



MINISTRY OF EDUCATION AND SCIENCE OF THE RUSSIAN FEDERATION
Federal state autonomous educational institution
of higher education
«Far Eastern Federal University»
(FEFU)

SCHOOL OF BIOMEDICINE

«AGREED»

Head of education program
«General medicine»



(signature) Khotimchenko Yu.S.
(Full name)
«09» of July 2019

«APPROVED»

Director of the Department of Clinical
Medicine



(signature) Geltser B.I.
(Full name)
«09» of July 2019



WORKING PROGRAM OF ACADEMIC DISCIPLINE (WPAD)

«Chemistry, Medical Chemistry»

Educational program

Specialty 31.05.01 «General medicine»

Form of study: full time

year 1 semester 1,2
lectures 36 hours
practical classes 36 hours
laboratory works 54 hours
total amount of in-classroom work 126 hours
independent self-work 90 hours
including exam preparation 36 hours
control works ()
credit 1 semester
exam 1 year, 2 semester

The working program is drawn up in accordance with the requirements of the Federal state educational standard of higher education (level of training), approved by the order of the Ministry of education and science of the Russian Federation from 09.02.2016 № 95.

The working program of the discipline was discussed at the meeting of the Department of fundamental and clinical medicine. Protocol No. 8, 09 of July 2019

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ABSTRACT

The discipline "Chemistry, Medical Chemistry" is intended for students enrolled in the educational program of higher education on 31.05.01 "General Medicine", is included in the basic part of the curriculum, is implemented on the 1st year in the 1st and 2nd semester. The total complexity of the discipline is 216 hours, 6 credits.

In developing the work program of the discipline, the Federal State Educational Standard of Higher Education in the specialty 31.05.01 "General Medicine" (specialty level) has been used.

The course program is based on the basic knowledge gained by students:
the ability to abstract thinking, analysis, synthesis (CC-1)

The content of the discipline covers a range of issues related to the study of the laws of thermodynamics and bioenergy, colligative properties of solutions, ionic equilibria, electrochemistry, chemical kinetics and catalysis, organic chemistry, analytical chemistry and physical and chemical methods of analysis. Mastering the discipline "Chemistry" is necessary for the subsequent study of such disciplines as "Pharmacology", "Medical Biotechnology" and "Biochemistry".

Purpose of studying the discipline is to master the future specialists in the basics of chemical and physicochemical knowledge, which are necessary for the study of processes occurring in a living organism, when they become qualitatively new physiological phenomena.

Objectives of the discipline:

- Master the skills of conducting scientific research to establish the relationship between the physicochemical properties of substances and their pharmacological activity. To study the basic laws of chemical kinetics and thermodynamics in order to determine the possibility of the occurrence and direction of bioenergy processes;

- Be able to apply the laws of chemical kinetics to increase the speed of the main and blocking side processes;

- To be able to apply physical and chemical methods for analytical and environmental purposes.

- Learn how to use the methods of inorganic, physical, analytical and organic chemistry to solve specific problems of biology and medicine.

As a result of studying this discipline, the students form the following professional competencies (elements of competencies).

Code and formulation of competence	Competence formulation phase	
the readiness to use basic physical and chemical, mathematical and other natural science concepts and methods in solving professional problems (CPC – 7)	Knowing	the basic computer databases about the structure and properties of organic compounds, including chemical and 3D computer graphics program
	Be able	using the rules of construction of chemical formulas, graphs, tables, using appropriate computer programs, including for the creation of computer presentations.
	Be master (skill)	by computer programs to build chemical and stereochemical formulas of organic compounds and other illustrative material.
the capacity for the assessment of morphological and physiological states and pathological processes in the human body for solving professional tasks (CPC – 9)	Knowing	physical and chemical basis of pathological processes in Human body; structure, significance and role of basic biogenic elements and they compounds in living systems; chemical methods of estimation of morph-functional and physiological states of living systems
	Be able	apply chemical concepts, laws and principals for estimation of morph-functional and physiological states, and of pathological processes in Human body
	Be master (skill)	by methods of estimation of morph-functional and physiological states of Human body including physical and chemical analysis methods

To form the above competencies will be applying the following methods of interactive learning: active reading, problem lectures, debriefing.

I. STRUCTURE AND CONTENT OF THEORETICAL PART OF COURSE

Unit I. General and inorganic Chemistry (18 hrs.)

Part 1. Base of General Chemistry (12 hrs.)

Item 1. Introduction to Chemistry. Periodic Table and Periodic trends. Energy changes in chemical reactions (2 hrs.).

Chemistry in the Modern World. The Atom. Isotopes and Atomic Masses. Introduction to the Periodic Table. Sizes of Atoms and Ions. Energetics of Ion Formation. The Chemical Families. Trace Elements in Biological Systems. The First Law of Thermodynamics. The Second Law of Thermodynamics. Entropy

Changes and the Third Law of Thermodynamics. Free Energy. Spontaneity and Equilibrium

Item 2. Basic chemical calculations (2 hrs.).

Mole, Avogadro's number. Molar mass. Units of Concentration. Substance concentration. Mass concentration. Per cent concentration. Equivalent Law. Calculations. Volumetric Analysis.

Item 3. Chemical kinetics. Chemical Equilibrium (2 hrs.).

Factors that affect reaction rates. Reaction rates and rate laws. Methods of determining reaction order. Using graphs to determine rate laws, rate constants, and reaction orders. The collision model of chemical kinetics. The concept of chemical equilibrium The equilibrium constant. Solving equilibrium problems. Non-equilibrium conditions. Factors that affect equilibrium. Controlling the products of reactions.

Item 4. Chemical bonds and structure of chemical compounds. The doctrine of the solutions. (2 hrs.).

Predicting the Geometry of Molecules and Polyatomic Ions. Localized Bonding and Hybrid Atomic Orbitals. Delocalized Bonding and Molecular Orbitals. Polyatomic Systems with Multiple Bonds. Aqueous Solutions. Stoichiometry of Reactions in Solution. Ionic Equations. Precipitation Reactions. Acid-Base Reactions.

Item 5. Electrolytes. Hydrolysis salts. Buffer system (2 hrs.).

The dissociation degree and the dissociation constant. pH and pOH of solutions. pH calculations. Aqua Hydrolysis salts. Factors that affect hydrolysis. Hydrolysis constant and Degree of hydrolysis. Calculations.

Item 6. Oxidation-Reduction reactions. Solubility and Complexation compounds. (2 hrs.).

Oxidation–Reduction Reactions in Solution. Factors Affecting Solution Formation. Effects of Temperature and Pressure on Solubility. Determining the Solubility of Ionic Compounds. Factors that affect Solubility. The Formation of Complex Ions. Solubility. Qualitative Analysis Using Selective Precipitation

Part 2. Inorganic Chemistry (6 hrs.)

Item 7. Periodic Trends and s-Block Elements (2 hrs.).

Overview of Periodic Trends. The Chemistry of Hydrogen. The Alkali Metals (Group 1). The Alkaline Earth Metals (Group 2). The s-Block Elements in Biology

Item 8. The p-Block Elements (2 hrs.).

The Elements of Group 13. The Elements of Group 14. The Elements of Group 15 (The Pnictogens). The Elements of Group 16 (The Chalcogens). The Elements of Group 17 (The Halogens). The Elements of Group 18 (The Noble Gases)

Item 9. The d-Block Elements (2 hrs.).

General Trends among the Transition Metals. A Brief Survey of Transition-Metal Chemistry. Metallurgy. Coordination Compounds. Crystal Field Theory. Transition Metals in Biology

Unit II. Organic Chemistry (18 hrs.)

Part 1. A Review of General Organic Chemistry: Electrons, Bonds, and Molecular Properties (4 hrs.).

Item 1. Structure of Organic Compounds. Classification of Organic Compounds (2 hrs.).

Organic Chemistry in the Modern World. Organic Molecules and Chemical Bonding. Types of Bonds. Orbitals and Molecular Shapes. Resonance Structures.

Classification of Organic Compounds. Functional Groups and Organic Nomenclature. Structural Formulas. Isomers.

Item 2. Introduction to Organic Reactivity and Catalysis. Acidity and Basicity of Organic Compounds (2 hrs.).

Electronic Effects. Steric Effects. Introduction to Reaction Mechanisms. Reaction Classification. Acidity and Basicity of Organic Compounds. The Brønsted-Lowry Definition of Acidity. The Lewis Definition of Acidity.

Part 2. Properties of Organic Compounds (8 hrs.).

Item 1. Hydrocarbons (2 hrs.).

Classes of Hydrocarbon. Aliphatic Hydrocarbons: Alkanes, Alkenes, Alkynes, Cycloalkanes. Nomenclature. Structure and Reactivity. Conformations of Cycloalkanes. Aromatic compounds. The Hückel Rule. Aromaticity. Benzene. Structure and Reactivity.

Item 2. Alcohols and Phenols (2 hrs.).

The Hydroxyl Group. Nomenclature of Alcohols. Physical properties of Alcohols. Chemical Properties of Alcohols: Acid-Base Reactions, Substitution Reactions, Dehydration, Oxidation. Phenols: Physical and Chemical Properties.

Item 3. Aldehydes and Ketones (2 hrs.).

The Carbonyl Group. Nomenclature of Aldehydes and Ketones. Physical properties of Aldehydes and Ketones. Chemical Properties of Aldehydes and Ketones: Oxidation-Reduction Reactions of Carbonyl Compounds, Addition Reactions, Formation of Acetals and Ketals, Reactivity of the α -carbon Atom, the Aldol Condensation.

Item 4. Carboxylic Acid and Carboxylic Acid Derivatives. Heterofunctional Organic Compounds (2 hrs.).

Carboxylic Acid and Acyl Groups. Nomenclature of Carboxylic Acid . Physical properties of Carboxylic Acid. Chemical Properties of Carboxylic Acid: Acidity, Nucleophilic Acyl Substitution, Reduction Acyl Derivatives. Heterofunctional Organic Compounds. Hydroxy acids. Configuration of molecules. Mirror Images and Chirality. Optical Activity. Fischer Projection Formulas.

Part 3. Biomolecules (6 hrs.).

Item 1. Carbohydrates (2 hrs.).

Classification of Carbohydrates. Monosaccharides. Chirality of Carbohydrates. Glucose. Fructose. Hemiacetals and Hemiketals. The Configuration of Glucose. Reduction of Monosaccharides. Oxidation of Monosaccharides. Glucosides. Disaccharides. Polysaccharides.

Item 2. Amino Acids, Peptides and Proteins (2 hrs.).

Amino Acids. Acid-Base Properties of α -amino Acids. Isoionic Point. Peptides. Protein. Protein Structure.

Item 3. Lipids (2 hrs.).

Classification of Lipids. Fatty Acids. Types of Fatty Acids. Glycerol. Triglycerides. Phosphoglycerates. Waxes. Glycolipids. Sterols.

II. STRUCTURE AND CONTENT OF PRACTICAL PART OF COURSE

Practical Classes (36 hrs.)

Unit I. General and inorganic Chemistry (18 hrs.)

Class 1. Mendeleev's Periodic Law and Periodic Table (2 hrs.)

1. Structures of Atoms and Ions.
2. Periodic trends
3. Energetics of Ion Formation.
4. The Chemical Families.

Class 2. Basic chemical calculations (2 hrs.)

1. Units of Concentration.
2. Substance concentration.
3. Mass concentration.
4. Per cent concentration.
5. Equivalent Law.

Class 3. Classes of inorganic compounds (2 hrs.)

1. Inorganic Nomenclature
2. Basic chemical properties of acids, bases, salts, oxides
3. Relationship between different classes of chemical compounds

Class 4. Chemical bonds and structure of chemical compounds (2 hrs.)

1. Localized bonding and hybrid atomic orbitals.
2. sp -, sp^2 -, sp^3 -hybridizations
3. Delocalized bonding and molecular orbitals.

Class 5. Chemical kinetics (2 hrs.)

1. Reaction rates and rate laws.
2. Methods of determining reaction order.
3. Using graphs to determine rate laws, rate constants, and reaction orders.
4. The equilibrium constant.

Class 6. pH. Buffers. (2 hrs.)

1. pH and pOH of solutions.
2. pH calculations.
3. Type of buffers systems
4. Calculations of buffers capacity

Class 7. Hydrolysis of salts (2 hrs.)

1. Hydrolysis salts.

2. Calculation of hydrolysis constant
3. Calculation of hydrolysis degree

Class 8. Oxidation-Reduction reactions (2 hrs.)

1. Type of oxidation-reduction reactions
2. Stoichiometry of oxidation-reduction reactions
3. Redox potential and direction of reaction
4. Calculation of redox potential

Class 9. Coordination compounds (2 hrs.)

1. Structure of coordination compounds
2. The Formation of Complex Ions.
3. Dissociation and basic chemical properties of complex compounds

Unit II. Organic Chemistry (18 hrs.)

Class 1. Structure of Organic Compounds (2 hrs.).

1. Types of Bonds.
2. Structural Formulas.
3. Nomenclature.
4. Isomers.

Class 2. Aliphatic Hydrocarbons (2 hrs.).

1. Alkanes. Structure and Reactivity.
2. Cycloalkanes. Conformations of Cycloalkanes.
3. Unsaturated Hydrocarbons Alkenes. Structure and Reactivity.
4. Unsaturated Hydrocarbons Alkynes. Structure and Reactivity.

Class 3. Aromatic Hydrocarbons (2 hrs.).

1. Conjugated and Aromatic Molecules
2. The Hückel Rule.
3. Benzene. Electrophilic Aromatic Substitution
4. Arenas. Directing effects of Substituents: ortho-para directors, meta-directors

Class 4. Alcohols and Phenols (2 hrs.).

1. Nomenclature of Alcohols.
2. Physical properties of Alcohols.
3. Chemical Properties of Alcohols: Acid-Base Reactions, Substitutions Reactions, Dehydration, Oxidation.
4. Phenols: Physical and Chemical Properties.

Class 5. Aldehydes and Ketones (2 hrs.).

1. Nomenclature of Aldehydes and Ketones, Physical properties
2. Chemical Properties of Aldehydes and Ketones: Oxidation-Reduction Reactions, Addition Reactions, Formation of Acetals and Ketals,
3. Reactivity of the α -carbon Atom, the Aldol Condensation.

Class 6. Carboxylic Acid (2 hrs.).

1. Carboxylic Acid and Acyl Groups. Nomenclature of Carboxylic Acid.
2. Physical and Chemical Properties of Carboxylic Acid: Acidity, Nucleophilic Acyl Substitution, Reduction Acyl Derivatives.
3. Hydroxy acids. Configuration of molecules.
4. Optical Activity. Fischer Projection Formulas.
5. Test 1

Class 7. Carbohydrates (2 hrs.).

1. Classification of Carbohydrates.
2. Monosaccharides. Glucose. Fructose. Hemiacetals and Hemiketals.
3. Chemical Properties of Monosaccharides.
4. Disaccharides.
5. Polysaccharides.

Class 8. Amino Acids, Peptides and Proteins (2 hrs.).

1. Amino Acids. Acid-Base Properties of α -amino Acids. Isoionic Point.
2. Structure and Reactivity.
3. Peptides.
4. Protein. Protein Structure.

Class 9. Lipids (2 hrs.).

1. Classification of Lipids.
2. Fatty Acids.
3. Triglycerides. Chemical Properties.
4. Test 2

Laboratory practical (54 hours.)

Introduction to Lab activity (4 hrs.)

Course Policies and information. Safety Rules. Laboratory Care and Waste Disposal. Laboratory technique, materials and fundamental operation.

Experiment № 1. The Density of Liquids and Solids (4 hrs.)

Part A: The Density of Water

Part B: The Density of Aluminum and the Thickness of Foil

Part C: Graphical Analysis of Mass and Volume Data of an Unknown Solid

Experiment № 2. Rates of Chemical Reactions - A Clock Reaction (4 hrs.)

Part A: The Effect of Concentration on the Reaction Rate

Part B: The Effect of Temperature on the Reaction Rate

Experiment № 3. Chemical Equilibrium and Le Chatelier's Principle (4 hrs.)

Part A: Saturated Sodium Chloride Solution

Part B: Acidified Chromate Solution

Part C: Aqueous Ammonia Solution

Part D: Cobalt(II) Chloride Solution

Experiment № 4. Types of Reactions (4 hrs.)

Part A: Combination Reactions

Part B: Decomposition Reactions

Part C: Single Displacement Reactions

Part D: Double Displacement (Exchange) Reactions

Experiment № 5. Calculation and preparation of accurate concentrations solutions (4 hrs.)

Part A: Preparation of standard $\text{Na}_2\text{B}_4\text{O}_7$ solution

Part B: Preparation of titrated HCl solution

Part C: Preparation of titrated NaOH solution

Part D: Estimation of working HCl solution concentration

Part E: Estimation of working NaOH solution concentration

Experiment № 6. Preparing of Buffers and Buffer Capacity (4 hrs.)

Part A: Ammonium acetate buffer solution + strong acid

Part B: Ammonium acetate buffer solution + strong base

Part C: Sodium bicarbonate buffer

Part D: Preparation of a buffer solution

Experiment № 7. Inorganic qualitative reaction (4 hrs.)

Part A: Detection of cations of Group 1a

Part B: Detection of cations of Group 1b

Part C: Detection of cations of Group 2

Part D: Detection of cations of Group 3

Part E: Detection of cations of Group 4

Part F: Detection of cations of Group 5

Experiment № 8. Titration of Vinegar (4 hrs.)

Part A: Setting up the burette and preparing the NaOH

Part B: Preparing the vinegar sample

Part C: Performing the titration

Part D: Calculations of molarity of Acetic Acid in Vinegar

Part E: Calculations of mass Percent of Acetic Acid in Vinegar

Experiment № 9. Purification of Liquids by Distillation (4 hrs.)

Part A: Simple Distillation of a Binary Mixture

Part B: Separation of a Binary Mixture by Using a Fractionating Column

Part C: Control of Purity Derived Substances

Part D: Lab Report

Experiment № 10. Purification of Solids by Recrystallization (4 hrs.)

Part A: Solubility tests.

Part B: Recrystallization of Phthalic Acid.

Part C: Drying to constant weight.

Part D: Determination melting point, % yield

Part E: Lab Report

Experiment № 11. Identification of Organic Unknowns by Methods of Qualitative Organic Analysis (4 hrs.)

Part A: Elemental analysis of organic compounds (C,H, Cl)

Part B: Identification of Functional Groups (Chemical Methods)

Part C: Identification of Functional Groups (IR- spectroscopy)

Part D: Lab Report

Experiment № 12. Separation Chemical Substance by Thin Layer Chromatography (TLC) (4 hrs.)

Part A: Separating Mixture of Aspirin, Acetaminophen, and Caffeine by TLC

Part B: Determination R_f

Part C: Control Sample Identification

Part D: Lab Report

Experiment № 13. (2 hrs.)

Lab Report № 9-12

III. EDUCATIONAL-METHODICAL SUPPORT OF STUDENTS' INDEPENDENT WORK

Methodical maintenance of independent work on the course «Chemistry, Medical Chemistry» is presented in Appendix 1 and includes:

- schedule of independent work on the Course, including the approximate time standards for each unit;
- characteristic of tasks for independent work and methodical recommendations for their performances;
- requirements for design and presentation of independent work results;
- criteria for evaluating of independent work

III. CONTROL OF THE COURSE GOALS ACHIEVEMENTS

No	Controlled sections / topics of disciplines	Codes and stages of competence formation	Assessment tools		
			Formative assessment	Midterm control / exam	
1	Unit 1. General and Inorganic Chemistry	the readiness to use basic physical and chemical, mathematical and other natural science concepts and methods in solving professional problems (GPC – 7)	Know	Poll Test control Presentation	Question for exam 26-50
			Can	task	assignment
			Master	test	assignment
2	Unit 2. Organic Chemistry	the capacity for the assessment of morphological and physiological states and pathological processes in the human body for solving professional tasks (GPC – 9)	Know	Poll Test control Presentation	Question for exam 26-50
			Can	task	assignment
			Master	test	assignment

Typical control tasks, teaching materials, criteria and indicators needed to assess the knowledge, skills and characterize the stages of formation of competences in the course of development of the educational program are provided in Appendix 2.

LIST OF TEXTBOOKS AND INFORMATIONAL-METHODOLOGICAL SUPPORT OF DISCIPLINE

Basic literature

(electronic and printed media)

1. European Federation of Clinical Chemistry and Laboratory Medicine, 2016
H. Baum https://link.springer.com/referenceworkentry/10.1007/978-3-662-49054-9_1064-1
2. Chemistry for Surgeons, 2018 Sukriti Rastogi
<https://link.springer.com/article/10.1007/s12262-018-1813-5>
3. Light Microscopy, 2016 Yolanda Markaki, Hartmann Harz
<https://link.springer.com/book/10.1007/978-1-4939-6810-7>

Additional literature

(electronic and printed media)

1. Chronic Obstructive Pulmonary Disease, 2016 Hiroyuki Nakamura, Kazutetsu Aoshiba
<https://link.springer.com/book/10.1007/978-981-10-0839-9>
2. Forensic Toxicology, 2016 <https://link.springer.com/journal/11419>

LIST OF INFORMATION TECHNOLOGIES AND SOFTWARE

The location of the computer equipment on which the software is installed, the number of jobs	List of licensed software
Multimedia auditorium Vladivostok Russian island, Ayaks 10, building 25.1, RM. M723 Area of 80.3 m2 (Room for independent work)	Windows Seven enterprice SP3x64 Operating System Microsoft Office Professional Plus 2010 office suite that includes software for working with various types of documents (texts, spreadsheets, databases, etc.); 7Zip 9.20 - free file archiver with a high degree of data compression; ABBYY FineReader 11 - a program for optical character recognition; Adobe Acrobat XI Pro 11.0.00 - software package for creating and viewing electronic publications in PDF; WinDjView 2.0.2 - a program for recognizing and viewing files with the same format DJV and DjVu.

In order to provide special conditions for the education of persons with disabilities all buildings are equipped with ramps, elevators, lifts, specialized places equipped with toilet rooms, information and navigation support signs

METHODICAL INSTRUCTIONS ON SUBJECT STUDYING

Learning outcomes or intended learning outcomes are statements of what a learner is expected to know, understand and/or be able to demonstrate after completion of a process of learning. Learning outcomes are determined by the teaching staff. Student workload is the time (expressed in hours) that it is expected that an average learner (at a particular cycle/level) will need to spend to achieve specified learning outcomes. This time includes all the learning activities which the student is required to carry out (e.g. lectures, seminars, practical classes, private study, professional visits, examinations, etc.).

In the determination of workload, the following factors have an important role:

- The total number of contact hours for the course unit;
- Preparation before and finalizing of notes after the attendance of the lecture/seminar;
- The amount of further independent work required to finish the course unit successfully.

The amount of independent work is the most difficult item to calculate and depends largely on the discipline concerned and the complexity of the topic. Independent work includes:

- The collection and selection of relevant material;
- Reading and study of that material;
- Preparation for an oral or written examination;
- Writing of a paper or dissertation;
- Independent work in a laboratory.

Planning student workload

The following three steps that would help to plan student workload are recommended: 1. Estimating student workload (teacher plan). The average student workload of a course unit/module depends on the total amount of learning

activities a student is expected to complete in order to achieve the foreseen learning outcomes. It is measured in work hours. Workload can be defined on the basis of the following educational activities:

- Contact studies. They include work with or under the guidance of a teacher: lecture, seminar, laboratory work, tutorial, practical class, practical session, internship, work placement.

- Independent studies: performance of tasks, writing of papers, reading of books and articles, project work, practicing technical or laboratory skills. This item is the most difficult one to calculate.

- Assessment: oral or written examination, essay, test, examples of works, report, thesis, presentation

Examples of determining student workload

The whole study time can be divided into three parts:

- student's preliminary work before contact hours
- contact hours
- student's independent work after contact hours. The scope of independent work can be linked with the teaching/learning approach.

The time allocated for the independent performance of tasks depends on the type of a task.

Written work. The work time is calculated according to the formula: 100 words/1 hour.

Oral presentation. Where the length of a presentation is 1 hour, preparing for it requires a minimum of 6 hours.

Reading of literature. Students should know whether the literature is additional or is necessary for the examination.

Reading a book comprises three phases: looking through a book; thorough reading by taking notes of the key issues; review (repeating). A text will only be understood well if read three times. A text of 100 pages that is easy to read requires 20 hours, while that which is difficult to read or a text in a second language needs 30 hours. The time needed for learning a text that is being read can also be

estimated according to the complexity and the number of words of the text. The time should be multiplied by a figure that shows the times the text was read (e.g. 3 times).

Where a student is reading a book to prepare a report, reading is more motivated and effective. Moreover, writing facilitates perception, so reading the book once may be sufficient. Where a book is used for preparing a survey of literature for a research paper, normally it is not read in its entirety, but only in separate portions. However, this kind of work with a book also takes up a lot of time which should be estimated correctly (estimation can be based on the amount of words and the book can be deemed to have been read once). Information and communication technology assisted teaching/learning When estimating the time needed for virtual learning, the following factors are of relevance:

- The time required to carry out the assigned tasks;
- The time for communicating with the teacher and peers;
- The time required to read the specified literature;
- The time required for the search of material;
- The time for contact distance learning;
- The time for getting familiar with the learning environment. If the student

is not familiar with the electronic learning environment, he or she may need 8 to 24 hours of additional work. The time foreseen to master new software should be 8 hours.

A student may also waste some time due to technical problems. Contact hours for virtual learning are estimated in the same manner as those for traditional teaching methods. Preparation for assessment Student workload also includes preparation for assessment and the assessment itself (e.g. examination). The time used depends on the assessment method. Traditional examination. The estimation of the time needed to prepare for an examination can be based on a rule that each week of studies (40 hours) corresponds to 8 hours of preparation for an examination. Project (report, portfolio) assessment A report may be oral or prepared in written form. Workload is estimated with account of the time used to

search for and read literature. Continuous assessment. The student is assessed for the performance of minor tasks and is under constant monitoring without a separate examination. In this case, the assessment does not require any special time. Continuous assessment enhances student activity and students usually make a very effective use of the time allocated for a course unit.

MATERIAL AND TECHNICAL EQUIPMENT OF SUBJECT

For carrying out practical work, as well as for organizing independent work, students have access to the following laboratory equipment and specialized classrooms that meet applicable sanitary and fire regulations, as well as safety requirements for educational and research and production work:

Name of equipped premises and rooms for independent work	List of basic equipment
Computer classroom School of Biomedicine. Laboratory building aud. L403, 15 seats	Screen with an electric drive 236 * 147 cm Trim Screen Line; DLP Projector, 3000 ANSI Lm, WXGA 1280x800, 2000: 1 EW330U Mitsubishi; The subsystem of specialized fixing equipment CORSA-2007 Tuarex; Video switching subsystem: DVI DXP 44 DVI Pro Extron matrix switcher; DVI extension cable for twisted pair DVI 201 Tx / Rx Extron; Audio switching and sound reinforcement subsystem; ceiling speaker system SI 3CT LP Extron; DMP 44 Extron digital audio processor; extension for the control controller IPL T CR48; Wireless LANs for students are provided with a system based on 802.11a / b / g / n access points 2x2 MIMO (2SS). Monoblock HP RgoOpe 400 All-in-One 19.5 (1600x900), Core i3-4150T, 4GB DDR3-1600 (1x4GB), 1TB HDD 7200 SATA, DVD +/- RW, GigEth, Wi-Fi, BT, usb kbd / mse, Win7Pro (64-bit) + Win8.1Pro (64-bit), 1-1-1 Wty
Multimedia audience	Monoblock Lenovo C360G-i34164G500UDK; Projection screen Projecta Elpro Electrol, 300x173 cm; Multimedia projector, Mitsubishi FD630U, 4000 ANSI Lumen, 1920x1080; Mortise interface with TLS TAM 201 Stan automatic cable retractor; Document Camera Avervision CP355AF; Sennheiser EW 122 G3 UHF range microphone microphone wireless system as part of a wireless microphone and receiver; Video conferencing codec LifeSizeExpress 220- Codeconly- Non-AES; Network camera Multipix MP-HD718; Dual LCD panels 47 ", Full HD, LG M4716CCBA; Audio switching and sound reinforcement subsystem; centralized uninterrupted power supply
Reading rooms of the FEFU Scientific Library with open access to the Foundation (Building A - Level 10)	HP RgoOpe 400 All-in-One 19.5 (1600x900), Core i3-4150T, 4GB DDR3-1600 (1x4GB), 1TB HDD 7200 SATA, DVD +/- RW, GigEth, Wi-Fi, BT, usb kbd / mse, Win7Pro (64-bit) + Win8.1Pro (64-bit), 1-1-1 Wty Internet access speed 500 Mbit / s. Jobs for people with disabilities are equipped with braille displays and printers; equipped with: portable devices for reading flat-printed texts, scanning and reading machines with a video optimizer with the ability to adjust color spectra; magnifying



	electronic loops and ultrasonic markers
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MINISTRY OF EDUCATION AND SCIENCE OF THE RUSSIAN FEDERATION
Federal State Autonomous Educational Institution of Higher Professional Education
«Far Eastern Federal University»
(FEFU)

SCHOOL OF BIOMEDICINE

«CONCURRENCE:»
Education Program Manager

(signature) (print full name)
« » _____ 20__ г.

«APPROVALS:»
Head of Department
Chemistry and Biosystems Engineering
(Department name)

(signature) (print full name)
« » _____ 20__ г.

**EDUCATIONAL AND METHODOLOGICAL SUPPORT INDEPENDENT WORK
OF STUDENTS
Chemistry, Medical Chemistry
Specialty 31.05.01 General Medicine
Full-time training**

**Vladivostok,
2018**

Independent work includes:

- 1) The library or homework with educational literature and lecture notes,
- 2) Preparation for practical training,
- 3) Preparation for testing and control interview.

Order of independent work of the students is determined by schedule plan of this work in the subject.

The schedule plan of independent work in the subject.

No	Date/time period	Type of Independence Work	The approximate time to complete the rules	Forms of control
1.	November, 15	Work with literature	8 hrs.	Library-research paper
2.	During the 1 st Term	Homework	7 hrs.	Report
3.	During the 1 st Term	Revising for prelab discussion	8 hrs.	Recitation/Quiz
4.	During the 1 st Term	Lab report preparation	8 hrs.	Lab report/ Participation
5.	December, 25	Revising for pass-fail exam	7 hrs.	Test
6.	During the 2 nd Term	Revising for prelab discussion	4 hrs.	Recitation/Quiz
7.	During the 2 nd Term	Lab report preparation	8 hrs.	Lab report/ Participation
8.	During the 2 nd Term	Homework	6 hrs	Report
9.	June, 15	Revising for exam	36 hrs.	Exam

Recommendations to Independence Work of Students

Pre-lab Assignment and Prelab Discussion

The pre-lab assignment is handed in at the beginning of the lab period. The pre-lab assignment contains questions about the upcoming lab including a section on safety. The pre-lab assignment is graded and worth 5%. Without it, you can not begin lab. Participation in the pre-lab discussion is required. Your presence will earn you 5%. During the pre-lab discussion, the lab instructor will relate important lab information, review the pre-lab assignment, discuss the upcoming lab, and give you an opportunity to ask questions. If you arrive late, the 5% cannot be recovered.

Laboratory work including taking notes

To a large extent, you control what you do in lab. Successful lab work takes active planning and thinking. What you do in lab is a reflection of the care and attention given to your laboratory work. Sometimes there is room for creative and critical thinking in lab work. For example, the lab procedures are, by design, sometimes incomplete. You may be asked to design an experiment to test a hypothesis. You may be able to design the experiment in more than one equally valid way. Scientists sometimes follow different methods to accomplish the same task and this has led to important scientific discoveries. You should use the lab notebook to describe how you approach tasks, answer questions posed in the lab procedure, make observations, and to record important information like the mass of a chemical. Trying to save pages in a lab notebook for next semester is a “penny wise - pound foolish” way of using your lab notebook. Lab instructors assign 10% of your grade from lab work. It is rather easy to earn the 10%. It is not so easy, however, to use your time in lab carefully and thoughtfully.

Post-Lab Discussion

By the end of lab, you will have gathered evidence. The post-lab discussion is about the process of thinking about this evidence and how it can be analyzed, through reasoning, to make valid scientific claims. Not all of the evidence you will gather will be easy to interpret. For example, sometimes mistakes in lab will make conclusions very difficult. However, there is no requirement to get a certain outcome in your lab work. What is required is that you work with the evidence you have generated (unless told by a lab instructor to do otherwise). The post-lab discussion can make writing up your results easier. Timing of the post lab discussion varies. Your lab instructor will appreciate your help in forming a group of 4-8 students to have a post lab discussion. Students can have this discussion without the lab instructor but touching base with the lab instructor should always occur before leaving lab. To document the post-lab discussion, you should write 5-6 summarizing sentences in your lab notebook, under the heading, Post-lab Discussion. The post-lab discussion carries 10 pts for participation. If your lab

instructor notices that you do not participate in the lab discussion and that you do not write the summary, these points will not be awarded.

Lab Report Rubric

Lab Report Rubric	Level of attainment		
	Level 0 (0-20 points)	Level 1 (21-40 points)	Level 2 (41-60 points)
Claim(s): Is a statement(s), derived from evidence, using scientific reasoning. The claim should be underlined on the first page following the introduction.	Does not make a claim, or makes an inaccurate claim.	Makes an accurate but incomplete claim.	Makes an accurate and complete claim.
Evidence: Scientific data that supports the claim. The data needs to be appropriate and sufficient to support the claim.	Does not provide evidence, or only provides inappropriate evidence (Evidence that does not support claim).	Provides appropriate, but insufficient evidence to support claim. May include some inappropriate evidence. Evidence from other published sources may be mentioned.	Provides appropriate and sufficient evidence to support claim. Communicates and describes evidence from other published sources that adds significance to gathered evidence.
Reasoning: Scientific arguments that use evidence and appropriate scientific principles to make claims.	Does not provide reasoning, or only provides reasoning that does not link evidence to claim..	Provides reasoning that links the claim and evidence. Repeats the evidence and/or includes some scientific principles, but is not sufficient. May identify and/or mention outside sources or studies.	Provides reasoning that links evidence to claim. Includes appropriate and sufficient scientific principles. Appropriately identifies and discusses outside sources or studies.

Samples for Library-Research Work

1. Compounds of calcium in bone tissue, similarity of Ca-ions and Sr-ions, isomorphous replacement (“bone seeker”, Strontium-90 problem).
2. Biological activity of complex compounds. Metalenzymes: concept about structure of active center.
3. Water as the most important hydrogen compound, its physical and chemical properties. Aquo-complexes and crystalline hydrates. Distilled and pyrogenic-free water: preparation and use in pharmacy.
4. Biological role of s-elements in mineral balance of Human body. Macro- and micro-s-elements.

5. Chemical bases use titanium, niobium and tantalum in surgery. Use of titanium dioxide and ammonium metavanadate in pharmacy.
6. Chemical mechanism of cadmium and mercury toxic activity. Chemical bases use mercury compounds in medicine and pharmacy.
7. Concept about chemistry of bactericidal action of chlorine and iodine. The use in medicine of bleach, bleach water, preparations of active chlorine and iodine.
8. Noble and inert gases and their compounds. Use of Noble gases in medicine.
9. Radioactive elements. Use of radioisotopes in medicine. Tracer method.
10. Radioactivity. The alpha- and beta-radioactivity of natural mineral water.
11. Chemical bases use of silver compounds as therapeutic agents.
12. Biological activity of boron. Antiseptic properties of boric acid and its salts.
13. The chemistry of the toxic effect of lead compounds. The use in medicine of lead drugs (Lead (II) acetate, lead (II) oxide).
14. Biological activity of oxygen. Chemical base use of oxygen and ozone in medicine.
15. Phenomenon of chemical catalysis. Basic principles and type of catalysis. The catalysts, promoters, catalyst poisons and inhibitors.



MINISTRY OF EDUCATION AND SCIENCE OF THE RUSSIAN FEDERATION
Federal State Autonomous Educational Institution of Higher Education
«Far Eastern Federal University»
(FEFU)

SCHOOL OF BIOMEDICINE

FUND OF ASSESSMENT TOOLS

Chemistry, Medical Chemistry

Specialty 31.05.01 General Medicine

Full-time training

**Vladivostok,
2018**

The passport is filled in accordance with the Regulations on the Funds of assessment tools of educational programs of higher education - undergraduate, specialist's and master's programs of Far Eastern Federal University, approved by order of the rector of 12.05.2015 №12-13-850.

Code and formulation of competence	Competence formulation phase	
the readiness to use basic physical and chemical, mathematical and other natural science concepts and methods in solving professional problems (GPC – 7)	Knowing	the basic computer databases about the structure and properties of organic compounds, including chemical and 3D computer graphics program
	Be able	using the rules of construction of chemical formulas, graphs, tables, using appropriate computer programs, including for the creation of computer presentations.
	Be master (skill)	by computer programs to build chemical and stereochemical formulas of organic compounds and other illustrative material.
the capacity for the assessment of morphological and physiological states and pathological processes in the human body for solving professional tasks (GPC – 9)	Knowing	physical and chemical basis of pathological processes in Human body; structure, significance and role of basic biogenic elements and they compounds in living systems; chemical methods of estimation of morph-functional and physiological states of living systems
	Be able	apply chemical concepts, laws and principals for estimation of morph-functional and physiological states, and of pathological processes in Human body
	Be master (skill)	by methods of estimation of morph-functional and physiological states of Human body including physical and chemical analysis methods

CONTROL OF THE COURSE GOALS ACHIEVEMENTS

No	Controlled sections / topics of disciplines	Codes and stages of competence formation	Assessment tools		
				Formative assessment	Midterm control / exam
1	Unit 1. General and Inorganic Chemistry	the readiness to use basic physical and chemical, mathematical and other natural science concepts and methods in solving professional problems (GPC – 7)	Know	Poll Test control Presentation	Question for exam 26-50
			Can	task	assignment
			Master	test	assignment
2	Unit 2. Organic Chemistry	the capacity for the assessment of morphological and physiological	Know	Poll Test control Presentation	Question for exam 26-50
			Can	task	assignment
			Master	test	assignment

		states and pathological processes in the human body for solving professional tasks (GPC – 9)			
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Scale of competence level assessment

Code and formulation of competence.	Stages of formation of competence		Criteria	Indicators	Points
the readiness to use basic physical and chemical, mathematical and other natural science concepts and methods in solving professional problems (GPC – 7)	Knows (entry level)	Theoretical bases of computer science, collection, storage, retrieval, processing, transformation, dissemination of information in medical and biological systems;	Knowledge of the theoretical foundations of computer science, storage, search; processing, transformation, dissemination of information in medical and biological systems;	Formed structured systematic knowledge of the theoretical foundations of informatics, storage, retrieval; processing, transformation, dissemination of information in medical and biological systems;	65-71
	Can (advanced level)	Conduct textual and graphic processing of medical data using standard operating system tools and common office applications, as well as application and special software;	Ability to conduct textual and graphic processing of medical data using standard operating system tools and common office applications, as well as application and special software;	Ready and able to carry out textual and graphic processing of medical data using standard operating system tools and common office applications, as well as application and special software tools;	71-84
	Master (high level)	Basic methods of statistical processing of clinical and experimental data with application of standard applied and special software;	Skill in applying basic methods of statistical processing of clinical and experimental data with standard application and special software techniques;	Systematic application of the skills of basic methods of statistical processing of clinical and experimental data with the methods of standard applied and special software;	85-100
the capacity for the assessment of morphological and	Knows	physical and chemical basis of pathological	Knowledge of physical and	Formed structured systematic	

physiological states and pathological processes in the human body for solving professional tasks (GPC – 9)	(entry level)	processes in Human body; structure, significance and role of basic biogenic elements and they compounds in living systems; chemical methods of estimation of morph-functional and physiological states of living systems	chemical basis of pathological processes in Human body; structure, significance and role of basic biogenic elements and they compounds in living systems; chemical methods of estimation of morph-functional and physiological states of living systems	knowledge of physical and chemical basis of pathological processes in Human body; structure, significance and role of basic biogenic elements and they compounds in living systems; chemical methods of estimation of morph-functional and physiological states of living systems	
	Can (advanced level)	apply chemical concepts, laws and principals for estimation of morph-functional and physiological states, and of pathological processes in Human body	Ability to apply chemical concepts, laws and principals for estimation of morph-functional and physiological states, and of pathological processes in Human body	Ready and able to carry out apply chemical concepts, laws and principals for estimation of morph-functional and physiological states, and of pathological processes in Human body	
	Master (high level)	by methods of estimation of morph-functional and physiological states of Human body including physical and chemical analysis methods	Skill in application of the methods of estimation of morph-functional and physiological states of Human body including physical and chemical analysis methods	Systematic application of the methods of estimation of morph-functional and physiological states of Human body including physical and chemical analysis methods	

GENERAL CHEMISTRY QUESTIONS

Electronic Structure and Periodic Table

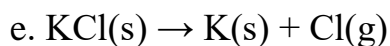
1. What value or values of ml are allowable for an orbital with $l = 2$?
 - a. 0
 - b. 2
 - c. -1
 - d. none of the above
 - e. all of the above

2. According to Bohr Theory, which of the following transitions in the hydrogen atom will give rise to the least energetic photon? Use the equation: $E_n = (-2.18 \times 10^{-18} \text{ J})(1/n^2)$
- $n = 5$ to $n = 3$
 - $n = 6$ to $n = 1$
 - $n = 4$ to $n = 3$
 - $n = 5$ to $n = 4$
 - $n = 6$ to $n = 5$
3. Consider a $3d_{xz}$ orbital. Which of the following statements is incorrect?
- The xz plane is a nodal surface.
 - The xz plane divides the electron probability distribution into two identical mirror-image halves.
 - The xy plane divides the electron probability distribution into two identical mirror-image halves.
 - The yz plane divides the electron probability distribution into two identical mirror-image halves.
 - The nucleus is located at a node.
4. The electronic configuration of the element whose atomic number is 26 is:
- $1s^2 2s^2 2p^6 3s^2 3p^6 4s^0 3d^8$
 - $1s^2 2s^2 2p^6 3s^2 3p^6 3d^6 4s^2$
 - $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$
 - $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^4 4p^2$
 - none of the above
5. Which of the following has the largest radius?
- F
 - N
 - C
 - O
 - Ne
6. Which of the following elements has the largest ionization energy?
- Na
 - Ne
 - F
 - K
 - Rb
7. Which of the following has the greatest electron affinity (most negative value)?
- Cl
 - K
 - He

- d. Na
 - e. Rb
8. Which of the following species is not isoelectronic with any of the others?
- a. V^{3+}
 - b. Ca^{2+}
 - c. Ar
 - d. Cl^-
 - e. S^{2-}
9. In Bohr's model of the hydrogen atom, the radius of an orbit
- a. is proportional to n^2 .
 - b. is smallest for the highest energy state.
 - c. increases when a photon of light is emitted from an excited atom.
 - d. can have any value that is larger than the ground-state radius.
 - e. none of the above
10. Which of the following atoms is not a one-electron system?
- a. H
 - b. He^+
 - c. Li^{2+}
 - d. Be^{2+}
 - e. O^{7+}
11. Which of the following statements about periodic properties is incorrect?
- a. Both electron affinity and ionization energy decrease down a group.
 - b. Atomic size increases to the right across a period.
 - c. Ionization energy increases to the right across a period.
 - d. Atomic size increases down a group.
 - e. Electron affinity increases to the right across a period.

Bonding

1. Which one of the following is most likely to be an ionic compound?
- a. HNF_2
 - b. H_2CO
 - c. N_2H_4
 - d. $CaCl_2$
 - e. CH_3Cl
2. In which of the following processes does the enthalpy change (ΔH) directly represent the magnitude of the lattice energy of $KCl(s)$?
- a. $Cl_2(g) + 2K(s) \rightarrow 2KCl(s)$
 - b. $KCl(s) \rightarrow K^+(aq) + Cl^-(aq)$
 - c. $KCl(s) \rightarrow K^+(g) + Cl^-(g)$
 - d. $KCl(s) \rightarrow K(s) + Cl(g)$



3. Order the following by increasing bond strength: $\text{N}\equiv\text{N}$, $\text{N}=\text{N}$, $\text{N}-\text{N}$
- $\text{N}\equiv\text{N}$, $\text{N}=\text{N}$, $\text{N}-\text{N}$
 - $\text{N}\equiv\text{N}$, $\text{N}-\text{N}$, $\text{N}=\text{N}$
 - $\text{N}-\text{N}$, $\text{N}=\text{N}$, $\text{N}\equiv\text{N}$
 - $\text{N}=\text{N}$, $\text{N}-\text{N}$, $\text{N}\equiv\text{N}$
 - $\text{N}=\text{N}$, $\text{N}\equiv\text{N}$, $\text{N}-\text{N}$
4. Which of the following compounds has the greatest bond polarity?
- PH_3
 - NH_3
 - HF
 - H_2S
 - CH_4
5. Which of the following is not planar?
- BCl_3
 - ClF_3
 - PCl_3
 - XeF_4
 - C_2H_4
6. Use VSEPR theory to predict the ideal bond angles around the two carbon atoms in acetaldehyde, CH_3CHO . (The first carbon has single bonds to three H atoms and one C atom; the second carbon has single bonds to C and H, and a double bond to O.)
- 109° , 109°
 - 109° , 120°
 - 120° , 109°
 - 120° , 90°
 - 105° , 105°
7. In a carbon-carbon triple bond, what is the nature of the bonding between the carbons?
- two 2s orbitals overlapping
 - two 2p orbitals overlapping
 - two sp orbitals overlapping, two 2py overlapping and two 2pz overlapping
 - an sp and sp² overlapping and 2p orbitals overlapping
 - an sp² and sp² overlapping and 2p orbitals overlapping
8. Which of the following molecules has sp³ hybridization and a dipole moment?
- SiH_4
 - BF_3
 - NH_3

d. BrF₃

e. PCl₅

9. In the molecular orbital description of bonding in benzene (C₆H₆), how many electrons occupy delocalized MOs?

a. 2

b. 3

c. 4

d. 5

e. 6

10. In which of the following species is the octet rule violated by the central atom?

a. CH₄

b. SF₄

c. PCl₄⁺

d. SO₂

e. NH₃

11. The number of electron dots in the Lewis symbol for an element equals the

a. number of outermost p electrons.

b. number of electrons needed to fill the outermost p orbital.

c. period number that contains the element.

d. number of outermost s and p electrons.

e. number of outermost s electrons.

Phases and Phase Equilibria

1. Calculate the pressure of 0.55 mol of NH₃ gas in a 2.00 L vessel at 25 °C, using the ideal gas law.

a. 2.5 atm

b. 6.7 atm

c. 0.6 atm

d. 7.5 atm

e. 3.4 atm

2. A steel tank contains carbon dioxide at 34 °C and is at a pressure of 13.0 atm. Determine the internal gas pressure when the tank and its contents are heated to 100 °C.

a. 10.7 atm

b. 9.4 atm

c. 38.2 atm

d. 1.9 atm

e. 15.8 atm

3. Deviations from the ideal gas law are less at:

a. high temperatures and high pressures

- b. high temperatures and low pressures
 - c. low temperatures and high pressures
 - d. low temperatures and low pressures
 - e. high volumes and low temperatures
4. A mixture of three gases has a pressure of 1380 mmHg at 298 K. The mixture is analyzed and is found to contain 1.27 mol CO₂, 3.04 mol CO, and 1.50 mol Ar. What is the partial pressure of Ar?
- a. 238 mm Hg
 - b. 302 mm Hg
 - c. 356 mm Hg
 - d. 1753 mm Hg
 - e. 8018 mm Hg
5. Which of the following exhibits the most hydrogen bonding?
- a. LiH
 - b. CH₄
 - c. NH₃
 - d. H₂S
 - e. CH₂F₂
6. Which of the following carbon compounds has the highest melting point?
- a. CF₄
 - b. CCl₄
 - c. CBr₄
 - d. CI₄
 - e. CH₄
7. Water has such a high specific heat because
- a. it has such a low molecular weight.
 - b. it is rather dense.
 - c. the O-H single bond has a high bond energy.
 - d. it has many relatively strong hydrogen bonds.
 - e. it dissolves both ionic and covalent compounds.
8. The triple point is
- a. an end to the liquid-gas line in a phase diagram.
 - b. the relationship between the boiling point, melting point and vapor pressure of a substance.
 - c. the point on a phase diagram where solid, liquid, and gas are in equilibrium.
 - d. the three pieces of data needed to solve the Clausius-Clapeyron equation.
 - e. the (P,V,T) coordinate of a point on a phase diagram.
9. The main forces responsible for the structure of DNA are
- a. ionic bonds and covalent bonds.

- b. covalent bonds and ionic bonds.
 - c. hydrogen bonds and dipole-dipole interactions.
 - d. covalent bonds and hydrogen bonds.
 - e. covalent bonds and dipole-dipole interactions.
10. Which of the following is not likely to exhibit hydrogen bonding?
- a. CH₃CH₂OH
 - b. CH₃NH₂
 - c. H₂O
 - d. NH₂OH
 - e. (CH₃)₃N

Stoichiometry

1. What is the mass of one mole of acetylsalicylic acid (aspirin), C₉H₈O₄?
- a. 29 g
 - b. 108 g
 - c. 196 g
 - d. 180. g
 - e. none of the above
2. Determine the number of moles of aluminum in 2.154 x 10⁻¹ kg of Al.
- a. 5816 mol
 - b. 7.984 mol
 - c. 6.02 x 10²³ mol
 - d. 4.801 mol
 - e. 8.783 mol
3. How many grams of zinc are there in 22.7 g of ZnCl₂?
- a. 0.35 g
 - b. 0.17 g
 - c. 10.9 g
 - d. 1476 g
 - e. 0.32 g
4. A compound with a composition of 87.5 % N and 12.5 % H was recently discovered. What is the empirical formula for this compound?
- a. NH₂
 - b. N₂H₃
 - c. NH
 - d. N₂H₂
 - e. N₂H
5. This equation is unbalanced: PCl₃ + H₂O → H₃PO₃ + HCl When it is correctly balanced, the coefficients are, respectively

- a. 1,3,1,1
- b. 1,1,1,3
- c. 1,3,1,3
- d. 2,3,2,3
- e. none of the above

6. Given 6 mol of each reactant, which one would be limiting in the following reaction? $4\text{Au} + 8\text{NaCN} + \text{O}_2 + 2\text{H}_2\text{O} \rightarrow 4\text{NaAu}(\text{CN})_2 + 4\text{NaOH}$

- a. Au
- b. NaCN
- c. O₂
- d. H₂O
- e. There is no limiting reactant.

7. In the direct reaction of silicon with Cl₂ the yield of SiCl₄ is 50. %. How many grams of silicon must be reacted with excess chlorine in order to obtain 17 g SiCl₄?

- a. 1.4 g
- b. 2.8 g
- c. 5.6 g
- d. 17 g
- e. 28 g

8. In the reaction of Fe₃O₄ with carbon to form carbon dioxide and iron, the number of moles of carbon required to convert 23 g of Fe₃O₄ to products is

- a. 0.05
- b. 0.1
- c. 0.2
- d. 0.3
- e. 0.4

9. A 20.0 mL sample of an element with a density of 3.0 g/mL contains 4×10^{23} atoms. What is the atomic weight of this element?

- a. 300
- b. 40
- c. 60
- d. 90
- e. none of the above

10. How many moles of oxygen gas will react with 12.4 mol aluminum? Equation: $4\text{Al} + 3\text{O}_2 \rightarrow 2\text{Al}_2\text{O}_3$

- a. 0.24 mol
- b. 0.42 mol
- c. 4.8 mol

- d. 9.3 mol
- e. 16.8 mol

11. Balance the following redox equation occurring in aqueous solution: $\text{KMnO}_4 + \text{KCl} + \text{H}_2\text{SO}_4 \rightarrow \text{MnSO}_4 + \text{K}_2\text{SO}_4 + \text{H}_2\text{O} + \text{Cl}_2$ What is the stoichiometric coefficient for chlorine (Cl_2) when the equation is balanced with smallest whole number coefficients?

- a. 1
- b. 3
- c. 5
- d. 8
- e. 10

Thermodynamics and Thermochemistry

1. Data: (1) $\text{H}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}(\text{g}) \quad \Delta H = -241.8 \text{ kJ}$ (2) $\text{H}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}(\text{l}) \quad \Delta H = -285.8 \text{ kJ}$

On the basis of the above data, which of the following statements is false?

- a. Reaction (1) is exothermic.
 - b. Reaction (2) is the formation reaction for $\text{H}_2\text{O}(\text{l})$.
 - c. The reverse of reaction (2) is endothermic.
 - d. The energy content of $\text{H}_2\text{O}(\text{g})$ is lower than $\text{H}_2\text{O}(\text{l})$.
 - e. ΔH for the reaction: $\text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2\text{O}(\text{g})$ is $+44 \text{ kJ/mol}$.
2. What is the amount of heat necessary to raise the temperature of 8.5 kg of water from $12.5 \text{ }^\circ\text{C}$ to $84 \text{ }^\circ\text{C}$?
- a. $3.0 \times 10^3 \text{ kJ}$
 - b. 36 J
 - c. $2.5 \times 10^3 \text{ kJ}$
 - d. $2.5 \times 10^6 \text{ kJ}$
 - e. 25 kJ
3. Data: ΔH°_f values: $\text{CH}_4(\text{g})$, -74.8 kJ ; $\text{CO}_2(\text{g})$, -393.5 kJ ; $\text{H}_2\text{O}(\text{l})$, -285.8 kJ . Using the ΔH°_f data above, calculate $\Delta H^\circ_{\text{rxn}}$ for the reaction below. Reaction: $\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l})$
- a. -604.2 kJ
 - b. 890.3 kJ
 - c. -997.7 kJ
 - d. -890.3 kJ
 - e. none of the above

4. Data: $2\text{Ba}(\text{s}) + \text{O}_2(\text{g}) \rightarrow 2\text{BaO}(\text{s}) \quad \Delta H^\circ = -1107.0 \text{ kJ}$ $\text{Ba}(\text{s}) + \text{CO}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{BaCO}_3(\text{s}) \quad \Delta H^\circ = -822.5 \text{ kJ}$ Given the data above, calculate ΔH° for the reaction below. Reaction: $\text{BaCO}_3(\text{s}) \rightarrow \text{BaO}(\text{s}) + \text{CO}_2(\text{g})$

- a. -1929.5 kJ
- b. -1376.0 kJ
- c. -284.5 kJ
- d. 269.0 kJ
- e. 537 kJ

5. Which of the following is not a state function?

- a. ΔE
- b. ΔH
- c. q
- d. P
- e. V

6. Two solutions (the system), each of 25.0 mL volume and at 25.0 °C, are mixed in a beaker. A reaction occurs between them, causing the temperature to drop to 20.0 °C. After the products have equilibrated with the surroundings, the temperature is again 25.0 °C and the total volume is 50.0 mL. No gases are involved in the reaction. Which one of the following relationships concerning the change from initial to final states (both at 25.0 °C) is correct?

- a. $\Delta E = 0$
- b. $\Delta H = 0$
- c. $\Delta E < 0$
- d. $w = 0$
- e. $q = 0$

7. Which one of the following processes is exothermic?

- a. $\text{H}_2(\text{l}) \rightarrow \text{H}_2(\text{g})$
- b. $\text{CO}_2(\text{s}) \rightarrow \text{CO}_2(\text{g})$
- c. $\text{H}_2\text{O}(\text{g}) \rightarrow \text{H}_2\text{O}(\text{l})$
- d. $16\text{CO}_2(\text{g}) + 18\text{H}_2\text{O}(\text{l}) \rightarrow 2\text{C}_8\text{H}_{18}(\text{l}) + 25\text{O}_2(\text{g})$
- e. $\text{H}_2(\text{g}) \rightarrow 2\text{H}(\text{g})$

8. Predict the signs of ΔH° , ΔS° , and ΔG° for the vaporization of liquid water at 150°C.

- a. $\Delta H^\circ > 0$, $\Delta S^\circ > 0$, $\Delta G^\circ > 0$
- b. $\Delta H^\circ < 0$, $\Delta S^\circ < 0$, $\Delta G^\circ < 0$
- c. $\Delta H^\circ > 0$, $\Delta S^\circ < 0$, $\Delta G^\circ > 0$
- d. $\Delta H^\circ > 0$, $\Delta S^\circ > 0$, $\Delta G^\circ < 0$
- e. none of the above

9. Which of the following substances has the lowest standard molar entropy (S°) at 25°C?

- a. $\text{CH}_3\text{OH}(\text{l})$
- b. $\text{CO}(\text{g})$
- c. $\text{MgO}(\text{s})$

- d. H₂O(l)
- e. CaCO₃(s)

10. When crystalline solid barium hydroxide octahydrate and crystalline solid ammonium nitrate are mixed in a beaker at room temperature, a spontaneous reaction occurs. The temperature of the beaker contents rapidly falls to below 0° C. Use this information to decide whether the reaction is exothermic or endothermic and what the signs of ΔH and ΔS are.

- a. endothermic; $\Delta H > 0; \Delta S > 0$
- b. exothermic; $\Delta H < 0; \Delta S > 0$
- c. endothermic; $\Delta H < 0; \Delta S < 0$
- d. endothermic; $\Delta H < 0; \Delta S > 0$
- e. exothermic; $\Delta H > 0; \Delta S < 0$

11. Sodium carbonate can be made by heating sodium hydrogen carbonate:
 $2\text{NaHCO}_3(\text{s}) \rightarrow \text{Na}_2\text{CO}_3(\text{s}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{g})$ For this reaction, $\Delta H^\circ = 128.9 \text{ kJ}$ and $\Delta S^\circ = 321 \text{ J/K}$. At approximately what temperature will $K = 1$?

- a. 401.6 K
- b. 401.6° C
- c. 33.1 K
- d. 33.1° C
- e. none of the above

Rate Processes in Chemical Reactions—Kinetics and Equilibrium

1. For the overall hypothetical reaction $\text{A} + 5\text{B} \rightarrow 4\text{C}$, the rate of appearance of C given by $\Delta[\text{C}]/\Delta t$ is the same as

- a. $\Delta[\text{A}]/\Delta t$
- b. $-(5/4)(\Delta[\text{B}]/\Delta t)$
- c. $-(4/5)(\Delta[\text{B}]/\Delta t)$
- d. $-(1/4)(\Delta[\text{A}]/\Delta t)$
- e. none of the above.

2. The initial rate of the reaction $\text{PCl}_5 \rightarrow \text{PCl}_3 + \text{Cl}_2$ is increased a factor of four when the concentration of PCl_5 is doubled. Therefore, the rate

- a. depends on the concentrations of PCl_3 and Cl_2 .
- b. is first order with respect to PCl_5 .
- c. is second order with respect to PCl_5 .
- d. is fourth order with respect to PCl_5 .
- e. is first order with respect to PCl_3 .

3. Consider the reaction $\text{A} \rightarrow \text{products}$. Which of the following plots is consistent with a zero-order reaction?

- a. $[\text{A}]$ plotted against time gives a horizontal, straight line.

- b. In $[A]$ plotted against time gives a straight line of negative slope.
- c. $1/[A]$ plotted against time gives a straight line of positive slope.
- d. $[A]$ plotted against time gives a straight line of negative slope.
- e. $[A]$ plotted against time gives a curved line of negative slope, decreasing in magnitude as time increases
4. The rate constant of a first-order reaction is $3.68 \times 10^{-2} \text{ s}^{-1}$ at 150°C , and the activation energy is 71 kJ/mol . What is the value of the rate constant at 170°C ?
- a. $9.2 \times 10^{-2} \text{ s}^{-1}$
- b. $3.7 \times 10^{-2} \text{ s}^{-1}$
- c. 2.49 s^{-1}
- d. $4.0 \times 10^{-2} \text{ s}^{-1}$
- e. none of the above
5. The reaction $3\text{ClO}^-(\text{aq}) \rightarrow \text{ClO}_3^-(\text{aq}) + 2\text{Cl}^-(\text{aq})$ has been proposed to occur by the following mechanism. $\text{ClO}^-(\text{aq}) + \text{ClO}^-(\text{aq}) \rightarrow \text{ClO}_2^-(\text{aq}) + \text{Cl}^-(\text{aq})$ (slow)
 $\text{ClO}_2^-(\text{aq}) + \text{ClO}^-(\text{aq}) \rightarrow \text{ClO}_3^-(\text{aq}) + \text{Cl}^-(\text{aq})$ (fast) Which rate law is consistent with this mechanism?
- a. $\text{rate} = k[\text{ClO}^-]$
- b. $\text{rate} = k [\text{ClO}^-]^3$
- c. $\text{rate} = k [\text{ClO}_2^-][\text{ClO}^-]$
- d. $\text{rate} = k [\text{ClO}^-]^2$
- e. $\text{rate} = k [\text{Cl}^-][\text{ClO}^-]^2$
6. A catalyst speeds up a reaction by
- a. increasing the number of high-energy molecules.
- b. increasing the temperature of the molecules in the reaction.
- c. increasing the number of collisions between molecules.
- d. increasing the activation energy for the reaction.
- e. providing a new reaction pathway for molecules.
7. Consider the following gas-phase equilibrium: $\text{H}_2(\text{g}) + \text{I}_2(\text{g}) \leftrightarrow 2\text{HI}(\text{g})$ At a certain temperature, the equilibrium constant K_c is 4.0 . Starting with equimolar quantities of H_2 and I_2 and no HI , when equilibrium was established, 0.20 moles of HI was present. How much H_2 was used to start the reaction?
- a. 0.10 mol
- b. 0.23 mol
- c. 0.20 mol
- d. 4.0 mol
- e. Need to know the volume of the reaction vessel.
8. At a certain temperature the equilibrium constant $K_p = 0.132$ for the reaction: $\text{PCl}_5(\text{g}) \leftrightarrow \text{PCl}_3(\text{g}) + \text{Cl}_2(\text{g})$ At equilibrium, the partial pressures of both PCl_5

and PCl_3 are 100. mmHg. What is the total pressure of the equilibrium system, in mmHg?

- a. 100. mmHg
- b. 200. mmHg
- c. 300. mmHg
- d. 400. mmHg
- e. 332 mmHg

9. Ammonium iodide dissociates reversibly to ammonia and hydrogen iodide: $\text{NH}_4\text{I}(\text{s}) \leftrightarrow \text{NH}_3(\text{g}) + \text{HI}(\text{g})$ At 400°C , $K_p = 0.215$. If 150 g of ammonium iodide is placed into a 3.00-L vessel and heated to 400°C , calculate the partial pressure of ammonia when equilibrium is reached.

- a. 0.22 atm
- b. 0.46 atm
- c. 0.11 atm
- d. 0.88 atm
- e. 1.2 atm

10. Consider the equilibrium reaction: $3\text{ClO}^-(\text{aq}) \leftrightarrow \text{ClO}_3^-(\text{aq}) + 2\text{Cl}^-(\text{aq})$ The equilibrium constant $K_c = 3.2 \times 10^3$. The following concentrations are present: $[\text{Cl}^-] = 0.50 \text{ mol/L}$; $[\text{ClO}_3^-] = 0.32 \text{ mol/L}$; $[\text{ClO}^-] = 0.24 \text{ mol/L}$. Is the mixture at equilibrium and, if not, in which direction will reaction proceed?

- a. The system is at equilibrium.
- b. The system is not at equilibrium; reaction will proceed left to right.
- c. The system is not at equilibrium; reaction will proceed right to left.
- d. The system cannot reach equilibrium since the ClO_3^- and Cl^- concentrations are not in the stoichiometric ratio.
- e. There is not enough information to tell.

11. Consider the following reaction in the gas phase: $\text{H}_2 + \text{I}_2 \leftrightarrow 2\text{HI}$ If the pressure increased by reducing the the volume of the flask,

- a. more HI will be produced.
- b. more H_2 and I_2 will be produced.
- c. the results will depend on what the amounts of each are.
- d. the amount of HI will remain the same.
- e. the equilibrium constant will change.

Solution Chemistry

1. Which of the following ions has an incorrect charge?

- a. N^{3-}
- b. Al^{3+}
- c. S^{2-}

- d. Cl⁻
- e. Mg²⁺

2. Which of the following pairs of elements would be most likely to form an ionic compound?

- a. P and Br
- b. Zn and K
- c. C and O
- d. Al and Rb
- e. F and Ca

3. What is the name of NaI?

- a. sodium iodide
- b. sodium(I) iodide
- c. sodium monoiodide
- d. sodious iodide
- e. sodium iodine

4. Which of the following combinations of names and formulas is incorrect?

- a. H₃PO₄ phosphoric acid
- b. HNO₃ nitric acid
- c. NaHCO₃ sodium carbonate
- d. H₂CO₃ carbonic acid
- e. KOH potassium hydroxide

5. Calculate the concentration of calcium ions in a saturated calcium phosphate solution. (K_{sp} for Ca₃(PO₄)₂ = 1.3×10^{-26})

- a. 1.2×10^{-5} mol/L
- b. 2.0×10^{-5} mol/L
- c. 6.6×10^{-6} mol/L
- d. 7.8×10^{-6} mol/L
- e. 8.3×10^{-6} mol/L

6. Calculate the molar solubility of silver carbonate in 1.0 M sodium carbonate solution. (K_{sp} for Ag₂CO₃ = 8.1×10^{-12})

- a. 8.1×10^{-12} M
- b. 2.8×10^{-6} M
- c. 1.4×10^{-6} M
- d. 1.4×10^{-8} M
- e. 2.0×10^{-4} M

7. Calculate the pH of a solution necessary to just begin the precipitation of Mg(OH)₂ when [Mg²⁺] = 0.001 M. (K_{sp} for Mg(OH)₂ = 1.2×10^{-11})

- a. 11
- b. 10

- c. 9
- d. 8
- e. 4

8. In qualitative analysis, the metals of Ion Group 1 can be separated from other ions by precipitating them as chloride salts. A solution initially contains Ag^+ and Pb^{2+} at a concentration of 0.10 M. Aqueous HCl is added to this solution until the Cl^- concentration is 0.10 M. What will the concentrations of Ag^+ and Pb^{2+} be at equilibrium? (K_{sp} for $\text{AgCl} = 1.8 \times 10^{-10}$; K_{sp} for $\text{PbCl}_2 = 1.7 \times 10^{-5}$)

- a. $[\text{Ag}^+] = 1.8 \times 10^{-11} \text{ M}$; $[\text{Pb}^{2+}] = 1.7 \times 10^{-6} \text{ M}$
- b. $[\text{Ag}^+] = 1.8 \times 10^{-7} \text{ M}$; $[\text{Pb}^{2+}] = 1.7 \times 10^{-4} \text{ M}$
- c. $[\text{Ag}^+] = 1.8 \times 10^{-11} \text{ M}$; $[\text{Pb}^{2+}] = 8.5 \times 10^{-5} \text{ M}$
- d. $[\text{Ag}^+] = 1.8 \times 10^{-9} \text{ M}$; $[\text{Pb}^{2+}] = 1.7 \times 10^{-3} \text{ M}$
- e. $[\text{Ag}^+] = 1.8 \times 10^{-9} \text{ M}$; $[\text{Pb}^{2+}] = 8.5 \times 10^{-6} \text{ M}$

9. Silver chloride is relatively insoluble in water (K_{sp} for $\text{AgCl} = 1.8 \times 10^{-10}$) but it is soluble in aqueous ammonia, due to the formation of the complex ion $\text{Ag}(\text{NH}_3)_2^+$. How many moles of AgCl will dissolve in 1.00 L of solution containing 6.0 moles of free NH_3 ? (K_{f} for $\text{Ag}(\text{NH}_3)_2^+ = 1.7 \times 10^7$)

- a. $9.1 \times 10^{-6} \text{ mol}$
- b. $2.9 \times 10^{-4} \text{ mol}$
- c. 0.0091 mol
- d. 0.084 mol
- e. 0.33 mol

10. What is the mass of $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ in 60.0 mL of 0.0880 M solution?

- a. 0.181 g
- b. 1.81 g
- c. 5.02 g
- d. 5.28 g
- e. none of the above

11. The freezing point of pure camphor is $178.4 \text{ }^\circ\text{C}$, and its molal freezing-point constant, K_{f} is $40.0 \text{ }^\circ\text{C}/\text{m}$. Find the freezing point of a solution containing 3.00 g of a compound of molar mass 125 g/mol in 45.0 g of camphor.

- a. $174.1 \text{ }^\circ\text{C}$
- b. $157.1 \text{ }^\circ\text{C}$
- c. $135.2 \text{ }^\circ\text{C}$
- d. $140.4 \text{ }^\circ\text{C}$
- e. $11.6 \text{ }^\circ\text{C}$

Acids and Bases

1. Calculate the hydroxide ion concentration of a solution if its pH is 6.389.

- a. 1.00×10^{-14} mol/L
 - b. 4.08×10^{-7} mol/L
 - c. 9.92×10^{-7} mol/L
 - d. 2.45×10^{-8} mol/L
 - e. none of the above
2. Which of the following is a correct description of the natural direction of a Brønsted-Lowry acid-base reaction?
- a. weaker acid + weaker base \rightarrow stronger acid + stronger base
 - b. weaker acid + stronger base \rightarrow stronger acid + weaker base
 - c. stronger acid + weaker base \rightarrow weaker acid + stronger base
 - d. stronger acid + stronger base \rightarrow weaker acid + weaker base
 - e. None of the above statements is always correct.
3. In a 0.100 M HF solution, the percent dissociation is determined to be 9.5%. Calculate the K_a for HF based on this data.
- a. 9.5×10^{-2}
 - b. 1.0×10^{-3}
 - c. 3.1×10^{-3}
 - d. 7.6×10^{-4}
 - e. 9.5×10^{-4}
4. What is the pH of a solution prepared from 0.250 mol of NH_3 dissolved in sufficient water to make 1.00 L of solution? ($K_b = 1.8 \times 10^{-5}$)
- a. 2.12
 - b. 2.67
 - c. 8.92
 - d. 11.33
 - e. 13.40
5. Which of the following reactions illustrate $\text{Al}(\text{OH})_3$ acting as a Lewis acid?
- a. $\text{Al}(\text{OH})_3 \rightarrow \text{Al}^{3+} + 3\text{OH}^-$
 - b. $\text{Al}(\text{OH})_3 + \text{OH}^- \rightarrow \text{Al}(\text{OH})_2\text{O}^- + \text{H}_2\text{O}$
 - c. $\text{Al}(\text{OH})_3 + \text{OH}^- \rightarrow \text{Al}(\text{OH})_4^-$
 - d. $\text{Al}(\text{OH})_3 + 3\text{H}^+ \rightarrow \text{Al}^{3+} + 3\text{H}_2\text{O}$
 - e. $\text{Al}^{3+} + 3\text{OH}^- \rightarrow \text{Al}(\text{OH})_3$
6. Which of the following pairs of species is not a conjugate acid-base pair?
- a. HCl and H^+
 - b. HSO_4^- and SO_4^{2-}
 - c. H_2SO_4 and HSO_4^-
 - d. H_2O and OH^-
 - e. NH_3 and NH_2^-
7. Consider each of the following pairs of acids. Which statement is correct?

- a. HClO_2 is a stronger acid than HClO_4 .
 - b. H_2SO_4 is a stronger acid than H_2SeO_4 .
 - c. H_2O is a stronger acid than HF .
 - d. H_2S is a stronger acid than H_2Se .
 - e. HS^- is a stronger acid than H_2S .
8. Consider the reaction $\text{CH}_3\text{NH}_2 + \text{H}_2\text{O} \rightarrow \text{CH}_3\text{NH}_3^+ + \text{OH}^-$ where CH_3NH_2 is methylamine and CH_3NH_3^+ is the methylammonium ion. Select the correct description of this reaction in terms of Lewis acid-base theory.
- a. Methylamine serves as a Lewis acid in the forward reaction and methylammonium ion serves as a Lewis base in the reverse reaction.
 - b. Water serves as a Lewis base in the forward reaction and the hydroxide ion serves as a Lewis base in the reverse reaction.
 - c. Methylamine serves as a Lewis base in the forward reaction and hydroxide ion serves as a Lewis acid in the reverse reaction.
 - d. Water serves as a Lewis acid in the forward reaction and methylammonium ion serves as a Lewis base in the reverse reaction.
 - e. Methylamine serves as a Lewis base in the forward reaction and hydroxide ion serves as a Lewis base in the reverse reaction.
9. What is the pH of a buffer prepared by adding 180 mL of 0.100 M NaOH to 200 mL of 0.100 M acetic acid? (K_a for $\text{CH}_3\text{COOH} = 1.8 \times 10^{-5}$)
- a. 3.79
 - b. 4.34
 - c. 4.74
 - d. 5.04
 - e. 5.70
10. Consider the titration of 50.00 mL of 0.1000 M HBr with 0.1000 M KOH. Calculate the pH after 49.00 mL of the base has been added to the 50.00 mL of HBr.
- a. 2.0
 - b. 3.0
 - c. 4.0
 - d. 6.0
 - e. 7.0
11. An aqueous solution of a weak acid, HA, is titrated with NaOH solution. The pH at the midpoint of the buffer region is 4.5. What is the K_a of the acid?
- a. 3.2×10^{-5}
 - b. 3.2×10^{-10}
 - c. 1.8×10^{-3}
 - d. 7.0×10^{-7}

Answers

Electronic Structure and Periodic Table

1. (e) 2. (e) 3. (a) 4. (c) 5. (c) 6. (b) 7. (a) 8. (a) 9. (a) 10. (d) 11. (b)

Bonding

1. (d) 2. (c) 3. (c) 4. (c) 5. (c) 6. (b) 7. (c) 8. (c) 9. (e) 10. (b) 11. (d)

Phases and Phase Equilibria

1. (b) 2. (e) 3. (b) 4. (c) 5. (c) 6. (d) 7. (d) 8. (c) 9. (d) 10. (e)

Stoichiometry

1. (e) 2. (b) 3. (c) 4. (a) 5. (c) 6. (b) 7. (c) 8. (c) 9. (d) 10. (d) 11. (c)

Thermodynamics and Thermochemistry

1. (d) 2. (c) 3. (d) 4. (d) 5. (c) 6. (d) 7. (c) 8. (d) 9. (c) 10. (a) 11. (a)

Rate Processes in Chemical Reactions—Kinetics and Equilibrium

1. (c) 2. (c) 3. (d) 4. (a) 5. (d) 6. (e) 7. (c) 8. (c) 9. (b) 10. (b) 11. (d)

Solution Chemistry

1. (e) 2. (e) 3. (a) 4. (c) 5. (d) 6. (c) 7. (b) 8. (d) 9. (e) 10. (b) 11. (b)

Acids and Bases

1. (d) 2. (d) 3. (b) 4. (d) 5. (c) 6. (a) 7. (b) 8. (e) 9. (e) 10. (b) 11. (c)

Evaluation tools for intermediate attestation

Questions List for a General and Inorganic Chemistry course (Pass-fail exam)

1. What 4 things determine how properties change?
2. What two properties determine the reaction tendency of an element?
3. What is atomic radius?
4. What are the patterns for atomic radius in the periodic table?
5. What two factors determine the force felt by the outer electrons?
6. What is the shielding effect?
7. What two factors cause atomic radius to increase in the elements of a column?
8. What causes the atomic radius to decrease across a row?
9. Why are the noble gas atoms larger than expected?
10. What configuration is found in ions?
11. Why is a positive ion smaller than its atom?
12. Why is a negative ion larger than its atom?
13. What is an oxidation number?
14. What is the pattern for oxidation numbers?
15. What property determines the reactivity of the metals?
16. What property determines the reactivity of the non-metals?
17. What is ionization energy?
18. What factors affect ionization energy?
19. What are the patterns for ionization energy in the periodic table?
20. Why do these patterns form?
21. Why are second and higher ionization energies higher?
22. Why do the alkali metals react violently with H₂O?
23. Why are the alkaline earth metals less reactive than the alkali metals?
24. What is electron affinity?
25. What factors influence electron affinity?
26. Why does electron affinity drop down a column?
27. What happens to electron affinity across a row and why?
28. Explain why Be, N, and Ne are exceptions to the trends.
29. Why does ionization energy have the greatest effect on the reactivity of metals?
30. Why does electron affinity have the greatest effect on the reactivity of non-metals?

31. An object of analytical chemistry. Quantitative and qualitative analysis. Its purposes and methods of performance.
32. Measurement of material's mass and quantity, measures.
33. Fundamental laws of stoichiometry.
34. Volumetric analysis. Methods of volumetric analysis.
35. Exact volume measurement vessels. Their preparation for analysis.
36. Definition of equivalent. Calculation of element, compound, oxidizer and reducer equivalents.
37. Definition of molar equivalent concentration, its calculation.
38. What solutions are named mono-, deci-, centinormal?
39. Definition of titer and its calculation.
40. Solutions, which are used in volumetric analysis. Define standard and titrated solutions.
41. What is the way of preparation of standard and work solutions? What is fixanals?
42. Calculation of material quantity, required to make a solution.
43. Calculation of solution's molar equivalent concentration (normality).
44. Recalculate solution's percentage concentration into molar equivalent concentration and vice versa.
45. Calculation of titrimetric analyses results.

***Questions List for an Organical Chemistry course
(Exam)***

1. Organic molecules and chemical bonding. Hybrid orbitals: sp^3 hybridization and tetrahedral bonding. Formation of π -bonds: sp^2 and sp hybridization
2. Classification of organic compounds. Common functional groups in organic compounds. Organic nomenclature. Structural formulas. Isomers.
3. Acidity and basicity of organic compounds. The Bransted-Lowry definition of acidity. The Lewis definition of acidity.
4. Reaction Mechanisms. Reaction Classification: acid-base (proton transfer) reaction, substitution reaction, elimination reaction, addition reaction.
5. Classes of Hydrocarbon. Alkanes. Cycloalkanes. Nomenclature. Structure and reactivity. Radical reactions. Conformations of Cycloalkanes.
6. Unsaturated Hydrocarbons: Alkenes. Alkynes. Nomenclature. Structure and Reactivity. Electrophilic addition reaction.
7. Aromatic Hydrocarbons. The Hückel Rule. Aromaticity. Benzene. Structure and Reactivity.

8. Arenes. Directing effects of Substituents: ortho-para directors, meta-directors.
9. Alcohols. Nomenclature of alcohols. Physical properties of alcohols. Chemical properties of alcohols: Acid-base reactions, substitutions reactions, dehydration, oxidation.
10. Phenols. Nomenclature. Physical and chemical properties of phenols.
11. Aldehydes and ketones. The carbonyl group. Nomenclature. Physical properties. Chemical properties: Oxidation-reduction reactions of carbonyl compounds, addition reactions.
12. Aldehydes and ketones. Formation of acetals and ketals, reactivity of the α -carbon atom, the aldol condensation.
13. Carboxylic Acid. Nomenclature. The structure of the acyl groups. Physical properties of carboxylic acid. Chemical properties of carboxylic acid: acidity, nucleophilic acyl substitution, reactivity of the α -carbon atom.
14. Derivatives of the carboxylic acid: amides, acid chlorides, anhydrides, esters.
15. Hydroxy acids. Configuration of molecules. Mirror Images and Chirality. Optical Activity. Fischer Projection Formulas. Enantiomers, diastereomers.
16. Carbohydrates. Classification of Carbohydrates. Monosaccharides. Glucose. Fructose. Chemical Properties.
17. Monosaccharides. Glucose stereoisomers. Structures of monosaccharides: open-chain and cyclic form. Glycosides.
18. Disaccharides. Types of glycosidic linkages for glucose-containing disaccharides. Reducing and non-reducing disaccharides. Maltose, cellobiose, lactose, sucrose.
19. Polysaccharides. Building blocks of polysaccharides. Classification. Starch. The structure, properties and hydrolysis. Amylose and amylopectin. Cellulose. The structure, chemical properties, practical use.
20. Amino acids. Acid-base properties of α -amino acids. Isoionic point. Chemical properties.
21. Peptides. Protein. Protein Structure. Qualitative reactions of proteins
22. Classification of Lipids. Fatty Acids. Types of Fatty Acids. Triglycerides.
23. Triglycerides. Chemical Properties.
24. Phosphoglycerides. Waxes. Glycolipids. Sterols.

Evaluation tools for current attestation

Lab Grades for Each Experiment Lab grades are assigned for each experiment considering these factors:

1. 10%: pre-lab assignment and pre-lab discussion.
2. 10%: safe behavior in lab.
3. 10%: laboratory work including taking notes.
4. 10%: post-lab discussion.
5. 60%: Laboratory Report.

Lab reports are graded on a 0-60 pt scale.

Pre-lab assignments are graded and returned. The remaining points (safety, lab work, post-lab) are presumed, and a lab instructor will inform students only if they do not earn full credit. Deducted points are communicated by the time lab reports are returned. Typically students lose points when they do not attend pre- or post-lab discussions, do something unsafe in lab, or do not use their lab notebooks properly.



MINISTRY OF EDUCATION AND SCIENCE OF THE RUSSIAN FEDERATION
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SCHOOL OF BIOMEDICINE

LECTURE NOTES
Chemistry, Medical Chemistry
Specialty 31.05.01 General Medicine
Full-time training

Vladivostok,
2018

Theme __. Topic name (___hours), using the active learning method - _____ (if the use of MAO is foreseen)

A description of the use of the method of active learning in this lecture (if available).

The following is the material of the lecture directly in one of the following options: the full text of the lecture, a short supporting summary, a detailed (detailed, containing points and sub-items) plan. A lecture can also be presented as a media presentation or video lecture. Media presentations are applied to the UMCD in a separate folder, video lectures are placed in the electronic training course of the discipline (ESM) in the LMS Blackboard FEFU.

Thus, materials of all lectures in accordance with the MEP should be presented in the teaching room.



MINISTRY OF EDUCATION AND SCIENCE OF THE RUSSIAN FEDERATION
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SCHOOL OF BIOMEDICINE

**PRACTICAL CLASSES MATERIALS
AND LABORATORY MANUAL**
Chemistry, Medical Chemistry
Specialty 31.05.01 General Medicine
Full-time training

**Vladivostok,
2018**

Practical Classes

(Total workload: 36 hrs., including 8 hrs. interactive learning)

Class 1. Mendeleev's Periodic Law and Periodic Table (2 hrs.)

1. Complete the following table concerning subatomic particles: Relative Charge Relative Mass (in amu) Location in Atom Proton Neutron Electron

2. The figure below represents a box from the Periodic Table. Identify what each number is.

3. Define the atomic number (Z) and mass number (A) of an atom.

Z =

A =

Which of these is used to identify an atom of a particular element?

4. Write the complete isotope symbol for the following atoms or ions, including the element symbol.

- An atom that has 37 protons, 50 neutrons and 37 electrons
- An ion that has 16 protons, 17 neutrons and 18 electrons
- An ion that has 82 protons, 125 neutrons and 78 electrons

5. Using the Periodic Table as a guide, determine how many protons and neutrons are in the following atoms:

^{27}Al ^{14}N ^{65}Cu

If an atom of Al is electrically neutral, how many electrons does it have?

If an atom of N has 10 electrons, what charge will it have?

If an atom of Cu has a +2 charge, how many electrons does it have?

6. Complete the following: Isotopes are atoms that have

- the same number of (or the same number)
- different numbers of (or different number) ^{37}Rb 85.47

8. Is the atomic mass of an element the same as the mass number of an atom?

Circle one: Yes / No Why or why not?

9. Two naturally occurring isotopes of gallium are ^{69}Ga (mass = 68.99 amu) and ^{71}Ga (mass = 71.01 amu). The atomic weight of gallium is 69.70 amu.

Which one of these isotopes is found in the greatest natural abundance? Explain your reasoning.

10. Calculate the atomic mass of silicon. Show your work clearly in the space provided.

Class 2. Basic chemical calculations (2 hrs.)

Mole Quantities, Particle Quantities and Mass

1. What is the numeric value of the mole? 1 mole = What is this value also known as?

2. A balloon contains 2.58×10^{24} molecules of methane, CH_4 . How many moles of methane are in this balloon?
3. An engagement ring contains 3.97×10^4 μmoles of gold. How many individual gold atoms are in this ring?
4. Exactly how is the formula mass of a substance different from its molar mass?
5. Calculate the molar masses of each of the following substances (don't forget your units):
 - a. silicon (in its natural state)
 - b. phosphorus pentachloride
 - c. lithium sulfite
6. Why does 1 mole of silicon have a different mass from 1 mole of phosphorus pentachloride (see 5a and 5b)? Don't they both consist of the same number of particles?
7. What is the mass (in grams) of 0.121 moles of phosphorus pentachloride?
8. A rock contains 3.64 kg of lithium sulfite. How many moles of lithium sulfite are in this rock?
9. What is the mass (in grams) of 4.78×10^{24} atoms of silicon?
10. Aspartame ($\text{C}_{14}\text{H}_{18}\text{N}_2\text{O}_5$) is an artificial sweetener marketed as Nutra-Sweet. A packet of Nutra-Sweet is found to contain 500. mg of aspartame. How many individual molecules of aspartame are in this packet?

Formulas as Conversion Factors

11. How many individual P atoms are found in 1.76×10^{22} molecules of tetraphosphorus trisulfide?
12. How many moles of O are in 0.575 moles of calcium perchlorate?
13. How many individual C atoms are found in 308 grams of benzene, C_6H_6 ?
14. What mass (in grams) of Na is found in 10.84 moles of sodium oxide?

Percent Composition of Compounds

15. Calculate the mass percent of each element in aluminum bicarbonate, $\text{Al}(\text{HCO}_3)_3$.
16. Arrange the following compounds in order of increasing mass % of phosphorus: H_3PO_4 , P_4O_{10} and PCl_3 .
17. When 54 grams of a compound containing the elements N and O (only) is analyzed, it is found to contain 40 grams of O. What mass of N would be found in a 2.36 gram sample of this same compound?

18. What is the mass (in grams) of hydrogen in 253 grams of propane, C_3H_8 ?

Empirical and Molecular Formulas

19. When a 34.61 gram sample of a "chromium oxide" compound is analyzed, it is found to contain 21.43 grams of Cr, the remainder being O. (a) Determine this empirical formula of this compound, and (b) state whether the chromium ions in this compound have a +3, +4 or a +6 charge.

20. Aspirin has a percent composition of 60.00% C, 4.48% H and 35.53% O. Determine its empirical formula.

21. The compound benzene is determined to have an empirical formula C_1H_1 . If the true molar mass of benzene is 78.0 g, what is the true chemical formula of benzene?

22. Adipic acid is used in the manufacture of nylon. When a 68.4 g sample of adipic acid is analyzed it is found to consist of 33.8 g C, 4.7 g H and 29.9 g O. a. What is the empirical formula of adipic acid?

b. If the molar mass of adipic acid is 146 grams, what is the true molecular formula of adipic acid?

Calculation of solution

- 13.9 g of solid KNO_3 are dissolved in enough water to make 175 mL of an aqueous potassium nitrate solution. What is the molarity of this solution?
- How many moles of HCl are dissolved in 25.0 mL of a 4.36 M hydrochloric acid solution? How many grams of HCl are dissolved in this solution?
- What volume (in mL) of a 0.650 M aqueous glucose solution ($C_6H_{12}O_6$) contains 5.48 grams of glucose?
- A solution is prepared by mixing 36.0 g of $CaCl_2$ with 271 g water. The density of this solution is 1.09 g/mL. a. What is the molarity of this solution?
b. Suppose 55.0 mL of water are now added to this solution. What is its new molarity?
- What volume of 18.0 M H_2SO_4 (aq) and water would you need to prepare 5.00 L of 2.65 M H_2SO_4 (aq)?
- Oxygen gas can be generated by the catalytic decomposition of an aqueous solution of hydrogen peroxide: $2 H_2O_2 (aq) \rightarrow 2 H_2O (l) + O_2 (g)$ If the complete decomposition of 48.2 mL of H_2O_2 (aq) yields 2.33 grams of oxygen, what is the molarity of this solution?
- Consider the reaction: $3 CuSO_4 (aq) + 2 Na_3PO_4 (aq) \rightarrow Cu_3(PO_4)_2 (s) + 3 Na_2SO_4 (aq)$ a. What volume (in milliliters) of a 2.93 M aqueous solution of sodium phosphate will completely react with 41.8 mL of a 1.65 M aqueous solution of copper(II) sulfate?
b. What mass of solid copper(II) phosphate precipitate (in grams) will be produced by this reaction?
- During a certain titration, 32.9 mL of 0.124 M aqueous barium hydroxide was required to completely neutralize 25.0 mL of nitric acid. Determine the molarity of the nitric acid.
- Consider the reaction: $Al_2(CO_3)_3 (s) + 6 HCl (aq) \rightarrow 3 CO_2 (g) + 3 H_2O (l) + 2 AlCl_3 (aq)$ What volume (in L) of carbon dioxide gas will be

collected at STP if 4.00 grams of aluminum carbonate are mixed with 34.7 mL of 6.00 M hydrochloric acid?

10. Consider the following scenarios and determine if the substances in question are ionic, polar covalent or nonpolar covalent. Substance A is insoluble in water, but dissolves readily in the solvent CS₂. Substance B dissolves readily in water. The resulting solution conducts electricity. Substance C dissolves readily in water. The resulting solution does not conduct electricity.
11. Consider the ionic compound magnesium fluoride MgF₂.
 - a. Sketch a diagram showing the interactions between the ions and water when this compound is dissolved.
 - b. Do you expect MgF₂ to be more or less soluble than MgCO₃ in water? Why?
 - c. Do you expect MgF₂ to be more or less soluble than NaF in water? Why?

Class 3. Classes of inorganic compounds (2 hrs.)

1. How are the following types of compounds recognized from their formulas? Ionic Covalent Acid
2. When do parentheses appear in the formulas of ionic compounds?
3. Do Greek Prefixes appear in the names of ionic or covalent compounds? Explain why they are used.
4. Do Roman Numerals appear in the names of ionic or covalent compounds? Explain why they are used.
5. What is the relationship between the number of hydrogens in an acid and the charge on the anion that they are combined with?

Class 4. Chemical bonds and structure of chemical compounds (2 hrs.)

4. Localized bonding and hybrid atomic orbitals.
5. sp-, sp²-, sp³-hybridizations
6. Delocalized bonding and molecular orbitals.

Class 5. Chemical kinetics (2 hrs.)

Reaction Rates

1. Consider the reaction: $\text{Mg (s)} + 2 \text{HCl (aq)} \rightarrow \text{MgCl}_2 \text{ (aq)} + \text{H}_2 \text{ (g)}$ Using collision theory, explain why only mild bubbling is observed when 0.1M acid is used, but vigorous bubbling is observed when 6.0M acid is used in this reaction.
2. Using collision theory, explain why the same quantity of water will evaporate (H₂O liquid → H₂O gas) more quickly out of a wide shallow dish than a tall narrow glass.

3. On Figure A, sketch in any changes that would occur if the reaction temperature was increased. On Figure B, sketch in any changes that would occur if a catalyst was added to the reaction. Would these changes cause the reactions to slow down or speed up? Are the reactions depicted exothermic or endothermic? .

Chemical Equilibrium

4. Consider a reversible reaction that achieved a state of equilibrium. Which of the following statements about the state of equilibrium are true and which are false?
- At equilibrium, the concentrations of the reactants and products are equal.
 - At equilibrium, the rates of the forward and reverse reactions are equal.
 - At equilibrium, the reversible reaction stops.
 - At equilibrium, the concentrations of the reactants and products remain constant. In the space below, explain why the false statements are false.
6. Consider the following reversible reaction: $2 \text{HBr}(\text{g}) \rightleftharpoons \text{H}_2(\text{g}) + \text{Br}_2(\text{g})$ The equilibrium constant (K_{eq}) for this reaction is 1.54×10^{-6} at room temperature. If this reaction is allowed to achieve equilibrium at RT in a sealed flask, what should be found in it?
- Only H_2 and Br_2 Explain your choice (briefly!).
 - Only HBr
 - A mixture containing more HBr than H_2 and Br_2
 - A mixture containing more H_2 and Br_2 than HBr
7. Write equilibrium expressions for the following reactions:
- $4 \text{NH}_3(\text{g}) + 7 \text{O}_2(\text{g}) \rightleftharpoons 4 \text{NO}_2(\text{g}) + 6 \text{H}_2\text{O}(\text{g})$
 - $2 \text{NO}(\text{g}) + \text{Cl}_2(\text{g}) \rightleftharpoons 2 \text{NOCl}(\text{g})$
8. Consider the following reversible reaction at equilibrium: $2 \text{CH}_4(\text{g}) \rightleftharpoons \text{C}_2\text{H}_2(\text{g}) + 3 \text{H}_2(\text{g})$ At 1700°C , the equilibrium concentrations of the reactants and products are measured as $[\text{CH}_4] = 0.0203 \text{ M}$, $[\text{C}_2\text{H}_2] = 0.0451 \text{ M}$ and $[\text{H}_2] = 0.112 \text{ M}$. Determine the value of the equilibrium constant (K_{eq}) at this temperature. Is the reaction reactant-favored or product-favored at this temperature?
9. Consider the following reversible reaction at equilibrium: $2 \text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2 \text{SO}_3(\text{g})$ $K_{\text{eq}} = 4.32$ at 600°C At 600°C , the following equilibrium concentrations are measured: $[\text{O}_2] = 0.0449 \text{ M}$ and $[\text{SO}_3] = 0.261 \text{ M}$. Determine the equilibrium concentration of SO_2 at this temperature.
10. Consider the following reversible reaction at equilibrium: $\text{N}_2(\text{g}) + 3 \text{H}_2(\text{g}) \rightleftharpoons 2 \text{NH}_3(\text{g})$ $K_{\text{eq}} = 0.0584$ at 775°C When this reaction achieves equilibrium in a 1.50 L sealed flask at 775°C , there are 0.180 moles of N_2 and 0.424 grams

of H₂ present in the flask. Calculate the pressure (in atm) exerted by NH₃ in this equilibrium mixture.

Applying Le Chatelier's Principle

11. Nitrogen monoxide and nitrogen dioxide produce dinitrogen trioxide in the following reversible, exothermic reaction, shown here at equilibrium: $\text{NO (g)} + \text{NO}_2 \text{ (g)} \rightleftharpoons \text{N}_2\text{O}_3 \text{ (g)}$ Predict the shifts that will occur (left, right or none) in accordance with Le Chatelier's Principle if the equilibrium system is disturbed in the following ways:
- removing NO
 - adding N₂O₃
 - decreasing the temperature
 - increasing the pressure
 - adding a catalyst
 - Briefly explain your answer to part c).
12. Suppose you are a chemical engineer in a plant that produces ethylene (C₂H₄) from the endothermic equilibrium reaction: $\text{C}_2\text{H}_6 \text{ (g)} \rightleftharpoons \text{C}_2\text{H}_4 \text{ (g)} + \text{H}_2 \text{ (g)}$ In order to obtain the highest possible yield of C₂H₄ (increase the amount), would you:
- add or remove hydrogen?
 - add or remove ethane (C₂H₆)?
 - increase or decrease the temperature?
 - increase or decrease the volume?
 - Briefly explain your answer to part d).

Class 6. pH. Buffers. (2 hrs.)

- How does the concentration of the buffer affect the buffer capacity?
- What differences would be observed if HCl were used in place of NaOH?
- Write equations to show how a buffer works.

Class 7. Hydrolysis of salts (2 hrs.)

- Hydrolysis salts.
- Calculation of hydrolysis constant
- Calculation of hydrolysis degree

Class 8. Oxidation-Reduction reactions (2 hrs.)

Write balanced chemical equations for each of the following reactions:

- Gaseous nitrogen dioxide reacts with water to produce nitric acid and gaseous nitrogen monoxide.
- Rubidium metal (s) + water (l) → rubidium hydroxide (aq) + hydrogen (g)
- Iodine (s) + chlorine (g) → iodine trichloride (l)

4. Sodium bicarbonate when heated will form solid sodium carbonate, carbon dioxide and water vapor.

5. Magnesium acetate (aq) + ammonium sulfide (aq) → magnesium sulfide (s) + ammonium acetate (aq).

6. C_6H_{14} (l) + oxygen (g) → carbon dioxide (g) + water (l) [C₆H₁₄ is named hexane]

7. Iron(III) oxide (s) + carbon (s) → iron (s) + carbon dioxide (g)

8. Potassium carbonate (aq) + nitric acid (aq) → potassium nitrate (aq) + carbon dioxide (g) + water (l)

9. Ammonium nitrite (s) → nitrogen (g) + water (l)

10. Gaseous sulfur dioxide reacts with oxygen forming gaseous sulfur trioxide.

11. Cobalt(III) bromide (aq) reacts with pure chlorine forming pure bromine and cobalt(III) chloride (aq).

12. Silver (s) + S₈ (s) → silver sulfide (s) [S₈ is a naturally occurring form of sulfur]

13. Ammonia (g) + oxygen (g) → nitrogen monoxide (g) + water (l)

14. Aqueous barium hydroxide reacts with phosphoric acid forming water and solid barium phosphate.

15. Aluminum (s) + copper(II) sulfate (aq) → aluminum sulfate (aq) + copper (s)

16. When heated, solid lithium chlorate decomposes to produce solid lithium chloride and oxygen gas.

17. Benzene liquid (C₆H₆) burns in oxygen to form carbon dioxide and water (and heat).

18. Silicon dioxide (s) + hydrofluoric acid (aq) → silicon tetrafluoride (g) + water (l)

Class 9. Coordination compounds (2 hrs.)

4. Structure of coordination compounds

5. The Formation of Complex Ions.

6. Dissociation and basic chemical properties of complex compounds

Laboratory practical

(Total workload: 54 hrs., including 24 hrs. interactive learning)

Experiment № 1. The Density of Liquids and Solids (4 hrs.)

Objectives

The objectives of this laboratory are:

- a) To determine the density of pure water;
- b) To determine the density of aluminum (applying the technique of water displacement) and to use this value to determine the thickness of a piece of aluminum foil;
- c) To measure the mass and volume (via measured dimensions) of several cylinders of an unknown material, and to determine the density of this material via graphical analysis of the collected data.

Procedure

Safety

Be especially careful when adding the aluminum to your graduated cylinder, as the glass could break. Tilt the graduated cylinder and allow the pellets to gently slide to the bottom.

Materials and Equipment

100-mL graduated cylinder, metric ruler, aluminum pellets, small beaker, aluminum foil, thermometer, electronic balance, distilled water, tube of unknown solid cylinders and graph paper.

Experimental Procedure

Part A: The Density of Water

1. Using the electronic balance, obtain the mass of your 100-mL graduated cylinder. Make sure it is dry before you weigh it.
2. Add 20-25 mL of distilled water to the graduated cylinder. Precisely measure this volume of water. Then measure the combined mass using the electronic balance.
3. Add another 20-25 mL of distilled water to the graduated cylinder. Again, precisely measure this volume of water, and then measure the combined mass using the electronic balance.

4. Repeat Step 3 to obtain a third set of mass and volume measurements.
5. Use your thermometer to record the temperature of the water in your graduated cylinder.
6. Analysis: Subtract the mass of the empty cylinder from each combined mass measurement to obtain three mass measurements of water. Use the three sets of mass and volume measurements to calculate three density values for water. Then take the average of these three density values. Finally, look up the true density of water at the temperature used, and evaluate the accuracy of your average density value by calculating your percent error.

Part B: The Density of Aluminum and the Thickness of Foil

The Density of Aluminum

1. Using the electronic balance, obtain the mass of a clean, dry small beaker.
2. Obtain a sample of aluminum from your instructor. Transfer all the pellets to the beaker, and measure the mass of the beaker and pellets.
3. Pour 30-35 mL of water into your 100-mL graduated cylinder. Precisely measure this volume.
4. Carefully add all the aluminum pellets to the water, making sure not to lose any water to splashing. Also make sure that the pellets are all completely immersed in the water. Measure the new volume of the water plus the pellets.
5. When finished, retrieve and dry the aluminum pellets and return them to your instructor.
6. Analysis: Use your measured mass and volume (obtained via water displacement) of the aluminum pellets to calculate the density of aluminum. Then look up the true density of aluminum and evaluate your accuracy by calculating your percent error.

The Thickness of Aluminum Foil

7. Now obtain a rectangular piece of aluminum foil from your instructor. Use the ruler to measure the length and width of the piece of foil.
8. Measure the mass of the foil using the electronic balance.
9. When finished, return the foil to your instructor and the ruler to the stockroom.
10. Analysis: Use these measurements along with the density of aluminum to calculate the thickness of the foil.

Part C: Graphical Analysis of Mass and Volume Data of an Unknown Solid

1. Check out a ruler from the stockroom and obtain a tube containing cylindrical pieces of an unknown solid material from your instructor. Record the ID Code of the unknown solid on your report form.

2. Using the ruler, measure the dimensions (diameter and height) of each cylindrical object. Start with the smallest object first and progress in order of increasing object size.

3. Measure the mass of each cylindrical object using an electronic balance. Again, start with the smallest object first and progress in order of increasing object size.

4. Replace all the objects in the tube and return the tube to your instructor.

5. Analysis: Use the measured dimensions to calculate the volume of each solid object. Then, on the graph paper supplied, plot the mass (Y) versus the volume (X) of each measured object. Add a bestfit line to this plot. Calculate the slope of this line, which is the density of the unknown solid. Then use this density to identify the unknown material analyzed. Your unknown material is one of the substances listed in the table below.

<i>Substance</i>	<i>Density (g/cm³)</i>
Polyvinylchloride (PVC)	1.35
Maple	0.77
Acrylic	1.16
Polytetrafluoroethylene (Teflon)	2.20
Polypropylene	0.90
Aluminum	2.71
Polyurethane	1.23

Experiment № 2. Rates of Chemical Reactions - A Clock Reaction (4 hrs.)

Objectives

This experiment is designed to study the kinetics of a chemical reaction. The reaction involves the oxidation of iodide ions by bromate ions in the presence of acid. The objective of this lab is to determine the rate law and activation energy for the redox reaction between the iodide ion (I⁻), bromate ion (BrO₃⁻) ion, and hydrogen ion in water

Experimental Procedure

A. Equipment needed

In lab: 100-1000µL digital pipets, 10mL beakers Class sets of hot plates and thermometers are stored in the lab. Ask your instructor for the location.

Chemicals in lab:

DI water 0.0100 M KI 0.100M HCl
 0.00100 M Na₂S₂O₃ 0.0400 M KBrO₃ or NaBrO₃ 1% starch solution.
 0.100M CuSO₄

From stock room: cotton swabs, stopwatches, bucket of ice, 1.5 dram (~5ml) glass vials with caps.

B. Waste Disposal

All solutions except rinses must be disposed of in the proper waste container in the hood.

C. Before Starting Experimental Work (Before Class)

1. In your notebook, enter the experiment title, date, your name and name of partner. Number each page.
2. Write the purposes of the lab. (There are several!) Include all of the different kinetic values that will be determined in this lab (for example, “x”, “y”, etc. ...)
3. Write out the overall and monitoring reactions.
4. Describe in your own words how the reaction rate is determined using the monitoring reaction
5. Write an executive summary of the procedures described in this lab. Include all major procedural steps that you will need to follow.
6. Record the information from Tables 1, 2 & 3 in your notebook or tape them in your notebook.

D. Procedure 1. Finding of the Reaction Rate Law in the absence of a catalyst. (Exp# A1 to A7)

Table 1: Experimental parameters for determining the reaction rate law.

Exp.	Initial Reagents					Final Reagent	Temp.
	KI 0.0100 M	DI H ₂ O	HCl 0.100 M	Starch 1%	Na ₂ S ₂ O ₃ 0.00100 M	KBrO ₃ [*] 0.0400 M	
A1	0.300 mL	0.600 mL	0.300 mL	0.150 mL	0.150 mL	0.300 mL	Room Temp
A2	0.600 mL	0.300 mL	0.300 mL	0.150 mL	0.150 mL	0.300 mL	Room Temp
A3	0.900 mL	---	0.300 mL	0.150 mL	0.150 mL	0.300 mL	Room Temp
A4	0.300 mL	0.300 mL	0.300 mL	0.150 mL	0.150 mL	0.600 mL	Room Temp
A5	0.300 mL	---	0.300 mL	0.150 mL	0.150 mL	0.900 mL	Room Temp
A6	0.300 mL	0.300 mL	0.600 mL	0.150 mL	0.150 mL	0.300 mL	Room Temp
A7	0.300 mL	---	0.900 mL	0.150 mL	0.150 mL	0.300 mL	Room Temp

*Note: NaBrO₃ may be used in place of KBrO₃.

1. In 6 separate clean 10mL beakers, obtain enough of the solutions listed in Table 1 to fill each beaker about 2/3 full.

2. You will need 7 pipet tips that fit the adjustable 100-1000 μ L digital pipets. Set them up as in the picture below so that you can dedicate one tip to each reagent that you will use during the experiment.

3. Measure the temperature of the starch solution. Assume that all reagents are at the same room temperature and record this temperature in Table A for runs A1 thru A7.

4. Obtain a clean 5mL glass vial and label it A1.

5. Fill the vial with the “Initial Reagents” for Experiment A1 using an adjustable digital pipet.

6. One partner should get ready to start a timer. Another partner should obtain the listed amount of KBrO₃ in a digital pipet.

7. At the same time, one partner should start the timer while the other one adds the KBrO₃ to vial A1. Cap the vial, shake to mix the reagents, and place the vial on a white surface to make it easy to see color changes.

8. When the solution becomes colored (usually blue, sometimes brown), stop the timer.

9. In Table A, record the volumes used and the time it took (in seconds, to the nearest second) for the solution to turn color.

10. Label a 250-500 mL beaker “Waste”. Empty the contents of the vial into the waste beaker. Rinse it 3-4 times with DI water. (The rinses can go down the sink.) Drain the vial upside down on a paper towel for 30 seconds and then dry the interior with a cotton swab.

11. Repeat steps 4-10 for vials A2 thru A7.

E. Procedure 2. The Effect of Copper Ion as a Catalyst. (Experiments B1-B-3)

Experiment B1 thru B3: These use the same reagents as experiment A1 with the addition of a Cu²⁺ catalyst and less water.

1. Repeat steps D4-D10 for vials B1 thru B3

Table 2: Experimental Parameters for Evaluation of a Catalyst.

Exp.	Initial Reagents						Final Reagent	Temp.
	KI 0.0100 M	DI H ₂ O	HCl 0.100 M	Starch 1%	Na ₂ S ₂ O ₃ 0.00100M	CuSO ₄ 0.100M	KBrO ₃ 0.0400 M	
B1	0.300 mL	0.450 mL	0.300 mL	0.150 mL	0.150 mL	0.150 mL	0.300 mL	RT
B2	0.300 mL	0.300 mL	0.300 mL	0.150 mL	0.150 mL	0.300 mL	0.300 mL	RT
B3	0.300 mL	0.150 mL	0.300 mL	0.150 mL	0.150 mL	0.450mL	0.300 mL	RT

F. Procedure 3. Determination of the Activation Energy via the Effect of Temperature on Reaction Rate. (Experiments C1 thru C2)

Table 3: Experimental Parameters for Determining the Activation Energy.

Exp.	Initial Reagents					Final Reagent	Temp.
	KI 0.0100 M	DI H ₂ O	HCl 0.100 M	Starch 1%	Na ₂ S ₂ O ₃ 0.00100 M	KBrO ₃ 0.0400 M	
C1	0.300 mL	0.600 mL	0.300 mL	0.150 mL	0.150 mL	0.300 mL	0-2°C
C2	0.300 mL	0.600 mL	0.300 mL	0.150 mL	0.150 mL	0.300 mL	39-41°C

Experiment C1: Use the same reagents as experiment A1, but at a lower temperature, between 0–2°C.

1. Obtain a clean 5mL glass vial and label it C1.
2. Fill the vial with the “Initial Reagents” for Experiment C1 using an adjustable digital pipet.
3. Place Vial C1 and the small beaker of 0.0400M KBrO₃ in an ice-water bath in a medium beaker. Keep the level low to prevent the beaker from tipping over. Allow them to come to temperature for at least 5 minutes.
4. One partner should get ready to start a timer. The second partner should obtain the listed amount of KBrO₃ in a digital pipet from the now cold reagent.
5. At the same time, one partner should start the timer while the other one adds the KBrO₃ to vial C1. Cap the vial, shake to mix the reagents, and place the vial back in the ice bath near the edge where the color can be easily observed.
6. When the solution becomes colored (blue to brown), stop the timer.
7. In Table A, record the volumes used and the time it took for the solution to turn dark.
8. Measure the temperature of the solution and record it in Table A.
9. Empty the contents of the vial into the waste container. Rinse it 3-4 times with DI water. (The rinses can go down the sink.) Drain the vial upside down on a paper towel for 30 seconds and then dry the interior with a cotton swab.

Experiment C2: Use the same reagents as experiment A1, but at a higher temperature, between 39-41°C.

1. Obtain a clean 5mL glass vial and label it C2.
2. Fill the vial with the “Initial Reagents” for Experiment C2 using an adjustable digital pipet.
3. Place Vial C2 and the small beaker of 0.0400M KBrO₃ in warm water bath in a medium beaker on a hot plate. (Start with hot water out of the sink tap. It will usually already be close to the correct temperature.) Keep the level low to

prevent the beaker from tipping over. Allow them to come to temperature for at least 5 minutes.

4. One partner should get ready to start a timer. The second partner should obtain the listed amount of KBrO_3 in a digital pipet from the now warm reagent.

5. At the same time, one partner should start the timer while the other one adds the KBrO_3 to vial C2. Cap the vial, shake to mix the reagents, and place the vial on a white surface to make it easy to see color changes.

6. When the solution becomes colored (blue to brown), stop the timer.

7. In Table A, record the volumes used and the time it took for the solution to turn dark.

8. Measure the temperature of the solution and record it in Table A.

9. Empty the contents of the vial into the waste container. Rinse it 3-4 times with DI water. (The rinses can go down the sink.) Drain the vial upside down on a paper towel for 30 seconds and then dry the interior with a cotton swab.

G. Cleanup

1. Rinse all pipet tips with water and share with your partner to use during the rest of the semester. Do not throw them away.

2. Rinse vials and caps with DI water and return to equipment area.

3. Transfer the contents of your waste beaker to the central waste container.

4. Clean all remaining items used and return them to your equipment locker.

5. Wipe down your work area with a damp paper towel.

IV. Further Instructions

A. Required Calculations and Graphs

1. *Data Sharing*: Obtain data for Experiments A1 thru C2 from 5 other student groups.

2. *Determination of average reaction times*: If the times recorded for a given experiment vary significantly between the groups, perform a Q-test on any suspect values. Average any values retained by the Q-test. Record values in Table 1 to one decimal place for use in further calculations in order to prevent rounding errors. Show a sample calculation for Q and the average for the first experiment that required a Q-test.

3. *Determination of initial rates of reactions*: Calculate the initial rate of reaction for each of the experiments that was performed. This will require that you first calculate the concentration of each reactant in the reaction mixture. Remember that all reagents dilute each other when you mix them. Create Table B in your

notebook. Record your calculated values in Table B. Show a sample calculation for experiment A1.

4. *Determination of rate law exponents:* Calculate the value for each exponent in the rate law (x , y , z , and p for the catalyst) using the method of initial rates. For each determination of exponents x , y , & z , there are three possible combinations of experiments that can be used to estimate the exponent. For the exponent of I^- (“ x ”), use experiments A1/A2, A1/A3, and A2/A3 to estimate the exponent three times and average the values you find. Then, choose the closest whole number. Show the full calculation for all of the values of “ x ” to 3 sig figs as shown on page 4, average the three values of the exponent and report the average rounded to the nearest whole number. Also, determine three estimates of the exponent of BrO_3^- (“ y ”) and H^+ , (“ z ”) to 3 sig figs and one value for the copper catalyst exponent, (“ p ”) to 3 sig figs. You do not have to show all explicit calculations, but include the ratio of the rates ($rate_2/rate_1$), the ratio of the concentrations ($conc.2/conc.1$), the natural logs of these two values, and the value of the exponent that you get by dividing the two natural log values. All values should be summarized in Table C. Determine the average value for x , y , z and p and then round the final values to the nearest whole number. (*Note: -0.5 to +0.5 rounds to 0; 0.5 to 1.5 rounds to 1; 1.5 to 2.5 rounds to 2.*)

5. *Determination of the rate constants:* Once the rate law is determined, any experiment can be used to determine the rate constant k by simply plugging in the appropriate values for an experiment.

a) Calculate the value of the rate constant without a catalyst for the first 7 experiments (A1-A7). Show the first calculation. Calculate and report the average of the calculated values for the rate constant k at room temperature.

b) Calculate the value of the rate constant without catalyst at low temperature. (Exp C1)

c) Calculate the value of the rate constant without catalyst at elevated temperature. (Exp C2)

d) Calculate the value of the rate constant with a catalyst for the 3 experiments that used the catalyst (B1 – B3). Show all calculations. Calculate and report the average of the calculated values for the rate constant k .

6. *Determination of the Rate Laws:*

a) Write the three rate laws for the overall reaction without the catalyst, using the values of x , y , z , and k determined above. (i.e., at low temp., room temp., and high temp.)

b) Write the rate law for the overall reaction with the catalyst at room temp, using the values of x , y , z , p , and k determined above.

7. *Determination of the Activation Energy*

Using the three k values obtained in calculations 5a (average), 5b and 5c, Fill in Table D.

Graph: Prepare an Arrhenius plot of $\ln k$ versus $1/T$. This should lead to a linear plot, with the slope being $-E_a/R$. You must prepare your plot in such a way (using Excel) that the computer fits a best-fit line through the data and reports the slope for you to determine E_a . This can be done by selecting Chart/Add Trendline from the menu. On the "Type" tab of the dialog box, select "linear". On the "options" tab, be sure to check the box to include the equation of the line on the graph. Do NOT set the intercept to zero. Determine the activation energy (in kJ/mol) from the linear least squares fit slope of the graph. Show your calculation. Do not use the two-point slope formula!

B. In-Class Work

1. Complete Table A
2. Complete the first row in Table B (A1), showing all calculations.
3. Complete the second row in Table B (A2). (You do not need to show the calculations for this row or any of the other rows.)
4. Calculation #3 for "x" using experiments A1 & A2.

C. At-Home Work

1. Complete all tables and calculations not completed during class. Since all groups were required to perform the experiments within a 2°C range, use the temperatures from your group to perform all calculations.
2. Prepare the Excel Graph and do the calculations associated with the determination of the activation energy.

D. Experimental Discussion and Summary:

After all calculations are complete:

1. Explain why the reaction rate changes as concentration changes.
2. Explain why the reaction rate changes as temperature changes.
3. Explain why the reaction rate changes when a catalyst is added.
4. In the method of initial rates that was used in this experiment, why do we pick the "initial" portion of the reaction to study?
5. Write an executive summary that addresses the purpose of this lab, describing what was done in general terms and what was learned in this experiment, including any issues that were encountered and dealt with.
6. Present a table that summarizes what was determined from this experiment. (e.g., rate laws, rate constants & activation energy under varying conditions.)

Experiment № 3. Chemical Equilibrium and Le Chatelier's Principle (4 hrs.)

Objectives

The objective of this lab is to observe the effect of an applied stress on chemical systems at equilibrium.

Procedure

Safety

All of the acids and bases used in this experiment (NH_3 , HCl , HNO_3 and NaOH) can cause chemical burns. In particular, concentrated 12M HCl is extremely dangerous! If any of these chemicals spill on you, immediately rinse the affected area under running water and notify your instructor. Also note that direct contact with silver nitrate (AgNO_3) will cause dark discolorations to appear on your skin. These spots will eventually fade after repeated rinses in water. Finally, in Part 4 you will be heating a solution in a test tube directly in a Bunsen burner flame. If the solution is overheated it will splatter out of the tube, so be careful not to point the tube towards anyone while heating.

Materials and Equipment

Equipment: 10 small test tubes, test tube rack, test tube holder, Bunsen burner, 2 medium-sized beakers (for stock solutions), 10-mL graduated cylinder, wash bottle, stirring rod, and scoopula.

Chemicals: solid NH_4Cl (*s*), saturated NaCl (*aq*), concentrated 12M HCl (*aq*), 0.1M FeCl_3 (*aq*), 0.1M KSCN (*aq*), 0.1M AgNO_3 (*aq*), 0.1M CoCl_2 (*aq*), concentrated 15M NH_3 (*aq*), phenolphthalein, 0.1M K_2CrO_4 (*aq*), 6M HNO_3 (*aq*), and 10% NaOH (*aq*).

Experimental Procedure

Record all observations on your report form. These should include, but not be limited to, color changes and precipitates. Note that solution volumes are approximate for all reactions below. Dispose of all chemical waste in the plastic container in the hood.

Part 1: Saturated Sodium Chloride Solution

- Place 3-mL of saturated NaCl (*aq*) into a small test tube.
- Carefully add concentrated 12M HCl (*aq*) drop-wise to the solution in the test tube until a distinct change occurs. Record your observations.

Part 2: Acidified Chromate Solution

- a. Place 3-mL of 0.1M K_2CrO_4 (aq) into a small test tube.
- b. Add an equal amount of 6M HNO_3 (aq) to this solution. Record your observations.
- c. Now add 10% $NaOH$ (aq) drop-wise until the original color is returned. Record your observations. Here the added sodium hydroxide is effectively removing acidic hydrogen ions from the equilibrium system via a neutralization reaction: $H^+ (aq) + OH^- (aq) \rightarrow H_2O (l)$.

Part 3: Aqueous Ammonia Solution

Instructor Prep: At the beginning of lab prepare a stock solution of aqueous ammonia. Add 4 drops of concentrated 15M NH_3 (aq) and 3 drops of phenolphthalein to a 150-mL (medium) beaker, top it up with 100-mL of distilled water, and mix with a stirring rod. Label the beaker and place it on the front desk. The entire class will then use this stock solution in Part 3.

- a. Place 3-mL of the prepared stock solution into a small test tube.
- b. Add a medium scoop of NH_4Cl powder to the solution in this test tube. Record your observations.

Part 4: Cobalt(II) Chloride Solution

- a. Place 3-mL of 0.1M $CoCl_2$ (aq) into 3 small test tubes. Label these test tubes 1-3.
- b. The solution in test tube #1 remains untouched. It is a control for comparison with other tubes.
- c. To the solution in test tube #2, carefully add concentrated 12M HCl (aq) drop-wise until a distinct color change occurs. Record your observations.
- d. To the solution in test tube #3, first add a medium scoop of solid NH_4Cl . Then heat this solution directly in your Bunsen burner flame (moderate temperature). Firmly hold test tube #3 with your test tube holder, and waft it back and forth through the flame (to prevent overheating and “bumping”) for about 30 seconds, or, until a distinct change occurs. Record your observations. Then cool the solution in test tube #3 back to room temperature by holding it under running tap water, and again record your observations.

Part 5: Iron(III) Thiocyanate Solution

Instructor Prep: At the beginning of lab prepare a stock solution of iron(III) thiocyanate. Add 1-mL of 0.1M $FeCl_3$ (aq) and 1-mL of 0.1M $KSCN$ (aq) to a 150-mL (medium) beaker, top it up with 100-mL of distilled water, and mix with a

stirring rod. Label the beaker and place it on the front desk. The entire class will then use this stock solution in Part 5.

a. Place 3-mL of the prepared stock solution into 4 small test tubes. Label these test tubes 1-4.

b. The solution in test tube #1 remains untouched. It is a control for comparison with other tubes.

c. To the solution in test tube #2, add 1-mL of 0.1M FeCl₃ (aq). Record your observations.

d. To the solution in test tube #3, add 1-mL of 0.1M KSCN (aq). Record your observations.

e. To the solution in test tube #4, add 0.1M AgNO₃ (aq) drop-wise until all the color disappears. A light precipitate may also appear. Record your observations. Here the added silver nitrate is effectively removing thiocyanate ions from the equilibrium system via a precipitation reaction: $\text{Ag}^+1 (\text{aq}) + \text{SCN}^-1 (\text{aq}) \rightarrow \text{AgSCN} (\text{s})$.

Experiment № 4. Types of Reactions (4 hrs.)

Objectives

The objectives of this laboratory are as follows:

- To perform and observe the results of a variety of chemical reactions.
- To become familiar with the observable signs of chemical reactions.
- To identify the products formed in chemical reactions and predict when a reaction will occur.
- To write balanced equations for the reactions studied.
- To use the results from the single replacement reactions to devise a partial activity series.

Procedure

Safety

1 Be especially cautious with the 6 M and 3 M acid solutions as they can burn your skin.

2 Do not stare directly at the magnesium when it burns as the light can hurt your eyes.

3 Skin discoloration will result from contact with silver nitrate.

4 Do not touch metals with your hands.

Materials and Equipment

Solids: Mg, CuSO₄•5H₂O, Ca, Cu, Zn, NaHCO₃

Solutions: 6M HCl, 6M NaOH, 6M H₂SO₄, 1M NH₄NO₃, and 0.1M solutions of CuSO₄, ZnSO₄, AgNO₃, NaCl, Ni(NO₃)₂, Pb(NO₃)₂ and K₂CrO₄

Equipment: crucible tongs, one large test tube, two small test tubes, ten small test tubes, test tube holder, test tube rack, 100-mL beaker, red litmus paper, Bunsen Burner

Supplies for Instructor Demonstrations: CaO solution (prepared in advance by stockroom), sucrose, 18M H₂SO₄, distilled water, Na, means of cutting Na and removing from its storage vessel, two 100-mL beakers, straw, red and blue litmus paper, glass stirring rod

Instructions for Performing each Reaction

Perform each of the following reactions except those that are to be demonstrated by your instructor. You will use 1 mL of solution in many of the reactions; estimate this by drops (typically 12-15 drops from the reagent bottle dispenser) or by measuring 1 mL once in your graduated cylinder and then transferring it to a test tube to determine how far it fills the test tube. Record your observations on the data page as you complete each reaction. Make sure that you observe the results of every reaction even if you didn't actually mix the chemicals yourself. Then write a balanced equation for each reaction. Be sure to include the states of all compounds in your equations (solid, liquid, aqueous, or gas). If no reaction occurs write the words "no reaction" (or NR) instead of the products in your balanced equation and indicate why you think there was no reaction. Unless otherwise indicated dispose of all waste in the waste container provided. *Do not put metal strips in the sink.*

A. Combination Reactions

1. Instructor Demonstration. Pour about 35 mL of a clear saturated solution containing calcium oxide into a 100 mL beaker. Allow the solution to stand for about 15 minutes. Observe. Use a straw to blow bubbles into the solution for a few seconds. Observe the solution again.

2. Hold a small strip of magnesium metal (used in flashbulbs and fireworks) in your crucible tongs and ignite the metal in the hot portion of a burner flame. Don't forget to note the color and composition of the residue left on the tongs.

B. Decomposition Reactions

1. Instructor Demonstration. Perform this reaction in the fume hood. Fill a 100 mL beaker about one-third full of granulated sugar (sucrose, C₁₂H₂₂O₁₁). Add about 20 mL of concentrated (18 M) sulfuric acid and stir until mixed well. Continue stirring until the mixture darkens. Observe. Do not touch the reaction

products or the beaker with your hands; use a stirring rod to guide the solid product that forms.

2. Place a small amount (an amount that will fit on the end of a spatula) of solid copper(II) sulfate pentahydrate in a medium test tube. Use a test tube holder to hold the tube at about a 45° angle and heat in a burner flame for a few minutes, remembering not to point the tube at anyone in the room. Note any changes in the appearance of the solid and anything else that appears in the test tube. Allow the solid to cool and add a few drops of water. Observe. Dispose of the copper compound in the waste container.

C. Displacement Reactions

Use 1 mL of each solution unless otherwise specified. For reactions involving metals, use just one piece of metal. Do not put the metal pieces in the sink. If no discernable initial change is noted, let the reaction mixture stand for at least five to ten minutes before observing again. Not all of the combinations will yield observable reactions. Repeat the reaction if there is any doubt about whether a reaction occurred or not.

1. Instructor Demonstration. Cautiously add a small piece of sodium metal to water. Test the resulting solution with red and blue litmus papers (red litmus paper will turn blue in the presence of a base; blue litmus will turn red in the presence of an acid).

2. Calcium metal and water (15 mL) in a large test tube
3. Zinc metal and water
4. Copper metal and 6 M hydrochloric acid
5. Zinc metal and 6 M hydrochloric acid
6. Zinc metal and 0.1 M copper(II) sulfate
7. Copper metal and 0.1 M zinc sulfate
8. Copper metal and 0.1 M silver nitrate

D. Exchange Reactions

Use 1 mL of each solution unless otherwise specified. Be sure to mix the solutions well.

1. 0.1 M silver nitrate and 0.1 M sodium chloride
2. 0.1 M nickel(II) nitrate and three drops of 6 M sodium hydroxide
3. 0.1 M lead(II) nitrate and 0.1 M potassium chromate
4. 1 M ammonium nitrate and 6 M sodium hydroxide. Warm the test tube gently by passing it back and forth through a burner flame. Hold a strip of moistened red litmus paper in the tube without letting it come in contact with the

sides of the tube and note any color changes to the paper. Remove the tube from the flame and quickly and cautiously note the smell.

5. Place 5 mL of 6 M hydrochloric acid in a 100 mL beaker. Carefully add several spatulas of solid sodium bicarbonate. Observe.

6. Combine about 5 mL each of 6 M sodium hydroxide and 6 M sulfuric acid in a large test tube. Mix with a stirring rod. Cautiously feel the outside of the test tube. If you cannot detect anything, make sure that you used the correct concentrations of acid and base. *Be very careful with the concentrated sulfuric acid; it is very caustic and can dissolve skin and clothing.*

Experiment № 5. Calculation and preparation of accurate concentrations solutions (4 hrs.)

Part 1. Preparation of standard $N_{A_2}B_4O_7$ solution

Objectives

To prepare 250 ml 0.1 n (0.1 normality) standard sodium tetraborate ($N_{A_2}B_4O_7$) solution.

Procedure:

The quantity of sodium tetraborate $N_{A_2}B_4O_7 \cdot 10H_2O$ (borax) m (g), which is required to prepare 250 ml 0.1 n solution,

You need to take clean, dry 100 ml capacity glass and, using the analytical balance, to scale close to calculated (4,7671 g) quantity of borax a (g). It is sluiced down with 40 - 70 ml of deionized water and is dissolved, while warming until 50 °C as well mixing with a glass stick. The solution is cooled (glass stick shouldn't be withdrawn), quantifiably transfused into 250 ml measurement flask, then attenuated with deionized water until the mark and well mixed. Prepared solution of borax is poured into the new cleaned glass vessel, which is theretofore rinsed with little volume of prepared borax solution. Sodium tetraborate solution's molar equivalent concentration is calculated.

You need to stick etiquette on the vessel with prepared standard solution of borax

Part 2. Preparation of titrated NaOH solution

Objectives

To make 500 ml titrated 0.1 n NaOH solution.

Procedure:

NaOH quantity m (g), which is required to prepare 500 ml 0,1 n solution.

Calculated NaOH quantity with an excess of 0.2 g is scaled on the technical balance in 100 ml capacity glass. Then substance is dissolved in 50-60 ml of deionized water, then poured into 500 ml capacity clean glass vessel (it is forbidden to use vessels with cut glass stoppers), also diluted until 500 ml and well mixed. An etiquette, where solution's concentration is not indicated, needs to be stuck on the vessel. Prepared titrated NaOH solution molar equivalent concentrations (normality) will be estimated later.

Part 3. Estimation of working HCl solution concentration

Objectives

Using titration with etalon borax solution estimate working HCl solution concentration. Write a report of analysis.

Procedure

Prepare titration dishes: burette, pipette, 3 Erlenmeyer flasks (250ml). Wash clean burette with small volume of borax solution. Fill it with etalon borax solution of known concentration. Make sure that air is absent in a burette. Fix filled burette in vertical position. Wash pipette with 10 ml of investigative solution (in this case HCl). Use pipette to add 10 ml of investigative HCl solution in to three Erlenmeyer flasks. Add 2-3 drops of methyl orange to each flask. Solution becomes red. Note burette data - starting volume for act of analysis (started from, ml). Put Erlenmeyer flask with investigative solution on ceramic tile or white paper to observe changes in color. Titrate with borax solution until yellow color appears. Note burette data for act of analysis (drop to, ml).

Park 4. Estimation of working NaOH solution concentration

Objectives

Estimate molar concentration of prepared NaOH solution by titration with HCl of known concentration. Write a report of analysis.

Procedure:

Fill burette with HCl solution of estimated normality. Add 10 ml of NaOH solution to three Erlenmeyer flasks using pipette, add 2-3 drops of indicator - methyl orange. Solution becomes yellow. Titrate with HCl until orange color. Note data of titration in table of act of analysis. Calculate normality of NaOH solution accordingly to titration data, write the report No 2. Put estimated molar concentration of solution on the label.

Experiment № 6. Preparing of Buffers and Buffer Capacity (4 hrs.)

Objectives

The purpose of this experiment is to prepare buffer solutions and to determine their buffer capacity.

Procedure

Safety:

- Always wear an apron and goggles in the lab.
- Report any spills so they may be cleaned up.

Equipment/Materials:

acetic acid (0.10, 0.30, 0.50M)	0.100 M NaOH
sodium acetate	burets
buret clamp	pH meter
standard buffer solution (pH 4 & 7)	magnetic stirrer (if available)
100 mL volumetric flasks	250 mL beaker
stir bar	

Experimental Procedure

1. Before preparing the buffer solutions, you must determine the amount of acetic acid and sodium acetate required. The K_a of acetic acid is 1.8×10^{-5} . In the space below, show the calculations for the preparation of 100 mL of an acetic acid – sodium acetate pH 5.0 buffer using:
 - a. 0.10 M acetic acid
 - b. 0.30 M acetic acid
 - c. 0.50 M acetic acid
2. Your lab group will be assigned to prepare one of the buffer solutions. Check your calculations with your instructor before proceeding.
3. The procedure for calibrating pH meters varies from instrument to instrument. Follow your instructor's directions for this step.
4. Once the pH has been calibrated, measure the pH of the buffer solution that your group prepared. Record the value in your data table.
5. Clean two burets. Fill one with the buffer solution that you prepared and fill the other with the 0.100 M NaOH. Make sure that the tips of the burets are filled and that the level of the liquids is at or below the 0.00 mL line.
6. Transfer 10.00 mL of the buffer solution to a 250 mL beaker and add some distilled water. If a magnetic stirrer is to be used, keep the tip of the electrode above the stir bar.
7. Begin adding the NaOH to the buffer solution in small increments. After each addition, record the total volume of NaOH added and the pH of the solution in the data table.
8. Continue adding the NaOH solution until the pH has risen at least one pH unit.
9. Repeat as time allows.

Experiment № 7. Inorganic qualitative reaction (4 hrs.)

Objectives

1. To carry out analytic reactions of cations and anions presented in the table; to put down and memorize the effects of reactions (analytic signals). To write down equations of analytic reactions in molecular, ionic and short ionic form.
2. perform reactions of flame coloration by Ca^{2+} , Ba^{2+} , K^+ , Na^+ cations.

Procedure

Carry out the reactions 1 to 15 (Table No. 4) transferring 1 ml of solutions 1 and 2 into clean test tubes (measuring not required). Put the observed effect of the reactions to Table No 4. For example, the first reaction for detection of Ag⁺ ion is to be carried out in the following way: Transfer 1 ml of AgNO₃ (solution No. 1) without measuring to a clean test tube (the height of the solution in test tube to be about 1 cm). Transfer about 1 ml 6N HCl solution (solution No. 2) to the same test tube. Observe the effect of the reaction, when white precipitate is formed; write it down to the Table.

At the end of the report write the equation of the reaction in molecular, ionic and short ionic form:

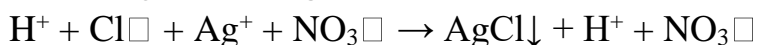
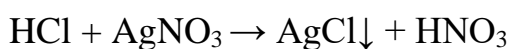
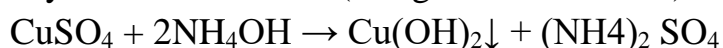
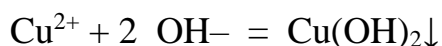


Table. Analytic ion reactions and their effects

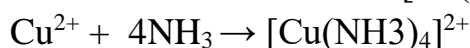
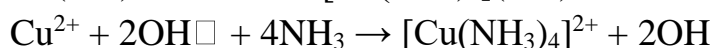
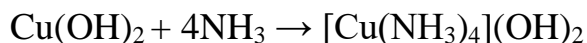
Serial No	Ion to detect	Solution No. 1 (with the ion to detect)	Solution No. 2 (the reagent)	Reaction effect (analytic signal)
1.	Ag ⁺	AgNO ₃	HCl	white precipitate
2.	Hg ₂ ²⁺	Hg ₂ (NO ₃) ₂	HCl	
3.	Hg ₂ ²⁺	Hg ₂ (NO ₃) ₂	KJ	
4.	Pb ²⁺	Pb(NO ₃) ₂	HCl	
5.	Pb ²⁺	Pb(NO ₃) ₂	KJ	
6.	Hg ²⁺	Hg(NO ₃) ₂	KJ (drops)	
7.	Cu ²⁺	CuSO ₄	NH ₃	
8.	Fe ²⁺	FeSO ₄	K ₃ [Fe(CN) ₆]	
9.	Fe ³⁺	FeCl ₃	NH ₄ CNS	
10.	NH ₄ ⁺	NH ₄ Cl	NaOH	
11.	CO ₃ ²⁻	K ₂ CO ₃	HCl	
12.	CO ₃ ²⁻	K ₂ CO ₃	BaCl ₂	
13.	SO ₄ ²⁻	Na ₂ SO ₄	BaCl ₂	
14.	Cl ⁻	NaCl	AgNO ₃ (drops)	
15.	NO ₃ ⁻	NaNO ₃	FeSO ₄ , H ₂ SO ₄	

Then carry out the analytic reaction of the next ion, observe, memorize, put down to the table the effect of the reaction (analytic signal), write down the equations of the reaction and so on. Analytic reaction for detection of Cu²⁺ ion (Table No. 4, reaction 7) is to be carried out in the following way: mix 1 ml of CuSO₄ solution with 1 ml of NH₄OH solution. At first, bluish white precipitate of copper hydroxide is formed (1 stage of the reaction):

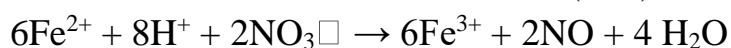
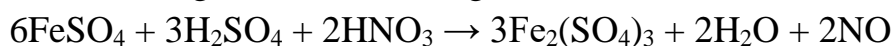




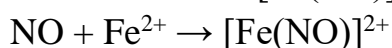
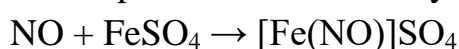
This precipitate of $\text{Cu}(\text{OH})_2$ is dissolved in excess ammonia, and tetraammonia ion is formed (2 stage of the reaction). The solution becomes blue; the color intensity depends on the concentration of Cu^{2+} :



Analytic reaction for detection of NO_3^- ion should be carried out in the following way: transfer 1 ml of NaNO_3 solution to a clean test tube, add a little amount (with a head of safety-matches) of FeSO_4 crystals, which are solved in the solution on the tube shaking. Carefully pour by a wall of the inclined tube in slight flow about 1 ml of concentrated H_2SO_4 . Between unmixed layers of the solutions of different density, the greenish brown „ring of nitrate” is formed. The reaction proceeds in two stages. In the first stage, ferrous sulfate reduces NO_3^- to NO :



In the second stage, NO and ferrous sulfate form greenish brown, unstable complex compound, ferrous nitrosyl sulfate:



I- and Br- ions interfere with this reaction, forming „rings” of another color.

For flame coloration reactions, use platinum wire after washing it by concentrated hydrochloric acid and heating on gas burner flame. Dip the heated platinum wire into analysis solution and then into the colorless flame of gas burner. After examination of flame color of ions, put down the data obtained to the table.

Experiment № 8. Titration of Vinegar (4 hrs.)

Objectives

The objectives of this laboratory are to determine the molarity and percent by mass of acetic acid in vinegar.

Procedure

Be especially careful when handling the sodium hydroxide base (NaOH), as it is corrosive and can cause chemical burns to the skin. If any NaOH spills on you,

rinse immediately under running water for up to 15 minutes and report the accident to your instructor.

Materials and Equipment

50-mL burette, 5-mL volumetric pipette*, pipette bulb*, ~ 0.1 M NaOH (aq), vinegar, phenolphthalein, burette stand, two 250-mL (or 125 mL) Erlenmeyer flasks, wash bottle with distilled water, funnel

Titration Procedure

Your instructor will demonstrate the correct use of the volumetric pipette and burette at the beginning of the lab session. Detailed instructions on how to use a pipette are also found on the last page of this handout. Note that three titrations must be performed.

1. Obtain a 50-mL burette, 5-mL volumetric pipette and a pipette bulb from the stockroom.

Setting up the burette and preparing the NaOH

2. Rinse the inside of the burette with distilled water. Allow the distilled water to drain out through the tip in order to ensure that the tip is also rinsed.

3. Now rinse the burette with a small amount of NaOH (aq). To do this, add about 5-mL of NaOH (aq) to the burette, then twirl the burette on its side (over the sink) to rinse its entire inner surface. Then allow the NaOH (aq) to drain out through the tip.

4. Fill the burette with NaOH (aq) up to the top, between 0-mL and 5-mL. Use a funnel to do this carefully, below eye-level, and preferably over the sink. After this you will need to flush the tip of the burette – your instructor will show you how to do this. Now measure the volume at the level of the NaOH precisely, and record it as the “Initial Burette Reading” on your report. Also record the exact molarity of the NaOH (aq), which is labeled on the stock bottle.

Preparing the vinegar sample

5. The volumetric pipette used in this lab is designed to measure and transfer exactly 5.00 mL of solution. First, rinse the inside of the volumetric pipette with distilled water. Using the pipette bulb, draw the water into the pipette up above the 5-mL mark, then allow it to drain out through the tip. You may want to do this several times for practice. Then perform a final rinse, but this time use vinegar.

6. Now use the volumetric pipette to transfer 5.00-mL of vinegar into a clean 250-mL Erlenmeyer flask (see instructions on page 4). Record this volume of vinegar (precise to two decimal places) on your report. Then add about 20-mL of distilled water and 5 drops of phenolphthalein to this Erlenmeyer flask.

Performing the titration

7. Begin the titration by slowly adding NaOH (aq) from the burette to the vinegar in the Erlenmeyer flask. Swirl Erlenmeyer flask as you add the base in order to efficiently mix the chemicals. Some pinkness may appear briefly in the flask as the base is added, but it will quickly disappear as the flask is swirled.

8. As the equivalence point is approached, the pink color will become more pervasive and will take longer to disappear. When this occurs, start to add the NaOH (aq) drop by drop. Eventually the addition of just one drop of NaOH (aq) will turn the solution in the Erlenmeyer flask a pale pink color that does not disappear when swirled. This indicates that the equivalence point has been reached. Do not add any more NaOH (aq) at this point. Measure this volume of NaOH (aq) precisely, and record it as the “Final Burette Reading” on your report. Then show the resulting solution in the flask to your instructor so s/he can record the final color on your report form.

9. Refill your burette with NaOH (aq), and then repeat this procedure for a second sample of vinegar, and then a third sample of vinegar. You do not need to flush the tip of the burette again. Note that if you use less than 25-mL of NaOH (aq) for the second titration, you do not need to refill the burette for the third titration; also that you will need to clean out and re-use one of your Erlenmeyer flasks for the third titration. You and your partner should take turns performing these titrations.

10. When finished, dispose of your chemical waste as instructed.

Pipetting Instructions

a. Get the appropriate amount of the solution you wish to pipette in a clean, dry beaker. Never pipette directly out of the stock bottles of solution. This creates a contamination risk.

b. Insert the tip of the pipette into the beaker of solution so that it is about a quarter inch from the bottom. Be sure not to press the tip against the bottom of the container.

c. If you are right handed, hold the pipette in your right hand, leaving your index finger free to place over the top of the pipette. With your left hand, squeeze the pipette bulb. Press it firmly over the top of the pipette, but **DO NOT INSERT THE PIPET DEEP INTO THE BULB!**

d. Release the pressure on the bulb and allow the solution to be drawn up into the pipette until it is above the volume mark. Do not allow the solution to be sucked into the bulb itself.

e. Quickly remove the bulb and place your index finger firmly over the top of the pipette. Then remove the pipette tip from the beaker of solution.

f. Slowly roll your finger to one side and allow the liquid to drain until the bottom of the meniscus is aligned with the volume mark. With practice you will be able to lower the liquid very, very slowly.

g. When the bottom of the meniscus is even with the volume mark, press your index finger firmly on the top of the pipette so no liquid leaks out. Touch the tip once to the side of the beaker to remove any hanging drops.

h. To transfer the solution, place the tip of the pipette against the wall of the receiving container at a slight angle. Then allow the liquid to drain from the pipette.

i. When the solution stops flowing, touch the pipette once to the side of the receiving container to remove any hanging drops. DO NOT blow out the remaining solution. The pipette has been calibrated to deliver the appropriate amount of solution with some remaining in the tip.

Calculations

Molarity of Acetic Acid in Vinegar

- First, using the known molarity of the NaOH (aq) and the volume of NaOH (aq) required to reach the equivalence point, calculate the moles of NaOH used in the titration.

- From this mole value (of NaOH), obtain the moles of HC₂H₃O₂ in the vinegar sample, using the mole-to-mole ratio in the balanced equation.

- Finally, calculate the molarity of acetic acid in vinegar from the moles of HC₂H₃O₂ and the volume of the vinegar sample used.

Mass Percent of Acetic Acid in Vinegar

- First, convert the moles of HC₂H₃O₂ in the vinegar sample (previously calculated) to a mass of HC₂H₃O₂, via its molar mass.

- Then determine the total mass of the vinegar sample from the vinegar volume and the vinegar density. Assume that the vinegar density is 1.000 g/mL (= to the density of water).

- Finally, calculate the mass percent of acetic acid in vinegar from the mass of HC₂H₃O₂ and the mass of vinegar.