Name of the Department:
Department of Medical Biochemistry, Faculty of Medicine, Semmelweis University

Name of the course: Medical Chemistry
code: AOKOBI001_1A
credits: 6
Director of the course:
Prof. László Tretter M. D., Ph. D., D. Sc.

Description of the curriculum

The principal aim of the course is to prepare students for the understanding of Biochemistry and Molecular Biology. This requires a firm knowledge of the basics of general, organic and inorganic chemistry.

I. General Chemistry

Structure of atoms, ions and molecules. Chemical bonds

Relation of atomic radius, ionization energy, electron affinity and electronegativity to the periodic table. Ionic bond, ion radius, ions. Covalent bonding, $\sigma$ and $\pi$ bonds, hybrid orbitals, hybridization of carbon. Electron pair repulsion, geometry of molecules, bond angle. Molecular orbital theory.
Polar covalent bonds. Molecules composed of more than two atoms. Coordinative bond.

Solutions, laws of aqueous solutions, their biological and medical aspects

Solute, solvent, solution. The solution process. Solubility of ions in water, dissociation.
Enthalpy of hydration. Concentration, % and molar concentration, normality, molality, molar fraction. Saturated solutions. Solubility, partition, solubility product.
Freezing point depression and boiling point elevation of aqueous solutions. Osmotic pressure, dependence on temperature, solute concentration and ionic dissociation.
Biological and medical importance of osmosis.
**Electrolytes**


**Electrochemistry**


**Thermodynamics**


**Chemical kinetics**

Reaction kinetics, rate of reaction, order and molecularity. Half-time of reactions. The van't Hoff rule. Activated complex, transition state, activation energy. The Arrhenius equation. Catalysis, catalysts. Reversible processes, the law of mass action, equilibrium constant and its relation to free energy change. Consecutive reactions, the importance of rate-limiting steps in metabolic processes.

**II. Inorganic chemistry**

**Properties of non-metals**

Group of halogens, their biological significance. Oxygen group, oxygen, free radicals containing oxygen, air, air pollution, ozone. Sulfur, its compounds. The nitrogen group.

Properties of metals

Alkali metals and their compounds. Alkali earth metals and their compounds, the biological significance of calcium and magnesium. Earth metals. Heavy metals and their biological importance. Precious metals. Medically important metals and metal-containing compounds.

III. Organic chemistry

General properties of organic compounds


Classification of hydrocarbons based on their carbon backbone


Functional groups. Classification and chemical characterization of compounds containing various functional groups

Classification of organic compounds according to their functional groups.

I. Halogenated hydrocarbons, their physicochemical properties.
II. Organic compounds containing hydroxyl groups. Classification. Alcohols, physical properties, chemical reactions. Enols and phenols, their chemical reactions. Synthesis of ethers, their reactions.
III. Oxo compounds: classification, nomenclature, physical properties. Chemical reactions of aldehydes and ketones, nucleophilic addition reactions. Condensation reactions of oxo-compounds, oxidation reduction, substitution on the carbon chain.


V. Organic compounds containing sulfur: thiols, thiophenols and thioethers, their synthesis and physicochemical properties.

VI. Organic compounds containing nitrogen: classification, physicochemical properties of nitro compounds. Amines, classification, synthesis, basicity. Important chemical reactions of amines (e.g. Schiff base formations). Amides of carbonic acids.

Lectures and practical lessons

Two lectures and a laboratory lesson (practical) are held every week; schedules can be found in separate uploaded files. Students are expected to keep records and write protocols on the performed experiments (suggested structure: aim of the experiment, applied methods/devices/reactions, results and evaluation). Hand-written protocols might be presented either at the end of the lab lesson or at the beginning of the next practical to the lab teacher. Students might get two points for each acceptable protocol, and points collected this way are added to the total score they achieve in the corresponding midterm exams. Thus, bonus points collected in weeks 2 – 4 (at most 6) are added to the scores of midterm I and those obtained in weeks 7 – 11 (at most 10) to midterm II, respectively. Importantly, these bonus points not only improve your midterm grades but might help you pass the midterm, too.

Requirements for acknowledgement of the semester

(1) Participation in the laboratory practicals is obligatory; students should sign the attendance sheets at the end of the practicals. In case of more than three absences from the practicals for any reason, the semester will not be acknowledged and the student is not going to be allowed to sit for the semifinal exam. Missed practicals can be completed only in the same week at another group; certificate from the host teacher should be presented by the student to the assigned teacher.

(2) It is compulsory to pass both midterm examinations; see next paragraph for details.

Midterm examinations

Two midterm written examinations will be held in weeks 6 and 12 of the semester, respectively, during regular laboratory practicals. Midterm tests consist of four theoretical questions and four problems (calculations). The material of midterm I covers that of lectures given in the first 5 weeks, while midterm II is based on the lecture material of weeks 6-11. Midterm tests will be evaluated by lab teachers.
and marked as 0, 2, 3, 4 or 5. These ‘midterm bonus points’ are added to the scores achieved at the semifinal exam (see below).

**Passing both midterms is a prerequisite to acknowledgement of the semester.**

Failed midterms might be retaken twice.
The first retake is written, comprising four theoretical questions and four calculations. It should be performed in week 7 (retake of midterm I) and week 13 (retake of midterm II), supervised by the student’s own lab teacher.

Students having failed the first retake might sit for the second retake in the last week of the semester. The second retake is an oral exam conducted by an examination committee. Students having failed the first retake of both midterms I and II will be examined in the material of both midterms at the same time.

**Semifinal examination**

Only those students who have fulfilled both acknowledgement criteria, thus obtained an official electronic Neptun signature, are entitled to sit for the semifinal exam.

The semifinal is a written exam that consists of two theoretical parts and a practical exam.

**First theoretical part (50 min):** drawing 10 structures within 20 min (both inorganic and organic, 1 point each), answering two short questions (providing definitions of two ‘important terms’ taken from the topic list; one point each) and solving four chemical calculations (2 points each).

The list of structures to be memorized can be found on the last page of this document. Please note that any inorganic base or salt might be asked that can be formed by combining any cations and anions provided there. Moreover, any normal or branched-chain alkane, alkene or alkyne (up to eight carbon atoms) can be asked such as 2,3-dimethyl-pentane, 3-methyl-2-hexene etc.

**Second theoretical part (80 min):** 40 multiple choice questions (1 point each).

**Lab exam (practical exam) (15 min):** writing an essay on a laboratory experiment performed during the semester (evaluation: 0, unacceptable; 1 point, minor mistakes; 2 points, clear, detailed and correct). Exact quantities (mass, volume of reagents, incubation times etc.) are not expected here.

Therefore, the maximal score is \(20 + 40 + 2 = 62\).

The exam is unsuccessful with
- 10 or less points in part 1, OR
- 20 or less points in part 2, OR
- 0 point from the practical exam.

Students who pass both part 1 AND part 2 but fail the practical essay have to retake only the practical essay when they repeat the semifinal exam. Those who want a better grade are entitled to rewrite the first 2 parts as well; however, risking that they might perform worse.

Students who pass the practical exam but fail either part 1 or part 2 (or both parts) are obliged to retake both theoretical parts but not the practical exam.
In case of successful exams, i.e. when both theoretical units and the practical exam are successfully completed (at least 11, 21 and 1 points are obtained in blocks 1, 2 and the practical essay, respectively), bonus points from the midterms (at most 10) are added to the scores acquired during the exam. Therefore, successful semifinals are evaluated as follows:

- 33-39 points = grade 2 (pass)
- 40-49 points = grade 3 (satisfactory)
- 50-59 points = grade 4 (good)
- 60-72 points = grade 5 (excellent).

It is possible to write the practical essay in week 14, in the first 15 minutes of the last laboratory practical of the semester. Students successfully completing this test (getting 1 or 2 points) are exempted from writing the practical exam at the semifinal exam. It is to note that this is an extra opportunity for passing the practical exam prior to the beginning of the exam period and in case of failure the semifinal exam should proceed as outlined above.

### Competition

Those students who have passed BOTH midterm examinations with a grade of 3 or better are entitled to participate in the competition. Eligible students should sign up at their lab teachers. The competition is organized in week 14 (the exact date and venue will be announced later). It is based on the whole material of the semester and has the same format as the written semifinal except that no lab essays will be asked. Students achieving at least 75% of the maximal score will be exempted from the semifinal exam.

### Exemption from the semifinal exam

Students who learned general, inorganic and organic chemistry at a university level prior to the commencement of their studies at Semmelweis University might sit for an exemption exam that takes place in the middle of September. Students are kindly asked to present their official documents (transcripts with exam results and a detailed syllabus on the courses they completed) to the tutor (Gergely Keszler, EOK building, room 2.132). The exemption exam encompasses parts 1 and 2 of the semifinal (structures, short definitions of important terms, calculations and multiple choice questions); lab essays will not be asked.

### Registration and modification of examination dates:

Electronically, via the Semmelweis University Neptun System. Retakes are not possible within 3 days following the exam.

All our examination rules comply with the official examination regulations of the Semmelweis University.

### Recommended textbooks, manuscripts, handouts:

- **General chemistry:** Ebbing-Gammon: General Chemistry, latest edition
- **Inorganic chemistry:** Tóth: Concise inorganic chemistry for Medical Students (manuscript)
**Laboratory:**  Hrabák: Selected Collection of Chemical Calculations and Biochemical Exercises (2007); Hrabák: Laboratory Manual - Medical Chemistry and Biochemistry (third edition, 2007)

Manuscripts and textbooks can be purchased in the bookshops of Semmelweis Publisher (on the ground floor of the NET and EOK buildings).

**TOPIC LIST AND IMPORTANT TERMS**
GENERAL CHEMISTRY TOPICS (1 – 37)

Note: Chapter numbers correspond to the 9th edition of D.D. Ebbing – S. D. Gammon: General Chemistry (2009)
Important terms are written in italics.

ATOMIC STRUCTURE
(Ebbing: Chapter 7. Quantum theory of the atom)
   Nucleus, electrons, proton, neutron, atomic number, mass number, atomic weight, isotopes, atomic orbitals, principal quantum number, angular momentum quantum number, magnetic quantum number, spin quantum number

ELECTRON CONFIGURATION OF ELEMENTS
(Ebbing: Chapter 8. Electron configurations and periodicity)
2. Electronic structures of atoms: electron configurations and orbital diagrams.
3. Periodic properties of the elements (atomic radius, ionization energy and electron affinity) and the electronic structure of main-group elements.
   Orbital diagram, Pauli exclusion principle, building-up (Aufbau) principle, Hund’s rule, noble gas core, pseudo-noble-gas core.
   Periodic law, effective nuclear charge, first ionization energy, electron affinity; electronegativity
   Skills: Writing the orbital diagram for the ground state of any atom if the mass number is given.

IONIC AND COVALENT BONDING
(Ebbing: Chapter 9. Ionic and Covalent bonding. Chapter 10: Molecular geometry and chemical bonding theory)
4. Formation of ionic bonding and description of ions.
5. The covalent bond. Transition between ionic and covalent bonding.
   Cation, anion, lattice energy, ionic radius.
   Bonding and non-bonding (lone) electron pairs, coordinate covalent bond, octet rule, multiple bonds, polar covalent bond, electronegativity, delocalized bonding, resonance, bond length (bond distance), covalent radius, bond energy.
   Skills: Writing the Lewis-electron-dot symbols and valence-shell electron configurations for the atoms of the second and third periods.

INTERMOLECULAR FORCING
(Ebbing: Chapter 11: States of matter: Liquids and solids/11.5. Intermolecular forces: explaining Liquid properties)
   Dipole-dipole forces, London (dispersion) forces, Van der Waals forces, hydrogen bonding.

CHEMICAL EQUILIBRIA
(Ebbing: Chapter 14: Chemical equilibrium)
7. Chemical equilibria (basic principles): The equilibrium constant. The law of mass action. Predicting the direction of a reaction. Changing the reaction conditions: LeChatelier principle.
Chemical equilibrium, equilibrium constant, law of mass action, homogenous equilibrium, heterogenous equilibrium, reaction quotient, LeChatelier principle.
Exergonic vs. endergonic reactions, reversible vs. irreversible reactions.

CONCENTRATIONS OF ACIDS AND BASES
(Ebbing: Chapter 3: Calculations with chemical formulas and equations: Mass and moles of substance; Chapter 4: Molar concentrations; Diluting solutions.
Chapters 15 and 16: Electrolytes; Acids and Bases, Neutralization. Equivalents and normality)

8. Concentrations (basic principles): The mole concept. Neutralization. Calculation of various concentrations (percentage concentrations, molarity and normality).
Molecular weight, formula weight, molar mass (mole, “mol”), Avogadro’s number, molar mass, mass percentage, molar concentration (molarity, M), titration

ACID-BASE CONCEPTS
(Ebbing: Chapter 15. Acid-base concepts; Chapter 16: Acid-base equilibria; Lecture)

15. Acid-base properties of salt solutions (hydrolysis). Anion-hydrolysis (example: acetate) and cation-hydrolysis (example: ammonium ion). pH of acidic salts (examples: NaHSO_4, NaHCO_3, NaH_2PO_4 and NaHPO_4).

Acid (Arrhenius theory), base (Arrhenius theory), self ionization of water, ion-product constant for water (water product, Kw), pH, pOH, the pH scale; acid-base titration curve, equivalence point.
Acid (Bronnsted-Lowry theory), base (Bronnsted-Lowry theory), conjugate acid-base pair, Lewis-acid, Lewis base.
Acid ionization (dissociation) constant, base ionization (dissociation) constant, degree of ionization.
common ion effect, buffer, Henderson-Hasselbalch equation
acid-base indicators, buffer capacity (acid capacity and base capacity
Skills: Drawing the titration curves of strong and weak (monoprotic and polyprotic) acids/bases

(Lecture): Intracellular and extracellular buffer systems of the body, average charge of phosphoric acid at various pH, components of the bicarbonate buffer in the blood, role of the ventilation in pH stabilization role of the red blood cells in pH stabilization, role of the kidney in pH stabilization, metabolic acidosis, metabolic alkalosis, respiratory acidosis, total acidity of the urine; anion-hydrolysis (example), cation-hydrolysis (example), cation
and anion hydrolysis (example), acidic salts with acidic pH (example), acidic salts with basic pH (example)

SOLUBILITY AND COMPLEX IONS
(Ebbing: Chapter 17: Solubility and complex-ion equilibria)


18. Unidentate, bidentate, ambidentate and polydentate ligands in complexes. Chelate complexes, complexometric titration. EDTA and biological complexes (heme, vitamin B12, calmodulin, EF hand). Elimination of heavy metal ions from the body.

Solubility, solubility product constant ($K_{sp}$), ion product (Q), conditions for precipitation
Complex salts, double salts, ligands, central ions, coordination number of complexes, unidentate-, bidentate-, ambidentate- and multidentate ligands (examples), chelate complexes (examples), Lewis acid-base theory, geometric isomerism, chiral isomerism, crystal field theory, high and low spin complex; structure of EDTA, biological complexes of iron and calcium, EF hand protein motif

SOLUTIONS
(Ebbing: Chapter 11: States of Matter; Liquids and Solids; Chapter 12: Solutions)


23. Boiling point and freezing point of solutions. Molal freezing point depression and boiling point elevation of aqueous solutions. Colligative properties. Anomalous behavior of ionic solutions, interionic attractions, van’t Hoff factor. Formula mass of ionic compounds. Determination of concentration or molar mass by freezing point depression measurements.

Change of state (phase transition), melting, freezing, vaporization, sublimation, condensation, vapor pressure, boiling point, freezing point, heat of vaporization, phase diagram, surface tension

Solute, solvent, hydration of ions, Lugol solution, Henry’s law, Bunsen (absorption) coefficient, colligative properties, molality, mole fraction, vapor-pressure lowering, Raoult’s law, boiling-point elevation, freezing point depression, osmosis, osmotic pressure, isotonic-, hypertonic-, hypotonic solutions

Partial pressure of gases, ppm, decompression sickness, artificial air

THERMODYNAMICS

(Ebbing: Chapter 6: Thermochemistry; Chapter 18: Thermodynamics and Equilibrium)

26. Entropy change, spontaneous and reversible processes, the second law of thermodynamics. The 2nd law of thermodynamics, absolute and standard entropies.
28. Entropy change, spontaneous and reversible processes, the II. law of thermodynamics. The 3rd law of thermodynamics, absolute and standard entropies.

Internal energy, work, heat, enthalpy change, standard enthalpy change.

Standard enthalpy of fusion/vaporization/sublimation/solution/solvation, lattice enthalpy, molar heat capacity.

Standard enthalpy of formation/combustion, average bond enthalpy.

Entropy change, standard and absolute entropy.; Standard free enthalpy change, exothermic-, endothermic-, exergonic-, endergonic reactions.

First, second and third laws of thermodynamics.

REACTION KINETICS

(Ebbing: Chapter 13: Rates of reaction)

30. Spontaneity and speed of chemical reactions. Reaction rate. Rate equation, rate law. Rate constant and its unit, initial rate. Collision and transition state theories of the mechanism of chemical reactions.
32. Reaction rate and temperature. Activation energy. Potential energy diagrams. Catalysis. Enzymes as biocatalysts; strong specificity of enzymes

reaction rate, rate law, rate constant, rate equation, initial rate, collision theory, transition-state theory, frequency factor, reaction order, molecularity, reaction mechanism, mono- bi- and termolecular reactions, first, pseudo-first, second, zero orders, overall order of a reaction, rate determining step, half-life.
catalysis, catalyst, activation energy, activated complex, Arrhenius equation; energy diagrams of catalysed and non-catalysed reactions, homogeneous and heterogeneous catalysis, chemisorption, enzyme, substrate, stereospecificity

ELECTROCHEMISTRY
(Ebbing: Chapter 19: Electrochemistry; Lectures)
33. Voltaic cells: Notation for a voltaic cell. Electrode potentials (reduction potentials) and the electromotive force. Normal and standard electrode potential. Calculation of equilibrium constants from the electromotive force.
34. Dependence of electrode potentials on concentrations: the Nernst equation. Concentration cells. The hydrogen electrode. Measurement of pH, the glass electrode.
36. Direction of redox reactions. Biologically important redox systems (examples for reversible and for irreversible redox reactions).
37. Specific and equivalent conductance. Determination of the degree of dissociation and the ionization constant by conductometry. (Practice book and lecture)
Voltaic (galvanic) cell, half cell, salt bridge, electromotive force, standard electrode potential, Nernst equation, concentration cell, hydrogen electrode, glass electrode, non- polarizable electrodes, calomel electrode, silver electrode; specific and equivalent conductance; Daniell element; non-polarizable electrode

INORGANIC CHEMISTRY TOPICS (1 – 14)

Important terms are written in italics.

1. Alkali and alkaline earth metals and their compounds.
structure of sodium and potassium chloride, hydroxide, alkali and alkali earth metal ions, structure of magnesium and calcium chloride, sulfate and carbonate, role of calcium in biological systems, structure and utilization of barium sulfate

2. Boron and aluminium family metals. Arsenic, antimony, bismuth and their compounds.
boric acid as a Lewis acid, Amphoteric hydroxides. Double salts of aluminium. Poisonous property of arsenic.

different hybridization of diamond and graphite, coordinative bond in CO, CO as a poison, structure of carbon dioxide, green house gases, equilibrium of carbonic acid, hardness of water caused by alkali earth metal hydrocarbonates; cyanides as poisonous compounds.

4. Silicon and derivatives. Tin and lead and their compounds.
silicon as semiconductor, poisonous effects of lead, removal of lead ions by EDTA, different oxidation states of Sn and Pb


6. Phosphorus and its compounds: allotropes, oxides, oxiacids, phosphates. different phosphoric acids, biological role of phosphates


8. Properties of water. surface tension, maximal density at 4°C, hydrogen bondings and their role in the high boiling point, constant and removable hardness of water

9. Sulfur and its compounds: allotropes, oxides, oxiacids, sulfides, sulfites, sulfates, thiosulfates. structures of sulfide, sulfite, sulfate, thiosulfate ions, practical aspects of the dilution of sulfuric acid

10. Characteristics of halogens. Fluorine, bromine, iodine and their compounds. electron configuration of halogens, H-bond formation of fluorine in compounds, fluorine in teeth, structures of the oxyanions of bromine and iodine, Lugol solution, reaction of iodine with starch, principles of iodometry

11. Chlorine and its compounds. structures of oxyanions of chlorine, formation of NaOCl, properties of HCl and NaCl


13. Transition elements. Manganese, iron, cobalt and their compounds. Copper, zinc, mercury and their compounds. Precious metals. role of KMnO4, different oxidation states of iron, organic iron compounds, poisonous effect of heavy metals, photosensitivity of silver halogenides, utilization of platinum electrodes

14. Nomenclature of inorganic compounds. system of the endings of differently oxidized salts of inorganic acids, nomenclature of acidic and basic salts, names of compounds containing more identical atoms or ions.
COVALENT BONDING IN ORGANIC COMPOUNDS
(Chapter 2)

1. The central role of carbon atoms in organic chemistry. Chemical bonds. Hybridization of atomic orbitals, the hybrid states of carbon, resonance and delocalization in organic compounds.

sp, sp2, sp3 hybridization, promotion of carbon, aromatic compound, antiaromatic compounds, benzenoid compound

DIPOLE MOMENTUM AND GENERAL ACID-BASE PROPERTIES OF ORGANIC COMPOUNDS
(Chapter 3)


Dipole momentum, Debye unit, polar covalent bond, resonance structure, resonance energy, ring strain, torsional strain

THE STERIC STRUCTURE OF ORGANIC MOLECULES. ISOMERISM AND TERMINOLOGY.
(Chapter 4)

3. Principles of constitution, configuration and conformation isomerism
4. Types of constitution isomerism: branching (backbone) isomerism, position isomerism and tautomerism.
5. Configuration in organic chemistry: geometric (cis-trans) and optical (stereo) isomerism. Chirality and prochirality, stereogeneic (chiral) centers, enantiomers and diastereomers. Racemic mixtures and meso compounds.
6. Terminology of chiral compounds: relative and absolute configuration, the D/L and the R/S systems. Stereochemical numbering.
7. Conformation in organic chemistry

CLASSIFICATION OF ORGANIC COMPOUNDS
(Chapter 5)
8. Classification of organic compounds according to the main functional groups
9. Reaction types and reaction mechanisms in organic chemistry.

**SN1 reaction, SN2 reaction, functional group, homologous series, homolytic bond breaking, heterolytic bond breaking, nucleophile reactant, electrophile reactant, electrophilic, nucleophilic, radical reactions, addition, substitution, elimination, Markownikow rule, 1-1 example for fundamental reaction types in organic chemistry (e.g. nucleophile addition), rearrangement reactions, regioselective reaction**

**MAJOR FUNCTIONAL GROUPS AND THEIR REACTIONS**

(Chapter 6)

10. Structure and reactions of alkanes: nomenclature, conformational analysis, radical reactions.
13. Structure and reactions of homoaromatic compounds. Benzene and polycyclic compounds. Resonance stabilization in aromatic compounds and the Huckel’s rule.
14. Mechanism of electrophilic substitution of aromatic compounds. Effect of substituents of the aromatic ring on the reaction rate and product formation in further electrophilic substitution type of reactions.
15. Classification, structure, physical and chemical properties and reactions of organic hydroxyl compounds (alcohols, enols, phenols). Formation of ethers and esters.
16. Classification and nomenclature of ethers (epoxides, hemiacetals, acetals). Electronic structure of open-chain and cyclic ethers; physical properties, miscibility, chemical reactivity: coordinative bonding, basicity, Zeisel test, nucleophilic attack of epoxydes, peroxyethers.
17. Structure, nomenclature, chemical-physical properties, biological role and characteristic reactions of carbonyl compounds (aldehydes, ketones). Important nucleophilic addition reactions (addition of simple inorganic molecules; dimerisation, polymerisation, aldol formation, acetal formation, formation of ketimines, oximes, hydrazones, Schiff’s bases).
20. Organic thio-compounds. Thiaoalcohols, thioethers, sulfinic and sulfonic acids.
dehydration, epoxide, hemiacetal, acetal, Zeisel test, peroxycethers, thiol, thion, thioether, disulfide, sulfide, sulfoxide, sulfone, sulfonate, carbonyl group, aldehyde, ketone, quinone, β-unsaturated carbonyl compounds, amphoterism, desmotropism, protective groups, dimerization, paraformaldehyde, glycosides, furanose, pyranose, Schiff-base, hydrazine, oxime, and mild oxidation of oxo-compounds; Tollens test; Fehling test; Cannizzaro reaction
ester, anhydride, amide, halogenide, azide, aminoacid, fatty acid, carboxylate, dimerization of carboxylic acids, nucleophilic acylation, alpha substitution, driving forces of esterification, cyclic esters, soaps, transesterification, transamination, polyesters nitroso, nitro, oxime, amine, imine, amide, nitrile, isonitrile, cyanate, isocyanate, imide, hydrazine, hydrazide, azo, electron distribution within the most important functional groups: acid-base properties (primary, secondary... amines), conjugation effect, amphoterism of imidazole, structure of the amide bond, restricted rotation and isomerism, tautomerism of nucleotide bases, Schiff-base formation, isocyanate formation

**TOPICS FOR THE LAB EXAM (1 – 19)**

1. The factor of titrating solutions; factorization of HCl
2. The factor of titrating solutions; factorization of NaOH
3. Titration of strong acids with NaOH
4. Titration of acetic acid with NaOH
5. Titration of gastric fluid
6. Principles of the electrometric titration of phosphoric acid and plotting the titration curve
7. Determination of Cl− concentration by means of precipitation titration
8. Permanganometry: principles, factorization of the titrating solution
9. Permanganometry: determination of Fe²⁺ concentration
10. Iodometry: principles, factorization of titrating solution
11. Iodometry: principles, determination of sodium hypochlorite concentration
12. Complexometric titration: determination of unknown Cu²⁺ concentration
13. Complexometric titration: determination of Ca²⁺ and Mg²⁺ concentration of the same solution
14. Conductometry: description of the conductometer, determination of the cell constant
15. Determination of the ionization constant of acetic acid by conductometry
16. Spectrophotometry: determination of the absorption spectrum of phenol red and plotting the calibration curve of the dissociated phenol red anion
17. Spectrophotometric determination of the ionization constant of phenol red
18. Electrochemistry: measurement of the electromotive force of the Daniell element; studying the effect of electrolyte concentration on the electromotive force
19. Electrochemistry: experiments with iron redox electrodes as well as with redox systems of biological relevance
The 10 structures asked at the semifinal exam will be selected from the following list

**Inorganic acids and other compounds:** sulfuric acid, sulfurous acid, nitric acid, nitrous acid, hydrochloric acid, hydrobromic acid, hypochlorous acid, chlorous acid, chloric acid, perchloric acid, hypobromous acid, bromous acid, bromic acid, perbromic acid, hydrogen cyanide, metaphosphoric acid, orthophosphoric acid, boric acid, carbonic acid, water, ammonia, hydrazine, hydroxylamine, hydrogen peroxide, superoxide anion, pyrophosphate anion, hydrogen sulfide, carbon monoxide, carbon dioxide, nitrous oxide, nitric oxide, sulfur dioxide, sulfur trioxide, hydroxyapatite, fluoroapatite, ferrous ammonium sulfate

**Any inorganic salts and bases consisting of the following cations and anions:**

- **Cations:** ammonium, sodium, potassium, magnesium, calcium, ferrous, ferric, cuprous, cupric, zinc, silver, aluminium, mercurous, mercuric, manganese
- **Anions:** hydroxide, oxide, fluoride, chloride, bromide, sulfide, sulfate, sulfite, hydrogen sulfate, thiosulfate, nitrate, nitrite, hypochlorite, chlorite, chlorate, perchlorate, hypobromite, bromite, bromate, perbromate, cyanide, phosphate, monohydrogen phosphate, dihydrogen phosphate, carbonate, hydrogen carbonate (bicarbonate), permanganate, chromate, ferricyanide

**Hydrocarbons:** alkanes, alkenes and alkynes (up to carbon number 8, both normal- and branched-chain isomers); 1,3-butadiene, 2-methyl-1,3-butadiene (isoprene)

**Aromatic rings:** benzene, naphthalene, phenanthrene, pyrrole, thiophene, furane, thiazole, oxazole, imidazole, pyrazole, pyridine, pyrane, pyrimidine, purine, indole, pteridine, acridine

**Small organic compounds:** methanol, ethanol, propanol, isopropanol, n-butanol, ethylene glycol, glycerol, inositol, phenol, diethylether, formaldehyde, acetaldehyde, acetone, mercaptoethanol, aniline, urea, guanidine

**Organic acids:** formic acid, acetic acid, propionic acid, butyric acid, valeric acid, caproic acid, oxalic acid, malonic acid, succinic acid, glutaric acid, maleic acid, fumaric acid, lactic acid, β-hydroxybutyric acid, pyruvic acid, acetoacetic acid, citric acid, cis-aconitic acid, isocitric acid, α-ketoglutaric acid, malic acid, oxaloacetic acid

**Types of bondings and derivatives:** ether, phenolether, thoether, ester, lactone, thioester, anhydride (including mixed and phosphoric acid anhydrides), hemiacetal, hemiketal (cyclic forms included), Schiff-base, oxime, hydrazone, hydroxamic acid, amide, thiol, sulfinic acid, sulfonic acid, sulfoxide, acyl chloride.