MARITIME UNIVERSITY OF SZCZECIN
MUS
POLAND

Institute of Mathematics, Physics and Chemistry

Department of Chemistry

Technical chemistry laboratory

Laboratory exercise

Solutions

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EXERCISE SHEET

<table>
<tr><td>1</td><td colspan="3">Relation to subjects: ESO/26, DiRMiUO/26, EOUNIE/26</td></tr>
<tr><td></td><td>Specialty/Subject</td><td>Learning outcomes for the subject</td><td>Detailed learning outcomes for the subject</td></tr>
<tr><td></td><td>ESO/ 25 Technical chemistry</td><td>EKP1
K_W01, K_W02, K_U05
EKP2
K_U08, K_U09</td><td>SEKP4 – Examination of physico-chemical properties of aqueous solutions, types of concentrations, the concept of solubility, the effect of temperature.
SEKP4 – Examination of electrolytic dissociation, writing dissociation equations, determining the dissociation constant and degree, influence of the dielectric constant and common-ion.</td></tr>
<tr><td></td><td>ESO/26 Chemistry of water, fuels and lubricants</td><td>EKP3
K_U014, K_U015, K_U016.</td><td>SEKP6 – Performing determinations of selected indicators of technical water quality;</td></tr>
<tr><td></td><td>DiRMiUO/26 Chemistry of water, fuels and lubricants</td><td>EKP3
K_U014, K_U015, K_U016.</td><td>SEKP6 – Performing determinations of selected indicators of technical water quality;</td></tr>
<tr><td></td><td>EOUNiE/26 Chemistry of water, fuels and lubricants</td><td>EKP3
K_U014, K_U015, K_U016.</td><td>SEKP6 – Performing determinations of selected indicators of technical water quality;</td></tr>
<tr><td>2</td><td colspan="3">Purpose of the exercise:
– mastering the knowledge of solutions and basic methods of expressing concentrations as well as ionic dissociation;
– consolidation of knowledge about the degree and constant of dissociation, strong and weak acids and bases;
– learning to write the dissociation equations of acids, bases and salts and the formula for the equilibrium constant K;
– mastering the basic mathematical relationships of concentration, constant K and degree of dissociation;
– experimental study of the influence of selected physicochemical factors on the course of the ion dissociation process in laboratory conditions.</td></tr>
<tr><td>3</td><td colspan="3">Prerequisites:
The student has a general chemical knowledge of the types of solutions, ways of expressing the concentration of solutions and general chemical knowledge of ionic dissociation (definition, constant and degree of dissociation).</td></tr>
<tr><td>4</td><td colspan="3">Description of the laboratory workplace:
Basic laboratory equipment - glass test tubes, water baths, micro-spatula for adding reagents, reagents: sodium chloride, potassium nitrate (V), calcium hydroxide in the form of a solution, ammonium nitrate (V), magnesium chloride, copper (II) chloride, acetone, acid acetic acid, methyl orange, sodium acetate, phenolphthalein, ammonium chloride.</td></tr>
</table>

5	**Risk assessment:** Contact with chemical reagents in solid form and diluted in liquid form – probability of chemical burn – very small, effects – minor Final assessment – **MEDIUM** **Security measures required:** a. protective gloves, b. protective glasses, c. lab coats.
6	**The course of the exercise:** 1. Getting to know the workplace manual (appendix 2), 2. Perform individual exercises in accordance with the workplace manual for exercises.
7	**Exercise report:** 1. Develop an exercise in accordance with the instructions contained in the workplace manual. 2. Solve the given task and/or answer the questions included in the set of tasks and questions to be completed by the student.
8	**Archiving of research results:** Submit the report on the exercise in the applicable form at the beginning of the next laboratory exercises
9	**Assessment method and criteria:** a. EKP1, EKP2 – control of the students' mastery of the basic chemical concepts of solutions, ways of expressing concentrations and concepts related to the ionic dissociation process will be carried out during the course, b. SEKP4 – the detailed learning outcome for an individual student will be assessed on the basis of the solutions to tasks and problems presented in the report, given for independent solution/development: – **mark 2,0** – the student has too little knowledge of solutions and electrolytic dissociation, but lacks the ability to solve simple tasks regarding solutions and ionic dissociation applications; – **mark 3,0** – has basic chemical knowledge of solutions and electrolytic dissociation, and has the ability to solve simple tasks (converting the concentration of solutions, distinguishes between strong and weak acids and bases, and is able to determine the pH of the solutions; – **mark 3,5 – 4,0** – has extended chemical knowledge in the field of solutions and electrolytic dissociation and has the ability to solve complex tasks in his specialty; – **mark 4,5 – 5,0** – has the ability to apply complex knowledge in the field of solutions and electrolytic dissociation to solve complex tasks and problems in his specialty.
10	**Literature:** 1. Stundis H., Trześniowski W., Żmijewska S.: *Ćwiczenia laboratoryjne z chemii nieorganicznej*. WSM, Szczecin 1995. 2. Kozłowski A., Ćwirko K., *Instrukcja stanowiskowa do ćwiczeń laboratoryjnych: Roztwory i dysocjacja jonowa*, AM Szczecin, 2013. 3. Cox P.A. translation of Z. Zawadzki: *Chemia nieorganiczna*. PWN. Warsaw 2006. 4. Drapała T.: *Chemia ogólna i nieorganiczna*. SGGW, Warsaw 1994. 5. Bielański A.: *Chemia ogólna i nieorganiczna*. PWN, Warsaw 1994. 6. Jones L., Atkins P., *Chemia ogólna*. Molecules, matter reactions, WN PWN, Warsaw 2004.

	7. Mastalerz P.: *Elementarna chemia nieorganiczna*. Wydawnictwo Chemiczne. Warsaw 2000. 8. Śliwa A.: *Obliczenia chemiczne. Zbiór zadań*. PWN. Warsaw 1994. 9. Pazdro M. *Zbiór zadań z chemii dla szkół średnich*. 10. Jones L., Atkins P., General chemistry. Molecules, matter reactions, WN PWN, Warsaw 2004. 11. Chemistry. Virtual textbook - basics and applications 12. Kozłowski A., *Materiały dydaktyczne z chemii technicznej*, developed for the purposes of auditorium classes (not published).	
11	Notes	

APPENDIX 1 – MANUAL

1. SCOPE OF THE EXERCISE

Issues and keywords:

- basic concepts of solutions (types of solutions, solvent, solute, solubility, saturated and unsaturated solutions, hydration, molar heat of dissolution);
- Le Chatelier-Braun principle (the equilibrium law);
- ways of expressing solution concentrations (percentage, molar, molar, normal, weight fraction, ppm concentration).

2. THEORETICAL INTRODUCTION TO THE EXERCISE

2.1. Solutions

2.1.1. Definition and basic concepts

A **solution** is a homogeneous mixture of solute and solvent. The **solvent** in the solution is that substance which is predominantly quantified. Solutions can exist in three states of aggregation, so they can be gaseous solutions (air, food storage solutions, mixtures used by deep sea divers), liquid and solid (metal, semi-metal and non-metal alloys, such as silicon with a small addition of phosphorus used in electronic industry, glass). Solid solutions are widely used. Liquid solutions are among the most common and of the greatest practical importance. They can be obtained by dissolving gases, other liquids or solids in a liquid – most often water.

Due to the particle size of substances dissolved in liquid solutions, these solutions are divided into:

– real solutions – particle size	– from 0.1 nm to 1 nm,
– colloids	– from 1 nm to 500 nm,
– suspensions	– bigger than 500 nm.

Dissolution processes

Dissolving any substance is possible only when the interaction between the solute and the solvent is strong enough and leads to the destruction of the crystal lattice. It is known from experience that substances with an ionic bond dissolve well in polar solvents (liquid ammonia, water), while the solubility of these substances in solvents with covalent bonds (benzene, carbon tetrachloride) does not occur or is minimal. In the case of polar solvents, their dipole molecules are attracted by electrostatic forces by the surface ions of the crystal lattice of a substance with ionic bonding, which leads to the weakening of this bond in the crystal and the removal of the ion from the lattice node with simultaneous hydration. This process is shown in Fig. 1.

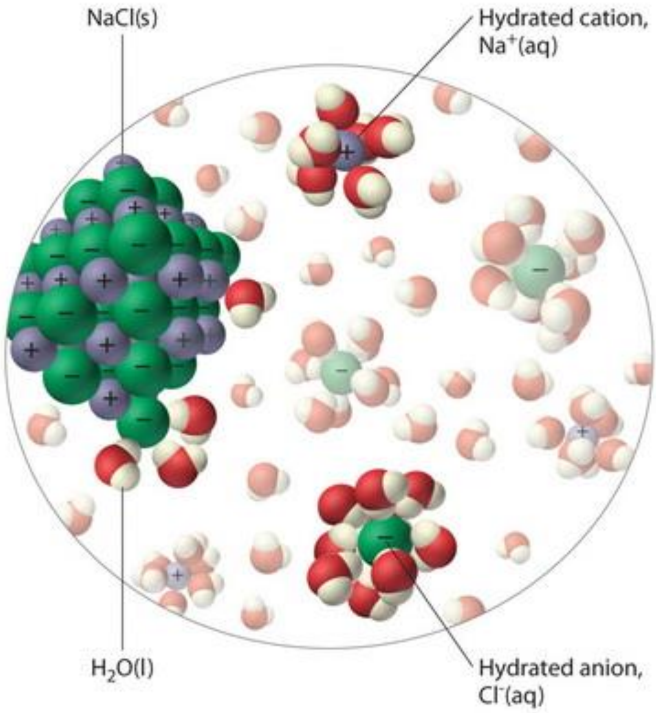

Fig. 1 Scheme of NaCl dissolution and formation of hydrated ions
source: https://socratic.org/questions/52f1151702bf34733dc14e7e)
(accessed, September, 29, 2021)

The effect of the solvent on the weakening of the bond in the crystal lattice is the greater the greater the value of the dielectric constant (ε) of the solvent. It is known from Coulomb's law that the force of mutual interaction of two charges (F) is

$$F = \frac{q_1 \cdot q_2}{r^2} \cdot \frac{1}{\varepsilon}$$

where:

q_1, q_2 – ion charges,
r – distance between charges,
ε – dielectric constant.

The value of the dielectric constant is determined from the ratio of the capacitance of the flat capacitor with the test substance between its plates to the capacitance of this capacitor. In the solutions depending on ε, the relationship was found that if:

$\varepsilon < 10$	dissociation does not occur,
$10 \leq \varepsilon \leq 40$	partial dissociation occurs,
$\varepsilon > 40$	complete dissociation occurs.

Due to its angular structure and the electronegativity difference between oxygen and hydrogen atoms, a water molecule has two distinct electric poles – we say it is a dipole. The measure of the spread of opposite charges is the dipole moment, which is expressed by the product of the absolute value of the charge of one of the poles and the distance. The unit of the dipole moment is coulomb times meter (C · m). Table – 1 shows the value of the dielectric constant of some solvents and the dipole moments of their molecules.

Due to the very high value of the dipole moment of water molecules, it facilitates the transition of the ions of the electrolyte crystal lattice into the solution and their hydration. The value of the bond strength in crystal networks for different compounds is different. For example, for NaF it is 916 kJ/mol, and for CsI 582,4 kJ/mol. Therefore, the solubility of CsI is almost ten times higher than that of NaF.

Table 1

Dipole moments (μ) and dielectric constant (ε) of some compounds

Substance	μ 10^{-30} [C · m]	ε	K
C_6H_6	0.00	2.28	293.0
CCl_4	0.00	2.24	298.0
CH_4 (ciekły)	0.00	1.70	100.0
NH_3	4.87	25.00	195.7
CH_3OH	5.67	32.63	298.0
C_2H_5OH	5.67	24.30	298.0
H_2O	6.13	78.54	298.0
H_2O (lód)	–	88.00	273.0
CH_3Cl	6.20	9.08	293.0
CH_3COOH	1.68	6.15	293.0
HCl	1.28	12.00	170.0
CH_3COCH_3	in benzene 2.76	20.74	298.0

High ion hydration is also conducive to increasing the solubility. This explains the greater solubility of $MgSO_4$ compared to $BaSO_4$, despite the lower binding strength in $BaSO_4$.

The dissolution processes are accompanied by **thermal effects**, which are most often the result of the hydration of the ions formed in the solution. Their size depends on the type and amount of the dissolved substance as well as the type and amount of solvent. If the dissolved substance has a crystalline structure, some of the heat of hydration is used to destroy the crystal lattice. Depending on which effect is greater – whether the energy of hydration or the destruction of the crystal lattice – the heat of dissolution will be either positive or negative. The negative heat of dissolution occurs when the energy of hydration is lower than the energy of the bonds in the crystal lattice. The solubility of such substances increases with temperature. If the heat of hydration predominates, then the solubility of such substances decreases with the temperature.

The molar heat of dissolution is the change in enthalpy accompanying the dissolution of 1 mole of substance molecules to form a very dilute solution.

Substances with covalent bonds do not dissolve in ionic solvents, but they readily dissolve in covalent solvents. There is a rule here that „**similar dissolves in similar**". We should remember this when selecting the appropriate solvent to dissolve the substance in question.

Dissolving the substance is a reversible process, so by evaporating the solvent the solute can be recovered.

When dissolving a substance, the concentration of the solution gradually increases until a certain limit is reached - the concentration of a saturated solution. A solution saturated at a constant temperature has a constant composition and is characterized by the fact that it is in equilibrium with even a minimal amount of substance dissolved in the solid phase that is in contact with the solution. The values of the heat of dissolution (Q_r) of selected salts in water are summarized in the Table 2.

Table 2

The heat of dissolution (Q_r) of salt in water

Compound	Q_r [kJ/mol]	Compound	Q_r [kJ/mol]
$MgCl_2$	– 151.90	$Na_2SO_4 \cdot 10\,H_2O$	+ 78.44
$MgCl_2 \cdot 6H_2O$	– 14.20	NH_4NO_3	+ 26.47
NaCl	+ 5.02	CH_3COONa	– 16.20
Na_2CO_3	– 23.31	$CH_3COONa \cdot 3H_2O$	+ 19.20

The solubility of most compounds generally increases with temperature (Fig. 2), although this is not necessarily the case (Table 3).

The **solubility of a substance** is the mass in grams of a substance that is dissolved in 100 grams of solvent (usually at 293 K).

The solubility of a substance depends on the type and size of interactions between solute molecules and solvent molecules. As the temperature rises, the solubility of most solids increases and the solubility of gases decreases. The solubility of gases in liquids increases with increasing pressure. Figure 2 shows examples of the dependence of solubility on temperature.

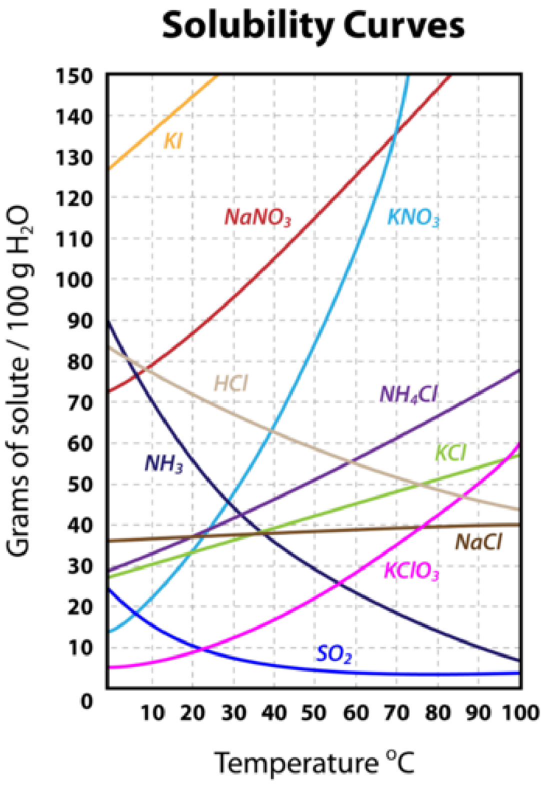

Fig. 2. Solubility curves for selected chemical compounds
source: https://courses.lumenlearning.com/cheminter/chapter/how-temperature-influences-solubility/
(accessed September, 29, 2021)

A **saturated solution** is a solution containing the greatest amount of a solute at a certain temperature (e.g. 20ºC) which is in equilibrium with that substance remaining in the solid phase. The points on the curve correspond to the concentrations of saturated solutions at a given temperature. Each point of the field below the solubility curve corresponds to an **unsaturated solution**, above – to a **supersaturated solution**. Supersaturated solutions are unstable and are usually obtained by supercooling saturated solutions. If a small crystal of the solute is introduced into such a solution, an excess of the solute will be released, exceeding that of the proper saturated solution.

Table 3

The solubility of the substance [g/100g] of the solvent

Compound	Temperature [K]			
	273 (0°C)	293 (20°C)	333 (60°C)	373 (70°C)
NaCl	35.7	36.0	37.3	39.8
KNO_3	13.3	20.9	85.5	202.0
$Ca(OH)_2$	0.185	0.165	0.116	0.077
CH_3COONa	36.3	46.5	139.0	170.3
KCl	27.6	34.0	45.5	56.7

2.1.2. Ways of expressing the concentration of a solution

1. The **percent concentration** of the solution denoted as $C_{\%}$ (expressed as a percentage by mass) is the number of grams of the dissolved substance in 100 g of the solution.

$$C_{\%} = \frac{m_s}{m_r} \cdot 100\%$$

where:

m_s – mass of solute in [g],

m_r – mass of the solution, which is the sum of the mass of the solute and the mass of the solvent, e.g. water, in [g].

The percent concentration of the solution can be expressed as a percentage by volume and it means the number of cubic centimetres of the substance dissolved in 100 cm^3 of the solution.

$$C_{\%} = \frac{V_s}{V_r} \cdot 100\%$$

where:

V_s – volume of solute in [cm^3];

V_r – the volume of the solution, which is made up of the sum of the volumes of the solute and solvent in [cm^3].

2. The **molar concentration** of the solution C_m is the number of moles of the dissolved substance in 1 dm^3 of the solution and is expressed in [mol/dm^3].

$$C_m = \frac{n_s}{V_r}$$

where:

n_s – number of moles of solute in [mole],

V_r – solution volume expressed in [dm^3].

3. The **molality** of the solution is the number of moles of the dissolved substance in 1 kg of the solvent and is expressed in [mol/kg].

$$C_L = \frac{n_s}{m_r}$$

where:

n_s – number of moles of solute in [mole],
m_r – mass of the solvent expressed in [kg].

Practical molality is applicable only for calculations related to Raoult's law (colligative properties of solutions).

4. The **mole fraction** is the concentration of the solution, expressed as the ratio of the number of moles of a given component to the sum of the number of moles of all components of the solution.

$$X_1 = \frac{n_1}{n_1 + n_2 + \ldots + n_n} \left[\frac{\text{mol}}{\text{mol}}\right]$$

where:

n_1 – number of moles of the first component in [mole],
$n_1 + \ldots + n_n$ – the number of moles of all components.

$$X_1 + X_2 + \ldots\ldots + X_n = 1$$

Similarly, we express the **weight and volume fraction** of a given component and express them in [kg/kg] and [m^3/m^3] respectively.

5. The **ppm concentration** is a frequently used method of expressing the concentration (from *parts per million*), which is the number of grams (or cm^3) of a substance in 1 000 000 grams (or m^3 of solution), respectively, depending on whether it is a so-called „ppm by weight” or „ppm by volume”.

If we have given ppm by weight, then: 1 ppm by weight = 1 μg/g or 1mg/kg, and if ppm by volume: 1 ppm by volume = 1 $\mu l/dm^3$ = 1 cm^3/m^3.
You can easily convert % to ppm and vice versa:

$$1\% = \frac{1}{100} = \frac{10^4}{10^6} = 10\,000 \text{ ppm}$$

$$1 \text{ ppm} = \frac{1}{10^6} = \frac{10^{-4}}{10^2} = 10^{-4}\% = 0{,}0001\%$$

2.2. Le Chatelier-Braun principle (the equilibrium law)

The Le Chatelier-Braun principle - says that if a stimulus acts on a system in dynamic equilibrium, changes occur in this system to counteract the stimulus. It allows to predict the behaviour of a given system under the influence of the considered stimulus, which may be, for example, a change in temperature, concentration, pressure. In quantitative terms, this rule is expressed by the law of mass action of Guldberg and Waage.

Example 1

Assess how the amount of sulphur oxide (VI) formed in the reaction:

$$2SO_2\,(g) + O_2\,(g) \rightleftharpoons 2SO_3\,(g) \qquad \Delta H_0 < 0$$

affects:

a) temperature increase,
b) pressure reduction,
c) use of a catalyst.

Solution: Since the reaction is exothermic ($\Delta H<0$), we can write the reaction as:

$$2SO_2\,(g) + O_2\,(g) \rightleftharpoons 2SO_3\,(g) + Q$$

The heat effect **Q** is on the product side, because exo-energetic reactions are characterized by the fact that heat is released into the environment.

a) Increasing the temperature will increase our Q. This means that there will be more products (because Q is on the products side). The system, in order to inhibit the action of the stimulus, will therefore shift the equilibrium towards the substrates (**to the left**), and thus the **amount of SO_3 generated will be reduced**.
b) The effect of the pressure change applies only to the reactions with the reactants and products in the gas phase. In the analysed reaction, on the side of the substrates we have **3 moles** of gases, and on the side of products only **2 moles** of gases. So, by lowering the pressure, we reduce the amount of reactants (remember that the change in pressure applies to the side of the reaction equation where the amount of gas is greater). The system, acting against it, will shift the equilibrium of the reaction towards the reactants (**to the left**), hence **the amount of SO_3 generated will decrease**.
c) The **catalyst** does not shift the equilibrium of the reaction, and thus the amount of SO_3 formed neither increases nor decreases, but causes a significant increase in the reaction rate.

Example 2

The equilibrium of the reaction is expressed by the equation:

$$2CO\,(g) + O_2\,(g) \rightleftharpoons 2CO_2\,(g) \qquad \Delta H = -566\ \text{kJ}$$

How will the equilibrium of the reaction shift (right or left) if we make the following changes to the system?

a) we will raise the temperature,
b) we will remove some of the CO_2,
c) we will increase the pressure,
d) we will add O_2.

Solution: As in the first example, **$\Delta H < 0$**, therefore we will add **Q** on the product side

$$2CO\,(g) + O_2\,(g) \rightleftharpoons 2\,CO_2\,(g) + Q$$

a) by increasing the temperature, the system will protect itself against heat increase and the reaction equilibrium shifts to the left (against our will), because the system shifts the equilibrium towards the substrates, where no heat is released;
b) by removing CO_2 we reduce the amount of products, so the system – wanting to inhibit the loss of CO_2 – will shift the equilibrium to the **right** to counteract the loss of CO_2;
c) by increasing the pressure – the number of substrates will increase (again there are more moles of gases) – the system will want to counteract the increase in pressure, so the equilibrium will shift to the **right** (there are fewer gas molecules there);
d) by adding O_2, we will thus increase the number of substrates – the system will counteract the increase in O_2, concentration, shifting the equilibrium towards the products, i.e. to the **right**.

3. PERFORMING THE EXERCISE

Experiment 1 – Preparation of a saturated potassium chloride (KCl) solution

Materials and reagents:

Rack with test tubes, measuring cylinder, micro-spatula, mortar, water bath, solids: potassium chloride (KCl).

Performance:

Dissolve about 1.5 g of potassium chloride (KCl) in 5 cm^3 of water at room temperature. To speed up the dissolving process, first crush the salt in a mortar, then shake the test tube after adding water. After complete dissolution, add 1 micro spatula (about 0.2 g) of the same salt, each time dissolve the salt until the solution is saturated.

Elaboration of the results:

1. Describe the course of the experiment, define saturated and unsaturated solutions.

Experiment 2 – Preparation of a supersaturated solution

Materials and reagents:

Test tube rack, measuring cylinder, micro spatula, water bath, solid: sodium acetate (CH_3COONa).

Performance:

Using a funnel, fill a dry test tube to about 3/4 of its height with sodium acetate (CH_3COONa), then add 5 cm^3 of distilled water to it. Heat the test tube in a water bath until the salt is completely dissolved. Then remove carefully and cool the bottom of it well in a stream of running water (do not cool the top layer of the solution). After cooling the solution, throw salt crystals (for example NaCl) into it and observe.

Elaboration of the results:

1. Describe the course of the experiment, give the definition and methods of obtaining saturated and supersaturated solutions.
2. What is the role of the salt crystal added to the cooled solution?
3. What is the use of crystallization in technology?

Experiment 3 – The influence of temperature on the solubility of substances

Materials and reagents:

Rack with test tubes, measuring cylinder, micro spatula, water bath, solids: sodium chloride ($NaCl$), potassium nitrate(V) (KNO_3), solutions: calcium hydroxide ($Ca(OH)_2$).

Performance:

Prepare approx. 5 cm^3 of a saturated solution of sodium chloride ($NaCl$) and potassium nitrate(V) (KNO_3). Heat a test tube with a saturated solution of sodium chloride ($NaCl$) in a water bath to a temperature of about 313K and, stirring, add salt in small portions (a pair of crystals) until a solution is obtained with a small amount of solid phase at the bottom at a given temperature. Then heat the solution to about 363K and again add salt in portions until saturation. Carry out the experiment analogously for the solution of potassium nitrate(V) (KNO_3). Pour the solution of calcium hydroxide ($Ca(OH)_2$) saturated at room temperature into a test tube (approx. 5 cm^3) and heat to boiling. Observe the occurring phenomenon.

Elaboration of the results:

2. Explain how an increase in temperature affects the solubility of individual salts.
3. Present on a graph the dependence of the solubility of the tested salts on the temperature.

Experiment 4 – Influence of temperature on solubility of substances – enthalpy of dissolution

Materials and reagents:

Rack with test tubes, measuring cylinder, micro spatula, water bath, solids: magnesium chloride ($MgCl_2$), ammonium nitrate(V) (NH_4NO_3).

Performance:

In two test tubes, prepare approx. 5 cm^3 of a saturated solution of ammonium nitrate(V) (NH_4NO_3) and magnesium chloride ($MgCl_2$). Compare the thermal effects in both test tubes while dissolving the salt. Then add 2 micro spatulas of ammonium nitrate(V) (NH_4NO_3) into the first test tube, and 2 micro spatulas of magnesium chloride ($MgCl_2$) into the second test tube. Gradually heat both test tubes in a water bath. Observe the salt dissolving effect in both cases.

Elaboration of the results:

1. Explain the occurring phenomenon with the Le Chatelier-Braun principle.
2. Plot the temperature dependence of the test salt solubility.
3. Read the molar heat of dissolution of the tested salts in the table, identify the exothermic and endothermic reaction and explain what this reaction is characterized by.

4. DEVELOPMENT OF THE EXERCISES

1. Prepare a report according to the guidelines in the experimental section.
2. Place the cover sheet as the first page of the report.
3. After the theoretical part has been concisely developed, include in the report the study of individual experiments and the solved additional task/tasks given by the academic teacher.

5. THE FORM AND CONDITIONS FOR PASSING THE LABORATORY EXERCISE

1. Passing the so-called „entry test” before starting the exercise.
2. Submission of a correct written laboratory report on the performed exercise in accordance with the guidelines for the preparation of the laboratory report, please see the link below: https://www.am.szczecin.pl/en/facilities/institute-of-mathematics-physics-and-chemistry/department-of-chemistry/technical-chemistry/tech-chemistry-lab-manuals/

I. Examples of tasks with a solution

Example 1

In 250 cm^3 of nitric acid solution there are 25 g of HNO_3. Calculate the molar concentration of this solution.

Data:

$V_{sol.}$ = 250 cm^3 = 0.25dm^3 (because in the definition the number of moles refers to 1dm^3),
m_s = 25 g.

Wanted:

c_m = ?

Solution:

Step I

We calculate the molar mass of HNO_3, because this value is necessary to convert the mass of the acid to the number of moles, and regardless of whether we use the formula for concentration or separately, we will convert the mass of acid to the number of moles.

$M_{\text{of acid}}$ = 63 g/mol.

Step II

Using the molar concentration formula we calculate the concentration of the acid solution.

$$c_m = \frac{m_s}{M \cdot V} = \frac{25\text{g}}{63\frac{\text{g}}{\text{mol}} \cdot 0{,}25\text{dm}^3} = 1{,}59\frac{\text{mol}}{\text{dm}^3}$$

Example 2

647 cm^3 of an unknown concentration of sulfuric acid solution was mixed with water to obtain 1.5 kg of a 55% solution. Calculate the molar concentration of the starting acid solution.

Data:

$V_{r\text{-}ru}$ = 647 cm^3,
m_r = 1.5 kg,
c_p = 55%.

Wanted:

c_m = ?

Solution:

Step I

Having the mass and percent concentration of the solution after dilution, we will calculate the mass of sulphuric acid.

1500 g of solution – is 100%
m_s g – is 55% $m_s = 825$ g

Step II

To calculate the molar concentration, it is necessary to know the molar mass (M) of sulfuric acid M = 98 g/mol.

Step III

Since we know the volume of the solution (V), the mass of the solute (m_s) and the molar mass of the acid (M), we can calculate the molar concentration c_m.

$$c_m = \frac{m_s}{M \cdot V} = \frac{825\mathrm{g}}{98\frac{\mathrm{g}}{\mathrm{mol}} \cdot 0{,}647\mathrm{dm}^3} = 13\frac{\mathrm{mol}}{\mathrm{dm}^3}$$

Note:

Such a value of the molar concentration indicates that strong sulfuric acid is very slightly volatile and therefore it can displace other strong acids from salt solutions, such as nitric acid (V) or hydrochloric acid.

Example 3

The solubility of a certain substance at the temperature of 363K is 42g, and at the temperature of 293K it is 12g. Calculate how much of the substance crystallizes out of 250g of a saturated solution after lowering the temperature from 363K to 293K.

Solution:

Data:

R_1 = 42 g,
R_2 = 12 g.

Wanted:

m_s = ?

If the solubility at the temperature of 363K is 42 g, it means that in 100 g of the solvent 42 g of the substance are dissolved, giving 142 g of the solution. We can write that at 363K we have:
142 grams of a solution in which there are 42 grams of the substance and in 250 grams of a solution there are x grams of the substance, i.e.:

$$x = 250\ \mathrm{g} \cdot 42\ \mathrm{g}/142\ \mathrm{g} = 73.9\ \mathrm{g}$$

That is, 250 g of the solution consists of 73.9 g of the substance and 250 g – 73.9 = 176.1 g of the solvent.

After cooling the solution, some of the substance will crystallize out, but there will still be 176.1 g of solvent in the solution. For a temperature of 293K, we can write: in 100 g of solvent, 12 g of a substance are dissolved, then in 176.1 g of solvent, x g of a substance is dissolved, i.e. $x = 176.1\ \mathrm{g} \cdot 12\mathrm{g}/100\ \mathrm{g} = 21.1$ g.

If there was originally 73.9 g of substance in the solution, and 21.1 g of substance remained, then 73.9 g crystallized out – 21.1 g = 52.8 g of substance.

Example 4

Calculate the concentration of Ni^{2+} ions in ppm units in a solution containing 0.100 g $NiSO_4$ in a volume of 1000 cm^3. The molar mass of Ni is 58.7 g/mol and the molar mass of $NiSO_4$ is 154.7 g/mol.

Solution:

First, we calculate the mass of Ni^{2+} in this solution:

$$m(\mathrm{Ni}^{2+}) = 0{,}100 \cdot 58.7/154.7 = 0.0379\ \mathrm{g} = 37.9\ \mathrm{mg} = 37.9 \cdot 10^{-6}\ \mathrm{kg}$$

We assume that the diluted solution has a mass numerically equal to its volume (the density equals 1 kg/ dm^3). We then use the formula:

$$1\,\mathrm{ppm} = \frac{1\mu\mathrm{g}}{1\mathrm{g}} = \frac{1\mathrm{mg}}{1\mathrm{kg}} = u_m \cdot 10^6$$

where:

u_m – stands for mass fraction.

Then

$$c(\mathrm{ppm}) = \frac{m(\mathrm{Ni}^{2+})[\mathrm{mg}]}{m(\mathrm{roztw.})[\mathrm{kg}]} = u_m(\mathrm{Ni}^{2+}) \cdot 10^6 = \frac{37{,}9 \cdot 10^{-6}}{1{,}00} 10^6 = 37{,}9\mathrm{ppm}$$

II. Tasks and questions to be completed by the student

1. Calculation tasks:

1. Some water was added to 20 g of solid NaOH and the concentration of the solution obtained was 5% (m/m). How many grams of water have been added?
2. Calculate the percent concentration of HNO_3 with a density of 1.08 g/ml, which in 250 ml of solution contains 37.8 g of HNO_3.
3. What volume of 15% (m/m) NaCl solution with a density of 1.14 g/ml contains 30 g of NaCl?
4. How many grams of Na_2CO_3 need to be weighed to make 100 ml of a solution with concentration of 0.1 mol/l?
5. Calculate the solubility if the concentration of the saturated solution is 25%.
6. The solubility of potassium iodide at 293K is 145g. How many grams of this salt can be dissolved in 0.5 dm^3 in water at a temperature of 293K.
7. What volume of 0.2M NaOH solution is required to neutralize 0.5 dm^3 of a 0.05 Molo solution of H_2SO_4?
8. 300 g of water was added to 50 g of a 30% $BaCl_2$ solution. Calculate the percentage of the solution obtained.
9. How much NaCl should be added to 250 grams of 15% NaCl solution to obtain a 25% solution?
10. How much water should be added to 25 cm^3 of a 27% solution of KOH with a density of 1.25 g/cm^3 to obtain a 0.2 molar solution?
11. What is the concentration % of a solution whose concentration is 12 ppm?
12. The concentration of a certain solution is 0.05%, express this concentration in ppm.
13. Express the concentration of 10 ppm in the following units: mg/kg, µg/g and %.

1.1. Complex tasks:

1. Prepare 500 ml of a 0.2 mol/l iron (II) sulphate (VI) solution. How many grams of technical $FeSO_4 \cdot 7H_2O$ containing 2% impurities should be weighed?
2. How many grams of solid NaI should be added to 200 ml of NaI solution at a concentration of 4% (m/m) and a density of 1.03 g/ml to obtain a solution at a concentration of 15% (m/m)?
3. What weight of Na_2CO_3 was prepared to check the titre of 0.1257 mol/l HCl solution, if 17.8 ml of this acid were used for its titration against methyl orange?