MARITIME UNIVERSITY OF SZCZECIN

Institute of Mathematics, Physics and Chemistry
Department of Chemistry

EXERCISE INSTRUCTION

Laboratory Exercise 3

pH of solutions
Hydrolysis of salts

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EXERCISE SHEET

<table>
<tr><td>1</td><td colspan="3">Relation to subjects: ESO/25, 27 DiRMiUO/25, 27 EOUNiE/25, 27</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/26 Chemistry of water, fuels and lubricants</td><td>EKP3
K_U014, K_U015, K_U016.</td><td>SEKP3 – Water quality indicators;
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>SEKP3 – Water quality indicators;
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>SEKP3 – Water quality indicators;
SEKP6 – Performing determinations of selected indicators of technical water quality;</td></tr>
<tr><td>2</td><td colspan="3">Purpose of the exercise:
mastering basic chemical concepts related to solutions of acids, bases and salts as well as acquiring practical skills in the field:
– determination of pH and measuring the pH of solutions,
– calculating the pH of solutions of strong and weak acids and bases,
– determining the reaction of individual types of salts after hydrolysis on the basis of the reaction,
– calculating the pH of specific salt solutions and buffer mixtures.</td></tr>
<tr><td>3</td><td colspan="3">Prerequisites:
general knowledge of pH, solution reaction, hydrolysis obtained from high school, knowledge of the key issues of ion dissociation acquired during the previous exercise, knowledge of the principles of work in a chemical laboratory</td></tr>
<tr><td>4</td><td colspan="3">Description of the laboratory workplace:
a set of laboratory glassware, a set of reagents and indicators for testing pH and hydrolysis,</td></tr>
<tr><td>5</td><td colspan="3">Risk assessment:
the likelihood of chemical burns from exposure to 0.2 M sulphuric acid is very small, and the effects are minor,
Final assessment – VERY SMALL THREAT
Security measures required:
1. Lab coats, gloves and safety glasses.
2. Health and safety cleaning products, paper towels.</td></tr>
<tr><td>6</td><td colspan="3">The course of the exercise:
1. Getting to know the workplace manual (appendix 1).
2. Carrying out the experiments provided for in the manual.</td></tr>
<tr><td>7</td><td colspan="3">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.</td></tr>
</table>

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 knowledge of basic chemical concepts of pH and hydrolysis in class, 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 a general knowledge of pH and salt hydrolysis, but is not able to use it in practice to solve basic problems, – mark 3,0 – has basic chemical knowledge of pH and salt hydrolysis and is able to use it to a small extent to solve potential problems in his specialty, – mark 3,5 – 4,0 – has extensive chemical knowledge in the field of pH, salt hydrolysis and is able to use it to a basic extent to determine the pH and calculate the pH of various electrolyte solutions and to solve problems on the ship, – mark 4,5 – 5,0 – has complete chemical knowledge in the field of pH and salt hydrolysis and is able to use in practice complex chemical knowledge to determine and calculate the pH of individual acid, alkali, salt and buffer solutions and to solve complex problems,
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: pH roztworów. Reakcje soli z wodą, 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:

- pH,
- salt hydrolysis.

pH

Water dissociation constant (K_w):

$$H_2O(l) + H_2O(l) \rightleftarrows H_3O^+_{(aq)} + OH^-_{(aq)} \text{ or simplified version:}$$

$$H_2O \rightleftarrows H^+ + OH^- \text{ (self-ionization of water)}$$

$K = [H_3O^+] \cdot [OH^-] / [H_2O] = 1{,}8 \cdot 10^{-16}$ at 25°C; molar concentration of water $[H_2O]$ has a constant value: 55.5 mol/dm^3:

$$[H_3O^+] \cdot [OH^-] / 55.5 = 1.8 \cdot 10^{-16}$$

$$[H_3O^+] \cdot [OH^-] = K_w \text{ (The } K_w \text{ value is constant!)} = 1.0 \cdot 10^{-14} \text{ in pure water}$$

$$\mathbf{[H_3O^+] = [OH^-] = 1.0 \cdot 10^{-7}}$$

The K_w value allows to convert from $[H_3O^+]$ to $[OH^-]$ and vice versa in any aqueous solution not just pure water.

The pH is a measure of the acidity of the solution, the acidity is related to the concentration of the hydronium ion $[H_3O^+]$ in the solution. The more acidic the solution the larger concentration of the hydronium ion.

The pH scale is based on the molar concentration of the hydronium ion $[H_3O^+]$ in the solution and ranges from **0 to 14** for most of the practical applications. There are values that are lower than 0 (very acidic solution) and higher than 14 (concentrated aqueous bases).

The pH is defined as the negative logarithm (log) of the $[H_3O^+]$:

$$\mathbf{pH = -\log [H_3O^+]; \quad [H_3O^+] = 10^{-pH}}$$

$$\mathbf{pOH = -\log [OH^-]; \quad [OH^-] = 10^{-pOH}}$$

$$\mathbf{pH + pOH = 14}$$

The pH of pure water equals 7 (**pH=7**): $[H_3O^+] = [OH^-] = 1.0 \cdot 10^{-7}$.

Acidic solution has a pH value lower than 7 (**pH < 7**); Basic solution has a pH value higher than 7 (**pH >7**).

pH indicators are substances that change the color in the presence of a base or an acid. Examples of pH indicators:

- litmus paper (when the solution is acidic litmus turns red; when the solution is basic litmus turns blue),

- phenolphthalein (clear and colorless when the solution is acidic, pink in a basic solution),
- methyl orange (red in acidic solution, yellow in basic solution),
- universal indicator (the colors from yellow to red indicate an acidic solution; the colors from blue to violet indicate a basic solution, green color indicates the neutral solution).

Buffer solutions (buffers) – prevent a change in pH that is caused by the addition of acid or base. Buffer must contain something that reacts with an acid – a base and also something that reacts with a base – an acid. Buffer is a solution of a weak acid and a salt of its conjugate base and a solution of a weak base and a salt of its conjugate acid.

In other words, an acidic buffer (for example acetic acid and sodium acetate) solution is made from a weak acid and one of its salts – usually a sodium salt. An alkaline buffer (for example ammonia plus chloride solution) solution is made from a weak base and one of its salts.

pH calculation of a buffer of weak acid and conjugate base (weak acid plus one of its salts):

$$[H_3O^+] = K_a \times \text{Molarity of a weak acid / Molarity of its salt}$$
$$pH = -\log [H_3O^+]$$

Example: Determine the pH of a buffer of 0.5 M of an acetic acid (HAc) and 0.3 M of sodium acetate (Ac^-). Given: $K_a = 1.7 \cdot 10^{-5}$.
$[H_3O^+] = 1.7 \cdot 10^{-5} \cdot 0.5 / 0,3 = 2.8 \cdot 10^{-5} mol/dm^3$
$pH = -\log [0.000028] = 4.55$

Henderson – Hassellbach Approximation can be also used for calculation pH of an acid – salt buffer:

$$pH = pK_a + \log ([A^-] / [HA])$$
$$pH = 4.77 + \log (0.3 / 0,5) = 4.55$$

pH calculation of a buffer of weak base and conjugate acid (weak base plus one of its salts):

$$[H_3O^+] = 10^{-14} \cdot \text{Molarity of a salt} / K_b \times \text{Molarity of a weak base}$$
$$pH = -\log [H_3O^+]$$

Hydrolysis of salts

Hydrolysis reaction is the reverse reaction to neutralization reaction, salts react with water to give back an acid and a base – a simplified definition.

In other words **hydrolysis of salts** is the reaction between water and aqueous ion, yields to a basic or acidic solution by forming H_3O+ (or simplified version H+) or OH^- ions.

$A^- + H_2O \leftrightarrow HA + OH^-$ (basic solution) salts formed from strong bases and weak acids
$M^+ + H_2O \leftrightarrow MOH + H^+$ (acidic solution) salts formed from strong acids and weak bases

Salts formed from strong acids and strong bases do not hydrolyze!
Salts formed from strong acids and weak bases, weak acids and strong bases and finally weak acids and weak bases hydrolyze!

Example 1 (A worked example): write a hydrolysis reaction of $MgCl_2$ (a salt formed from a strong acid and a weak base):

Solution:

a) the first step: write a dissociation reaction of a given salt: $MgCl_2 \rightarrow Mg^{2+} + 2Cl^-$

b) the second step: pick an ion that is a weak part of the given salt: Mg^{2+}

c) write a hydrolysis reaction: $\mathbf{Mg^{2+} + 2H_2O \leftrightarrow Mg(OH)_2 + 2H^+}$ acidic solution

Example 2 (A worked example): write a hydrolysis reaction of Na_2CO_3 (a salt formed from a weak acid and a strong base):

Solution:

a) the first step: write a dissociation reaction of a given salt:

$$Na_2CO_3 \rightarrow 2Na^+ + CO_3^{2-}$$

b) the second step: pick an ion that is a weak part of the given salt: CO_3^{2-}

c) write a hydrolysis reaction: $\mathbf{CO_3^{2-} + 2H_2O \leftrightarrow H_2CO_3 + 2OH^-}$ basic solution

Example 3 (A worked example): write a hydrolysis reaction of $MgCO_3$ (a salt formed from a weak acid and a weak base):

Solution:

a) the first step: write a dissociation reaction of a given salt:

$$MgCO_3 \rightarrow Mg^{2+} + CO_3^{2-}$$

b) the second step: pick an ion that is a weak part of the given salt: CO_3^{2-}; Mg^{2+}

c) write a hydrolysis reaction: $\mathbf{CO_3^{2-} + 2H_2O \leftrightarrow H_2CO_3 + 2OH^-}$

$\mathbf{Mg^{2+} + 2H_2O \leftrightarrow Mg(OH)_2 + 2H^+}$

Examples of selected calculations related to concentration of $[H_3O^+]$, $[OH^-]$ and pH (with a solution)

Example 1 What are the $[H_3O^+]$ and $[OH^-]$ of a solution with pH of 4?

$$pH = 4 \quad [H_3O^+] = 10^{-4}\ mol/dm^3$$
$$pOH = 14 - 4 = 10\ [OH^-] = 10^{-10}\ mol/dm^3$$

Example 2 What are the pH and pOH of a solution with a $[OH^-]$ of $4.5 \cdot 10^{-2}$ mol/dm^3?

$$pOH = -\log [4.5 \cdot 10^{-2}] = 1.35$$
$$pH = 14 - 1.35 = 12.65$$

Example 3 Determine the pH of a 0.1 mol/dm^3 solution of HCl, a strong acid.

$$HCl \longrightarrow H^+ + Cl^-$$

Strong acid: $[H_3O^+]$ = Given concentration of strong acid solution

$$[H_3O^+] = 0.1\ mol/dm^3$$
$$pH = -\log [0.1] = 1$$

Example 4 Determine the pH of a 0.01 mol/dm^3 solution of NaOH and $Ba(OH)_2$, strong bases. Strong base: $[OH^-]$ = coefficient in front of the hydroxide ion (OH^-) x given concentration of strong base solution

$$NaOH \longrightarrow Na^+ + OH^- \quad [OH^-] = 1 \cdot 0.01\ mol/dm^3 = 0.01\ mol/dm^3\ pOH = -\log[0.01] = 2;$$
$$pH = 14 - 2 = 12$$
$$Ba(OH)_2 \longrightarrow Ba^{2+} + 2OH^- \quad [OH^-] = 2 \cdot 0.01\ mol/dm^3 = 0.02\ mol/dm^3$$
$$pOH = -\log[0.02] = 1.70;\ pH = 14 - 1.70 = 12.3$$

Example 5 Determine the pH of a solution of CH_3COOH (acetic acid), weak acid. Given: molar concentration of acetic acid (c; M) = 0.4 mol/dm^3; $K_a = 1.8 \cdot 10^{-5}$

c/K = 0.4/0.000018 = 22222 simplified version of law of dilution:

$K = c \cdot \alpha^2$ $\alpha = (K/c)^{0.5} = 0.0067$

α = the molar concentration of dissociated electrolyte / the total molar concentration of electrolyte [mol/dm^3; mol/l].

$0.0067 = x$ / 0.4 mol/dm^3 $x = 0.00268$ mol/dm^3 pH = – log [x] = – log [0.00268] = 2.57

Tasks and questions to be performed by the student:

1. What are the $[H_3O^+]$ and $[OH^-]$ of a solution with pH of 6? Answer: $[H_3O^+]=10^{-6}$mol/dm^3; $[OH^-]=10^{-8}$mol/dm^3
2. What are the pH and pOH of a solution with a $[H_3O^+]$ of $2.1 \cdot 10^{-2}$ mol/dm^3? Answer: pH=1,7; pOH=12,3
3. What are the pH and pOH of a solution with a $[OH^-]$ of $5.5 \cdot 10^{-5}$ mol/dm^3? Answer: pH=9,74; pOH=4,26
4. Determine the pH of a 0.005 mol/dm^3 solution of $HClO_4$. Answer: pH=2,30
5. Determine the pH of a 0.01 mol/dm^3 solution of KOH. Answer: pH=12,0
6. Determine the pH of a 0.01 mol/dm^3 solution of $Sr(OH)_2$. Answer: pH=12,3
7. Determine the pH of a 0.01 mol/dm^3 solution of CH_3COOH; $K_a = 1.8 \cdot 10^{-5}$. Answer: pH=3,37
8. Determine the pH of a 0.01 mol/dm^3 solution of NH_4OH; $K_b = 1.8 \cdot 10^{-5}$. Answer: pH=10,63
9. Write a hydrolysis reaction of $BaCO_3$ and state whether a salt solution is acidic, basic or neutral.
10. Write a hydrolysis reaction of K_2S and $ZnSO_4$ and state whether salt solutions are acidic, basic or neutral.

2. INSTRUCTION 3 – LABORATORY EXERCISE 3

Experiment 1 – Acid / Base indicators (Identifying acids and bases with indicators)

Materials and reagents:

Glass test tube set, measuring cylinder, acetic acid solution (0.1M CH_3COOH), ammonium hydroxide solution (0.1M NH_4OH), pH indicators: methyl orange, methyl red, litmus, phenolphthalein, universal indicator.

Experimental procedure:

Pour 4 cm^3 of acetic acid solution (0.1M CH_3COOH) to each of five test tubes, into the next five test tubes pour 4 cm^3 of ammonium hydroxide solution (0.1M NH_4OH). Test tubes with the ammonium hydroxide solution should be placed behind the test tubes with acetic acid solution on the test tube rack. Add to each pair of test tubes (acid and base) 3 drops of indicator given in the table.

Test tube	pH indicator	The colour of the acetic acid solution (0.1M CH_3COOH)	The colour of the ammonium hydroxide solution (0.1M NH_4OH)	pH ranges of pH indicators
1.	Methyl orange			
2.	Methyl red			
3.	Litmus			
4.	Phenolphthalein			
5.	Universal			

Data analysis (after the experiment):

1. Fill in the given table for the pH indicators.
2. Calculate the pH of acetic acid solution. Given: $K = 1.8 \cdot 10^{-5}$; $[CH_3COOH] = 0.01$ M.
3. Calculate the pH of ammonium hydroxide solution. Given: $K = 1.8 \cdot 10^{-5}$; $[NH_4OH] = 0.01$ M.

Experiment 2 – pH of salt solutions

Materials and reagents:

Glass test tube set, measuring cylinder, micro spatula, pH indicators: universal, phenolphthalein, selected salt solutions: sodium sulfide (2M Na_2S), zinc chloride (2M $ZnCl_2$), solids: copper (II) sulfate ($CuSO_4$), sodium carbonate (Na_2CO_3), potassium nitrate (KNO_3), sodium sulfate (Na_2SO_4), sodium acetate (CH_3COONa), ammonium chloride (NH_4Cl).

Experimental procedure:

Pour 4 cm^3 of distilled water to each of nine test tubes. Add to each of the test tube 5 drops of the universal indicator solution. The first test tube treat as a control sample and add to the rest of the test tubes small amount of salts listed in the given table. Do not mix it. Record the color change of the universal indicator. Repeat the procedure using phenolphthalein as a pH indicator.

Test tube	Salts plus control sample	The colour change of the universal indicator	pH value (see the table below)	Acidic/Basic/neutral solution (state whether solution is acidic, basic or neutral)	The colour change of the phenolphthalein
1.	Control sample				
2.	Copper (II) sulfate ($CuSO_4$)				
3.	Sodium carbonate (Na_2CO_3)				
4.	Potassium nitrate (KNO_3)				
5.	Sodium sulfide (2M Na_2S)				
6.	Zinc chloride (2M $ZnCl_2$)				
7.	Sodium sulfate (Na_2SO_4)				
8.	Sodium acetate (CH_3COONa)				
9.	Ammonium chloride (NH_4Cl)				

pH value	The colour of the universal indicator (according to Yamada – based on pH value)
4.0	Red
5.0	Orange
6.0	Yellow
7.0	Green
8.0	Blue
9.0	Dark blue
10.0	Violet

Data analysis (after the experiment):

1. Fill in the given table.
2. For each given salt identify the acid and the base from which this salt was formed; write the neutralization reaction (acid + base = salt + water) and balance it.
3. Write the hydrolysis reactions (salt + water) – molecular and ionic form of the reaction (molecular and ionic equation).
4. What type of salts hydrolyze?

Experiment 3 – The effect of temperature on hydrolysis reaction

Materials and reagents:

Glass test tube set, measuring cylinder, test tube holder, solution of sodium acetate (0.1 M CH_3COONa), phenolphthalein.

Experimental procedure:

Pour 4 cm^3 of sodium acetate (0.1 M CH_3COONa) solution into the test tube and add 2 – 3 drops of phenolphthalein. Heat the solution in the laboratory water bath and then cool it down. Record the obtained colors of the solution.

Data analysis (after the experiment):

1. Write the hydrolysis reaction (CH_3COONa + water) – molecular and ionic equation.
2. Explain the effect of temperature on the reaction of interest.

Experiment 4 – pH of a buffer solution

Materials and reagents:

Glass beakers, measuring cylinder, solutions of sodium acetate (0.1 M CH_3COONa), acetic acid (0.1M CH_3COOH), ammonium chloride (0.1M NH_4Cl), ammonium hydroxide (0.1M NH_4OH), pH indicators: litmus strips, pH color strips, universal pH indicator.

Experimental procedure:

Pour 4 cm^3 of ammonium chloride (0.1M NH_4Cl) and 8 cm^3 of ammonium hydroxide (0.1M NH_4OH) into the first beaker. Pour 4 cm^3 of sodium acetate (0.1 M CH_3COONa) and 8 cm^3 of acetic acid (0.1M CH_3COOH) into the second beaker. Determine the pH of the buffer solutions using pH color strips (different pH ranges: pH:1 – 7; pH: 7 – 14), pH meter and also determine the color of the buffer solutions after adding a few drops of universal indicator and also after immersing the litmus strip in the buffer solution for 2 seconds.

How to use pH color strips? Immerse the strip in the test solution for a few seconds then compare the developed color to the sequence chart on the package and read the pH.

Data analysis (after the experiment):

Fill in the table below

Testing methods for measuring pH values	Beaker I (alkaline buffer solution) colour of a buffer solution	Beaker I pH	Beaker II (acidic buffer solution) colour of a buffer solution	Beaker II pH
Litmus strip (litmus paper)		–		–
pH color strips (pH ranges: 1 – 7 and 7 – 14)	–		–	
Universal pH indicator				
pH meter	–		–	
Calculation of the pH value (obtained buffer solution)	–		–	

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.