MINISTRY OF HEALTH OF UKRAINE
Zaporizhzhya State Medical University
Analytical Chemistry Department

MEDICAL CHEMISTRY

Module 1

ACID-BASE EQUILIBRIA IN BIOLOGICAL FLUIDS

Topic module 2

Manual
for teachers and students
of the 1-st course of the Medical Faculty
speciality "General Medicine"

Zaporizhzhya 2020

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The Manual is considered and approved at the meeting of The Committee of Physical and Chemical Subjects of Zaporizhzhya State Medical University (record № ______ of _______), at the meeting of The Central Methodical Council of Zaporizhzhya State Medical University (record № ______of ________)

PREFACE

Medical chemistry is studied according to the approved standard of academic curriculum of 2005 year for students of the Universities of the III-IV levels of accreditation of Ukraine for specialty 7.110101 "General medicine" field of study 1101 "Medicine" is established in accordance with the educational and skill characteristics (ESC) and education and professional programs (EPP) training approved by order of the Ministry of Education of Ukraine from 16.04.03 № 239.

Organization of studying process is proceed according to the requirements of European Credit Transfer System (ECTS).

Study of Medical Chemistry is carried out during 1 semester of 1st year of training according to the academic curriculum.

The curriculum consists of a discipline module, including 4 blocks of topic modules:

1. Chemistry of biogenic elements. Complex formation in biological fluids.
2. Acid – base equilibriums in biological fluids.
3. Thermodynamic and kinetic processes regularities and electrokinetic phenomena in biological systems.
4. Physics and chemistry of surface effects. Lyophobic and lyophilic disperse systems.

TOPIC MODULE 2

ACID-BASE EQUILIBRIA IN BIOLOGICAL FLUIDS

Tutorial # 4

1. THEME: Characteristics of the quantitative composition of solutions. Solution preparation.

2. PURPOSE: To consolidate knowledge on solution concentration expressing.

3. OBJECTIVES:

3.1. To get skills of calculation of substance sample weight, solutions concentration, conversion of solutions concentrations from one to another.

3.2. Learning to use glassware for solutions preparation.

4. PLAN AND ORGANIZATIONAL STRUCTURE OF THE TUTORIAL:

4.1. Organizational part …………………………,,,,,,,,…....................5 min

4.2. Goal-setting and motivation of the theme studying (opening address of the teacher)………..........……………...5 min

4.3. Determination of initial knowledge level ……….….......................15 min

4.4. Correction of the initial knowledge level.……..............................25 min

4.5. Organization of independent work of students (target teacher guidance, rules of accident prevention) ……………………………………………….5 min

4.6. Laboratory work ………………………………………………..45 min

4.7. Final comtrol: cheking of the laboratory work results and protocols ………………………....…..10 min

4.8. Concluding remarks of the teacher, guidance to the next lesson ………………...……………………………...………….…........................3 min

5. REFERENCE FOR SELF- STUDY:

5.1. Questions for self-study

1. Solutions concentration and ways of its expression: the mass fraction, molar concentration, molar concentration of equivalent, mole fraction, molal concentration (molality), the titre of the solute.
2. The concept of equivalent, equivalence factor, molar mass of equivalent.
3. Chemical glassware for solution preparation: volumetric flasks, pipettes, burettes, flasks, cylinders, beakers, etc.
4. The role of fluids in vital activity.

5.2. Solve the tasks:

Task 1. Calculate the mass fraction of sodium chloride, if 20 g of sodium chloride is dissolved in 180 g of water.

Answer: 10%.

Task 2. 500 ml of the solution contains 10,6 g of sodium carbonate. Calculate the molar concentration of the prepared solution.

Answer: 0,2 mol/l.

Task 3. Sodium hydroxybutyrate is used for the introduction of anesthesia,. It is available as 20% solution in 10 ml ampoules. The dose of the drug is 70 mg per 1 kg of body weight. How many ml of sodium hydroxybutyrate must be injected to the patient of 60 kg weighs?

Answer: 21 ml

5.3. Examples of the solved tasks:

Task 1. Calculate the mass fraction of sodium chloride in the solution if in 200 g of water was dissolved 50 g of sodium chloride..

Solution. We determine the mass of the solution:

m = a(NaCl) + a(H_2O) = 200 + 50 = 250 g ;

and find the mass fraction of sodium chloride in the solution:

$$\omega\% = \frac{a_x}{m} \cdot 100 = \frac{50}{250} \cdot 100 = 20\%.$$

Answer: 7,88.

Task 2. Calculate the molar concentration of 20% solution of $FeSO_4$ if its density is 1,21 g / ml.

Solution. First we calculate how many grams of $FeSO_4$ are contained in 1 liter of the solution.

$$\omega\%_{(FeSO_4)} = \frac{a_{(FeSO_4)}}{m} \cdot 100 = \frac{a_{(FeSO_4)}}{V \cdot \rho} \cdot 100, \text{ so}$$

$$a_{(FeSO_4)} = \frac{\omega_{(FeSO_4)} \cdot V \cdot \rho}{100} = \frac{20 \cdot 1000 \cdot 1{,}21}{100} = 242\,g$$

Recalculate the mass found in moles of $FeSO_4$

$$n_{(FeSO_4)} = \frac{a_{(FeSO_4)}}{M_{(FeSO_4)}} = \frac{242}{151{,}9} = 1{,}59\ \text{mol.}$$

As this number of moles is contained in 1 liter of the solution, the molar concentration of the solution is 1,59 mol/l.

Answer: 1,59 mol/l.

Task 3. 200 ml of solution contains 4,9 g of sulfuric acid. Calculate the molar concentration of sulfuric acid equivalent.

Solution. We calculate the equivalent molar mass of sulfuric acid:

$M(H_2SO_4) = 98$ g/mol;

$f_e(H_2SO_4) = 1/2$;

$M_e(H_2SO_4) = M(H_2SO_4) \cdot f_e(H_2SO_4) = 98 \cdot 1/2 = 48$ g/mol.

Then we can find the equivalent molar concentration of sulfuric acid:

$$C_{Me(H_2SO_4)} = \frac{a_{(H_2SO_4)}}{M_{e(H_2SO_4)} \cdot V} = \frac{4{,}9}{49 \cdot 0{,}2} = 0{,}5\ \text{mol/l.}$$

Answer: 0.5 mol/l.

5.4. Study tests:

1. The mass fraction of a substance in solution is:

A. The ratio of the mass of the substance to the mass of the solution.*

B. The ratio of the solution mass to the mass of the substance.

C. The difference between the mass of the solution and the mass of the substance.

2. What mass of NaCl is required for preparing of 2000 g of 5% solution?

A. 10g.

B. 100 g.*

C. 150 g.

D. 5 g.

E. 200 g.

3. NaCl (sodium chloride) as isotonic (0.9%) solution is used intravenously, subcutaneously, rectally at poisoning, toxic dyspepsia, acute blood loss, shock. How much water and sodium chloride (in grams) should be taken to prepare 200 g of isotonic solution?

A. 18 g of water and 192g of sodium chloride.

B. 198,2 g of water and 1,8 g of sodium chloride.*

C. 153g of water и 57 g of sodium chloride.

D. 192,8 g of water и 7,2 g of sodium chloride.

E. 99,1 g of water и 0,9 g of sodium chloride.

4. Molar mass of equivalent of acid M is equal to M/Z, where Z is:

A. The acidity of the base.

B. Basicity of the acid.*

C. The number of metal cations.

D. Valence of the metal

E. The number of acid residues.

5. Molar mass of equivalent of base M is equal to M/Z, where Z is:

A. The acidity of the base.*

B. Basicity of the acid.

C. The number of metal cations.

D. Valence of the metal.

E. The number of acid residues.

6. The molar mass of equivalent of salt M is M/Z, where Z is:

A. The number of metal cations.

B. Product of the number of metal cations and the valence of the metal.*

C. Valence of the acid residue.

D. Valence of the metal.

E. The number of acid residues.

7. Equivalent molar mass of chromium sulfate (III) (M = 392 g/mol) is:

A. 392 • 6 = 2352.

B. 392 : 6 = 65.*

C. 392 • 2= 784.

D. 392 : 3 = 130.

E. 392.

8. Select the formula to calculate the mass fraction:

A. $\frac{a_x}{M_x \cdot V}$.

B. $\frac{a_x}{m} \cdot 100$.*

C. $\frac{a_x}{M_{ekv.x} \cdot V}$.

D. $\frac{a_x}{V}$.

9. Select the formula to calculate the molar concentration:

A. $\frac{a_x}{M_x \cdot V}$.*

B. $\frac{a_x}{m} \cdot 100$.

C. $\frac{a_x}{M_{ekv.x} \cdot V}$.

D. $\frac{a_x}{V}$.

10. Select the formula to calculate the molar concentration of equivalent:

A. $\frac{a_x}{M_x \cdot V}$.

B. $\frac{a_x}{m} \cdot 100$.

C. $\frac{a_x}{M_{ekv.x} \cdot V}$.*

D. $\frac{a_x}{V}$.

11. Select a formula to calculate the titre:

A. $\frac{a_x}{M_x \cdot V}$.

B. $\frac{a_x}{m} \cdot 100$.

C. $\frac{a_x}{M_{ekv.x} \cdot V}$.

D. $\frac{a_x}{V}$.*

12. What is molar concentration?

A. The ratio of the amount of a solute to the volume of the solution.*

B. The ratio of the amount of a solute to the mass of the solution.

C. The ratio of the amount of a solute to the mass of the solvent.

13. What is molar concentration of equivalent?

A. The ratio of equivalents of a solute to volume of the solution.*

B. The ratio of equivalents of a solute to the mass of the solution.

C. The ratio of the amount of a solute to the volume of the solution.

14. What is molal concentration ?

A. The ratio of the amount of a solute to the mass of the solvent.

B. The ratio of the amount of a solute to the mass of the solution.

C. The ratio of the amount of a solute to solvent mass.*

15. What is mole fraction?

A. The ratio of the amount of a solute to the total amount of the components of the solution.*

B. The ratio of the amount of a solute to the amount of the solvent.

C. The ratio of the amount of a solute to solvent mass.

16. What is the titre of the solution ?

A. Mass of a solute in 1 ml of the solution.*

B. Amount of the dissolved substance in 1 ml of the solution.

C. Mass of a solute in 100 ml of the solution.

17. What is molar mass of equivalent ?

A. The product of the equivalence factor and molar mass of the substance.*

B. The product of the equivalence factor and weight of the substance.

C. The ratio of the molar mass of the substance to the factor of equivalence.*

18. Indicate the equivalence factor of permanganate ion in half reaction

$MnO_4- + 8H^+ + 5ē \leftrightarrow Mn^{2+} + 4H_2O$:

A. 1/5.*

B. 5.

C. 8.

19. Indicate the equivalence factor of permanganate ion in half reaction

$MnO_4^- + ē \leftrightarrow MnO_4^{2-}$:

A. 1/5.

B. 5.

C. 8.

D. 1/3.

E. 1.*

20. Indicate the equivalence factor of sulfuric acid in the reaction:

$2NaOH + H_2SO_4 \leftrightarrow Na_2SO_4 + 2H_2O$

A. 1/2.*

B. 2.

C. 1.

D. 1/3.

Literature:

1. Medical Chemistry : textbook / V. A. Kalibabchuk [and ol.] ; ed. by V. A. Kalibabchuk. - K. : Medicine, 2010. - 224 p.

2. Levitin Ye.Ya. General and Inorganic Chemistry : textbook for students of higher schools / Ye. Ya. Levitin, I. A. Vedernikova. - Kharkiv : Golden Pages, 2009. - 360 p.

6. LABORATORY WORK

6.1. Preparation of solutions of a given concentration.

Calculate the weight of the samples and prepare solutions of NaCl and Na_2CO_3: 200 ml of 0.1 M, 250 ml of 0.2 N., 100 ml of 3%.

Weigh the calculated sample of the substance on the balance, put it into a volumetric flask (for 0.1 M and 0.2 N. solutions) or into a cylinder (3% solutions), dissolve in a small amount of distilled water and bring to the volume with the same solvent, stir thoroughly and transfer the solution into a bottle for storage.

7. TRAINING AND CONTROL MEANS:

- cards for determination of initial level of knowledge and skills;
- control questions;
- tests.

Tutorial # 5

1. THEME: Acid-base equilibrium in a body. Hydrogen ion exponent (pH) of biological fluids

2. PURPOSE: To learn to assess and predict the processes, the flow of which depends on the change of the reaction medium.

3. OBJECTIVE GOALS

3.1. To use knowledge of the concentration of hydrogen ions to assess the properties of solutions, direction of chemical reactions.

3.2. To gain the skills of calculation of strong and weak electrolytes solutions pH

3.3. To gain the skills of determination of solutions pH by means of indicators.

4. PLAN AND ORGANIZATIONAL STRUCTURE OF THE TUTORIAL:

4.1. Organizational part …………………………,,,,,,,,…....................5 min

4.2. Goal-setting and motivation of the theme studying (opening address of the teacher)………..............……………...5 min

4.3. Determination of initial knowledge level ………….......................15 min

4.4. Correction of the initial knowledge level.…….….............................25 min

4.5. Organization of independent work of students (target teacher guidance, rules of accident prevention) …………………………………………….5 min

4.6. Laboratory work ……………………………………………..45 min

4.7. Final comtrol: cheking of the laboratory work results and protocols ……………………...…..10 min

4.8. Concluding remarks of the teacher, guidance to the next lesson …………………...……………………………...……………........................3 min

4.8. Concluding remarks of the teacher, guidance to the next lesson……………………………………………………...3 min

5. REFERENCE FOR SELF-STUDY:

5.1. Questions for self-study

1. Electrolyte solutions. The dissociation of strong and weak electrolytes.
2. Dissociation of water. The ionic product of water
3. Protolytic acid-base theory. The main types of protolytic reactions.
4. Calculation of the pH of solutions of strong and weak electrolytes.
5. Biological pH
6. Determination of pH using indicators. Methyl orange, phenolphthalein.
7. Hydrolysis of salts. The degree of hydrolysis, hydrolysis constant. The role of hydrolysis in the biochemical processes.

5.2. Solve the tasks:

Task 1. Calculate the pH of 0,1 M solutions of hydrochloric acid and sodium hydroxide. To show how it changes when the solutions are diluted 100 times.

Answer: 1; 13; 3; 11.

Task 2. Calculate the pH of 1 M solution of phenol.

Answer: 5

Task 3. Calculate the constant and the degree of hydrolysis of 0,01 M solution of HCOOK when K(HCOOH) = $1{,}8\cdot10^{-4}$.

Answer: $5{,}5\cdot10^{-11}$; $7{,}45\cdot10^{-3}$%.

5.3. Examples of the solved tasks:

Task 1. Calculate the pH of 0,01 M solution of HCOOK.

Solution. Formate ion is an anion base, so we use the formula to calculate the pH of weak bases:

$$pH = 7 + ½pK_a + ½lgC_b$$

Reference data: pK_a(HCOOH) = 3,75.

pH = 7 + ½ ·3,75 + ½ lg0,01 = 7 + 1,88 + ½ · (–2) = 7,88.

Answer: 7,88.

Task 2. Calculate the pH of a 0.01 M aqueous solution of ammonium acetate.

Solution. Ammonium acetate is a salt formed by a weak base and a weak acid. Therefore we use the formula:

$$pH = ½ (pK_w + pK_a – pK_b)$$

Reference data:

$pK_a(CH_3COOH)$ = 4,76; $pK_b(NH_4OH)$ = 4,76.

pH = ½ (14 + 4,76 – 4,76) = 7.

Answer: 7.

Task 3. Calculate the constant and the degree of hydrolysis of 0.01 M solution of sodium arsenite $NaAsO_2$, when $K(HAsO_2) = 5{,}8\cdot10^{-10}$.

Solution. $NaAsO_2$ is a salt formed by a strong base and a weak acid, so:

$$K_h = \frac{K_w}{K_a} = \frac{10^{-14}}{5{,}8\cdot10^{-10}} = 1{,}72\cdot10^{-5};\ \alpha_h = \sqrt{\frac{K_w}{K_a\cdot C_b}} = \sqrt{\frac{10^{-14}}{5{,}8\cdot10^{-10}\cdot0{,}01}} = 4{,}15\cdot10^{-2}.$$

or α_h = 4,15%.

Answer: $1{,}72\cdot10^{-5}$; 4,15%.

5.4. Study tests:

1. Indicate the most correct definition of "electrolytes". These are substances :

A Current conductors .

B. Decomposed by an electric current.

C. Solutions and melts which conduct electricity. *

2. Dissociation - is:

A. Decomposition of salt by water with the formation of a weak electrolyte.

B. Decomposition of electrolyte by an electric current.

C. Decomposition of electrolyte into ions under the influence of a polar solvent. *

D. Destruction of red blood cell as a result of osmosis in hypotonic solution.

3.In solution weak electrolyte exists mostly in the form of:

A. Ions.

B. Molecules. *

C. Equal number of ions and molecules.

4. Strong acids include:

A. H_2S.

B. H_2SO_3.

C. H_2SO_4.*

D. HNO_2.

E. HNO_3.*

5. The dissociation constant characterizes the dissociation of:

A a strong electrolyte.

B. a weak electrolyte. *

C. a nonelectrolyte.

6. Which of the following substances in aqueous solution dissociates reversible?

A. $BeC1_2$.

B. HNO_3.

C. H_2CO_3.*

D. $MgCl_2$.

7. What particles concentration is most important in aqueous solution of H_3PO_4 ($\alpha = 17\%$):

A. H^+.

B. H_3PO_4.*

C. PO_4^{3-}.

D. $H_2PO_4^-$.

E. HPO_4^{2-}

8. Using the values of the dissociation constants Ka indicate which acid is the strongest:

A. HNO_2 ($К_а$=5·10^{-4}).*

B. HВгO ($К_а$=2·10^{-9}).

C. CH_3COOH ($К_а$=2·10^{-5}).

D. HCOOH ($К_а$=2·10^{-4}).

E. C_6H_5COOH ($К_а$=6·10^{-5}).

9. Which aqueous solution is a process of dissociation in:

A. NH_3.*

B. CO_2.*

C. Cl_2.

D. SO_2.*

E. N_2.

10. What substances will decrease the degree of dissociation of acetic acid CH_3COOH:

A. HCl.*

B. H_2O.

C. CH_3COONa.*

D. CH_3COOK.*

11. What is the pH of a solution with concentration of H^+ equal to 9.10 mol / l?

A. 5.

B. 7.

C. 9.*

D. 11.

E. 14.

12. What is the concentration of hydrogen ions in a solution with pH of 10?

A. 10^{-6}.

B. 10^{-4}.

C. 10^{-10}.*

D. 10^{-1}.

E. 1.

13. What is the concentration of hydrogen ions in a solution with pOH of 10?

A. 10^{-5}.

B. 10^{-4}.*

C. 10^{-10}.

D. 10^{-1}.

E. 1.

14. pH of the first solution is 6, of the second - 3. How many times the concentration of hydrogen ions in the first solution is less than the second?

A. 2 -fold.

B. 10-fold.

C. 100 -fold.

Д. 1000 -fold.*

E. 10000 -fold.

15. The ionic product of water is:

A. 10^{-7}.

B. 10^{-14}.*

C. 10^{-10}.

D. 7.

E. 14.

16. Select the concentration of OH-ions that corresponds to an acid solution:

A. 10^{-7}.

B. 10^{-9}.*

C. 10^{-4}.

D. 10^{-3}.

E. 10^{-6}.

17. The process of salt destruction under the action of water with forming of a weak electrolyte is called:

A Hemolysis.

B. Hydrolysis. *

C. Electrolysis.

D. Dissociation.

E. Plasmolysis.

18. Salts are not hydrolyzed if they are formed by:

A a strong base and a weak acid.

B. a weak base and a weak acid.

C. a strong base and a weak acid.

D. a strong base and a strong acid. *

19. Hydrolysis is:

A Decomposition of electrolyte into ions.

B. Destruction of erythrocyte in hypotonic solution.

C. Decomposition of electrolyte by electric current.

D. Decomposition of a salt by water with the formation of a weak electrolyte. *

20. What salt is formed after hydrolysis of $AlCl_3$ mainly:

A. $Al(OH)_3$.

B. $Al(OH)_2Cl$.

C. $AlOHCl_2$.*

21. What salt is formed after hydrolysis of Na_3BO_3 mainly:

A. H_3BO_3.

B. NaH_2BO_3.*

C. Na_2HBO_3.

22. Hydrolysis of $FeCl_3$ is mainly:

A. In one stage with the formation of $FeOHCl_2$. *

B. In two stages with the formation of $Fe(OH)_2Cl$.

C. In three stages with the formation $Fe(OH)_3$.

23. pH of Na_2CO_3 solution is:

A Neutral.

B. Acidic.

C. Alkaline. *

24. pH of CH_3COONH_4 aqueous solution is:

A. $pH < 7$.

B. $pH > 7$.

C. $pH \approx 7$.*

25. What salt has $pH < 7$?

A. $ZnCl_2$.*

B. Na_2S.

C. $CHCOONH_4$.

D. $CaCl_2$.

26. What salt aqueous solution has $pH = 7$?

A. Na_2CO_3.

B. Na_2S.

C. Na_2SO_4.*

D. $NaHCO_3$.

E. $CuSO_4$.

27. Indicate which of these salts is hydrolyzed to the greatest extent.

A. $FeCl_2$.

B. NaCl.

C. Na_2SO_3.

D. $CaCl_2$.

E. $FeSO_3$.*

28. What factors increase the hydrolysis of salts?

A. Increased pressure.

B. Increasing of the salt concentration.

C. Dilution. *

D. Cooling.

29. What substance increases $AlCl_3$ hydrolysis when adding it any to the solution?

A. HCl.

B. H_2O.*

C. $AlCl_3$.

D. NaOH.*

30. In which of these solutions is hydrolysis reduced by adding NaOH?

A. NaCl.

B. $(NH_4)_2SO_4$.

C. Na_2CO_3.*

D. $NaNO_3$.

31. In which case will the color of phenolphthalein in a solution of sodium acetate CH3COONa be more intense?

A. When the solution is heated. *

B. Upon cooling of the solution.

C. With vigorous stirring.

32. Which of the following salts is hydrolyzed by the cation?

A. Sodium sulfate.

B. Sodium sulfite.

C. Ammonium sulphate. *

D. Sodium sulphide.

33. Which of the following salts is hydrolyzed by the anion?

A. Sodium sulfate.

B. Potassium Sulphate.

C. Ammonium sulphate.

D. Sodium sulphide. *

34. Choose only the acids:

A. $H_2PO_4^-$.

B. PO_4^{3-}.

C. H_3PO_4.*

D. NH_3.

E. NH_4^+.*

35. Choose only the bases:

A. $H_2PO_4^-$.

B. PO_4^{3-}.*

C. H_3PO_4.

D. NH_3.*

E. NH_4^+.

36. Choose ampholytes:

A. $H_2PO_4^-$.*

B. PO_4^{3-}.

C. H_3PO_4.

D. NH_3.

E. NH_4^+.

37. The ionic product of water is:

A. Sum of the concentrations of hydrogen ions and hydroxide ions.

B. The ratio of the concentration of hydrogen ions and hydroxide ions.

C. The product of the concentration of hydrogen ions and hydroxide ions. *

38. At pH = 2 the ionic product of water is:

A. 10^{-2}.

B. 10^{-4}.

C. 10^{-14}.*

D. 10^{-10}.

E. 10^{-12}.

39. pH is:

A. The natural logarithm of the hydrogen ion concentration.

B. The negative natural logarithm of the hydrogen ion concentration.

C. The negative decimal logarithm of the hydrogen ion concentration. *

D. The negative decimal logarithm of the acid concentration.

E. The natural logarithm of the acid concentration.

40. pOH is:

A. The natural logarithm of the concentration of hydroxyl ions.

B. The negative natural logarithm of the concentration of hydroxyl ions.

C. The negative decimal logarithm of the concentration of hydroxyl ions. *

D. The negative decimal logarithm of the base concentration.

E. The natural logarithm of the alkali concentration.

41. Pepsin of gastric juice acts in:

A the neutral media.

B. the alkaline media.

C. the acidic media. *

42 . pH of 0,001 M nitric acid solution is:

A. 3.*

B. 4.

C. 11.

D. 10.

43. pH of 0,0001 M hydrochloric acid solution is:

A. 3.

B. 4.*

C. 11.

D. 10.

44. pH of 0,001 M sodium hydroxide solution is:

A. 3.

B. 4.

C. 11.*

D. 10.

45. pOH of 0,001 M potassium hydroxide solution is:

A. 3.*

B. 4.

C. 11.

D. 10.

Literature:

1. Medical Chemistry : textbook / V. A. Kalibabchuk [and ol.] ; ed. by V. A. Kalibabchuk. - K. : Medicine, 2010. - 224 p.

2. Levitin Ye.Ya. General and Inorganic Chemistry : textbook for students of higher schools / Ye. Ya. Levitin, I. A. Vedernikova. - Kharkiv : Golden Pages, 2009. - 360 p.

3. Inorganic Chemistry : textbook / ed.: D. F. Shriver, P. W. Atkins. - 3th ed. - [s. l.] : Oxford University Press, 2004. - 763 p.

4. Petrucci R.H., Harwood W.S., Herring F.G. General Chemistry. Principles and Modern Applications. – 8th ed. – Prentice Hall, 2002. – 1160 p.

6. LABORATORY WORK

6.1. The study of methyl orange (MO) and phenolphthalein (FF) indicators colouring in acidic and alkaline media

Put test solutions and proposed indicators in a series of test tubes according to the table.

Put down the results of the experiments into the table:

Test tube	Solution+indicator	Colour
1	1 ml HC1 + 2 drops MO.	
2	1 ml NaOH + 2 drops MO.	
3	1 ml HC1 + 2 drops FF.	
4	1 ml NaOH + 2 drops FF.	

6.2. Determination of solutions pH with universal indicator paper

Apply the test solution on a strip of universal indicator paper with a glass rod. Compare the colour of the paper with a standard colour scale of pH. Put down the results of the experiments into the table:

Test	Solution	Colour	pH
1	1 мл HC1		
2	1 мл NaOH		

6.3. Determination of solution pH with ionomer (pH meter)

The cell for pH determination is made of glass indicator electrode and silver chloride reference electrode. Universal ionomer ЭВ-74 must be set up according to the instructions. Electrodes must be thoroughly washed with distilled water before dipping and wiped with filter paper. pH is measured in separate portions of the sample solution. Carry out several parallel measurements. Put down the results of the experiments into the table:

Solution	pH		
	pH_1	pH_2	$pH_{cp.}$
Stock solution CH_3COOH			
Dilution 1:2			
Dilution 1:4			

7. TRAINING AND CONTROL MEANS:

– cards to identify initial knowledge and skills;

– control questions;

– tests.

Tutorial # 6

1. THEME: Buffer systems

2. PURPOSE: To learn to assess the effect of the buffer systems to maintain a certain pH.

3. OBJECTIVES

3.1. To gain the skills to prepare buffers with given pH.

3.2. To use the knowledge about the mechanism of action of the buffer systems for predicting biochemical processes in the body fluids when the reaction medium changes.

3.3. To learn to identify the boundaries of the buffer system with acids and alkalis.

4. PLAN AND ORGANIZATIONAL STRUCTURE OF THE TUTORIAL:

4.1. Organizational part …………………………,,,,,,,,,…....................5 min

4.2. Goal-setting and motivation of the theme studying (opening address of the teacher)……….............……………..5 min

4.3. Determination of initial knowledge level ………………......................15 min

4.4. Correction of the initial knowledge level.……….............................25 min

4.5. Organization of independent work of students (target teacher guidance, rules of accident prevention) ……………………………………………….5 min

4.6. Laboratory work ………………………………………………..45 min

4.7. Final comtrol: cheking of the laboratory work results and protocols …………………………...…...10 min

4.8. Concluding remarks of the teacher, guidance to the next lesson ……………………...……………………………...…………….......................3 min

5. REFERENCE FOR SELF-STUDY:

5.1. Questions for self- study

1. Buffer systems, protolytic equilibrium in them.
2. Buffer capacity, the factors on which it depends.
3. Practical definition of buffer capacity by acid and alkali.
4. Examples of the buffer systems: phosphate, hydrocarbonate. Show the mechanism of action by the reaction equation.
5. Calculation of the buffer system pH.
6. Buffer systems of the human body: protein, hemoglobin. The mechanism of action.
7. The significance of the buffer systems to organisms.
8. Acid-base status of the body. Acidosis, alkalosis (respiratory, metabolic).

5.2. Solve the tasks:

Task 1. Calculate the pH of a buffer mixture containing 0,01 mole of formic acid and 0,2 M of potassium formate.

Answer: 5,05.

Task 2. Calculate the required volumes of solutions of CH_3COONa (C_{Me}=0,1 mol/l) and CH_3COOH (C_{Me}=0,1 mol/l) for preparing of 3 l of acetate buffer, pH 5,24.

Answer: 2250 ml; 750 ml.

Task 3. The titration of 100 mL of serum (pH = 7,36) took 3.6 mL of 0.05 M HCl. Calculate the buffering capacity of the blood, if the pH changed to 7,0.

Answer: 0,005 mol/l.

5.3. Examples of the solved tasks:

Task 1. Calculate the concentration of H +, OH-, and pH of the solution obtained by mixing 50 ml of 0,2 M acetic acid and 30 ml of 0,1 M sodium acetate solution.

Solution. As the solution contains a weak acid and its conjugate base, it is a buffer system, the pH of which can be calculated according to the formula:

$$pH = pK_a + \lg\frac{C_b}{C_a}$$

Reference data: $pK_a(CH_3COOH) = 4,76$.

$$pH = pK_a + \lg\frac{C_b}{C_a} = 4,76 + \lg\frac{0,1 \cdot 30}{0,2 \cdot 50} = 4,76 + \lg 0,3 = 4,24;$$

$$[H^+] = 10^{-pH} = 10^{-4,24} = 5,79 \cdot 10^{-5};$$

$$[OH^-] = \frac{K_w}{[H^+]} = \frac{10^{-14}}{5,79 \cdot 10^{-5}} = 1,73 \cdot 10^{-10}.$$

Answer: 4,24; $5,79 \cdot 10^{-5}$; $1,73 \cdot 10^{-10}$.

Task 2. Calculate the buffering capacity of acetate buffer containing 7,00 ml of acetic acid(C_{Me}=0,1 mol/l) and 3,00 ml of sodium acetate(C_{Me}=0,1 mol/l). The titration of the buffer system with phenolphthalein took 2,00 ml NaOH (C_{Me}=0,5 mol/l).

Solution. Buffer capacity for alkali calculated as:

$$B_{(NaOH)} = \frac{C_{Me\,(NaOH)} \cdot V_{(NaOH)}}{V_{buff.sol.} \cdot (pH_1 - pH_0)}$$

$pK_a(CH_3COOH) = 4{,}76$.

$$pH_0 = pK_a + \lg\frac{C_b}{C_a} = 4{,}76 + \lg\frac{0{,}1 \cdot 3}{0{,}1 \cdot 7} = 4{,}76 - 0{,}37 = 4{,}39.$$

pH_1 – corresponds to the lower boundary of the color change range of phenolphthalein indicator (8,2).

$$B_{(NaOH)} = \frac{0{,}5 \cdot 0{,}002}{0{,}01 \cdot (8{,}2 - 4{,}39)} = 0{,}026\ \text{mol/l}$$

Answer: 0,026 mol/l.

5.4. Study tests:

1. Buffer systems is aqueous solutions which maintain a constant:

A. Osmotic pressure.

B. Temperature.

C. pH value. *

D. Potential difference.

2. The buffer system is:

A. a redox pair.

B. an acid-base pair. *

S. a galvanic pair.

3. Acidic buffer solutions consist of:

A A strong acid and a weak base.

B. A weak acid and a strong base.

C. A weak acid and its conjugate base. *

D. A strong acid and its conjugate base.

4. Basic buffer solutions consist of:

A A weak base and a strong acid.

B. A strong base and a weak acid.

C. A weak base and its conjugate acid. *

D. A strong base and its conjugate acid.

5. What substance solution can be added to an aqueous solution of CO_2 to form a buffer system?

A. H_2CO_3.

B. $NaHCO_3$.*

C. NaOH.

D. Na_2CO_3.

E. HCl.

6. What substance solution can be added to an aqueous solution of NH_3 to form a buffer system?

A. NH_4Cl.*

B. HNO_3.

C. NaOH.

D. NaCl.

E. HCl.

7. What substance binds strong acid added to the buffer solution:
A Strong base.
B. Proton acceptor. *
C. Strong acid.
D. Proton donor.

8. What substance binds alkali added to the buffer solution:
A Strong base.
B. Proton acceptor. *
C. Strong acid.
D. Proton donor. *

9. Excess H + or OH-ions are bound components of the buffer solution forming:
A Strong electrolyte.
B. Ampholyte.
C. Weak electrolyte. *
D. Nonelectrolyte.

10. When adding HCl to a buffer solution pH does not changes because of forming of:
A A strong base.
B. A weak acid. *
C. A strong acid.
D. A weak base.
E. Water.

11. When adding NaOH to a buffer solution pH does not changes because of forming of:

A A strong base.

B. A weak acid.

C. A strong acid.

D. A weak base.

E. Water. *

12. What is pH of acetate buffer with concentration ratio of components equal to 1?

A. 1.00.

V. 4.76. *

C. 7.00.

D. 9.25.

E. 11,000.

13. In phosphate buffer, add acid is bound by:

A. HPO_4^{2-} - ion.*

B. $H_2PO_4^{-}$- ion.

C. Acid H_3PO_4..

D. $PO_4{}^{3-}$ - ion.

14. Strong alkalis added to phosphate buffer are bound by:

A. HPO_4^{2-} - ion.

B. $H_2PO_4^{-}$- ion. *

C. Acid H_3PO_4..

D. $PO_4{}^{3-}$ - ion.

15. Buffer system of blood does not include:

A. Aacetate. *

B. Protein.

S. Bicarbonate.

D. Phosphate.

E. Hemoglobin.

16. In physiological media CO2 forms:

A Carbonate.

B. Hydrocarbons. *

S. Carbonates and bicarbonates.

17. What equilibrium are in the red blood cells after CO_2 absorption?:

A. $CO_2 + H_2O \leftrightarrow 2H^+ + CO_3^{2-}$.

B. $CO_2 + H_2O \leftrightarrow H^+ + HCO_3^-$.

C. $CO_2 + H_2O \leftrightarrow H_2CO_3 \leftrightarrow H^+ + HCO_3^-$.*

18. Which of the proposed buffer systems of the body has a function of conjugate base (proton acceptor):

A. HHb .

B. H_3N^+ - COOH.

C. $NaHCO_3$.*

D. KH_2PO_4.

E. $HHbO_2$.

19. In the tissue capillaries excess of acid (H^+ ions) is bound by:

A. HHb .

B. Hb^-.

C. HCO_3^-.

D. $HHbO_2$.

E. HbO_2^-.*

20. Pulmonary capillary blood pH remains constant as excess of H^+ ions is bound by:

A. HHb .

B. Hb^-.

C. HCO_3^-.*

D. $HHbO_2$.

E. HbO_2^-.

21. Select one that is proton donor (a weak acid) of the buffer systems of the body,:

A. $NaHCO_3$.

B. HPO_4^{2-}.

C. H_3N^+- Prot-COO^-.*

D. HbO_2^-.

22. What is the weak acid in acidic protein buffer:

A. H_2N-Prot-COO^-.

B. H_3N^+- Prot-COO^-.

C. H_3N^+- Prot-COOH.*

23. What is the base in basic protein buffer:

A. H_2N- Prot-COO^-.*

B. H_3N^+- Prot-COO^-.

C. H_3N^+- Prot-COOH.

24. In which of the buffer systems is acid buffer capacity the maximum:

A. 100 ml 0,1 M HNO_2 + 50 ml 0,2 M KNO_2.*

B. 20 ml 0,1 M HNO_2 + 10 ml 0,1 M KNO_2.

C. 50 ml 0,2 M HNO_2 + 20 ml 0,1 M KNO_2.

25. How will buffer solution pH change after dilution?

A. Increase.

B. Decrease.

C. Do not change. *

26. How will buffer capasity change after dilution?

A. Increase.

B. Decrease. *

C. Do not change. *

27. Choose the two most powerful buffers of red blood cells:

A. Acetate.

B. Amino acid.

S. Bicarbonate. *

D. Phosphate.

E. Hemoglobin. *

28. Acetate buffer contains:

A. $CH_3COOH+CH_3COOC_2H_5$.

B. $CH_3COOH+NaOH$.

C. $CH_3COOH+CH_3COONa$.*

29. Components ratio of bicarbonate buffer of the blood plasma at $pH = 7.36$ should be:

A. 15:1.

B. 20:1.*

C. 7:1.

30. Phosphate buffer contains:

A. $H_3PO_4+ NaH_2PO_4$.

B. $NaH_2PO_4+Na_2HPO_4$.*

C. Na_3PO_4+ Na_2HPO_4.

31. Acidosis is the shift of blood pH to:

A. The acidic side. *

V. The alkaline side.

C. To pH = 7.

32. Alkalosis is the shift of blood pH to:

A. The acidic side.

V. The alkaline side. *

C. To pH = 7.

33. Blood pH is:

A. 7,36.*

B. 1,86.

C. 6,02.*

D. 5,56.

34. Gastric pH is:

A. 7,1-8,3.

B. 3,1-4,4.

C. 0,9-1,5.*

D. 8,0-10,5.

35. Urine pH is:

A. 5,0-6,5.*

B. 8,0-10,5.

C. 3,1-4,4.

D. 7,1-8,3.

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2. Levitin Ye.Ya. General and Inorganic Chemistry : textbook for students of higher schools / Ye. Ya. Levitin, I. A. Vedernikova. - Kharkiv : Golden Pages, 2009. - 360 p.

3. Inorganic Chemistry : textbook / ed.: D. F. Shriver, P. W. Atkins. - 3th ed. - [s. l.] : Oxford University Press, 2004. - 763 p.

4. Petrucci R.H., Harwood W.S., Herring F.G. General Chemistry. Principles and Modern Applications. – 8th ed. – Prentice Hall, 2002. – 1160 p.

6. LABORATORY WORK

6.1. Effect of acid and alkali on the pH of the buffer system

Put 5 ml CH_3COOH (C_{Me} = 0,1 mol/l) and 5 ml CH_3COONa (C_{Me} = 0,1 mol/l) into the test tube. Pour the resulting buffer system equally into three tubes. Add 3 drops of HC1 (C_{Me} = 0,1 mol/l) into the first test tube. Add 3 drops of NaOH (C_{Me} = 0,1 mol/l) into the second test tube. Add 2 drops of methyl orange indicator into each tube. Compare the color of the solutions, write down the reactions and draw conclusions.

6.2. Effect of dilution on the buffer system pH

Prepare the buffer system (as 6.1.) and pour equally into two test tubes. Add 6 ml of distilled water to the first test tube. Add 2 drops of methyl orange indicator to each tube. Compare the color of the solutions and draw conclusions.

6.3. Determination of buffer capacity of acetate buffer system

10,00 ml of acetate buffer system, which contains 7,00 ml of acetic acid (C_{Me} = 0,1 mol/l) and 3,00 ml of sodium acetate (C_{Me} = 0,1 mol/l), should be prepared in a titration flask. Add 2 drops of phenolphthalein to this solution and titrate with alkali (C_{Me} = 0,1 mol/l) until pink colou appearsr. Calculate the buffer capacity by alkali:

$$B_{(NaOH)} = \frac{C_{Me\,(NaOH)} \cdot V_{(NaOH)}}{V_{buff.sol.} \cdot (pH_1 - pH_0)}, \text{ where}$$

B – buffer capacity of the solution by alkali;

$C_{Mэ(NaOH)}$ – equivalent molar concentration of alkali;

$V_{(NaOH)}$ – volume of alkali;

$V_{buff.sol.}$ – volume of the buffer system;

pH_0 – initial pH value of the buffer solution;

pH_1 – final pH value of buffer solution.

7. TRAINING AND CONTROL MEANS:

– cards to identify initial knowledge and skills;
– control questions;
– tests.

Tutorial #7

1. THEME: Fundamentals of titrimetric analysis. Acid-base titration

2. PURPOSE: To learn the basics of quantitative analysis, to study the general principles of titrimetric methods and techniques of acid-base titration

3. OBJECTIVE GOALS

3.1. To understand the objectives, principles and methods of quantitative analysis.

3.2. To learn the basics of titrimetric methods of analysis, learn how to use the analytical measuring glassware.

3.3. To gain skills of alkalimetric titration.

3.4. To learn how to interpret the results of the analysis in clinical practice.

4. PLAN AND ORGANIZATIONAL STRUCTURE OF THE TUTORIAL:

4.1. Organizational part …………………………,,,,,,,,…....................5 min

4.2. Goal-setting and motivation of the theme studying (opening address of the teacher)……….............……………...5 min

4.3. Determination of initial knowledge level ……………...............................15 min

4.4. Correction of the initial knowledge level.…………..............................25 min

4.5. Organization of independent work of students (target teacher guidance, rules of accident prevention) ……………………………………………….5 min

4.6. Laboratory work ………………………………………………..45 min

4.7. Final comtrol: cheking of the laboratory work results and protocols …………………………....…...10 min

4.8. Concluding remarks of the teacher, guidance to the next lesson …………………...…………………………...…………….........................3 min

5. REFERENCE FOR SELF-STUDY:

5.1. Questions for self- study

1. The essence of the quantitative analysis.
2. The essence of titrimetric methods of analysis. Basic concepts. Classification of titrimetric methods on the type of chemical reaction.
3. Volume measurement. Measuring glassware.
4. Standard solutions (titrants) and methods of their preparing. Primary and secondary standards. Ways of expressing of standard solutions concentration.
5. What reactions underlie the acid-base titration? What is the essence of acidimetry and alkalimetry?
6. Titration curves, drastic change, equivalence point.
7. Acid-base indicators. Transition interval of indicator, the principle of indicator selecting for titration.
8. Application of alkalimetry in clinical analysis, in sanitary and hygienic studies.

5.2. Solve the tasks:

Task 1. Calculate the equivalent molar concentration of sodium hydroxide if the titration of 20,00 ml of 0,05075 M solution of oxalic acid consumed 19,50 ml of this solution.

Answer: 0,1041 mol/l.

Task 2. Calculate the sample of sodium tetraborate decahydrate ($MNa_2B_4O_7{\cdot}10H_2O$ = 381,4 г/моль) to prepare 200,0 ml of 0,05000 N. solution.

Answer: 1,907 g.

5.3. Examples of the solved tasks:

Task 1 The titration of 20,00 ml of 0,09218 N. solution of sodium tetraborate consumed
19,42 ml of solution of hydrochloric acid. Calculate the equivalent molar concentration and the titre of HCl solution.

Solution. According to the law of equivalents:

$$C_{Me}(Na_2B_4O_7) \cdot V(Na_2B_4O_7) = C_{Me}(HCl) \cdot V(HCl),$$

then

$$C_{M_Э(HCl)} = \frac{C_{M_Э(Na_2B_4O_7)} \cdot V_{(Na_2B_4O_7)}}{V_{(HCl)}} = \frac{0,09218 \cdot 20,00}{19,42} = 0,09493\,\text{моль/л.}$$

Calculate the titre of HCl solution according to the formula:

$$T = C_{Me}(HCl) \cdot Me'(HCl);$$

M(HCl) = 36,46 g/mol;
fe(HCl) = 1;
Me′(HCl) = M(HCl) · fe(HCl) : 1000 = 0,03646.

T = 0,09493 · 0,03646 = 0,003461.

Answer: 0,003461 g/ml; 0,09493 mol/l.

5.4. Study tests:

1. Is a solution prepared by sample accurately weighed a solution with a

known titer?

A. It is possible, because it is used for titration.

B. No. it is not.

C. It can be, because the title of the solution is easy to calculate. *

2. Titrant (standard solution, working solution) is called reagent solutions:

A Of known concentration C and used in the titration. *

B.Of determined concentration.

C. Of known titre.

D. Of known concentration.

E. Used in the titration of the analyte.

3. Specify how to prepare standard solution:

A. Prepared and determined titre. *

B. Qualitative and quantitative.

C. Direct and indirect.

D. Pipetting and separate samples.

E. Indication and instrumental.

4. Specify the primary way of titrants concentration expressing:

A Equivalent molar concentration. *

B. Titre.

C. The titer for the analyte.

D. Percentage concentration.

E. Molar concentration.

5. The reactions used in titrimetric analysis methods must meet the requirements:

A. proceeding to the stoichiometric ratios. *

B. being fast enough. *

C. It must be possible to fix the point of equivalence. *

D. Should not go to adverse reactions. *

E. The reaction should proceed quantitatively. *

6. Which of the following laws is the theoretical basis of titrimetric methods of analysis:

A. The law of equivalents. *

B. The law of conservation of mass substances.

C. The law of multiple proportions.

D. The volume ratio law.

7. The end point in the method of acid-base titration is determined by the pH indicator. What should be taken into account in the selection of pH indicators:

A. The pH at the equivalence point. *

B. Nature of the titrant.

C. The properties of the reaction products.

D. The change in pH during the titration.

E. The nature of the analyte.

8. Primary standards must meet the requirements of:

A. To be chemically pure and meet the chemical formula.

B. To be stable when stored in solid form and in solution.

C. Have a large equivalent molar mass.

D. To be easily purified by recrystallization or otherwise.

E. All of the above. *

9. In the method of acid-base titration volumetric solution (titrant) is prepared:

A.By the method of prepared titre. *

B. By the method of determined titre. *

C. From the standard titres. *

D. By the method of pipetting.

E. By the method of separate portions.

10. The equipment that is necessary to perform the titration method:

A. Volumetric glassware for accurate volume measuring.

B. Volumetric glassware for rough volume measuring.

C. Indicators for fixing the equivalence point.

D. Titrants.

E. All the above. *

11. What methods of acid-base titration allow to determine the contents in of sulfuric acid solution:

A. alkalimetry. *

B. acidimetry.

C. Either one or the other.

D. Neither one nor the other.

12. Titration curve is:

A Graphic representation of the pH changes in the titration. *

B. The sharp change in pH near the equivalence point.

C. 0.1% before the equivalence point.

D. Strictly equivalent amount of titrant is added..

E. 0.1% after the equivalence point.

13. Titration drastic change - is:

A Sudden change in pH near the equivalence point. *

B. Graphic representation of pH changes in the titration.

C. 0.1% before the equivalence point.

D. Strictly equivalent amount of titrant is added..

E. 0.1% after the equivalence point.

14. The start of the drastic change is:
A. 0.1% before the equivalence point.*
B. Graphic representation of pH changes in the titration.
C. The sharp change in pH near the equivalence point.
D. Strictly equivalent amount of titrant is added.
E. 0.1% after the equivalence point.

15. The start of the drastic change is:
A. 0.1% after the equivalence point.*
B. Graphic representation of pH changes in the titration.
C. 0.1% before the equivalence point.
D. Strictly equivalent amount of titrant is added.
E. The sharp change in pH near the equivalence point

16. The equivavalence point is:
A. Strictly equivalent amount of titrant is added. *
B. Graphic representation of pH changes in the titration.
C. 0.1% before the equivalence point.
D. The sharp change in pH near the equivalence point.
E. 0.1% after the equivalence point.

17. Find the definition of " indicator transition interval ":
A. The pH at which there is a visible change in color of the indicator. *
B. The value of the pH at which titration with the indicator is completed.
C. The amount of titrant is strictly equivalent to the amount of analyte.
D. Time of the titration, when the pH is 7.
E. Endpoint of the titration.

18. For the titration curve construction should be calculated pH at different titration points:

A. Of the stock solution *

B. At a half-neutralization point. *

C. At the start of the drastic change. *

D. At the equivalence point. *

E. At the end of the drastic change. *

19. In the selection of pH-indicators value the most significant is:

A. pH at the equivalence point. *

B. Nature of the titrant.

C. The properties of the reaction products.

D. The nature of the analyte.

E. Change in pH during the titration.

20. What requirements should the reactions in titrimetric analysis meet:

A. Adequate reaction rate. *

B. Stoichiometric ratio of the components. *

C. Ability to fix the point of equivalence. *

D. Reaction products must be soluble.

E. The reaction products should be colorless.

21. Specify how pH indicators are selected:

A. Qualitative and quantitative. *

B. Pipetting and separate samples.

C. Direct and indirect.

D. Prepared and determined titre.

E. Indication and instrumental.

22. What indicators are used in cases when titrating NH_3 with HCl solution:

A. Phenolphthalein.

B. Methyl orange. *

C. Either one or the other.

D. Neither one nor the other.

23. Choose the right procedure if the substance reacts with the titrant stoichiometrically with sufficient speed:

A. Direct titration.*

B. Back titration.

C. Substitution titration.

D. Separate samples technique.

24Choose the right l procedure if the substance reacts with the titrant stoichiometrically but slowly:

A. Back titration. *

B. Direct titration.

C. Substitution titration.

D. Separate samples technique.

25. Choose the right procedure if the substance reacts with the titrant quickly but not stoichiometrically:

A. Substitution titration.*

B. Direct titration.

C. Back titration

D. Separate samples technique.

26. Choose the right procedure if the substance does not directly interact with the titrant:

A. Substitution titration.*

B. Direct titration.

C. Back titration

D. Separate samples technique.

27. Specify titration methods:
A. Direct and indirect. *
B. Qualitative and quantitative.
C. Pipetting and separate portions.
D. Prepared and determined titre.
E. Indication and instrumental.

28. Specify titration techniques:
A. Pipetting and separate portions. *
B. Qualitative and quantitative.
C. Direct and indirect.
D. Prepared and determined titre.
E. Indication and instrumental.

29. What law are calculations in volumetric analysis based on:
A. Law of equivalents. *
B. Law of mass conservation.
C. Law of constant composition.
D. Avogadro's law
E. Rault’s Law.

30. What is necessary for the preparation of burette for titration:
A. Rinse it with a titrant. *
B. Rinse it with analyzed solution.
C. Rinse it with distilled water.

31. The substances used to prepare titrants must:
A. Be solid.
B. Not change the composition during the storage in the solid state and in

solution. *

C. Be colorless.

D. Not have odor.

32. What belongs to volumetric glassware for accurate measuring:

A. Burette. *

B. Graduated cylinder.

C. Beaker.

D. Volumetric tube.

33. HCl with exactly known concentration can be prepared in a volumetric flask only by:

A. taking the accurate sample of HCl at analytical balance.

B. taking the accurate amount of HCl with a pipette.

C. from fixanal. *

34. The volume of a sample solution for titration is metered by:

A. Burette.

B. Graduated cylinder.

C. Pipette. *

D. Volumetric tube.

35. Before each titration it is necessary to:

A. Bring the titrant volume to zero. *

B. Record the division of the burette from which the titration starts.

C. Bring the titrant volume to the nearest whole number.

36. The solutions of exact concentration are prepared in the:

A. Burette.

B. Graduated cylinder.

C. Beaker.

D. Volumetric flask. *

37. How the last drop can be removed from a pipette?

A. Touch the end of the pipette to the wall of the flask. *

B. Blown last drop.

C. Shake the pipette.

38. Titration process is carried out using:

A. Burette. *

B. Graduated cylinder.

C. Beaker.

D. Volumetric flask.

39. What condition is incorrect: "Reactions underlying titrimetric analysis should be ...":

A. Fast.

B. Not be accompanied by side effects.

C. Exothermic. *

D. Irreversible.

40. When transferring the sample pipette should be held:

A. Vertically. *

B. Horizontally.

C. Turning up the spout.

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1. Quantitative Chemistry Analysis : textbook / D. C. Harris. – 7th ed. : W.H Freeman and Company, 2007. – 808 p.

2. Basic Concepts of Analytical Chemistry : textbook / S.M. Khopkar. – 2nd ed.. : New Age International., 1998. – 500 p.

6. LABORATORY WORK

6.1. Determination of hydrochloric acid mass fraction

Method – direct titration

Pipetting technique

Chemical reaction:

$$HCl + NaOH \leftrightarrow NaCl + H_2O$$

$$H^+ + OH^- \leftrightarrow H_2O$$

Preliminary calculations:

M(HCl) = 36,46 g/mol;

fe(HCl) = 1;

Me(HCl) = fe(HCl) · M(HCl) = 1 · 36,46 = 36,46 g/mol;

Me′(HCl) = Me(HCl) : 1000 = 0,03646.

General procedure: put 5,00 ml of pharmacopoeia hydrochloric acid (ρ=1,04 g/sm^3) into 100,0 ml volumetric flask, bring to the mark with distilled water and mix. Then transfer 5,00 ml of the resulting solution with measuring pipette into titration flask, add 1-2 drops of phenolphthalein indicator and titrate from the burette with 0,1 M solution of sodium hydroxide until a stable (not disappearing within 30 sec) pale pink color of the solution achived. Repeat the titration until three reproducible results achived.

Mass fraction (in %) of hydrochloric acid in the preparation is calculated by the average amount of sodium hydroxide according to the formula:

$$\omega_{(HCl)} = \frac{V_{(NaOH)} \cdot C_{M\,(NaOH)} \cdot M_{e(HCl)} \cdot V_f \cdot 100\%}{a_{(HCl)} \cdot V_p \cdot 1000}$$

7. TRAINING AND CONTROL MEANS:

– cards to identify initial knowledge and skills;

– control questions;

– tests.

Tutorial # 8

1. THEME: Acid-base titration. Acidimetry, alkalimetry

2. PURPOSE: To understand the essence of acidimetry and learn how to use this method in clinical analysis, in sanitary and hygienic studies

3. OBJECTIVES

3.1. To gain the skills of acidimetric titration.

3.2 To learn how to interpret the results of the analysis in clinical practice.

4. PLAN AND ORGANIZATIONAL STRUCTURE OF THE TUTORIAL:

4.1. Organizational part …………………………,,,,,,,,,…....................5 min

4.2. Goal-setting and motivation of the theme studying (opening address of the teacher)……….............……………...5 min

4.3. Determination of initial knowledge level ………….......................15 min

4.4. Correction of the initial knowledge level.……….............................25 min

4.5. Organization of independent work of students (target teacher guidance, rules of accident prevention) ……………………………………………….5 min

4.6. Laboratory work ………………………………………………..45 min

4.7. Final comtrol: cheking of the laboratory work results and protocols ……………………....…...10 min

4.8. Concluding remarks of the teacher, guidance to the next lesson …………………...……………………………...…………….........................3 min

5. REFERENCE FOR SELF-STUDY:

5.1. Questions for self- study

1. Acid-base titration, the basic equation of the method..
2. Acidimetry. Titrants and methods of their preparing. Primary and secondary standards.
3. Titration curves.
4. Acid-base indicators.
5. Application of acidimetry in clinical analysis, in sanitary and hygienic studies.

5.2. Solve the tasks:

Task 1. Calculate sample of sodium carbonate (fe = 1/2) to prepare 400,0 ml of 0,05000 N. solution.

Answer: 1,06 g.

Task 2. The sample of 1,.083 g of sodium carbonate is dissolved in a volumetric 200,0 ml flask. Calculate the molar concentration and the titre of the resulting solution.

Answer: 0,005415 g/ml; 0,05110 mol/l.

5.3. Examples of the solved tasks:

Task 1 25.00 ml of detergent placed in 250,0 ml flask and brought with water to the mark. The titration of 5,00 ml of this solution took 4,00 ml of 0,025 M solution of hydrochloric acid. Calculate the ammonia content in g/l, assuming that the alkalinity of the sample is determined only by the component.

Solution. Calculate how many grams of ammonia are contained in 1000.0 ml:

$$m = \frac{V_{(HCHC} \cdot C_{(HCl)} \cdot M_{e(NH_3)} \cdot V_f \cdot 1000{,}0}{1000 \cdot V_p \cdot a}$$

$$m = \frac{4{,}00 \cdot 0{,}02500 \cdot 17{,}03 \cdot 250{,}0 \cdot 1000{,}0}{1000 \cdot 5{,}00 \cdot 25{,}00} = 3{,}40\ \mathrm{g/l}$$

Answer: 3,40 g/l.

5.4. Study tests (p. 149-155)

Literature:

1. Quantitative Chemistry Analysis : textbook / D. C. Harris. – 7th ed. : W.H Freeman and Company, 2007. – 808 p.

2. Basic Concepts of Analytical Chemistry : textbook / S.M. Khopkar. – 2nd ed.. : New Age International., 1998. – 500 p.

6. LABORATORY WORK

6.1. Determination of mass content of ammonia in aqueous ammonia solution

Method – Back titration (titration by excess)

Pipetting technique

Chemical reaction:

$$NH_3 \cdot H_2O + HCl \leftrightarrow NH_4Cl + H_2O$$

$$NH_3 \cdot H_2O + H^+ \leftrightarrow NH_4 + H_2O$$

$$HCl_{(res.)} + NaOH \leftrightarrow NaCl + H_2O$$

$$H^+_{(res.)} + OH^- \leftrightarrow H_2O$$

Preliminary calculations:

$M(NH_3) = 17{,}03$ g/mol;

$fe(NH_3) = 1$;

$Me(NH_3) = fe(NH_3) \cdot M(NH_3) = 17{,}03$ g/mol.

General procedure: put the sample into the 50,00 ml volumetric flask quantitatively, bring with distilled water to the mark and mix thoroughly. Transfer 10,00 ml of the resulting solution into a titration flask in which 20,00 ml 0,1000 M solution of hydrochloric acid should be put beforehand. Add 1-2 drops of methyl orange indicator and titrated with 0,1000 M sodium hydroxide to change the colour of the solution from pink to orange. Repeat the titration until three reproducible results achieved.

The gram amount of ammonia in aqueous ammonia solution is calculated by the average amount of sodium hydroxide according to the formula:

$$m_{(NH_3)} = \frac{(C_{M\,(HCl)} \cdot V_{(HCl)} - C_{M\,(NaOH)} \cdot V_{(NaOH)}) \cdot M_{e(NH_3)} \cdot V_f}{V_p \cdot 1000}$$

7. TRAINING AND CONTROL MEANS:

– cards to identify initial knowledge and skills;

– control questions;

– tests.

Tutorial # 9

1. THEME: Colligative properties of solutions

2. PURPOSE: To learn how to use data on the osmotic pressure in medical practice

3. OBJECTIVES

3.1. To gain the skills of calculation of osmotic pressure, osmotic concentration, isotonic coefficient.

3.2. To learn how to interpret the results of the analysis in clinical practice.

4. PLAN AND ORGANIZATIONAL STRUCTURE OF THE TUTORIAL:

4.1. Organizational part …………………………,,,,,,,,…....................5 min

4.2. Goal-setting and motivation of the theme studying (opening address of the teacher)………..............……………...5 min

4.3. Determination of initial knowledge level ……………......................15 min

4.4. Correction of the initial knowledge level.……….............................25 min

4.5. Organization of independent work of students (target teacher guidance, rules of accident prevention) ……………………………………………….5 min

4.6. Laboratory work ……………………………………………..45 min

4.7. Final comtrol: cheking of the laboratory work results and protocols ………………………...…..10 min

4.8. Concluding remarks of the teacher, guidance to the next lesson …………………...……………………………...………….......................3 min

5. REFERENCE FOR SELF-STUDY:

5.1. Questions for self- study

1. Colligative properties of diluted solutions. Rault’s Law and its corollaries. Cryometry (depression), ebulliometry.
2. Diffusion and osmosis in solutions.
3. Van't Hoff's Law of osmotic pressure. Isotonic coefficient.
4. Isotonic, hypertonic and hypotonic solutions.
5. The role of osmosis and osmotic pressure in biological systems. Oncotic pressure, turgor, hemolysis, plasmolysis.

5.2. Solve the tasks:

Task 1. Calculate the molar mass of the non-electrolyte, if the osmotic pressure of 100 ml of solution, in which 7,5 g of this substance dissolved, is equal to 212,8 kPa at 25 ° C.

Answer: 873,2 g/mol.

Task 2. Human albumin has a molar mass of 65 kg/mol. Calculate the osmotic pressure of the solution of 2 g of protein per 100 ml of water at 27 ° C. Count the solution density is 1,0 g/ml.

Answer: 767 Pa.

Task 3. What is boiling temperature of the solution containing 23,5 g of non-electrolyte with molar mass of 94 g/mol in 250,0 g of water?

Answer: 100,52°C

5.3. Examples of the solved tasks:

Task 1. Calculate the molar mass of the solute, if the vapor pressure above the solution containing 5 g of the substance in 72,0 g of water at 298 K is equal to 3,145 kPa, and the vapor pressure of water at the same temperature is equal to 3,167 kPa.

Solution. The lowering of the vapor pressure above the solution is:

$\Delta p = p_0 - p = 3{,}167 - 3{,}145 = 0{,}022$ кПа.

The molar mass of the solvent (water) is equal to 18 g/mol. We calculate the molar mass of the solute according to the formula:

$$M_{(sub)} = \frac{p_0 \cdot m_{(sub)} \cdot M_{(H_2O)}}{\Delta p \cdot m_{(H_2O)}} = \frac{3{,}167 \cdot 5 \cdot 18}{0{,}022 \cdot 72} = 180\ \text{g/mol.}$$

Answer: 180 g/mol.

Task 2. The solution of 20 g of hemoglobin in 1 liter of water has an osmotic pressure of 0,762 kPa at 298 K. Calculate the molar mass of hemoglobin.

Solution. Molar mass of hemoglobin can be calculated from Van't Hoff's Law of osmotic pressure:

$$M = \frac{m \cdot R \cdot T}{V \cdot \pi} = \frac{20 \cdot 8{,}314 \cdot 298}{1 \cdot 0{,}762} = 65028\ \text{g/mol}$$

Answer: 65028 g/mol.

Task 3. Calculate the osmotic pressure of 5,85% sodium chloride solution at 273 K. The dissociation degree of sodium chloride is 0,96 and the density of the solution is 1.04 g/ml.

Solution. The osmotic pressure of the electrolyte solution is calculated by the formula:

$$\pi = i \cdot C_M \cdot R \cdot T$$

Previously we calculate the molar concentration of sodium chloride solution and the isotonic coefficient:

$$C_M = \frac{\omega \cdot \rho \cdot 10}{M} = \frac{5{,}85 \cdot 1{,}04 \cdot 10}{58{,}5} = 1\,\text{mol/l.}$$

$i = 1 + \alpha\ (n – 1) = 1 + 0{,}96\ (2 – 1) = 1{,}96$.

The osmotic pressure is:

$$\pi = 1{,}96 \cdot 1 \cdot 8{,}314 \cdot 273 = 445\,\text{kPa}$$

Answer: 445 kPa.

5.4. Study tests

1. Colligative properties of the solutions depend on:
A. The number of particles in solution. *
B. The chemical nature of the solute.
C. The chemical nature of the solvent.
2. Osmosis is a spontaneous process of diffusion of :
A. Solvent towards a concentrated solution. *
B. Solvent towards a diluted solution.

C. Dissolved substances towards concentrated solution.

D. Solute towards a diluted solution.

3. Isotonic solutions have the same:

A. pH.

B. Osmotic pressure.*

C. $T_{b.}$

D. $T_{f.}$

4. What do red blood cells placed in 9% NaCl solution undergo:

A Hemolysis.

B. Plasmolysis. *

C. Do not change.

5. What transition is through a semipermeable membrane during red blood cells hemolysis:

A. of the solvent from the cell.

B. of the solvent into the cell. *

C. of solute from the cell.

D. of the solute into the cell.

6. What process occurs with erythrocytes in a hypotonic solution:

A. Hemolysis. *

B. Plasmolysis.

C. No changes.

7. Isotonic coefficient for electrolyte solutions is:

A. $i = l$.

B. $i > l$.*

C. $i < l$.

8. What osmotic pressure do solutions with the same molar concentration of NaCl and glucose have:

A. The same.

B. $\pi_{gluc.} < \pi_{NaCl}$.*

C. $\pi_{gluc.} >> \pi_{NaCl}$.

9. There are solutions with the same molar concentration of 0.01 mol/l: glucose, sodium chloride and calcium chloride. Which solution will be of the greatest osmotic pressure?

A. $CaCl_2$.*

B. Glucose.

C. NaCl.

10. When canning foods in syrup microorganisms become unviable due to:

A Hemolysis.

B. Plasmolysis. *

C. Denaturation.

11. Oncotic pressure is:

A. The osmotic pressure of the blood of cancer patients.

B. The part of the blood osmotic pressure due to the presence of low molecular weight electrolytes in it.

C. The part of the blood osmotic pressure due to plasma proteins. *

D. Integral part of the blood osmotic pressur caused by the relation: hemoglobin / oxy-hemoglobin.

12. Boiling point of solutions compared to the pure solvent is:

A. Higher. *

B. Lower.

C. The same.

D. Depends on the nature of the solute.

13. Freezing point of solutions compared to the pure solvent is:

A. Higher.

B. Lower. *

C. The same.

D. Depends on the nature of the solute.

14. Solutions of glucose and sugar which have the same molar concentration boil:

A. At the same temperature.*

B. $T_{b\ (gluc)} < T_{b\ (sugar)}$.

C. $T_{b\ (gluc)} > T_{b\ (sugar)}$.

15. The solubility of oxygen in the blood plasma will increase when:

A Decreasing of the concentration of salts in the plasma. *

B. Increasing of the salt concentration in the plasma.

C. Increasing of the pressure. *

D. Rising of the temperature.

16. Which of the following solutions with the same concentration has a maximum boiling point?

A Potassium chloride.

B. Urea.

C. Calcium chloride. *

D. Sodium chloride.

17. Which of the following solutions with the same concentration has a minimum boiling point?

A Potassium chloride.

B. Urea. *

C. Calcium chloride.

D. Sodium chloride.

18. Will the osmotic pressure increasing the same with increasing of temperature if diluted solutions of glucose andacetic acid have the same concentration?

A. Yes. it will.

B. Osmotic pressure of glucose <acetic acid. *

C. The osmotic pressure of glucose> acetic acid.

19. There are three solutions with the same concentration - 0.1 mol / l - sodium chloride (A), calcium chloride (B) and glucose (C). Which of the solutions will be hypertonic in relation to each other?

A. A in relation to C.

B. A in relation to B.

C.B in relation to A and C *

D. C in relation to A and B.

20. The semipermeable membrane passes molecules of:

A Solvent. *

B. Solution.

C. Dissolved substance.

21. Choose a natural semipermeable membrane:

A. Polyethylene.

B. The membrane of animal cells. *

C. Paper.

D. Cellophane.

22. Choose a synthetic semipermeable membrane:

A. Polyethylene.

B. The membrane of animal cells.

C. Paper.

D. Cellophane. *

23. Osmotic pressure is:

A. The excessive atmosphere pressure which stops osmosis.

B. The excessive hydrostatic pressure which stops osmosis. *

C. Hydrostatic pressure which directs osmosis in reverse.

24. Osmotic pressure is measured:

A. Stalagmometer.

B. Osmometer. *

C. Viscometer.

25. According to Van't Hoff’s law osmotic pressure of the diluted nonelectrolyte solution equal to the pressure :

A. of a solution in gaseous state and the volume of the solution.

Q. of a solvent in gas phase and the volume of the solution at a given temperature.

C. of the solute in gas phase and the volume of the solution at a given temperature. *

26. The mathematical expression of Van't Hoff’s law for nonelectrolytes is:

A. $\pi = i \cdot C_M \cdot R \cdot T$.

B. $\pi = C_M \cdot R \cdot T$.*

C. $\pi = n \cdot C_M \cdot R \cdot T$.

27. The mathematical expression of Van't Hoff’s law for electrolytes is:

A. $\pi = i \cdot C_M \cdot R \cdot T$.*

B. $\pi = C_M \cdot R \cdot T$.

C. $\pi = n \cdot C_M \cdot R \cdot T$.

28. Isotonic coefficient indicates:

A. As far the osmotic pressure of the electrolyte is more than the osmotic pressure of nonelectrolyte.

Q. How many times the osmotic pressure of the electrolyte is more than nonelectrolyte osmotic pressure at the same molar concentration. *

C. How many times the osmotic pressure of the electrolyte is less than nonelectrolyte osmotic pressure at the same molar concentration.

29. Hypotonic solution is the solution which osmotic pressure is:

A. Higher than the osmotic pressure of another solution.

B. Lower than the osmotic pressure of another solution. *

C. Equal to the osmotic pressure of another solution.

30. Hypertonic solution is the solution which osmotic pressure is:

A. Higher than the osmotic pressure of another solution. *

B. Lower than the osmotic pressure of another solution.

C. Equal to the osmotic pressure of another solution.

31. Isotonic solution is the solution which osmotic pressure is:

A. Higher than the osmotic pressure of another solution.

B. Lower than the osmotic pressure of another solution.

C. Equal to the osmotic pressure of another solution. *

32. Physiological or izotonic solution is:

A. 0,9% solution of Na_2SO_4.

B. 0,9% solution of NaCl.*

C. 0,9% solution of $NaNO_3$.

33. Physiological or izotonic solution is:

A. 0,9% solution of Na_2SO_4.

B. 0,9% solution of NaCl.*

C. 0,9% solution of $NaNO_3$.

34. In medicine it is used hypertonic NaCl solution:

A. 1%.

B. 10%.*

C. 0,9%.

35. The isoosmotic state of human body is supported by:

A. Kidneys and lungs. *

B. Bones and skin.

C. Brain and liver.

36. Hemolysis is:

A. The destruction of cells in hypotonic solution. *

B. Wrinkling of cells in a hypertonic solution.

C. Constant state of the cell.

37. Plasmolysis is:

A. The destruction of cells in hypotonic solution.

V. Wrinkling of cells in a hypertonic solution. *

C. Constant state of the cell.

D. Increasing of cell volume.

38. Oncotic pressure of the blood is:

A. 1atm.

B.0,4 atm.

C. 0,04 atm.*

D. 0,2 atm.

39. Oncotic pressure is due to:

A Proteins. *

B. Electrolytes.

C. Nonelectrolyte.

40. 25% solution of $MgSO_4$ causes laxative effect because that is:

A Hypertonic solution. *

B. Hypotonic solutions.

C. Isotonic solution.

41. Saturated vapor pressure above the solution compared to the vapor pressure of the solvent is:

A. Higher.

B. Lower. *

C. The same.

42. The relative decrease of vapor pressure of the solvent over a solution is:

A. Proportional to mole fraction of a solute. *

B. Proportional to mole fraction of a solvent.

C. Inversely proportional to mole fraction of a solute.

43. Aqueous solutions freeze at temperature:

A. 0 ºC.

B. Higher 0 ºC.

C. Lower 0 ºC.*

44. Depression of the solution freezing point depends on:

A. The concentration of dissolved substances. *

B. Nature of the solute.

C. Nature of the solvent. *

D. Volume of the solution.

45. Elevation of the solution boiling point depends on:

A. The concentration of dissolved substances. *

B. Nature of the solute.

C. Nature of the solvent. *

D. Volume of the solution.

46. Cryoscopic constant depends on:

A. The concentration of dissolved substances.

B. Nature of the solute.

C. Nature of the solvent. *

D. Temperature.

47. Ebullioskopic constant depends on:

A. The concentration of dissolved substances.

B. Nature of the solute.

C. Nature of the solvent. *

D. Temperature.

48. Elevation of the solution boiling point is described by the relationship:

A. $\Delta T_f = k_{f.} \cdot C$.

B. $\Delta T_b = k_b \cdot C$.*

C. $\Delta T_f = k_{f.} : C$.

D. $\Delta T_b = k_b : C$.

49. Depression of the solution freezing point is described by the relationship:

A. $\Delta T_f = k_{f.} \cdot C$. *

B. $\Delta T_b = k_b \cdot C$.

C. $\Delta T_f = k_{f.} : C$.

D. $\Delta T_b = k_b : C$.

50. Isotonic coefficient value depends on:

A. The electrolyte dissociation degree. *

B. The number of ions that electrolyte are dissociated into. *

C. The nature of the solvent.

Literature:

1. Medical Chemistry : textbook / V. A. Kalibabchuk [and ol.] ; ed. by V. A. Kalibabchuk. - K. : Medicine, 2010. - 224 p.

2. Applied Colloid and Surface Chemistry : textbook / R. Pashley, M. Karaman. – 2nd ed.. : John Wiley and Sons, 2004. – 190 p.

6. LABORATORY WORK

6.1. Determination of the molecular weight of the solute by the cryoscopic method

Fill the external glass of cryostat with freezing salt (a mixture of crushed ice and 20-40 g of sodium chloride).

Pour 25 ml of water into the internal glass of cryostat. Put the internal glass into freezing salt and immerse an electronic thermometer in it.

Monitor the thermometer; mixing the solution to avoid overcooling of it. Note the crystal haze of solvent temperature ($t_{cr.0}$). Repeat the experiment three times and write down the average crystallization temperature of the solvent.

Weigh 0,5 g of glucose.

Take the internal glass with crystal solvent out of the freezing salt and melt the crystals of the solvent heating the glass with your hand.

Put glucose into the solvent, stir until fully dissolved. Determine of crystallization temperature of solution ($t_{cr.}$). Repeat the experiment three times and write down the average crystallization temperature of solution.

Calculate freezing point depression of the solution in comparison with a pure solvent:

$\Delta t_{кр.} = t_{кр.} – t_{кр.0} =$

Calculate the molecular mass of the substance according to the formula:

$$M_{(sub)} = \frac{k_f \cdot m_{(sub)} \cdot 1000}{m_{(solv)} \cdot \Delta T_f}$$

k_f – cryoscopic constant of water 1,86°.

Compare the experimentally found value of molecular mass of glucose with the tabular value.

6.2. Preparation of inorganic semi-permeable membrane

Pour 2 ml of $CuSO_4$ solution into the test tube, add crystals of $K_4[Fe(CN)_6]$ (do not stir!). Describe the external effect after 20 minutes. Write down the equation of the reaction. Indicate which compound is a semipermeable membrane.

6.3. Arborescent formations

Pour 3 ml of K_2SiO_3 solution into the test tube and add crystals of $MnC1_2$, $CoSO_4$, $NiCl_2$ (do not stir!). Write down the equations of the reactions. Specify which compounds are semi-permeable membranes, describe the external effect, draw conclusions.

7. TRAINING AND CONTROL MEANS:

- cards to identify initial knowledge and skills;
- control questions;
- tests.

Tutorial # 10

1. THEME: Seminar «Chemistry of biogenic elements. Complexing in biological fluids. Acid-base equilibriua in biological fluids»

2. PURPOSE: To summarize the material, check the students' knowledge on the topic and the ability to apply this knowledge in biomedical research.

3. OBJECTIVES

3.1. Implementing of the final control of knowledge and skills : written card control, a personal interview with each student on written work and computer testing..

3.2. To pay attention to the issues and tasks that cause difficulties for students

3.3. To check laboratory work protocols.

4. PLAN AND ORGANIZATIONAL STRUCTURE OF THE TUTORIAL:

4.1. Organizational time …………………………….............................5 min

4.2. Goal-setting and motivation of studying of the theme (opening address of the teacher)………..……………….......................................5 min

4.3. determination of knowledge level (written card control)……...30 min

4.4. Test control ……………………………………………....................10 min

4.5. Personal interview with each student and checking of laboratory work protocols…..…...60 min

4.6 Concluding remarks of the teacher, guidance to the next lesson………...3 min

5. REFERENCE FOR SELF -STUDY:

5.1. Revise theoretical material on topics 1-9 and tasks solution.

5.2. Revise the questions on studied sections of medicinal chemistry using the recommended literature and lectures.

Control questions

1. Chemistry of biogenic elements.
Complexing in biological fluids.

1. General characteristics of s-, p-, d-elements and their compounds according to their position in the periodic table of Mendeleev. Their chemical properties. Electronic configuration of atoms and ions.

2. Biological and medical significance of s-, p-, d-elements and their important compounds.

3. Coordination compounds. Modern concept of the coordination compounds structure and their classification.

4. Instability and stability constants of coordination compounds. Chelate compounds.

2. Acid – base equilibria in biological fluids

1. Modern ideas of solutions. The role of fluids in vital activity.

2. The solubility of gases in liquids. The solubility of liquids and solids in liquids.

3. Equilibria in electrolytes solutions. Dissociation of strong and weak electrolytes.

4. Dissociation of water. Ionic product of water. pH of body fluids.

5. Calculation of pH of solutions of strong and weak electrolytes.

6. Protolytic acid-base theory. The main types of protolytic reactions.

7. Hydrolysis of salts. The degree and the constant of hydrolysis. The role of hydrolysis in biochemical processes.

8. Buffer systems, protolytic equilibrium in them. Buffer capacity. Examples of the buffer systems: phosphate, hydrocarbonate. Calculation of buffer system pH.

9. Buffer systems of the human body: protein, hemoglobin. The mechanism of action. Significance of buffer systems for living organisms.

10. Fundamentals of titrimetric methods of analysis.

11. Acid-base titration. Titrants. Acid-base indicators, and the principles of their selection.

12. Application of acid-base titration methods in clinical analysis, in sanitary and hygienic studies.

13. Colligative properties of diluted solutions. Rault’s Law and its corollaries. Cryometry (depression), ebulliometry.

14. Diffusion and osmosis. Osmotic pressure. Van't Hoff's Law.

15. Plasmolysis and hemolysis.

16. Colligative properties of diluted solutions. Isotonic coefficient.

17. Isotonic, hypertonic and hypotonic solutions in medicinal practice.

5. The role of osmosis and osmotic pressure in biological systems.

5.3. Study tests

Literature:

1. Medical Chemistry : textbook / V. A. Kalibabchuk [and ol.] ; ed. by V. A. Kalibabchuk. - K. : Medicine, 2010. - 224 p.

2. Levitin Ye.Ya. General and Inorganic Chemistry : textbook for students of higher schools / Ye. Ya. Levitin, I. A. Vedernikova. - Kharkiv : Golden Pages, 2009. - 360 p.

3. Quantitative Chemistry Analysis : textbook / D. C. Harris. – 7th ed.: W.H Freeman and Company, 2007. – 808 p.

4. Inorganic Chemistry : textbook / ed.: D. F. Shriver, P. W. Atkins. - 3th ed. - [s. l.] : Oxford University Press, 2004. - 763 p.

2. Applied Colloid and Surface Chemistry Textbook / R. Pashley, M. Karaman. – 2nd ed.. : John Wiley and Sons, 2004. – 190 p.

5. Petrucci R.H., Harwood W.S., Herring F.G. General Chemistry. Principles and Modern Applications. – 8th ed. – Prentice Hall, 2002. – 1160 p.

6. Basic Concepts of Analytical Chemistry : textbook / S.M. Khopkar. – 2nd ed.. : New Age International., 1998. – 500 p.

CONTENT