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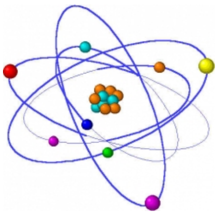
ACID-BASE BALANCE. BUFFER SYSTEMS

Teaching and methodical manual
for foreign students of Zaporozhye State Medical University



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It is recommended by Methodic commission on chemical sciences as a textbook for students of medical faculty (Minutes №3, 27.11.2014)

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CONTENTS

PREFACE

Medicinal Chemistry is one of the most rapidly developing areas within the discipline of Chemistry, both globally and locally. It is the study of the design, biochemical effects, regulatory and ethical aspects of drugs for the treatment of disease.

The aim of this discipline is to produce graduates with an appropriate background in biology and pharmacology, built upon a strong chemistry foundation.

Methodical recommendation of Medicinal Chemistry is designed to equip students with strong grounding in biological and chemical technique which is relevant to the pharmaceutical world.

The discipline gives an in-depth coverage of the chemical techniques required and relates these to the relevant pharmacology, anatomy, biochemistry and molecular biology.

The whole course of Medical chemistry which consists of ten topics is studied by students-physicians during the first year. Lecturer staff of department has prepared an educational and methodical recommendation in which the theoretical material is stated in the concise and available form.

The distribution of material on each of ten topics that are studied is set according to training program, the thematic plan of lectures and practical training.

The material of each topic is stated in such way that performance of practical work and the solution of situational tasks are preceded by theoretical part in which questions of medicine and biological value and also connection with other disciplines (biological chemistry, normal physiology, pathophysiology and others) are included.

Offered laboratory works and situational tasks will give students the chance to understand theoretical material fully and to use this knowledge in practice.

The experience of teaching medical chemistry shows that it is not always possible to coordinate an order of laboratory works realization with sequence of lecture course statement. That is why students usually have to prepare for practical work performance independently before the lesson. Therefore the theoretical part (in which the necessary volume of knowledge for conscious performance of experiment is given) precedes to each section of these Methodical recommendations.

Increasing of level of seminar and laboratory works is reached by use of such forms of occupations which open and consolidate theoretical knowledge, train scientific thinking, develop creative initiative and impart skills of handling devices and chemicals, chemical ware.

The structures, figures and schemes are clear and easy to follow and color is used well, highlighting main points without being distracting.

Chapters are helpfully signposted throughout, informing the reader how topics are related, which is especially important in such a multidisciplinary subject.

Topics are also presented clearly and with a logical progression culminating in the main points, questions and reading sections at the beginning of each chapter.

An assortment of case studies is provided and the authors work through each one in great detail, giving an overall perspective on the science.

Finally, very useful and informative appendices and a glossary are provided together with a comprehensive index that is good enough to rival any search engine!

There are many books that describe medicinal chemistry and its uses, but these methodological recommendations present medicinal chemistry and its related topics in a clear, informative and interesting way that really demonstrates the application and impact of this fundamental subject in society.

INTRODUCTION

Purpose: to study the theory of acids and bases, strong and weak electrolytes, determination methods pH, notions of buffer solutions.

Targets:

- explore the theory of acids and bases, as well as the properties of strong and weak acids;
- Learn how to perform calculations, dissociation constants, the pH of the solution (including buffer), buffer capacity;
- master the methods of determining the pH of solutions, including body fluids;
- Learn how to assess the reliability of the results;
- learn a test material on the topic.

The student should know:

- -theory of acids and bases;
- -parameters describing the acid-base balance;
- properties of strong and weak acids;
- -a method of determining the concentration of acids and bases;
- -characterization of buffer solutions;
- -the use of acid-base balance to living organisms

The student should be able to:

- -to determine the concentration of acids or bases-titrimetric method;
- -use the burette setting, the magnetic stirrer.
- -carry out calculations of pH of strong and weak acids and bases;
- -to prepare buffer solutions;
- -to carry out the calculation of buffer capacity;
- -determine the validity of the measurements and calculations.

CONCISE THEORETICAL MATERIAL

The constancy of the acid-base equilibrium, as it is known, is one of the basic conditions of normal activity of organism. From pH depends on the stability of membranes, enzymes, dissociation of electrolytes, neuromuscular excitability and conductivity, complexation, etc. processes.

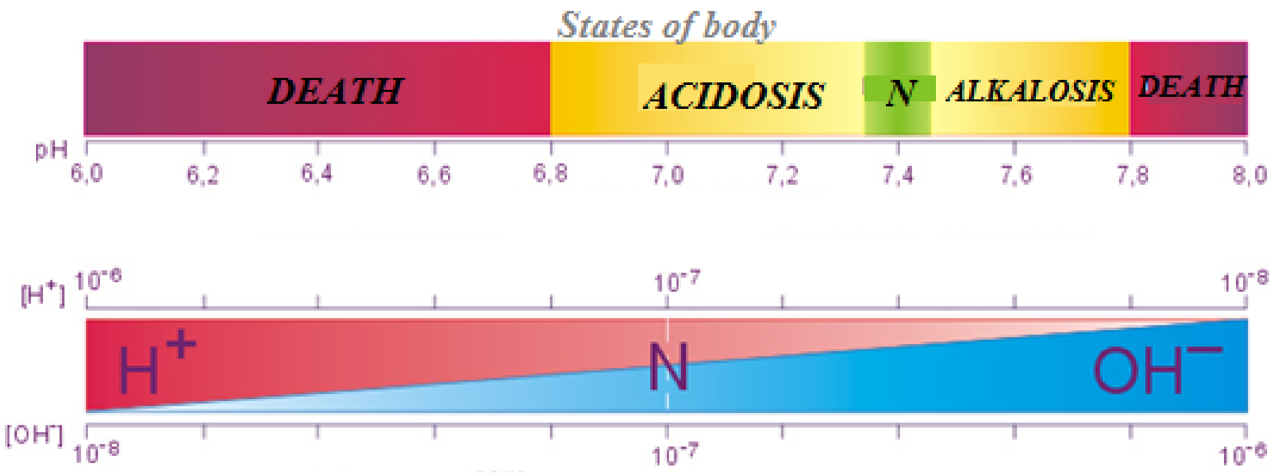
Acid-base system represents the ratio between the concentration of the active masses of hydrogen and hydroxyl ions (major). It is characterized by using a Ph-negative decimal logarithm of the hydrogen ion concentration. Shift the pH ± 0.1 compared to physiological norm leads to respiratory and circulatory disorders, at ± 0.3, loss of consciousness, and in the range of ± 0.4- the death of the body.

During the life of the organism are formed as acidic and alkaline products of metabolism, and first formed nearly 20 times greater than the second. Therefore, mechanisms to ensure the maintenance of the constancy of the acid-basic body systems are aimed at neutralizing and removing, especially acidic products of metabolism.

There are 2 variants of acid-base balance:

Acidosis-a typical pathological process characterized by an absolute or a relative increase in the body burden of acid and alkaline substances. Gas acidosis occurs when increasing the pCO_2 in arterial blood (for example, with alveolar hypoventilation), non-gaseous acidosis-during the accumulation in the body acidic foods woven metabolism (e.g., diabetes mellitus).

Alkalosis is a typical pathological process characterized by an absolute or a relative increase of alkaline in the body or lower acidic substances. Gas alkalosis develops excessive allocation of carbon dioxide (e.g. altitude sickness), non-gas alkalosis, excessive flow of alkaline or acidic substances allocation (for example, uncontrollable vomiting).



Mechanisms of regulation of the acid-basic body systems are very efficient and are able to compensate for the significant changes in Ph. Maintaining the acid-basic body systems is ensured by buffer systems, blood and tissue and the physiological mechanisms of compensation: the lungs, kidneys, liver, blood, bone, and skin.

Therefore, study of acid-base equilibria, and buffer systems is relevant.

1 Theory of solutions of weak electrolytes

In 1887, s. Arrhenius created the theory of electrolytic dissociation (TED).

Basic provisions.

1. dissolution or meltdown Electrolytes decompose into ions.

2. in a solution of electrolytes (hydrated) solvated ions are moving chaotically. When passing through the solution of electric current cations move to the cathode (-), and anions to the anode (+).

3. Dissociation (ionization) is reversible.

Weak electrolytes in solutions dissociate completely.

Electrolytic dissociation degree α shows the proportion of molecules dissolved ions.

Electrolytic dissociation degree is calculated as the ratio of the number of molecules dissolved ions (n), to the total number of solute molecules (N_o):

$$\alpha = \frac{n}{N_0} \cdot 100\ \%$$

The degree of dissociation is expressed in fractions of units or percent.

For example, if α = 30 %, this means that out of every 100 molecules of electrolyte ions break up 30(α = 0, 3)..

For Example,

$$CH_3COOH \underset{V_2}{\overset{V_1}{\rightleftarrows}} CH_3COO^- + H^+$$

According to the law of mass action the speed of direct response

$V_1 = k_1[CH_3COOH]$, and speed feedback $V_2=k_2[CH_3COO^-]\cdot[H^+]$.

In the solution is a balance between the processes of dissociation and Association: V1 = V2, so the equilibrium constant, dissociation of weak electrolytes, is called the constant of dissociation (C k)

$$K = \frac{[CH_3COO^-]\cdot[H^+]}{[CH_3COOH]} = K_d$$

Here in the numerator are the concentration of ions-products of dissociation, and the denominator is the concentration of undissociated molecules.

According to the TED main characteristics of weak electrolyte are the degree of dissociation α and Dissociation constant (ionization) Kd.

The degree of dissociation of electrolyte depends on the nature of the solvent, temperature, and the presence of other ions in solution and the concentration.

Dissociation constant depends only on the nature of the solvent, electrolyte and temperature, but is independent of concentration. Therefore, to describe a weak electrolyte is more convenient to use a constant of dissociation. The more Compact the better the electrolyte decays into ions, the stronger the electrolyte.

Between the Cd and the α relationship. If you indicate the concentration of electrolyte, decaying into two ions, through c, and the degree of dissociation of this solution throughα, the concentration of each ion will be the $C \cdot \alpha$, and concentration of undissociated molecules (1-α). Then the equation becomes: dissociation constants

$$K_d = \frac{C\alpha^2}{1-\alpha}.$$

This equation is the mathematical expression of the breeding of Ostwald.

For fluids in which the dissociation of electrolyte is very small ($\alpha \ll 1$), Ostwald law equation simplifies to:

$K_д \approx C\alpha^2$ or $\alpha \approx \sqrt{K_д / C}$

$\alpha \approx \sqrt{K_д \cdot V}$

where C- is the concentration of the binary electrolyte (mol/l), $V = \frac{1}{C}$ is breeding (l/mol). Ostwald dilution law is formulated as follows: "based on dilution (dilution) weak electrolyte solution by its degree of dissociation is increased.

In practice, for the characteristics of weak electrolyte often use rate of dissociation constants. RC =-lg Kd. The more of the weaker electrolyte.

2 Theory of solutions of strong electrolytes

Designed in 1923, p. Debye and Hukkel.

Basic provisions.

1. strong electrolytes in aqueous solutions fully dissociate, i.e. the degree of dissociation of $\alpha = 1$ or 100%. In solutions of electrolyte ions interact with polar solvent molecules and formation of solvate shell (the shell if hydrated solvent is water). The hydrate shell, increase the size of the ions and therefore decreases the ability of ion transfer electric current, participate in chemical reactions.

2. the ions interact with each other and around each of hydrated ion is "ion atmosphere of hydrated ions of opposite sign, which inhibits the action of each ion.

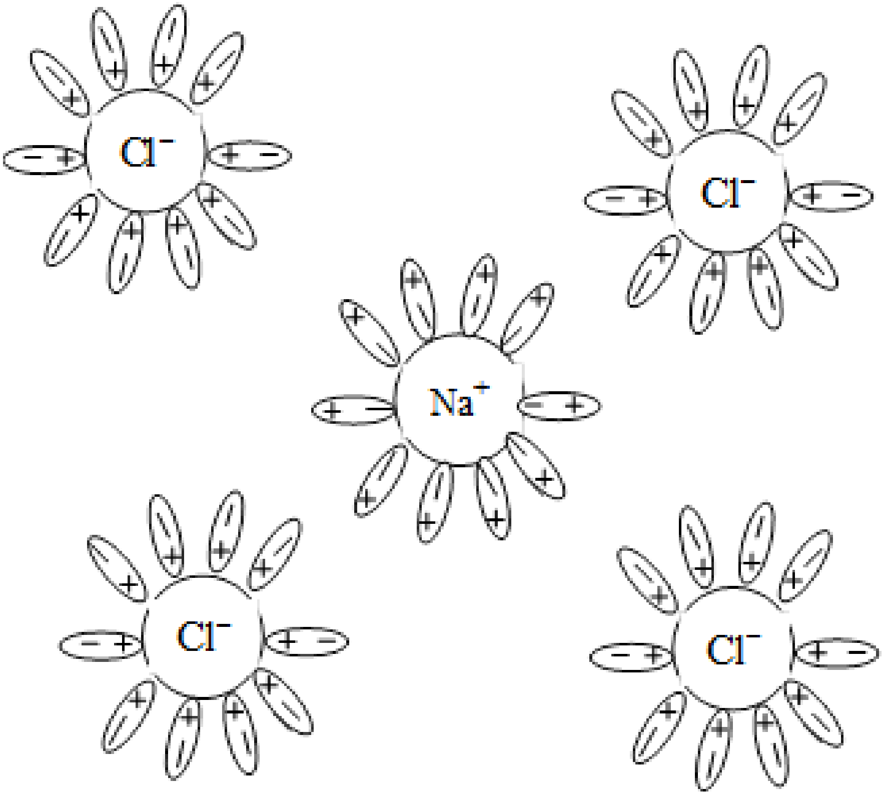


Fig. 1. Formation of hydrates in aqueous solution of NaCl

The emergence of hydrate membrane and ion atmospheres can only speak about the apparent degree of dissociation, because it does not correspond to the actual extent of electrolyte ions. In the case of solutions of strong electrolytes degree is called activity (active concentration).

Electrolyte activity understand conditional effective concentration, in which the electrolyte is in chemical reactions, the collegiate properties of solutions, if you have electric charges. The activity is associated with the true concentration of solute- $\mathbf{a = f_a \cdot C}$, C- the apparatus with which the analytical concentration, mol/l; a-activity of electrolyte, mol/l; $\mathbf{f_a}$-activity rate (dimensionless).

$\alpha \mathbf{f_a} = \alpha$ seeming

FA activity coefficient characterizes deviation of some properties of real strong electrolyte solution with a concentration on the properties of solution at infinite dilution or perfect solution, i.e. in the absence of ion interactions.

For example, if $\mathbf{f_a} = 1$, then the movement of ions in solution of the free ion and a lack of cooperation in this case, a = c; If the $\mathbf{f_a}\alpha$ 1, there are electrostatic interactions between ions.

FA activity coefficient depends on the size of the charge of the ion of nature of solvent, temperature, and concentration of ions.

Other quantitative characteristics of interionic electrostatic interactions is the ionic strength of solution (I):

$$\mathbf{I = {}^{1}/_{2}\ (C_1z_1^2 + C_2 z_2^2 + ...C_i z_i^2)}$$

where: c is the concentration of this ion in MOL/kg, z-charge of each ion.

Between the ionic strength of the solution (I) and the coefficient of activity of FA there is a connection: $\lg f_a = -0,5\ z^2\sqrt{I}$,

where z is the ion charge. The greater the ionic strength of solution and the magnitude of the charge of ions, the less activity coefficient.

Human blood plasma ionic strength close to 0.15 MOL/kg.

3. Classification of acids and bases

The first theory, has divided the substance on acid and Arrhenius was offered bases. In accordance with Arrhenius analysis, acid is a substance in which the dissociation in aqueous solution form hydrogen ions H + and Foundation-substance, hydroxyl ions are formed by the dissociation of OH-.

However, this theory is untenable, since it does not take into account the interaction of particles of solute with solvent. So, for example, could not have been based on the Arrhenius theory, explain why some dissolved salts in the water environment may be acid or alkaline, and during the transition to non-aquatic solvents, many substances are completely changed their acid-base properties.

Resolve data conflicts to many scientists. At the beginning of the 20th century almost simultaneously two theories: theory of acids and bases Brensteda-Lowry, and Lewis theory, developed by Pearson. The most common theory of acids and bases is the theory of Usanovica. However, the range of events faced by analytical chemistry, most satisfactorily explains the protolitic theory Brensteda-Lowry.

This theory successfully explains virtually all processes in solution, and in non-aqueous systems the theory has received universal recognition.

4. Main theories of acids and bases

According Lewis theory acid is a substance which accepts a pair of electrons, which are substances, directing it. The formation of Covalent bonds. In the class of acids are molecules formed by atoms with unfilled eight electronic sheath (BF_3, SO_3), cation-complexing agents (Fe^{3+}, Co^{2+}, etc.), halides with unsaturated bonds ($TiCl_4$), a molecule with polarized double bonds (CO_2, SO_2); in class basis-molecules with free electronic pair (NH_3, H_2O), organic anions, with double and triple bonds, aromatic compounds.

Expanding the classes of substances, the Lewis theory could not give Chief for Chemistry: quantitative evaluation criterion of acids and bases, which could serve as a basis for analytical calculation of acid-base equilibria, for example, the definition of Ph.

Theory Of Pearson. Between acid-Acceptor pairs of electrons to the electron pair donor does not have to be covalent bond and ionic and coordination can occur. According to the theory of acid-base reactions are reactions of chelation.

The Theory Of Usanovica. Acid are substances that cations or anions host (or electrons), Foundation-substances that anions (or electrons) and the host cation. The wording in the classes of acids or bases are Lewis bases and acids, oxidizers and recovery.

5. Protolitic theory of acids and bases

Brønsted-Lowry acid-base theory gives the most general idea of acids and bases. According to this theory, acids are substances or ions which can give the hydrogen ion (proton), Foundation-substances or ions which can accept protons. There are substances which can be and donors, and with Ms time resolution of protons, they are called ampholytes. Acids and bases may be molecular, cationic and anionic (table 4.1, 4.2).

Acid (Brønsted-Lowry theory)

Type of acids	Example	The reaction of acid-base interactions in aqueous solution
Molecular acids	HCl, H_2CO_3, HNO_3, HCOOH, H_2S	$H_2S + H_2O \rightleftarrows HS^- + H_3O^+$ acid 1 base 2 base 1 acid 2
Cationic acids	NH_4^+, $C_2H_5OH_2$, $[Zn(H_2O)_6]^{2+}$, H_3O^+	$NH_4^+ + H_2O \rightleftarrows NH_3 + H_3O^+$ acid 1 b*ase* 2 base 1 acid 2
Anionic acids	HSO_4^-, HCO_3^-, $HC_2O_4^-$	$HCO_3^- + H_2O \rightleftarrows CO_3^{2-} + H_3O^+$ acid 1 base 2 *base* 1 acid 2
Molecular bases	NH_3, CH_3NH_2, $C_6H_5NH_2$	$NH_3 + H_2O \rightleftarrows NH_4^+ + OH^-$ base1 acid2 acid1 base2
Cationic bases	$[Al(H_2O)_5OH]^{2+}$ $[Zn(H_2O)_5OH]^+$	$[Zn(H_2O)_5OH]^+ + H_2O$ $\rightleftarrows [Zn(H_2O)_5]^+ + OH^-$
Anionic bases	Cl^-, NO_2^-, $HCOO^-$, CO_3^{2-}, HCO_3^-	$NO_2^- + H_2O \rightleftarrows HNO_2 + OH^-$ base1 acid2 acid1 base 2

All the reactions acid-base interaction on the theory of Brensted-Lowry is a reversible transfer of a Proton from the acid to the base. As a result of this migration is a pair of new particles, one of which is capable to give proton and the other to take it. Thus, the interaction of acid and base are related acids and bases:

$HA + B \leftrightarrows HB^+ + A^-$

For Example,

$HCOOH + H_2O \leftrightarrows H_3O^+ + HCOO^-$

$H_2O + NH_3 \leftrightarrows NH_4^+ + OH^-$

6 The Ionic product of water and pH

Water is a very weak electrolyte dissociates and is insignificant. Dissociation of H_2O is a protolitic reaction:

$H_2O + H_2O \rightleftarrows H_3O^+$

or simply:

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

Dissociation constant of water at 298K, some method of electric conductivity is equal to:

$$K_d(H_2O) = \frac{[H^+][OH^-]}{[H_2O]} = 1{,}8 \cdot 10^{-16} mol/L$$

Water is present in great excess, its concentration can be $[H_2O]$ standing at 55.6 moles/l (1000 g: 18 g/mol = 55.6 moles). Combine the two constants K_d (H_2O) and [H2O] into one, we get:

$$K_{H_2O} = [H^+][OH^-] = 1{,}8 \cdot 10^{-16} \cdot 55{,}6 = 10^{-14}$$

The amount is called ion product of water. This value is constant at a given temperature. With increasing temperature Ionic product of water increases.

If [H +] = [OH-] = 10^{-7} mol/l, then it's a neutral environment. If [H +] [OH-]>i.e. [H +] > 10^{-7}, the solution is acidic environment. If $[[H^+]<[OH^-]$, т.е. $[H^+] <10^{-7}$, then the solution is alkaline.

7. pH.

In practice, the use of hydrogen ion concentration [H +] to describe credy is not very convenient. Therefore, for this purpose, use a negative base-10 logarithm of the activity of the (concentration) hydrogen ions, called hydrogen index Ph:

$pH = – \lg a(H^+)$ or $pH = –\lg[H^+]$

Similarly hydroxyl is pOH =-lg a (OH-) or pOH =-lg [OH-]

For example, if the [H +] = 10-2 mol/l (acidic) pH = 2, and when [H +] = 10-9 mol/l (alkaline environment), pH = 9. In a neutral environment [H +] = 10-7 mol/l and pH = 7. From these examples, it follows that:

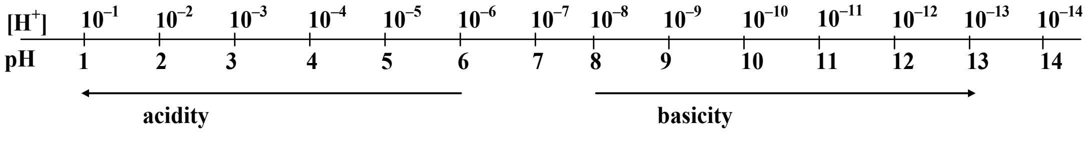
If pH = 7, then it's a neutral environment;

If the pH is 7 < acidic environment;

If the pH is 7 > alkaline environment.

Taking the expression [H +] [OH-] = 10-14 and following mathematical conversions, we get: pH + Ron = 14.

Scale [h +] and Ph



Calculate the pH of solutions of strong and weak electrolytes

1. in the case of solutions of strong acids and bases:

a) $H_2SO_4 \rightarrow 2H^+ + SO_4^{2-}$, $[H^+] = C_H$ (acid)f_a, where SN is the molar concentration equivalent, f_a is activity coefficient;

for diluted solutions $f_a \approx 1$.

$pH = -\lg[H^+] = -\lg C_H$ (acid) f_a.

б) $Ba(OH)_2 \rightarrow Ba^{2+} + 2OH^-$, $[OH^-] = C_H$ (base) f_a

$pOH = -\lg[OH^-] = -\lg C_H$ (base) f_a

$pH = 14 - pOH$

2. in the case of solutions of weak acids and bases:

а) $CH_3COOH \rightleftarrows CH_3COO^- + H^+$, $[H^+] = C_H$ (acid)$\cdot\alpha$,

where C_H (acid)$\cdot\alpha$ – number of dissociated molecules weak acid

$pH = -\lg[H^+] = -\lg C_H$ (acid)$\cdot\alpha$.

б) $NH_4OH \rightleftarrows NH_4^+ + OH^-$, $[OH^-] = C_H$ (base)$\cdot\alpha$,

where C_H (base)$\cdot\alpha$ – number of dissociated molecules weak base

$pOH = -\lg[OH^-] = -\lg C_H$ (base.) $\cdot\alpha$

$pH = 14 - pOH$.

The role of hydrogen ions in biological processes

Body fluids contain strong and weak acids: HCl, H_2CO_3, pyruvic, lactic acid and others.

There are three types of acidity in biological fluids:

1. total acidity is the total concentration of strong and weak acids. The total acidity is usually defined by acid-base titration.

2. active acidity (concentration) is equal to the free hydrogen ions in a solution. Active acidity measure is the value of the pH of the solution.

3. potential acidity is equal to the concentration of the weak acid molecules and undissociated is calculated from the difference of the values of total and active kislotnostej.

Any biological fluid normally has a definite value of active acidity, pH.

Table

Interval of pH values of critical biological fluids

Gastric juice	0,9 – 2,0
Urine	5,0 – 8,0
Saliva	5,6 – 7,9
Blood plasma	7,36 – 7,44
Lacrimal fluid	7,6 – 7,8
Pancreatic juice	8,6 – 9,0

A number of pathological processes occurring in the body, may cause changes in the pH of body fluids. Therefore, determination of the pH of the body fluids (gastric juice, urine, etc.) are used in the diagnosis and monitoring the effectiveness of therapy.Determination of the reaction of environment and knowledge of the h + ion concentration in biological liquids frequently is necessary in biochemical studies (study of enzyme activity).

Determination of pH

Colorimetric determination of pH based on changing the color, acid-base indicator color depending on the Ph.

Indicators can be one-color, with color only in alkaline medium, and in acidic medium is colorless (nitro phenols, phenolphthalein), and two that have different coloration in acidic and alkaline media (methyl orange, phenol red, etc.).

Each indicator is an indicator titration and interval (a) transition.

The titration rate RT is a pH value within an interval of transition in coloring, observed the most drastic color change indicator.

The transition interval coloring the indicator is called interval values pH (pH d) within which occurs a distinct eye color change indicator. Border transition interval is approximately equal to T indicator ± 1. In determining the pH of the solution can be used only one indicator, the coloring of the transition interval of pH of the solution.

Table

Indicator	$pT_{ind.}$	The transition interval coloring		
		coloring I	ΔpH	coloring II
Methyl Orange	3,7	red	3,1–4,4	yellow
Methyl Red	5,7	red	4,2–6,3	yellow
Litmus	7,0	red	5,0–8,0	blue
Phenolphthalein	9,2	colorless	8,2–10,0	pink

Usually, first determine the approximate pH value using universal indicator. Universal indicator is a mixture of several indicators with different, but adjacent to each other at intervals color transition, covering the range of pH from 1 to 14, with a mixture of indicators has a certain color to different pH values.

Universal indicator paper is a filter paper impregnated with universal indicator. It shall be accompanied by colour scale pH values for each color. Accuracy is not more than 0.5 units. Ph. On the approximate value of the pH indicator picked up for a more precise definition.

Potentiometric (ionometric) definition of pH is based on measuring electromotive force (EMF) of a galvanic circuit composed of the indicator half-cell (electrode), whose potential depends on the pH of the medium (glass, hydrogen, quinhydrone) and electrode (silver chloride, calomel), a permanent capacity.

Measuring scale ionomer (pH-meter) graded as in mV, and u. Ph. Accuracy up to 0.01 units. Ph. Can be used to determine the pH of the muddy and colored liquids.

8. Buffer solutions

Buffer solutions are solutions of pH, which little changes when you add a small amount of strong acids or bases, and dilution.

Proton theory in terms of a simple buffer solution consists of a weak acid and its conjugate base or a weak base and its conjugate acid. In this case, the buffer solution is characterized by the presence of the acid-base balance:

$$HA \rightleftarrows H^+ + A^-$$

$$B + H^+ \rightleftarrows BH^+$$

Formed by conjugate acid-base pair on/a and/BH + is called the buffer systems.

Classification of buffer systems

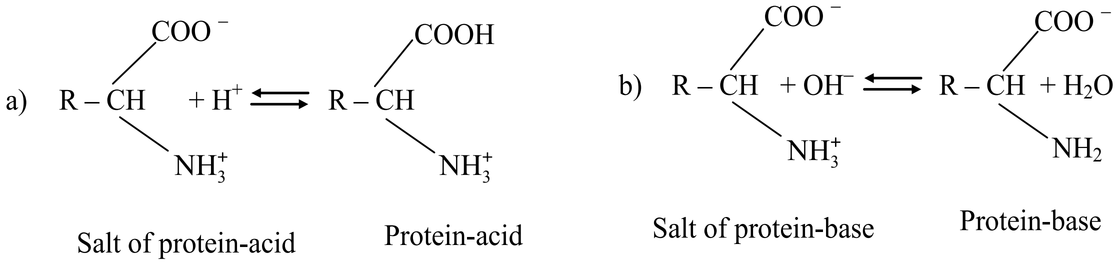
1. acid. Consist of weak acids and salts of the acid. For example, acetate buffer ($CH_3COOH + CH_3COONa$), buffer ($H_2CO_3 + NaHCO_3$).

2. main. Consist of a weak base and its salts. For example, ammonium buffer system ($NH_3 = H_2O + NH_4Cl$).

3. Salt. Consist of acidic and high salt or two of acid salts. For example, the carbonate buffer ($NaHCO_3 + Na_2CO_3$), phosphate buffer system ($KH_2PO_4 + K_2HPO_4$).

4. the amino acid and protein. If the total charge of the molecule amino acid or protein is equal to zero (isoelectric), the solutions of these compounds

are not a buffer. Their buffer effect begins to occur when they add a quantity of acid or alkali. Then part of the protein (amino acids) transforms from a State in the form of "isoelectric protein-acid" or in the form of "protein-base". A mixture of two forms of protein: a) weak "protein-acid" + this weak acid salt; b) weak "protein-base" + salt of this weak Foundation:



where R is macromolecular protein balance.

Calculate pH of buffer systems

To calculate the pH of a buffer solution in acetate buffer, consider the processes occurring in it and their impact on each other.

Sodium acetate dissociates almost completely the ions, the acetate ion is hydrolysis as a weak acid ion:

$$CH_3COONa \rightarrow Na^+ + CH_3COO^-$$

$$CH_3COO^- + HOH \rightleftarrows CH_3COOH + OH^-$$

Acetic acid, also member of the buffer dissociates only slightly:

$$CH_3COOH \rightleftarrows CH_3COO^- + H^+$$

Low dissociation CH_3COOH is more suppressed in the presence of CH_3COONa, so the concentration of acetic acid undissociated are almost equal to its initial concentration:

$$[CH_3COOH] = [acid]$$

On the other hand, also depressed by the presence of salt hydrolysis in acid solution. Therefore, we can assume that the concentration of acetate ions in the buffer mixture is almost equal to the initial salt concentration without considering the concentration of acetate ions from the dissociation of the acid:

$$[CH_3COO^-] = [salt]$$

According to the law of mass action, a balance between the products of dissociation of acetic acid and not dissociating molecules obeys the equation:

$$К_д = \frac{[H^+][CH_3COO^-]}{[CH_3COOH]}.$$

Substituting the total concentration of acid and salt dissociation constants in the equation, we get:

$$[H^+] = К_д \frac{[acid]}{[salt]},$$

here for acid buffer systems: $\mathbf{pH = pK_{(acid)} + lg \frac{[salt]}{[acid]}}$**.**

This equation is called the equation of Gendersona-Gasselbah.

After a similar output to the main buffer systems:

$$\mathbf{pOH = pK_{(base)} + lg \frac{[salt]}{[base]}, \quad pH = 14 - pK_{(base)} - lg \frac{[salt]}{[base]}}$$

where pK (acid) , pK (base) - negative logarithm of electrolytic dissociation constant of a weak acid , weak base , [salt] - salt concentration , [acid] - acid concentration , [base] - concentration of base .

From these equations it is evident that the pH of the acid (base) of the buffer depends on the nature of the weak electrolyte (pKa (acid) , pK (base)) and the ratio of the concentrations of salts and acid (base).

It should be noted that the buffer system effective to maintain the pH in the range : pKa (acid) acid systems 1 to 14 - (pK (base) 1) for the core system .

The mechanism of action of buffer systems:

1. Dilution. Dilution water decreases the concentration of both components in the buffer system in the same way, so the value of their relationship will not change. RK (acid) and RC (grounds) are constant at a given temperature and dilution. Indeed, the simultaneous decrease of the concentrations of acid and salt in acetate buffer system from 0 m to 0, 001M with dilution water changes the pH of the buffer solution with a 4.63 to 4.73 (a negligible change in pH when the dilution buffer solution in a 100 times due to some changes in the activity of salt). Therefore, the dilution of ultimately little changes the pH of buffer systems.

2. addition of acids and bases. When you add a small amount of strong acids or bases the pH of buffer systems varies slightly. For example, consider an acetate buffer:

CH_3COOH / CH_3COO^-

acid component -component-

a) when it is added to the acetic buffer a small amount of HCl, H^+ ions interact with the main component of the buffer solution:

$$H^+ + CH_3COO^- \rightleftarrows CH_3COOH.$$

The degree of dissociation of Ch3cooh is small and the concentration [H +] practically does not change. the pH of buffer solution will be reduced, but only slightly.

Thus, if the acetic buffer add x mol/l HCl, the equation to calculate the pH of a buffer system becomes:

$$\mathbf{pH = pK(acid) + lg}\ \frac{[\text{salt}] - X}{[\text{acid}] + X}$$

b) when you add a small amount of NaOH, neutralized acid component ions buffer solution:

$$OH^{-} + CH3COOH \rightleftarrows CH3COO- + H2O.$$

As a result, added strong basis is replaced with an equivalent amount of weak conjugate base (CH3COO–), which to a lesser extent, affects the reaction the pH of buffer solution is increased, but only slightly.

Thus, if the acetate buffer to add the mol/l NaOH, the pH of the buffer equation to calculate the system becomes:

$$\mathbf{pH = pK(acid) + lg}\ \frac{[\text{salt}] + У}{[\text{acid}] - У}$$

Buffer tank

The ability of buffer solution pH value is maintained when you add strong acids or alkalis approximately constant characterizes the buffer tank.

Buffer tank (a) is the number of moles of a strong acid or alkali to be added to 1 l of buffer solution to move its pH per unit.

Buffer tank system shall be determined in relation to the added acid (Vacid.) or base (alkali) (Vbas) and is calculated using the formulae:

$$B_{acid.} = \frac{C_H(HA) \cdot V(HA)}{|pH - pH_0| \cdot V(б.р.)}, \quad B_{bas.} = \frac{C_H(B) \cdot V(B)}{|pH - pH_0| \cdot V(б.р.)},$$

Buffer tank with respect to acid (V_{acid}) is determined by the concentration (number of equivalents) component with key properties; buffer tank in relation to the base (V_{base}) is determined by the concentration (number of equivalents) with acidic component in buffer solution.

Maximum buffer capacity while adding strong acids and bases is achieved with the component ratio buffer solution equal to 1 When pH = pK, $V_{base\cdot}$ = max. (fig. 4.1).

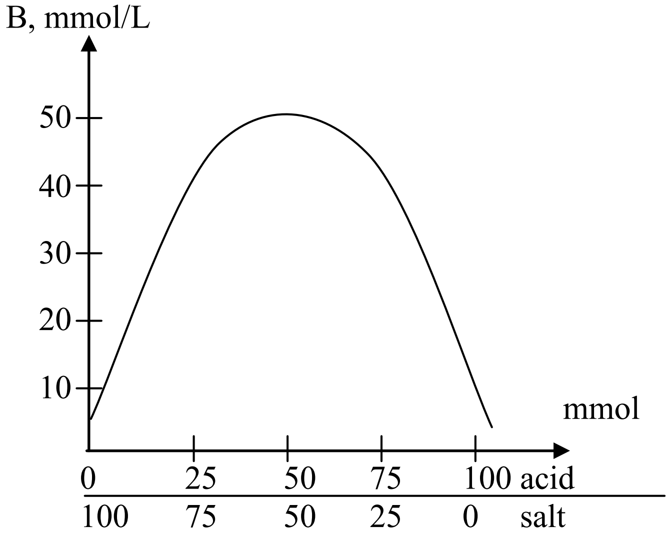


Fig. 1. changing the buffer capacity depending on the size of the [salt]/[acid].

Therefore, the use of any buffer mixture is limited to a certain area of pH (buffer area), namely:

$pH = pK_{(acid)} \pm 1$ for acidic systems, or

$pH = 14 - (pK_{(base)} \pm 1)$ for basic systems.

Buffer tank depends not only on the relationship of the concentrations of buffer solution, but also on the total concentration of the buffer mixture.

For example, given two buffer solution, one of which contains 100 and the other is 10 mmol of acetic acid and sodium acetate. Compare how to change their Ph when added to 1 l each 5 mmol solution of hydrochloric acid.

The added acid will react with sodium acetate, and this relationship is the first solution to 0.9, and 0.33. As a result, the first solution of salt/acid ratio and therefore the pH changed less. From here, you can see that the first buffer solution has a greater buffer capacity.

Thus, the buffer tank is basically depends on the ratio of the concentrations of the components and their absolute concentrations, and hence from dilution.

Buffer systems of the body

The principal source of hydrogen ions in the body is carbon dioxide produced by the metabolism (metabolism) and it takes about ≈ 15000 mmol/day.

Hydration of carbon dioxide leads to the formation of carbonic acid:

$$CO_2 + H_2O \rightleftarrows H_2CO_3 \rightleftarrows HCO_3^- + H^+$$

To a lesser extent, the number of ions h + (30-80 mmol/day) due to the arrival of the body, as well as the formation of acids such as sulfuric acid (as a result of the exchange of sulphur containing amino acids), phosphorous (phosphate compounds metabolism), organic acids formed by incomplete oxidation of lipids and carbohydrates.

The body is released from the acids due to the processes of respiration and urinary output, i.e. in the body there is a correlation between metabolic processes and gas exchange. In the evaluation of acid-base state of the body, it is important to not only the definition of pH values, but also the characteristics of the mechanisms that regulate this parameter.

If the body was not immediate buffer mechanisms and respiratory (respiratory) compensation, then even normal daily load of acids have been accompanied by significant fluctuations in Ph.

Consistency ROP liquid environments organism is supported in living organisms buffer systems. Chief among these are hydro, hemoglobin, phosphate and protein. All buffer systems in the body related substance that provides biological fluids constant pH value. In humans and animals are buffer systems in the blood (plasma and erythrocytes), intercellular spaces in the cells and other tissues.

Buffer systems of blood are plasma systems buffer and buffer systems of erythrocytes. Buffer systems, hydro-plasma protein and phosphate, the latter role is insignificant. They account for≈ 44% of it blood buffer capacity. Erythrocyte- hemoglobin buffer systems, hydro, organic phosphate (phosphate). They accounted for ≈56% of it blood buffer capacity.

Table

Blood buffers separate buffer tank

The name of the buffer system	% relative buffer capacity
Hemoglobin and the oxyhemoglobin	35%
Organic phosphate	3%
Inorganic phosphates	2%
Plasma proteins	7%
Plasmatic hydrogen carbonate	35%
Hydrogen cells	18%

The most important buffer is a buffer system with about 55% of the buffering capacity of the blood. Moreover, it occupies a central position among all other important mechanisms of homeostasis of hydrogen ions, including the hemoglobin buffer system (which provides 35% buffer capacity of blood), as well as hydrogen ion secretion in the kidney. Directly measure very low concentrations of carbonic acid in the blood. At equilibrium with the dissolved CO2 in the equation instead of [H_2CO_3] enter [Co2]. Gendersona-Gasselbah equation takes the following form:

$$pH = 6{,}1 + \lg\frac{[HCO_3^-]}{[CO_2]}, \text{ where } pK = -\lg \mathbf{K}_{d_1}(H_2CO_3) = 6{,}1$$

Practically in blood measure the partial pressure of carbon dioxide CO2. The concentration of dissolved in plasma Co2 is calculated P_{CO_2} by multiplying by a constant value of solubility of CO2. If P_{CO_2} expressed kPa, the constant equal to 0.23, if in mm. the Republic of Tajikistan. art. -0.03.

Therefore, if P_{CO_2} expressed in kPa, the equation takes the following form:

$$pH = 6{,}1 + \lg \frac{[HCO_3^-]}{[\mathbf{P}_{CO_2} \cdot 0{,}23]}$$

The partial pressure of Co2 in the blood plasma of normal is ~ 5.3 kPa (40 mmHg), which corresponds to the CO2 concentration of ~ 1.2 mmol/l. Maintain consistency at this level depends on the balance between the release of Co2 as a result of the reactions of metabolism and its losses from the body via the alveoli.

In the cells of the renal tubules and in the erythrocytes of the trapped light, Co2 used for the formation of bicarbonate ions. The kidneys play a key role in maintaining a constant concentration of bicarbonate in the circulating blood. Red blood cells carry the fine regulation of bicarbonate in the blood plasma.

When blood plasma 5.3 kPa these two fabrics are normally continuous extracellular concentration of hydrogen-ions 24 mmol/l Ratio in the extracellular fluid [NS]/[Co2] (both measurements are in mmol/l) is 20: 1. Gendersona-Gasselbah equation for the ratio corresponds to the largest plasma pH equal to 7.4:

$$pH = 6{,}1 + \lg\frac{24}{1{,}2} = 6{,}1 + \lg 20 = 6{,}1 + 1{,}3 = 7{,}4$$

Thus, the strong response of arterial blood plasma in healthy people corresponds to pH = 7.40.

Reduction ratio $[HCO_3^-]$ / $[CO_2]$ < 20 acidosis. Acidosis can be caused by increased formation of hydrogen ions h + or enhanced allocation of hydrocarbons.

Raising the ratio [$[HCO_3^-]$ / $[CO_2]$> 20 leads to alkalosis.

As in plasma central role in binding of the h + ion is hydrogen-anion concentrations in the plasma makes up the alkalinity of the blood.

Phosphate buffer system is contained in both the blood and cellular fluids of other tissues, particularly in the kidneys.

In cells, it is represented by KH_2PO_4 and K_2HPO_4. In blood plasma, and intercellular space of NaH_2PO_4и Na_2HPO_4. The main role in the mechanism of action of this system is the ion:

$$H_2PO_4^- \rightleftarrows H^+ + HPO_4^{2-}$$

Increase in the concentration of h + resulted in a shift to the left, i.e., the

reaction to the formation of acid: $HPO_4^{2-} + H^+ \rightleftarrows H_2PO_4^-$

Protein buffer systems are ampholytes , because it is composed of α-aminoacids containing groups with acidic properties (-COOH and-) and basic properties (-COO-and-NH_2). The mechanism of action of this buffer system can be represented as follows:

acid buffer system

а) $H_3N^+ - R - COOH + OH^- \rightleftarrows H_3N^+ - R - COO^- + H_2O$

protein-acid

б) $H_3N^+ - R - COO^- + H^+ \rightleftarrows H_3N^+ - R - COOH$

protein-acid salt

the base buffer system

а) $H_2N - R - COO^- + H^+ \rightleftarrows H_3N^+ - R - COO^-$

protein – basis

б) $H_3N^+ - R - COO^- + OH^- \rightleftarrows H_2N - R - COO^- + H_2O$

protein - basis

(conjugate acid)

where R is the remainder of macromolecular proteins.

The role of plasma proteins in thyroid of hydrogen ions is quite small.

Hemoglobin buffer system is only in erythrocytes. Its mechanism of action is related to the accession and the impact of oxygen. The hemoglobin (Hb) is oxidised and $\text{HHв}O_2$ restored Hhb form.

$\text{HHв} + O_2 \rightleftarrows \text{HHв}O_2 \rightleftarrows H^+ + \text{Hв}O_2^-$

$\text{Hhb} \rightleftarrows + \text{Hв}^-$

The mechanism of action is based on the reactions:

$HbO_2^- + H^+ \rightleftarrows HHbO_2 \rightarrow HHb + O_2$

$HHbO_2 + OH^- \rightleftarrows HbO_2^- + H_2O$

$HHb + OH^- \rightleftarrows Hb^- + H_2O$

$Hb^- + H^+ \rightarrow HHb$

The above schematic reaction shows that the addition of a strong acid or strong alkali calls the defensive response of the buffer system to maintain constant pH value that is bound to add to the h + and Oh-, and education little dissociated of electrolytes.

Hemoglobin buffer system in the body functions efficiently only in combination with mineralized system. Because the aerobic metabolism in erythrocytes did not occur, they generate relatively little CO2. From blood plasma, in accordance with the concentration gradient of CO2 diffuses into the blood, where the enzyme carbonic anhydrase catalyzes its interaction with water, leading to the formation of carbonic acid.

As the dissociation of H^+ ions am H_2CO_3 mainly interact with hemoglobin as a buffer system. Increases concentration in erythrocytes of bicarbonate ions that diffuse into the extracellular fluid, in accordance with the concentration gradient.

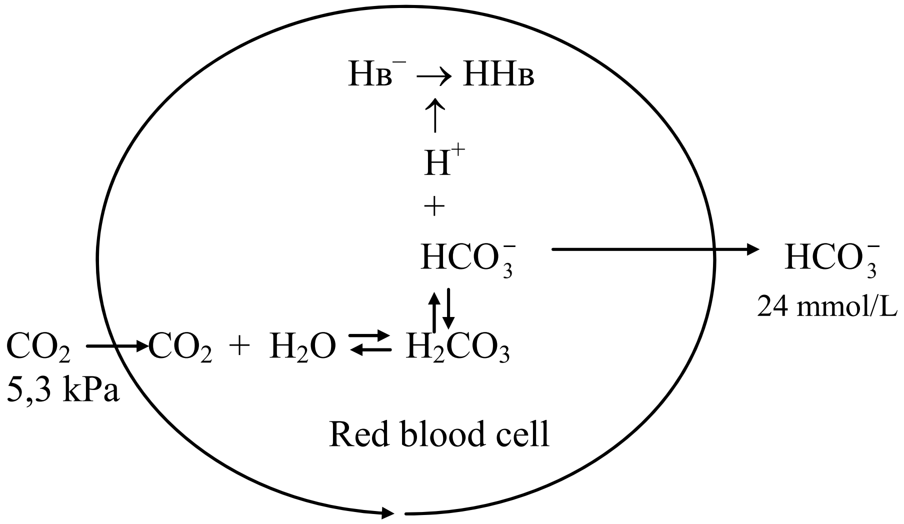


Fig. 2. Formation of hydrocarbons in erythrocytes

Thus, most of the Co2 into the blood appears to run in the plasma is not in the form of acid, as well as hydrogen ions. It is present in erythrocytes of the who hemoglobin buffer system and carbonic anhydrase allows erythrocytes do the trick.

All buffer systems of the body are interrelated.

Received from outside or formed during the metabolism of H^+ ions are weakly dissociated compounds in body fluids is therefore considerably less free ion H^+ than is there.

However, in diseases of the respiratory system, circulatory system, liver, kidneys, in case of poisoning, starvation, burn disease, uncontrollable vomiting, debilitating diarrhea, etc. may be a violation of acid-base balance. It may be accompanied by an increase in the concentration of hydrogen ions in body fluids and such a condition called acidosis, or decrease in the concentration of hydrogen ions, and such a condition called alkalosis.

Protection circuit against the acidosis

Acidosis occurs more commonly, as in the body of many substances are formed from the breakdown of acids. As noted above, blood and kidney buffer systems stabilize blood pH and thus the internal environment of the body.

Introduced ion of hydrogen in hydrogen-ion neutralizes acidosis buffer blood systems to form weak carbonic acid waders migrations:

$H^+ + HCO_3^- \rightleftarrows H_2CO_3$

The plethora of recent splits to H_2O and CO_2

$H_2CO_3 \rightarrow H_2O + CO_2$

Remove CO_2 through the light means the substitution of hydrogen ions and water molecules by hydrogen ions. It drains a buffer system, the kidney, however, is the new number of hydrogen ions.

QUESTIONS FOR SELF-TRAINING

1. Theories of acids and bases;
2. dissociation constants, the pH of the solution (including buffer), buffer capacity;
3. Parameters that characterize the acid-base equilibrium;
4. Properties of strong and weak acids;
5. Water dissociation
6. Dissociation constants
7. The method for determining the concentration of acids and bases;
8. methods of determining the pH of solutions, including body fluids;
9. The use of the acid-base equilibrium applying to living organisms

TASKS

1. Calculate the pH and Ron sulfuric acid solution, if in 1 l of solution containing 0.049 g of H_2SO_4 (sulfuric acid equivalence factor is equal to ½) *. Answer: pH = 3, pOH = 11.

2. Calculate the pH of 0.001 m solution acetic acid dissociation degree, if it is equal to 0.134. Answer: pH = 3.87

3. How change the pH by adding 30 ml of a 0.2 M solution of sodium hydroxide to 300 ml water? Answer: increase by 5.26 pH units

4. How many times the hydrogen ion concentration in the blood than in the cerebrospinal fluid ¬ STI? (pH (blood) = 7.36 pH (cerebrospinal fluid), A = 7.53: about 1.5 times.

5. Define the pH of the buffer solution containing 1 liter of 18.4 g of formic acid and 68 g of sodium formate , if pK (HCOOH) = 3,75. How to change the pH of the solution when diluted 50 times ? * Answer: pH = 4.15 . Practically unchanged.

6. Define acetate buffer pH of the mixture prepared from 100 mL of 0.1 M solution and 200 ml of CH3COOH 0.2M CH3COONa, if kd (CH3COOH) = 1,75 □ 10-5. How to change the pH of the buffer solution by the addition thereto of 30 ml of 0.2M NaOH? Answer: pH = 5.36 , pH increased by 0.46 units.

7. Define pH formic acid, half-neutralized with alkali (pK (HCOOH) = 3,75). Answer: pH = 3.75.

8. To 100 mL of blood for pH changes from 7.36 to 7.00 is necessary to add 36 ml of 0.05M HCl. Calculate the buffering capacity of blood acid (mol/L pH).). Answer: 0.05 (mol/L ·pH).

THE STANDARD ANSWERS

Example # 1

Calculate the [H +] and pH 0.003 m HCl at 298K.

It Is Given:	Answer
C_M (HCl) = 0,003 mol/L pH - ? $[H^+]$ - ?	Hydrochloric acid is a strong electrolyte which in aqueous solution are almost completely dissociates into ions. Because the concentration of HCl is low, then activity coefficient (fa) is approximately 1 and activity (s) is equal to the concentration. Then, the activity of hydrogen ions ($a(H^+)$ or $[H^+]$) p: $[H^+] = C_{\text{н}}$ (HCl)

1. Define $[H^+]$: $[H^+] = C_M(HCl) = 0,003$

$(C_M (HCl) = C_{\text{н}} (HCl))$

2. Define pH: $pH = - \lg[H^+] = - \lg 0,003 = 2,52$.

Answer: $[H^+] = 0,003$ mol/L; pH = 2,52.

Example # 2

Calculate the pH of 0.01 m solution of NH4OH at 298K if the degree of dissociation of ammonium hydroxide is 0.042.

It Is Given:	Answer:
C_M (NH_4OH) = 0,01 mol/L $\alpha = 0,042$ pH - ?	$NH_4OH \rightleftarrows NH_4^+ + OH^-$

1In the diluted solution weak electrolyte activity of hydroxide ions is equal to:

$[OH^-] = C_н \cdot \alpha = 0{,}01 \cdot 0{,}042 = 4{,}2 \cdot 10^{-4} c C_н (NH_4OH) = C_M (NH_4OH)$

2. $pOH = -\lg [OH^-] = -\lg 4{,}2 \cdot 10^{-4} = 3{,}38$

3. $pH = 14 - pOH = 14 - 3{,}38 = 10{,}62$.

Response: $pH = 10{,}62$.

Example #3

Calculate the degree of dissociation of lactic acid, [H +] and pH of 0.1 m solution of lactic acid at a temperature of 298K, if lactic acid dissociation constant (KD) = 1.38· 10-4.

It Is Given:

C_M(acid)= 0,1 mol/L

K_d (acid) = $1{,}38 \cdot 10^{-4}$.

α - ? $[H^+]$ - ?

pH - ?

Answer

Lactic acid is a weak acid and Mono-basic dissociation by schema:

$$CH_3CH(OH)COOH \rightleftarrows CH_3CH(OH)COO^- + H^+$$

1. determine the degree of dissociation:

For diluted solutions of weak binary electrolyte formula applies:

$\alpha = \sqrt{\frac{K_д}{C}}$ (Simplified expression Ostwald dilution law).

Then, $\alpha = \sqrt{\frac{1{,}38 \cdot 10^{-4}}{0{,}1}} = \sqrt{13{,}8 \cdot 10^{-4}} = 3{,}7 \cdot 10^{-2} = 0{,}037$

2. Define $[H^+]$: $[H^+] = C_н \cdot \alpha = 0{,}1 \cdot 0{,}031 = 0{,}0037$ mol/L

$C_M(CH_3CH(OH)COOH) = C_н (CH_3CH(OH)COOH)$

3. Define pH:pH = – lg $[H^+]$ = –lg 0,0037 = 2, 43

Response: α = 0,037, $[H^+]$ = 0,0037 mol/L, pH = 2,43

Example # 4

Calculate the degree of dissociation and the concentration of acetic acid, the concentration of hydrogen ions in acetic acid, the pH is equal to 3.87. Dissociation constant of acetic acid at a temperature of 298K is equal to1.75 $\cdot 10^{-5}$.

It Is Given:

pH = 3,87

$К_{Д} = 1,75 \cdot 10^{-5}$

$[H^+]$ - ? C_M - ?

α - ?

Answer:

1.Define $[H^+]$:$[H^+] = 10^{-pH} = 10^{-3,87} = 0,000135$

2. Define C_MAcetic acid dissociates according to the scheme:

$CH_3COOH \rightleftarrows CH_3COO^- + H^+$.

Dissociation constant expressed by the ratio: $К_{Д} = \frac{[H^+] \cdot [CH_3COO^-]}{[CH_3COOH]}$

$[H^+] = [CH_3COO^-]$, a $[CH_3COOH]$ in a dilute solution of a weak binary electrolyte can be taken as C_M. When: $К_{Д} = \frac{[H^+]^2}{C_M}$

Thence : $C_M = \frac{[H^+]^2}{К_{Д}} = \frac{(1,35 \cdot 10^{-4})^2}{1,75 \cdot 10^{-5}} = 0,00104.$

3. Define α: For dilute solutions of weak electrolytes binary applicable formula: $\alpha = \sqrt{\frac{K_д}{C}} = \sqrt{\frac{1{,}75 \cdot 10^{-5}}{0{,}00104}} = 0{,}13$.

Response: $[H^+] = 0{,}000135$ mol/L; $C_M = 0{,}00104$ mol/L; $\alpha = 0{,}13$.

Example # 5

To 2 liters of 0.1M CH3COOH was added 49.2 g CH3COONa. Calculate the pH of the buffer solution (KD (CH3COOH) = $1{,}75 \cdot 10^{-5}$)

It is given	Answer
V(solution) = 2 L	
$C_M(CH_3COOH) = 0{,}1$ mol/L	$CH_3COOH \rightleftarrows CH_3COO^- + H^+$
$m(CH_3COONa) = 49{,}2$ g	
$K_д\ (CH_3COOH) = 1{,}75 \cdot 10^{-5}$	
	$CH_3COONa \rightarrow CH_3COO^- + Na^+$
	1Calculate the concentration of the sodium acetate solution:
pH - ?	$C_M(CH_3COONa) = m(CH_3COONa)/M(CH_3COONa) \cdot V = 49{,}2 / 82 \cdot 2 = 0{,}3$

2. Calculate the pH of acetate buffer solution:

$$pH = -lgK_д + lg\,\frac{[CH_3COONa]}{[CH_3COOH]} = -lg1{,}75 \cdot 10^{-5} + lg\,\frac{0{,}3}{0{,}1} = 4{,}75 + 0{,}48 = 5{,}23.$$

Answer: pH = 5,23

Example # 6

What is the pH of the buffer solution containing 1 l of 0.1 mol NH4OH and NH4Cl (pK (NH4OH) = 4,75)? How to change the pH of the solution when diluted with water by 10 times?10 раз?

It is given:

$C_M(NH_4OH) = 0{,}1$ mol/L

$C_M(NH_4Cl) = 0{,}1$ mol/L

V(so;ution) = 1 л

pК $(NH_4OH) = 4{,}75$

pH_1 - ? pH_2 - ?

1. Calculate the pH1 of the initial solution :

$$pH_1 = 14 – pK(NH_4OH) – \lg\frac{[NH_4Cl]}{[NH_4OH]} = 14 – 4{,}75 – \lg\frac{0{,}1}{0{,}1} = 9{,}25$$

2. Calculate the pH2 solution after dilution. When the solution diluted 10 times the concentration of salt and base are also reduced by 10 times: $pH_2 = 14 – 4{,}75 – \lg\frac{0{,}01}{0{,}01} = 9{,}25$*

Answer: $pH_1 = 9{,}25$; $pH_2 \approx 9{,}25$.

*Note: in fact, the pH changes upon dilution of several (in this case increased by approximately 0.07 units, depending on the change of the activity coefficients of ions due to a decrease in the ionic strength of the solution upon dilution).

Example # 7

To change the pH by one to 10 ml of an acetate buffer solution required to add 0,52 ml of a 1M solution of NaOH. Find in alkali buffer capacity (mol / l .pH) of the buffer solution.

It is given:

$\Delta pH = 1$

V = 10 ml = 0,01 L

$C_M(NaOH) = 1$ mol/L

V(base) = 0,52 ml=

Buffer capacity for alkali can be defined by the formula:

$$B_{bas.} = \frac{C_H(NaOH) \cdot V(NaOH)}{\Delta pH \cdot V} =$$

$0{,}52 \cdot 10^{-3}$ L

$B_{base.}$ - ?

$$= \frac{1 \cdot 0{,}52 \cdot 10^{-3}}{1 \cdot 0{,}01} = 0{,}052 \text{ mol/L. pH;}$$

$(C_н\ (NaOH) = C_M\ (NaOH))$

Answer: 0,052 mol/L. pH

Example # 8

To 16 mL 0.1 M Na2HPO4 solution was added 40 ml of 0.04 M solution NaH2PO4. Define:

a) The pH of the buffer solution (Kd $(H_2PO_4^-) = 1{,}6 \cdot 10^{-7}$;

b) as to change the pH of the solution by adding thereto 6 mL of 0.1 M sol. of HCl;

c) is it possible to prepare a phosphate buffer solution with pH = 8.5

It is given:

$C_M\ (Na_2HPO_4) = 0{,}1$ mol/L

V(р-ра Na_2HPO_4) = 16 ml

$C_M\ (NaH_2PO_4) = 0{,}04$ mol/L

$V(NaH_2PO_4) = 40$ ml

$C_M(HCl) = 0{,}1$ моль/л

$V\ (HCl) = 6$ мл

$K_д\ (H_2PO_4^-) = 1{,}6 \cdot 10^{-7}$

a) pH – ? б) ΔpH–?

a) Calculate the pH phosphate buffer solution

In the phosphate buffer solution as an acid ion, dissociating the following scheme: $H_2PO_4^- \rightleftarrows H^+ + HPO_4^{2-}$.

Since the dissociation constant of this process is low, we can assume that the concentration is equal to the concentration of NaH2PO4, and the concentration is equal to the concentration of Na2HPO4. Then:

$$pH = -\lg K_д\ (H_2PO_4^-) + \lg \frac{[HPO_4^{2-}]}{[H_2PO_4^-]}$$

$C_1 \cdot V_1 = C_2 \cdot V_2$. Then a new concentration is equal NaH_2PO_4

$[NaH_2PO_4] = \dfrac{C_M(NaH_2PO_4)1 \cdot V(NaH_2PO_4)}{V}$

The final concentration of Na2HPO4 will be:

$[Na_2HPO_4] = \dfrac{C_M(Na_2HPO_4)1 \cdot V(Na_2HPO_4)}{V}$

$C_M\ (NaH_2PO_4)_1 \cdot V(NaH_2PO_4) = n\ (NaH_2PO_4) = 0{,}040 \cdot 0{,}04 = 0{,}0016$

$C_M\ (Na_2HPO_4)_1 \cdot V(Na_2HPO_4) = n\ (Na_2HPO_4) = 0{,}016 \cdot 0{,}1 = 0{,}0016$

Then $\dfrac{[HPO_4^{2-}]}{[H_2PO_4^-]} = \dfrac{n(Na_2HPO_4) \cdot V}{n(NaH_2PO_4) \cdot V} = \dfrac{n(Na_2HPO_4)}{n(NaH_2PO_4)}$

Then: $pH = -\lg 1{,}6 \cdot 10^{-7} + \lg\dfrac{n(Na_2HPO_4)}{n(NaH_2PO_4)} = 6{,}8 + \lg\dfrac{0{,}0016}{0{,}0016} = 6{,}8 + \lg 1 = 6{,}8$

b) Calculate the change in pH when added to the buffer solution a solution of HCl.

By adding 6 ml of 0.1 M solution HCl (which is 0.0006 mol), added acid will react with 0.0006 mole Na2HPO4 forming 0.0006 mol NaH2PO4: $Na_2HPO_4 + HCl = NaH_2PO_4 + NaCl$

Then the amount of Na2HPO4 decrease by 0.0006 mol:

$n\ (Na_2HPO_4) = 0{,}0016 - 0{,}0006 = 0{,}0010$

A number of NaH2PO4 increase by 0.0006 mol:

$n\ (NaH_2PO_4) = 0{,}0016 + 0{,}0006 = 0{,}0022$

$pH = -\lg 1{,}6 \cdot 10^{-7} + \lg\dfrac{n(Na_2HPO_4)}{n(NaH_2PO_4)} = 6{,}8 + \lg\dfrac{0{,}0010}{0{,}0022} = 6{,}46$

$\Delta pH = 6{,}8 - 6{,}46 = 0{,}34$

c) Prepare a phosphate buffer solution with pH = 8.5 is impossible, since the buffer zone of the effective action of the system is given by □ pH = pK □ 1. For the phosphate buffer solution pK = 6.8, and the zone of effective buffering action in the pH range lies 5,8 -7,8.

Answer: a) pH = 6,8; b) decrease by 0.34 units. pH, c) is not possible.

EXPERIMENTAL PART

Laboratory work 1: Titrimetric analysis

Fill the clean burette by NaOH of known concentration and determine the level of the solution to the "zero" level (the lower meniscus). Then put 10 ml of the acid solution of unknown concentration into the titration flask using a pipette. Add 2-3 drops of phenolphthalein solution and titrate from the burette until slightly pink coloration of the solution appears.

Using the amount of alkali calculate the titer and normality of the acid solution for each of the three tasks by the formula:

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

$$T_{HCl} = N_{HCl} \cdot Э_{HCl}$$

Laboratory work 2: Composition and Preparation of Common Buffers and Solutions

Preparation of Bicarbonate-Carbonate Buffer (pH 9.2–10.8)

To create 100ml of a 0.1M bicarbonate buffer solution, mix sodium bicarbonate and sodium carbonate, decahydrate, as given below.

Solution A: 0.1M sodium bicarbonate ($NaHCO_3$ MW = 84.0) (MW = molecular weight)

Solution B: 0.1M sodium carbonate, decahydrate ($Na_2CO_3 \cdot 10H_2O$ FW = 286.2) (FW = formula weight)

Table 1. Bicarbonate-Carbonate Buffer.

pH at 20°C	**pH at 37°C**	**Solution A (ml)**	**Solution B (ml)**
9.4	9.1	80	20
9.5	9.4	70	30
9.8	9.5	60	40

9.9	9.7	50	50
10.1	9.9	40	60
10.3	10.1	30	70
10.5	10.3	20	80
10.8	10.6	10	90

Preparation of Citrate Buffer (pH 3.0–6.2)

To create 100ml of a 0.1M citrate buffer, mix citric acid, monohydrate, and trisodium citrate dehydrate as given below.

Solution A: 0.1M citric acid monohydrate ($C_6H_8O_7 \bullet H_2O$ FW = 210.14)

Solution B: 0.1M trisodium citrate, dihydrate $C_6H_5O_7Na_3 \bullet 2H_2O$ FW = 294.12)

Table 2. Citrate Buffer.

pH	**Solution A (ml)**	**Solution B (ml)**
3.0	82.0	18.0
3.2	77.5	22.5
3.4	73.0	27.0
3.6	68.5	31.5
3.8	63.5	36.5
4.0	59.0	41.0
4.2	54.0	46.0
4.4	49.5	50.5
4.6	44.5	55.5
4.8	40.0	60.0

5.0	35.0	65.0
5.2	30.5	69.5
5.4	25.5	74.5
5.6	21.0	79.0
5.8	16.0	84.0
6.0	11.5	88.5
6.2	8.0	92.0

Preparation of Phosphate Buffer (pH 5.8–8.0 at 25°C)

To create 100ml of a 0.1M phosphate buffer, mix sodium phosphate, dibasic dihydrate and sodium phosphate monobasic monohydrate, as given below, and dilute to 100ml with water.

Note: The dibasic stock sodium phosphate may be somewhat harder to dissolve; adding a little heat may help.

Solution A: 0.2M sodium phosphate, dibasic dihydrate ($Na_2HPO_4 \bullet 2H_2O$ FW = 178.05)

Solution B: 0.2M sodium phosphate, monobasic, monohydrate ($NaH_2PO_4 \bullet H_2O$ FW = 138.01)

Table 3. Phosphate Buffer.

pH at 25°C	**Solution A (ml)**	**Solution B (ml)**
5.8	4.0	46.0
6.0	6.15	43.85
6.2	9.25	40.75
6.4	13.25	36.75

6.6	18.75	31.25
6.8	24.5	25.5
7.0	30.5	19.5
7.2	36	14
7.4	40.5	9.5
7.6	43.5	6.5
7.8	45.75	4.25
8.0	47.35	2.65

TESTS

What is buffer effect:

a) *solution resists the change of pH when acids or alkalis or water are added to solution
b) solution resists the change of pH when acids or alkalis are added to solution
c) solution resists the change of pH when water are added to solution
d) solution resists the change of pH when acids are added to solution
e) solution resists the change of pH when alkalis are added to solution

What composition of buffer solutions are?

a) *mixture of either a weak acid and its conjugate base, or weak base and its conjugate acid
b) mixture of weak acid and its salt
c) mixture of weak base and its salt
d) mixture of two salts
e) mixture of two acids

How do buffer solutions classify?

a) *Acid, base, salt buffers
b) Solid, liquid, gaseous buffers
c) Stable and non-stable buffers
d) Strong and weak buffers
e) Resistant and non-resistant

Choose range of pH of acetate buffer:

a) *3,6 - 5,6
b) 8,3 - 10,3
c) 9,24 - 11,0
d) 9,2 - 10,8
e) 5,8 - 8,0

Choose range of pH of ammonia buffer:

a) *8,3 - 10,3
b) 3,6 - 5,6
c) 9,24 - 11,0
d) 9,2 - 10,8
e) 5,8 - 8,0

Choose range of pH of borate buffer:

a) *9,24 - 11,0
b) 8,3 - 10,3
c) 3,6 - 5,6
d) 9,2 - 10,8
e) 5,8 - 8,0

Choose range of pH of carbonate buffer:

a) *9,2 - 10,8
b) 8,3 - 10,3
c) 9,24 - 11,0

d) 3,6 - 5,6

e) 5,8 - 8,0

Choose range of pH of phosphate buffer:

a) *5,8 - 8,0

b) 8,3 - 10,3

c) 9,24 - 11,0

d) 9,2 - 10,8

e) 3,6 - 5,6

Choose acid buffer:

a) *$H_2 CO_3/NaHCO_3$

b) NH_4OH/NH_4Cl

c) NaH_2PO_4/ Na_2HPO_4

d) $Na_2B_4O_7/NaOH$

e) $N(CH_3)_2/[NH(CH_3)_2]Cl$

Choose salt buffer:

a) *NaH_2PO_4/ Na_2HPO_4

b) H2 $CO_3/NaHCO_3$

c) NH_4OH/NH_4Cl

d) $Na_2B_4O_7/NaOH$

e) $N(CH_3)_2/[NH(CH_3)_2]Cl$

Choose base buffer:

a) *NH_4OH/NH_4Cl

b) $H_2\ CO_3/NaHCO_3$

c) $H_2\ SO_3/NaHSO_3$

d) $Na_2B_4O_7/NaOH$

e) $NaH_2PO_4/\ Na_2HPO_4$

Choose mechanism of buffer effect of acetate buffer solution at presence of alkalis:

a) *$CH_3COOH + OH^- = CH_3COO^- + H_2O$

b) $CH_3COO^- + H^+ = CH_3COOH$

c) $CH_3COOH = CH_3COO^- + H^+$

d) $CH_3COO^- + H_2O = CH_3COOH + OH^-$

e) $CH_3COOH^{2+} + 2OH^- = CH_3COO^- + H_2O$

Choose mechanism of buffer effect of acetate buffer solution at presence of acids:

a) *$CH_3COO^- + H^+ = CH_3COOH$

b) $CH_3COOH + H^+ = CH_3COOH_2^+$

c) $CH_3COOH = CH_3COO^- + H^+$

d) $CH_3COO^- + H_2O = CH_3COOH + OH^-$

e) $CH_3COOH^{2+} + 2OH^- = CH_3COO^- + H_2O$

Choose mechanism of buffer effect of bicarbonate buffer solution at presence of alkalis:

a) *$H_2CO_3 + OH^- = HCO_3^- + H_2O$

b) $HCO_3^- + OH^- = CO_3^{2-} + H_2O$

c) $HCO^{3-} + H^+ = H_2CO_3$

d) $HCO_3^- + H_2O = H_2CO_3 + OH^-$
e) $CO_3^{2-} + H^+ = HCO_3^-$

Choose mechanism of buffer effect of bicarbonate buffer solution at presence of acids:

a) *$HCO^{3-} + H^+ = H_2CO_3$
b) $CO_3^{2-} + H^+ = HCO^{3-}$
c) $HCO^{3-} + OH^- = CO_3^{2-} + H_2O$
d) $H_2CO_3 = HCO_3^- + H^+$
e) $HCO_3^- + H_2O = H_2CO_3 + OH^-$

Choose mechanism of buffer effect of ammonia buffer solution at presence of alkalis:

a) *$NH_4^+ + OH^- = NH_3 + H_2O$
b) $NH_4OH + H^+ = NH_4^+ + H_2O$
c) $NH_4OH = NH_4^+ + OH^-$
d) $NH_3 + H_2O = NH_4OH$
e) $NH_3 + H^+ = NH_4^+$

Choose mechanism of buffer effect of ammonia buffer solution at presence of acids:

a) *$NH_3 + H^+ = NH_4^+$
b) $NH_4^+ = NH_3 + H^+$
c) $NH_4OH = NH_4^+ + OH^-$
d) $NH_3 + H_2O = NH_4OH$
e) $NH_4^+ + OH^- = NH_3 + H_2O$

Choose mechanism of buffer effect of phosphate buffer solution at presence of alkalis:

a) *$H_2PO_4^- + OH^- = HPO_4^{2-} + H_2O$
b) $HPO_4^{2-} + OH^- = PO_4^{3-} + H_2O$
c) $H_2PO_4^- + H^+ = H_3PO_4$
d) $H_2PO_4^- = HPO_4^{2-} + H^+$
e) $H_3PO_4 + OH^- = H_2PO_4^- + H_2O$

Choose mechanism of buffer effect of phosphate buffer solution at presence of acids:

a) *$HPO_4^{2-} + H^+ = H_2PO_4^-$
b) $H_2PO_4^- + H^+ = H_3PO_4$
c) $PO_4^{3-} + H^+ = HPO_4^{2-}$
d) $HPO_4^{2-} + OH^- = PO_4^{3-} + H_2O$
e) $H_2PO_4^- + H_2O = H_3PO_4 + OH^-$

Choose formula of calculation of pH of buffer solutions:

a) *pH=pKa+ lg(Cb/Ca)
b) pH=1/2 pKa-1/2 lgCa
c) pH=7+ 1/2 pKa+ 1/2 lgCb
d) pH = -lg[H+]
e) pH = 14 - pOH

What is buffering capacity?

a) *Number of moles of a strong monoprotonate base or acid required to be added to 1 l of buffer solution to rais its pH by 1
b) Number of moles of a strong monoprotonate base or acid required to be added to 1 l of buffer solution to rais its pH by 5
c) Number of moles of a strong monoprotonate base or acid required to be added to 1 kg of buffer solution to rais its pH by 1
d) Number of moles of a strong base or acid required to be added to 1 kg of buffer solution to rais its pH by 1
e) Number of moles of a strong biprotonate base or acid required to be added to 1 kg of buffer solution to rais its pH by 1

Choose buffer solution which presence in the human body:

a) *Bicarbonate
b) Acetate
c) Borate
d) Ammonia
e) Tartrate

Choose buffer solution which presence in the human body:

a) *Phosphate
b) Acetate
c) Borate
d) Ammonia
e) Tartrate

Choose buffer solution which presence in the human body:

a) *Protein
b) Acetate
c) Borate
d) Ammonia
e) Tartrate

Choose buffer solution which presence in the human body:

a) *Hemoglobin
b) Acetate
c) Borate
d) Ammonia
e) Tartrate

Choose buffer solution which presence in the human body:

a) *Oxyhemoglobin
b) Acetate
c) Borate
d) Ammonia
e) Tartrate

What normal range of pH of blood:

a) *7,36-7,42
b) less than 6,90
c) more than 7,80
d) ~7,0
e) 6,90-7,80

What normal ratio [HCO_3^-]/[H_2CO_3] inplasma?

a) *20

b) 10

c) 3

d) 4

e) 15

What concentrations of HCO_3^- and H_2CO_3in plasma?

a) *28 and 1,4mM

b) 15 and 1,6mM

c) 28 and 15 mM

d) 1,4 and 28 mM

e) 1,6 and 15 mM

What is alkali reserve?

a) *HCO_3^- - content in plasma

b) HPO_4^{2-} - content in plasma

c) CH_3COO^- - content in plasma

d) Pr^-- content in plasma

e) NH_3 - content in plasma

What place of biological importance of carbonate buffer?

a) *Extracellular fluids

b) Intracellular fluids

c) Neither extracellular nor intracellular

d) Intracellular and extracellular fluids

e) In erytrothytes

What normal ratio [HPO_4^{2-}]/[$H_2PO_4^-$]?

a) *4

b) 20

c) 15

d) 2

e) 10

Choose effect of hypoventilation:

a) *respiratory acidosis

b) respiratory alkalosis

c) metabolic acidosis

d) metabolic alkalosis

e) pH of blood not change

Choose effect of hyperventilation:

a) *respiratory alkalosis

b) respiratory acidosis

c) metabolic alkalosis

d) pH of blood not change

e) metabolic acidosis

How organism usually compensate acidosis or alkalosis?

a) *Pulmonary elimination of CO_2 or urinary elimination ofHCO_3^-

b) Urinary elimination of CO_2
c) Pulmonary elimination of O_2
d) Pulmonary elimination of HCO_3^-
e) Urinary elimination of proteins

What happens to the pH of a buffer system if one halves the concentration of both the acid and the salt?

a) *Nothing
b) pH goes up because there is less total acid in the solution.
c) pH goes down because there is less conjugate base to mask the presence of the acid.
d) It depends upon the original concentration of acid and salt.
e) It is impossible to predict.

The body's water volume is closely tied to the level of which of the following ions?

a) calcium ions
b) potassium ions
c) hydrogen ions
d) *sodium ions

The term hypotonic hydration refers to ________.

a) the feeling one might have after a long swim
b) the unpleasant feeling people have after drinking too much liquor
c) *a condition that may result from renal insufficiency or drinking extraordinary amounts of water

d) a condition that is caused by high levels of sodium in the extracellular fluid compartment

Hypoproteinemia is a condition of unusually low levels of plasma proteins. This problem is often characterized by ________.

a) *tissue edema
b) extreme weight loss
c) extreme weight gain
d) nerve damage

Which of the following hormones is important in the regulation of sodium ion concentrations in the extracellular fluid?

a) antidiuretic hormone
b) erythropoietin
c) *aldosterone
d) renin

Atrial natriuretic peptide is a hormone that is made in the atria of the heart. The influence of this hormone is to ________.

a) enhance atrial contractions
b) activate the renin-angiotensin mechanism
c) prevent pH changes caused by organic acids
d) *reduce blood pressure and blood volume by inhibiting sodium and water retention

Respiratory acidosis can occur when ________.

a) a person consumes excessive amounts of antacids
b) *a person's breathing is shallow due to obstruction
c) a runner has completed a very long marathon
d) the kidneys secrete hydrogen ions

Which of the following two organs function as the most important physiological buffer systems?

a) *the lungs and the kidneys
b) the adrenal glands and the testes
c) the thyroid gland and the heart
d) the stomach and the liver

Which of the choices below is not an essential role of salts in the body?

a) neuromuscular activity
b) membrane permeability
c) secretory activity
d) * anabolism of proteins

Which of the choices below exerts primary control over sodium levels in the body?

a) ADH
b) * aldosterone
c) water levels
d) glucocorticoids

The fluid link between the external and internal environment is ________.

a) *plasma
b) intracellular fluid
c) interstitial fluid
d) cerebrospinal fluid

Newborn infants have a relatively higher ________ content in their ECF than do adults.

a) iron
b) *sodium
c) magnesium
d) bicarbonate

Whereas sodium is found mainly in the extracellular fluid, most ________ is found in the intracellular fluid.

a) iron
b) chloride
c) * potassium
d) magnesium

Which of the following describes the distribution of sodium and potassium between cells and body fluids?

a) K+ mainly in the cells, Na+ in the body fluids
b) Na+ mainly in the cells, K+ in the body fluids
c) equal amounts of each ion in the cells and body fluids
d) little of either in the cells, but large amounts of each in the body fluids

Problems with fluid, electrolyte, and acid-base balance are particularly common in infants because of their ________.

a) *inefficient kidneys
b) comparatively low metabolic rates
c) low rate of insensible water loss
d) low daily rate of fluid exchange

The single most important factor influencing potassium ion secretion is ________.

a) the potassium ion content in the renal tubule cells
b) the pH of the ICF
c) intracellular sodium levels
d) *potassium ion concentration in blood plasma

The term alkaline reserve is used to describe the ________ buffer system.

a) phosphate
b) hemoglobin
c) *bicarbonate
d) protein

A falling blood pH and a rising partial pressure of carbon dioxide due to pneumonia or emphysema indicates ________.

a) *respiratory acidosis
b) respiratory alkalosis
c) metabolic acidosis
d) metabolic alkalosis

The movement of fluids between cellular compartments ________.

a) requires active transport

b) *is regulated by osmotic and hydrostatic forces

c) requires ATP for the transport to take place

d) involves filtration

What hormone reduces blood pressure and blood volume by inhibiting nearly all events that promote vasoconstriction and sodium ion and water retention?

a) ADH

b) aldosterone

c) *atrial natriuretic peptide

d) thyroxine

Which of the following is not a method for regulating the hydrogen ion concentration in blood?

a) chemical buffer systems

b) * diet

c) respiratory changes

d) renal mechanism

Which of the following is not a chemical buffer system?

a) bicarbonate

b) phosphate

c) *nucleic acid

d) protein

Extracellular fluid in the human body is composed of all of the following except ________.

a) lymph and interstitial fluid
b) blood plasma
c) cerebrospinal fluid
d) *glucose

Which of the following statements is true regarding fluid shifts?

a) Nonelectrolytes are the controlling factor in directing fluid shifts.
b) Electrolytes are not as important as proteins in regulating fluid shifts in the body.
c) * Electrolytes have greater osmotic power than nonelectrolytes and therefore have the greatest ability to cause fluid shifts.
d) There are always more positive electrolytes than negative in a solution; it is therefore impossible to follow fluid shifts.

Which of the following hormones is important in stimulating water conservation in the kidneys?

a) aldosterone
b) thymosin
c) *antidiuretic hormone
d) atrial natriuretic peptide

The maintenance of the proper pH of the body fluids may be the result of ________.

a) *the control of respiratory ventilation

b) the operation of the various buffer systems in the stomach

c) the active secretion of OH- into the filtrate by the kidney tubule cells

control of the acids produced in the stomach

Which of the following is not a disorder of water balance?

a) *excessive hydration due to excess ADH secretion

b) hypotonic hydration, in which sodium content is normal but water content is high

c) edema or tissue swelling, which is usually due to an increased capillary hydrostatic pressure

d) excess water in interstitial spaces due to a low level of plasma proteins

The regulation of sodium ________.

a) is due to specific sodium receptors in the hypothalamus

b) *is linked to blood pressure

c) involves aldosterone, a hormone that increases sodium excretion in the kidneys

d) involves hypothalamic osmoreceptor detection of ion concentration

Select the correct statement about renal mechanisms of acid-base balance.

a) The kidneys are not able to excrete phosphoric acid.

b) Excreted hydrogen ions are unbound in the filtrate.

c) *Kidney tubule cells are able to synthesize bicarbonate ion.

d) The kidneys are the most important mechanism for eliminating all bicarbonate ions.

Blood analysis indicates a low pH, and the patient is breathing rapidly. Given your knowledge of acid-base balance, which of the following is most likely?

a) respiratory acidosis
b) *metabolic acidosis
c) metabolic alkalosis
d) respiratory alkalosis

A patient is breathing slowly and blood pH analysis indicates an abnormally high value. What is the likely diagnosis?

a) respiratory acidosis
b) metabolic acidosis
c) *metabolic alkalosis
d) respiratory alkalosis

One of the major physiological factors that triggers thirst is ________.

a) a dry mouth from high temperatures
b) becoming overly agitated
c) drinking caffeinated beverages
d) *a rise in plasma osmolality

Annie has just eaten a large order of heavily salted french fries, some pickled eggs, and some cheese. How will consuming this much salt affect her physiology?

a) It will increase the osmolality of the blood.
b) *There will be a temporary increase in blood volume.

c) She will experience hypotension.
d) There will be a shift in the pH of her body fluids to the higher side of the pH scale.

The most important force causing net water flow across capillary walls is ________.

a) osmotic pressure of plasma proteins
b) *hydrostatic pressure of capillary blood
c) hydrostatic pressure of interstitial fluid
d) intracellular hydrostatic pressure

Which of the following does not depend on the presence of electrolytes?

a) membrane polarity
b) neuromuscular excitability
c) maintenance of osmotic relations between cells and ECF
d) *amount of body fat

The regulation of potassium balance ________.

a) is not linked to sodium balance
b) includes renal secretion, but never absorption
c) is accomplished mainly by hepatic mechanisms
d) *involves aldosterone-induced secretion of potassium

Atrial natriuretic peptide is a hormone that is made in the atria of the heart. The influence of this hormone is to ________.

e) enhance atrial contractions

f) activate the renin-angiotensin mechanism
g) prevent pH changes caused by organic acids
h) *reduce blood pressure and blood volume by inhibiting sodium and water retention

Respiratory acidosis can occur when ________.

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f) *a person's breathing is shallow due to obstruction
g) a runner has completed a very long marathon
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h) * anabolism of proteins

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h) glucocorticoids

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f) *sodium
g) magnesium
h) bicarbonate

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g) * potassium
h) magnesium

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