Institute of Mathematics, Physics and Chemistry

Department of Chemistry

Technical chemistry laboratory

Laboratory exercise

Oxidation and reduction reactions in solutions

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EXERCISE SHEET

1	**Relation to subjects:** ESO/25, 27 DiRMiUO/25, 27 EOUNiE/25, 27		
	Specialty/Subject	**Learning outcomes for the subject**	**Detailed learning outcomes for the subject**
	ESO/26 Chemistry of water, fuels and lubricants	EKP3 K_U014, K_U015, K_U016.	SEKP3 – Water quality indicators; 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.	SEKP3 – Water quality indicators; SEKP6 – Performing determinations of selected indicators of technical water quality;
	Chemistry of water, fuels and lubricants	EKP3 K_U014, K_U015, K_U016.	SEKP3 – Water quality indicators; SEKP6 – Performing determinations of selected indicators of technical water quality;
2	**Purpose of the exercise:** deepening and expanding the chemical knowledge of the oxidation and reduction processes as well as the ability to balance redox reactions and identify the oxidant and reducer.		
3	**Prerequisites:** general knowledge, obtained from high school, on the principles of determining the oxidation states of elements, balancing redox reactions and recording half reactions, knowledge of the principles of work in a chemical laboratory.		
4	**Description of the laboratory workplace:** a set of laboratory glassware, a set of reagents for carrying out redox reactions, indicators.		
5	**Risk assessment:** the likelihood of chemical burns from exposure to 0.2 M sulfuric 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.		
6	**The course of the exercise:** 1. Getting to know the workplace manual (appendix 1), 2. Performing a redox reaction in solution.		
7	**Exercise report:** 1. Prepare a report 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 knowledge of basic chemical concepts related to oxidation and reduction processes during classes;		

	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 about the redox reaction in the solution, or is unable to use it in practice; – mark 3,0 – has basic chemical knowledge of redox reaction in solution and is able to use it to a small extent; – mark 3,5 – 4,0 – has an extensive knowledge of chemical and redox reactions in solution and is able to use it to solve basic potential technical problems arising in his specialty; – mark 4,5 – 5,0 – has the ability to apply complex chemical knowledge of oxidation and reduction processes and is able to use it to solve complex technical problems.
10	**Literature:** 1. Kozłowski A., Gabriel-Półrolniczak U., Workplace instruction for laboratory exercises *Reakcje redoks w roztworze*, AM Szczecin, 2013. 2. Stundis H., Trześniowski W., Żmijewska S.: *Ćwiczenia laboratoryjne z chemii nieorganicznej*. WSM, Szczecin 1995. 3. Jones L., Atkins P., *Chemia ogólna, Cząsteczki, materia reakcje*, WN PWN, Warsaw 2004. 4. Bielański A., *Chemia ogólna i nieorganiczna*, PWN, Warsaw, 1996. 5. Śliwa A., *Obliczenia chemiczne*, PWN, Warsaw 1987. 6. Kozłowski A., *Materiały dydaktyczne z chemii technicznej*, developed for the purposes of auditorium classes (not published). 7. Multimedia presentation „Redox reactions" from the resources of the AGH e-learning center in Krakow: http://zasoby1.open.agh.edu.pl/dydaktyka/chemia/a_e_chemia/filmy/wmv/ 8. Cox P.A., Short lectures. Inorganic Chemistry, PWN, Warsaw, 2006. 9. http://www.zmnch.pl/files/chlasts/Wykaz_niezgodnych_substancji_chemicznych.pdf
11	Notes

APPENDIX 1 – MANUAL

1. SCOPE OF THE EXERCISE

Issues and keywords:

- oxidation and reduction reactions (definition of the oxidation and reduction process as well as oxidizing agent and a reducing agent (oxidant and reducer), rules for determining the oxidation state of atoms in molecules and ions);
- balancing redox reactions (molecular and ionic), determining the oxidizing and reducing agent;
- oxidation-reduction potential (definition, Nernst equation);
- typical oxidizing and reducing agents and their application in life and technology.

2. THEORETICAL INTRODUCTION TO THE EXERCISE

2.1. Oxidation and reduction reactions

Redox reactions, i.e. oxidation-reduction reactions, consist in changing the oxidation state of the elements that make up the reacting substances as a result of the exchange of electrons between the oxidizing substance and the reducing substance.

Initially, the term oxidation defined the process of adding oxygen, and reduction – the loss of oxygen by the substance. The similarities between the joining reaction with oxygen, the joining reactions with halides (F_2, Cl_2, Br_2, I_2), sulphur and other non-metals meant that these processes were also called oxidation.

In the new broader approach, oxidation is understood as a process in which the oxidation state of an element increases (associated with the donation of electrons), while reduction as a process in which a decrease in the oxidation state of an element (associated with the collection of electrons) occurs. An oxidant is a substance that is reduced by taking electrons from the oxidized substance. A reductant is a substance that oxidizes as a result of donating electrons of the reduced substance.

The oxidation of one substance must always be accompanied by the reduction of another substance, and vice versa. Hence, the oxidation and reduction reactions constitute a coupled system, called the redox system, where the number of electrons attached by the oxidant is equal to the number of electrons given off by the reducer.

$$\begin{aligned} &\text{reducer 1} && = \text{oxidant 1} + ne^- \\ &\text{oxidant 2} + ne^- && = \text{reducer 2} \end{aligned}$$

$$\text{reducer 1} + \text{oxidant 2} = \text{oxidant 1} + \text{reducer 2}$$

Substances containing elements in intermediate oxidation states may exhibit oxidizing properties in the presence of a stronger reducing agent, and reducing properties in the presence of a stronger oxidant. Thus, the redox properties depend not only on the nature of the substance itself, but also on the environment and the abundance of other substances with redox properties in it.

The oxidation reaction is always accompanied by the reduction reaction, but for the sake of convenience, you can write separately the oxidation and reduction reactions by means of oxidation and reduction half reactions.

For example, the reaction:

$$Fe + 2HCl \longrightarrow FeCl_2 + H_2\uparrow$$

can be written using half reactions:

$$Fe_{(s)} \longrightarrow Fe^{2+}_{(aq)} + 2e^- \qquad \text{oxidation reaction}$$

$$2H^+_{(aq)} + 2e^- \longrightarrow H_{2(g)}\uparrow \qquad \text{reduction reaction}$$

Correct recording of the redox reaction requires balancing the changes in the oxidation states of all elements (the sum of the oxidation states on both sides of the reaction must be the same).

2.2. Oxidation state

The oxidation state of an element in a chemical compound is a hypothetical charge assigned to an atom of the element, assuming that the bonds formed are ionic. The following rules apply when determining the oxidation state:

1. the sum of the oxidation states of all the atoms making up the neutral molecule is zero;
2. the sum of the oxidation states of all the atoms that make up the ion equals the charge of the ion;
3. elements in the free state take the oxidation state 0, regardless of whether the substance is in the form of single atoms or in the form of molecules, e.g. Fe^0, N_2^0, O^0, O_2^0, O_3^0;
4. fluorine in all compounds occurs in the oxidation state –I;
5. combined oxygen most often occurs in the oxidation state –II, with the exception of peroxides, e.g. H_2O_2, in which it takes the oxidation state –I and OF_2, in which oxygen takes the oxidation state II;
6. hydrogen usually assumes oxidation state I in compounds; exceptions are alkali and alkaline earth hydrides in which its oxidation state is equal to –I;
7. the oxidation state of alkali metals is I, and alkaline earths is II in chemical compounds.

2.3. Balancing redox reactions

The easiest way to balance redox reactions is to use half reactions. The following strategy is used:

1. the redox reaction equation is written in skeleton form (giving the reagents without stoichiometric coefficients);
2. the ionic reaction equation is written from the skeleton equation in molecular form;
3. analysing the changes in the oxidation state of elements, the reducer and the oxidant are determined and the oxidation and reduction reactions are recorded;
4. balances the number of electrons replaced in both half reactions by multiplying them by a factor that allows obtaining the same number of electrons transferred in the oxidation-reduction reaction;
5. the coefficients in the half-equations are transferred to the molecular skeletal equation of the reaction;
6. the factors for the remaining reagents are completed to balance the number of atoms of all elements on both sides of the equation; if necessary, products made of elements that do not change the oxidation state should be added.

2.4. Oxidants and reducers (oxidizing and reducing agents)

Oxidants are substances that take electrons from the substance to be oxidized. The group of typical oxidants includes the most electronegative elements: fluorine F_2, chlorine Cl_2, bromine Br_2, and oxygen O_2. The role of oxidants is also played by those compounds in which elements adopting different oxidation states occur in the highest oxidation states, e.g., nitric acid(V) HNO_3, potassium nitrate(V) KNO_3, potassium manganate(VII) $KMnO_4$, potassium chromate(VI) $K_2Cr_2O_7$, iron(III) chloride $FeCl_3$, hydrogen peroxide H_2O_2.

Reducers are substances that have the ability to donate electrons to the reduced substance. The group of typical reducing agents includes the most electropositive elements, e.g., metals from groups I and II of the periodic table, such as sodium, Na or magnesium, Mg, as well as hydrogen and carbon. Chemical compounds made of metals or non-metals with a lower oxidation state possible for a given element are often used as reducing agents, e.g. iron(II) chloride $FeCl_2$, sulphuric acid(IV) H_2SO_3, sodium nitrate(III) $NaNO_2$, potassium iodide KI, carbon monoxide CO.

Many chemical compounds can act as an oxidant or a reducing agent, depending on what chemicals are in its environment. An example would be hydrogen peroxide H_2O_2:

$$2MnO_4^- + 5H_2O_2 + 6H_3O^+ \longrightarrow 2Mn^{2+} + 5O_2 + 14H_2O$$
$$2I^- + H_2O_2 + 2H_3O^+ \longrightarrow I_2 + 4H_2O$$

In the first reaction, hydrogen peroxide acts as a reducing agent in relation to the stronger oxidant, which is potassium manganate(VII). In the second reaction, hydrogen peroxide acts as an oxidant in relation to the iodide ions I^-. This is due to the presence of oxygen in the oxidation state –I in hydrogen peroxide, which can be reduced to the oxidation state –II, or oxidized to molecular oxygen, changing to the oxidation state 0. What role each factor in the reaction takes depends mainly on the value of the redox potentials of both factors. Similarly, the atoms of some elements, depending on the oxidation state in a given compound, may only act as an oxidant or only a reducing agent in the reactions, or be a reducing agent in some reactions, and an oxidant in others, as shown by the example of sulphur atoms in various oxidation states.

Table 1 gives examples of common oxidants and reducers.

Table 1

Common oxidants and reducers

Oxidants (oxidizing agents)	**Reducers (reducing agents)**
High electronegativity in the free state (F_2, O_2, O_3, Cl_2, Br_2), PbO_2, HNO_3, H_2SO_4 concentrated, $HClO_3$, $HClO_4$, $KMnO_4$, $H_2Cr_2O_7$, metal peroxides, H_2O_2 (in this case, the oxidation-reduction properties depend on the reaction medium and the remaining reactants)	metals with low electronegativity (Na, Mg), C, CO, hydrogen, some ions (Cl^-, Br^-, SO_3^{2-})

2.5. Oxidation-reduction potential (redox potential)

Redox potential is a measure of a molecule's ability to accept or donate electrons. The difference in potentials of the two redox pairs (consisting of an oxidized and reduced particle) determines the direction of the redox reaction in a solution. The redox potential of a system is given by the formula:

$$E = E_0 + \frac{RT}{nF} \ln \frac{C_{utl.}}{C_{red.}}$$

where:

E_0	–	standard redox potential of a system,
$C_{utl.}$	–	concentration of the oxidized form,
$C_{red.}$		concentration of the reduced form,
T	–	temperature [K],
R	–	gas constant: 8.31 [J/mol · K],
F	–	Faraday constant 96500 [C/mol].

In a wide range of concentrations, we can closely compare directly the standard redox potentials of the system (oxidant-reducer pairs). This redox pair, which has a lower standard potential E_0, will oxidize, i.e., it will act as a reducing agent.

2.6. The importance of redox processes in life and technology

We can observe redox processes in everyday life. Some of them are undesirable or even harmful. An example of such processes is the corrosion of metals, i.e., the process of destroying them under the influence of external factors, or the rancidity of fats. Others are simply irreplaceable, without which it is difficult to imagine our daily life - they are the basis of many biochemical processes taking place in organisms, e.g., photosynthesis and respiration. Other such processes include combustion reactions, e.g., of natural gas or coal, as well as liquid fuels. Redox processes are widely used and have great advantages in technology and industry. An example is the use of oxidizing (bleaching) agents, cleaning and disinfecting water and sewage, and many others. One application of redox processes is rocket technology.

Liquid or solid fuels may be used to propel rocket engines. The most common ingredients for the production of fuels are: hydrazine or its derivatives, kerosene (kerosene), liquid oxygen, liquid hydrogen, hydrogen peroxide (ignolin), nitric acid, boranes and other. The rocket is filled with fuel (liquid or solid) and an oxidizer. The fuel and oxidizer mix and burn in the combustion chamber. When fuel is burned, a redox reaction occurs. The oxidizing agent is necessary in the rocket in order for the fuel to be burned most efficiently. The large amount of exhaust gas generated during the combustion of engine exhaust nozzles at a speed of more than 2000 m/s, according to the law of conservation of momentum, causes acceleration and movement of the rocket in the opposite direction to the gas outlet.

Another example is the use of bleaching fabrics in the textile and paper industries. Bleaching is the process of removing the natural colour of textile fibres, yarns, fabrics, pulp and other products. Some dyes can be removed by oxidizing agents such as bleach. The most commonly used bleaching agents are chlorine compounds, hydrogen peroxide, sodium perborate, and potassium permanganate. Most of these dyes can be removed with a reducing agent such as sulphur dioxide.

Redox reactions are often used in the wastewater treatment process. At the waste water treatment site, they undergo several subsequent processes to reduce pollutants and toxic substances. The removal of organic substances present in sewage is achieved by accelerating biological processes occurring in nature in the reactions of biological oxidation carried out with the use of the so-called „Activated sludge”, which contains many aerobic bacteria.

Redox processes are also used in water treatment to remove iron(II), manganese(II), hydrogen sulphide, sulphides, substances responsible for the colour, taste and smells of water, phyto- and zooplankton, organic compounds, both of anthropogenic and natural origin. In addition, they allow for the destabilization of colloids or the elimination of the risk of microbial contamination by disinfecting the water. Oxidation can be used in several stages of the process line as preliminary, intermediate and final oxidation. Typical oxidants used in water treatment include air (more precisely its oxygen content), chlorine (and chlorine compounds) and ozone.

Free radicals are very reactive atoms or groups of atoms that contain an unpaired electron. They are assigned an important role in the development of a number of diseases, such as arthritis, heart disease, etc. Free radicals can damage living cells, attack adipose tissue, proteins and nucleic acids. To prevent damage caused by free radicals, so-called "antioxidants" are used. It is a type of molecule that can neutralize free radicals by providing them with electrons (redox process). Examples of commonly used antioxidants are vitamin C (ascorbic acid), vitamin E, and beta-carotene.

Antioxidants are also often used in industry to increase the shelf life. These include, among others: BHT (butylated hydroxytoluene) and BHA (butylated hydroxyanisole), which is added to food in a concentration below 1%.

In terms of the transport and storage of oxidizing and reducing substances, it should be remembered that the combination of reducing and oxidizing agents may cause violent reactions, such as self-ignition, production of toxic gases or explosion. In this context, it is important to prevent the contact of incompatible substances through their separation and appropriate storage. Examples of incompatible substances, accidental contact of which may cause spontaneous combustion or explosion, include iodine and ammonia, ammonium nitrate and powdered metals, concentrated nitric acid and sulphides or metals, copper and hydrogen peroxide.

3. PERFORMING THE EXERCISE

Materials and reagents (for all experiments 1 – 4):

Rack with test tubes, micro spatula, polyethylene plastic pipettes, measuring cylinder with a capacity of 50 cm^3, solutions: potassium hexacyanoferrate(III) (0.1 M $K_3[Fe(CN)_6]$), iron(II) sulphate(VI) (5% $FeSO_4$), hydrogen peroxide (3% H_2O_2), potassium manganate(VII) (0.1 M $KMnO_4$), potassium bromide (5% KBr), potassium iodide (5% KI), potassium chloride (5% KCl), sodium sulphide (2 M Na_2S), iron(III) sulphate(VI) (0.1 M $Fe_2(SO_4)_3$), iodine solution in potassium iodide (I_2 w KI), sulphuric(VI) acid (H_2SO_4) (1 : 3), sodium hydroxide (1 M NaOH), chloroform (trichloromethane) ($CHCl_3$), solids: sodium sulphate(IV) (Na_2SO_3).

Experiment 1 – Determining the oxidant (oxidizing agent) and reducer (reducing agent) based on the characteristic features of the reaction products

Performance:

Pour into three test tubes successively:

a) test tube one – 4 cm^3 of potassium manganate(VII) (0.1 M $KMnO_4$) and acidify it with dilute sulphuric(VI) acid (H_2SO_4) (1 : 3),
b) test tube two – 4 cm^3 of solution of iodine in potassium iodide (I_2 w KI),
c) test tube three – 4 cm^3 of potassium iodide solution (5% KI) acidified with dilute sulphuric acid(VI) (H_2SO_4) (1 : 3).

Then divide the solutions in all test tubes into two parts.
To the first part (test tubes with numbers 1, 2, 3) add a few drops of the following solutions successively until a clear reaction occurs:

- to the first– solution of iron(II) sulphate(VI) (5% $FeSO_4$),
- to the second – sodium sulphide solution (2 M Na_2S),
- to the third – hydrogen peroxide solution (3% H_2O_2) (attention – 1 drop).

Leave the remaining parallel test tubes (1a, 2a, 3a) as reference test tubes.

Elaboration of the results

1. Write the balanced chemical equations of the three performed redox reactions.
2. Fill in the Table 2.

Table 2

Table with the results of experiment 1

	Test tube 1	Test tube 2	Test tube 3
Reagent I	$KMnO_4$	I_2 w KI	KI
Reagent II	$FeSO_4$	Na_2S	H_2O_2
Reaction			
Initial colour/ colour source			
Final colour/ colour source			
Oxidant (oxidizing agent)			
Reducer (reducing agent)			

Experiment 2 – Oxidation – reduction reactions in acidic/basic/neutral solution on the example of potassium manganate(VII) $KMnO_4$

Performance:

Pour about 4 cm^3 of potassium manganate(VII) (0.1 M $KMnO_4$) into four test tubes. Add to the next three:

- to the first tube – 1 cm^3 of sodium hydroxide (1 M NaOH),
- to the second – 1 cm^3 of distilled water,
- to the third – 1 – 3 drops of sulphuric acid(VI) (H_2SO_4) (1 : 3),

Leave the fourth test tube as a reference test tube.

In the first test tube, the solution is alkaline (basic), in the second – neutral, and in the third – acidic. To the prepared solutions (test tubes 1, 2, 3) add a pinch of solid sodium sulphate(IV) (Na_2SO_3). After the reaction, compare the final colour with the colour of the reference test tube.

Elaboration of the results

Write balanced chemical equations of the reactions, knowing that in an alkaline environment manganese changes its oxidation number from VII to VI, in neutral from VII to IV and finally in acidic from VII to II.

Experiment 3 – Investigation of the influence of the oxidation-reduction potential on the course of the redox reaction

To two test tubes, add successively 4 cm^3 of:

- to test tube 1 – potassium manganate(VII) solution (0.1 M $KMnO_4$),
- to test tube 2 – iron(III) sulphate(VI) solution (0.1 M $Fe_2(SO_4)_3$).

Then, acidify each of the solutions in the test tube with dilute sulphuric(VI) acid (H_2SO_4) (1 : 3) and divide into three parts (3 additional test tubes) and, to each of them, add successively the solution of: potassium iodide (5% KI), potassium bromide (5% KBr) and potassium chloride (5% KCl), until the colour changes significantly.

Elaboration of the results

Write the ionic chemical equation of the redox reaction if it is known that the oxidants (oxidizing agents) are reduced to Mn^{2+}, Fe^{2+} ions, and the reducing agents are oxidized to I_2, Br_2 and Cl_2.

List the results of the experiment in a table by inserting „+” where the reaction takes place and „–” where it does not:

Reagents	KJ	KBr	KCl
$KMnO_4$			
$Fe_2(SO_4)_3$			

$E_{MnO_4^-/Mn^{2+}}$	= 1.52 V	$E_{Cl_2/2Cl^-}$	= 1.40 V
$E_{Br_2/2Br^-}$	= 1.08 V	$E_{Fe^{3+}/Fe^{2+}}$	= 0.75 V
$E_{I_2/2I^-}$	= 0.58 V		

On the basis of the values of the oxidation-reduction potential given above, justify why only some of the reactions took place.

Experiment 4 – Investigation of the influence of the oxidant (oxidizing agent) on the course of the redox reaction

Performance:

Pour about 4 cm^3 of potassium iodide solution (5% KJ) into a test tube and acidify with dilute sulphuric acid(VI) (H_2SO_4) (1 : 3) Then add one drop of hydrogen peroxide (3% H_2O_2) and about 3 cm^3 of chloroform ($CHCl_3$). Shake the contents of the test tube so that the released iodine goes to the chloroform layer.

Elaboration of the results

1. What role (oxidant/reducer) does hydrogen peroxide H_2O_2 play in the above reaction?
2. Justify your conclusion by writing the appropriate redox reaction.
3. What part of the tube is the chloroform layer in, what does that show?
4. How to justify the easy transfer of iodine to the chloroform layer (solubility)?
5. What is the name of the observed process of iodine transfer from the aqueous layer to chloroform and where is it applicable?

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. Examples of tasks with a solution

Example 1

Calculate the oxidation state of each element:

b) in the molecule $KMnO_4$
c) in ion HPO_4^{2-}

Solution:

a) Potassium, in accordance with rule 7, will assume the oxidation state in the analysed compound, and oxygen, in accordance with rule 5, in the oxidation state-II. It remains for us to calculate the degree of oxidation of the manganese:

b)

$$\overset{\mathrm{I}}{K}\ \overset{x}{Mn}\ \overset{-\mathrm{II}}{O_4}$$

Since, according to rule 1, the sum of the oxidation states of the atoms in the neutral molecule is zero, determining the manganese oxidation state in the analysed molecule comes down to solving the equation:

$1 \cdot 1 + 1 \cdot x + 4 \cdot (-2) = 0$
$x = 7$,
so

$$\overset{\mathrm{I}}{K}\ \overset{\mathrm{VII}}{Mn}\ \overset{-\mathrm{II}}{O_4}$$

c) Hydrogen according to rule 6 will have an oxidation state of I, and oxygen according to rule 5 will have an oxidation state of –II. So, we need to calculate the oxidation state of the phosphorus:

$$\overset{\mathrm{I}}{H}\ \overset{x}{P}\ \overset{-\mathrm{II}}{O_4}{}^{2-}$$

The entry „2–" in the upper right index says that we are dealing with an ion whose total charge is –2. To find the oxidation state of phosphorus in this ion, we need to solve the equation:

$1 \cdot 1 + 1 \cdot x + 4 \cdot (-2) = -2$
$x = 5$,
so

$$\overset{\mathrm{I}}{H}\ \overset{\mathrm{V}}{P}\ \overset{-\mathrm{II}}{O_4}{}^{2-}$$

Example 2

Balance the reaction:

$$KMnO_4 + FeSO_4 + H_2SO_4 \longrightarrow MnSO_4 + Fe_2(SO_4)_3 + H_2O$$

Solution:

1. the given skeletal reaction is written in molecular form; we write it in ionic form:

$$MnO_4^- \ + \ Fe^{2+} \ + \ H_3O^+ \longrightarrow Mn^{2+} \ + \ Fe^{3+} \ + \ H_2O$$

2. we identify the reducer and the oxidant. These are Fe^{2+} and MnO_4^-.

$$\overset{II}{Fe^{2+}} \rightarrow \overset{III}{Fe^{3+}} \qquad \text{reducer (oxidation)}$$

$$\overset{VII}{MnO_4^-} \rightarrow \overset{II}{Mn^{2+}} \qquad \text{oxidant (reduction)}$$

3. we write down the oxidation and reduction reactions:

$$\overset{II}{Fe^{2+}} = \overset{III}{Fe^{3+}} + e^- \qquad \text{oxidation}$$

$$\overset{VII}{MnO_4^-} + 5e^- = \overset{II}{Mn^{2+}} \qquad \text{reduction}$$

4. We multiply the equations by a factor that gives the same number of electrons transferred in the oxidation-reduction reaction:

$$\overset{II}{Fe^{2+}} = \overset{III}{Fe^{3+}} + e^- \quad / \cdot 5$$

$$\overset{VII}{MnO_4^-} + 5e^- = \overset{II}{Mn^{2+}} \quad / \cdot 1$$

We obtain:

$$5\,\overset{II}{Fe^{2+}} = 5\,\overset{III}{Fe^{3+}} + 5\,e^-$$

$$\overset{VII}{MnO_4^-} + 5e^- = \overset{II}{Mn^{2+}}$$

5. We enter the calculated coefficients into the molecular reaction

$$KMnO_4 \ + \ 5FeSO_4 \ + \ H_2SO_4 \longrightarrow MnSO_4 \ + \ 5Fe_2(SO_4)_3 \ + \ H_2O$$

Since after entering the coefficients on the right side of the equation we have 10 iron atoms (2 x more than calculated in the half equations), the remaining coefficients calculated in the half equations are also multiplied by 2.

$$2KMnO_4 \ + \ 10FeSO_4 \ + \ H_2SO_4 \longrightarrow 2MnSO_4 \ + \ 5Fe_2(SO_4)_3 \ + \ H_2O$$

6. In the obtained molecular equation, we must include among the products the remaining ions that do not change the oxidation state, i.e. K^+ and SO_4^{2-}, adding K_2SO_4 on the right side of the reaction

$$2KMnO_4 \ + \ 10FeSO_4 \ + \ H_2SO_4 \longrightarrow K_2SO_4 \ + \ 2MnSO_4 \ + \ 5Fe_2(SO_4)_3 \ + \ H_2O$$

7. We select the coefficients of the remaining reactants so that the number of atoms of all elements is the same on both sides

$$2KMnO_4 + 10FeSO_4 + 8H_2SO_4 \longrightarrow K_2SO_4 + 2MnSO_4 + 5Fe_2(SO_4)_3 + 8H_2O$$

Example 3

Will Fe^{3+} ions oxidize Br^- ions?

Solution

Standard redox potentials for both redox pairs are:

E_0 for Br_2 /Br^- + 1,09V
E_0 for Fe^{3+} /Fe^{2+} + 0,77V

Thus, Fe^{3+} ions will not oxidize Br^- ions; while Br_2 can be an oxidant for Fe^{2+} ions.

II. TASKS AND QUESTIONS TO BE COMPLETED BY THE STUDENT

1. Calculate the oxidation state of the elements in compounds:

   a) H_2SO_4, H_2SO_3, H_3PO_4, H_2S, HCl, HClO, $HClO_3$;
   b) Na_2SO_4, K_2SO_3, P_2O_5, Na_2S, KCl, $HClO_4$, $NaClO_3$;
   c) NH_3, N_2O, NO, SO_2, SO_3, P_2O_3, NH_4Cl;
   d) $NaAlO_2$, Al_2O_3, NaOH, H_2O, HIO_3, Na_2HPO_4;
   e) MnO_2, $KMnO_4$, K_2CrO_4, $K_2Cr_2O_7$, K_2MnO_4, Cr_2O_3;
   f) NaH, MgH_2, AlH_3, CH_4, NH_3, H_2S, H_2O_2.

2. Calculate the oxidation state of the elements in ions:

   a) SO_4^{2-}, MnO_4^-, PO_4^{3-}, ClO_4^-, CO_3^{2-}, NO_2^-;
   b) AsO_2^-, AsO_3^{3-}, AsO_4^{3-}, MnO_3^{2-}, $H_2PO_4^-$;
   c) ClO_3^-, HPO_4^{2-}, $Cr_2O_7^{2-}$, SO_3^{2-}, ClO^-, IO_3^-;
   d) CN^-, BO_3^{3-}, NO_2^-, SiO_3^{2-}, SCN^-, HSO_3^-.

3. Balance the following redox equations in the solutions; write down the appropriate half-reactions (equations). Indicate the oxidizing and reducing agent.

   a. $HI + H_2SO_4 \longrightarrow I_2 + H_2S + H_2O$
   b. $FeCl_3 + SnCl_2 \longrightarrow FeCl_2 + SnCl_4$
   c. $H_2SO_3 + Cl_2 + H_2O \longrightarrow H_2SO_4 + HCl$
   d. $Ca + H_2O \longrightarrow Ca(OH)_2 + H_2$
   e. $Al + HCl \longrightarrow AlCl_3 + H_2$
   f. $H_2S + O_2 \longrightarrow SO_2 + H_2O$
   g. $NH_3 + O_2 \longrightarrow NO + H_2O$
   h. $NH_3 + O_2 \longrightarrow H_2O + NO_2$
   i. $HNO_3 + NO \longrightarrow H_2O + NO_2$
   j. $H_2S + HNO_3 \longrightarrow H_2SO_4 + NO + H_2O$
   k. $H_2S + H_2SO_3 \longrightarrow S + H_2O$
   l. $As_2S_3 + HNO_3 + H_2O \longrightarrow H_3AsO_4 + H_2SO_4 + NO$
   m. $Ag_2S + HNO_3 \longrightarrow AgNO_3 + S + NO + H_2O$
   n. $Sb_2S_3 + O_2 \longrightarrow Sb_2O_3 + SO_2$
   o. $NaNO_2 + FeSO_4 + H_2SO_4 \longrightarrow Na_2SO_4 + Fe_2(SO_4)_3 + NO + H_2O$
   p. $NaNO_2 + KI + H_2SO_4 \longrightarrow I_2 + NO + Na_2SO_4 + K_2SO_4 + H_2O$
   q. $NaNO_3 + Zn + NaOH \longrightarrow NH_3 + Na_2ZnO_2 + H_2O$
   r. $KOH + Br_2 \longrightarrow KBrO_3 + KBr + H_2O$
   s. $HClO_3 \longrightarrow ClO_2 + HClO_4 + H_2O$
   t. $HNO_2 \longrightarrow HNO_3 + NO + H_2O$
   u. $Sn^{2+} + Hg^{2+} \longrightarrow Sn^{4+} + Hg_2^{2+}$
   v. $Br^- + SO4^{2-} + H^+ \longrightarrow Br_2 + SO_3^{2-} + H_2O$
   w. $NO_2^- + I^- + H_3O^+ \longrightarrow NO + I_2 + H_2O$
   x. $PbS + NO_3^- + H_3O^+ \longrightarrow Pb^{2+} + NO + S + H_2O$
   y. $AsO_3^{3-} + ClO^- \longrightarrow AsO_4^{3-} + Cl^-$
   z. $SO_3^{2-} + HPO_3^{2-} + H_3O^+ \longrightarrow S + HPO_4^{2-} + H_2O$
   aa. $AsO_4^{3-} + S^{2-} + H_3O^+ \longrightarrow AsO_3^{3-} + S + H_2O$
   bb. $AsO3^{3-} + BrO_3^- \longrightarrow AsO_4^{3-} + Br^-$

cc. $Bi + NO_3^- + H_3O^+ \longrightarrow Bi^{3+} + NO + H_2O$
dd. $Bi_2S_3 + NO_3^- + H_3O^+ \longrightarrow Bi^{3+} + NO + S + H_2O$

4. Balance each of the following redox reactions:
   a) $NH_3 + O_2 \longrightarrow NO + H_2O$
   b) $NH_3 + O_2 \longrightarrow H_2O + N_2$
   c) $HNO_3 + NO \longrightarrow H_2O + NO_2$
   d) $H_2S + HNO_3 \longrightarrow H_2SO_4 + NO + H_2O$
   e) $PbO_2 + HCl \longrightarrow PbCl_2 + Cl_2 + H_2O$
   f) $Pb + H_3PO_4 \longrightarrow Pb_3(PO_4)_2 + H_2$
   g) $HClO_4 + H_2SO_3 \longrightarrow HCl + H_2SO_4$
   h) $S + HNO_3 \longrightarrow H_2SO_4 + NO$

5. Balance each of the following redox reactions:

   1) $S^{2-} + I_2 \longrightarrow S + I^-$
   2) $NH_4^+ + NO_2^- \longrightarrow N_2 + H_2O$
   3) $Mg + H^+ + NO_3^- \longrightarrow Mg^{2+} + N_2O + H_2O$
   4) $Br^- + SO_4^{2-} + H^+ \longrightarrow Br_2 + SO_3^{2-} + H_2O$
   5) $NO_2^- + I^- + H^+ \longrightarrow NO + I_2 + H_2O$
   6) $SO_4^{2-} + Zn + H^+ \longrightarrow S^{2-} + Zn^{2+} + H_2O$
   7) $AsO_3^{3-} + ClO^- \longrightarrow AsO_4^{3-} + Cl^-$
   8) $SO_3^{2-} + HPO_3^{2-} + H^+ \longrightarrow S + HPO_4^{2-} + H_2O$