



**Institute of Mathematics,  
Physics and Chemistry**

**Department of Chemistry**

**Technical chemistry laboratory**

**Laboratory exercise**

**Ionic dissociation**

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## EXERCISE SHEET

1	<b>Relation to subjects:</b> ESO/26, DiRMiUO/26, EOUNiE/26		
	<b>Specialty/Subject</b>	<b>Learning outcomes for the subject</b>	<b>Detailed learning outcomes for the subject</b>
	ESO/ 25 Technical chemistry	EKP1 K_W01, K_W02, K_U05  EKP2 K_U08, K_U09	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.
	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	<b>Purpose of the exercise:</b> <ul style="list-style-type: none"> <li>– mastering the knowledge of solutions and basic methods of expressing concentrations as well as ionic dissociation;</li> <li>– consolidation of knowledge about the degree and constant of dissociation, strong and weak acids and bases;</li> <li>– learning to write the dissociation equations of acids, bases and salts and the formula for the equilibrium constant K;</li> <li>– mastering the basic mathematical relationships of concentration, constant K and degree of dissociation;</li> <li>– experimental study of the influence of selected physicochemical factors on the course of the ion dissociation process in laboratory conditions.</li> </ul>		
3	<b>Prerequisites:</b> 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	<b>Description of the laboratory workplace:</b> 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	<p><b>Risk assessment:</b>  Contact with chemical reagents in solid form and diluted in liquid form – probability of chemical burn – very small, effects – minor  Final assessment – <b>MEDIUM</b>  <b>Security measures required:</b></p> <ol style="list-style-type: none"> <li>protective gloves,</li> <li>protective glasses,</li> <li>lab coats.</li> </ol>
6	<p><b>The course of the exercise:</b></p> <ol style="list-style-type: none"> <li>Getting to know the workplace manual (appendix 2),</li> <li>Performing individual exercises in accordance with the workplace instruction for exercises.</li> </ol>
7	<p><b>Exercise report:</b></p> <ol style="list-style-type: none"> <li>Develop an exercise in accordance with the instructions contained in the workplace manual.</li> <li>Solve the given task and/or answer the questions included in the set of tasks and questions to be completed by the student.</li> </ol>
8	<p><b>Archiving of research results:</b>  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</p>
9	<p><b>Assessment method and criteria:</b></p> <ol style="list-style-type: none"> <li>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,</li> <li>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: <ul style="list-style-type: none"> <li>– <b>mark 2,0</b> – 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;</li> <li>– <b>mark 3,0</b> – 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);</li> <li>– <b>mark 3,5 – 4,0</b> – has extended chemical knowledge in the field of solutions and electrolytic dissociation and has the ability to solve complex tasks in his specialty;</li> <li>– <b>mark 4,5 – 5,0</b> – has the ability to apply complex knowledge in the field of solutions and electrolytic dissociation to solve complex tasks and problems in his specialty.</li> </ul> </li> </ol>
10	<p><b>Literature:</b></p> <ol style="list-style-type: none"> <li>Stundis H., Trzeźniowski W., Żmijewska S.: <i>Ćwiczenia laboratoryjne z chemii nieorganicznej</i>. WSM, Szczecin 1995.</li> <li>Kozłowski A., Ćwirko K., <i>Instrukcja stanowiskowa do ćwiczeń laboratoryjnych: Roztwory i dysocjacja jonowa</i>, AM Szczecin, 2013.</li> <li>Cox P.A. translation of Z. Zawadzki: <i>Chemia nieorganiczna</i>. PWN. Warsaw 2006.</li> <li>Drapała T.: <i>Chemia ogólna i nieorganiczna</i>. SGGW, Warsaw 1994.</li> <li>Bieleński A.: <i>Chemia ogólna i nieorganiczna</i>. PWN, Warsaw 1994.</li> <li>Jones L., Atkins P., <i>Chemia ogólna</i>. Molecules, matter reactions, WN PWN, Warsaw 2004.</li> </ol>

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

## APPENDIX 1 – MANUAL

### 1. SCOPE OF THE EXERCISE

#### Issues and keywords:

- basic concepts of ion dissociation (electrolytic dissociation equations of acids, bases and salts, constant K, degree of dissociation, strong and weak electrolytes);
- factors affecting the degree of dissociation (common-ion effect, Ostwald's dilution law).

### 2. THEORETICAL INTRODUCTION TO THE EXERCISE

#### 1. Ionic dissociation

In the second half of the 19th century, the Swedish chemist S.A. Arrhenius proved experimentally that chemicals can be divided into two groups.

Group I includes those whose aqueous solutions conduct electricity, and group II includes those which practically do not conduct electricity under the same conditions. Arrhenius called group I substances electrolytes, group II substances non-electrolytes.

The ability to conduct electricity through aqueous solutions of electrolytes, Arrhenius explained by the presence of particles in them capable of carrying electric charges, and therefore endowed with charges.

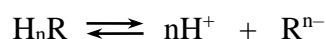
Arrhenius theory can be summarized in the form of four basic experiential assumptions:

- electrolytes, i.e. acids, bases and salts, when dissolved in water, decompose into electrically charged elements, i.e. they undergo the so-called electrolytic dissociation. These elements were called ions. Positively charged ions are called cations, negative ions are called anions;
- the sum of the electric charges of cations and anions resulting from electrolytic dissociation of electrolytes is always equal to zero;
- non-electrolytes, i.e. substances that do not conduct electric current in solutions and in the molten state, do not undergo electrolytic dissociation;
- the chemical properties of ions are quite different from those of neutral atoms and molecules.

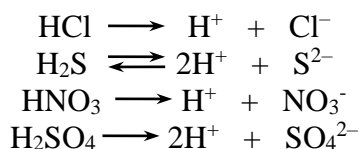
In 1887 Arrhenius gave the following definition of acids, hydroxides, and salt:

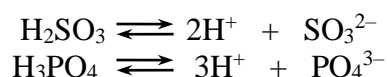
**Acids** are chemical compounds which, when dissolved in water, dissociate completely or partially into hydrogen cations and anions of acidic residues.

The electrolytic dissociation of acids is represented by the equation:



e.g.:

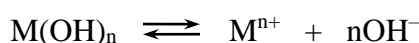




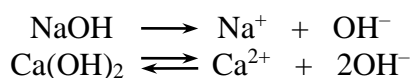
Acids such as  $\text{HNO}_3$ , which in their molecule have one hydrogen atom that can be detached during dissociation, are called monoprotic acids. Acids with two hydrogen atoms in the molecules capable of dissociating two hydrogen cations are called diprotic acids, and those with three hydrogen atoms – triprotic acids.

**Hydroxides** are chemical compounds which, when dissolved in water, dissociate completely (bases) or partially into hydroxide anions  $\text{OH}^-$  and metal cations.

The electrolytic dissociation of hydroxides is represented by the equation:



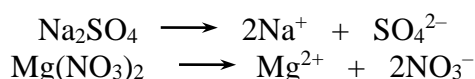
e.g.:



**Salts** are the reaction products of acids with bases. At room temperature, these compounds generally exist in a solid, crystalline state and have an ionic structure, i.e. metal cations or an ammonium cation and anions of acidic residues. Therefore, the process of dissolving the salt in water consists in switching to the solution of the cations and anions already existing in the crystal. The ionic structure of the salt also explains the electric conductivity of the molten salts. The water-soluble salt dissociates into the metal cation  $\text{M}^{n+}$  and the anion of the acid residue  $\text{R}^{n-}$ .



e.g.:



Research on electrolytes has shown that despite the same molar concentration of different electrolytes in the solution, they differ in their ability to conduct electricity.

Such electrolytes that conduct electric current well in aqueous solutions are called strong electrolytes, and electrolytes that at the same concentration as strong electrolytes show much weaker electrical conductivity are called weak electrolytes.

Strong electrolytes include almost all salts (except some Hg salts), some inorganic acids (e.g.:  $\text{HCl}$ ,  $\text{HNO}_3$ ,  $\text{HClO}_4$ ,  $\text{H}_2\text{SO}_4$ ,  $\text{HBr}$ ,  $\text{HI}$ ) and all alkali metals hydroxides (e.g.:  $\text{LiOH}$ ,  $\text{KOH}$ ,  $\text{NaOH}$ ) and some alkaline earth hydroxides (e.g.:  $\text{Ba}(\text{OH})_2$ ,  $\text{Sr}(\text{OH})_2$ ).

The weak electrolytes are: acids –  $\text{HF}$ ,  $\text{HNO}_2$ ,  $\text{H}_2\text{S}$ ,  $\text{CH}_3\text{COOH}$ ,  $\text{H}_2\text{CO}_3$ , hydroxides:  $\text{Mg}(\text{OH})_2$  and metal hydroxides of the other groups (except I and II),  $\text{NH}_4\text{OH}$ .

Examples of medium-strength electrolytes are:  $\text{H}_2\text{SO}_3$ ,  $\text{H}_3\text{PO}_4$ ,  $\text{H}_3\text{AsO}_4$  and  $\text{Ca}(\text{OH})_2$ .

### 1.1. The degree of dissociation

The degree of dissociation, usually denoted by the symbol  $\alpha$ , is a quantity that quantitatively describes the dissociation of electrolytes in solution. This concept was first introduced by Arrhenius and is often used to compare the strength of electrolytes. The degree of dissociation of the electrolyte  $\alpha$  in the first approximation determines the amount of dissociation. It is expressed as the ratio of the number of moles of molecules dissociated into ions to the number of moles of solute molecules.

$$\alpha = \frac{n_z}{n_0} \cdot 100\%$$

where:

- $n_z$  – number of moles of molecules dissociated into ions,
- $n_0$  – number of moles of solute molecules.

The value of the degree of dissociation is less than or equal to one, it can also be expressed as a percentage. If the degree of dissociation is equal to or close to one, then the electrolyte is almost completely dissociated into ions and the electrical conductivity of such a solution is significant. It turns out that strong electrolytes completely dissociate ( $\alpha = 1$ ), while weak electrolytes have  $\alpha < 1$ . For weak electrolytes, the dissociation degree values are several percent. The degree of dissociation depends on:

- a. type of electrolyte and type of solvent,
- b. solution concentration,
- c. temperature,
- d. presence of other electrolytes in the solution.

#### – Type of electrolyte

The basic factor determining the value of the degree of dissociation is the binding strength in the molecule that undergoes dissociation. An example would be the HF molecule. Although the H – F bond is much more polar than the H – Cl bond, HF is a weak electrolyte and HCl is a strong electrolyte. Part of this is due to the very strong H – F bond, as a result of which hydrogen is difficult to separate.

#### – Type of solvent

The solvent can significantly affect the value of the degree of dissociation of a given electrolyte. For example, under comparable conditions, hydrochloric acid (HCl) is 100% dissociated in water, and in benzene, the dissociation rate of this acid is  $\alpha < 1\%$ .

#### – Solution concentration

The concentration has an important influence on the degree of dissociation. Measurements showed that  $\alpha$  increases with dilution of the electrolyte and in very dilute solutions all electrolytes show a degree of dissociation close to one, i.e. almost all molecules or groups of solute ions dissociate into ions. The above is due to Ostwald's law of dilution.

#### – Temperature

The degree of dissociation increases with increasing temperature, but the effect is small.

#### – Common-ion effect

If we add another compound to the electrolyte solution that is in a certain ionic equilibrium, which during dissociation in the solution will produce an ion already present in this equilibrium (there may be more of these common ions), then the equilibrium state will be disturbed, and in accordance with the law of mass action of Guldberg and Waage and Le Chatelier's and Braun's equilibrium law (both laws will be hereinafter referred to as the laws of chemical equilibrium – the law of masses describes the behaviour of a system in a mathematical manner, the equilibrium law presents the same qualitatively in a descriptive

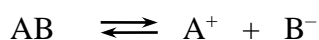
manner), the system must be adapt to the new situation. The adaptation of the system to the changing concentration of ions in the ionic equilibrium is called the common-ion effect. However, it should be remembered that the common-ion effect is simply due to the correct application of the laws of chemical equilibrium.

The common-ion effect occurs in electrolyte solutions whenever we are interested in the pH of the solution (because the solvent – water – also supplies  $\text{H}_3\text{O}^+$  oxonium ions to the solution), but in a large part of the calculations we ignore it completely. This effect should be taken into account in part of calculations concerning ionic equilibria in electrolyte solutions, and in particular in calculations of:

- pH or degree of protolysis of weak electrolytes in buffer solutions,
- the degree of protolysis of weak electrolytes in mixtures with strong acids or hydroxides (e.g.  $\text{CH}_3\text{COOH}$  and  $\text{HCl}$ ,  $\text{NH}_3 \cdot \text{H}_2\text{O}$  and  $\text{NaOH}$ ),
- pH or the degree of protolysis of components in mixtures of weak acids or weak hydroxides (e.g.  $\text{HCOOH}$  and  $\text{CH}_3\text{COOH}$ ,  $\text{CH}_3\text{NH}_2 \cdot \text{H}_2\text{O}$  and  $\text{NH}_3 \cdot \text{H}_2\text{O}$ ),
- pH or the degree of protolysis of very dilute acids and hydroxides (by taking into account the proportion of oxonium ions derived from the autoprotolysis of water in the calculations).

## 1.2. The dissociation constant

The dissociation of weak electrolytes can be viewed as a reversible chemical reaction. The weak electrolyte of the general formula AB undergoes dissociation in aqueous solution according to the equation:



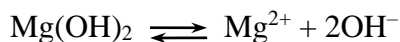
Since electrolytic dissociation is a reversible process, there is an equilibrium in the electrolyte solution for which you can write the expression for the equilibrium constant ( $K_c$ )

$$K_c = \frac{[\text{A}^+] \cdot [\text{B}^-]}{[\text{AB}]} = \text{const.}, \text{ if } T = \text{const.}$$

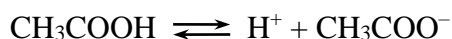
where:

- $[\text{A}^+]$  –,  $[\text{B}^-]$  – actual ion concentration at equilibrium,  
 $[\text{AB}]$  – the actual concentration of undissociated molecules.

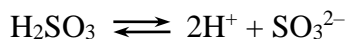
### Examples:



$$K_c = \frac{[\text{Mg}^{2+}] \cdot [\text{OH}^-]^2}{[\text{Mg}(\text{OH})_2]}$$



$$K_c = \frac{[\text{H}^+] \cdot [\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]}$$



$$K_c = \frac{[\text{H}^+]^2 \cdot [\text{SO}_3^{2-}]}{[\text{H}_2\text{SO}_3]}$$



The constant  $K_c$  of electrolytes at a given temperature is practically constant, **independent of concentration**. Therefore, it can be used as an exact measure of the power of electrolytes, much more accurate than the degree of dissociation ( $\alpha$ ), which depends on the concentration.

**The stronger the electrolyte, the greater the value of the constant  $K_c$ .**

Table 4

$K_c$  dissociation constants values for selected acids at 298K

Type of acid	Dissociation constant
<b>Strong acids</b>	
HI	$K = 1.0 \cdot 10^{10}$
HBr	$K = 3.0 \cdot 10^9$
HCl	$K = 1.0 \cdot 10^7$
H <sub>2</sub> SO <sub>4</sub>	$K_1 = \text{about } 10^3$ $K_2 = 1.2 \cdot 10^{-2}$
HClO <sub>3</sub>	$K = 5 \cdot 10^2$
HNO <sub>3</sub>	$K = 1$
HClO <sub>4</sub>	$K = \sim 1.0 \cdot 10^{10}$
<b>Medium-strength acids</b>	
HClO <sub>2</sub>	$K = 1.0 \cdot 10^{-2}$
H <sub>2</sub> SO <sub>3</sub>	$K_1 = 1.6 \cdot 10^{-2}$ $K_2 = 6.3 \cdot 10^{-8}$
H <sub>3</sub> PO <sub>3</sub>	$K = 1.6 \cdot 10^{-3}$
H <sub>3</sub> PO <sub>4</sub>	$K_1 = 7.5 \cdot 10^{-3}$ $K_2 = 6.3 \cdot 10^{-8}$ $K_3 = 1.3 \cdot 10^{-12}$
<b>Weak acids</b>	
HNO <sub>2</sub>	$K = 2.0 \cdot 10^{-4}$
H <sub>2</sub> CO <sub>3</sub>	$K_1 = 4.5 \cdot 10^{-7}$ $K_2 = 4.7 \cdot 10^{-11}$
H <sub>2</sub> S	$K_1 = 6 \cdot 10^{-8}$ $K_2 = 10^{-14}$
HClO	$K = 2.9 \cdot 10^{-8}$
H <sub>3</sub> BO <sub>3</sub>	$K = 5.5 \cdot 10^{-9}$
HCN	$K = 7.5 \cdot 10^{-10}$
H <sub>3</sub> BO <sub>3</sub>	$K_1 = 5.8 \cdot 10^{-10}$ $K_2 = 1.8 \cdot 10^{-13}$ $K_3 = 1.6 \cdot 10^{-14}$

Summing up, in the case of acids, the division of acids shown in Fig. 3 can be assumed, but due to the value of the constant  $K_c$  sulphuric acid (IV) – H<sub>2</sub>SO<sub>3</sub> and orthophosphoric acid (V) – H<sub>3</sub>PO<sub>4</sub> according to some authors, are classified as medium-strength acids.

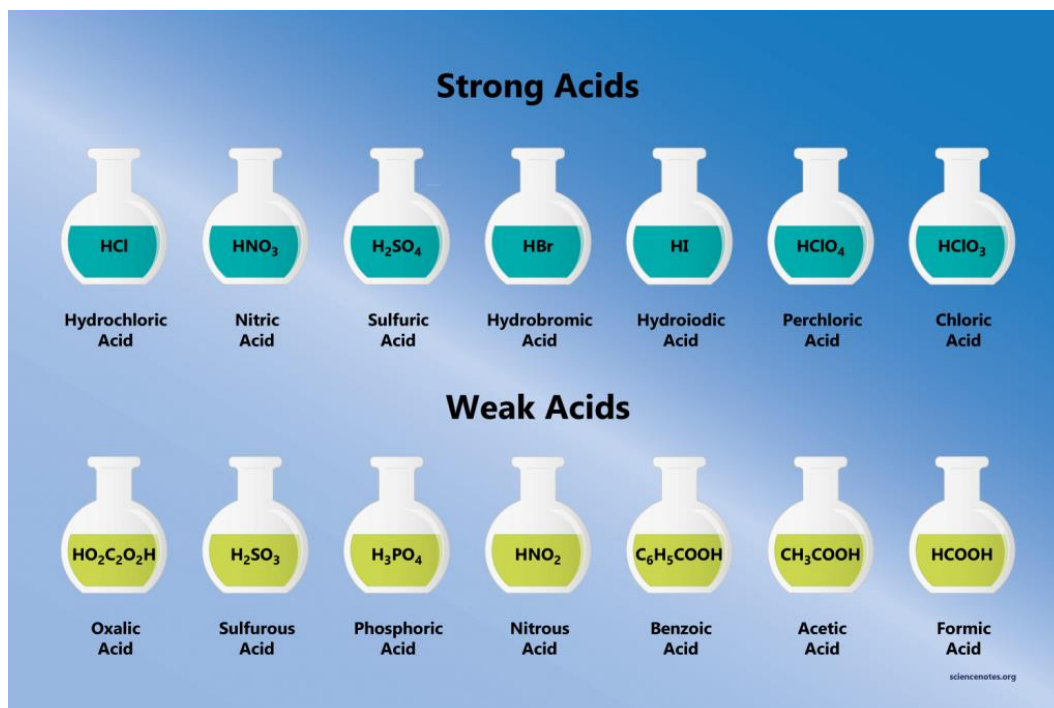


Fig. 3. Examples of selected strong and weak acids  
 source: <https://sciencenotes.org/list-of-common-strong-and-weak-acids/>  
 (accessed September, 29, 2021)

Table 5 shows the values of the  $K_b$  constants for selected common hydroxides

Table 5

$K_b$  dissociation constants values for selected hydroxides at 298K

Type of hydroxide	Dissociation constant
Strong hydroxides	
NaOH	$K > 10^{14}$
KOH	$K > 10^{14}$
LiOH	$K > 10^{14}$
RbOH	$K > 10^{14}$
Weak hydroxides	
$Mg(OH)_2$	$K = 2.5 \cdot 10^{-3}$
$Fe(OH)_2$	$K = 1.3 \cdot 10^{-4}$
$Cu(OH)_2$	$K = 1.3 \cdot 10^{-4}$
$Zn(OH)_2$	$K = 4.0 \cdot 10^{-5}$
$Al(OH)_3$	$K = 1.4 \cdot 10^{-9}$

To sum up, in the case of hydroxides, it can be assumed that strong hydroxides form only metals in the **I and II main groups**, and not all of them, because calcium hydroxide  $Ca(OH)_2$  by some authors, due to the value of the constant  $K_b$  is considered a medium-strength hydroxide, and magnesium hydroxide  $Mg(OH)_2$  and beryllium hydroxide are weak hydroxides. The hydroxides of the metals of the remaining groups form weak hydroxides. Among other common hydroxides, a weak hydroxide is also ammonium hydroxide  $NH_4OH$ , the constant of which has the value  $K_b = 1.8 \cdot 10^{-5}$ .

### 1.3. Ostwald's law of dilution

Ostwald's law captures the relationship between the degree of dissociation and the dissociation constant. For e.g. weak acetic acid  $\text{CH}_3\text{COOH}$ , with initial concentration  $c$ , dissociation is as follows:



After the dissociation of acetic acid, the equilibrium state, the concentrations of hydrogen, acetate and undissociated acid in the solution are respectively:

$$\begin{aligned} [\text{H}^+] &= \alpha C_o \\ [\text{CH}_3\text{COO}^-] &= \alpha C_o \\ [\text{CH}_3\text{COOH}] &= (1 - \alpha) \cdot C_o \end{aligned}$$

Substituting these values into the formula for the constant  $K_c$  we obtain:

$$K_c = \frac{\alpha \cdot C_o \cdot \alpha \cdot C_o}{C_o \cdot (1 - \alpha)} = \frac{\alpha^2 \cdot C_o}{1 - \alpha}$$

and we get the so-called Ostwald's law of dilution.

$$K_c = \frac{c \cdot \alpha^2}{1 - \alpha}$$

where:

- $c$  – output molar concentration of electrolyte [ $\text{mol}/\text{dm}^3$ ],
- $\alpha$  – degree of dissociation,
- $K_c$  – dissociation constant.

Since for weak electrolytes the value of the dissociation degree is very small  $\alpha \ll 1$ , the result of subtraction  $[1 - \alpha]$  is close to 1 and we can write:

$$K_c = c \cdot \alpha^2$$

hence:

$$\alpha = \sqrt{\frac{K_c}{c}}$$

From Ostwald's law of dilution, it follows that the degree of dissociation of a weak electrolyte is inversely proportional to the square root of the molar concentration of this electrolyte and directly proportional to the square root of its dissociation constant. This means that at an infinitely great dilution, even weak electrolytes can be completely dissociated into ions.

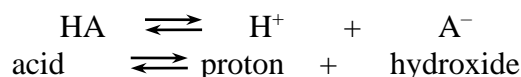
The law of dilution is valid only for weak and very dilute electrolyte solutions.

## 1.4. Theories of acids and hydroxides

The Arrhenius electrolytic dissociation theory introduced the definition of an acid as a stripping compound in aqueous solutions of hydrogen cations  $H^+$  and hydroxide of the stripping compound under these conditions hydroxyl anions  $OH^-$ . These definitions explain satisfactorily most of the phenomena occurring with the participation of acids or bases in aqueous solutions. However, observations have been made that revealed the shortcomings of the Arrhenius theory. For example, it cannot convincingly explain the basic properties of ammonia and amines in water, especially the behaviour of electrolytes in non-aqueous solutions. It is also impossible to explain why the hydrogen ion in aqueous solutions does not exist as a free proton, but as bound to at least one water molecule and exists as a **hydronium ion  $H_3O^+$** . These and many other facts led in 1923 the Danish chemist Brønsted, and independently of him, the English chemist Lowry, to a more general theory called **the proton theory of acids and hydroxides**.

According to Brønsted, **acids** are chemical compounds (or ions) capable of donating protons, and hydroxides are chemical compounds (or ions) capable of attaching protons.

An acid molecule or ion, after donating a proton, becomes a hydroxide molecule or ion, and conversely, a hydroxide molecule, after attaching a proton, becomes an acid molecule or ion.



Compound  $A^-$  is therefore a hydroxide conjugated with the appropriate acid **HA**, and the acid **HA** is conjugated with the hydroxide  $A^-$ .

Table 6

Examples of acids and hydroxides so defined are as follows

Acid		Hydroxide		Acid		Hydroxide
HCl	+	$H_2O$	$\rightleftharpoons$	$H_3O^+$	+	$Cl^-$
$H_2O$	+	$NH_3$	$\rightleftharpoons$	$NH_4^+$	+	$OH^-$
$H_3O^+$	+	$OH^-$	$\rightleftharpoons$	$H_2O$	+	$H_2O$
$HCO_3^-$	+	$H_2O$	$\rightleftharpoons$	$H_3O^+$	+	$CO_3^{2-}$
$H_2O$	+	$HCO_3^-$	$\rightleftharpoons$	$H_2CO_3$	+	$OH^-$

According to the proton theory of acids and hydroxides, an acid can show its acidic properties only towards the hydroxide which accepts a proton, and vice versa, the hydroxide can only be acidic towards the acid.

The presented examples show that water can behave ambiguously, sometimes it can react as acid and other times as hydroxide, depending on the type of the dissolved substance. Many other solvents behave similarly to water, including alcohol and acetic acid, which donate a proton.

### 3. PERFORMING THE EXERCISE

#### Experiment 1 – Influence of the dielectric constant on the ionic dissociation process

##### Materials and reagents:

Rack with test tubes, measuring cylinder, copper(II) chloride solution (1M CuCl<sub>2</sub>), acetone (CH<sub>3</sub>COCH<sub>3</sub>).

##### Performance:

Pour about 2 cm<sup>3</sup> of copper(II) chloride solution (1M CuCl<sub>2</sub>) into two test tubes. Add acetone (CH<sub>3</sub>COCH<sub>3</sub>) to the first one until the colour changes, to the second one the same amount of distilled water. Note down the colour of the solutions obtained. Then add distilled water to the test tube with acetone (CH<sub>3</sub>COCH<sub>3</sub>) until the colour changes. Compare the colours of the solutions in both test tubes.

##### Elaboration of the results:

Explain the cause of the colour change of the copper(II) chloride solution (1M CuCl<sub>2</sub>) based on the value of the dielectric constant of the solvent used. Present the results of the experiment in the form of a properly completed Table 2, which should be attached to the exercise report.

Table 7

Summary of the results for experiment 4 and conclusions

No of test tube	Content	Type of solvent added, dielectric constant value of the dominant solvent	The colour of the test solution after adding the solvent (acetone/water)	Copper(II) chloride dissociation reaction in the tested environment (* or **)	The cause of the colour of the solution and its change, explain on the basis of Coulomb's law
1.	Copper(II) chloride, water, acetone				
2.	Copper(II) chloride, water				

\*  $\text{CuCl}_2 \rightleftharpoons \text{Cu}^{2+} + 2\text{Cl}^-$  shifted equilibrium towards ions

\*\*  $\text{Cu}^{2+} + 2\text{Cl}^- \rightleftharpoons \text{CuCl}_2$  shifted equilibrium towards undissociated molecules

## Experiment 2 – Reactions with cobalt(II) chloride

### Materials and reagents:

Rack with test tubes, measuring cylinder, cobalt(II) chloride solution (1M  $\text{CoCl}_2$ ), concentrated hydrochloric acid (HCl).

### Performance:

Introduce 2  $\text{cm}^3$  (about 1/5 volume) of the cobalt(II) chloride ( $\text{CoCl}_2$ ) solution into the test tube. Observe the colour of the solution. Then add the concentrated hydrochloric acid dropwise (until the solution changes colour). Re-dilute the solution with distilled water and observe the colour change once more.

### Elaboration of the results:

Write the cobalt(II) chloride dissociation equation and write the equilibrium (dissociation) constant of this process. Explain why addition of common chloride ion reverses cobalt(II) chloride dissociation.

## Experiment 3 – Common-ion effect

### Materials and reagents:

Beakers with a capacity of 50  $\text{cm}^3$ , measuring cylinder, iron(III) sulphate(VI) (0.1M  $\text{Fe}_2(\text{SO}_4)_3$ ), potassium thiocyanate solution (0.3 M KSCN).

### Performance:

In a 50  $\text{cm}^3$  beaker, mix 2 – 3 drops of iron(III) sulphate(VI) solution (0.1M  $\text{Fe}_2(\text{SO}_4)_3$ ) with the same amount of potassium thiocyanate solution (0.3M KSCN). Add approx. 5  $\text{cm}^3$  of distilled water, mix and leave the beaker on a sheet of white paper. Observe the colour of the solution. Then dilute the solution with distilled water until the blood-red colour disappears. Divide the resulting straw-coloured solution into two parts. To one part of the solution, add 10 drops of iron(III) sulphate(VI) solution (0.1M  $\text{Fe}_2(\text{SO}_4)_3$ ), to the other, the same amount: 10 drops of potassium thiocyanate solution (0.3M KSCN). Observe the changes in the colour of the solutions in both beakers against the white background.

### Elaboration of the results:

1. Write the equation of the reaction taking place in the beaker leading to the formation of a blood-red colour.
2. What is the colour change after diluting the solution?
3. Write the dissociation reaction of the obtained salt and the formula for the constant K.
4. How do you explain the straw-coloured solution returning to red after adding both KSCN and  $\text{Fe}_2(\text{SO}_4)_3$ ?
5. Which of the added reagents gave rise to a more intense red colour, despite adding the same amount of  $\text{SCN}^-$  and  $\text{Fe}^{3+}$  ions?  
Explain why.

## Experiment 4 – Acid and base dissociation reaction – shifting the equilibrium of the reaction

### Materials and reagents:

Test tube rack, measuring cylinder, micro spatula, acetic acid (0.1M CH<sub>3</sub>COOH), solid: sodium acetate (CH<sub>3</sub>COONa), ammonium hydroxide (0.1M NH<sub>4</sub>OH), solid: ammonium chloride (NH<sub>4</sub>Cl), methyl orange, phenolphthalein.

### Performance:

Pour about 4 cm<sup>3</sup> of acetic acid solution (0.1M CH<sub>3</sub>COOH) into two test tubes. Then add two drops of methyl orange solution to each test tube. Leave one test tube with acetic acid as a control sample, to the other add 1 micro spatula of sodium acetate (CH<sub>3</sub>COONa) and mix. Compare the colour of the obtained solution with the colour of the solution in the control sample. Pour about 4cm<sup>3</sup> of ammonium hydroxide solution (0.1M NH<sub>4</sub>OH) into the next two test tubes. To each of them, add one drop of phenolphthalein solution. Leave one sample with ammonium hydroxide as a control sample, add a micro spatula of ammonium chloride (NH<sub>4</sub>Cl) into the other one and mix the solution. Compare the colour of the obtained solution with the colour of the control sample.

### Elaboration of the results:

After writing the dissociation equations and dissociation constants for acetic acid and ammonium hydroxide, state how the equilibrium of the reaction in the acetic acid solution will shift after adding sodium acetate. Also state how the equilibrium of the reaction in ammonium hydroxide solution is shifted after the addition of ammonium chloride. Why does the colour change occur in both cases? Present the results in the form of a table, filling it in and attaching to the exercise report.

Table 8

Summary of results and conclusions for the experiment 6

No of test tube	Content	Chemical formula of the tested acid/base, dissociation reaction	Chemical formula of added salt and its dissociation reaction	Common-ion	Colour of the solution	1. How has the equilibrium of the reaction shifted? 2. Why has the colour changed in the solution?
1.	acetic acid + methyl orange * Control sample		–	–		–
2.	acetic acid + methyl orange + sodium acetate					
3.	ammonium hydroxide + phenolphthalein * Control sample		–	–		–
4.	ammonium hydroxide + phenolphthalein + ammonium chloride					

#### **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 task/additional 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/>



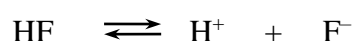
## I. Example of a task with a solution

### Example

Calculate the degree of dissociation of hydrofluoric acid in a 0.1 – molar solution if the  $\text{H}^+$  ion concentration is  $0.005 \text{ mol/dm}^3$ .

### Solution:

The degree of dissociation  $\alpha$  is the ratio of the number of  $N_Z$  dissociated particles to the number of  $N_W$  introduced particles (it can also be expressed as a percentage): It then informs us about the percentage of molecules that have decayed into ions. Hydrofluoric acid dissociates according to the equation:



One  $\text{H}^+$  ion is formed from each dissociating molecule. If the concentration of  $\text{H}^+$  ions is  $0.005 \text{ mol/dm}^3$ , it means that in each  $\text{dm}^3$  of solution  $0.005 \text{ mol}$  of HF molecules has dissociated. The number of particles introduced into  $1 \text{ dm}^3$  of the solution is  $0.1$  and the number of dissociated particles is  $0.005$ , hence:

$$\alpha = \frac{N_Z}{N_W} \times 100\% = \frac{0,005}{0,1} \times 100\% = 5\%$$

Answer: the dissociation degree is 5%

## II. Tasks and questions to be completed by the student

### Calculations regarding dissociation:

1. Calculate the degree of dissociation of hydrofluoric acid in a 0.1 molar solution, if the concentration of  $\text{H}^+$  ions is  $0.008 \text{ mol/dm}^3$ .
2. Calculate the degree of dissociation of 0.01 molar organic acid, in which the concentration of hydronium ions is  $5.2 \cdot 10^{-3} \text{ mol/dm}^3$ .
3. The degree of dissociation of hydrocyanic acid at a concentration of  $0.01 \text{ mol/dm}^3$  is  $2.7 \cdot 10^{-4}$ . Calculate the value of the acid constant of hydrogen cyanide.
4. Calculate the molar concentration of undissociated molecules in a monoprotic acid solution with concentration  $0.05 \text{ mol/l}$  ( $\alpha = 1.9\%$ ).
5. Calculate the degree of dissociation of monoprotic acid ( $K = 4 \cdot 10^{-4}$ ) in the development of the concentration of  $\text{H}^+$  ions equal to  $0.1 \text{ mol/dm}^3$ .

### Complex tasks regarding dissociation:

1. How much water should be added to  $0.5 \text{ dm}^3$  of 0.2 molar acetic acid solution ( $K = 1.8 \cdot 10^{-5}$ ) to double the degree of acid dissociation?
2. As the water evaporates from the electrolyte solution, does the degree of dissociation:
  - a) decrease,
  - b) increase,
  - c) remain constant?

### Applied chemistry task:

What is a better means for de-icing roads: sodium chloride or magnesium chloride – present a short comparative analysis of the problem using the knowledge and skills acquired during the exercise and explain why.