GDAŃSK UNIVERSITY
OF TECHNOLOGY

Subject card

Subject name and code	Basic Chemistry, PG_00037478		
Field of study	Biotechnology		
Date of commencement of studies	October 2020	Academic year of realisation of subject	2020/2021
Education level	first-cycle studies	Subject group	Obligatory subject group in the field of study
Mode of study	Full-time studies	Mode of delivery	at the university
Year of study	1	Language of instruction	Polish
Semester of study	1	ECTS credits	6.0
Learning profile	general academic profile	Assessment form	exam
Conducting unit	Department of Inorganic Chemistry -> Faculty of Chemistry		
Name and surname of lecturer (lecturers)	Subject supervisor	prof. dr hab. inż. Anna Dołęga	
	Teachers	dr inż. Aleksandra Wiśniewska prof. dr hab. inż. Anna Dołęga dr inż. Mateusz Daśko	

Lesson types and methods of instruction	Lesson type	Lecture	Tutorial	Laboratory	Project	Seminar	SUM
	Number of study hours	30.0	30.0	0.0	0.0	0.0	60
	E-learning hours included: 0.0						
	Podstawy Chemii dla kierunku Biotechnologia - Moodle ID: 6712 https://enauczanie.pg.edu.pl/moodle/course/view.php?id=6712						

Learning activity and number of study hours	Learning activity	Participation in didactic classes included in study plan	Participation in consultation hours	Self-study	SUM
	Number of study hours	60	10.0	80.0	150

Subject objectives	A knowledge of principal concepts in general chemistry.

Learning outcomes

Course outcome	Subject outcome	Method of verification
K6_U02	Student can apply the basic chemical principles in problem solving.	[SU1] Assessment of task fulfilment
K6_W02	Student knows basic chemical nomenclature and principles.	[SW1] Assessment of factual knowledge

Subject contents	Lecture: Basic concepts and definitions: basic chemical laws, balanced chemical equations, ionic equations, nomenclature of chemical compounds. The three states of matter: gases, liquids and solids. Crystalline and non-crystalline (amorphous) solids. Equations of state: ideal gas law, cubic and virial equations of state, Dalton"s law of partial pressures, the kinetic theory of gases. Characteristics and structure of liquids, solutions. Anisotropy and isotropy in solids, crystal lattice, polymorphism and isomorphism, crystal defects, non-stoichiometric compounds, reactions in solid state. Atomic structure: atomic nucleus, atomic and mass numbers, mass deficiency and nuclear energy, isotopes, nucleus stability, spontaneous disintegration of nuclei, radio decay rate, half-life period, thermonuclear reactions. Atomic structure: electrons in atoms, Bohr model, Heisenberg uncertainty principle, electron density, quantum numbers, atomic orbitals, Pauli exclusion principle, Hund’s rule. Periodic table of elements: periodicity of chemical and physical properties of atoms, periods, groups and blocks of elements, atomic, ionic and van der Waals radii. Chemical bonds: valence electrons, octet rule, electronegativity, electron affinity, energies of chemical bonds, Molecular orbitals: LCAO (MO) method, sigma and pi orbitals, hybridization of atomic orbitals, hybridizations type and their geometric consequences. Lewis structures (diagrams), VSEPR Strong chemical bonds and their types, ionic, metallic and covalent bonds, physiochemical properties of molecular and ionic compounds, metals, alloys. Weak interactions: hydrogen bonds, van der Waals forces. Solutions. Properties and functions of solvent, water as a solvent, solvation, autodissotiation of water, donor and acceptor solvents, melted salts. Electrolytes: weak and strong electrolytes, a the dissociation constant, the degree of ionization. Basic thermochemistry, the enthalpy of a reaction (the heat of a reaction). Hess’ law. Chemical equilibrium: the law of mass action. Basic kinetics, reaction rates, rows and mechanisms. Redox reactions, oxidation number, reducing and oxidizing agents. Galvanic cells. Standard electrode potentials. The galvanic series. Descriptive chemistry: hydrogen, oxygen and water, chemistry of carbon compounds. Classes: Basic concepts and chemical laws (atomic and molecular masses, mole Avogadro number, molar mass, isotopes). Ideal gas law. Electrons configurations. Composition stoichiometry. Formulas. Composition from formulas. Determination of a chemical formula, empirical (simplest) and molecular formulas. Composition of mixtures. Lewis diagrams. Solutions – expressing the concentration – mass concentration, molar concentration, number concentration, volume concentration. Concentration conversion. Dilution and mixing of solutions Balacing equations (including redox equations). Reaction stoichiometry, excess and limiting reagent, parallel reactions, reaction yield. Reactions in solutions.
Prerequisites and co-requisites	No requirements

Assessment methods and criteria

Subject passing criteria	Passing threshold	Percentage of the final grade
Written tests -three times during semester	60.0%	33.0%
Written exam	60.0%	67.0%

Recommended reading

Basic literature	1. L. Jones, P. Atkins "Chemia ogólna"; PWN, 2004, or more recent issues (Polish translation from English "General Chemistry" original) 2. A. Bielański „Podstawy chemii nieorganicznej” (PWN) – recent issues; 3. P.A. Cox „Krótkie wykłady, chemia nieorganiczna”, PWN, 2003; (Polish translation from English "Instant Notes in Inorganic Chemistry" original)
Supplementary literature	
eResources addresses	

Example issues/ example questions/ tasks being completed	1. Write the formulas of salts with the following systematic names: copper (II) sulfate (IV), ammonium nitrate (III), potassium chlorate (V), magnesium bromide. 2. Give the systematic names of the following compounds: $H2SO_4$, H_2SO_3, CuCl, $FePO_4$. 3. How many moles of calcium oxide can be obtained from 500 grams of calcium carbonate? 4. How many molecules of nitrogen is found in 7 g of N_2? 5. Define the concept of mole. How many moles of oxygen atoms are in 16 grams of oxygen. 6. What is the law of conservation of mass? Based on this right, solve the following task. By heating 10 grams of silver nitrate, 3.67 grams of gaseous products and metallic silver were obtained. How many grams of silver were obtained? 7. What is the Avogadro number? Calculate the absolute mass of the gold atom (in grams) knowing that the relative atomic mass of gold is 196.97 u and the Avogadro number is 6,023´1023 mol-1. 8. Write in the molecular and ionic form any equation for the double exchange reaction. 9. Write down the actual and empirical formula of the dinitrogen tetroxide. Calculate the molecular weight and weight percentage of this compound. 10. What is termite and termite reaction? 11. Describe briefly the three states of matter. 12. In what state it occurs in nature: a) nitrogen; b) copper; c) sugar? Can these substances occur in other states? What conditions are necessary for this? 13. What is: a) isochoric gas conversion; b) gas compressibility factor; c) absolute zero temperature? 14. What is the compressibility factor of gases - give the value of the compressibility factor for perfect gases. 15. What are the assumptions of the kinetic theory of gases? 16. Express: a) 100,000 Pa in tracks; b) 0.75 bar in pascals; c) 2 at in mmHg. 17. Express the mass of the atom of lead in grams. 18. Calculate the weighted average atomic mass of carbon if it is known that in nature carbon occurs in the form of three isotopes of the given masses (in brackets the distribution in% of the number of atoms is given): 12C - 12.000000 u. (98.93%), 13C - 13, 003355 u. (1.07%), 14C - 14.003241989 (0.00%). 19. Explain the phenomenon of natural radioactivity? Give two examples of radionuclides. 20. Characterize three main types of radiation emitted by radioactive elements. 21. Using the VSEPR model, describe the geometry of the molecules: a) CCl_4; b) SF_4 22. Define the dipole moment. 23. Do the given molecules show a dipole moment: a) CO_2; SO_2? Justify the answer by specifying the structure of these molecules and the component dipole moments of bonds. 24. Which of the following molecules shows a higher dipole moment: NH_3 or NF_3? How is that explained? 25. Draw a diagram of the molecular orbitals of the F_2 molecule and calculate the bond order. 26. Use the molecular orbital diagram to explain why B_2 and O_2 are paramagnetic. 27. What is the heat capacity of a substance? 28. What is a vapor? 29. What is the vapor pressure of a substance? 30. How do the vapor pressures of more volatile and less volatile substances differ (measured at the same temperature)? 31. Give an example of a substance with an ionic and molecular structure and compare the crystal properties of these substances. Write briefly what they result from. 32. Compare the structure of two allotropic forms of carbon. 33. How are covalent crystals built? Give an example of a substance that forms covalent crystals. 34. Why do metals conduct electricity and heat well? 35. Write down two equations of reactions used for obtaining oxygen in the laboratory 36. Draw the Lewis formula of ozone. Does the ozone molecule have a dipole moment. 37. Write the equations for the formation of oxide, peroxide, superoxide and ozonide. 38. Provide equations of chemical reactions illustrating ozone's high chemical reactivity. 39. Give an example of ionic and covalent oxide. What is the physical state of these oxides at room temperature? 40. What is smog? 41. Draw Lewis formulas of nitric oxide and nitrogen dioxide . Why are these particles very reactive? 42. Give the equations of two reactions that can be used to obtain hydrogen under laboratory conditions. 43. Describe two industrial methods for obtaining hydrogen. 44. Which element is most common in the Earth's crust and which in the universe. 45. Write the half and total equation of the water electrolysis reaction. 46. Give two equations of chemical reactions for obtaining: a) covalent hydrides; b) salt type hydrides. 47. Give an example of one ionic hydride and one covalent. Write the equations of these hydrides with water. 48. Why does water reach a maximum density at 4 ° C? 49. Write two equations of reactions of water: a) with non-metal oxides; b) with metal oxides 50. What is: a) water of crystallization; b) heavy water; c) hydrogen peroxide; d) aqua regia 51. What is a saturated solution? 52. What is the effect of temperature on the solubility of: a) solids; b) gases. 53. What are electrolytes? What is the difference between weak and strong electrolytes? 54. Give the relationship between the constant and the degree of dissociation of the weak electrolyte. 55. What are colloids? 56. What is the phenomenon of osmosis? 57. What is the difference between transient and permanent hardness of water? 58. Suggest elemental reactions for the HCl + Br_2 = HBr + ClBr reaction 59. What is activation energy? 60. Define the speed of the chemical reaction 61. List the factors affecting the reaction rate. 62. What are catalysts? Explain how the catalyst works. 63. What is the reaction order? 64. Knowing that the reaction NO + O_3 = NO_2 + O_2 is an elementary reaction (one-step) write down its kinetic equation. 65. What effect does temperature have on the reaction rate?
Work placement	Not applicable