Ministry of Health of the Republic of Belarus
Educational institution
"Gomel State Medical University"

Department of General and Bioorganic Chemistry

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METHODOLOGICAL MANUAL
for conducting the laboratory class
in the academic discipline "Medical chemistry"
for students
the first-year of the Faculty of International Students (FIS),
studying in the specialty 7-07-0911-01 "General medicine"

Topic 9: Colligative properties of solutions

Time: 2 hours

Approved at the meeting of the Department
of General and Bioorganic Chemistry
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2026

THE TRAINING AND EDUCATIONAL GOAL, TASKS, MOTIVATION TO STUDY THE TOPIC

Training purpose:

– formation of students' basic professional competence for solving diagnostic, research and other tasks of professional activity based on knowledge about colligative properties of solutions of electrolytes and nonelectrolytes.

Educational goal:

– to develop your personal, spiritual potential;

– to form the qualities of a patriot and a citizen who is ready to actively participate in the economic, industrial, socio-cultural and social life of the country;

– learn to observe academic and labor discipline, the norms of medical ethics and deontology;

– to realize the social significance of their future professional activities.

Tasks:

As a result of the training session, the student should

know:

– the laws describing the colligative properties of solutions of electrolytes and nonelectrolytes;

be able to:

– solve computational problems to determine the colligative properties of solutions of various substances;

possess:

– skills in experimental determination of the molar masses of nonelectrolytes by cryoscopic method.

Motivation to study the topic:

The study of the colligative properties of solutions of substances allows us to form an idea about the processes of mass transfer in the human body, as well as about the properties of solutions used in medicine.

Osmosis regulates the inflow of water into the cells and intercellular structures. Osmotic pressure is important in biological systems because cell walls are semipermeable membranes. In particular, when a person is receiving intravenous fluids, the osmotic pressure of the fluid needs to be approximately the same as blood serum; otherwise bad things can happen.

NECESSARY EQUIPMENT

1. Methodological manual for students on the topic "Colligative properties of solutions".

2. Training tables:

a) Periodic Table of chemical elements by D.I. Mendeleev;

b) table of solubility of acids, bases and salts.

3. Reference materials of basic physico-chemical constants.

4. Chemical reagents and equipment necessary for laboratory work.

CONTROL QUESTIONS ON THE TOPIC OF THE CLASS

1. Colligative properties of solutions of electrolytes and nonelectrolytes. Vapor pressure lowering is the First Raoult’s Law.

2. Cryoscopic and Ebullioscopic Raoult's Laws.
3. Osmosis and osmotic pressure.

COURSE OF THE CLASS

The theoretical part

1. COLLIGATIVE PROPERTIES OF SOLUTIONS OF ELECTROLYTES AND NONELECTROLYTES. VAPOR PRESSURE LOWERING IS THE FIRST RAOULT'S LAW

Colligative properties are properties of dilute solutions that depend on the number of solute's particles present in the solution and are independent on the solute's nature.

Four important colligative properties that we will examine here are follows:

- *vapor pressure lowering;*
- *boiling point elevation;*
- *freezing point depression;*
- *osmotic pressure.*

Molecular compounds (nonelectrolytes) separate into individual molecules when they are dissolved, so for every 1 mol of dissolved molecules, we get 1 mol of particles. On the contrary, ionic compounds (electrolytes) separate into their constituent ions when they dissolve, so 1 mol of an ionic compound will produce more than 1 mol of dissolved particles. For example, each mole of NaCl that dissolves give 1 mol of Na^+ ions and 1 mol of Cl^- ions, for a total of 2 mol of particles in solution. Thus, the effect of the dissolution of NaCl on the properties of the solution can be twice as large as the effect of the dissolution of the same number of moles of glucose ($C_6H_{12}O_6$) [1-3].

The First Raoult's Law

Vapor pressure is what you get when the liquid is in a closed container at a certain temperature. At this temperature, an equilibrium is established between the number of particles leaving the surface and the number of particles returning to the surface again. In the 1980s, Raoult discovered that when a non-volatile solute (such as salt or sugar) is dissolved in a solvent, the vapor pressure decreases. The lowering of vapor pressure depends on the mole fraction of the solute and the vapor pressure of the pure solvent.

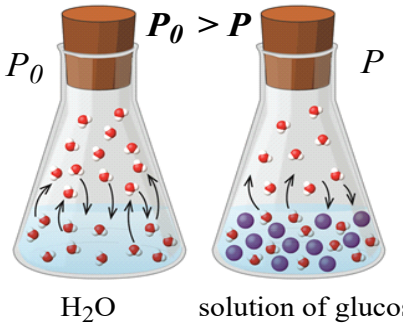


Figure 1 – The presence of non-volatile solute lowers the vapor pressure of a solution, reducing the evaporation of solvent molecules.

The First Raoult's Law may be formulated as:

- ***the vapor pressure of solution containing non-volatile solute(s) is always lower than the vapor pressure of pure solvent.***

Processes of evaporation and condensation are smaller in solution because nonvolatile solute bind solvents' molecules thus retarding evaporation process (decreasing a fraction of solvent molecules with energy sufficient to enter the vapor phase).

Mathematical description of the First Raoul's Law for a binary system:

For nonelectrolytes:

$$\frac{P_0 - P}{P_0} = \frac{n(X)}{n(X) + n(solvent)} = \chi(X)$$

▪ ***The relative lowering of vapor pressure of a dilute solution containing non-volatile solute is equal to the mole fraction of the solute – χ(X).***

P_0 – vapor pressure of a pure solvent;
P – vapor pressure of a solution;

$\frac{P_0 - P}{P_0}$ — relative lowering of vapor pressure;

$n(X)$ – number of moles of solute, mol;
$n(solvent)$ – number of moles of solvent, mol.

For electrolytes:

$$\frac{P_0 - P}{P_0} = \frac{i \times n(X)}{i \times n(X) + n(solvent)}$$

i – isotonic coefficient (Van't Hoff's factor), that characterizes ionization of the electrolyte;

$$\alpha = \frac{i - 1}{n - 1}$$

α – ionization percent;
n – number of ions contained in the molecule:

NaCl	$n = 2$
Na_2SO_4	$n = 3$
Na_3PO_4	$n = 4$.

2. CRYOSCOPIC AND EBULLIOSCOPIC RAOULT'S LAWS

Solutions freeze at lower temperatures than pure solvents. Seawater freezes at a lower temperature than pure water, and so the Arctic and Antarctic oceans remain unfrozen even at temperatures below 0 °C (as do the body fluids of fish and other cold-blooded sea animals that live in these oceans).

Cryoscopic Raoult's Law: ***freezing point of solution containing non-volatile solute(s) is always lower than freezing point of pure solvent.***

In other words, ***the decrease in freezing point of a dilute solution is directly proportional to the molal concentration of the solute.***

Freezing point of a liquid is the particular temperature at which a liquid transforms into a solid or the temperature at which the vapor pressure of solid is equal to the vapor pressure of liquid.

Solution has lower vapor pressure than pure solvent and hence freezes at lower temperature than pure solvent.

Mathematical description of Cryoscopic Law for a binary system:

For nonelectrolytes:

$$\Delta T_f = K_f \times C_m(X) = \frac{K_f \times m(X) \times 1000}{M(X) \times m(solvent)}$$

For electrolytes:

$$\Delta T_f = i \times K_f \times C_m(X)$$

$\Delta T_f = T_{f\,(Solvent)} - T_{f\,(Sln)}$ – freezing point depression (difference between the freezing point of pure solvent and freezing point of solution containing non-volatile solute);

$C_m(X)$ – molal concentration (molality) of the non-volatile solute in the solvent, mol/kg;

K_f – the molar freezing point depression constant of a solvent or cryoscopic constant (1.86°×kg/mol for H_2O);

m(X), m(solvent) – mass of the non-volatile solute (solvent), g.

For blood plasma of people ΔT_f equals to 0.56°, for blood plasma of mammals ΔT_f is a bit higher (0.58°).

Problem 1. Which solution's freezing point deviates more from that of pure water – a 1 M solution of NaCl or a 1 M solution of $CaCl_2$?

Solution:

Colligative properties depend on the number of dissolved particles, so the solution with the greater number of particles in solution will show the greatest deviation. When NaCl dissolves, it dissociates into two ions, Na^+ and Cl^-. But when $CaCl_2$ dissolves, it dissociates into three ions – one Ca^{2+} ion and two Cl^- ions. Thus, $CaCl_2$ will have 50 % more impact on freezing point depression than NaCl.

Ebullioscopic Raoult's Law: ***boiling point of solution containing non-volatile solute(s) is always higher than boiling point of pure solvent.***

In other words, ***the increase in boiling point of a dilute solution is directly proportional to the molal concentration of the solute species.***

The boiling point of a liquid is a temperature under which its vapor pressure equals the external atmospheric pressure.

Since the vapor pressure of a solution is lowered due to the presence of non-volatile solutes, it stands to reason that the solution's boiling point will subsequently be increased. Compared to pure solvent, a solution, therefore, will require a higher temperature to achieve any given vapor pressure, including one equivalent to that of the surrounding atmosphere.

Mathematical description of Ebullioscopic c Law for a binary system:

For nonelectrolytes:

$$\Delta T_b = K_b \times C_m(X) = \frac{K_b \times m(X) \times 1000}{M(X) \times m(solvent)}$$

For electrolytes:

$$\Delta T_b = i \times K_b \times C_m(X)$$

$\Delta T_b = T_{b\,(Sln)} - T_{b\,(Solvent)}$ – boiling point elevation (difference in boiling points of a solution and a pure solvent);

$C_m(X)$ – molal concentration (molality) of the non-volatile solute in the solvent, mol/kg;

K_b – the molar boiling point elevation constant of a solvent or ebullioscopic constant (0.52°×kg/mol for water);

m(X), m(solvent) – mass of the non-volatile solute (solvent), g.

Problem 2. Calculate the freezing point depression and boiling point elevation for a solution containing 1.49 g of KCl in 360 ml of water.

Solution:

$\Delta T_f = i \times K_f \times C_m(KCl)$; $\Delta T_b = i \times K_b \times C_m(KCl)$

$$C_m(KCl) = \frac{moles\,of\ solute:\ n(KCl)}{mass\,of\ solvent(kg):\ m(H_2O)}$$

$n\,(KCl) = m / M = 1.49 / 74.5 = 0.02$ mol; $m\,(H_2O) = 0.360$ kg

$C_m\,(KCl) = \frac{0.02}{0.360} = 0.055\,mol/kg$

KCl is a strong electrolyte, hence, $\alpha(KCl) = 100\ \%$ or 1; $KCl \rightarrow K^+ + Cl^-$

$\alpha = \frac{i-1}{n-1}$ (n – number ions contained in a molecule), hence,

$i = \alpha \times (n-1) + 1 = 1 \times (2-1) + 1 = 2$

$\Delta T_f = 1.86 \times 2 \times 0.055 = 0.2046°$; $\Delta T_b = 0.52 \times 2 \times 0.055 = 0.0570°$

Answer: $\Delta T_f = 0.2046°$; $\Delta T_b = 0.0570°$

3. OSMOSIS AND OSMOTIC PRESSURE

Osmosis (from the Greek *osmós*, meaning "push") is the spontaneous and unidirectional flow of solvent molecules through a semipermeable membrane, into the solution or flow of solvent from a solution of lower concentration to the solution of higher concentration through a semipermeable membrane.

Semipermeable membrane is a membrane which allows the solvent molecules, but not the solute molecules, to pass through it. Cellulose, cellulose nitrate, animal bladder, etc. are used as semipermeable membranes.

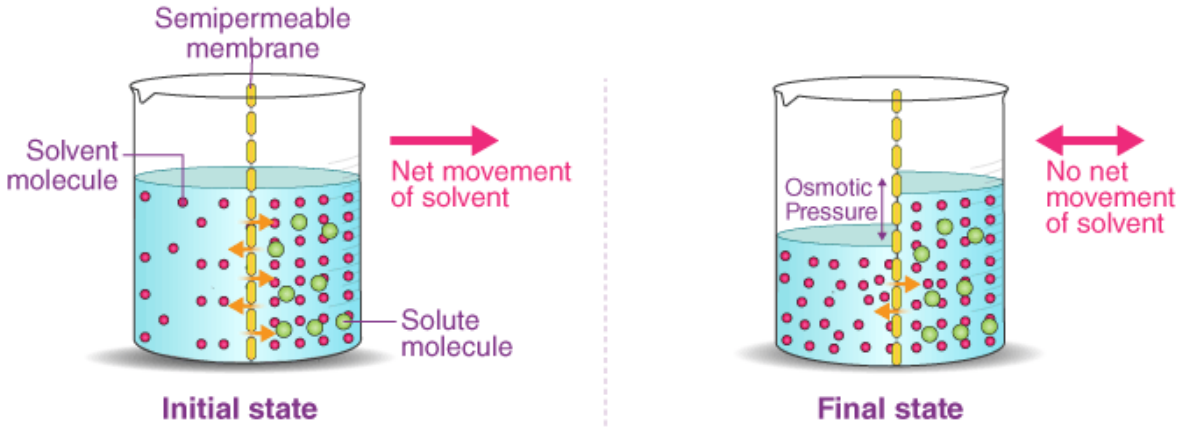


Figure 2 – Scheme of the osmosis

Osmosis describes a net movement of water molecules through a semipermeable membrane from the side where water concentration is higher/solute concentration is lower to the side where water concentration is lower/solute concentration is higher. This can be viewed as an attempt to equalize the concentrations of all substances on both sides of the membrane, maximizing entropy (disorder) and reaching a lower energy state.

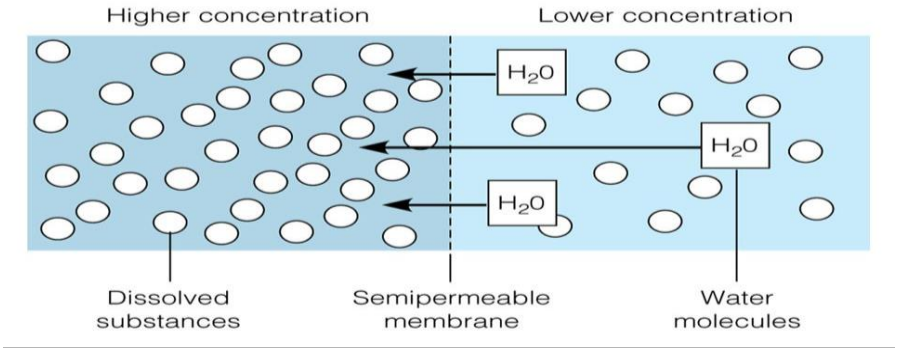


Figure 3 – The movement of water molecules during osmosis

Diffusion of water through the membrane creates an overpressure in the solution, called osmotic pressure. Osmotic pressure is a colligative property of solutions that is observed when using a semipermeable membrane.

Osmotic pressure (π, kPa) is the amount of pressure required to prevent the osmosis. Osmotic pressure is calculated by the *Van't Hoff's equations:*

For nonelectrolytes: $\pi = R \times T \times C_M(X)$,
For electrolytes: $\pi = i \times R \times T \times C_M(X)$

T – temperature, K;
R – ideal gas constant (8.31 J/mol×K);
i – Van't Hoff's factor;
$C_M(X)$ – molar concentration of the solute in the solution, mol/l.

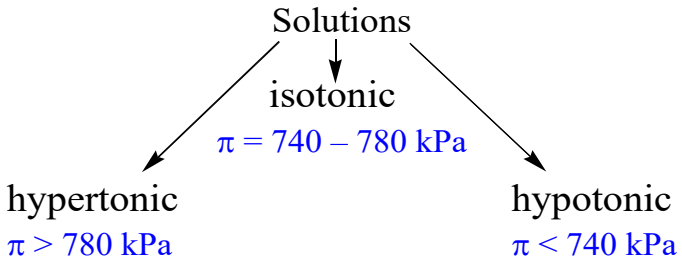
Types of solution on the basis of osmotic pressure:

- *isotonic solutions* are solutions having the same osmotic pressure; for example, 0.05 M urea solution and 0.05 M sucrose solution are isotonic because their osmotic pressures are the same; if these solutions are separated by a semipermeable membrane, there is no flow of solvent in either direction;
- solution whose osmotic pressure is higher than that of another solution is considered *hypertonic* in relation to this solution; e.g. 0.1 M urea solution has higher osmotic pressure than 0.05 M sucrose solution, hence, 0.1 M urea solution is hypertonic to 0.05 M sucrose solution; if these solutions are separated by a semipermeable membrane, the solvent flows from sucrose to urea as sucrose is having low concentration;
- solution having osmotic pressure lower than that of another is said to be *hypotonic* solution with that solution; for example, 0.05 M sucrose solution has osmotic pressure lower than that of 0.1 M urea solution, therefore 0.05 M sucrose solution is hypotonic with 0.1 M urea solution.

In biological systems osmosis is very essential as cell membranes are semipermeable, so the osmotic pressures of the body's fluids have important biological consequences.

Cell membranes of plants and animals are permeable for water molecules and some small ions. Normal osmotic pressure of blood plasma is 740-780 kPa (37 ºC) or 7.3-7.7 atm (1 atm = 101.3 kPa).

Depending on the value of the osmotic pressure of the solutions in relation to the value of the osmotic pressure of the blood plasma, the solutions can be classified:



Solute concentrations are especially important when solutions are injected into the body. Solutions injected into the body must have the same value of osmotic pressure as blood plasma; that is, they should be *isotonic* with respect to the blood plasma. Their injection into the blood vessel doesn't disturb osmotic homeostasis of a body. *Isotonic* solutions used in medicine for intravenous injections are:

a) 0.9 % by mass NaCl solution;

b) 4.5 % by mass glucose solution.

When a more concentrated solution, a *hypertonic solution (π > 780 kPa),* is injected, intracellular water tends to be removed from the cell by osmosis into a more concentrated surrounding solution. This process, known as *plasmolysis*, causes the cell to shrink and eventually stop functioning.

If a less concentrated solution, a *hypotonic solution (π < 740 kPa)*, is injected, water tends to move into the cell; the cell swells. This phenomenon is called *deplasmolysis.* Animal cells are rapidly destroyed in such a solution, because the membrane cannot withstand high osmotic pressure and breaks. This phenomenon is called *cytolysis.* Special cases of cytolysis are the destruction of red blood cells – *hemolysis* (occurs at 260-300 kPa), while hemoglobin enters the blood plasma and turns it red.

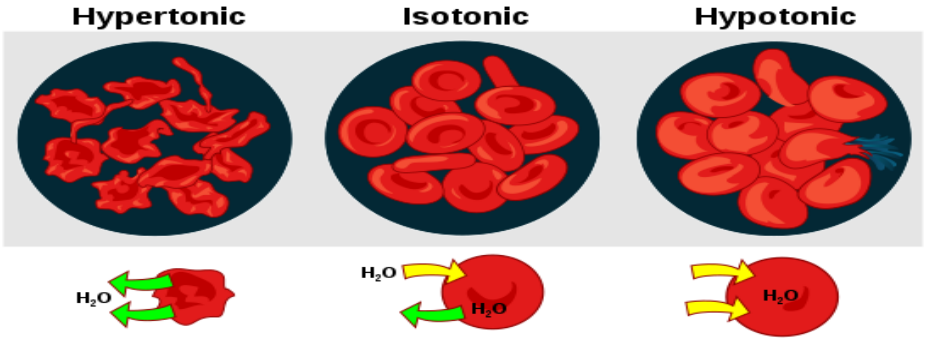


Figure 4 – Diagram of erythrocytes in hypertonic solution (cell shrinks), isotonic solution (cell remains unchanged) and hypotonic solution (cell swells and bursts)

In medical practice, the use of hypotonic solutions is limited to their use as solvents for water-soluble drugs.

Hypertonic solutions are applied in medicine:

• for the treatment of glaucoma;

• intravenously for pulmonary edema and internal bleeding;
• 10 % NaCl solution is used for healing purulent wounds;
• 2-5 % NaCl solutions are used for gastric lavage in case of poisoning with silver nitrate;
• 25 % $MgSO_4$ solution is used as a hypotensive agent.

Mainly electrolytes are responsible for high osmotic pressure of blood plasma and other biological fluids. Colloidal particles of proteins not able to pass through cell membranes also contribute into osmotic pressure of blood plasma, but their *oncotic pressure* is rather low: 3-4 kPa.

Significance of osmosis:
▪ influences the transport of nutrients and the release of metabolic waste products;
▪ It stabilizes the internal environment of a living organism by maintaining the balance between water and intercellular fluid levels;
▪ It maintains the turgidity of cells [1-3].

Problem 3. 0.63 M solution of HNO_3 is used in laboratory as an oxidizing agent. What is its osmotic pressure at t = 20 °C?

Solution:

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

HNO_3 is a strong electrolyte, hence, $\alpha(HNO_3) = 100$ % or 1.

$\alpha = \frac{i-1}{n-1}$ *(where n is a number of ions contained in a molecule);*

$HNO_3 \rightarrow H^+ + NO_3^-$; $i = \alpha \times (n-1) + 1 = 1 \times (2-1) + 1 = 2$

Hence, $\pi = 2 \times 8.31 \times (273 + 20) \times 0.63 = 3067.8$ kPa

Answer: $\pi = 3067.8$ *kPa*

The practical part
Safety instructions before laboratory work.

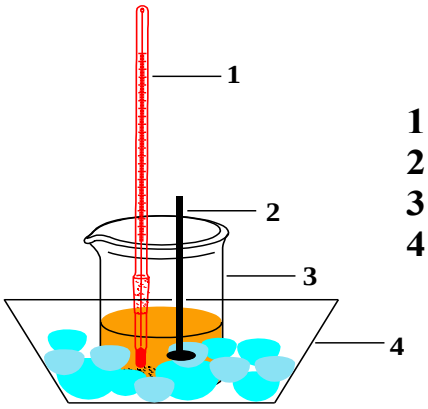
LABORATORY WORK

Determination of the molar mass of a medicinal substance (nonelectrolyte) by cryoscopic method

Equipment of the experiment: a crystallizer, a glass with a volume of 50-100 ml, measuring cylinders for 25 ml, a thermometer with scale divisions of 0.1-0.2 ^oC, a stirrer, a metal stand, a cooling mixture, a glucose solution.

The method of performing the experiment: the determination of the molar mass of a substance by lowering the freezing point of its solution in comparison with a pure solvent is performed using a cryoscope.

1. The scheme of a simplified cryoscope is shown in Figure 4.



1 - a thermometer
2 - a stirrer
3 - a glass with a solution
4 - a crystallizer

Figure 4 – Simplified cryoscope

Prepare a cooling mixture consisting of small pieces of ice and some crystals of table salt, and place it in a crystallizer. The temperature of the cooling mixture should be approximately –6 °C.

Pour 25 ml of a glucose solution with known mass percent of glucose (the density of the solution is considered equal to 1 g/ml) into a chemical glass and place the glass in the cooling mixture. Dip the thermometer into the solution and look after the temperature decreasing. Don't forget to mix the solution to avoid its local overcooling. Mark a temperature when the first ice crystals appear in the glass.

2. Write down the experimental data:

Volume of a glucose solution	V_{sln} = 25 ml
Percent by mass of glucose	ω_{gl} = 5 %
Mass of glucose	$m_{gl} = V_{sln} \times \rho_{sln} \times \omega_{gl}$ =
Mass of water	$m_{water} = m_{sln} - m_{gl}$ =
Freezing point of water	$t_{f\,(water)}$ = 0 °C
Freezing point of a solution	$t_{f\,(sln)}$ =
Freezing point depression	$\Delta t_f = t_{f\,(water)} - t_{f\,(sln)}$ =

3. Using cryoscopic Raoult's Law calculate glucose molar mass:

$$\Delta t_f = K_f \times \frac{m(C_6H_{12}O_6) \times 1000}{M(C_6H_{12}O_6) \times m(H_2O)};$$

$$M(C_6H_{12}O_6) = K_f \times \frac{m(C_6H_{12}O_6) \times 1000}{\Delta t_f \times m(H_2O)} =$$

The molar freezing point depression constant of water is 1.86 °·kg/mol.

4. Calculate the absolute and relative errors of the experiment, taking in account that the true molar mass of glucose is 180 g/mol.

$$D = |M_{gl}^{\,true} - M_{gl}^{\,exp}| =$$

$$D_0 = \frac{D}{M_{gl}^{\,true}} \times 100\% =$$

Control over the assimilation of the topic

It is conducted in the form of an oral conversation with students.

METHODOLOGICAL RECOMMENDATIONS FOR THE ORGANIZATION AND PERFORMING OF INDEPENDENT WORK OF STUDENT (IWS)

The time allotted for independent work can be used by students for:

– preparation for laboratory classes;
– taking notes of educational literature;
– performing tasks for self-control of knowledge;
– preparation of thematic reports, abstracts, presentations.

The main methods of organizing independent work:

– studying topics and problems that are not covered in the classroom;
– writing an abstract and making a presentation;
– performing tasks for self-control of knowledge.

List of tasks of IWS:

1. Vapor pressure of a water is 25 kPa (65 ºC). Calculate the vapor pressure of the solution at the same temperature, which contains 13.62 g of sugar in 90 g of water.

Answer: 24.8 kPa

2. How many grams of glucose should be dissolved in 270 g of water for:
a) reducing the freezing point of the solution by one degree;
b) increasing the boiling point of the solution by one degree?

Answer: 26.15 g; 93.54 g

3. How many degrees will the boiling point of the solution increase compared to the boiling point of water if 9 g of glucose $C_6H_{12}O_6$ is dissolved in 100 g of water?

Answer: at 0.26°

4. The isotonic coefficient of an aqueous solution of hydrochloric acid is 1.66 (ω = 6.8 %). Calculate the freezing point of this solution.

Answer: –6.15°C

5. At 25 °C, the osmotic pressure of a solution containing 2.8 g of high molecular compound in 200 ml of the solution is 0.70 kPa. Calculate the molar mass of the solute.

Answer: 49530 g/mol

6. Calculate the osmotic pressure of a solution containing 16 g of sucrose ($C_{12}H_{22}O_{11}$) and 350 g of water at 293 K (the density of the solution should be taken equal to 1 g/cm^3).

Answer: 311 kPa

The control of the IWS is carried out in the form of:

– evaluation of an oral answer to a question, message, report or presentation;
– individual conversation.

LIST OF SOURCES USED

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2. Основы химии для иностранных студентов = Essential chemistry for foreign students : учебно-методическое пособие / С. В. Ткачёв [и др.]. – 5-е изд. – Минск : БГМУ, 2018. – С. 83-88. – Режим доступа: http://rep.bsmu.by:8080/handle/BSMU/21054.

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