MARITIME UNIVERSITY OF SZCZECIN

Institute of Mathematics, Physics and Chemistry
Department of Chemistry

EXERCISE INSTRUCTION

Laboratory Exercise 2

Solutions and dissociation of electrolytes

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EXERCISE SHEET

1	**Relation to subjects**: ESO/26, DiRMiUO/26, EOUNIE/26		
	Specialty/Subject	**Learning outcomes for the subject**	**Detailed learning outcomes for the subject**
	ESO/ 25 Technical chemistry	EKP1 K_W01, K_W02, K_U05 EKP2 K_U08, K_U09	SEKP4- examination of physicochemical 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.
	ESO/26 Chemistry of water, fuels and lubricants	EKP3 K_U014, K_U015, K_U016.	SEKP6 – Performing determinations of selected indicators of technical water quality;
	DiRMiUO/26 Chemistry of water, fuels and lubricants	EKP3 K_U014, K_U015, K_U016.	SEKP6 – Performing determinations of selected indicators of technical water quality;
	EOUNiE/26 Chemistry of water, fuels and lubricants	EKP3 K_U014, K_U015, K_U016.	SEKP6 – Performing determinations of selected indicators of technical water quality;
2	**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.		
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).		
4	**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.		

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. Performing individual exercises in accordance with the workplace instruction 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:** Report on exercises - prepared in accordance with the rules applicable in the laboratory - should be submitted in writing to the academic teacher during the next classes
9	**Assessment method and criteria:** a. EKP1, EKP2 – checking 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	**References** 1. https://assets.openstax.org/oscms-prodcms/media/documents/Chemistry2e-WEB.pdf (accessed 15.07.22). 2. A. Kozłowski, A. Kalbarczyk-Jedynak, M. Ślączka-Wilk, K. Ćwirko, C. Wiznerowicz, G. Gorzycka, Instrukcje stanowiskowe do ćwiczeń laboratoryjnych: Roztwory i dysocjacja jonowa, AM Szczecin 2022 (in Polish). 3. J. E. McMurry, R. C. Fay, J. K. Robinson, Chemistry, 7th edition, global edition, publisher: Pearson, 2016.

	4. A. Blackman, S. Bottle, S. Schmid, M. Mocerino, U. Wille, Chemistry, 2nd edition, publisher: John Wiley&Sons, 2012. 5. G. Curran, Chemistry, publisher: The Career Press, 2011. 6. J. T. Moore, Chemistry for Dummies, publisher: Wiley Publishing, 2015. 7. D. Kealy, P.J. Haines, Analytical Chemistry, publisher: BIOS Scientific Publishers Limited, 2002. 8. Sparkcharts Chemistry, 2002 Spark Publishing, A Division of Barnes & Noble, Canada 2014. 9. M. D. Jackson, Chemistry, 2015 BarCharts, Inc. (Quickstudy.com). 10. M. Charmas, English for Students of Chemistry, Maria Curie-Skłodowska University Press, Lublin 2012. 11. Stundis H., Trześniowski W., Żmijewska S.: Ćwiczenia laboratoryjne z chemii nieorganicznej. WSM, Szczecin 1995 (in Polish). 12. M. Wesołowski, K. Szefer, D. Zimna, Zbiór zadań z analizy chemicznej, Wydawnictwa Naukowo – Techniczne, Warszawa 1997 (in Polish).	
11	Notes	

1. THEORY

KEYWORDS:

- solutions,
- chemical equilibrium,
- acids and bases.

Solutions

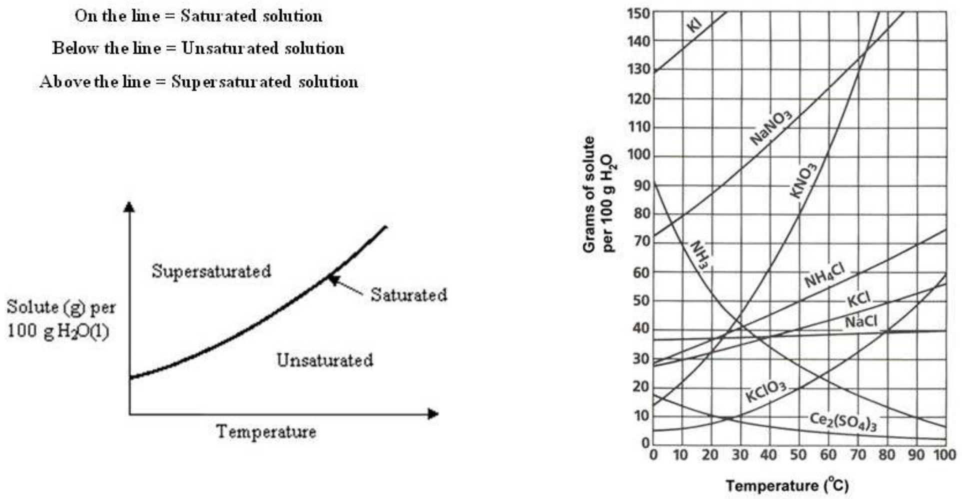
Solution – a homogeneous mixture of more than one substance in one phase. A solution is composed of a solvent and one or more solutes.

Solvent – major component of a solution (it is present in the largest amount).

Solute – minor component of a solution (it is present in the lesser amount).

There are solutions of liquids (for example salt in water, alcohol in water, oxygen dissolved in lake), gases (our atmosphere: nitrogen is considered the solvent, and the oxygen, carbon dioxide, other gases are considered the solutes), solids (for example alloys – solutions of one metal in another metal – brass – solution of zinc and copper).

Solubility – is the ability of a substance to dissolve in another substance also is the maximum amount of solute that will dissolve in a given amount of a solvent at a specified temperature [g/100cm^3], see the solubility graph shown below (Fig. 1):



Source:http://www.kentchemistry.com/links/Kinetics/SolubilityCurves.htm

Source:https://sites.prairiesouth.ca/legacy/chemistry//chem30/graphics/4_graphics/sol_curve.jpg

Fig. 1. The solubility graph (source: http://www.kentchemistry.com/links/Kinetics/SolubilityCurves.htm) (accessed 19.07.22)

Solubility [g/100g H_2O] of selected chemical compounds:

Chemical Compound	273 K (0°C)	293 K (20°C)	333 K (60°C)	373 K (100°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

Substance is insoluble when it does not dissolve in another substance.

A saturated solution – is a solution that contains the maximum amount of dissolved solute at a given temperature.

An unsaturated solution – is a solution that has less than the maximum amount of dissolved solute at a given temperature.

A supersaturated solution – is a solution that is holding more dissolve solute (more than its maximum amount) at a given temperature.

General rule of solubility: „like-dissolves-like”: polar solvents (for example water, ethanol) dissolve polar solutes; non-polar solvents (for example benzene, dichloromethane) dissolve non-polar solutes. Solubility is also related to the temperature. Normally solubility increases with increasing temperature (example: solids dissolving in liquids) in case of gases dissolving in liquids, for example oxygen dissolving in river/lake/sea/ocean, the solubility goes down with increasing temperature which affects the aquatic life.

Solution Concentration, see the table below:

WAYS TO MEASURE THE CONCENTRATION OF A SOLUTION	FORMULA	SOLUTION CONCENTRATION UNITS
PERCENT COMPOSITION: **MASS/MASS PERCENTAGE**	m/m% = grams of solute/grams of solution x 100% grams of solution = grams of solute [g] + grams of solvent [g] grams of solution = volume of solution [cm^3, dm^3, ml, l] x density of solution [g/cm^3; g/ml; kg/m^3]	[%]
PERCENT COMPOSITION: **MASS/VOLUME PERCENTAGE**	m/v%=grams of solute/milliliters of solution x 100%	[%]
PERCENT COMPOSITION: **VOLUME/VOLUME PERCENTAGE**	v/v%=milliliters of solute/milliliters of solution x 100% volume of solution = volume of solute [ml; cm^3] + volume of solvent [ml; cm^3]	[%]
MOLARITY	M = moles of solute / liter of solution; $M = n/V$ or $C = n/V$ n – moles of solute; V – volume of solvent; C – concentration number of moles (n) = mass of sample / molar mass	[mol/l]; [mol/dm^3]
MOLALITY	m = mol of solute / kilogram of solvent	[mol/kg]
MOLE FRACTION	x = mole of solute (mole of one component) / total moles (total moles in the mixture)	

Examples of the pollution units: parts per million (ppm), parts per billion (ppb):

ppm; ppb

ppm (parts per million) = $1 \cdot 10^{-6} = 1 \cdot 10^{-4}$ % = mg/kg/l = µg/g/ml
ppb (parts per billion) = $1 \cdot 10^{-9} = 1 \cdot 10^{-7}$% = µg/kg/l = ng/g/ml

kg – kilogram, mg-milligram, µg – microgram, ng – nanogram
ppm (V/V) = $1 \cdot 10^{-6}$(V/V) = $1 \cdot 10^{-4}$ %(V/V) = µL/L = nL/mL
ppb (V/V) = $1 \cdot 10^{-9}$(V/V) = $1 \cdot 10^{-7}$%(V/V) = nL/L = pL/mL

l – liter, ml – milliliter, µL – microliter, nL – nanoliter, pL – picoliter

Unit conversions

mg to µg = · 1000

ppm to mg/l or mg/kg = equivalent
% to ppm = · 10,000
1 l = 1000 ml
1 dm^3 = 1000 cm^3
1 l = 1 dm^3
1 ml = 1 cm^3
°C to Kelvin = + 273

Chemical Equilibrium

Equilibrium – typical for reversible reaction – two opposite reactions (forward and reverse) are taking place at the same time in the same place and with equal rates. Represented in the chemical balanced equation by $\rightleftarrows$:

$a\,A + b\,B \rightleftarrows c\,C + d\,D$ at equilibrium, the process is described by constant (K_c; equilibrium constant):

$K_c = [C]^c \cdot [D]^d / [A]^a \cdot [B]^b$ [ ] refers to molarity ($M = mol/l$; mol/dm^3)

The products (the right side of the arrow) always appear on the top line of the expression (the numerator), and the reactants (the left side of the arrow) on the bottom line of the expression (the denominator)

It is said that chemical equilibrium is dynamic, it means that reactants and products are continually being formed and re-formed.

Le Chatelier's principle:

When a system at dynamic equilibrium is perturbed (disturbed, changed), the system will adjust itself in order to minimize the effects of the perturbation (the effect of changes).

Summary of Le Chatelier's principle:

Change	Equilibrium shift	Effect on eqilibrium constant
Concentration	Adding reactant/removing product – equilibrium shifts in the forward direction. Removing reactant/adding product – equilibrium shifts in the reverse direction.	No change
Temperature	Increasing temperature – equilibrium shifts in the forward direction for endothermic reactions, in the reverse direction for exothermic reactions. Decreasing temperature – equilibrium shifts in the reverse direction for endothermic reactions, in the forward direction for exothermic reactions.	Change
Catalysts	No shift. Catalyst reduces the time of reaction. It does not affect the composition of the equilibrium mixture.	No change
Pressure	Increasing pressure – equilibrium shifts toward the side of the reaction with fewer moles of gas. Decreasing pressure – equilibrium shifts toward the side of the reaction with more moles of gas.	No change

Enthalpy (H) is the measure of energy that can be released as heat; ΔH – reaction enthalpy (the enthalpy change).

When $\Delta H < 0$, the reaction is exothermic (heat is released); when $\Delta H > 0$, the reaction is endothermic (heat is consumed).

Acids and Bases

Common properties and characteristics of acids and bases:

ACID	**BASE**
is a an electrolyte (electrolyte dissolves in water to form a solution that conducts electricity), tastes sour, burn the skin, reacts with certain metals in order to produce hydrogen gas, reacts with carbonates to produce carbon dioxide, causes pH indicators to change color (for example an acid turns a litmus paper red), reacts with bases to produce a salt and water, acidic solutions have a pH value of less than 7.	is an electrolyte, tastes bitter, feels slippery on the skin, reacts with acids to produce a salt and water, causes pH indicators to change color (it turns litmus paper blue), basic solutions have a pH value of greater than 7.

Binary acids are acids containing hydrogen (H) and only one other element, generally a nonmetal. **Oxoacids** contain hydrogen, oxygen and some other element, usually a nonmetal.

Electrolytes (acids, bases, salts) – substances that conduct electricity when dissolved in water or conduct electricity in the molten state. Second definition: electrolytes are substances that break into ions (ELECTROLYTIC DISSOCIATION) when dissolved in water. The ions positively charged are called cations, the ions negatively charged are called anions. Dissociation of electrolytes depends on dielectric constant of the solvent (ε) that is a measure of the polar nature of the solvent:

$\varepsilon < 10$	no dissociation (ionization),
$10 \leq \varepsilon \leq 40$	partial dissociation (ionization),
$\varepsilon > 40$	complete dissociation (ionization).

SELECTED SOLVENT	DIELECTRIC CONSTANT (ε)
H_2O	78.54
CH_3COCH_3	20.74
C_2H_5OH	24.30
C_6H_6	2.28

Nonelectrolytes (for example: fats, sugars, alcohols) – substances that do not conduct electricity when dissolved in water. These substances do not ionize – do not form ions.

Summary of the two main acids and bases theories:

THE ARRHENIUS THEORY (the first acid-base theory proposed by a Swedish chemist in 1887)	THE BRONSTED – LOWRY THEORY (proposed by the Danish chemist Bronsted and English chemist Lowry, independently in 1923)
Acid – is a substance that releases hydrogen proton H^+ ions in aqueous solution, for example: $HCl \longrightarrow H^+ + Cl^-$ simplified version $HCl\,(g) + H_2O(l) \longrightarrow Cl^-(aq) + H_3O^+\,(aq)$ more accurate version H_3O^+ is called the hydronium ion (proton H+ bonds to the water molecule and forms hydronium ion). Base – is a substance that releases OH^- (hydroxide) ions in aqueous solution, for example: $NaOH \longrightarrow Na^+ + OH^-$ Arrhenius also classified the reaction between an acid and a base as a neutralization reaction: acid + base $\longrightarrow$ salt + water, for example: $HCl + NaOH \longrightarrow H_2O + NaCl$ $NaCl \longrightarrow Na^+ + Cl^-$	Acid – is a proton (H^+) donor Base – is a proton (H^+) acceptor According to Bronsted – Lowry theory, acid-base reaction is a competition for a proton, for example: $NH_3 + H_2O \rightleftarrows NH_4^+ + OH^-$, where ammonia is a base because it accepts the proton, and water is an acid because it donates the proton in the forward reaction. In the reverse reaction, the ammonium ion is an acid, the hydroxide ion is a base.

Strong acids and strong bases dissociate completely in water, see the table below:

EXAMPLES OF COMMON STRONG ACIDS	DISSOCIATION OF COMMON STRONG ACIDS
HCl – hydrochloric acid HBr – hydrobromic acid HI – hydroiodic acid HNO_3 – nitric acid $HClO_4$ – perchloric acid H_2SO_4 – sulfuric acid (first ionization only)	strong acids = complete dissociation = single arrow: $\longrightarrow$ $HA \rightarrow H^+ + A^-$ $HCl \rightarrow H^+ + Cl^-$ $HBr \rightarrow H^+ + Br^-$ $HI \rightarrow H^+ + I^-$ $HNO_3 \rightarrow H^+ + NO_3^-$ $HClO_4 \rightarrow H^+ + ClO_4^-$ $H_nA \rightarrow nH^+ + A^{n-}$ $H_2SO_4 \rightarrow 2H^+ + SO_4^{2-}$
EXAMPLES OF COMMON STRONG BASES	DISSOCIATION OF COMMON STRONG BASES
NaOH – sodium hydroxide KOH – potassium hydroxide LiOH – lithium hydroxide RbOH – rubidium hydroxide CsOH – cesium hydroxide $Ca(OH)_2$ – calcium hydroxide $Ba(OH)_2$ – barium hydroxide $Sr(OH)_2$ – strontium hydroxide	strong bases = complete dissociation = single arrow: $\longrightarrow$ $MOH \rightarrow M^+ + OH^-$ $NaOH \rightarrow Na^+ + OH^-$ $KOH \rightarrow K^+ + OH^-$ $LiOH \rightarrow Li^+ + OH^-$ $RbOH \rightarrow Rb^+ + OH^-$ $CsOH \rightarrow Cs^+ + OH^-$ $M(OH)_n \rightarrow M^{n+} + nOH^-$ $Ca(OH)_2 \rightarrow Ca^{2+} + 2OH^-$ $Ba(OH)_2 \rightarrow Ba^{2+} + 2OH^-$ $Sr(OH)_2 \rightarrow Sr^{2+} + 2OH^-$
DISSOCIATION OF SALTS, examples: $M_nA_m \rightarrow nM^{m+} + mA^{n-}$ $NaCl \rightarrow Na^+ + Cl^-$ $Na_2SO_4 \rightarrow 2\,Na^+ + SO_4^{2-}$ $Fe_2(SO_4)_3 \rightarrow 2Fe^{3+} + 3SO_4^{2-}$	

Weak acids and weak bases dissociates (ionize) partially. Weak acids and bases react with water to establish an equilibrium system – double arrow: forward and reverse reactions. The degree to which an acid and a base dissociates is shown by acid dissociation constant (K_a) and base dissociation constant (K_b). These constants are a quantitative measure of the strength of acid an base in solution and depend on temperature only! These constants do not depend on concentration or pressure.

Acid dissociation constant:

$HA + H_2O \rightleftarrows A^- + H_3O^+$ $K_a = [A^-] \cdot [H_3O^+] / [HA]$, [] – refers to Molarity [mol/dm^3; mol/l]; [HA] – the molar concentration of HA at equilibrium

Base dissociation constant:

$$B + H_2O \rightleftarrows BH^+ + OH^- \qquad K_b = [BH^+] \cdot [OH^-] / [B]$$

Simplified version: $HA \rightleftarrows H^+ + A^-$ $K_a = [H^+] \cdot [A^-] / [HA]$

$$M(OH)_n \rightleftarrows M^{n+} + nOH^- \quad K_b = [Mn^{n+}] \cdot [OH^-]^n / [M(OH)_n]$$

Example of a weak acid: $H_2CO_3 \rightleftarrows 2H^+ + CO_3^{2-}$ $K_a = [H^+]^2 \cdot [CO_3^{2-}] / [H_2CO_3]$
Example of a weak base: $Mg(OH)_2 \rightleftarrows Mg^{2+} + 2OH^-$ $K_b = [Mg^{2+}] \cdot [OH^-]^2 / [Mg(OH)_2]$

Law of dilution by Wilhelm Ostwald is a relationship between a dissociation constant (K) and the degree of dissociation (α) of a weak acid or base:

$$\mathbf{K = \alpha^2 \text{ x } C_0 / 1-\alpha}$$

Simplified version of the Law of dilution ($\alpha < 0.05$; $C/K \geq 400$):

$$\mathbf{K = C_0 \cdot \alpha^2}$$

where, C_0 – is the total molar concentration of electrolyte [mol/dm^3; mol/l].

The degree of dissociation (α):

$$\alpha = C/C_0$$

where, C – the molar concentration of dissociated electrolyte; C_0 – the total molar concentration of electrolyte [mol/dm^3; mol/l]. Strong electrolyte: $\alpha = 1$; weak electrolyte: $\alpha < 1$.

The degree of dissociation (α) depends on the concentration of the solution (the higher the concentration of the solution the lower α), temperature (the higher the temperature the higher α) and also it depends on electrolyte.

Examples of acids and bases ionizing partway:

HF – hydrofluoric acid
H_2S – hydrosulfuric acid
H_2SO_3 – sulfurous acid
H_2CO_3 – carbonic acid
H_3BO_3 – boric acid
HClO – hypochlorous acid
H_3PO_4 – phosphoric acid
HCN – hydrocyanic acid (prussic acid)
HSCN – sulfocyanic acid (sulfocyanide)
CH_3COOH – acetic acid HNO_2 – nitrous acid

EXAMPLES OF BASES IONIZING PARTWAY
$Mg(OH)_2$ – magnesium hydroxide
$Al(OH)_3$ – aluminum hydroxide
$Cu(OH)_2$ – copper (II) hydroxide (cupric hydroxide)
$Fe(OH)_3$ – iron (III) hydroxide
$Fe(OH)_2$ – iron (II) hydroxide (ferrous hydroxide)
$Zn(OH)_2$ – zinc hydroxide
NH_4OH – Ammonia

The Common-Ion effect is an application of the Le Chatelier's principle. It is used to describe the effect on an equilibrium involving a substance that adds an ion to the solution that is already a part of the equilibrium. In other words addition of a common ion causes a shift in equilibrium – when additional product (common ion/ions) is added, the system shifts the equilibrium towards the reactants. The common ion effect states that any ionic salt is less soluble in the presence of a common ion, an ion that is a part of the salt.

Examples of selected calculations (with a solution)

Example 1

What is the mass/mass percentage when you dissolve 5 grams of sodium chloride in 50 grams of water? m/m % = 5g / 55 g · 100% = 9.1 %

Example 2

Calculate the mass/mass percentage of H_2SO_4. Given: $d_{H2SO4} = 1.175$ g/ cm^3; vol (volume) = 500 cm^3; mass of acid = 147 g. m/m% = grams solute/grams solution x 100%
grams solution = density · volume [g/ cm^3] m = d · vol = d · v = 1.175 g/ cm^3 · 500 ml = 587.5g
m/m% = 147g / 587.5g · 100% = 25%

Example 3

Calculate the molarity (M) of a 2 l solution made with 3 moles of KOH.

M (molarity) = moles of solute / liters of solution

M = n/V or C =n/V n – moles of solute; V – volume of solvent; c – concentration

M = 3moles/2 dm^3 = 1,5 M or c = 3moles/2l = 1,5 mol/l (mol/dm^3)

Example 4

Calculate the molarity of 2.5 l solution made with 200 g of calcium chloride ($CaCl_2$).

moles of solute (n) = mass of solute (m) / molar mass of solute (M) n = m/M

$CaCl_2$ molar mass (the mass of one mole of a substance) = atomic mass of calcium [g] + atomic mass of chlorine [g] = 40g + 2 · 35.5g = 111.1g/mol

n = 200g/111.1 = 1.8 moles

Molarity (M) = 1.8 moles/2.5 dm^3 = 0.72 M or 0.72 mol/dm^3 (mol/ dm^3) M = mol/dm^3

Example 5

How many grams of KCl would be required to produce a 3.0 M (3 mol/dm^3) solution with a volume of 500 cm^3?

KCl molar mass = 39g + 35.5g = 74.5g/mol

c = n/V n = c · V = 3 mol/ dm^3 · 0.5dm^3 = 1.5 moles of KCl

n = m/M = moles of solute / molar mass of solute

m = c x M · V = 3.0 mol/dm^3 · 74.5 g/mol · 0.5dm^3 = 111.75g

Additional tasks and questions to be performed by the student:

1. Calculate the mass/mass percentage. Given: mass of solute = 25g of NaOH; mass of solution = 500g. Answer: 5%
2. Calculate the mass of solution. Given: mass of solute = 36g of NaOH; m/m% percentage = 18%. Answer: 200 grams
3. Calculate the mass of solute. Given: volume of HCl solution = 1000 cm^3; m/m% percentage of HCl = 30%; d_{HCl} = 1.15 g/cm^3. Answer: 345 grams
4. Calculate the mass of solvent – mass of water. Given: mass of solute = 20g of NaOH; m/m% percentage = 5%. Answer: 380 grams
5. Calculate the mass/mass percentage of $NaNO_3$. Given: d = 1,175 g/ cm^3; vol (volume) = 100cm^3; mass of solute = 28.2g. Answer: 24%
6. Calculate the volume/volume percentage. Given: volume of solute = 50cm^3 of hexane; volume of benzene (solvent) = 150cm^3. Answer: 25%
7. Determine the molarity (M) of KOH. Given: volume of solution = 500cm^3; mass of solute = 5,62g of KOH. Answer: 0,2 mol/dm^3
8. Determine the molarity (M) of NaOH. Given: volume of solution = 200cm^3; mass of solute = 160mg. Answer: 0,02 mol/dm^3
9. How many grams of $KMnO_4$ would be required to produce a 0,02M (0,02 mol/dm^3) solution with a volume of 1000 cm^3? Answer: 3,16 grams
10. How many grams of Na_2CO_3 would be required to produce a 0,1 (0,1 mol/dm^3) solution with a volume of 100 cm^3? Answer: 1,06 grams

2. INSTRUCTION 2 – LABORATORY EXERCISE 2

Experiment 1 – The effect of temperature on solubility of selected substances

Materials and reagents:

Glass test tube set, measuring cylinder, micro spatula, test tube holder, laboratory water bath, salts: sodium chloride (NaCl), potassium nitrate (KNO_3), saturated solution of calcium hydroxide ($Ca(OH)_2$).

Experimental procedure:

Add 5 cm^3 of saturated solution of calcium hydroxide into the test tube and heat it in the laboratory water bath.

Prepare a saturated solution of sodium chloride (NaCl): add 5 cm^3 of distilled water into the first test tube and add quite a bit of solid sodium chloride (NaCl), stir it. Keep adding sodium chloride and stirring. The solution is saturated if it has dissolved all of the sodium chloride it can at the temperature (sodium chloride falls to the bottom of the test tube – the solution that is on top of sodium chloride is saturated).

Check how the solubility of sodium chloride changes with temperature: heat the saturated solution of sodium chloride (by putting the test tube into the laboratory water bath), stir it and add again solid sodium chloride until it falls to the bottom of the test tube, etc.

Do the same with potassium nitrate – second test tube: make a saturated solution of potassium nitrate and then check how the solubility of potassium nitrate changes with temperature – the same procedure as for the sodium chloride.

Data analysis (after the experiment):

1. How the solubility of sodium chloride, potassium nitrate and calcium hydroxide changes with temperature?
2. Construct the solubility graph for sodium chloride, potassium nitrate and calcium hydroxide.

Solubility [g/100g H_2O] of selected chemical compounds:

Chemical compound	Temperature, K/ºC						
	283/10	293/20	303/30	313/40	323/50	333/60	363/90
$BaCl_2$	33,3	35,7	38,2	40,7	43,6	46,4	Not available
KCl	31,0	34,0	37,0	40,0	42,6	45,5	54,0
KNO_3	20,9	31,6	45,8	63,9	85,5	110,0	202,0
$Ca(OH)_2$	0,176	0,165	0,153	0,141	0,128	0,116	0,085
NaCl	35,8	36,0	36,3	36,6	37,0	37,3	39,0
$NaNO_3$	80,0	88,0	96,0	104,0	114,0	124,0	Not available
saccharose	190,0	204,0	225,0	245,0	278,0	294,0	Not available

Experiment 2 – The supersaturated solution

Materials and reagents:

Glass test tube set, measuring cylinder, micro spatula, laboratory water bath, solid substance: sodium acetate (CH_3COONa).

Experimental procedure:

Fill the dry test tube with solid sodium acetate (CH_3COONa) up to three quarters (3/4) of the test tube, add 5 cm^3 of distilled water and heat it in the laboratory water bath until the salt is completely dissolved. Remove the test tube carefully and cool it down (the lower part of the test tube) in the stream of running water. Then add just a crystal of salt and observe what happens.

Data analysis (after the experiment):

1. Write the definition of supersaturated solution and give an example of saturated solution.
2. Write the definition of crystallization.

Experiment 3 – The effect of temperature on solubility of selected substances – reaction enthalpy.

Materials and reagents:

Glass test tube set, measuring cylinder, micro spatula, test tube holder, laboratory water bath, salts: magnesium chloride ($MgCl_2$), ammonium nitrate (NH_4NO_3).

Experimental procedure:

Prepare a saturated solution of magnesium chloride ($MgCl_2$) – first test tube and a saturated solution of ammonium nitrate (NH_4NO_3) – second test tube. Compare the heat effects while dissolving the salts. Add two micro spatulas of magnesium chloride ($MgCl_2$) into the first test tube and two micro spatulas of ammonium nitrate (NH_4NO_3) into the second test tube. Heat both test tubes in the laboratory water bath. Observe the salts solubility effect in both cases.

Data analysis (after the experiment):

1. Explain the phenomenon using the Le Chátelier’s principle.
2. Construct the solubility graph for magnesium chloride and ammonium nitrate.
3. Identify the endothermic (heat is consumed) and exothermic (heat is released) reaction.

Experiment 4 – The effect of dielectric constant on dissociation

Materials and reagents:

Glass test tube set, measuring cylinder, copper (II) chloride (1M $CuCl_2$) solution, acetone (CH_3COCH_3).

Experimental procedure:

Pour about 2 cm^3 of copper (II) chloride (1M $CuCl_2$) solution to each of two test tubes. Keep adding acetone (CH_3COCH_3) to the first test tube until it changes the color. Add the same amount of distilled water as acetone (added to the first test tube) into the second test tube. Record the color of obtained solutions in both test tubes. Add distilled water to the first test tube and observe what happens with the color.

Data analysis (after the experiment):

Fill in the table below:

Test tube	Content	Solvent added; dielectric constant of the dominant solvent	The colour of the solution after adding of solvent (acetone/distilled water)	Dissociation reaction of copper (II) chloride (in acetone/in water); chemical equilibrium	Conclusions based on dielectric constant
1.	$CuCl_2$ plus acetone-dominant solvent; water				
2.	$CuCl_2$ plus water				

Experiment 5 – The common ion effect

Materials and reagents:

Glass beakers, measuring cylinder, iron (III) sulfate solution (0.1M $Fe_2(SO_4)_3$), potassium thiocyanate solution (0.3M KSCN).

Experimental procedure:

Add 2 – 3 drops of iron (III) sulfate solution (0.1M $Fe_2(SO_4)_3$) and 2 – 3 drops of potassium thiocyanate solution (0.3M KSCN) to a beaker, stir it. Then add 5 cm^3 of distilled water to the same beaker. Record the color of the solution. Dilute the obtained solution in the beaker with distilled water until the blood color disappears (dilute to get the yellowish color of the solution). Divide the obtained solution into two parts (two parts = two beakers). Add 10 drops of iron (III) sulfate solution (0.1M $Fe_2(SO_4)_3$) into the first beaker and 10 drops of potassium thiocyanate solution (0.3M KSCN) into the second beaker. Record the color of the solution in both beakers.

Data analysis (after the experiment):

1. Write the chemical reaction leading to a blood red color (balanced chemical equation between iron (III) sulfate and potassium thiocyanate).
2. Write the dissociation reaction of the obtained salt and the dissociation constant (K).
3. How do you explain the return of the blood red color after adding solutions of iron (III) sulfate and potassium thiocyanate?

Experiment 6 – Acids and Bases – dissociation reaction – equilibrium shifts

Materials and reagents:

Glass test tube set, measuring cylinder, micro spatula, acetic acid solution (0.1M CH_3COOH), ammonium hydroxide solution (0.1M NH_4OH), solid substances: sodium acetate (CH_3COONa) and ammonium chloride (NH_4Cl), methyl orange, phenolphthalein.

Experimental procedure:

Pour 4 cm^3 of acetic acid solution (0.1M CH_3COOH) to each of two test tubes. Add to both test tubes 2 drops of methyl orange. Treat the first test tube as a control sample, add to the second test tube solid sodium acetate (CH_3COONa) – one micro spatula and stir it. Record and compare the color in both test tubes, the control sample – first test tube, and the sample – second test tube.

Pour 4 cm^3 of ammonium hydroxide solution (0.1M NH_4OH) into the next two test tubes (test tubes 3 and 4). Add one drop of phenolphthalein solution to each of two test tubes. Leave one test tube as a control sample – test tube 3, add to the second test tube solid ammonium chloride (NH_4Cl) – one micro spatula and stir it – test tube 4. Record and compare the color in both test tubes.

Data analysis (after the experiment):

Fill in the table below:

Test tube	Content	Chemical formula of acid/base; dissociation reaction	Chemical formula of added salt; dissociation reaction	The common ion	The colour of the solution	Equilibrium shifts (toward the product/products or reactants)
1.	Acetic acid+methyl orange – control sample		–	–		–
2.	Acetic acid+methyl orange+sodium acetate – sample					
3.	Ammonium hydroxide+phenolphthalein – control sample		–	–		–
4.	Ammonium hydroxide+phenolphthalein+ammonium chloride – sample					

3. GUIDELINES FOR WRITING THE FINAL LABORATORY REPORT

1. First page of the report – The Laboratory Report Cover Sheet found on our website: https://www.am.szczecin.pl/en/facilities/institute-of-mathematics-physics-and-chemistry/department-of-chemistry/chemistry-lab-manuals/
2. Second page of the report – „The Theoretical Part" – on a maximum of one page including brief description of keywords.
3. Third page of the report – „The Experimental Part" – including all performed experiments with titles, raw data, reactions, calculations, tables, graphs, etc. It should be written in accordance with „Data analysis (after the experiment)".
4. Additional task/tasks given by the academic teacher.
5. References.

4. IN ORDER TO PASS THE LABORATORY EXERCISE STUDENTS MUST PASS „THE ENTRY TEST" AND SUBMIT THE FINAL LABORATORY REPORT AT THE NEXT LABORATORY MEETING. THE LAB REPORT MUST BE ACCEPTED BY THE ACADEMIC TEACHER.