Determine the temperature at which the reaction will occur spontaneously.

4. Determine the temperature at which the reaction of trypsin denaturation will occur spontaneously, if $\Delta H_{\text{reaction}}^0 = 283 \text{ kJ/mol}$, $\Delta S_{\text{reaction}}^0 = 288 \text{ J/(mol·K)}$. _____

5. What is a thermodynamic process? What are types of processes by constancy of one of parameters? _____

6. What is meant by the term "a state of a system"? What are possible states of a system? _____

7. What variables are called state functions? List and describe state functions. _____

8. What is enthalpy? What are its measurement units? _____

9. What is entropy? What are its measurement units? _____

10. What is Gibbs energy? How can it be calculated? What can be found from the value of the function?

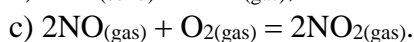
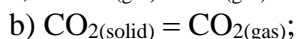
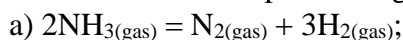
11. What reactions are called exergonic and what reactions are called endergonic?

12. Formulate the First Law of Thermodynamics. Explain equivalence of heat and work.

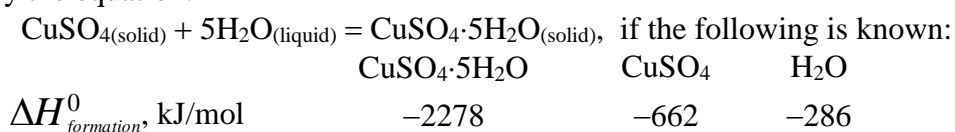
13. Formulate the Hess law and its consequences. What is the standart enthalpy of formation (combustion) of substances?

14. Formulate the Second Law of Thermodynamics. Under what conditions is a process spontaneous in an isolated system?

15. Without performing calculations, determine the sign of ΔS^0 of the following processes:



16. Calculate the heat effect of the reaction of formation of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ crystalline hydrate, which is obtained by the equation:



Literature: P. 24-26

https://saylordotorg.github.io/text_general-chemistry-principles-patterns-and-applications-v1.0/s22-chemical-thermodynamics.html

https://saylordotorg.github.io/text_general-chemistry-principles-patterns-and-applications-v1.0/s22-01-thermodynamics-and-work.html

https://saylordotorg.github.io/text_general-chemistry-principles-patterns-and-applications-v1.0/s22-02-the-first-law-of-thermodynamic.html

https://saylordotorg.github.io/text_general-chemistry-principles-patterns-and-applications-v1.0/s22-03-the-second-law-of-thermodynamic.html

https://saylordotorg.github.io/text_general-chemistry-principles-patterns-and-applications-v1.0/s22-04-entropy-changes-and-the-third-law.html

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PRACTICAL CLASS № 11.

SUBJECT: Principles of kinetics of biochemical processes

IMPORTANCE OF THE SUBJECT: Chemical kinetics is the basis for studies of biochemical processes, pharmacokinetics of medicinal substances in clinical diagnosis. The study of concepts of chemical kinetic is an important component for understanding processes of metabolism and energy exchange in the organism at the cellular level. Almost all processes occurring in living organisms are catalytic. Chemical processes in living organisms proceed with the help of biological catalysts - enzymes. Knowledge of the fundamentals of catalytic (enzymatic) processes, formation of metabolites, absorption and transformation of medicinal substances, mechanisms of enzymatic reactions is an important part in understanding processes of metabolism and energy exchange in the organism at the cellular level.

AIM: be able to theoretically substantiate and practically determine the dependence of the speed of reaction on various factors, explain and practically determine effects of catalysts on reaction rates.

I. Theoretical part:

1. Basic concepts of chemical kinetics: momentary and average rate of a chemical reaction, simple and complex reactions; homogeneous and heterogeneous reactions; molecularity and order of reactions.

2. Kinetics of complex reactions: parallel, sequential, conjugate, reversible, chain. The concept of antioxidants. Free radical reactions in the living organism.

3. Dependence of the reaction rate on: a) nature of the reactants; b) concentration of reactants (the law of mass action, the physical meaning of the constant of reaction rate); c) temperature (a concept of the theory of active collisions, role of activation energy, the Arrhenius equation, the van't Hoff rule). A concept of the theory of the transition state (activated complex).

4. Chemical kinetics as the basis for studying the rates and mechanisms of biochemical processes. Peculiarities of kinetics of biochemical processes.

5. Catalysis and catalysts. Types of catalysis. Mechanism of action of the catalyst.

6. Homogeneous, heterogeneous and microheterogeneous catalysis. Acid-base catalysis. Autocatalysis. Promoters and catalytic poisons.

7. Enzymes as biological catalysts. Basics of kinetics of enzymatic reactions. Features of action of enzymes: selectivity, efficiency, effect of pH and temperature on enzymatic reactions. Effect of concentrations of the enzyme and substrate on the rate of an enzymatic process.

II. Practical part

1. Laboratory work

1. Effect of concentrations of reactants on the reaction rate.

Prepare solutions of sodium thiosulfate of different concentrations:

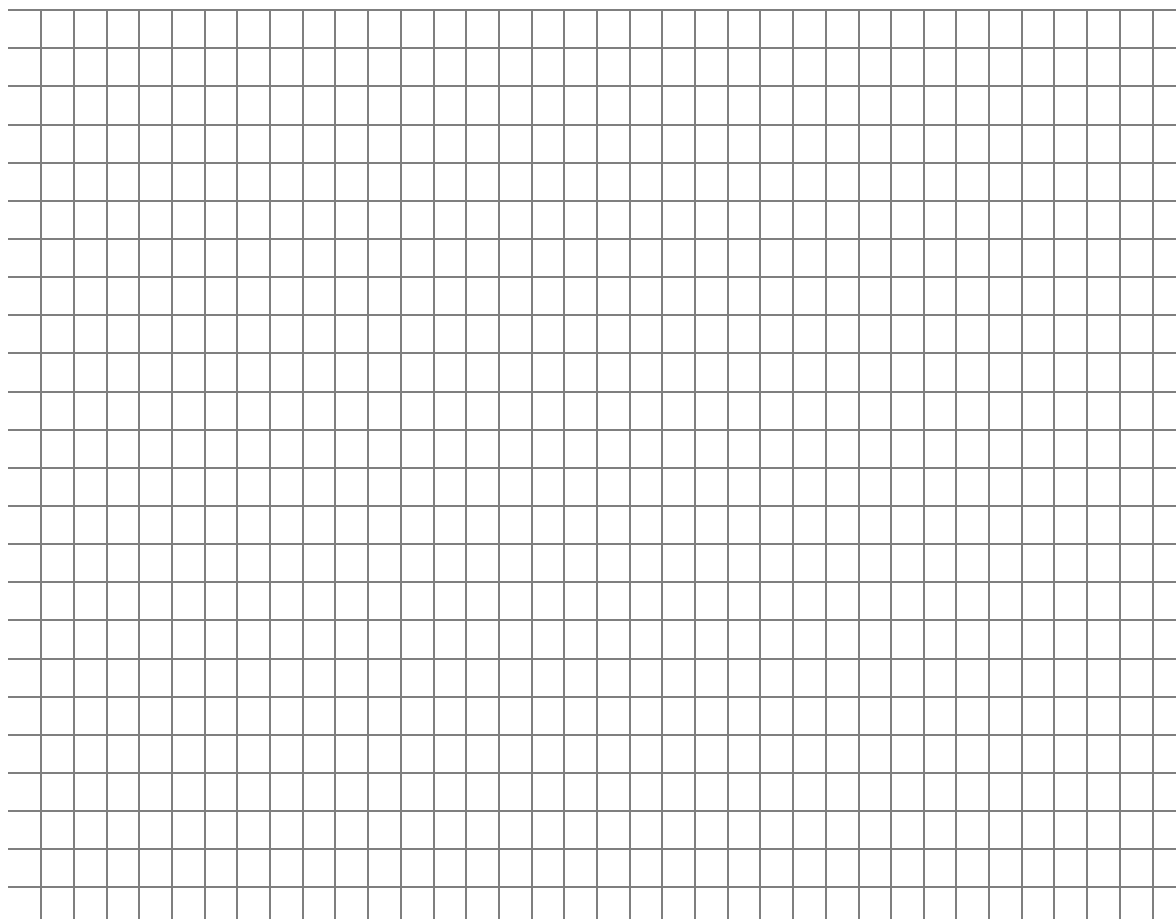
	Test tube 1	Test tube 2	Test tube 3
$\text{Na}_2\text{S}_2\text{O}_3$ -0.5M	5 drops	10 drops	15 drops
H_2O	10 drops	5 drops	-

Add 1 drop of 0.25 M H_2SO_4 solution to the first test tube and determine the time of the end of the reaction by an appearance of turbidity of the reaction mixture. Similarly, perform the experiment with the second and third test tubes. Fill in the data of the experiment in the table.

Table 1

№ of test tube	Number of drops of $\text{Na}_2\text{S}_2\text{O}_3$	Number of drops of H_2O	Time taken by the process, s.	Relative velocity 1/s
1	5 drops	10 drops		
2	10 drops	5 drops		
3	15 drops	0		

Construct a graph of the effect of concentrations of the reactants on the reaction rate, with the concentration of $\text{Na}_2\text{S}_2\text{O}_3$ (number of drops) on the x-axis, and the relative reaction rate on the y-axis. Write the reaction equation and draw conclusions.



2. Enzymatic catalysis

Place 5 drops of 0.5% starch solution in 2 test tubes. Add equal amount of your own saliva to one test tube and mix thoroughly. After 1-2 minutes, add one drop of a very dilute solution of iodine

in potassium iodide (light yellow solution) to the both test tubes.

In conclusion, note the phenomena that are observed in the test tubes number 1 and number 2. In which of the products there is no blue color when adding iodine and why? What enzyme in saliva catalyzes the reaction of starch hydrolysis? _____

2. Control tasks

1. What is the rate of a chemical reaction? _____

2. List factors that effect the rate of a chemical reaction. _____

3. Formulate the law of mass action for the rate of chemical reactions. The rate constant.

4. Effect of concentrations on the reaction rate: _____

5. Effect of temperature on the reaction rate. The van't-Hoff's rule. Peculiarities of the temperature coefficient of reaction rate for biochemical processes _____

6. How does the reaction rate depend on the activation energy? Write the Arrhenius equation, explain it. _____

7. What is an activated complex? Why do reactions proceed through the stages of formation of activated complexes? _____

8. The order of reaction. Kinetic equations of reactions of the first, second and zero order

9. Molecularity of reactions. Provide examples. _____

10. Consider the rates the forward and reverse gas phase reactions: of

$$v_{\text{forward}} = k_1 \cdot C^2(\text{A}) \cdot C(\text{B}); \quad v_{\text{reverse}} = k_2 \cdot C^2(\text{C}),$$

How do the rates change:

a) when the pressure in the system increases 2 times? _____

b) when the volume of the gases increases 2 times? _____

11. At 150°C, the reaction proceeds to completion in 10 minutes. Assuming that the temperature coefficient γ is equal to 2, calculate in how many minutes would the reaction proceed to completion at 170°C. _____

12. The reaction rate is expressed by the equation: $v = k \cdot C(\text{A}) \cdot C^2(\text{B})$. How many times will the rate of the reaction change upon increasing concentration of the reactants 3 times? _____

13. The rate of a reaction is expressed by the equation: $\text{A}(\text{s}) + 2\text{B}(\text{gas}) = \text{C}(\text{s})$. How will the reaction rate change if the concentration of substance B is increased 2 times? _____

14. The rate of a reaction increases 8 times with increasing temperature from 40°C to 70°C. Calculate the temperature coefficient γ of the rate of this reaction. _____

15. During the Chernobyl disaster (1986), a Cs-137 radionuclide was emitted, with a half-life of 30 years. Calculate what fraction of the radionuclide remains in the organism at present. _____

16. The shelf life of a drug at 20°C is 0.5 years, and at 10°C - 1.5 years. Calculate the temperature coefficient of the rate of this reaction. _____

17. What is catalysis, autocatalysis, catalyst? _____

18. What types of catalysis exist? Explain. _____

19. What causes an increase in the reaction rate upon an addition of a catalyst? Explain the answer. _____

20. Explain the mechanism of heterogeneous catalysis _____

21. What are activators, reactivators and inhibitors of catalysis? _____

22. What are promoters and catalytic poisons _____

23. Peculiarities of action of enzymes. Explain _____

24. Enzymes as biological catalysts. _____

25. Dependence of speed of enzymatic processes on concentration of enzyme and substrate. _____

26. When using a laundry detergent containing enzymes why is it important to soak the linen in its solution instead of boiling it? _____

27. Explain the difference between the effects of a significant temperature increase on the activity of inorganic catalysts and enzymes _____

28. Homogeneous catalysis. Explain. _____

29. What is the cause of the specificity and selectivity of a catalyst? _____

Literature: P. 37-40

https://saylordotorg.github.io/text_general-chemistry-principles-patterns-and-applications-v1.0/s18-00-chemical-kinetics.html

https://saylordotorg.github.io/text_general-chemistry-principles-patterns-and-applications-

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PRACTICAL CLASS № 12.

SUBJECT: Formation of coordination complexes in heterogeneous systems

IMPORTANCE: Coordination compounds have specific structure and bond type. A great many compounds in nature are classified as coordination compounds because of their structure, properties and biological activity. Metal-containing enzymes, hemoglobin, myoglobin, vitamin B12 are biologically active coordination compounds.

AIM: to be able to explain properties and practically determine presence of coordination compounds using qualitative reactions.

I. Theoretical part:

1. Reactions of formation of complexes. A. Werner's coordination theory and modern ideas of structure of complex compounds.
2. Classification of coordination complexes by the charge of the internal sphere and by the nature of the ligands. Denticity.
3. Intracoordination complexes (chelates).
4. Iron -, cobalt-, copper- and zinc-containing biocomplex compounds. Metal-ligand homeostasis. Disturbances of metal-ligand homeostasis.
5. Complexons and their applications in medicine as antidotes at poisoning with heavy metals (chelation therapy) and as antioxidants during storage of medicinal preparations.

II. Practical part

1. Laboratory work

Experiment 1. *Obtaining and properties of a complex ion containing Fe^{2+}*

Potassium ferricyanide $K_3[Fe(CN)_6]$ (red blood salt) forms a blue precipitate of “Turnbull's Blue” in reactions with Fe^{2+} ions.



The precipitate is insoluble in acids, but is decomposed in alkalis.



To a test tube, add 3-4 drops of a solution of a Fe^{2+} salt, 1-2 drops of hydrochloric acid (to prevent salt hydrolysis) and 2-3 drops of a potassium ferricyanide solution. A blue precipitate of “Turnbull's Blue” is obtained. Check whether the obtained precipitate is soluble in alkalis.

Conclusions: _____

Experiment 2. *Obtaining and properties of a complex ion containing Fe^{3+}*

Potassium ferrocyanide $K_4[Fe(CN)_6]$ forms a dark blue precipitate “Prussian Blue” in reactions with Fe^{3+} cation.



The precipitate is not soluble in acids, but decomposes in alkalis.



To a test tube, add 3-4 drops of a FeCl_3 solution, 2-3 drops of hydrochloric acid (to prevent salt hydrolysis) and 3-4 drops of a potassium ferrocyanide $\text{K}_4[\text{Fe}(\text{CN})_6]$ solution. A dark blue precipitate is obtained. Add 6-8 drops of a sodium hydroxide solution to the precipitate until it dissolves.

Conclusions: _____

Experiment 3. *Obtaining ammonia-containing coordination complex of Cu^{2+} ion*

An excess of ammonium hydroxide forms a blue-violet complex ion in reactions with Cu^{2+} ion.



Add 3-4 drops of a Cu^{2+} salt solution to a test tube. Add an ammonia solution dropwise until a light blue precipitate is obtained. Add an excessive amount of ammonia solution until a bright blue transparent solution is obtained.

Conclusions _____

2. Control tasks:

1. What chemical compounds are called coordination complexes? Write the classification of coordination compounds by the nature of the ligands and provide appropriate examples.

2. Write the classification of coordination compounds by the charge of the internal sphere and provide appropriate examples.

3. Classify coordination compounds by the nature of the ligands.

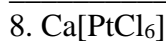
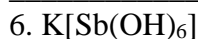
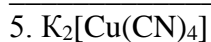
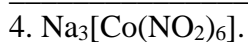
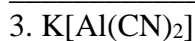
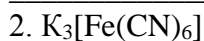
4. Classify coordination compounds by the charge of the complex ion.

5. Intracoordination complexes (chelates): definitions, examples, roles.

6. Provide examples of iron-, cobalt, copper and zinc-containing biocomplex compounds in the human body, indicate their functions.

7. What is chelation therapy? What are complexons? Provide examples of complexons and their applications in dentistry.

For the coordination compound, give the name and determine the charge of the complex ion, the degree of oxidation and the coordination number of the complexing agent:



Write the chemical formula and write down the first degree of dissociation of the coordination complex:

1) sodium dicyanoargentate

2) potassium tetrahydroxoplumbate (II)

3) sodium tetraiodinemercurate (II)

4) potassium hexacyanochromate (III)

5) hexaamminocobalt (III) bromide

6) sodium hexanitrocobaltate (III)

7) diaquatetraamminonickel (II) nitrate

8) sodium tetrahydroxyzincate (II)

Literature: P. 16-23

https://saylordotorg.github.io/text_general-chemistry-principles-patterns-and-applications-v1.0/s21-solubility-and-complexation-eq.html

https://saylordotorg.github.io/text_general-chemistry-principles-patterns-and-applications-v1.0/s21-00-solubility-and-complexation-eq.html

https://saylordotorg.github.io/text_general-chemistry-principles-patterns-and-applications-v1.0/s21-02-factors-that-affect-solubility.html

https://saylordotorg.github.io/text_general-chemistry-principles-patterns-and-applications-v1.0/s21-03-the-formation-of-complex-ions.html

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PRACTICAL CLASS № 13.

SUBJECT: Reactions of precipitation and dissolution

IMPORTANCE: Understanding processes of homogeneous and heterogeneous equilibrium in the body and their relation to other types of processes in solutions allows to analyze conditions for formation and dissolution of precipitates, for example, the formation of stones in the kidneys (nephrocalcinosis) or in the gall bladder (cholelithiasis); plays an important role in forming a holistic approach to considering both general homeostasis of the organism and homeostasis of organs and tissues.

AIM: in an experiment, study effects of various factors on shifts of chemical equilibrium in a thermodynamic system.

I. Theoretical part:

1. Chemical equilibrium. The constant of chemical equilibrium and ways of its representation.
2. Shifts of chemical equilibrium upon changes of temperature, pressure, concentrations of substances. The Le Chatelier's principle.
3. Reactions of precipitation and dissolution. The solubility product. Conditions for formation and dissolution of precipitates.
4. Role of heterogeneous equilibrium with the participation of salts in the overall homeostasis of the organism.
5. Chemical basis of mineralization of bone and dental tissues.

II. Practical part

1. Laboratory work

1. Effect of concentrations of substances on a shift of chemical equilibrium

Add 50 mL water to a flask, add 3 drops of a saturated FeCl_3 solution and 3 drops of a saturated NH_4SCN solution. Mix the resulting solution thoroughly and divide it into four test tubes.

Add 2 drops of the saturated solution of FeCl_3 to the first test tube. Add 2 drops of the saturated solution of NH_4SCN to the second test tube. Add NH_4Cl crystals on the tip of a spatula to the third test tube. Leave the fourth test tube for comparison.

Observe the changes and fill in the table with the results of the experiment. Write an equation of the reaction, the expression of its constant of chemical equilibrium, make conclusions about the effect of concentration on shifts of chemical equilibrium.

Table 6

№ of test tube	Added reagent	Color	Conclusions (direction of the equilibrium shift)
1	FeCl ₃		
2	NH ₄ SCN		
3	NH ₄ Cl (crystals)		

Conclusions: _____

2. Effect of temperature on shifts of chemical equilibrium

To two test tubes, add 5 mL of a starch solution and 1 drop of a solution of iodine. Heat one of the test tubes until the color changes, and then cool it. Leave the second test tube for comparison. Describe your observations. Make conclusions.

Conclusions: _____

2. Контрольні завдання

1. What processes are called irreversible. Give examples. _____

2. What processes are called reversible. Give examples. _____

3. What is chemical equilibrium? Why is it called dynamic? Explain using an example.

4. The Le Chatelier's principle _____

5. What is the kinetic description of chemical equilibrium?

6. How do temperature and pressure effect chemical equilibrium? Explain with an example.

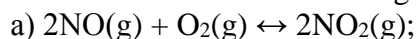
7. Explain the constant of chemical equilibrium of a heterogeneous system. Example

8. The solubility product _____

9. Condition for formation of a precipitate _____

10. Condition for dissolution of a precipitate _____

11. Write expressions of the equilibrium constants of the following reactions:



12. In which direction will the chemical equilibrium of the following reversible reaction shift:

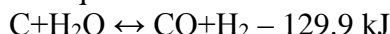


upon an increase in pressure; upon an increase of temperature? How will the concentrations of system components change? _____

13. Determine the direction of the shift of the equilibrium of the following reversible reactions as the temperature rises and the concentration of the reaction product increases:



14. In which direction will the chemical equilibrium of the following reversible reaction shift:



upon an increase in pressure; upon an increase of temperature? How will the concentrations of system components change? _____

15. How should the oxygen concentration be changed so that the rate of the homogeneous simple reaction: $2\text{NO}(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{NO}_2(\text{g})$ would not change upon a twofold decrease in concentration of nitric oxide (II)? _____

16. $\text{SP}_{\text{CaSO}_4} = 2.5 \cdot 10^{-5}$. Calculate solubility of CaSO_4 in mol/L. _____

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PRACTICAL CLASS № 14.
SUBJECT: Measurement of electrode potentials

IMPORTANCE: All reactions occurring in a living organism are accompanied by electrochemical phenomena. These include three types of bioelectric potentials (diffusion, membrane and phase potentials), as well as oxidation-reduction potentials, which arise in a transfer of electrons between molecules and produce energy necessary for the life of the organism.

AIM: in an experiment, determine concentration of an acid with the method of potentiometric titration.

I. Theoretical part:

1. Electrode potentials and mechanism of their formation.
2. The Nernst equation. Normal (standard) electrode potential.
3. Standard hydrogen electrode.
4. Measurement of electrode potentials. Indicator electrodes and reference electrodes. Silver-silver chloride electrode. Ion-selective electrodes. Glass electrode.
5. Galvanic cells.
6. Diffusion potential. Membrane potential. Rest potential. Action potential.
7. The oxidation-reduction (redox) potential as a measure of oxidative and reducing capacity of systems. The Nernst-Peters equation. Standard redox potential.
8. Prediction of direction of redox reactions from values of redox potentials. Values of redox potentials in the mechanism of biological oxidation.
9. Potentiometry. Potentiometric titration.

II Practical part

1. Laboratory work

Measurement of concentration of a strong acid with potentiometric titration.

The method is based on the fact that the equivalence point of a reaction is determined not by observing a change of the color of an indicator, as in a regular titration, but by an abrupt jump of the value of a potential of the indicator (ion-selective) electrode at the equivalence point.

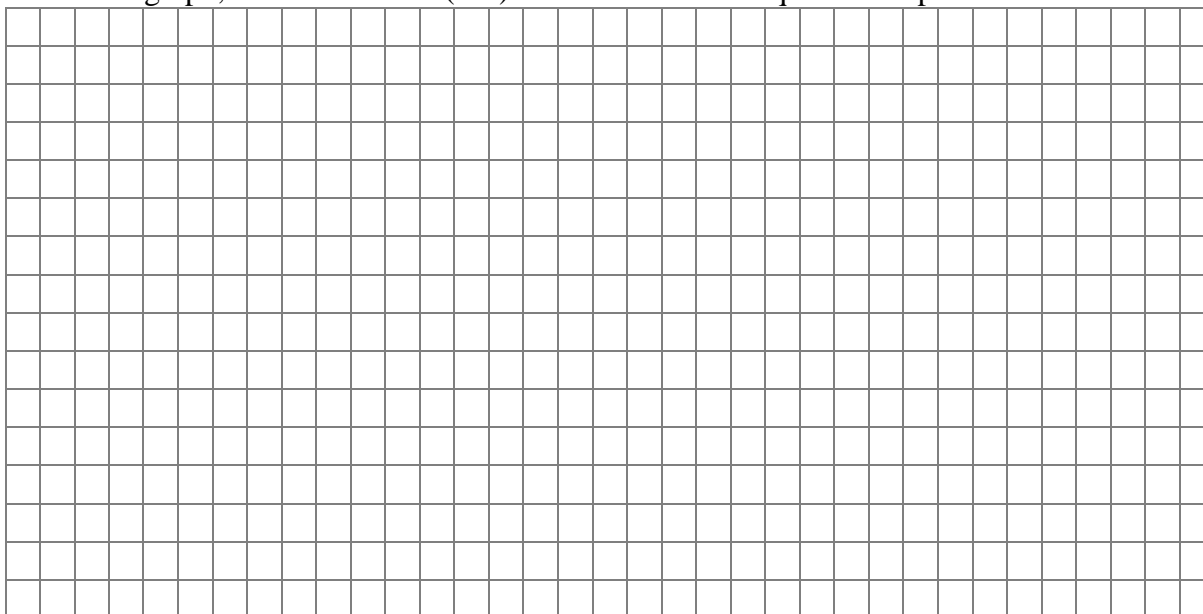
Add 20 mL of an analyzed HCl solution to a 50 mL beaker, measure the pH and record it in the table. Using a burette, add sequentially the following volumes of a 0.1N NaOH solution to the HCl solution in the beaker: 4 mL – two times; 1 mL – two times; 0.5 mL – two times. Stir the solution in the beaker after each addition of alkali, and measure the pH. After an abrupt jump of pH, add 4 mL of the NaOH solution – three times, measuring pH each time. The pH initially changes slowly, then it undergoes a sharp change close to the equivalence point, and then it again changes slowly. Stop the titration after three different pH values are obtained. Fill in the table with the results:

Table 7.

№ of measurement	Volume of added alkali, mL	pH	№ of measurement	Volume of added alkali, mL	pH
1			5		
2			6		
3			7		
4			8		

Write equations of the reaction in the molecular and ionic form:

Based on the obtained data, plot a graph, with the volume (mL) of NaOH on the x-axis, and the pH values on the y-axis. The equivalence point is located in the middle of the titration curve jump. From the graph, find the volume (mL) of the alkali at the equivalence point.



Calculation: With the values of volumes of NaOH and HCl and normality of NaOH, calculate N_{HCl} (the normality of the analyzed solution of HCl) from the formula:

$$N_{\text{NaOH}} \cdot V_{\text{NaOH}} = N_{\text{HCl}} \cdot V_{\text{HCl}}$$

Conclusion: write down the obtained value of N_{HCl} : _____

2. Control tasks

1. Describe the mechanism of formation of the electrode potential _____

Definition:

- electrode _____

- electrode potential _____

2. Nernst equation (in a full and simplified form) _____

3. The standard hydrogen electrode _____

4. What is a standard metal electrode? _____

5. Galvanic cell. Principle of action _____

6. Ion-selective electrodes _____

7. Indicator electrodes. Examples _____

8. Reference electrodes. Examples _____

9. The glass electrode. Principle of action _____

10. The silver-silver chloride electrode. Principle of action. _____

11. Diffusion potential _____

12. Membrane potential _____

13. Calculate the electrode potential of the magnesium electrode in a solution of its salt, if the concentration of the Mg^{2+} is 0.01 mol/L at 25°C _____

14. The EMF of a circuit composed of saturated calomel and hydrogen electrodes immersed in gastric juice at a temperature of 18°C is equal to 0.332V. The potential of the saturated calomel electrode measured against the normal hydrogen electrode at 18°C is equal to $E = 0.250V$. Calculate the pH of the gastric juice.

17. What is the direction of electron movement in the galvanic cell $Mg / Mg^{2+} // Pb^{2+} / Pb$?

18. Write equations of reactions that occur in a galvanic cell consisting of zinc and silver plates immersed in solutions of their salts with a concentration of cations equal to 1 mol/L.

19. Determine which of the electrodes is the cathode in the galvanic cell formed by standard electrodes: Ag/Ag^+ or Mn/Mn^{2+} ; ($E^0_{Ag|Ag^+} = 0.799\text{ V}$; $E^0_{Mn|Mn^{2+}} = -1.179\text{ V}$)

20. Based on the standard electrode potentials, determine which of the following galvanic cells has the largest EMF: a) $Zn|Zn^{2+} || Ni^{2+}|Ni$; b) $Cd|Cd^{2+} || Ni^{2+}|Ni$

21. Calculate the potential of a zinc electrode in 200 mL of a solution that contains 0.2 g $ZnSO_4$, at the temperature of 298 K. _____

22. Redox electrodes _____

23. The Nernst-Peters equation. Explain _____

24. Standard redox potential _____

25. In what direction will the reaction proceed spontaneously at standard conditions:
 $2NaCl + Fe_2(SO_4)_3 \rightleftharpoons 2FeSO_4 + Cl_2 + Na_2SO_4$, if $\varphi^0_{Cl_2|2Cl} = 1.36\text{V}$, and $\varphi^0_{Fe^{3+}|Fe^{2+}} = 0.771\text{V}$?

Literature: P. 92-95

https://saylordotorg.github.io/text_general-chemistry-principles-patterns-and-applications-v1.0/s23-00-electrochemistry.html

https://saylordotorg.github.io/text_general-chemistry-principles-patterns-and-applications-v1.0/s23-01-describing-electrochemical-cel.html

https://saylordotorg.github.io/text_general-chemistry-principles-patterns-and-applications-v1.0/s23-02-standard-potentials.html

https://saylordotorg.github.io/text_general-chemistry-principles-patterns-and-applications-v1.0/s23-03-comparing-strengths-of-oxidant.html

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PRACTICAL CLASS № 15

SUBJECT: Adsorption processes and ion exchange in biosystems. Chromatography

Importance: Among the processes occurring on the surface of the phase interface in heterogeneous systems, the phenomena of sorption (absorption and adsorption of substances) are of great importance.

The study of sorption of substances allows for adsorption therapy (hemisorption, plasma sorption, lymph sorption, liquor sorption, enterosorption, application therapy), purification of vitamins and antibiotics, application of immobilized preparations (polymer-fixed enzymes, hormones, antibiotics).

Chromatographic analysis is used for diagnosis of various diseases, clinical control of treatment, control of the process of detoxification of the body at poisoning. Chromatography is used in toxicology, chemistry, forensic medicine, forensic science and hygiene.

AIM: Explain principles of adsorption of substances, calculate adsorption capacity and evaluate quantitative characteristics of sorbents, explain physicochemical basis of adsorption therapy methods.

Explain methods of chromatographic analysis and their role in medical and biological research. Be able to separate mixtures using adsorption chromatography and paper chromatography.

I. Theoretical part:

1. Surface phenomena and surface tension. Significance in biology and medicine.
2. Classification of substances according to their ability to change surface tension of water. Surface tension isotherm. Duclos-Traube rule.
3. Orientation of surfactant molecules in the surface layer. Structure of biological membranes.
3. Sorption, types of sorption and their characteristics.
4. Gibbs, Langmuir and Freundlich adsorption isotherms. Their analysis.
5. Adsorption of electrolytes: selective and ion exchange. Panet-Fajans rule. Ion exchangers
6. Physical and chemical principles of adsorption therapy (hemisorption, plasma sorption, lymph sorption, enterosorbition, application therapy). Immunosorbents.
7. What is chromatography? Applications of chromatography in biology and medicine.
8. Classification of chromatography methods by the aggregate state of phases.
9. Classification of chromatography methods by the technique.
10. Classification of chromatography methods by the mechanism of distribution.

II. Practical part

1. Laboratory work

Experiment 1. Adsorption of dyes from solutions on activated charcoal.

Add 5 mL of 0.02% of solutions of iron (III) chloride, potassium dichromate and kongo red to a series of test tubes. Grind finely some activated charcoal in a porcelain mortar. Add 0.2 g of activated charcoal to each test tube, shake for 2-3 minutes and filter. What do you observe?

Has the color (or color intensity) of the solutions (filtrate) changed? Why?

Conclusion: _____

II. Separation of mixtures with chromatography methods

1. Separation of Fe^{3+} and Cu^{2+} cations using adsorption column chromatography

Separation of cations Fe^{3+} and Cu^{2+} is done with adsorption column chromatography. Prepare the chromatography column as following: take a dry glass tube (length 12-15 cm, diameter 1 cm) with a narrow end. Place some cotton wool into the narrow end. Fill the tube with 4-5 cm of aluminium oxide powder, tapping the tube gently to avoid gaps in the powder. Fix the prepared column in a stand. Take 3 mL ferric chloride solution and 3 mL cupric sulfate solution of same molarity, mix them in a test tube. Pour carefully the obtained mixture into the column. Place an empty beaker under the column. After some time, colored layers appear in the column. Draw the column with the colored layers of the cations in the protocol of the practical work. Make a conclusion about dependence of adsorption of cations on aluminium oxide on the charge of the cation.

After all the mixture has gone through the column, wash the adsorbent with a small amount of water and then add a developer into the column for better spectacularity. The developer is a diluted

solution of potassium ferrocyanide $K_4[Fe(CN)_6]$. The top layer becomes dark blue, the bottom one becomes brown.

2. Partition chromatography of amino acids on paper

On one side of a rectangular chromatography paper draw a line at a distance of 1 cm from the edge. Mark and number three points (start points) on the line as follows:

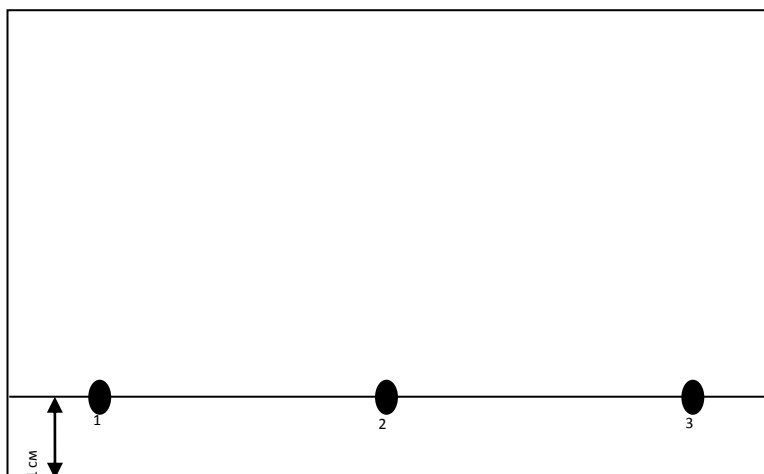


Fig.1 Chromatography paper with the start points

At the start points, place small drops (3 - 4 mm) of solutions of amino acids: glycine, alanine and mixtures thereof. Dry the paper.

Pour the solvent into a glass beaker to a height of 0.5 cm. Immerse the chromatography paper in it, with the start line at the bottom. Leave the paper in the beaker for some time. When the solvent reaches almost the edge of the filter, remove the paper, dry it and treat it with a developer - a solution of ninhydrin. Dry the chromatogram to reveal colored stains of the amino acids.

Draw an obtained chromatogram:

Using the chromatogram, determine the retention factors of amino acids glycine (R_f glycine) and leucine (R_f leucine). To do this, measure the distance from the start point to the middle of the colored spot for each of the amino acids (r_1 and r_2) and the distance from the start point to the solvent front line (r_{solvent}).

Calculate the retention factors (R_f) of the amino acids with the formulas:

$$R_{f \text{ glycine}} = \frac{r_1}{r_{\text{solvent}}}, \quad R_{f \text{ leucine}} = \frac{r_2}{r_{\text{solvent}}},$$

where r_1 is the distance from the start point to the middle of the spot of glycine, cm;

r_2 is the distance from the start point to the middle of the spot of leucine, cm;

r_{solvent} is the distance from the starting point to the front of the solvent, cm.

Make a conclusion about effect of polarity of the molecules on the distribution of amino acids.

Conclusion: _____

2. Control tasks

2. What is a surface tension isotherm? Draw surface tension isotherms for surfactants and surface-inactive substances.

2. Draw a structure of the cell membrane.

3. Formulate principles and assumptions of the Langmuir theory of adsorption, draw a Langmuir adsorption isotherm.

4. What principles of adsorption are described in the Shilov's rule? Give an example.

5. What principles of adsorption are described in the Rebinder's rule? Give an example.

6. Draw a scheme of selective adsorption of ions on the surface of a AgCl crystal, explain it using the Panet-Fajans rule.

7. Write equations of the reactions that characterize the use of ion exchangers to remove NaCl from water.

8. Give examples and explain methods of adsorption therapy. Give an example of applications of adsorbents in medicine.

9. What is the stationary phase in chromatography methods? Give examples. What is the mobile phase in chromatography methods? Give examples..

10. Describe gas-liquid chromatography.

11. How to identify components of a gas mixture in gas chromatography? Draw an example of a chromatogram.

12. How to identify components of a mixture in paper chromatography. Draw an example of a chromatogram. Mark “witnesses” and explain what they are.

13. Describe gel filtration (molecular sieve method). Draw a scheme of it.

14. Describe affinity chromatography. Draw a scheme of it. Give examples of its applications in medicine.

Literature: P. 124-128, 128-131

“Accepted” “ _____ ” _____ 20 _____. _____

PRACTICAL CLASS № 16

SUBJECT: Preparation, purification and properties of colloidal solutions

IMPORTANCE: a study of colloidal solutions and methods of their preparation allows for creation models of cells, cell membranes, nerve fibers all of which are colloidal systems.

A study of the properties of colloidal solutions and methods of their purification is necessary in application of such methods of diagnosis and treatment as electrophoresis, compensatory dialysis, vividialysis and “artificial kidney”.

AIM: be able to analyze principles of methods for preparation and purification of colloidal solutions; prepare colloidal solutions using different methods; compose and write a structure of the micelle; determine the charge sign of the dispersion phase particles.

I. Theoretical part:

1. Classification of dispersion systems by the degree of dispersion.
2. Colloid state. Lyophilic and lyophobic colloidal systems.
3. Structure of colloidal particles.
4. Double electric layer. Electrokinetic potential of the colloidal particle.
5. Methods of preparatiuon of colloidal solutions.
6. Methods of purification of colloidal solutions:

- a) dialysis;
- b) electro dialysis;
- c) compensatory dialysis;
- d) vividialysis;
- e) ultrafiltration;
- e) hemodialysis and “artificial kidney”.

7. Molecular-kinetic properties of colloidal systems (Brownian motion, diffusion, osmotic pressure).

8. Optical properties of colloidal systems.

9. Electrophoresis, its applications in research and clinical laboratory practice. The Helmholtz-Smoluchovsky equation.

II Practical part

1. Laboratory work

Experiment 1. *Preparation of a sol of sulfur with the method of solvent exchange*

Add 1 mL of a solution of sulfur in ethanol (obtained by long-term infusion) to 10 mL distilled water in a test tube. What do you observe? Sulfur forms a true solution in ethanol, but can not dissolve in water thus creating a colloidal system.

Conclusions _____

Experiment 2. *Preparation of a ferric hydroxide sol with the method of hydrolysis*

Add 50 mL distilled water to a conical flask and bring it to a boil. Measure out 5 mL 5% FeCl₃ (with a measuring test tube) and add it gradually to the boiling water.

Write an equation of the hydrolysis reaction of ferric chloride:

Products of hydrolysis partially react with each other to form oxoferrous chloride (FeOCl) and water.

Write an equation of the reaction:

Oxoferrous chloride is the stabilizer of colloidal particles.

Write the formula of the micelle of Fe(OH)₃:

Mark the color of the resulting sol:

Experiment 3. *Preparation of a Prussian blue sol with the method of double replacement*

Determination of the charge of colloidal particles with the method of capillary analysis

Preparation of a Prussian blue sol with the method of double replacement

1) Add 3 mL of K₄[Fe(CN)₆] and 1 mL of FeCl₃ solutions to a test tube.

2) Add 1 mL of K₄[Fe(CN)₆] solution and 3 mL of FeCl₃ solution to another test tube.

A colloidal solution of Prussian blue is obtained. Mark the color, write an equation of the reaction:

1) Write a formula of the micelle of the sol obtained in the first test tube. Explain it.

2) Write a formula of the micelle of the sol obtained in the second test tube. Explain it.

Determination of the charge of colloidal particles by the method of capillary analysis

The method is based on the dependence of adsorption of sol particles on the charge sign of the adsorbent surface. We use filter paper as an adsorbent. When immersed in water, cellulose capillaries of paper become negatively charged, and the water rises up the capillaries in the paper. If there are charged particles in the water, they can only move with the water if they are not attracted to the capillary walls, i.e. when the particles are charged negatively. If the particles are charged positively, they do not move with the water because they become adsorbed on the capillary walls. From this, we can determine the sign of the particle charge.

Using a pipette, place a drop of the sol obtained in the first test tube onto filter paper and observe its behavior.

Then place a drop of the sol obtained in the second test tube onto filter paper and observe its behavior.

If the colored stain moves along with water, the colloidal particles have a negative charge.

If the stain remains in place and water only spreads, the colloidal particles are positively charged.

Draw a position on the filter paper of the stain of the sol obtained in the first test tube, make a conclusion about the charge of its colloidal particles, compare with the charge in the micelle formula that you have written above:

Draw a position on the filter paper of the stain of the sol obtained in the second test tube, make a conclusion about the charge of its colloidal particles, compare with the charge in the micelle formula that you have written above:

Experiment 4. *Preparation of ferric hydroxide sol with the method of peptization*

Add 50 mL distilled water to a flask. Add 2 mL 5 % FeCl_3 solution to the flask. Next, add gradually 5 % ammonia solution until a strong ammonia odor is felt and a precipitate appears.

Write an equation of the reaction of obtaining a precipitate of ferric hydroxide:

After sedimentation of the precipitate carefully decant off the upper layer of the liquid. Add approximately 30 mL distilled water to the precipitate, shake it well, allow it to sedimentate and decant again. Repeat the procedure three times.

Add two small portions of the washed sediment (about 1 mL) in two test tubes. Add 10 mL water to the first test tube and 3 mL water and 2 mL of a 5% FeCl_3 solution to the second test tube.

Write a formula of the micelle of the sol obtained after peptization in the second test tube.

Determine the charge sign of the disperse phase particles: _____

2. Control tasks

1. What disperse systems are called colloid? What is the size of colloidal particles?

3. How is dialysis performed? Draw a diagram of the structure of the dialyzer.

4. How is electro dialysis performed? Draw a diagram of the structure of the electro dialyzer.

4. How is hemodialysis performed? What is “artificial kidney”, what is its purpose?

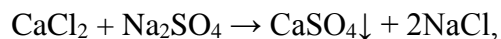
5. Briefly describe properties of colloidal solutions.

7. What is the micelle? What is its structure? What is the aggregate and granula? Explain the structure of the micelle with an example. What are potential-determining ions and counterions?

8. What is the formula of the micelle of a silver chloride sol obtained in a reaction of potassium chloride with an excess of silver nitrate?

9. What is the formula of the micelle of a silver iodide sol obtained in a reaction of silver nitrate with an excess of potassium iodide?

10. Explain the mechanism of formation of micelles of different sols that can be obtained in the reaction:



a) with an excess of CaCl_2 as a stabilizer:

b) with an excess of Na_2SO_4 as a stabilizer:

11. A sol was obtained in mixing 3 mL 0.0007M $\text{Fe}(\text{NO}_3)_3$ and 5 mL 0.01M Na_3PO_4 solutions. Write an equation of the reaction and a formula of the micelle of the resulting colloidal solution. Prove it by performing appropriate calculations.

12. A sol was obtained in mixing 15.0 mL 0.025M KCl and 85.0 mL 0.005M AgNO_3 solutions. Write an equation of the reaction and a formula of the micelle of the resulting colloidal solution. Prove it by performing appropriate calculations.

Determine the sign of the charge of the particles of the dispersed phase: _____

13. Write a formula of the micelle of a silver sol bromide obtained in a reaction of a solution of AgNO_3 with an excess of NaBr . To which electrode will the sol particles move? Explain.

14. With the Helmholtz - Smoluchovsky equation, calculate the rate of movement of colloidal particles of Prussian blue in electrophoresis in water, if the ζ -potential is equal to 0.058 V, the electric field gradient $H = 5 \cdot 10^{-2}$ V/m, viscosity of the dispersion medium $\eta = 1 \cdot 10^{-3}$ Pa·s, dielectric constant $\epsilon = 81$, electric constant $\epsilon_0 = 8.85 \cdot 10^{-12}$ F/m.

Literature: P. 134-151

“Accepted” “ _____ ” _____ 20 _____. _____

PRACTICAL CLASS № 17
SUBJECT: Coagulation of colloidal solutions

Importance: studying physicochemical principles of coagulation of colloidal solutions is necessary to determine the coagulation threshold, which opens up a possibility of preventing coagulation of hydrophobic particles, red blood cells and proteins, preventing formation of blood clots, determination of ESR (erythrocyte sedimentation rate), preservation and storage of blood, creation of modern thromboresistant materials, manufacture of sols of silver as medicines (collargol, protargol, etc).

AIM: explain physicochemical principles of coagulation of colloidal solutions, be able to measure the coagulation threshold.

I. Theoretical part:

1. Stability of disperse systems. Stability factors.
2. Coagulation of hydrophobic colloids. Mechanism of coagulating action of electrolytes.
3. Coagulation threshold, its measurement. The Schulze-Hardy rule.
4. Colloidal protection.
5. Aerosols: classification, methods of preparation, properties. Applications in clinic and sanitary-hygienic practice.

6. Toxic effects of some aerosols. Powders.
7. Suspensions: methods of preparation, properties. Pastes, their medical applications.
8. Emulsions: types, methods of preparation, properties. Applications in clinical practice.
9. Emulsifiers. Biological role of emulsification.
10. Semicolloidal soaps, detergents. Micelle formation in solutions of semicolloids.

II. Practical part

1. Calculation and situation tasks performed by students in the lesson

Task №1

A silver iodide sol is obtained in the reaction: $\text{KI} + \text{AgNO}_3 \rightarrow \text{AgI}\downarrow + \text{KNO}_3$ with KI taken in an excess. The sol is coagulated with solutions of potassium sulfate and calcium acetate. Which electrolyte's coagulating effect is stronger? Write a formula of the micelle of the sol.

Task

№2

A hydrosol of $\text{Al}(\text{OH})_3$ is stabilized with an excess of AlCl_3 . What volume (in liters) of 0.005M K_2CrO_4 solution should be added to 0.05 L of the sol to cause its apparent coagulation? Which ion of the electrolyte has the coagulating effect? The coagulation threshold of $\text{Al}(\text{OH})_3$ is 0.15 mmol/L.

Task № 3

Calculate the coagulation threshold if 4.0 mL of a 0.15M solution of $\text{Al}_2(\text{SO}_4)_3$ was used for coagulation of colloidal particles contained in 230.0 mL of sewage.

2. Control tasks

1. What process is called coagulation? What factors can cause coagulation of a liophobic colloidal system? What happens at the end of the coagulation process?

2. What is the coagulation threshold and coagulating ability of an electrolyte? What does coagulating ability of an electrolyte depend on? Formulate the Schulze-Hardy rule.

3. Write formulas of micelles of the following sols:

a) $\text{Al}(\text{OH})_3$ sol stabilized with AlCl_3 :

b) SiO_2 sol stabilized with H_2SiO_3 :

Which ions of electrolytes cause coagulation of these sols (the Schulze-Hardy rule)?

a)

b)

4. Пороги коагуляції золю електролітами NaCl та CaCl₂ однакові. Який висновок можна зробити щодо заряду колоїдних частинок?

5. Determine the charge sign of sol particles, if in coagulation of the sol with electrolytes the following values of coagulation thresholds (mmol/L) were obtained: $C_{\text{threshold}}(\text{NaCl}) = 300$; $C_{\text{threshold}}(1/2 \text{MgCl}_2) = 320$; $C_{\text{threshold}}(1/3 \text{Na}_3\text{PO}_4) = 0.6$; $C_{\text{threshold}}(1/2 \text{Na}_2\text{SO}_4) = 20$. Explain your reasoning.

6. What is the phenomenon of colloidal protection? What is its significance in biology and pharmacy? Give examples. What is the role of colloidal protection in the living organism?

7. The respective “gold” numbers (in mg) of the macromolecular compounds are as follows: gelatin - 0.01; egg albumin - 2.5; starch - 25.0; hemoglobin - 0.25; saponin - 115. Which of the substances has the highest protective effect for a hydrophobic sol? Explain using the meaning of the “gold” number.

8. What are gels? _____

9. How are emulsions classified according to the concentration of the dispersed phase? Emulsifiers. Biological role of emulsifiers.

Literature: P. 159-163

“Accepted” “ _____ ” _____ 20 _____. _____

PRACTICAL CLASS № 18

SUBJECT: Physicochemical properties of solutions of biopolymers

Importance: studying proteins that are components of all body tissues (proteins catalyze biochemical processes, perform regulatory, contractile, protective, support and structural functions, participate in coagulation of blood, are reserve nutrients, support the oncotic pressure of blood plasma) allows for understanding their physicochemical properties as well as the causes and consequences of

protein-related disorders (e.g. changing isoelectric state of proteins at a change of blood pH), causes and effects of edema. This opens up a possibility of using artificial and synthetic polymers in medicine and pharmacy that can be applied to make living tissue mimics, joint prosthetics, blood plasma substitutes, blood vessels, bones, dentures etc.

AIM: explain physicochemical properties of proteins; measure the isoelectric point of a protein solution with the method of precipitation; draw conclusions about the charge of protein molecules from the isoelectric point value.

I. Theoretical part:

1. High-molecular compounds (polymers) are the chief components of living organisms.
2. Comparative characteristics of solutions of high-molecular compounds, true solutions and colloidal solutions.
3. Mechanism of swelling of polymers, its dependence on various factors. The role of swelling in physiology of organisms.
4. Ionic state of biopolymers in aqueous solutions. Isoelectric state of proteins.
5. Isoelectric point of a protein, methods of its determination.
6. Gelification of solutions of high-molecular compounds, properties of gels.
7. Anomalous viscosity of solutions of high-molecular compounds. Viscosity of blood.
8. Donnan membrane equilibrium.

II. Practical part

1. Laboratory work

Determination of the isoelectric point of gelatin with maximum precipitation

Procedure

In five test tubes, combine solutions of acetic acid (CH_3COOH) and sodium acetate (CH_3COONa) in the amounts indicated in the table so as to obtain 10 mL of acetate buffer solutions with different pH values in each of the five test tubes.

Table

№	Volume of solutions, mL		pH	№ of test tube with maximum precipitation	pH of the solution with maximum precipitation	pI
	CH_3COOH (0.2 mol/L)	CH_3COONa (0.2 mol/L)				
1	9.75	0.25	3.17			
2	8.90	1.10	3.85			
3	5.35	4.65	4.70			
4	1.70	8.30	5.45			
5	0.25	9.75	6.35			

Add 0.5 mL of a gelatin solution to each test tube. Mix thoroughly.

Then add 2 mL of ethanol to each test tube, mix vigorously and leave the test tubes for 10 minutes. Measure the height of the layer of turgescient gelatin in each test tube with a ruler. Indicate in which test tube and at which pH value the maximum precipitation is observed. How can this be explained? Determine the isoelectric point of gelatin.

Fill in the chart with the results.

Conclusions: _____

2. Control tasks

1. Give a definition of high-molecular compounds. Give examples of biopolymers in the human body, indicate their functions.

2. Explain structural features of different classes of high-molecular compounds. Write down their formulas.

3. What is swelling of high-molecular compounds? Describe types of swelling and factors that effect the process. Give examples of swelling in the body.

4. What are causes and mechanisms of formation of “hungry” and “renal” edema?

5. What is the mechanism of edema development at insect bites (ants, wasps, etc.)?

6. Describe the ionic state of biopolymers in aqueous solutions, write a diagram of the structure.

7. What is the isoelectric point (pI) of a protein?

8. Explain determination of the isoelectric point of a protein solution with the method of precipitation.

9. Explain determination of the isoelectric point of a protein solution with the method of electrophoresis.

10. Explain determination of the isoelectric point of a protein solution with the method of swelling.

11. Describe methods of obtaining gels and their specific properties. Give examples of gels in the human body.

12. In the human body, an example of a gel is the skin. Aging of the skin occurs as a result of compaction of the structural network of the gel due to the separation of the liquid phase. What is this process called?

Literature: P. 168-170

“Accepted” “___” _____ 20 ____ . _____

PRACTICAL CLASS № 19

SUBJECT: Solving calculational tasks. Control of practical skills

AIM: to consolidate students’ knowledge of theoretical concepts, improve students’ ability to solve situational and computational problems, control mastery of practical skills acquired by students in practical classes in the discipline "Basics of Medical Chemistry".

II. Practical part:

Types of calculational and situational tasks: see the control tasks for each practical class.

“Accepted” “ ____ ” _____ 20 ____ . _____

PRACTICAL CLASS № 20

SUBJECT: Final control of module “Basics of Medical Chemistry”

AIM: to assess students’ knowledge of theoretical concepts, evaluate students’ ability to solve situational and computational problems in the discipline “Basics of Medical Chemistry”.

LIST OF QUESTIONS

Content module 1. Homogeneous equilibria in biological fluids

1. Electronic structure of biogenic elements. Typical chemical properties of elements and their compounds (reactions without change in degree of oxidation, with change in degree of oxidation, formation of coordination complexes). Relationship between positions of *s*-, *p*-, *d*-block elements in the Periodic System and their content in the body.

2. Solutions of coordination complexes. Modern ideas of structure of coordination complexes. Classification of coordination complexes (by nature of ligands and charge of the internal sphere).
3. Internal coordination complexes (chelate compounds). Polynuclear complexes. Coordination complexes in biological systems. Structure of hemoglobin.
4. Solutions in life. Enthalpy and entropy factors in dissolution and their association with the dissolution mechanism.
5. Solubility of gases in liquids and its dependence on various factors. The Henry-Dalton law. Effect of electrolytes on solubility of gases (the Sechenov's law). Solubility of gases in blood.
6. Solubility of solids and liquids. Distribution of substances between two immiscible liquids. The Nernst law of distribution, its importance in permeability of biological membranes.
7. Equilibrium in solutions of electrolytes.
8. Dissociation of water.
9. Ionic product of water. pH of biological fluids. Types of protolytic reactions. Neutralization, hydrolysis and ionization reactions.
10. Hydrolysis of salts. Degree of hydrolysis, its dependence on concentration and temperature. Hydrolysis constant.
11. Basics of titrimetric analysis. Methods of acid-base titration. Acid-base indicators and principles for their selection.
12. Buffer systems and their classification. pH of buffer solutions.
13. Mechanism of action of buffer systems.
14. Buffer capacity and factors on which it depends. Buffer systems of blood.
15. Colligative properties of dilute solutions: freezing point depression, boiling point elevation. The Raoult's laws. Cryometry and ebulliometry.
16. A colligative property of dilute solutions: osmosis. Osmotic pressure. The van't Hoff's law. Plasmolysis and hemolysis.
17. Colligative properties of dilute solutions of electrolytes. Isotonic coefficient. Hypo-, hyper- and isotonic solutions in medical practice. Role of osmosis in biological systems.

Content module 2. Heterogeneous equilibria in biological fluids

1. ATP as a universal source of energy for biochemical reactions. Macroergic compounds.
2. The First Law of Thermodynamics. Enthalpy. Standard heat of formation and combustion of substances.
3. The Hess law.
4. The Second Law of Thermodynamics. Entropy. Thermodynamic potentials: Gibbs energy, Helmholtz energy. Thermodynamic conditions of equilibrium. Criteria of direction of spontaneous processes.
5. Rate of chemical reactions. Effect of concentration on reaction rate. The law of mass action for rates of chemical reactions. Rate constant.
6. Kinetics of complex reactions: sequential, parallel, coupled, reversible, chain.
7. Order of reaction. Kinetic equations of reactions of first, second and zero order. Half-life period.
8. Effect of temperature on reaction rate. The van't-Hoff's rule. Characteristics of temperature coefficient of reaction rate for biochemical processes.
9. The Arrhenius equation. Activation energy. The theory of active collisions and the theory of transition state.
10. Action of catalysts. Homogeneous, heterogeneous and microheterogeneous catalysis. Acid-base catalysis. Autocatalysis and its role in processes of metabolism.
11. Enzymes as catalysts of biochemical reactions. Features of enzymic action: selectivity, efficiency, dependence of enzymic action on concentrations of enzyme and substrate, temperature and pH of the medium. Mechanism of action of enzymes.
12. Chemical equilibrium. Equilibrium constant, ways of expressing it. Shifts of chemical equilibrium with changes in temperature, pressure, concentration of substances. The Le Chatelier principle.

13. Reactions of precipitation and dissolution. Solubility product.
14. Electrode processes and mechanism of their formation. The Nernst equation. Normal (standard) electrode potential.
15. Standard hydrogen electrode.
16. Measurement of electrode potentials. Indicator electrodes and reference electrodes. The silver - silver chloride electrode. Ion-selective electrodes. The glass electrode.
17. Oxidation-reduction potentials as a measure of oxidizing and reducing capabilities of a system. The Peters equation. Standard oxidation-reduction potentials.
18. Prediction of direction of oxidation-reduction reactions with values of oxidation-reduction potentials.
19. Potentiometric measurement of ion activity. Potentiometric titration.
20. Diffusion and membrane potentials, their biological role. Resting potential. Action potential.
21. Surface phenomena and their significance in biology and medicine. Surfactants and surface-inactive compounds. Surface activity. The Duclos-Traube rule.
22. Adsorption at the liquid-gas and liquid-liquid interfaces. The Gibbs equation. Orientation of surfactant molecules in the surface layer. Structure of biological membranes.
23. Adsorption at the solid-gas interface. The Langmuir equation.
24. Adsorption from solutions on the solid surface. The Freundlich equation.
25. Physicochemical principles of adsorption therapy (hemisorption, plasma sorption, lymph sorption, enterosorption, application therapy).
26. Adsorption of electrolytes (selective and ion exchange). The Paneth-Fajans rule.
27. Natural and synthetic ion exchangers
28. Chromatography. Classification of chromatography methods based on mechanism of distribution of substances, aggregate state of phases and technique. Applications of chromatography in medical and biological studies.
29. Disperse systems and their classification by degree of dispersion. Methods of preparation and purification of colloidal solutions. Dialysis, electrodialysis, ultrafiltration, compensatory dialysis, vividialysis. "Artificial kidney".
30. Structure of colloidal particles.
31. Molecular kinetic properties of colloidal systems. Brownian motion, diffusion, osmotic pressure). Optical properties of colloidal systems.
32. Electrokinetic potential of colloidal particles. The Helmholtz-Smoluchowski equation. Electrophoresis, its use in dentistry and biomedical research.
33. Kinetic (sedimentation) and aggregative stability of disperse systems. Stability factors. Coagulation. Mechanism of coagulating action of electrolytes.
34. Coagulation threshold. The Schulze-Hardy rule. Coagulation processes in treatment of drinking water and wastewater. Colloidal protection.
35. Aerosols, suspensions, emulsions. Preparation, properties, applications. Toxicity of some aerosols. Types of emulsions. Emulsifiers. Biological role of emulsification.
36. High-molecular compounds as chief components of living organisms. Globular and fibrous proteins. Comparative characteristics of solutions of high-molecular compounds, true solutions and colloidal solutions.
37. Swelling and dissolution of polymers. Mechanism of swelling. Effect of pH of the medium, temperature and electrolytes on swelling. Role of swelling in physiology of the body.
38. Gelification of solutions of high-molecular compounds. Thixotropy. Syneresis. Precipitation of biopolymers out of solutions.
39. Anomalous viscosity of solutions of high-molecular compounds. Blood viscosity.
40. The Donnan's membrane equilibrium.
41. Isoelectric state of proteins. Isoelectric point and methods for its determination. Ionic state of biopolymers in aqueous solutions.

ASSESSMENT CRITERIA FOR FINAL MODULE CONTROL

Question 1 (theory) – from 0 to 33.5 points;

Question 2 (practical skills) – from 0 to 33 points;

Task 1 – 5 points;

Task 2 – 5 points;

One multiple choice task – 0.5 points; there are 7 multiple choice tasks.

The final module control is considered to be successfully completed if a student scored at least **50** points. The maximum sum of points in a final module control is **80**.

Students who completed all types of the work previously provided by the curriculum are permitted to take the final module control if their score for classwork is not less than the minimum (**72** points).

The total score for a module is a total of the classwork score and final module control score. The passing score for classwork ranges from 72 to 120 points. Thus, the minimum number of points needed to complete a module is: **72 + 50 = 122 points**. The maximum number is: **120 + 80 = 200 points**.

Assessment criteria for final module controls:

“EXCELLENT” (80-72 points) grade is awarded to a student who produced at least 90% correct answers to standardized test assignments, solved situational tasks without errors and gave comprehensive answers to the theory question. A student demonstrates profound knowledge of the course content, comprehensive theoretical knowledge and practical skills. A student understands importance of the subject, its connection to professionally-oriented and clinical subjects.

“GOOD” (71-64 points) grade is given to a student who produced at least 75% correct answers to standardized test assignments, made some minor mistakes in solving situational problems, gave complete answers to the theory question with minor mistakes. A student demonstrates solid knowledge of the course content, good theoretical knowledge and practical skills. A student understands importance of the subject, its connection to professionally-oriented and clinical subjects.

“SATISFACTORY” (63-50 points) grade is given to a student who produced at least 55% correct answers to standardized test assignments, made errors in solving situational problems, produced an incomplete answer to the theory question or made significant mistakes in it. A student only demonstrates knowledge of basics of the course content, has not mastered practical skills, can not explain connection of medical chemistry to professionally-oriented and clinical subjects.

“UNSATISFACTORY” (less than 50 points) is given to a student who produced less than 55% correct answers to standardized test assignments, made gross mistakes in answering the theory question or did not answer it at all. A student demonstrates lack of systematic knowledge, does not have practical skills, makes fundamental errors in answering the theory question and solving situational problems.

References

Medical Chemistry: textbook / V.O. Kalibabchuk, V.I. Halynska, V.I. Hryshchenko et al.; edited by Prof. V.O. Kalibabchuk – Kyiv: “Medicine”, 2010 – 224 p.

https://saylordotorg.github.io/text_general-chemistry-principles-patterns-and-applications-v1.0/index.html

APPENDICES

SOLUBILITY OF SALTS, ACIDS AND BASES IN WATER

	<i>Cl</i>	<i>OH</i>	<i>NO₃⁻</i>	<i>S²⁻</i>	<i>SO₃²⁻</i>	<i>SO₄²⁻</i>	<i>PO₄³⁻</i>	<i>SiO₃²⁻</i>	<i>CO₃²⁻</i>	<i>CH₃COO⁻</i>
<i>H⁺</i>	S	S	S	S	S	S	S	I	S	S
<i>NH₄⁺</i>	S	S	S	S	S	S	S	–	S	S
<i>Na⁺</i>	S	S	S	S	S	S	S	S	S	S
<i>K⁺</i>	S	S	S	S	S	S	S	S	S	S
<i>Ba²⁺</i>	S	S	S	S	ss	I	I	I	I	S
<i>Ca²⁺</i>	S	S	S	S	ss	ss	I	ss	I	S
<i>Mg²⁺</i>	S	I	S	S	ss	S	I	–	I	S
<i>Al³⁺</i>	S	I	S	I	–	S	I	–	–	S
<i>Fe²⁺</i>	S	I	S	I	ss	S	I	I	I	S
<i>Zn²⁺</i>	S	I	S	I	ss	S	I	I	–	S
<i>Ag⁺</i>	I	–	S	I	ss	ss	I	I	ss	S
<i>Hg²⁺</i>	S	–	S	I	–	–	I	–	I	S
<i>Cu²⁺</i>	S	I	S	I	–	S	I	–	–	S
<i>Pb²⁺</i>	SS	I	S	I	I	I	I	I	I	S

S – soluble
ss – slightly soluble
I – insoluble

CONSTANTS OF DISSOCIATION OF SOME WEAK ELECTROLYTES AT 25°C

Electrolyte	K	Electrolyte	K
Nitric acid HNO ₂	$K=4 \cdot 10^{-4}$	Hydrofluoric acid HF	$K=6.6 \cdot 10^{-4}$
Orthosilicic acid H ₂ SiO ₃	$K_1=2.2 \cdot 10^{-10}$ $K_2=1.6 \cdot 10^{-12}$	Hydrogen cyanide acid HCN	$K=7.9 \cdot 10^{-10}$
Formic acid HCOOH	$K=1.8 \cdot 10^{-4}$	Water H ₂ O	$K=1.8 \cdot 10^{-16}$
Sulfurous acid H ₂ SO ₃	$K_1=1.6 \cdot 10^{-2}$ $K_2=6.3 \cdot 10^{-8}$	Ammonium hydroxide NH ₄ OH	$K=1.8 \cdot 10^{-5}$
Hydrogen sulfide acid H ₂ S	$K_1=6.0 \cdot 10^{-8}$ $K_2=1.0 \cdot 10^{-14}$	Aluminum hydroxide Al(OH) ₃	$K=1.4 \cdot 10^{-9}$
Carbonic acid H ₂ CO ₃	$K_1=4.5 \cdot 10^{-7}$ $K_2=4.7 \cdot 10^{-11}$	Zinc hydroxide Zn(OH) ₂	$K_1=4.4 \cdot 10^{-5}$ $K_2=1.5 \cdot 10^{-9}$
Acetic acid CH ₃ COOH	$K=1.8 \cdot 10^{-5}$	Copper hydroxide (II) Cu(OH) ₂	$K_1=3.4 \cdot 10^{-7}$
Hypochlorous acid HClO	$K=5.0 \cdot 10^{-8}$	Iron hydroxide (II) Fe(OH) ₂	$K_1=1.3 \cdot 10^{-4}$
Orthophosphoric acid H ₃ PO ₄	$K_1=7.5 \cdot 10^{-3}$ $K_2=6.3 \cdot 10^{-8}$ $K_3=1.63 \cdot 10^{-12}$	Iron hydroxide (III) Fe(OH) ₃	$K_1=1.8 \cdot 10^{-11}$ $K_2=1.3 \cdot 10^{-12}$

**SOLUBILITY PRODUCT (SP)
OF SOME COMPOUNDS SPARINGLY SOLUBLE IN WATER**

Compound	SP	Compound	SP
MgF ₂	$7.1 \cdot 10^{-9}$	PbSO ₄	$1.6 \cdot 10^{-8}$
CaF ₂	$3.4 \cdot 10^{-11}$	Ca ₃ (PO ₄) ₂	$1.0 \cdot 10^{-25}$
BaF ₂	$1.7 \cdot 10^{-6}$	Mg(OH) ₂	$1.2 \cdot 10^{-11}$
AgCl	$1.6 \cdot 10^{-10}$	Ca(OH) ₂	$5.5 \cdot 10^{-6}$
AgBr	$4.0 \cdot 10^{-13}$	Cd(OH) ₂	$6.0 \cdot 10^{-15}$
AgJ	$9.7 \cdot 10^{-17}$	Cr(OH) ₃	$5.4 \cdot 10^{-31}$
PbJ ₂	$8.1 \cdot 10^{-9}$	Fe(OH) ₂	$1.7 \cdot 10^{-15}$
MgCO ₃	$2.0 \cdot 10^{-4}$	Fe(OH) ₃	$1.1 \cdot 10^{-36}$
CaCO ₃	$1.2 \cdot 10^{-8}$	Al(OH) ₃	$5.1 \cdot 10^{-33}$
BaCO ₃	$8.1 \cdot 10^{-9}$	CuS	$3.2 \cdot 10^{-38}$
Ag ₂ CrO ₄	$4.1 \cdot 10^{-12}$	Ag ₂ S	$5.7 \cdot 10^{-57}$
CaCrO ₄	$2.3 \cdot 10^{-2}$	CdS	$1.2 \cdot 10^{-28}$
BaCrO ₄	$1.6 \cdot 10^{-10}$	HgS	$4.0 \cdot 10^{-53}$
PbCrO ₄	$1.8 \cdot 10^{-14}$	PbS	$3.6 \cdot 10^{-29}$
CaSO ₄	$6.1 \cdot 10^{-5}$	MnS	$7.0 \cdot 10^{-16}$
BrSO ₄	$2.8 \cdot 10^{-7}$	FeS	$3.7 \cdot 10^{-19}$
BaSO ₄	$1.1 \cdot 10^{-10}$	ZnS	$8.0 \cdot 10^{-26}$

PERIODIC TABLE OF ELEMENTS

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About Chemistry

1A 1 H 1.00794 Hydrogen																	8A 2 He 4.002602 Helium
3 Li 6.941 Lithium	4 Be 9.012182 Beryllium											5 B 10.811 Boron	6 C 12.0107 Carbon	7 N 14.0067 Nitrogen	8 O 15.9994 Oxygen	9 F 18.9984032 Fluorine	10 Ne 20.1797 Neon
11 Na 22.989769 Sodium	12 Mg 24.3050 Magnesium	3B	4B	5B	6B	7B	8B		1B	2B	13 Al 26.9815386 Aluminum	14 Si 28.0855 Silicon	15 P 30.973762 Phosphorus	16 S 32.065 Sulfur	17 Cl 35.453 Chlorine	18 Ar 39.948 Argon	
19 K 39.0983 Potassium	20 Ca 40.078 Calcium	21 Sc 44.955912 Scandium	22 Ti 47.867 Titanium	23 V 50.9415 Vanadium	24 Cr 51.9961 Chromium	25 Mn 54.938045 Manganese	26 Fe 55.845 Iron	27 Co 58.933195 Cobalt	28 Ni 58.6934 Nickel	29 Cu 63.546 Copper	30 Zn 65.38 Zinc	31 Ga 69.723 Gallium	32 Ge 72.64 Germanium	33 As 74.92160 Arsenic	34 Se 78.96 Selenium	35 Br 79.904 Bromine	36 Kr 83.798 Krypton
37 Rb 85.4678 Rubidium	38 Sr 87.62 Strontium	39 Y 88.90585 Yttrium	40 Zr 91.224 Zirconium	41 Nb 92.90638 Niobium	42 Mo 95.96 Molybdenum	43 Tc [98] Technetium	44 Ru 101.07 Ruthenium	45 Rh 102.90550 Rhodium	46 Pd 106.42 Palladium	47 Ag 107.8682 Silver	48 Cd 112.411 Cadmium	49 In 114.818 Indium	50 Sn 118.710 Tin	51 Sb 121.760 Antimony	52 Te 127.60 Tellurium	53 I 126.90447 Iodine	54 Xe 131.293 Xenon
55 Cs 132.9054519 Cesium	56 Ba 137.327 Barium	57-71 Lanthanides	72 Hf 178.49 Hafnium	73 Ta 180.94788 Tantalum	74 W 183.84 Tungsten	75 Re 186.207 Rhenium	76 Os 190.23 Osmium	77 Ir 192.217 Iridium	78 Pt 195.084 Platinum	79 Au 196.966569 Gold	80 Hg 200.59 Mercury	81 Tl 204.3833 Thallium	82 Pb 207.2 Lead	83 Bi 208.98040 Bismuth	84 Po [209] Polonium	85 At [210] Astatine	86 Rn [222] Radon
87 Fr [223] Francium	88 Ra [226] Radium	89-103 Actinides	104 Rf [267] Rutherfordium	105 Db [268] Dubnium	106 Sg [271] Seaborgium	107 Bh [272] Bohrium	108 Hs [270] Hassium	109 Mt [276] Meitnerium	110 Ds [281] Darmstadtium	111 Rg [280] Roentgenium	112 Cn [285] Copernicium	113 Uut [284] Ununtrium	114 Fl [289] Flerovium	115 Uup [288] Ununpentium	116 Lv [293] Livermorium	117 Uus [294] Ununseptium	118 Uuo [294] Ununoctium

Lanthanides	57 La 138.90547 Lanthanum	58 Ce 140.116 Cerium	59 Pr 140.90765 Praseodymium	60 Nd 144.242 Neodymium	61 Pm [145] Promethium	62 Sm 150.36 Samarium	63 Eu 151.964 Europium	64 Gd 157.25 Gadolinium	65 Tb 158.92535 Terbium	66 Dy 162.500 Dysprosium	67 Ho 164.93032 Holmium	68 Er 167.259 Erbium	69 Tm 168.93421 Thulium	70 Yb 173.054 Ytterbium	71 Lu 174.9668 Lutetium
Actinides	89 Ac [227] Actinium	90 Th 232.03806 Thorium	91 Pa 231.03588 Protactinium	92 U 238.02891 Uranium	93 Np [237] Neptunium	94 Pu [244] Plutonium	95 Am [243] Americium	96 Cm [247] Curium	97 Bk [247] Berkelium	98 Cf [251] Californium	99 Es [252] Einsteinium	100 Fm [257] Fermium	101 Md [258] Mendelevium	102 No [259] Nobelium	103 Lr [262] Lawrencium

Alkali Metals	Alkaline Earth	Transition Metal	Basic Metal	Semi Metal	Non Metal	Halogen	Noble Gas	Lanthanides	Actinides
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s-BLOCK, p-BLOCK, d-BLOCK ELEMENTS

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	I	II											III	IV	V	VI	VII	VIII O
I	H																	He
II	Li	Be											B	C	N	O	F	Ne
III	Na	Mg											Al	Si	P	S	Cl	Ar
IV	K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
V	Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
VI	Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
VII	Fr	Ra	Ac	Unq	Unp	Unh	Uns	Uno	Une	Uun								

d - block

p - block

s - block

ELECTRONEGATIVITIES OF ELEMENTS (by Pauling)

0.7 3.98

1																	18	
1	H 2.20																	He
2	Li 0.98	Be 1.57											B 2.04	C 2.55	N 3.04	O 3.44	F 3.98	Ne
3	Na 0.93	Mg 1.31											Al 1.61	Si 1.90	P 2.19	S 2.58	Cl 3.16	Ar
4	K 0.82	Ca 1.00	Sc 1.36	Ti 1.54	V 1.63	Cr 1.66	Mn 1.55	Fe 1.83	Co 1.88	Ni 1.91	Cu 1.90	Zn 1.65	Ga 1.81	Ge 2.01	As 2.18	Se 2.55	Br 2.96	Kr
5	Rb 0.82	Sr 0.95	Y 1.22	Zr 1.33	Nb 1.6	Mo 2.16	Tc 2.10	Ru 2.2	Rh 2.28	Pd 2.20	Ag 1.93	Cd 1.69	In 1.78	Sn 1.96	Sb 2.05	Te 2.1	I 2.66	Xe
6	Cs 0.79	Ba 0.89	La 1.10	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2	Rn
7	Fr 0.7	Ra 0.9	Ac 1.1	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Uub	Uut	Uuq	Uup			
			Lanthanides	Ce 1.12	Pr 1.13	Nd 1.14	Pm	Sm 1.17	Eu	Gd 1.20	Tb	Dy 1.22	Ho 1.23	Er 1.24	Tm 1.25	Yb	Lu 1.0	
			Actinides	Th 1.3	Pa 1.5	U 1.7	Np 1.3	Pu 1.3	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr	

REACTIVITY SERIES OF METALS

Metal	Oxidation Reaction
Lithium	$\text{Li}(s) \longrightarrow \text{Li}^+(aq) + e^-$
Potassium	$\text{K}(s) \longrightarrow \text{K}^+(aq) + e^-$
Barium	$\text{Ba}(s) \longrightarrow \text{Ba}^{2+}(aq) + 2e^-$
Calcium	$\text{Ca}(s) \longrightarrow \text{Ca}^{2+}(aq) + 2e^-$
Sodium	$\text{Na}(s) \longrightarrow \text{Na}^+(aq) + e^-$
Magnesium	$\text{Mg}(s) \longrightarrow \text{Mg}^{2+}(aq) + 2e^-$
Aluminum	$\text{Al}(s) \longrightarrow \text{Al}^{3+}(aq) + 3e^-$
Manganese	$\text{Mn}(s) \longrightarrow \text{Mn}^{2+}(aq) + 2e^-$
Zinc	$\text{Zn}(s) \longrightarrow \text{Zn}^{2+}(aq) + 2e^-$
Chromium	$\text{Cr}(s) \longrightarrow \text{Cr}^{3+}(aq) + 3e^-$
Iron	$\text{Fe}(s) \longrightarrow \text{Fe}^{2+}(aq) + 2e^-$
Cobalt	$\text{Co}(s) \longrightarrow \text{Co}^{2+}(aq) + 2e^-$
Nickel	$\text{Ni}(s) \longrightarrow \text{Ni}^{2+}(aq) + 2e^-$
Tin	$\text{Sn}(s) \longrightarrow \text{Sn}^{2+}(aq) + 2e^-$
Lead	$\text{Pb}(s) \longrightarrow \text{Pb}^{2+}(aq) + 2e^-$
Hydrogen	$\text{H}_2(g) \longrightarrow 2\text{H}^+(aq) + 2e^-$
Copper	$\text{Cu}(s) \longrightarrow \text{Cu}^{2+}(aq) + 2e^-$
Silver	$\text{Ag}(s) \longrightarrow \text{Ag}^+(aq) + e^-$
Mercury	$\text{Hg}(l) \longrightarrow \text{Hg}^{2+}(aq) + 2e^-$
Platinum	$\text{Pt}(s) \longrightarrow \text{Pt}^{2+}(aq) + 2e^-$
Gold	$\text{Au}(s) \longrightarrow \text{Au}^{3+}(aq) + 3e^-$

Ease of oxidation increases

**STANDARD REDUCTION POTENTIALS AT 25⁰ C
FOR SOME COMMON HALF-REACTIONS**

Half-Reaction	E° (V)	Half-Reaction	E° (V)
$F_2 + 2e^- \rightarrow 2F^-$	2.87	$O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$	0.40
$Ag^+ + e^- \rightarrow Ag$	1.99	$Cu^{2+} + 2e^- \rightarrow Cu$	0.34
$Co^{3+} + e^- \rightarrow Co^{2+}$	1.82	$Hg_2Cl_2 + 2e^- \rightarrow 2Hg + 2Cl^-$	0.27
$H_2O_2 + 2H^+ + 2e^- \rightarrow 2H_2O$	1.78	$AgCl + e^- \rightarrow Ag + Cl^-$	0.22
$Ce^{4+} + e^- \rightarrow Ce^{3+}$	1.70	$SO_4^{2-} + 4H^+ + 2e^- \rightarrow H_2SO_3 + H_2O$	0.20
$PbO_2 + 4H^+ + SO_4^{2-} + 2e^- \rightarrow PbSO_4 + 2H_2O$	1.69	$Cu^{2+} + e^- \rightarrow Cu^+$	0.16
$MnO_4^- + 4H^+ + 3e^- \rightarrow MnO_2 + 2H_2O$	1.68	$2H^+ + 2e^- \rightarrow H_2$	0.00
$2e^- + 2H^+ + IO_4^- \rightarrow IO_3^- + H_2O$	1.60	$Fe^{3+} + 3e^- \rightarrow Fe$	-0.036
$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$	1.51	$Pb^{2+} + 2e^- \rightarrow Pb$	-0.13
$Au^{3+} + 3e^- \rightarrow Au$	1.50	$Sn^{2+} + 2e^- \rightarrow Sn$	-0.14
$PbO_2 + 4H^+ + 2e^- \rightarrow Pb^{2+} + 2H_2O$	1.46	$Ni^{2+} + 2e^- \rightarrow Ni$	-0.23
$Cl_2 + 2e^- \rightarrow 2Cl^-$	1.36	$PbSO_4 + 2e^- \rightarrow Pb + SO_4^{2-}$	-0.35
$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$	1.33	$Cd^{2+} + 2e^- \rightarrow Cd$	-0.40
$O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$	1.23	$Fe^{2+} + 2e^- \rightarrow Fe$	-0.44
$MnO_2 + 4H^+ + 2e^- \rightarrow Mn^{2+} + 2H_2O$	1.21	$Cr^{3+} + e^- \rightarrow Cr^{2+}$	-0.50
$IO_3^- + 6H^+ + 5e^- \rightarrow \frac{1}{2}I_2 + 3H_2O$	1.20	$Cr^{3+} + 3e^- \rightarrow Cr$	-0.73
$Br_2 + 2e^- \rightarrow 2Br^-$	1.09	$Zn^{2+} + 2e^- \rightarrow Zn$	-0.76
$VO_2^+ + 2H^+ + e^- \rightarrow VO^{2+} + H_2O$	1.00	$2H_2O + 2e^- \rightarrow H_2 + 2OH^-$	-0.83
$AuCl_4^- + 3e^- \rightarrow Au + 4Cl^-$	0.99	$Mn^{2+} + 2e^- \rightarrow Mn$	-1.18
$NO_3^- + 4H^+ + 3e^- \rightarrow NO + 2H_2O$	0.96	$Al^{3+} + 3e^- \rightarrow Al$	-1.66
$ClO_2 + e^- \rightarrow ClO_2^-$	0.954	$H_2 + 2e^- \rightarrow 2H^-$	-2.23
$2Hg^{2+} + 2e^- \rightarrow Hg_2^{2+}$	0.91	$Mg^{2+} + 2e^- \rightarrow Mg$	-2.37
$Ag^+ + e^- \rightarrow Ag$	0.80	$La^{3+} + 3e^- \rightarrow La$	-2.37
$Hg_2^{2+} + 2e^- \rightarrow 2Hg$	0.80	$Na^+ + e^- \rightarrow Na$	-2.71
$Fe^{3+} + e^- \rightarrow Fe^{2+}$	0.77	$Ca^{2+} + 2e^- \rightarrow Ca$	-2.76
$O_2 + 2H^+ + 2e^- \rightarrow H_2O_2$	0.68	$Ba^{2+} + 2e^- \rightarrow Ba$	-2.90
$MnO_4^- + e^- \rightarrow MnO_4^{2-}$	0.56	$K^+ + e^- \rightarrow K$	-2.92
$I_2 + 2e^- \rightarrow 2I^-$	0.54	$Li^+ + e^- \rightarrow Li$	-3.05
$Cu^+ + e^- \rightarrow Cu$	0.52		