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ВВЕДЕНИЕ В ОБЩУЮ ХИМИЮ

INTRODUCTION TO THE GENERAL CHEMISTRY

Практикум

3-е издание



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PREFACE

The book provides an introduction into the General Chemistry. It is necessary for foreign students who are going to pass the Chemistry exam into the Medical Universities in English.

The book is a revised and extended version of several books on the same topic written by the authors of this book together with Professor Eugene Victorovich Barkovsky (18.05.1946–26.12.2015) back in 2014 and 2015.

Actually, this book is a kind of compromise between translation of chemistry textbooks from Russian to English and popularizing material from original American sources. Authors hope that they combined the best Belarusian traditions with the best international points of view on Chemistry teaching. Even though entrance exam into universities usually requires some knowledge on very specific set of rules mostly based on simplifications and overestimations, this book not just deals with those “rules of thumb” but also provides modern explanations and interpretations.

Authors are looking forward to receive any feedback from readers and colleagues regarding style and content of the book. Previous editions have been carefully read by talented students who helped the Authors to brush the book.

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LESSON 1

1.1 PHYSICAL AND CHEMICAL PROPERTIES OF SUBSTANCES: WHAT IS THE DIFFERENCE?

The question from the title of this subsection is very important. It can be rephrased in the following ways. What is chemistry? What is the subject of the discipline you are starting (or, hopefully, continuing) to study? What is the difference between chemistry and physics?

Physical properties of a substance are: state of matter (solid, liquid, gas and plasma), density (the ratio between the mass and the volume), color (pink, blue, green, etc.), taste (sour, sweet, bitter, salty), solubility (the mass of substance that can be dissolved in water or other liquids at a given temperature), boiling temperature (the temperature of vaporization), freezing temperature (the temperature of crystallization).

Chemical properties of a substance are described as its abilities to form other substances in different conditions.

In physical processes a substance changes at least one of its conditions: its volume, its shape, its position in the space, etc., while new substances are not formed. Phase transitions are also physical processes. There are several traditional examples of such physical processes: the melting of the ice and freezing (crystallization) of the water, boiling of the water and condensation of the vapor.

Chemical processes are described by chemical reactions. A chemical reaction is a process that leads to the transformation of one set of chemical substances to another. Substances from the first set are called “reactants”. Substances from the second set are called “products”. For example, sulfur and oxygen (the reactants) may react with each other and form sulfur dioxide (the product).

Chemical equation for this reaction is as follows: $S + O_2 = SO_2$.

The rearrangement of atoms happens in chemical reactions, while atoms themselves stay the same. Nuclear reactions are not chemical reactions, even though new substances are formed in them. In those nuclear reactions atoms of one chemical element turn to atoms of another chemical element. So, nuclear reactions are studied in the course of physics and not chemistry. Remember that it is hard to classify all the processes into purely chemical and purely physical ones. In the pre-university course you can use a simplified rule to claim a process to be chemical: “atoms are the same, substances are different”.

There are several signs of chemical reactions: the thermal change (in some cases the heat is produced in chemical reaction, in other cases the heat is adsorbed from surroundings during the chemical reaction); the smell (for example, hydrogen sulfide has a smell of rotten eggs); formation of a gas without any characteristic smell; the change in color; precipitation (formation of insoluble substance). All this signs are not absolutely specific for chemical processes. For example, the heat is released or absorbed during phase transitions. So, physics and chemistry are mutually connected with each other.

Finally, chemistry is the science on the interactions of matter with other matter and with energy.

Questions:

- a. List several physical properties of water, sugar and salt.
- b. A piece of chalk has been dissolved in hydrochloric acid. Was it a physical or a chemical process?
- c. What is the difference between reactants and products?
- d. What are the signs of chemical reactions? How can you know that a chemical reaction happened?
- e. What is the subject of chemistry?

1.2 ATOMS AND MOLECULES

Atom is the smallest piece of an element that maintains the identity of that element.

There are many substances that exist as two or more atoms connected together. These combinations are called molecules. A molecule is the smallest part of a substance that determines physical and chemical properties of that substance.

Some elements exist in form of molecules. For example, hydrogen and oxygen exist as two-atom (diatomic) molecules. Sulfur may exist as an eight-atom molecule, S_8 , while phosphorus may exist as a four-atom molecule, P_4 . Other elements, such as carbon (C), exist as individual atoms, rather than molecules. So, we can classify substances into two groups: molecular substances and nonmolecular substances (atomic, ionic or metallic).

A **chemical compound** is a chemical substance consisting of **two or more different chemical elements**. Chemical compounds can be molecular compounds held together by covalent bonds, salts held together by ionic bonds, intermetallic compounds held together by metallic bonds, or complexes held together by coordinate covalent bonds.

In general, when nonmetal connects with other nonmetal, the compound typically exists as a molecule. As an exception we can mention SiO_2 that is an atomic substance and not molecular one. When nonmetal connects with metal, the resulting compound usually has an ionic structure. When two metals connect with each other, the resulting compound is classified as metallic one.

Pure chemical elements are not considered chemical compounds, even if they consist of molecules that contain multiple atoms of a single element (such as H_2 , S_8 etc.), which are called diatomic molecules or polyatomic molecules. Pure nonmetals may have molecular or atomic structure. Pure metals consist of atoms, positively charged ions and free electrons (electron “gas”), and they have metallic structure.

Substances composed from atoms of the same element are historically called “simple substances”. So, the term “pure chemical element” is a synonym of the term “simple substance”.

Allotropy is the property of some chemical elements to exist in **two or more different forms**, known as allotropes of these elements.

Coming back to carbon, the allotropes of that element include diamond (where the carbon atoms are bonded together in a tetrahedral lattice arrangement) and graphite (where the carbon atoms are bonded together in sheets of a hexagonal lattice). The term allotropy is used for pure chemical elements only, and not for chemical compounds. Allotropy refers only to different substances which exist as pure chemical elements within the same phase (i. e. different solid, liquid or gas substances).

Usually, the most common substance of a given element has the same name as the chemical element itself. For example, the word “oxygen” may refer to both atoms of oxygen, and to oxygen as a diatomic molecule O_2 . In a sentence like “glucose contains oxygen” we mean that there are atoms of oxygen in the molecule of glucose ($C_6H_{12}O_6$). In a sentence like “we breathe in oxygen and breathe out carbon dioxide”, we mean that the aim of breathing is to enrich our blood by a substance called “oxygen” (by O_2 molecules).

Questions:

- a. What is atom?
- b. What is molecule?
- c. List some substances which consist of molecules.
- d. List some substances of nonmolecular structure.
- e. Give a definition of pure chemical element.
- f. Give a definition of chemical compound.
- g. Give a definition of simple substance.
- h. What is allotropy?
- i. Are oxygen (O_2) and ozone (O_3) allotropes?
- j. Compose a sentence in which iron atoms are discussed.
- k. Compose a sentence in which iron as a substance is discussed.

1.3. SYMBOLS OF CHEMICAL ELEMENTS

In the Table 1.1 we placed the names of several chemical elements that are frequently used in the course of pre-university chemistry. Students should know both English and Latin names of these elements to be able to write formulas of chemical compounds. Some elements have almost identical names in both languages, while others have different names.

Except chemical symbols, there are usually special numbers before and inside chemical formulas.

The coefficient always goes before the compound or molecule, not after. Coefficient shows the number of molecules, compounds or moles.

The subscript is written in small numbers by the bottom right corner of the symbol. Subscript shows the number of certain atoms or groups of atoms in a given compound or molecule.

English names of chemical elements and their symbols

Element Name	Symbol	Element Name	Symbol
Aluminum	Al	Mercury	Hg
Argon	Ar	Molybdenum	Mo
Arsenic	As	Neon	Ne
Barium	Ba	Nickel	Ni
Beryllium	Be	Nitrogen	N
Bismuth	Bi	Oxygen	O
Boron	B	Palladium	Pd
Bromine	Br	Phosphorus	P
Calcium	Ca	Platinum	Pt
Carbon	C	Potassium	K
Chlorine	Cl	Radium	Ra
Chromium	Cr	Radon	Rn
Cobalt	Co	Rubidium	Rb
Copper	Cu	Scandium	Sc
Fluorine	F	Selenium	Se
Gallium	Ga	Silicon	Si
Germanium	Ge	Silver	Ag
Gold	Au	Sodium	Na
Helium	He	Strontium	Sr
Hydrogen	H	Sulfur	S
Iodine	I	Tantalum	Ta
Iridium	Ir	Tin	Sn
Iron	Fe	Titanium	Ti
Krypton	Kr	Tungsten	W
Lead	Pb	Uranium	U
Lithium	Li	Xenon	Xe
Magnesium	Mg	Zinc	Zn
Manganese	Mn	Zirconium	Zr

For example, $4\text{H}_2\text{O}$ means four (coefficient) molecules of water. Water consists of two (subscript) hydrogen atoms and a single oxygen atom.

Chemical formula $2\text{Al}_2(\text{SO}_4)_3$ means two (coefficient) compounds of aluminum sulfate. Aluminum sulfate consists of two (subscript) aluminum ions and three (subscript behind the brackets) sulfate anions. Each sulfate anion consists of a single sulfur atom and four (subscript inside the brackets) oxygen atoms.

In the Table 1.2 we placed the most frequently used names of compounds (acids and salts). Students should learn those names and formulas to understand the meaning of the questions in exercises.

English names of several acids and their salts

Acid	Formula	Salt	Formula
Sulfuric acid	H ₂ SO ₄	Sodium sulfate	Na ₂ SO ₄
Sulfurous acid	H ₂ SO ₃	Sodium sulfite	Na ₂ SO ₃
Hydrosulfuric acid	H ₂ S	Sodium sulfide	Na ₂ S
Nitric acid	HNO ₃	Potassium nitrate	KNO ₃
Nitrous acid	HNO ₂	Potassium nitrite	KNO ₂
Carbonic acid	H ₂ CO ₃	Calcium carbonate	CaCO ₃
Silicic acid	H ₂ SiO ₃	Calcium silicate	CaSiO ₃
Phosphoric acid	H ₃ PO ₄	Lithium phosphate	Li ₃ PO ₄
Hydrochloric acid	HCl	Zinc chloride	ZnCl ₂
Hydrobromic acid	HBr	Zinc bromide	ZnBr ₂
Hydroiodic acid	HI	Zinc iodide	ZnI ₂

Notice that salts that are made from two elements have the same ending of “-ide”. This rule comes from the nomenclature of binary compounds. So, a compound made from two elements that is not a salt has the same name, as if it was a salt: ZnO is zinc oxide, HF is hydrogen fluoride.

Salts that are made from three elements (including oxygen) may have an ending of “-ate” or “-ite”. The last one is used when the number of oxygen atoms is lower in a given salt than that in a similar salt with an ending of “-ate”. The same situation exists for the endings of acids. The ending of “-ous” shows that the number of oxygen atoms in a given acid is lower than that in a similar acid with an ending of “-ic”. The prefix “hydro-” is used to highlight that an acid is made just from hydrogen and another nonmetal.

The nomenclature of metal hydroxides is easy to understand: one needs to name an element and add the word “hydroxide” to it.

Exercises:

- What does the coefficient show?
- What does the subscript show?
- Write chemical formulas of compounds made up from i) single iron atom and three chlorine atoms; ii) two aluminum atoms and three oxygen atoms; iii) single calcium atom, single carbon atom and three oxygen atoms.
- Read the names of the following salts: CuSO₄, CuSO₃, CuS, Mg(NO₃)₂, Mn(NO₂)₂.
- Read the names of the following hydroxides: Fe(OH)₃, Ca(OH)₂, Ba(OH)₂, KOH, NaOH.
- Read the names of the following acids: H₂SO₄, H₃PO₄, HNO₃, HNO₂, HCl.

TEST FOR CLASSWORK

1. Choose physical processes:
 - a. the melting of the ice
 - b. the boiling of water
 - c. the burning of wood
 - d. the oxidation of sulfur
2. Choose chemical processes:
 - a. production of ammonia from nitrogen and hydrogen
 - b. the dissolving of glucose in water
 - c. the dissolving of calcium carbide in water
 - d. the dissolving of sodium bicarbonate in acetic acid
3. Choose pure chemical elements:
 - a. chlorine gas
 - b. iron
 - c. steel
 - d. sodium chloride
4. Choose compounds:
 - a. sulfur dioxide
 - b. white phosphorus
 - c. methane
 - d. oxygen
5. Choose allotropic modifications of carbon:
 - a. graphite
 - b. diamond
 - c. propane
 - d. carbon dioxide
6. Choose oxygen containing compounds:
 - a. CaO
 - b. O₂
 - c. O₃
 - d. H₂SO₄
7. Choose chemical elements which have absolutely different names in English and Latin:
 - a. Ag
 - b. Ni
 - c. K
 - d. Fe
8. Choose chemical elements which have similar names in English and Latin:
 - a. B
 - b. W
 - c. N
 - d. Pb
9. Which compounds are composed of three atoms?
 - a. NO₂
 - b. HCN
 - c. HNO₃
 - d. N₂O₅
10. Which compounds are made from atoms of two chemical elements?
 - a. SO₃
 - b. N₂
 - c. P₂O₅
 - d. H₂SiO₃

TEST FOR HOMEWORK

1. Choose physical processes:
 - a. the mixing of flour with sugar
 - b. condensation of water
 - c. the burning of magnesium in CO₂
 - d. the rusting of iron
2. Choose chemical processes:
 - a. production of the distilled water
 - b. the dissolving of sodium chloride in water
 - c. the dissolving of sodium sulfide in water
 - d. the dissolving of aluminum chloride in water

3. Choose pure chemical elements:
 - a. lime water c. nitrogen
 - b. marble d. mercury
4. Choose compounds:
 - a. pyrite c. silver
 - b. gold d. bronze
5. Choose allotropic modifications of oxygen:
 - a. oxide c. oxygen
 - b. ozone d. ozonide
6. Choose phosphorus containing compounds:
 - a. F_2 b. NaF c. P_4 d. H_3PO_4
7. Choose chemical elements which have absolutely different names in English and Latin:
 - a. Au b. Pt c. Na d. F
8. Choose chemical elements which have similar names in English and Latin:
 - a. Br b. I c. Sn d. Sb
9. Which compounds are composed of four atoms?
 - a. H_2O_2 b. SO_3 c. H_3O^+ d. P_4
10. Which compounds are made from atoms of three chemical elements?
 - a. CCl_4 b. O_2 c. H_3PO_4 d. KCN

EXERCISES FOR CLASSWORK

1. Write the formula of sodium sulfide: _____
2. Write the formula of zinc sulfide: _____
3. Write the formula of aluminum sulfide: _____
4. Write the formula of lithium sulfite: _____
5. Write the formula of calcium sulfate: _____
6. Write the formula of sodium nitrite: _____
7. Write the formula of potassium nitrate: _____
8. Write the formula of aluminum nitrite: _____
9. Write the formula of strontium phosphate: _____
10. Write the formula of calcium carbonate: _____
11. Write the formula of barium silicate: _____
12. Write the formula of strontium oxide: _____
13. Write the formula of calcium hydroxide: _____
14. Write the formula of sulfuric acid: _____
15. Write the formula of nitric acid: _____

EXERCISES FOR HOMEWORK

1. Write the name of NaOH: _____
2. Write the name of H₂SO₄: _____
3. Write the name of FeO: _____
4. Write the name of KCl: _____
5. Write the name of HBr: _____
6. Write the name of NaNO₂: _____
7. Write the name of NH₄NO₃: _____
8. Write the name of CO₂: _____
9. Write the name of H₃PO₄: _____
10. Write the name of K₂SiO₃: _____

LESSON 2

2.1 ATOMIC MASS AND MOLECULAR MASS

The atomic mass unit (u) is defined as one-twelfth of the mass of a carbon-12 atom, an isotope of carbon that has six protons and six neutrons in its nucleus. By this scale, the mass of a proton is 1.00728 u, the mass of a neutron is 1.00866 u, and the mass of an electron is 0.000549 u (1835 times lower than the mass of a proton, 1837 times lower than the mass of a neutron, and 1821 times lower than the atomic mass unit). Approximate mass of an atom may be estimated simply — by counting the total number of protons and neutrons in the nucleus (that number is known as mass number).

The molecular mass is the sum of mass numbers of all the atoms in a molecule.

In different books you can find such expressions as “relative atomic mass” and “relative molecular mass”. An adjective “relative” means that a given value is the ratio between the concrete mass and another mass taken as an etalon (1/12 of the mass of carbon-12). So, relative atomic (and molecular) mass is dimensionless — it has no units.

In biochemistry instead of atomic mass units they use Daltons to measure molecular mass. One Dalton is equal to one atomic mass unit (1 Da = 1 u).

2.2 THE MOLE AND MOLAR MASS

Chemists usually deal with millions, billions, and trillions of atoms and molecules at a time. Mole is a unit of amount that relates quantities of substances on a scale that is easy to interact with.

Chemistry uses a unit of amount called “mole”. A mole is a *number of things* equal to the number of atoms in exactly 12 g of carbon-12. Experimental measurements have determined that this number is very large:

$$1 \text{ mol} = 6.02214179 \cdot 10^{23} \text{ things } (N_A = \text{Avogadro's number})$$

In chemical calculations we usually use two digits after the point: $N_A = 6.02 \cdot 10^{23}$.

Once again, a mole means a number of things ($6.02 \cdot 10^{23}$ — a very big number), just like a dozen means a certain number of things (twelve).

Molar mass is the mass of 1 mole of a given substance. Molar mass is measured in gram per mole.

The number of moles (n) is equal to the ratio between mass (m , measured in grams) and molar mass (M , measured in gram/mol).

$$n = m / M$$

Relative molar mass is a dimensionless value that shows the ratio between the molar mass of a given substance and the molar mass of carbon-12.

The numbers representing molar mass and molecular mass of the given substance are exactly the same, even though molar mass is measured in gram per mol, while molecular mass is measured in atomic mass units. It works well because of the specific value of Avogadro's number. If one divides a certain number of things (molecules — for molecular substances, units — for ionic substances, atoms — for atomic substances) by the number of things in 1 mole, the number of moles in a sample will be calculated.

$$n = N / N_A$$

Because of this formula, Avogadro's number is usually replaced by Avogadro's constant that is measured in mol^{-1} .

We can combine two equations from this section together and use the resulting equation in numerous exercises. Namely, with this equation you can calculate a mass (if molar mass and the number of molecules are given), a molar mass (if mass and the number of molecules are given), and a number of molecules (if mass and molar mass are given).

$$m / M = N / N_A$$

From this equation we can also express the mass of an atomic mass unit measured in gram.

$$m = N / (M \cdot N_A) = 1 / (1 \cdot N_A) = 1 / 6.02 \cdot 10^{23} = 1.66 \cdot 10^{-24} \text{ gram.}$$

So, an atomic mass unit has a mass of $1.66 \cdot 10^{-24}$ gram if we consider that its molar mass is equal to 1 g/mol.

In each compound atoms of different elements are presented at certain molar ratios. For example, in 1 mol of NO_2 there is 1 mol of nitrogen atoms and 2 mol of oxygen atoms. If you know the number of moles of oxygen atoms in a sample of that compound, you can find out the number of moles of its molecules: it is exactly 2 times lower (in case with NO_2).

Questions:

- a. Give the definition of the atomic mass unit.
- b. What is the molecular mass?
- c. How to determine a mass number of an atom?
- d. How to determine molecular mass of a molecule?
- e. What is Avogadro's number?
- f. What is the meaning of mole?
- g. Give a definition of molar mass.
- h. Why numbers representing molecular and molar masses are always the same?
- i. What is the meaning of a relative atomic mass?
- j. What is the meaning of relative molar mass?

EXERCISES FOR CLASSWORK

1. Calculate molecular masses of the following substances:

HNO_3 (nitric acid) _____

KOH (potassium hydroxide) _____

ZnSO_4 (zinc sulfate) _____

Fe_2O_3 (iron (III) oxide) _____

MgCO_3 (magnesium carbonate) _____

$\text{Mg}_3(\text{PO}_4)_2$ (magnesium phosphate) _____

$\text{Al}_2(\text{SO}_4)_3$ (aluminum sulfate) _____

2. How many molecules are there in:

0.5 mol of carbon dioxide (CO_2)

0.0002 mol of water (H_2O)

30 mol of chlorine gas (Cl_2)

5 mol of nitrogen (I) oxide (N_2O)

3. Calculate the number of moles for

$18.06 \cdot 10^{23}$ molecules of H_2O

$1.8 \cdot 10^{22}$ atoms of C

$3.3 \cdot 10^{23}$ units of Na_2S

$1 \cdot 10^{15}$ cations of Na^+

- 4. Calculate the number of moles in:
40 g of sodium hydroxide (NaOH)**

250 g of water (H_2O)

120 g of sulfuric acid (H_2SO_4)

5 g of calcium carbonate (CaCO_3)

- 5. What is the mass of
3.6 mol of sulfuric acid (H_2SO_4)?**

5 mol of aluminum oxide (Al_2O_3)

0.3 mol of hydrogen (H_2)

2.4 mol of magnesium cations (Mg^{2+})

- 6. Determine the mass (in gram) of a
single N_2 molecule**

single NH_3 molecule

single SO_4^{2-} anion

single PO_4^{3-} anion

7. How many molecules are there in
128 g of oxygen

50 g of chlorine gas

20 g of nitrogen (II) oxide

0.4 g of ozone

8. Find the molar mass of a substance if 5.5 mol of it has a mass of 435 g?

9. How many molecules are there in 3 L of liquid water (density is equal to 1 g/ml)?

10. Find the number of moles in a sample of a substance if there are $3.7 \cdot 10^{26}$ molecules in that sample.

11. Calculate the mass of $4.9 \cdot 10^{25}$ molecules of carbon dioxide (CO_2).

12. Find the molar mass of a substance if $3.01 \cdot 10^{23}$ molecules of it have a mass of 49 g.

13. How many moles of oxygen atoms are there in 100 g of barium sulfate (BaSO_4)?

14. Find the mass of a sample of ammonium nitrate if you know that there are 0.47 moles of nitrogen atoms in that sample.

EXERCISES FOR HOMEWORK

1. Calculate molar masses of the following substances:

MgO (magnesium oxide) _____

HCl (hydrochloric acid) _____

FeSO_4 (iron (II) sulfate) _____

$\text{Al}(\text{OH})_3$ (aluminum hydroxide) _____

NH_4NO_3 (ammonium nitrate) _____

$\text{Cu}(\text{NO}_3)_2$ (copper (II) nitrate) _____

BaSO_4 (barium sulfate) _____

2. How many units are there in

3 mol of iron (II) sulfate (FeSO_4)

0.1 mol of sodium hydrogen phosphate (Na_2HPO_4)

0.15 mol of sodium dihydrogen phosphate (NaH_2PO_4)

2.43 mol of potassium iodide (KI)

3. Calculate the number of moles for
 $3.01 \cdot 10^{22}$ molecules of HNO_3

$5.4 \cdot 10^{25}$ atoms of Si

$3.3 \cdot 10^{23}$ units of Na_2S

$0.5 \cdot 10^{30}$ anions of SO_4^{2-}

4. Calculate the number of moles in:
50 g of potassium hydroxide (KOH)

0.5 g of FeSO_4

18 g of hydrogen sulfide (H_2S)

42 g of ammonia (NH_3)

5. What is the mass of
0.67 mol of nitric acid (HNO_3)

0.004 mol of acetic acid (CH_3COOH)

0.3 mol of barium carbonate (BaCO_3)

25 mol of neon (Ne)

6. Determine the mass (in gram) of two O₂ molecules

two N₂ molecules

ten SO₃²⁻ anions

twenty HPO₄²⁻ anions

7. How many units are there in 523 g of zinc chloride

30 g of sodium hydroxide

44 g of potassium sulfate

0.06 g of calcium phosphate

8. Find the molar mass of a substance if 2.5 moles of it has a mass of 150 g?

9. How many molecules are there in 4 L of liquid ethanol (the density of C₂H₅OH is equal to 0.8 g/ml)?

10. Find the number of moles in a sample of a substance if there are $8.9 \cdot 10^{22}$ molecules in that sample.

11. Calculate the mass of $2.6 \cdot 10^{23}$ molecules of hydrogen peroxide (H_2O_2).

12. Find the molar mass of a substance if $1.2 \cdot 10^{24}$ molecules of it have a mass of 126 g.

13. How many moles of sodium ions are there in 200 g of sodium carbonate (Na_2CO_3)?

14. Find the mass of a sample of zinc hydroxide if you know that there are 0.27 moles of oxygen atoms in that sample.

LESSON 3

3.1 VALENCE

Valence (it also may be written as valency or valence number) is the number of chemical bonds a given atom has formed in a given molecule.

The number of bonds formed by a given element was originally thought to be a fixed chemical property. In fact, in most cases this is not true. For example, phosphorus often has a valence of three, but can also have other valences.

Nowadays the definition of valence has become quite different from the classic one. The current International Union of Pure and Applied Chemistry (IUPAC) version of that term, adopted in 1994: *“The maximum number of univalent atoms (originally hydrogen or chlorine atoms) that may combine with an atom of the element under consideration, or with a fragment, or for which an atom of this element can be substituted”*.

Valence is written in Roman numbers which has no sign (no plus or minus). For example, the valence of hydrogen is always equal to “I”. In the most of the substances the valence of oxygen is equal to “II”.

Each chemical bond is represented by a line in diagrams. Total number of lines near the given atom is equal to its valence. For example, the valence of sulfur in sulfur trioxide is equal to six. The valence of phosphorus in phosphoric acid is equal to five (Figure 3.1).

Notice that there may be single, double or even triple lines between atoms. They are used to show single, double and triple covalent bonds, respectively.

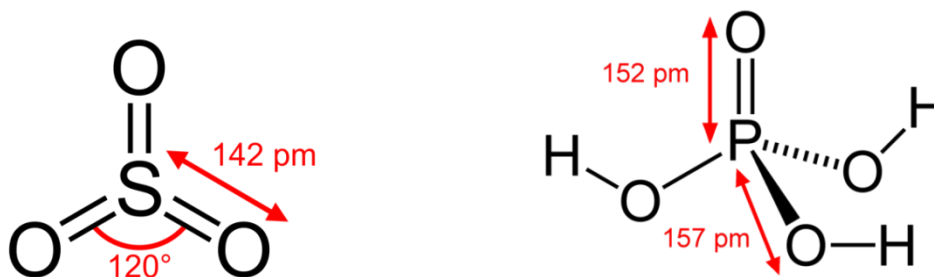


Figure 3.1. Structures of SO_3 and H_3PO_4

In easy cases one can predict the valence of a given element in a compound, if valences of other elements are known. However, that kind of prediction sometimes shows wrong valences. The only way to find out how exactly the atoms are connected together is to perform crystallographic experiment (X-ray or NMR analysis).

There are a few chemical elements that show constant valence in all their compounds. Here is a list of such elements: H (I), Li (I), Na (I), K (I), Rb (I), Cs (I), F (I), Be (II), Mg (II), Ca (II), Sr (II), Ba (II), Zn (II), Al (III). Also we should say that oxygen in the most of its compounds (except CO) shows the valence of II. So, if there is a binary compound featuring one of these elements, you can predict the valence of its partner.

The valence of sulfur in K_2S must be equal to the sum of valences for all potassium atoms (i.e. it is equal to II).

The valence of chlorine atom in ZnCl_2 must be equal to the valence of zinc divided by the number of chlorine atoms (i.e. it is equal to I).

The valence of phosphorus in Mg_3P_2 must be equal to the sum of valences for magnesium atoms divided by the number of phosphorus atoms (i.e. it is equal to III).

In general, the valence of a nonmetal in its binary compound with a metal is equal to the number of its subgroup in the periodic table.

In compounds made from three elements including oxygen the sum of valences of two other elements is equal to the sum of valences of all the oxygen atoms. For example, in H_2SeO_4 the valence of selenium atom is equal to VI ($2 \cdot 4 - 1 \cdot 2 = 6$).

3.2 CHEMICAL EQUATIONS AND THEIR BALANCING

The law of conservation of matter says that matter cannot be created or destroyed. In chemical equations, the number of atoms of each element in the reactants must be the same as the number of atoms of each element in the products. The mass of all the products should be the same as the mass of all reactants (if we are not taking into account the famous Einstein's equation $E = mc^2$ — the change in energy characteristic to every chemical reaction should lead to some little changes in the mass of substances).

Coefficients are used to balance a chemical equation.

Coefficient is a number in a chemical equation indicating the number of molecules (or moles) of each substance.

The sense of the balancing is to make the law of conservation of matter obey. In other words, you need to use coefficients to make the number of atoms of each element identical among reactants and products.

TEST FOR CLASSWORK

1. What is the valence of carbon in carbon dioxide?
a. II b. III c. IV d. V
2. In which compounds the valence of phosphorus is equal to V?
a. P_2O_5 b. P_2O_3 c. H_3PO_4 d. PCl_5
3. Calculate the sum of all coefficients in the following chemical reaction:
 $Zn + HCl \rightarrow ZnCl_2 + H_2$
a. 4 b. 5 c. 6 d. 7
4. Calculate the sum of coefficients before reactants in the following chemical reaction:
 $NaOH + H_2SO_4 \rightarrow Na_2SO_4 + H_2O$
a. 2 b. 3 c. 4 d. 5
5. Calculate the sum of coefficients before products in the following chemical reaction:
 $Al + HCl \rightarrow AlCl_3 + H_2$
a. 3 b. 4 c. 5 d. 6
6. In which molecules there are three single or a single triple covalent bond?
a. O_2 b. H_2O c. CO d. H_2O_2
7. Calculate the sum of all coefficients in the following chemical reaction:
 $Ca + H_2O \rightarrow Ca(OH)_2 + ?$
a. 3 b. 4 c. 5 d. 6
8. Calculate the sum of coefficients before reactants in the following chemical reaction:
 $AgNO_3 + ? \rightarrow AgCl + KNO_3$
a. 2 b. 3 c. 4 d. 5
9. Calculate the sum of coefficients before products in the following chemical reaction:
 $Ca(OH)_2 + H_3PO_4 \rightarrow Ca_3(PO_4)_2 + ?$
a. 1 b. 3 c. 5 d. 7
10. How many atoms are connected to the nitrogen atom in HNO_2 molecule?
a. 0 b. 1 c. 2 d. 3

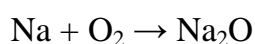
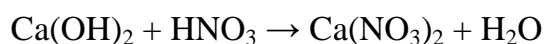
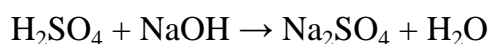
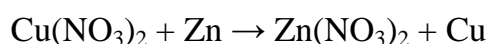
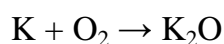
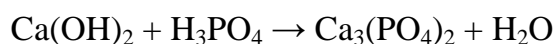
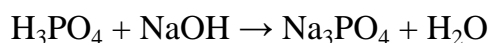
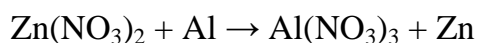
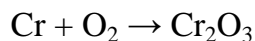
TEST FOR HOMEWORK

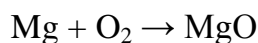
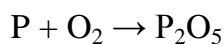
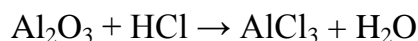
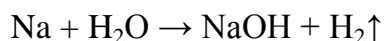
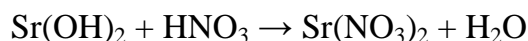
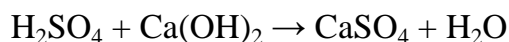
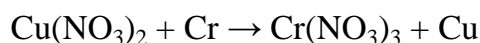
1. What is the valence of sulfur in sulfur trioxide?
a. II b. III c. IV d. VI

2. In which compounds the valence of silicon is equal to IV?
a. SiO₂ b. H₂SiO₃ c. K₂SiO₃ d. SiO
3. Calculate the sum of all coefficients in the following chemical reaction:
Fe + H₂SO₄ → FeSO₄ + H₂
a. 4 b. 5 c. 6 d. 7
4. Calculate the sum of coefficients before reactants in the following chemical reaction:
Al(OH)₃ + H₂SO₄ → Al₂(SO₄)₃ + H₂O
a. 2 b. 3 c. 4 d. 5
5. Calculate the sum of coefficients before products in the following chemical reaction:
Fe₂O₃ + HNO₃ → Fe(NO₃)₃ + H₂O
a. 3 b. 4 c. 5 d. 6
6. In which molecules there are four single or two double covalent bonds?
a. C₂H₂ b. NH₃ c. SO₂ d. CH₄
7. Calculate the sum of all coefficients in the following chemical reaction:
K + H₂O → ? + H₂
a. 4 b. 5 c. 6 d. 7
8. Calculate the sum of coefficients before reactants in the following chemical reaction:
CaCO₃ + ? → CaSO₄ + CO₂ + H₂O
a. 2 b. 3 c. 4 d. 5
9. Calculate the sum of coefficients before products in the following chemical reaction:
Ca(OH)₂ + HNO₃ → ? + H₂O
a. 1 b. 3 c. 5 d. 7
10. How many atoms are connected to the chlorine atom in HClO₄ molecule?
a. 2 b. 3 c. 4 d. 5

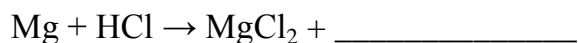
EXERCISES FOR CLASSWORK

1. Balance the following chemical equations:





2. Finish chemical equations and then balance them:



3. What are the products of the reaction between iron (II) oxide and hydrochloric acid?

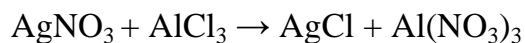
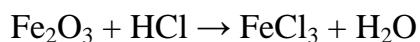
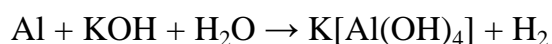
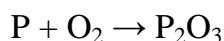
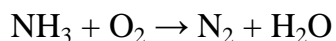
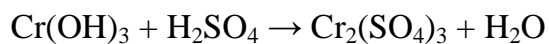
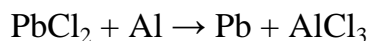
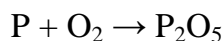
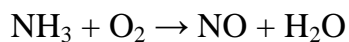
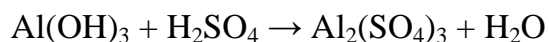
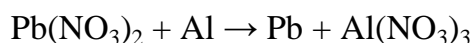
4. What are the products of the reaction between iron (III) oxide and nitric acid?

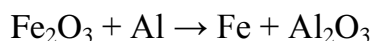
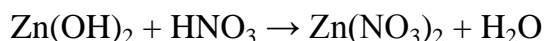
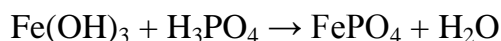
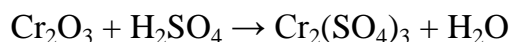
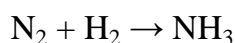
5. To produce zinc sulfate we need to put _____ in sulfuric acid.

6. To produce potassium hydroxide we should put _____ in water.

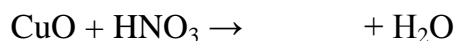
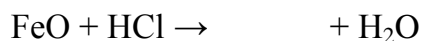
EXERCISES FOR HOMEWORK

1. Balance the following chemical equations





2. Finish chemical equations and then balance them:



3. The products of the reaction between aluminum and hydrochloric acid are:

4. Copper forms _____ in the reaction with sulfur.

5. To produce zinc phosphate we need to put zinc in _____ acid.

6. To produce iron (III) oxide we should burn _____.

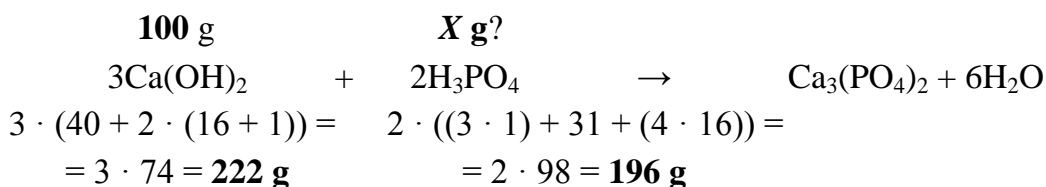
LESSON 4

4.1 CALCULATIONS USING CHEMICAL EQUATIONS

Here is the example of a simplest chemical calculation. Consider the following problem: “What is the mass of phosphoric acid (H_3PO_4) required for the complete neutralization of 100 g of calcium hydroxide ($\text{Ca}(\text{OH})_2$)?”

There are at least two mathematical ways to find the answer.

The first way allows calculations without direct referring to the quantity of matter (number of moles). The given mass of one of the reactants or products may be written just upon that chemical substance. The molar mass may be written under that substance. One should refer to the Periodic table to find atomic masses in case if he or she does not remember them by heart. The molar mass of the substance with unknown mass may also be written under the formula. Remember that molar masses of substances should be multiplied by coefficients from the balanced equation. To find out the unknown mass of the substance A (H_3PO_4 in our case) one has to multiply the known mass of another substance B ($\text{Ca}(\text{OH})_2$) by the molar mass of the unknown substance A and by the coefficient before that substance. Then the result of the multiplication has to be divided by the molar mass of the substance B (previously multiplied by the coefficient before the substance B).



$$X = m(\text{H}_3\text{PO}_4) = (100 \cdot 196) / 222 = 88.288 \text{ g} \approx 88.29 \text{ g}$$

The second way to find the same answer is based on the direct usage of numbers of moles. At first one has to divide the mass of substance A by its molar mass to find out the number of moles. Then one has to multiply the number of moles of substance A by the coefficient before the substance B and divide the result by the coefficient before the substance A. The result of this calculation is the number of moles of substance B. The last step is to multiply the number of moles of substance B by its molar mass.

$$1) n(\text{Ca(OH)}_2) = m(\text{Ca(OH)}_2) / M(\text{Ca(OH)}_2) = 100 / 74 = 1.35 \text{ mol}$$

$$2) n(\text{H}_3\text{PO}_4) = (1.35 \cdot 2) / 3 = 0.9 \text{ mol}$$

$$3) m(\text{H}_3\text{PO}_4) = n(\text{H}_3\text{PO}_4) \cdot M(\text{H}_3\text{PO}_4) = 0.9 \cdot 98 = 88.2 \text{ g}$$

In some cases it is easier to use the first method (when you just need to calculate the mass). In other cases when you should estimate the quantity of matter the second method should be chosen.

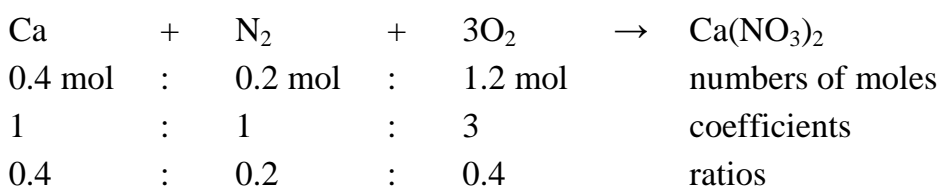
Both methods are based on a fact that substances react with each other at certain molar ratios. In our case the molar ratio between Ca(OH)_2 and H_3PO_4 is 3 over 2. If you take any amount of Ca(OH)_2 , the amount of H_3PO_4 will be equal to 2/3 of it. So, there is also a certain mass ratio between these reactants: for every 222 g of Ca(OH)_2 exactly 196 g of H_3PO_4 will be spent.

4.2 CASES WITH LIMITING REACTANT

There can be an excess of one of the reactants in the reaction mixture. In this case one should use the amount (or the mass) of the limiting reactant to calculate the mass of the product (or products).

It is easy to recognize the limiting reactant. Once again, there are at least two ways to do it. In the first way one should determine an amount of one product (either moles or mass) assuming all of each reactant reacts. Whichever reactant gives the *lesser* amount of product is the limiting reactant. The lowest number is the correct answer.

The second way is to calculate numbers of moles for each reactant and to compare those numbers with coefficients from the chemical equation. To make a comparison numbers of moles should be divided by coefficients. For the limiting reactant that ratio will be the lowest.



In this case the limiting reactant is N_2 . There is an excess of Ca ($0.4 - 0.2 = 0.2$ moles will be left after the end of the reaction) and O_2 ($1.2 - 0.2 \cdot 3 = 0.6$ moles will not take part in the reaction), while N_2 will react completely. The quantity of $Ca(NO_3)_2$ will be equal to 0.2 moles.

4.3 THE YIELD OF THE REACTION

Remember that reactants may not react completely. Because of different reasons (reaction may be stopped preliminarily, or the chemical equilibrium may be established) some part of a limiting reactant may not participate in a reaction. Alternatively, some part of a limiting reactant may participate in a parallel reaction and may not give the desired product. Anyway, the percentage of a reactant that did give the desired product is equal to the yield of a reaction. In other words, the ratio between the actual amount of a product and its theoretical amount is called the yield (relative yield). If you know both the amount of a reactant and the amount of a second product of the same reaction, use the amount of a second product to calculate the amount of the first product. That is how you will consider the possibility of the yield that is less than 100 %.

If the yield is equal to 100 %, and there is no limiting reactant (reactants are taken at stoichiometric ratios), you can use the rule that the final mass of all the products is equal to the initial mass of all the reactants.

EXERCISES FOR CLASSWORK

1. What is the mass of sulfur reacted with oxygen and produced 3.6 g of sulfur dioxide:

balance the equation: $S + O_2 \rightarrow SO_2$

2. What is the mass of oxygen reacted with copper and produced 8.5 g of copper (II) oxide:

balance the equation: $Cu + O_2 \rightarrow CuO$

3. Calculate the mass of phosphorus (V) oxide produced from 6.2 g of phosphorus and the excess of oxygen?

balance the equation: $P + O_2 \rightarrow P_2O_5$

- 4. What is the mass of zinc chloride produced in the reaction between 9 g of zinc and hydrochloric acid?**

balance the equation: $\text{Zn} + \text{HCl} \rightarrow \text{ZnCl}_2 + \text{H}_2$

- 5. What is the mass of silver chloride produced in the reaction between 10 g of silver nitrate and barium chloride?**

balance the equation: $\text{AgNO}_3 + \text{BaCl}_2 \rightarrow \text{AgCl} + \text{Ba}(\text{NO}_3)_2$

- 6. Calculate the mass of water needed to produce 20 g of lithium hydroxide from lithium.**

balance the equation: $\text{Li} + \text{H}_2\text{O} \rightarrow \text{LiOH} + \text{H}_2$

- 7. How many moles of oxygen are needed to produce 2 g of calcium nitrate according to the following equation.**

balance the equation: $\text{Ca} + \text{N}_2 + \text{O}_2 \rightarrow \text{Ca}(\text{NO}_3)_2$

- 8. What mass of sodium bicarbonate should be thermally decomposed to produce 0.35 mol of carbon dioxide?**

balance the equation: $\text{NaHCO}_3 \rightarrow \text{Na}_2\text{CO}_3 + \text{CO}_2 + \text{H}_2\text{O}$

- 9. Find the mass of zinc oxide formed in the reaction of zinc hydroxide decomposition. The mass of zinc hydroxide was equal to 75 g, the mass of water produced is equal to 10 g.**

balance the equation: $\text{Zn}(\text{OH})_2 \rightarrow \text{ZnO} + \text{H}_2\text{O}$

- 10. What is the mass of iron (III) chloride formed in the reaction between iron and chlorine gas? The mass of iron is equal to 18.2 g, the mass of chlorine gas is equal to 16.4 g.**

balance the equation: $\text{Fe} + \text{Cl}_2 \rightarrow \text{FeCl}_3$

EXERCISES FOR HOMEWORK

- 1. What is the mass of carbon reacted with oxygen and produced 5.2 g of carbon dioxide:**

balance the equation: $\text{C} + \text{O}_2 \rightarrow \text{CO}_2$

- 2. What is the mass of aluminum that can completely reduce 100 g of iron (III) oxide to the pure iron?**

balance the equation: $\text{Al} + \text{Fe}_2\text{O}_3 \rightarrow \text{Fe} + \text{Al}_2\text{O}_3$

- 3. Calculate the mass of orthophosphoric acid produced from 9.1 g of phosphorus (V) oxide and the excess of water?**

balance the equation: $\text{P}_2\text{O}_5 + \text{H}_2\text{O} \rightarrow \text{H}_3\text{PO}_4$

- 4. What is the mass of sodium iodide produced in the reaction between 9 g of sodium and the excess of iodine?**

balance the equation: $\text{Na} + \text{I}_2 \rightarrow \text{NaI}$

- 5. What is the mass of barium sulfate produced in the reaction between 6 g of barium hydroxide and potassium sulfate?**

balance the equation: $\text{Ba}(\text{OH})_2 + \text{K}_2\text{SO}_4 \rightarrow \text{BaSO}_4 + \text{KOH}$

- 6. Calculate the mass of strontium needed to produce 30 g of strontium hydroxide in the reaction with water.**

balance the equation: $\text{Sr} + \text{H}_2\text{O} \rightarrow \text{Sr}(\text{OH})_2 + \text{H}_2$

- 7. Find the mass of sodium sulfate formed in the reaction between a water solution containing 0.2 mol of sulfuric acid and the excess of sodium hydroxide.**

balance the equation: $\text{NaOH} + \text{H}_2\text{SO}_4 \rightarrow \text{Na}_2\text{SO}_4 + \text{H}_2\text{O}$

- 8. Find the number of moles of acetic acid that reacted with the excess of potassium hydroxide and produced 2.4 g of potassium acetate.**

balance the equation: $\text{KOH} + \text{CH}_3\text{COOH} \rightarrow \text{CH}_3\text{COOK} + \text{H}_2\text{O}$

- 9. Find the mass of aluminum oxide formed in the reaction of aluminum hydroxide decomposition. The mass of aluminum hydroxide was equal to 110 g, the mass of water produced is equal to 15 g.**

balance the equation: $\text{Al}(\text{OH})_3 \rightarrow \text{Al}_2\text{O}_3 + \text{H}_2\text{O}$

- 10. What is the mass of calcium carbonate formed in the reaction between calcium oxide and carbon dioxide? The mass of calcium oxide is equal to 33.3 g, the mass of carbon dioxide is equal to 12.2 g.**

balance the equation: $\text{CaO} + \text{CO}_2 \rightarrow \text{CaCO}_3$

LESSON 5

5.1 MOLAR VOLUME OF GASES

Kinetic theory of gases is the fundamental model that describes the physical properties of gases. It is based on the following statements (assumptions).

1. Gases consist of tiny particles of matter that are in constant motion.
2. Gas particles are constantly colliding with each other and the walls of a container. These collisions are elastic; that is, there is no net loss of energy from the collisions.
3. Gas particles are separated by large distances, with the size of a gas particle much less than the distances that separate them.
4. ***There are no interactive forces*** (i. e., attraction or repulsion) ***between the particles of a gas.***
5. The average speed of gas particles is dependent on the temperature of the gas.

An ideal gas exactly follows the statements of the kinetic theory. However, real gases are not ideal. Many gases deviate slightly from agreeing perfectly with the kinetic theory of gases. However, most gases adhere to the statements so well that the kinetic theory of gases is well accepted by the scientific community.

For ideal gases, the molar volume is given by the ideal gas equation ($p \cdot V = n \cdot R \cdot T$): this is a good approximation for many common gases at standard temperature and pressure.

The molar volume of an ideal gas at 1 atmosphere of pressure and at 0 °C (in so called normal conditions) is equal to 22.4 L/mol. According to the ideal gas equation, molar volume (V/n) is equal to $(R \cdot T)/p$. The higher the temperature, the higher the molar volume. The higher the pressure, the lower the molar volume. One may say that the molar volume of any gas in normal conditions is very close to 22.4 L/mol.

Standard conditions (standard temperature and pressure = STP) are often confused with normal conditions because of the lack of agreement between different scientific societies and organizations. Moreover, different organizations consider different temperatures and pressures as standard ones. The authors of this book prefer to use 25 °C and 1 atmosphere of pressure as STP. However, in problems from pre-university chemistry course the volumes of all the gases are thought to be measured in normal conditions, and all the gases are thought to behave as ideal gases. It means that 1 mole of each gas occupies exactly 22.4 L.

Molar volume (V_m) is equal to the volume divided by the number of moles, just like molar mass is equal to the mass divided by the number of moles.

$$V_m = V / n$$

$$M = m / n$$

Molar mass varies greatly for different substances (including gases), while molar volume for ideal gases is always the same (22.4 L/mol in normal conditions).

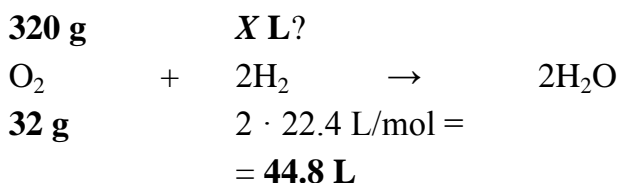
To calculate a volume of the gas with known mass one may use the following formula:

$$V = (V_m \cdot m) / M$$

The above written formula may also be combined with the formula for the calculation of the number of particles in a sample. Finally, you have to memorize the combined equation.

$$n = m / M = V / V_m = N / N_A$$

To make a “volume from mass” calculation using chemical equation one should write molar volume under the formula of a gas and multiply it by the coefficient. Here is the example.



$$X = V(\text{H}_2) = (320 \cdot 44.8) / 32 = 448\ \text{L}$$

The same problem can be solved in several steps. The number of moles of oxygen (O_2) is equal to 10 mol ($m / M = 320 / 32 = 10$). 1 mol of O_2 reacts with 2 mol of H_2 . So, 10 mol of O_2 react with 20 mol of H_2 ($10 \cdot 2 = 20$). The volume of 20 mol of H_2 in normal conditions is equal to $20 \cdot 22.4 = 448\ \text{L}$.

To make a “volume from volume” calculation you may not even use molar volume, since it will be crossed out anyway. Such calculation works for ideal gases in any conditions. For example, to calculate the volume of H_2 reacted with 2 L of O_2 you need just to multiply the volume of oxygen by the coefficient before the volume of hydrogen ($2 \cdot 2 = 4\ \text{L}$).

5.2 RELATIVE DENSITIES OF GASES

Density is the ratio between mass and volume ($\rho = m / V$). Actually, it is equal to the ratio between molar mass and molar volume as well ($\rho = M / V_m$).

Relative density is the ratio of the density of a given substance to the density of a certain reference material. In case of gases, relative density of one gas per another gas equals to the ratio between the molar masses of those gases, since their molar volumes are the same.

$$\rho_1 / \rho_2 = (M_1 / V_1) / (M_2 / V_2) = (M_1 \cdot V_m) / (M_2 \cdot V_m) = M_1 / M_2$$

For example, the density of oxygen per hydrogen is equal to the ratio between molar mass of oxygen and molar mass of hydrogen.

$$D_{\text{H}_2}(\text{O}_2) = M(\text{O}_2) / M(\text{H}_2) = 32 / 2 = 16$$

Average molar mass of the dry air is equal to 29 g/mol. So, density of nitrogen per dry air equals to molar mass of nitrogen divided by 29 g/mol.

$$D_{\text{dry air}}(\text{N}_2) = M(\text{N}_2) / M(\text{dry air}) = 28 / 29 = 0.97$$

Average molar mass for a mixture of gases is calculated as the sum of multiples of molar masses and molar fractions for all the components. For example, if our mixture is

made from 78 mol of nitrogen and 22 mol of oxygen, molar fractions of N_2 and O_2 are equal to 0.78 and 0.22, respectively. The average molar mass is calculated as follows.

$$M_{av} = M(N_2) \cdot 0.78 + M(O_2) \cdot 0.22 = 28 \cdot 0.78 + 32 \cdot 0.22 = 28.88$$

The average molar mass cannot be higher than the molar mass of a heaviest gas in a mixture. It cannot be also lower than the molar mass of a lightest gas in a mixture.

Using the relative density of the first gas per the second gas you can calculate the molar mass of the first gas: the relative density should be multiplied by the molar mass of the second gas.

EXERCISES FOR CLASSWORK

1. How many moles are there in:

11.7 L of nitrogen (in normal conditions)

14.3 L of oxygen (in normal conditions)

5.2 L of carbon dioxide (in normal conditions)

0.8 L of ammonia (in normal conditions)?

2. What volume is occupied by:

4.5 mol of chlorine gas (in normal conditions)?

0.05 mol of argon (in normal conditions)?

1.33 mol of methane (in normal conditions)?

15.75 mol of nitrogen (IV) oxide (in normal conditions)?

3. How many molecules are there in:

2 L of hydrogen (in normal conditions)?

6 L of ethylene (in normal conditions)?

3.45 L of hydrogen chloride (in normal conditions)?

2 L of ozone (in normal conditions)?

4. Find the molar mass of a gas if 10 g of it occupy a volume of 5.1 L (in normal conditions)?

5. Calculate the mass of ammonia which has a volume of 133.6 L (in normal conditions)?

6. Calculate the volume of oxygen required to burn down 14.2 L of methane.

7. What is the volume of carbon monoxide produced from 37 L of carbon dioxide in its reaction with coal (in normal conditions)?

8. Calculate the density per dry air for:

oxygen _____

ammonia _____

chlorine gas _____

phosgene (COCl_2) _____

9. Calculate the volume of an unknown gas which has a mass equal to 7 g and the relative density per oxygen which is equal to 0.625?

10. What is the density of unknown gas per nitrogen if its density per hydrogen is equal to 17?

11. Calculate the volume of hydrogen sulfide produced in the reaction between 0.5 mol of sodium sulfide and a water solution that contained 40 g of hydrogen chloride?

12. Find the volume of carbon dioxide (in normal conditions) produced in the reaction between 10 g of calcium carbonate and 8 g of hydrochloric acid.

EXERCISES FOR HOMEWORK

1. How many moles are there in:

1.8 L of ozone (in normal conditions)?

7.9 L of sulfur dioxide (in normal conditions)?

1.2 L of nitrogen dioxide (in normal conditions)?

21.2 L of nitrogen monoxide (in normal conditions)?

2. What volume is occupied by:

3.2 mol of hydrogen (in normal conditions)?

4.8 mol of acetylene (in normal conditions)?

1.6 mol of neon (in normal conditions)?

13.3 mol of ethane (in normal conditions)?

3. How many molecules are there in:

7.7 L of ammonia (in normal conditions)?

2.2 L of phosphine (PH₃) (in normal conditions)?

0.04 L of dinitrogen monoxide (in normal conditions)?

16.6 L of fluorine (in normal conditions)?

4. Find the molar mass of a gas if 5 g of it occupy a volume of 2.2 L (in normal conditions)?

5. Calculate the mass of sulfur (IV) oxide which has a volume of 43.7 L (in normal conditions)?

6. Calculate the volume of oxygen required to burn down 5.2 L of acetylene.

7. What is the volume of sulfur (IV) oxide produced from 4.5 g of hydrogen sulfide in its reaction with the excess of oxygen (in normal conditions)?

8. Calculate the molar mass of a gas if its density per oxygen is 1.875:

its density per ammonia is 1.765:

its density per chlorine gas is 0.704:

its density per hydrogen is 2:

9. Calculate the volume of an unknown gas which has a mass equal to 5 g and the relative density per dry air which is equal to 1.25?

10. What is the density of unknown gas per dry air if its density per ozone is equal to 1.5?

11. Find the volume of carbon dioxide produced in the reaction between a solution containing 0.2 mol of sulfuric acid and 2 g of sodium carbonate.

12. Find the volume of ammonia (in normal conditions) produced in the reaction between 7 g of ammonium nitrate and 9 g of potassium hydroxide.

LESSON 6

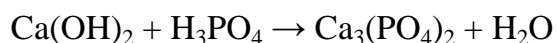
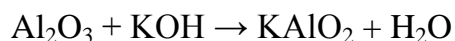
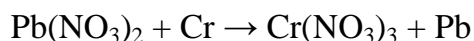
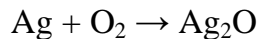
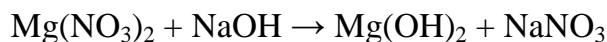
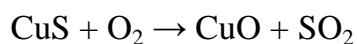
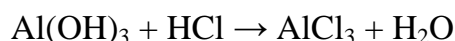
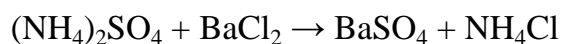
SAMPLE TICKET FOR CONTROL TASK #1

ON MAIN CONCEPTS AND LAWS OF CHEMISTRY

1. Calculate the mass of ammonia which has a volume of 33.6 L (in normal conditions)?

2. How many molecules are there in 1 L of liquid water and in water vapor (in normal conditions)?

3. Balance the following chemical equations:



4. How many moles of oxygen react with 36 g of carbon?

5. What is the mass of sulfur reacted with oxygen and produced 5.6 L of SO_2 (in normal conditions)?

6. Calculate the molar mass of unknown gas if the mass of 3 L of that gas is equal to 9.51 g

7. What is the volume of hydrogen sulfide formed in the reaction between hydrogen and sulfur? The mass of hydrogen is equal to 8.5 g, the mass of sulfur is equal to 3.4 g.

8. What is the mass of AgCl produced in the reaction between 1.34 g of AgNO₃ and 1.34 g of ZnCl₂?

9. Calculate the volume of oxygen required to burn down 4.2 L of methane (CH₄).

10. Calculate the volume of hydrogen gas required for the reaction with 3.3 L of oxygen at 35 °C and at the pressure of 1.2 atmospheres. Water is the product of that reaction.

LESSON 7

7.1 THE PERIODIC TABLE OF ELEMENTS

A periodic table is a tabular display of the chemical elements, organized on the basis of their atomic numbers, electron configurations, and recurring chemical properties. Elements are presented in order of increasing atomic number (number of protons in nucleus). Number of protons in nucleus determines electronic configuration. Electronic configuration determines chemical properties.

Historically, D. I. Mendeleev arranged elements in the order of the increase of their atomic masses. In the most of the cases atomic mass correlates well with the number of protons in nucleus, while several exceptions are known. For example, atomic number of K (19) is higher than atomic number of Ar (18), while the mass number of K (39) is lower than the mass number of Ar (40). One also may be interested in similar situation with Te and I.

A *period* is a horizontal row in the periodic table. Metallic properties of elements are decreasing from left to right of each period, while nonmetallic properties are increasing.

In this terminology metallic properties mean the ability to lose electron(s). Nonmetallic properties mean the ability to gain electron(s).

The first, second and third periods are usually referred to as “short periods”, while other periods (from fourth to the seventh one) are called “long periods” because they include d-elements.

A **group** is a vertical column in the periodic table. Elements within the same group generally have the same electron configuration in their valence shell. Consequently, elements in the same group tend to have a shared chemistry and exhibit a clear trend in properties with increasing atomic number.

There are four **blocks** in the periodic table. Two first groups (and He from the group 18) are known as s-elements and they build together an s-block. Elements from groups 3–12 are known as d-elements and they form a d-block. Elements from groups 13–18 are known as p-elements and they make a p-block. At the very bottom of the table one can see two lines with f-elements that make together an f-block of the table.

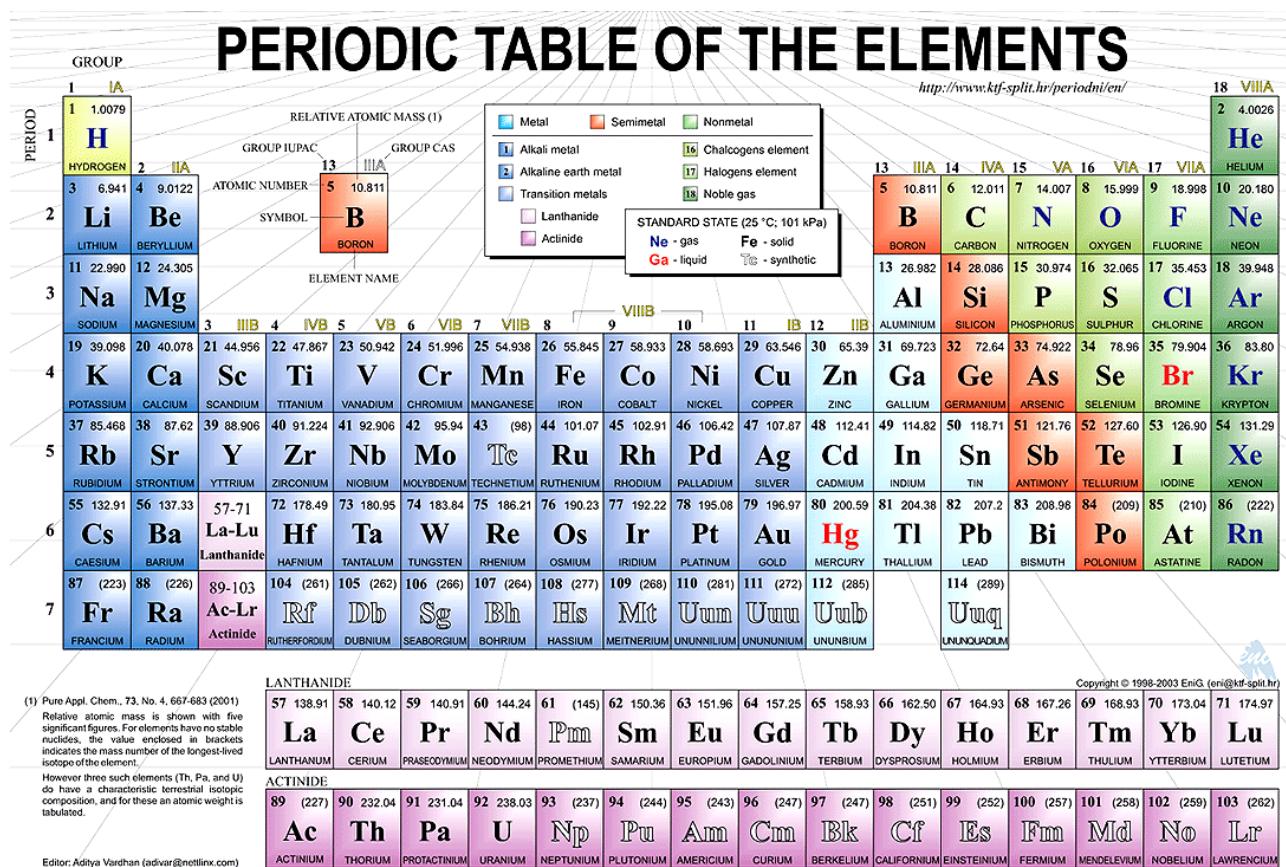


Figure 7.1. The Periodic table of chemical elements

Under an international naming convention, the groups are numbered numerically from 1 to 18 from the leftmost column (the alkali metals) to the rightmost column (the noble gases). Previously, they were known by roman numerals. The roman numerals were followed by either an “A” if the group was in the s- or p-block, or a “B” if the group was in the d-block. The roman numerals used correspond to the last digit of today's naming

convention (e. g. the group 4 elements were subgroup IVB, and the group 14 elements were subgroup IVA). In addition, groups 8, 9 and 10 used to be treated as one triple-sized group, known collectively as subgroup VIIIB. In 1988, the new IUPAC naming system was put into use, and the old group names were deprecated. However, in the most of the periodic tables both classifications are present.

For the elements from s- and p-blocks the valence in the highest oxide is equal to the number of a subgroup (that is written in Roman numbers). This rule has some deviations for elements from the second period. The valence in the binary compound with hydrogen is equal to the number of a subgroup for elements from subgroups IA–IVA. For elements from subgroups VA–VIIA that valence is equal to 8 minus the number of a subgroup.

Sometimes such term as a family is used to refer to several elements that are not making up a group or a subgroup, but have some similar chemical properties. For example, iron, cobalt and nickel are known as an iron family.

From the top to the bottom in a group, the atomic radii of the elements increase. Since there are more filled energy levels, valence electrons are found at longer distances from nuclei. It is easier to remove an electron since the atoms are less tightly bound. That is why metallic properties are increasing from the top to the bottom of each group, while nonmetallic properties are decreasing. Notice that these trends work well only in A-subgroups (for s- and p-elements), and not in B-subgroups (for d-elements).

The most metallic (those with the strongest metallic properties) elements (such as cesium and francium) are found in the bottom left corner of the periodic table and the most nonmetallic elements (oxygen, fluorine, chlorine) are found in the top right corner. The combination of horizontal and vertical trends in metallic character explains the stair-shaped dividing line between metals and nonmetals found in some types of periodic tables (see the Appendix 1), and the practice of sometimes categorizing several elements adjacent to that line as metalloids or semimetals. In other words, such elements as B, Si, Ge, As, Sb, Te and Po are classified as semimetals and metalloids. However, B, Si, As and Te are rather nonmetals than metals, while Ge, Sb and Po are rather metals than nonmetals.

7.2 HOW TO USE THE PERIODIC TABLE?

To determine the number of protons one should find the **ATOMIC NUMBER** of the element. The atomic number of the element is also equal to the number of electrons.

ATOMIC MASS given in the periodic table is the weighted average mass for all the isotopes with the respect of their abundance in Nature. Isotopes are atoms with equal numbers of protons in their nuclei but with different numbers of neutrons. Since chemical properties are determined by the number of electrons in the valence shells, isotopes demonstrate identical chemical properties and slightly different atomic masses.

For example, boron exists as about 20 % boron-10 (five protons and five neutrons in the nuclei) and about 80 % boron-11 (five protons and six neutrons in the nuclei). The weighted average atomic mass of boron is calculated as follows:

$$0.20 \times 10 \text{ u} = 2.0 \text{ u}$$

$$0.80 \times 11 \text{ u} = 8.8 \text{ u}$$

Sum = 10.8 u = the average weighted atomic mass of Boron

Thus, 10.8 u is the atomic mass of boron in the Periodic table.

The mass of electrons does not make a significant contribution into the atomic mass. So, atomic masses of isotopes are usually very close to integer numbers. However, in heavy elements the atomic mass is slightly lower than the sum of the numbers of protons and neutrons because of so-called “weight to energy” transitions according to the Einstein’s equation ($E = mc^2$).

Questions:

- a. Give the definition of the term “period”.
- b. Give the definition of the term “group”.
- c. How do metallic properties of elements change when you move from the left to the right of the same period?
- d. How do nonmetallic properties of elements change when you move from the top to the bottom of the same group?
- e. Is calcium a more active metal than strontium?
- f. Is phosphorus a more active nonmetal than chlorine?
- g. What elements are known under the name “metalloids”?
- h. Give the definition of the term “isotopes”.
- i. What element has 12 protons in its nucleus?
- j. What element has 25 electrons?

TEST FOR CLASSWORK

1. Metallic properties of chemical elements from A subgroups increase from:
 - a. left to right
 - b. right to left
 - c. top to bottom
 - d. bottom to top
2. Nonmetallic properties of chemical elements from A subgroups increase from:
 - a. left to right
 - b. right to left
 - c. top to bottom
 - d. bottom to top
3. Atomic radii of chemical elements from A subgroups increase from:
 - a. left to right
 - b. right to left
 - c. top to bottom
 - d. bottom to top
4. Electronegativity of chemical elements from A subgroups increases from:
 - a. left to right
 - b. right to left
 - c. top to bottom
 - d. bottom to top

5. Ionization energy of chemical elements from A subgroups increases from:
a. left to right c. top to bottom
b. right to left d. bottom to top
6. Choose s-elements:
a. Na b. Sr c. Al d. Sc
7. Choose p-elements:
a. Cu b. Sb c. Cl d. He
8. Choose d-elements:
a. Ba b. Fe c. Si d. Mn
9. Which properties are usually identical for elements from the same subgroup?
a. number of protons c. number of electrons on the outer layer
b. highest valence d. electronegativity
10. Which properties are identical for elements from the same period?
a. atomic radius c. the number of electron layers
b. ionization energy d. chemical properties

TEST FOR HOMEWORK

1. Metallic properties of chemical elements from A subgroups decrease from:
a. left to right c. top to bottom
b. right to left d. bottom to top
2. Nonmetallic properties of chemical elements from A subgroups decrease from:
a. left to right c. top to bottom
b. right to left d. bottom to top
3. Atomic radii of chemical elements from A subgroups decrease from:
a. left to right c. top to bottom
b. right to left d. bottom to top
4. Electronegativity of chemical elements from A subgroups decreases from:
a. left to right c. top to bottom
b. right to left d. bottom to top
5. Ionization energy of chemical elements from A subgroups decreases from:
a. left to right c. top to bottom
b. right to left d. bottom to top
6. Choose s-elements:
a. He b. Cl c. H d. Ar
7. Choose p-elements:
a. Al b. S c. Hg d. C

8. Choose d-elements:
 - a. Be
 - b. Co
 - c. Ni
 - d. Mg
9. Which properties are usually identical for elements from the same subgroup?
 - a. the formula of the highest oxide
 - b. atomic radius
 - c. the formula of the pure chemical element
 - d. the formula of the binary compound with hydrogen
10. Which properties are identical for elements from the same period?
 - a. electronegativity
 - b. molecular mass
 - c. the line in the Periodic table
 - d. the column in the Periodic table

EXERCISES FOR CLASSWORK

1. Find out numbers of period and group for:

K _____, Na _____, Ca _____, Al _____,
 Fe _____, Cr _____, Zn _____, C _____.

2. Calculate numbers of neutrons for the following atoms:

^{15}N _____; ^{119}Sn _____;
 ^{235}U _____; ^{137}Cs _____.

3. Write the names of metals from the group IIIA of the periodic table:

4. Write the names of metals from the group IVA of the periodic table:

5. Write the names of nonmetals from the group VA of the periodic table:

6. Write the names of nonmetals from the group VIA of the periodic table:

7. Write the names of metals from the 3rd period of the periodic table:

8. Arrange these elements in the order of the increase of their metallic properties (Al / Na / Mg / Si / Cs / C):

9. Arrange these elements in the order of the decrease of their nonmetallic properties (B / Br / Cl / F / Al / I):

10. Write the formulas of the highest oxides of elements from the VIA subgroup of the periodic table starting from the 3rd period:

11. Write the formulas of the binary compounds with hydrogen for elements from the IIIA group of the periodic table:

12. Write metalloids that are classified as metals if we divide all the elements into metals and nonmetals?

13. Calculate the number of protons in 20 g of phosphoric acid.

14. Calculate the atomic mass of an element that have three isotopes with the following atomic masses and corresponding abundances: 271 u (15 %), 273 u (48 %), 277 u (37 %).

15. The percent of ^{35}Cl in the sample of HCl is 80 %, the percent of ^{37}Cl is 20 %. Find the volume of hydrogen produced in the reaction between 100 g of that HCl sample with sodium.

EXERCISES FOR HOMEWORK

1. Find out numbers of period and group for:

P _____, N _____, S _____, O _____,

Cl _____, Si _____, Cr _____, Mn _____.

2. Calculate numbers of neutrons for the following atoms:

^{31}P _____; ^{23}Na _____;

^{122}Sb _____; ^{200}Hg _____.

3. Write the names of nonmetals from the group IIIA of the periodic table:

4. Write the names of nonmetals from the group IVA of the periodic table:

5. Write the names of metals from the group VA of the periodic table:

6. Write the names of metals from the group VIA of the periodic table:

7. Write the names of metals from the 2nd period of the periodic table:

8. Arrange these elements in the order of the increase of their metallic properties (Ca / Sr / Al / Mg / Ba / B):

9. Arrange these elements in the order of the decrease of their nonmetallic properties (F / O / S / C / Si / Be):

10. Write the formulas of the highest oxides of elements from the VA group of the periodic table starting from the 3rd period:

11. Write the formulas of the binary compounds with hydrogen for elements from the VIIA group of the periodic table:

12. Write metalloids that are classified as nonmetals if we divide all the elements into metals and nonmetals?

13. Calculate the number of protons in 60 g of calcium phosphate.

14. Calculate the atomic mass of an element that have three isotopes with the following atomic masses and corresponding abundances: 285 u (85 %), 282 u (5 %), 287 u (10 %).

15. The percent of ²H in the sample of H₂O is 2 %, the percent of ¹H is 98 %. Find the volume of hydrogen produced in the reaction between 10 g of that H₂O sample with potassium.

LESSON 8

8.1 QUANTUM NUMBERS

In the quantum-mechanical model of an atom, the state of an electron is described by four quantum numbers. The first quantum number is called the *principal quantum number* (n). The principal quantum number largely determines the energy of an electron. Electrons in the same atom that have the same principal quantum number are said to occupy the same electron shell (layer) of the atom. The principal quantum number can be any nonzero positive integer: 1, 2, 3, 4, 5, 6, 7,

Within a shell, there may be multiple possible values of the next quantum number, the *angular momentum quantum number* (ℓ). The ℓ quantum number has a minor effect on

the energy of the electron but also affects the spatial distribution of the electron in three-dimensional space — that is, the shape of an electron's distribution in space. The value of the ℓ quantum number can be any integer between 0 and $n - 1$:

$$\ell = 0, 1, 2, \dots, n - 1$$

Thus, for a given value of n , there are different possible values of ℓ :

If n equals to	ℓ can be in numbers	in letters
1	0	s
2	0, 1	s, p
3	0, 1, 2	s, p, d
4	0, 1, 2, 3	s, p, d, f

Electrons within a shell that have the same value of ℓ are said to occupy the same subshell (sublayer) in the atom. Commonly, instead of referring to the numerical value of ℓ , a letter represents the value of ℓ (to help distinguish it from the principal quantum number).

Any s orbital is spherically symmetric, and there is only one orbital in any s subshell. Any p orbital has a two-lobed, dumbbell-like shape; because there are three of them, we normally represent them as pointing along the x -, y -, and z -axes. Four of the d orbitals are four-lobed rosettes, while the fifth one contains two lobes and a torus around the center of an atom. Shapes of seven f -orbitals are even more complicated than those of d -orbitals (Figure 8.1).

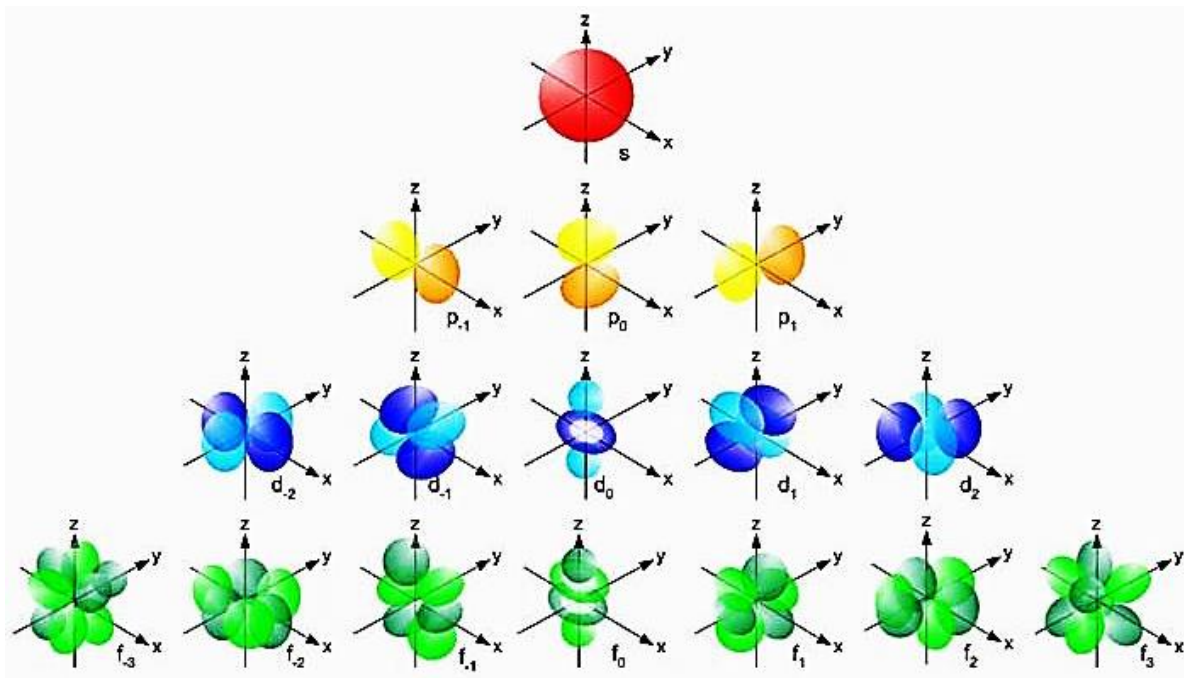


Figure 8.1 Shapes of electron orbitals

The next quantum number is called the *magnetic quantum number* (m_ℓ). For any value of ℓ , there are $2\ell + 1$ possible values of m_ℓ , ranging from $-\ell$ to ℓ :

$$-\ell \leq m_\ell \leq \ell$$

The following explicitly lists the possible values of m_ℓ for the possible values of ℓ :

If ℓ equals	The m_ℓ values can be
0	0
1	-1, 0, +1
2	-2, -1, 0, +1, +2
3	-3, -2, -1, 0, +1, +2, +3

The particular value of m_ℓ dictates the orientation of an electron's distribution in space. When ℓ is zero, m_ℓ can be only zero, so there is only one possible orientation. When ℓ is 1, there are three possible orientations for an electron's distribution. When ℓ is 2, there are five possible orientations of electron distribution. This goes on and on for other values of ℓ , but we need not consider any higher values of ℓ here. Each value of m_ℓ designates a certain orbital. Thus, there is only one orbital when ℓ is zero, three orbitals when ℓ is 1, five orbitals when ℓ is 2, and so forth. The m_ℓ quantum number has no effect on the energy of an electron.

The final quantum number is the *spin quantum number* (m_s). Electrons and other subatomic particles behave as if they are spinning (we cannot tell if they really are, but they behave as if they are). Electrons themselves have two possible spin states, and because of mathematics, they are assigned the quantum numbers $+1/2$ and $-1/2$. These are the only two possible choices for the spin quantum number of an electron.

8.2 ELECTRON CONFIGURATIONS OF ATOMS

Pauli exclusion principle: no two electrons in an atom can have the same set of four quantum numbers.

If you follow the arrows in order, they pass through the subshells in the order that they are filled with electrons in atoms (Figure 8.2). After the $3p$ subshell is filled, filling the $4s$ subshell first actually leads to a lesser overall energy than filling the $3d$ subshell.

The rule represented in figure 8.2 can be formulated in mathematical manner. According to the **Klechkovski's rule**, electrons first go to the orbital with lower sum of principal and angular momentum quantum numbers ($n + l$). If these values are equal, electrons go to the subshell with lower principal quantum number. So, if $n=4$ and $l=1$, the sum is 5. If $n=3$ and $l=2$, the sum is also 5. Electrons go to the $3d$ subshell and not to the $4p$ subshell, because of the lower value of n .

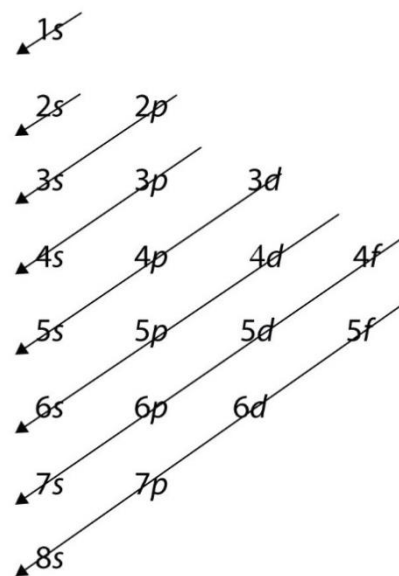
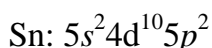


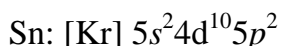
Figure 8.2. The scheme of the filling of electron subshells

Here is an example of complete electron configuration for Sn, which has 50 electrons.
 Sn: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^2$

Short electron configuration includes only the valence subshells.



The same electron configuration can be written as follows.



The first part of that form of electron configuration ([Kr]) means that Sn has the same electron configuration as Kr (the first noble gas with atomic number lower than that of the given element), while the second part ($5s^2 4d^{10} 5p^2$) shows electrons which are present in Sn and absent in Kr.

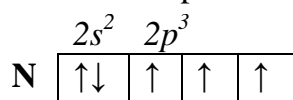
The shape of the periodic table mimics the filling of the subshells with electrons.

The periodic table is separated into blocks depending on which subshell is being filled for the atoms that belong in that section.

The electrons in the highest-numbered shell, plus any electrons in the last unfilled subshell, are called valence electrons; the highest-numbered shell is called the valence shell, while the inner electrons are called *core electrons*. The valence electrons largely control the chemistry of an atom.

Elements in each column of the Periodic table usually have the same valence shell electron configurations, and the elements have some similar chemical properties. This is strictly true for all elements in the *s* and *p* blocks. In the *d* and *f* blocks, because there are exceptions to the order of filling of subshells with electrons, similar valence shells are not absolute in these blocks. However, many similarities do exist in these blocks, so a similarity in chemical properties is expected.

Hund's Rule of Maximum Multiplicity is an observational rule which states that the atom is more stable when it has the highest number of unpaired electrons. Accordingly, it can be taken that if two or more orbitals of equal energy are available, electrons will occupy them singly before filling them in pairs. For example, electrons occupy three *p*-orbitals in Nitrogen atom, while they could occupy just two of them. In the last case there would be one pair of electrons. According to the Hund's rule that configuration is unstable, unlike the one represented below.



As one can see, electrons “enjoy their freedom”, just like people do when they are choosing their seats in the bus. There is a rule that helps to fill in cells (they represent orbitals) by arrows (they represent electrons) in a diagram shown above: first fill each cell in a subshell with an arrow looking up (that kind of arrow represents an electron with a positive spin), then fill the same cells with arrows looking down (with electrons of negative spin). That is how atoms usually behave in their ground state. Excitation (absorbance of energy) may cause deviations from all the rules of electron orbital fulfillment, except Pauli's principle (the maximal number of electrons in the same orbital is equal to 2). Also, the number of orbitals in each subshell cannot be changed even in the excited state. In other

words, in the excited state an electron may be shifted to the subshell of a higher energy. Such shifts are accompanied by pairing and unpairing of electrons.

Interestingly, ground state electron configurations of some d-elements, including chromium ($4s^13d^5$ and not $4s^23d^4$) and copper ($4s^13d^{10}$ and not $4s^23d^9$), look like electron configurations in the excited state. The first exception is usually explained by the Hund's rule: the state with 6 unpaired electrons is energetically favorable for chromium. The second exception is usually explained by the stability of completely filled d-block of copper.

Formation of an ion from a neutral atom is associated with the change of its electron configuration. Metals tend to lose all the electrons from their outer shell. For example, Na ($[\text{Ne}]3s^1$) turns to Na^+ ($[\text{Ne}]3s^0$). Nonmetals tend to gain electrons until the completion of their outer shells. For example, O ($[\text{He}]2s^22p^4$) turns to O^{2-} ($[\text{He}]2s^22p^6$, that is the same as $[\text{Ne}]$).

Questions:

- What does the principal quantum number characterize?
- What are the possible values of the principal quantum number?
- What does the angular momentum quantum number characterize?
- What are the possible values of the angular momentum quantum number?
- What does the magnetic quantum number characterize?
- What are the possible values of the magnetic quantum number?
- What does the spin quantum number characterize?
- What are the possible values of the spin quantum number?
- Formulate the Pauli exclusion principle.
- Formulate the Hund's rule.
- Formulate the Klechkovski's rule.

TEST FOR CLASSWORK

- How many energetic sublevels are there on the 4th energetic level?
a. 1 b. 2 c. 3 d. 4
- How many electron orbitals are there on the p-sublevel?
a. 1 b. 2 c. 3 d. 4
- What is the maximal number of electrons which can occupy all orbitals of the same d-sublevel?
a. 2 b. 6 c. 10 d. 14
- How many unpaired electrons are there in the nitrogen atom (in its normal state)?
a. 1 b. 2 c. 3 d. 4
- What is the number of electrons on the outer shell (level) of the chlorine atom:
a. 17 b. 8 c. 18 d. 7
- What is the maximal number of electrons on the same orbital?
a. 1 b. 2 c. 6 d. 10

7. Choose the correct order of electron orbitals fulfillment:
- a. $1s/2s/2p/3s/3d$ c. $1s/2s/3s/2p/3d$
 b. $1s/2s/3p/3s/3d$ d. $1s/2s/2p/3s/3p$
8. Choose the electron configuration of the nitrogen atom (in its normal state):
- a. $1s^2 2s^2 2p^3$ c. $1s^2 2s^3 2p^2$
 b. $1s^2 2s^1 2p^4$ d. $1s^1 2s^2 2p^4$
9. What element has the following electron configuration: $1s^2 2s^2 2p^6 3s^2 3p^1$
- a. Na b. Mg c. Al d. Si
10. Choose possible electron configurations (in both normal and excited states) for the carbon atom:
- a. $1s^2 2s^2 2p^2$ c. $1s^3 2s^2 2p^1$
 b. $1s^2 2s^1 2p^3$ d. $1s^2 2s^2 2p^1$

TEST FOR HOMEWORK

1. How many energetic sublevels are there on the 3rd energetic level?
- a. 1 b. 2 c. 3 d. 4
2. How many electron orbitals are there on the d-sublevel?
- a. 1 b. 3 c. 5 d. 7
3. What is the maximal number of electrons which can occupy all orbitals of the same f-sublevel?
- a. 2 b. 6 c. 10 d. 14
4. How many unpaired electrons are there in the oxygen atom (in its normal state)?
- a. 1 b. 2 c. 3 d. 4
5. What is the number of electrons on the outer shell (level) of the copper atom:
- a. 1 b. 2 c. 8 d. 18
6. How many d-electrons are there on the 3rd sublevel of the chromium atom (in its normal state)?
- a. 1 b. 3 c. 4 d. 5
7. Choose the correct order of electron orbitals fulfillment:
- a. ... $3s/3p/4s/3d$ c. ... $3s/3d/3p/4s$
 b. ... $3s/3d/3p/4s$ d. ... $3p/3d/3s/4s$
8. Choose the electron configuration of the zinc atom (in its normal state):
- a. $[Ar]4s^2 4p^6$ c. $[Ar]4s^1 4d^{10}$
 b. $[Ar]4s^2 3d^{10}$ d. $[Kr]5s^2 5d^{10}$
9. What element has the following electron configuration: $[Ne]3s^2 3p^4$
- a. S b. P c. Cl d. Ar

10. Choose possible electron configurations (in both normal and excited states) for the nitrogen atom:
- a. $1s^2 2s^2 2p^3$ c. $1s^3 2s^2 2p^2$
b. $1s^2 2s^1 2p^4$ d. $1s^2 2s^3 2p^2$

EXERCISES FOR CLASSWORK

1. Write the complete electron configuration for beryllium:

2. Write the complete electron configuration for bromine:

3. Write the complete electron configuration for scandium:

4. Write the short electron configuration for magnesium:

5. Write the short electron configuration for germanium:

6. Write the short electron configuration for titanium:

7. Draw the diagram with cells and arrows for the outer shell of carbon:

8. Draw the diagram with cells and arrows for the outer shell of sulfur:

9. Draw the diagram with cells and arrows for the outer shell of copper:

10. Find the mass of the product of the reaction between two pure chemical elements. The first one has an electronic configuration $1s^2 2s^2 2p^5$ and the volume of 5 L (in normal conditions). The second one has an electronic configuration $1s^2 2s^2 2p^6 3s^2$ and the mass of 0.5 g.

11. Which particles from this line (Na^+ / Mg^{2+} / F^- / Cl^- / Al^{3+} / Ne / Ar) have the same electronic configuration (if we ignore empty orbitals). Write these configurations.

12. Arrange the atoms (in their ground state) from this line (Ne / Cr / C / N / K / Mn) in the order of the increase of the number of unpaired electrons.

EXERCISES FOR HOMEWORK

1. Write the complete electron configuration for calcium:

2. Write the complete electron configuration for phosphorus:

3. Write the complete electron configuration for vanadium:

4. Write the short electron configuration for lithium:

5. Write the short electron configuration for arsenic:

6. Write the short electron configuration for cobalt:

7. Draw the diagram with cells and arrows for the outer shell of chrome:

8. Draw the diagram with cells and arrows for the outer shell of fluorine:

9. Draw the diagram with cells and arrows for the outer shell of nickel:

10. Find the mass of the product of the reaction between two pure chemical elements. The first one has an electronic configuration $1s^2 2s^2 2p^4$ and the volume of 7 L (in normal conditions). The second one has an electronic configuration $1s^2 2s^2 2p^6 3s^2 3p^4$ and the mass of 1.5 g.

11. Which particles from this line (K^+ / Ca^{2+} / I^- / Cl^- / S^{2-} / Xe / Ar) have the same electronic configuration (if we ignore empty orbitals). Write this configuration.

12. Arrange atoms and ions (in their ground state) from this line (He / Fe / Cu / P / Si / Mn^{2+}) in the order of the increase of the number of unpaired electrons.

LESSON 9

9.1 TYPES OF CHEMICAL BONDS

Ionic bonding typically occurs when it is easy for one atom to lose one or more electrons and for another atom to gain one or more electrons.

Electrons can move from one atom to another; when they do, species with overall electric charges are formed. Such species are called *ions*. Species with overall positive charges are termed *cations*, while species with overall negative charges are called *anions*. Metals tend to form cations, while nonmetals tend to form anions.

The magnitude of the charge (the number of lost or gained electrons) is listed as a right superscript next to the symbol of the element. If the charge is a single positive or negative one, the number 1 is not written (K^+ , Cl^-); if the magnitude of the charge is greater than 1, then the number is written usually *before* the “+” or “-” sign (Mn^{2+} , S^{2-}).

When electrons are shared between two atoms, they make a bond called a **covalent bond**. The equal sharing of electrons in a covalent bond is called a **nonpolar covalent bond**. For example, in H_2 molecule the two atoms involved in the covalent bond are both hydrogen atoms, each nucleus attracts both electrons equivalently. Thus the electron pair is equally shared by the two atoms.

A covalent bond between different atoms that attract the shared electrons by different amounts and cause an imbalance of electron distribution is called a **polar covalent bond**.

Covalent bond may be formed by the pair of electrons from one atom and an empty electron orbital of the second atom. In that case the first atom (with two paired electrons) is called “donor” and the second atom (with an empty orbital) is called acceptor. That kind of covalent bond is known as “donor-acceptor bond”, or “dative” bond.

Sigma (σ) bonds are the strongest type of covalent bonds due to the direct overlap of orbitals. They are formed by head-on overlapping between atomic orbitals (Figure 9.1).

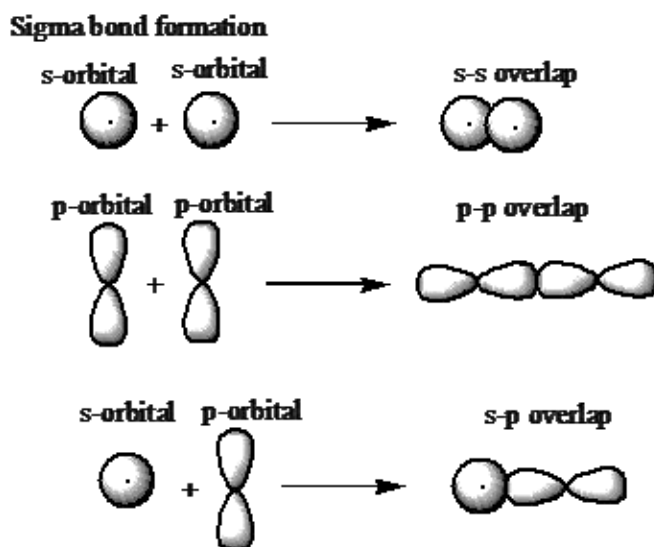


Figure 9.1. Formation of sigma bonds

Pi (π) bonds are usually weaker than sigma bonds. From the perspective of quantum mechanics, this bond's weakness is explained by significantly less overlap between the component p-orbitals due to their parallel orientation.

A typical double bond consists of one sigma bond and one pi bond; for example, the C=C double bond in ethylene (Figure 9.2).

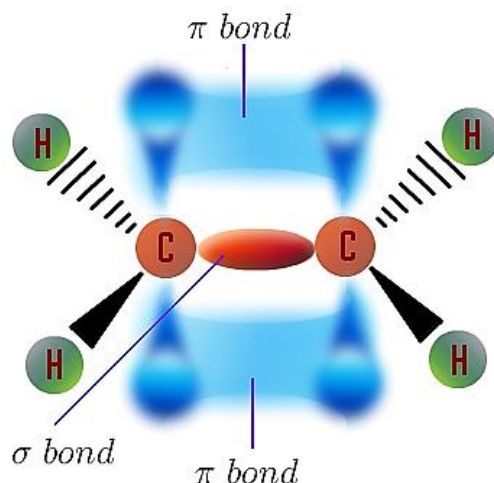


Figure 9.2. Sigma and pi bonds in the molecule of ethylene (H₂C = CH₂)

A typical triple bond, for example in acetylene (Figure 9.3), consists of one sigma bond and two pi bonds in two mutually perpendicular planes containing the bond axis. Two pi bonds are the maximum that can exist between a given pair of atoms.

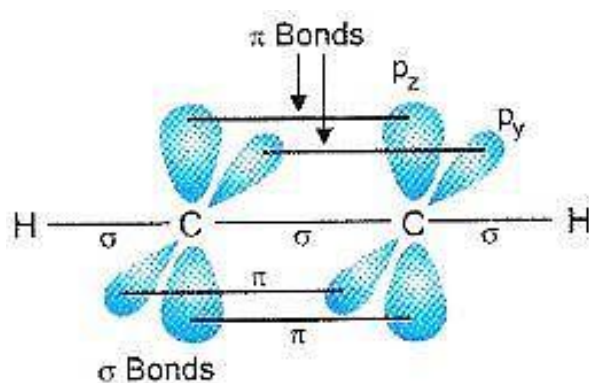


Figure 9.3. Sigma and pi bonds in the molecule of acetylene (HC \equiv CH)

A pi bond is weaker than a sigma bond, but the combination of pi and sigma bond is stronger than either bond by itself. For bonds that involve the same two elements, a double bond is stronger than a single bond, and a triple bond is stronger than a double bond.

A **hydrogen bond** is the electromagnetic attractive interaction of a polar hydrogen atom in a molecule or chemical group and an electronegative atom, such as nitrogen, oxygen or fluorine, from another molecule or chemical group. It is not a covalent chemical bond. The hydrogen atom has an attraction to another electronegative atom. These attractions can occur between molecules (*intermolecularly*), or within different parts of

a single molecule (*intramolecularly*). The hydrogen bond is weaker than covalent or ionic bonds. This type of bond occurs in both inorganic molecules such as water and organic molecules like DNA and proteins.

The electronegative atom attracts the electron cloud from around the hydrogen nucleus and, by decentralizing the cloud, leaves the atom with a positive partial charge. A hydrogen bond results when the positive charge of H^+ attracts a lone pair of electrons on another heteroatom, which becomes the hydrogen-bond *acceptor*. Once again, the same hydrogen atom should make polar covalent bond with fluorine, oxygen or nitrogen, and hydrogen bond with another electronegative atom (fluorine, oxygen or nitrogen).

Metallic bonding constitutes the electrostatic attractive forces between the delocalized electrons, called conduction electrons, gathered in an electron cloud, and the positively charged metal ions.

Questions:

- a. Describe the ionic bond
- b. Describe the covalent nonpolar bond
- c. Describe the covalent polar bond
- d. What is the difference between sigma and pi covalent bonds?
- e. Describe covalent double and triple bonds
- f. How does the fourth bond in NH_4^+ ion form?
- g. What is known about the mechanism of hydrogen bond formation?
- h. Describe bonds which are present in metals.

9.2 ELECTRONEGATIVITY

Any covalent bond between two different elements is polar. However, the degree of polarity is different for different bonds. A covalent bond between two different elements may be so slightly imbalanced that the bond is, approximately, nonpolar. A bond may be so polar that an electron actually transfers from one atom to another, forming a true ionic bond.

Electronegativity is a chemical property that describes the tendency of an atom or a functional group to attract electrons (or electron density) towards itself. The higher the electronegativity number, the more an element or a group of atoms attracts electrons towards it.

The covalent polar bond between two different atoms (A–B) is stronger than would be expected by taking the average of the strengths of the A–A and B–B bonds. This additional stabilization of the bond between atoms of two different elements is due to the contribution of ionic component to the bonding. There is an additional energy that comes from ionic factors, i. e. from the polar character of the bond.

In general, electronegativity increases from left to right along each period, and decreases from the top to the bottom of every group. Fluorine is the most electronegative of the elements, while cesium and francium are the least electronegative ones. In Figure 9.1 we

show the electronegativity scale created by Linus Pauling. In the Appendix I one can see edited electronegativity scale with some deviations from the originally proposed one.

H																
2.1																
Li	Be											B	C	N	O	F
1.0	1.5											1.5	2.5	3.0	3.5	4.0
Na	Mg											Al	Si	P	S	Cl
0.9	1.2											1.5	1.8	2.1	2.5	3.0
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br
0.8	1.0	1.3	1.5	1.6	1.6	1.5	1.8	1.9	1.8	1.9	1.6	1.6	1.8	2.0	2.4	2.8
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I
0.8	1.0	1.2	1.4	1.6	1.8	1.9	2.2	2.2	2.2	1.9	1.7	1.7	1.8	1.9	2.1	2.5
Cs	Ba		Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At
0.7	0.9		1.3	1.5	1.7	1.9	2.2	2.2	2.2	2.4	1.9	1.8	1.9	1.9	2.0	2.2
Fr	Ra															
0.7	0.9															

Figure 9.1. The table of chemical elements with their electronegativity levels

The polarity of a covalent bond can be judged by determining the difference of the electronegativities of the two atoms involved in the covalent bond, as summarized in the following table.

Table 9.1

Thresholds in the difference in electronegativity levels between two atoms

Difference in electronegativity between two atoms	Bond Type
0	nonpolar covalent
$0 < 0.4$	slightly polar covalent
0.4–1.9	definitely polar covalent
> 1.9	likely ionic

The real nature of a bond in a given compound must be estimated in several different experiments (X-ray crystallography, Raman spectroscopy, conductometry, etc.). The nature of a bond sometimes changes with the increase or decrease of temperature. However, you can predict the nature of a bond using some simple rules. A bond between two identical atoms of a nonmetal or between two identical parts of a large molecule is covalent nonpolar. A bond between two different nonmetals or between two nonidentical parts of a large molecule is covalent polar. A bond between metal and nonmetal or between cation and anion is ionic. A bond between identical of a metal or between different metals is metallic. Remember that there may be different types of bonds in the same compound. For example, in Na_2SO_4 the bonds between oxygen and sulfur are covalent polar, while the bonds between Na^+ cations and SO_4^{2-} anions are ionic.

TEST FOR CLASSWORK

- Choose binary compounds with ionic bonds:
a. CCl_4 b. KCl c. ZnO d. SiO
- Choose compounds with both ionic and covalent polar bonds:
a. NaCl b. KNO_3 c. NO_2 d. KOH
- Choose binary compounds with covalent polar bonds:
a. PCl_3 b. Na_3N c. K_2O d. KH
- In which compounds one can find at least one covalent nonpolar bond?
a. H_2 b. C_2H_6 c. H_2O_2 d. H_2O
- Choose substances with the metallic bonding:
a. AgCl b. KAlO_2 c. AgAu d. Cu_9Zn
- In which compounds there are just sigma covalent bonds?
a. CH_4 b. C_2H_2 c. C_2H_4 d. C_5H_{12}
- Which compounds contain a double bond?
a. O_2 b. N_2 c. H_2 d. C_2H_4
- Which compounds contain a triple bond?
a. C_2H_2 b. CO c. N_2 d. O_3
- Choose a compound with the most polar covalent bond:
a. NaF b. HF c. H_2O d. H_2S
- Indicate the possible schemes of hydrogen bond formation:
a. $\text{N-H}\cdots\text{O}$ c. $\text{F-H}\cdots\text{F}$
b. $\text{O-H}\cdots\text{N}$ d. $\text{N-H}\cdots\text{C}$

TEST FOR HOMEWORK

- Choose compounds with ionic bonds:
a. Ba(OH)_2 b. H_2SO_4 c. KNO_3 d. SiO_2
- Choose compounds with covalent polar bonds:
a. Cl_2 b. ZnSO_4 c. NH_3 d. ZnO
- Choose diatomic molecules with covalent nonpolar bonds:
a. P_4 b. N_2 c. O_2 d. S_8
- In which compounds one can find at least one covalent polar bond?
a. Br_2 b. CH_3Cl c. N_2O d. LiCl
- Choose substances with metallic bonding:
a. NaH b. SnCu_4 c. K_2ZnO_2 d. KNa
- In which compounds there are pi-bonds?
a. C_3H_8 b. C_3H_4 c. C_2H_4 d. O_2

7. Which compounds contain a double bond?
 a. I₂ b. P₄ c. SO₂ d. SO₃
8. Which compounds contain a triple bond?
 a. CaC₂ b. CO₂ c. KCN d. S₈
9. Choose a compound with the most polar bond:
 a. KCl b. LiF c. HF d. BF₃
10. Indicate the possible schemes of hydrogen bond formation:
 a. N-H.....F c. O-H.....O
 b. P-H.....P d. N-H.....N

EXERCISES FOR CLASSWORK

1. Write 5 samples of substances with just covalent nonpolar bonds:

2. Write 5 samples of substances with just ionic bonds:

3. Write 5 samples of substances with just covalent polar bonds:

4. Write 5 samples of substances with metallic bonding:

5. Arrange substances in this line (HF / H₂O / BH₃ / CH₄ / H₂S) in the order of the increase of the polarity of a bond between atoms of different elements:

6. Describe all the bonds in the following compounds:
 CO: _____
 HI: _____
 H₂S: _____
 OF₂: _____
 CH₄: _____
 H₂SO₄: _____
 NaOH: _____
 KMnO₄: _____
 SbSn: _____

EXERCISES FOR HOMEWORK

1. Write 5 samples of substances with both covalent polar and ionic bonds:

2. Write 5 samples of substances with both covalent nonpolar and ionic bonds:

3. Write 5 samples of substances with both covalent polar and nonpolar bonds:

4. Write 5 samples of substances which can make intermolecular hydrogen bonds:

5. Write 5 samples of substances with a triple bond:

6. Describe all the bonds in the following compounds:

SiH₄: _____

FeCl₃: _____

KCl: _____

Al₂O₃: _____

LiH: _____

HNO₃: _____

Ca(OH)₂: _____

K₂Cr₂O₇: _____

Cu₃₁Sn₃: _____

LESSON 10

10.1 OXIDATION STATE

The *oxidation state* is an indicator of the degree of oxidation of an atom in a chemical compound. The formal oxidation state is the *hypothetical charge* that an atom would have if all bonds to atoms of different elements were 100 % ionic. Oxidation states are typically represented by integers, which can be positive, negative, or equal to zero.

Oxidation state of pure chemical element is equal to zero.

Hydrogen has an oxidation state of +1 and oxygen has an oxidation state of -2 when they are present in the most of compounds. Exceptions to this rule are hydrides of active metals (such as LiH, NaH, CaH₂), silane (SiH₄), boranes (B₂H₆ and others), peroxides (H₂O₂, Na₂O₂), superoxides (KO₂) and compounds made from fluorine and oxygen (F₂O and F₂O₂).

Alkali metals (Li, Na, K, Rb, Cs, Fr) have an oxidation state of +1, while metals from the IIA group (Be, Mg, Ca, Sr, Ba) and Zinc (Zn) have an oxidation state of +2 in almost all of their compounds. Aluminum always demonstrates +3 oxidation state.

The main rule to calculate oxidation states of each of the atoms in a compound is as follows. The algebraic sum of oxidation states of all atoms in a compound must be equal to zero, since compounds have no charge. Ion has either positive or negative charge. That is why the algebraic sum of the oxidation states of the atoms from a given ion is equal to the total charge of the ion.

Oxidation state historically was thought to reflect the number of electrons transferred from one atom to the other. The number of electrons is an integer. However, there are many compounds in which oxidation states for some atoms are not integers (for example, in the Fe_3O_4 compound). That compound can really be represented by the formula $\text{FeO}\cdot\text{Fe}_2\text{O}_3$.

Most elements have more than one possible oxidation state. Carbon has nine integer oxidation states, and there are also molecules in which the average degree of oxidation of several carbons is fractional. For example, try to calculate oxidation state of carbon in propane molecule (C_3H_8).

10.2 OXIDATION STATE, OXIDATION NUMBER AND VALENCE

The term “**oxidation number**” has a different meaning than the term “**oxidation state**”. The term “oxidation number” is used *in coordination chemistry only*. It reflects the charge that central atom would have if all the ligands were removed along with the electron pairs that were shared with the central atom.

In certain compounds the module of the oxidation state is equal to the valence of an atom. One should realize that valence and oxidation state may be different for a given atom. For example, the valence of nitrogen in its diatomic molecule is equal to three (III), since there is a triple bond between two atoms. The oxidation state for both atoms from N_2 is equal to 0. Strange situation is known for CO molecule in which oxidation state of oxygen is equal to -2 , while its valence is equal to III (third bond is formed by donor-acceptor mechanism).

The term “oxidation state” is good for the description of the structure of binary ionic compounds in which electrons from much less electronegative atoms are really gained by the much more electronegative atoms. The term “valence” is good for the description of the structure of molecules with covalent bonds. However, there is a sense to use the oxidation state for the description of molecules and polyatomic ions to show the correspondence between oxides, hydroxides and salts, as well as for the balancing of redox reactions.

Maximal oxidation state is usually equal to the number of the subgroup (written in Roman numbers) in which the given element is located in the Periodic table.

Minimal oxidation state of the metal is equal to zero. Minimal oxidation state of the nonmetal is equal to the eight minus the number of its subgroup (written in Roman numbers).

Questions:

- Give the definition of the term “oxidation state”
- Give the definition of the term “oxidation number”
- Give the definition of the term “valence”
- How to calculate oxidation state in a compound?
- How to calculate oxidation state in an ion?

TEST FOR CLASSWORK

- Determine the oxidation state of phosphorus in H_3PO_4 :
a. +5 b. +3 c. +1 d. -3
- Choose compounds in which the oxidation state of nitrogen is equal to +3:
a. NH_3 b. NaNO_3 c. N_2O_3 d. HNO_2
- Choose anions in which the oxidation state of phosphorus is equal to +5:
a. PO_4^{3-} b. HPO_4^{2-} c. H_2PO_4^- d. H_2PO_3^-
- Calculate the charge of the ion made from three oxygen atoms and one silicon atom in its maximal oxidation state:
a. -2 b. +2 c. -3 d. +3
- Choose oxidation states possible for hydrogen atoms:
a. +1 b. -1 c. +2 d. 0
- For which compounds the term «oxidation state» is more applicable than the term «valence»?
a. K_2O b. NaCl c. PH_3 d. SiH_4
- For which compounds the term «valence» is more applicable than the term «oxidation state»?
a. PCl_3 b. N_2 c. LiF d. C_3H_8
- Choose the minimal oxidation state for sulfur:
a. 0 b. +6 c. +4 d. -2
- Choose the maximal oxidation state for chlorine:
a. -1 b. +1 c. +3 d. +7
- Which chemical elements demonstrate a single possible oxidation state in compounds?
a. H b. K c. F d. O

TEST FOR HOMEWORK

- Determine the oxidation state of sulfur in H_2SO_4 :
a. +4 b. +6 c. 0 d. -2
- Choose compounds in which the oxidation state of oxygen is equal to -2:
a. NO b. K_2O c. K_2O_2 d. KO_2
- Choose anions in which the oxidation state of carbon is equal to +4:
a. CO_3^{2-} b. HCOO^- c. HCO_3^- d. $\text{C}_2\text{O}_4^{2-}$
- Calculate the charge of the ion made from four oxygen atoms and one phosphorus atom in its maximal oxidation state:
a. -2 b. +2 c. -3 d. +3
- Choose oxidation states possible for nitrogen atoms:
a. +1 b. +3 c. +5 d. +7

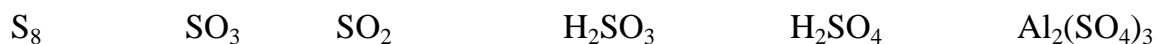
6. For which compounds the term «oxidation state» is more applicable than the term «valence»?
- a. NaBr b. Cl₂ c. CH₄ d. CaO
7. For which compounds the term «valence» is more applicable than the term «oxidation state»?
- a. NH₃ b. BaO c. NaF d. H₂S
8. Choose the minimal oxidation state for carbon:
- a. 0 b. -6 c. -4 d. -2
9. Choose the maximal oxidation state for manganese:
- a. -1 b. +1 c. +3 d. +7
10. Which chemical elements demonstrate a single possible oxidation state in compounds?
- a. Ca b. Zn c. Cr d. Mn

EXERCISES FOR CLASSWORK

1. Write oxidation states upon all the elements in the following compounds:



2. Write oxidation states upon all the elements in the following compounds:



3. Write oxidation states upon all the elements in the following compounds:



4. Write oxidation states upon all the elements in the following compounds:



5. Write oxidation states upon all the elements in the following compounds:



6. Write oxidation states upon all the elements in the following ions:



7. Arrange compounds in this line (CsCl / MgO / CrO₃ / Al₂O₃ / Mn₂O₇) in the order of the increase of the oxidation state of a metal:
-

8. Write 5 samples of substances in which oxygen demonstrates oxidation state different from -2:
-

EXERCISES FOR HOMEWORK

1. Write oxidation states upon all the elements in the following compounds:

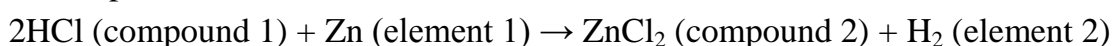


2. Write oxidation states upon all the elements in the following compounds:
 Cr K₂Cr₂O₇ CrO BaCrO₄ NaCrO₂ Cr₂O₃
 3. Write oxidation states upon all the elements in the following compounds:
 NaCl NaClO NaClO₂ NaClO₃ NaClO₄ Cl₂
 4. Write oxidation states upon all the elements in the following compounds:
 Fe FeO Fe(OH)₃ Na₂FeO₄ Fe₃O₄ Fe₂(SO₄)₃
 5. Write oxidation states upon all the elements in the following compounds:
 HI I₂ HIO₃ NaIO₄ KIO Ca(IO₂)₂
 6. Write oxidation states upon all the elements in the following ions:
 NO₂⁻ SiO₃²⁻ [Al(OH)₆]³⁻ MnO₄⁻ NH₄⁺ SO₃²⁻
 7. Arrange compounds in this line (SO₃ / OF₂ / SiO₂ / P₂O₅ / As₂O₃) in the order of the increase of the oxidation state of a nonmetal (which is not oxygen):
-
8. Write 5 samples of substances in which hydrogen demonstrates oxidation state different from +1:
-

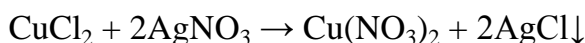
LESSON 11

11.1 SEVERAL WAYS TO CLASSIFY CHEMICAL REACTIONS

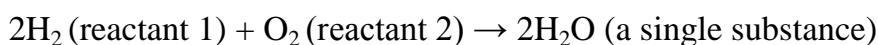
A single-replacement reaction is a chemical reaction in which one element is substituted for another element in a compound, generating a new element and a new compound as products.



A double-replacement reaction occurs when parts of two compounds are exchanged, making two new compounds. A characteristic of a double-replacement equation is that there are two compounds as reactants and two different compounds as products.



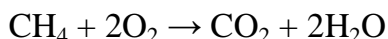
A composition reaction (*combination reaction* or a *synthesis reaction*) is a chemical reaction in which a single substance is produced from multiple reactants. A single substance as a product is the key characteristic of the composition reaction. There may be a coefficient other than one for the substance, but if the reaction has only a single substance as a product, it can be called a composition reaction.



A decomposition reaction starts from a single substance and produces more than one substance. One substance as a reactant and more than one substance as the products is the key characteristic of a decomposition reaction.

2NaHCO_3 (a single substance) \rightarrow Na_2CO_3 (product 1) + CO_2 (product 2) + H_2O (product 3)

A **combustion reaction** is a chemical reaction in which a reactant combines with oxygen to produce oxides of all other elements from that compound as products. Combustion is not any reaction between a substance and oxygen, but only the one that is associated with the fast burning process (there must be flames of fire, smoke, the release of big amount of heat and light). Some substances can slowly react with oxygen without any visible signs of the burning. Such processes are usually called oxidation or, sometimes, corrosion, but not combustion.

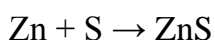


An **exothermic reaction** is a chemical reaction that releases energy in the form of heat. It is the opposite of an **endothermic reaction** in which the system absorbs energy from its surroundings in the form of heat.

A **reversible reaction** is a chemical reaction that results in an equilibrium mixture of reactants and products. If the concentrations of the reactants at equilibrium are very small, then such a reaction may be considered to be an (almost) **irreversible reaction**, although in reality small amounts of the reactants are still expected to be present in the reacting system. A truly irreversible chemical reaction is usually achieved when one of the products completely exits the reacting system in form of gas or insoluble substance.

11.2 REDOX REACTIONS

If the oxidation state of at least one element changes due to chemical reaction, that reaction is called reduction-oxidation (**redox**) reaction. In the simplest redox reaction there is a single substance called **oxidizer** which contains atoms gaining electrons, and a single substance called **reducer** which contains atoms losing electrons. We may say that atoms of oxidizer take electrons from atoms of reducer. In the example below sulfur takes electrons from zinc. Oxidation state of sulfur changes from 0 to -2 . Oxidation state of zinc changes from 0 to $+2$.



So, Sulfur is an oxidizer, while Zinc is a reducer.

Questions:

- Describe key characteristic of a single-displacement reaction
- Describe key characteristic of a double-displacement reaction
- Describe key characteristic of a composition reaction.
- Describe key characteristic of a decomposition reaction.
- Describe key characteristic of a combustion reaction.
- What is the difference between exothermic and endothermic reactions?
- What is the difference between reversible and irreversible reactions?
- How can you understand that a given reaction is redox reaction?

TEST FOR CLASSWORK

- Choose composition reactions:
 - $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$
 - $2\text{Na} + 2\text{H}_2\text{O} \rightarrow 2\text{NaOH} + \text{H}_2$
 - $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$
 - $\text{Zn} + \text{S} \rightarrow \text{ZnS}$
- Choose decomposition reactions:
 - $\text{NH}_3 + \text{HCl} \rightarrow \text{NH}_4\text{Cl}$
 - $2\text{KNO}_3 \rightarrow 2\text{KNO}_2 + \text{O}_2$
 - $\text{N}_2\text{O}_4 \rightarrow 2\text{NO}_2$
 - $2\text{Ag}_2\text{O} \rightarrow 4\text{Ag} + \text{O}_2$
- Choose single replacement reactions:
 - $2\text{K} + 2\text{H}_2\text{O} \rightarrow 2\text{KOH} + \text{H}_2$
 - $2\text{Al} + 6\text{HCl} \rightarrow 2\text{AlCl}_3 + 3\text{H}_2$
 - $\text{KCl} + \text{AgNO}_3 \rightarrow \text{AgCl} + \text{KNO}_3$
 - $\text{Zn} + \text{CuSO}_4 \rightarrow \text{ZnSO}_4 + \text{Cu}$
- Choose double replacement reactions:
 - $2\text{KOH} + \text{H}_2\text{SO}_4 \rightarrow \text{K}_2\text{SO}_4 + 2\text{H}_2\text{O}$
 - $\text{CaO} + 2\text{HCl} \rightarrow \text{CaCl}_2 + \text{H}_2\text{O}$
 - $\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$
 - $\text{ZnCl}_2 + 2\text{AgNO}_3 \rightarrow 2\text{AgCl} + \text{Zn}(\text{NO}_3)_2$
- Choose combustion reactions:
 - $4\text{NH}_3 + 3\text{O}_2 \rightarrow 2\text{N}_2 + 6\text{H}_2\text{O}$
 - $\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}$
 - $\text{Mg} + \text{S} \rightarrow \text{MgS}$
 - $2\text{Fe} + \text{O}_2 \rightarrow 2\text{FeO}$
- Which of the equations written below represent reduction-oxidation (redox) reactions?
 - $2\text{HNO}_3 + \text{Ca}(\text{OH})_2 \rightarrow \text{Ca}(\text{NO}_3)_2 + 2\text{H}_2\text{O}$
 - $2\text{H}_2\text{S} + 3\text{O}_2 \rightarrow 2\text{SO}_2 + 2\text{H}_2\text{O}$
 - $\text{Cl}_2 + \text{H}_2\text{O} \rightarrow \text{HCl} + \text{HClO}$
 - $3\text{CaO} + \text{P}_2\text{O}_5 \rightarrow \text{Ca}_3(\text{PO}_4)_2$
- In which reactions hydrogen atoms act as reducers?
 - $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$
 - $\text{H}_2 + 2\text{K} \rightarrow 2\text{KH}$
 - $2\text{HCl} + \text{Zn} \rightarrow \text{ZnCl}_2 + \text{H}_2$
 - $\text{HCl} + \text{NH}_3 \rightarrow \text{NH}_4\text{Cl}$
- In which reactions sulfur atoms act as oxidizers:
 - $\text{H}_2 + \text{S} \rightarrow \text{H}_2\text{S}$

- b. $\text{Ca} + \text{S} \rightarrow \text{CaS}$
 c. $\text{H}_2\text{SO}_4 \text{ (dilute)} + \text{Zn} \rightarrow \text{ZnSO}_4 + \text{H}_2$
 d. $2\text{H}_2\text{SO}_4 \text{ (concentrated)} + \text{Zn} \rightarrow \text{ZnSO}_4 + \text{SO}_2 + 2\text{H}_2\text{O}$
9. Choose appropriate characteristics of the following reaction: $3\text{S} + 2\text{H}_2\text{O} \rightarrow 2\text{H}_2\text{S} + \text{SO}_2$
 a. redox reaction
 b. single displacement reaction
 c. double displacement reaction
 d. disproportioning reaction for sulfur atoms
10. Choose appropriate characteristics of the following reaction: $\text{Zn} + \text{CuSO}_4 \rightarrow \text{ZnSO}_4 + \text{Cu}$
 a. redox reaction
 b. single displacement reaction
 c. double displacement reaction
 d. composition reaction

TEST FOR HOMEWORK

1. Choose composition reactions:
 a. $2\text{Ca} + \text{O}_2 \rightarrow 2\text{CaO}$
 b. $2\text{CaS} + 3\text{O}_2 \rightarrow 2\text{CaO} + 2\text{SO}_2$
 c. $\text{CO}_2 + \text{H}_2\text{O} + \text{CaCO}_3 \rightarrow \text{Ca}(\text{HCO}_3)_2$
 d. $\text{Zn} + \text{H}_2\text{S} \rightarrow \text{ZnS} + \text{H}_2$
2. Choose decomposition reactions:
 a. $\text{NH}_4\text{OH} \rightarrow \text{NH}_3 + \text{H}_2\text{O}$
 b. $\text{Ba}(\text{OH})_2 \rightarrow \text{BaO} + \text{H}_2\text{O}$
 c. $2\text{CH}_4 \rightarrow \text{C}_2\text{H}_2 + 3\text{H}_2$
 d. $2\text{CO} + \text{O}_2 \rightarrow 2\text{CO}_2$
3. Choose single replacement reactions:
 a. $\text{SiO}_2 + \text{CaCO}_3 \rightarrow \text{CaSiO}_3 + \text{CO}_2$
 b. $\text{Cl}_2 + 2\text{KI} \rightarrow \text{I}_2 + 2\text{KCl}$
 c. $\text{Mg} + \text{H}_2\text{O} \text{ (t}^\circ\text{)} \rightarrow \text{MgO} + \text{H}_2$
 d. $\text{Al}(\text{OH})_3 + \text{KOH} \text{ (t}^\circ\text{)} \rightarrow \text{KAlO}_2 + 2\text{H}_2\text{O}$
4. Choose double replacement reactions:
 a. $\text{BaCl}_2 + \text{H}_2\text{SO}_4 \rightarrow \text{BaSO}_4 + 2\text{HCl}$
 b. $\text{NaOH} + \text{HCl} \rightarrow \text{NaCl} + \text{H}_2\text{O}$
 c. $\text{I}_2 + \text{H}_2 \rightarrow 2\text{HI}$
 d. $\text{Zn}(\text{OH})_2 + \text{HCl} \rightarrow \text{ZnOHCl} + \text{H}_2\text{O}$
5. Choose combustion reactions:
 a. $4\text{Fe} + 3\text{O}_2 \rightarrow 2\text{Fe}_2\text{O}_3$
 b. $2\text{Mg} + \text{CO}_2 \rightarrow 2\text{MgO} + \text{C}$
 c. $\text{CaO} + \text{H}_2\text{O} \rightarrow \text{Ca}(\text{OH})_2$
 d. $2\text{NO} + \text{O}_2 \rightarrow 2\text{NO}_2$

6. Which of the equations written above represent reduction-oxidation (redox) reactions?
- $\text{H}_2\text{SO}_3 \rightarrow \text{SO}_2 + \text{H}_2\text{O}$
 - $4\text{HNO}_3 \rightarrow 4\text{NO}_2 + \text{O}_2 + 2\text{H}_2\text{O}$
 - $\text{K}_2\text{O} + \text{H}_2\text{O} \rightarrow 2\text{KOH}$
 - $2\text{NO}_2 + 2\text{KOH} \rightarrow \text{KNO}_2 + \text{KNO}_3 + \text{H}_2\text{O}$
7. In which reactions nitrogen atoms act as reducers?
- $3\text{H}_2 + \text{N}_2 \rightarrow 2\text{NH}_3$
 - $4\text{NH}_3 + 5\text{O}_2 \xrightarrow{\text{catalyst}} 4\text{NO} + 6\text{H}_2\text{O}$
 - $\text{NH}_3 + \text{HNO}_3 \rightarrow \text{NH}_4\text{NO}_3$
 - $(\text{NH}_4)_2\text{Cr}_2\text{O}_7 \xrightarrow{\text{t}^\circ} \text{N}_2 + 4\text{H}_2\text{O} + \text{Cr}_2\text{O}_3$
8. In which reactions oxygen atoms act as oxidizers:
- $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$
 - $3\text{O}_2 \rightarrow 2\text{O}_3$
 - $\text{K}_2\text{Cr}_2\text{O}_7 + 2\text{KOH} \rightarrow 2\text{K}_2\text{CrO}_4 + \text{H}_2\text{O}$
 - $2\text{H}_2\text{O}_2 + \text{S} \rightarrow \text{SO}_2 + 2\text{H}_2\text{O}$
9. Choose appropriate characteristics of the following reaction: $\text{NH}_4\text{NO}_2 \rightarrow \text{N}_2 + 2\text{H}_2\text{O}$
- redox reaction
 - composition reaction
 - decomposition reaction
 - conproportioning reaction for nitrogen atoms
10. Choose appropriate characteristics of the following reaction:
 $\text{Ca}(\text{HCO}_3)_2 \rightarrow \text{CaO} + 2\text{CO}_2 + \text{H}_2\text{O}$
- redox reaction
 - single displacement reaction
 - double displacement reaction
 - decomposition reaction

EXERCISES FOR CLASSWORK

1. Write 3 samples of combination (composition) reaction for chlorides:

2. Write 3 samples of decomposition reaction for salts:

3. Write 3 samples of single displacement reaction with HCl:

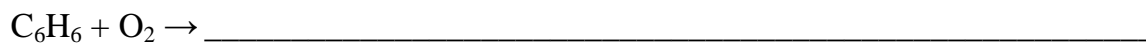
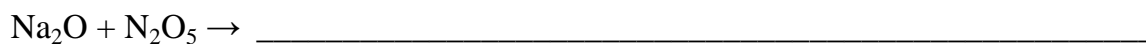
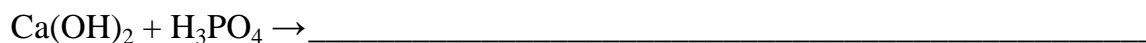
4. Write 3 samples of double displacement reaction with AgNO_3 :

5. Write 3 samples of neutralization reaction with KOH :

6. Write 3 samples of combustion reaction in which CO_2 is a product:

7. Write 2 samples of redox reaction in which hydrogen atoms are oxidizers:

8. Finish chemical reactions and classify them:



EXERCISES FOR HOMEWORK

1. Write 3 samples of combination (composition) reaction for sulfides:

2. Write 3 samples of decomposition reaction for acids:

3. Write 3 samples of single displacement reaction with H_2SO_4 :

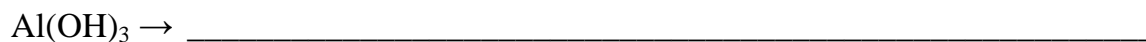
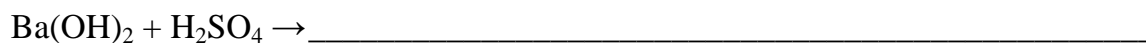
4. Write 3 samples of double displacement reaction with BaCl_2 :

5. Write 3 samples of neutralization reaction with $\text{Ba}(\text{OH})_2$:

6. Write 3 samples of combustion reaction in which SO_2 is a product:

7. Write 2 samples of redox reaction in which hydrogen atoms are reducers:

8. Finish chemical reactions and classify them:



LESSON 12

12.1 DEFINITIONS OF REDUCTION AND OXIDATION

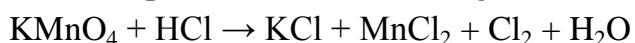
The increase in oxidation state of an atom (when it becomes *more positive*) through a chemical reaction is known as an **oxidation**; a decrease in oxidation state of an atom (when it becomes *more negative*) is known as a **reduction**.

Oxidation-reduction (*redox*) reactions involve the formal transfer of electrons: a net gain in electrons being a reduction and a net loss of electrons being an oxidation. The process of losing and gaining electrons occurs simultaneously. However, mentally we can separate the two processes. The total number of electrons being lost must be equal to the total number of electrons being gained for a redox reaction to be balanced.

12.2 BALANCING REDUCTION-OXIDATION REACTIONS

To balance redox reaction one should write each oxidation and reduction reaction separately, listing the number of electrons explicitly in each. Individually, the oxidation and reduction reactions are called half-reactions. Then one should multiply each half-reaction by the certain integer number until the number of electrons on each side cancels completely. Then one should use those two integers as coefficients in the whole reaction.

The example of redox reaction is given below.

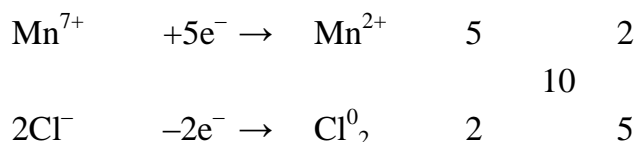


The first thing one should do is to calculate oxidation states:

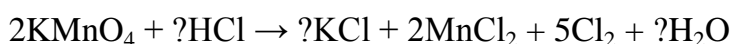


The second step is to notice those atoms for which the oxidation state has been changed (they are written in bold underlined font).

The third step is to write down two half reactions. The first half reaction is the reaction of reduction (Mn^{7+} has gained 5 electrons). The second half reaction is the reaction of oxidation (two Cl^- have lost 2 electrons). Mn^{7+} acted as oxidizer, while Cl^- acted as reducer. There is a cool mnemonic phrase which helps not to confuse oxidation and reduction. “*LEO the lion says GER*” — Loss of Electrons is Oxidation, Gain of Electrons is Reduction.



The number of gained electrons should be equal to the number of lost electrons. If we multiply the first half-reaction by 2 then 10 electrons will be gained. If we multiply the second half-reaction by 5 then 10 electrons will be lost. 2 is the coefficient before Mn^{7+} (KMnO_4) and Mn^{2+} (MnCl_2). 5 is the coefficient before Cl_2 and not HCl , since only a part of Cl^- has been reduced to form Cl_2 .



As one can see, electron balance helped to find out 3 from 6 coefficients in the reaction. Other coefficients can be found easily.

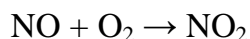


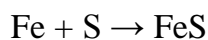
Remember that electron balance is *just a mathematical method* to find out coefficients in chemical reactions. It has a little (if any) chemical or physical sense.

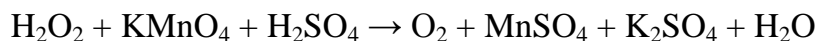
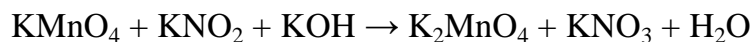
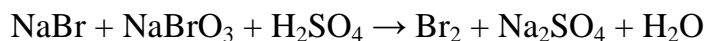
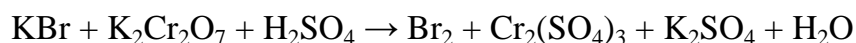
There are special types of redox reactions: disproportioning reactions (when atoms of the same element have the same oxidation state in reactants and different oxidation states in products) and conproportioning reactions (when atoms of the same element have different oxidation states in reactants and the same oxidation state in products). Different elements from the same compound may also act as reducer and oxidizer.

EXERCISES FOR CLASSWORK

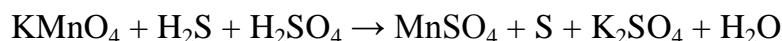
1. **Balance the following redox reactions using electron balancing, provide half-reactions of oxidation and reduction:**







2. What is the coefficient before KMnO_4 in the following reaction?

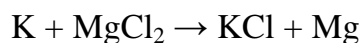
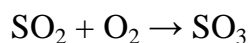


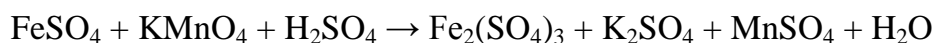
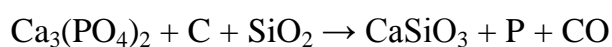
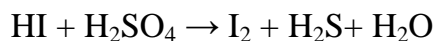
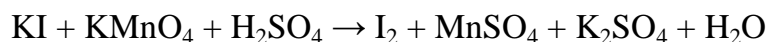
3. What part of the total H_2SO_4 amount participated in the following reaction really acted as an oxidizer?



EXERCISES FOR HOMEWORK

1. Balance the following redox reactions using electron balancing, provide half-reactions of oxidation and reduction:

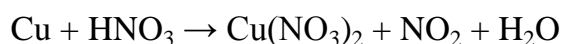




2. What is the coefficient before $\text{K}_2\text{Cr}_2\text{O}_7$ in the following reaction?



3. What part of the total HNO_3 participated in the following reaction really acted as an oxidizer?

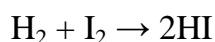


LESSON 13

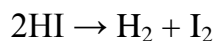
13.1 CHEMICAL EQUILIBRIUM

Chemical equilibrium is the state in which both reactants and products are present at concentrations which have no further tendency to change with time. Usually, this state results when the forward reaction proceeds at the same rate as the reverse reaction. This process is called *dynamic equilibrium*.

Consider the following reaction occurring in a closed container (so that no material can go in or out):

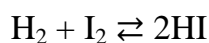


The way the equation is written, we are led to believe that the reaction goes to completion, that all the H_2 and the I_2 react to make HI. However, this is not the case. The reverse chemical reaction is also taking place:



It acts to undo what the first reaction does. Eventually, the reverse reaction proceeds so quickly that it matches the speed of the forward reaction. When that happens, any continued overall reaction stops: the reaction has reached chemical equilibrium (sometimes just spoken as *equilibrium*; plural *equilibria*), the point at which the forward and reverse processes balance each other's progress.

Because two opposing processes are occurring at once, it is conventional to represent an equilibrium using a double arrow:



The double arrow implies that the reaction is going in both directions.

Questions:

1. Define *chemical equilibrium*. Give an example.
2. Explain what is meant when it is said that chemical equilibrium is dynamic.

13.2 THE LAW OF MASS ACTION

The law of mass action relates the amounts of reactants and products at equilibrium for a chemical reaction.

The equilibrium constant is written as K_{eq} .

The K_{eq} is a characteristic numerical value for a given reaction at a given temperature; that is, each chemical reaction has its own characteristic K_{eq} .

Coming back to $\text{H}_2 + \text{I}_2 \leftrightarrow 2\text{HI}$ reaction, the K_{eq} is expressed as:

$$K_{\text{eq}} = [\text{HI}]^2 / ([\text{H}_2] \cdot [\text{I}_2])$$

$[\text{HI}]$, $[\text{H}_2]$ and $[\text{I}_2]$ are concentrations of HI, H_2 and I_2 , respectively.

In other words, the numerator of the expression for K_{eq} has the concentrations of every product, while the denominator of the expression for K_{eq} has the concentrations of every reactant, leading to the common ***products over reactants*** definition for the K_{eq} .

However, the law of mass action is valid only for concerted one-step (elementary) reactions that proceed through a single transition state and **is not valid in general** because rate equations do not, in general, follow the stoichiometry of the reaction. It means that there are usually several steps in each reaction. Maximal number of particles which may appear in the same point of space, collide with each other simultaneously and so react is equal to three. So, we can be sure that the maximal number of reacting particles (atoms, molecules or ions) in each step of the reaction is equal to three.

Adding a catalyst will affect both the forward reaction and the reverse reaction in the same way and will not have an effect on the equilibrium constant. The positive catalyst will speed up both reactions thereby increasing the speed at which equilibrium is reached.

13.3 LE CHATELIER'S PRINCIPLE

Chemical equilibrium can be shifted by changing the conditions that the system experiences. When we stress the equilibrium, the chemical reaction is no longer at equilibrium, and the reaction starts to move back toward equilibrium in such a way as to decrease the stress. The formal statement is called Le Chatelier's principle: "If an equilibrium is stressed, then the reaction shifts to reduce the stress".

There are several ways to stress an equilibrium. One way is to add or remove a product or a reactant in a chemical reaction at equilibrium. When additional reactant is added, the equilibrium shifts to reduce this stress: it makes more product. When additional product is added, the equilibrium shifts to reactants to reduce the stress. If reactant or product is removed, the equilibrium shifts to make more reactant or product, respectively, to make up for the loss.

When reactants or products are added or removed, *the value of the K_{eq} does not change*. The chemical reaction simply shifts, in a predictable fashion, to reestablish concentrations so that the K_{eq} expression reverts to the only one possible value in given conditions.

Pressure changes do not markedly affect the solid or liquid phases. However, pressure strongly impacts the gas phase. Le Chatelier's principle implies that a pressure increase shifts an equilibrium to the side of the reaction with the fewer number of moles of gas, while a pressure decrease shifts an equilibrium to the side of the reaction with the greater number of moles of gas. If the number of moles of gas is the same on both sides of the reaction, pressure has no effect.

Because temperature is a measure of the energy of the system, increasing temperature can be thought of as adding energy. The reaction will react as if a reactant or a product is being added and will act accordingly by shifting to the other side. For example, if the temperature is increased for an endothermic reaction, essentially a reactant is being added, so the equilibrium shifts toward products. Decreasing the temperature is equivalent to decreasing a reactant (for endothermic reactions) or a product (for exothermic reactions), and the equilibrium shifts accordingly.

EXERCISES FOR CLASSWORK

1. Write how different stresses will affect the equilibrium of the given system.



NOCl concentration increase: _____

NOCl concentration decrease: _____

NO concentration increase: _____

NO concentration decrease: _____

Cl₂ concentration increase: _____

Cl₂ concentration decrease: _____

Increase in pressure: _____

Decrease in pressure: _____

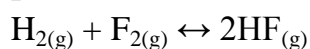
Increase in volume: _____

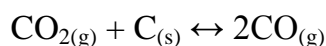
Decrease in volume: _____

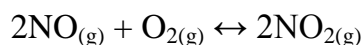
Increase in temperature: _____

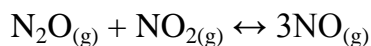
Decrease in temperature: _____

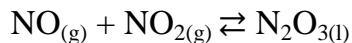
2. Write the expression of the constant of chemical equilibrium for the following processes.





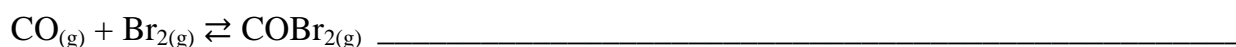
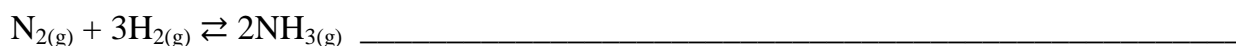




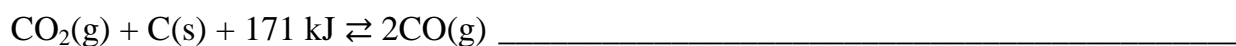


3. Write the equilibrium equation between elemental hydrogen and elemental chlorine as reactants and hydrogen chloride as the product (all the substances are gases).

4. What is the effect on this equilibrium if pressure is increased?

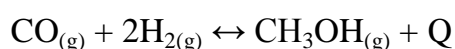


5. Predict the effect of decreasing the temperature on this equilibrium.



EXERCISES FOR HOMEWORK

1. Write how different stresses will affect the equilibrium of the given system.



CO concentration increase: _____

CO concentration decrease: _____

H₂ concentration increase: _____

H₂ concentration decrease: _____

CH₃OH concentration increase: _____

CH₃OH concentration decrease: _____

Increase in pressure: _____

Decrease in pressure: _____

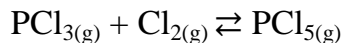
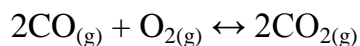
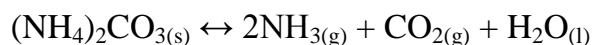
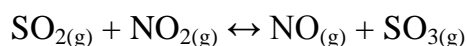
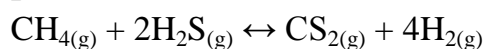
Increase in volume: _____

Decrease in volume: _____

Increase in temperature: _____

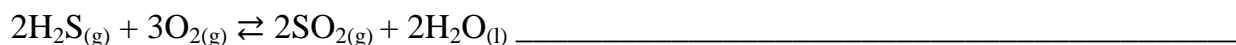
Decrease in temperature: _____

- 2. Write the expression of the constant of chemical equilibrium for the following processes.**

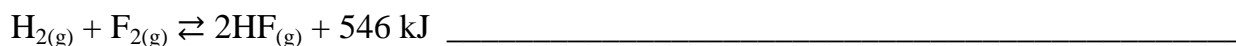
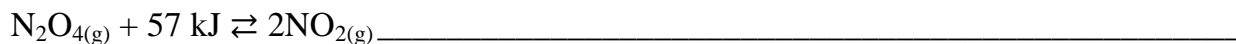


- 3. Write the equilibrium equation between iron (II) carbonate (solid substance) as the reactant and iron (II) oxide (solid substance) and carbon dioxide (gas) as the products.**
- _____

- 4. What is the effect on this equilibrium if pressure is decreased?**



- 5. Predict the effect of decreasing the temperature on this equilibrium.**



LESSON 14

14.1 THE RATE OF CHEMICAL REACTION

The **reaction rate** (rate of reaction) or **speed of reaction** for a reactant or product in a particular reaction is intuitively defined as how fast or slow a reaction takes place. The rate of chemical reaction may be measured in mol/L·sec or mol/L·min.

In other words the rate of chemical reaction (r) is the difference in concentration of the given reactant or product (ΔC , measured in mol/L) during the given time (τ , measured in seconds or minutes).

$$r = \Delta C / \tau$$

14.2 FACTORS INFLUENCING RATE OF REACTION

The nature of the reaction. Some reactions are naturally faster than others. The number of reacting species, their physical state (the particles that form solids move much more slowly than those of gases or those in solution), the complexity of the reaction and other factors can greatly influence the rate of a reaction.

Concentration. Reaction rate increases with concentration, as described by the rate law and explained by collision theory. As reactant concentration increases, the frequency of collisions for its particles increases, and so the rate of the reaction grows.

Pressure. The rate of gaseous reactions increases with pressure, which is, in fact, equivalent to an increase in concentration of the gas. The rate of both forward and backward reactions increases with the increase in pressure. However, for one of these processes the rate may grow steeper than for an opposite one. That is why the increase in pressure may have different effects (or no effect at all) on the chemical equilibrium. For condensed-phase reactions, the pressure dependence is weak.

Temperature. With the increase of temperature more of the colliding particles will have the necessary activation energy resulting in more successful collisions (when bonds are formed between reactants). Moreover, the frequency of collisions themselves is also growing with the temperature.

A catalyst. The presence of a catalyst increases the reaction rate (in both the forward and reverse reactions) by providing an alternative pathway with a lower activation energy.

Surface area. In reactions on surfaces, which take place for example during heterogeneous catalysis, the rate of reaction increases as the surface area does. That is because more particles of the solid are exposed and can be hit by reactant molecules.

In 1864, Peter Waage and Cato Guldberg pioneered the development of chemical kinetics by formulating *the law of mass action*, which states that the speed of a chemical reaction is proportional to the concentration of the reacting substances.

For a chemical reaction $n A + m B \rightarrow x C + y D$, the kinetic (rate) equation or rate law is a mathematical expression used in chemical kinetics to link the rate of a reaction to the concentration of each reactant. It is of the kind:

$$r = k_f[A]^{n'}[B]^{m'}$$

In this equation k_f is the *reaction rate coefficient* or *rate constant* of the forward reaction, although it is not really a constant, because it includes all the parameters that affect reaction rate, except for concentration, which is explicitly taken into account. Of all the parameters described before, temperature is normally the most important one.

The exponents n' and m' are called reaction orders and depend on the reaction mechanism. In the original Waage and Guldberg equation those exponents were nothing but the stoichiometric coefficients from the reaction. Original *law of mass action* works well only in case if there is only a single act in the whole chemical reaction.

As one can understand, the rate of the backward reaction is calculated in similar way.

$$r = k_b[C]^{x'}[D]^{y'}$$

Rates of forward and backward reactions are equal to each other in the state of equilibrium.

Exercise. Express the equilibrium constant (that is the ratio between k_f and k_b) using two equations for the rates of forward and backward reactions calculation (they are given above).

14.3 TEMPERATURE COEFFICIENT OF CHEMICAL REACTION

The Q_{10} **temperature coefficient** is a measure of the rate of change of a biological or chemical system as a consequence of increasing the temperature by 10 °C.

The Q_{10} is calculated as:

$$Q_{10} = \left(\frac{R_2}{R_1}\right)^{10/(T_2-T_1)}$$

where R_1 is the rate of reaction at the temperature 1 (T_1); R_2 is the rate of reaction at the temperature 2 (T_2).

Temperature in this case may be measured either in Celsius degrees or in Kelvins.

The same equation may be re-written in the following way.

$$R_2 = R_1 \cdot Q_{10}^{(T_2-T_1)/10}$$

Using this equation one may calculate the rates of reaction at temperature 2 knowing the rates of that reaction at temperature 1 and the Q_{10} coefficient.

If a question wants you to calculate how many times the rate has changed, you may modify the previous equation.

$$R_2/R_1 = Q_{10}^{(T_2-T_1)/10}$$

Q_{10} is a dimensionless quantity, as it is the factor by which a rate changes, and is a useful way to express the temperature dependence of a process.

For most biological systems, the Q_{10} value is ~ 2 to 3 . However, the “rule of thumb” that the rate of any chemical reactions doubles for every 10°C temperature rise is a common misconception. This has been generalized from the special cases of biological systems.

TEST FOR CLASSWORK

- Which actions can shift the equilibrium of the following process towards reactants?
 $2\text{H}_2\text{S}_{(g)} + 3\text{O}_2_{(g)} \leftrightarrow 2\text{SO}_2_{(g)} + 2\text{H}_2\text{O}_{(g)}$
 - pressure increase
 - addition of O_2
 - addition of SO_2
 - volume decrease
- The increase of pressure will shift the equilibrium of $\text{N}_2_{(g)} + \text{O}_2_{(g)} \leftrightarrow 2\text{NO}_{(g)}$ reaction:
 - towards reactants
 - towards products
 - it will not affect the equilibrium
- The decrease of the volume of the gas container will shift the equilibrium of $2\text{NO}_{(g)} + \text{O}_2_{(g)} \leftrightarrow 2\text{NO}_2_{(g)}$ process:
 - towards reactants
 - towards products
 - it will not affect the equilibrium
- How many times the velocity of $\text{CO}_2_{(g)} + \text{CaO}_{(s)} \leftrightarrow \text{CaCO}_3_{(s)}$ will grow in case of 3 times increase in CO_2 concentration?
 - 2
 - 3
 - 4
 - 9
- Indicate the change of the velocity of $2\text{CO}_{(g)} + \text{O}_2_{(g)} \leftrightarrow 2\text{CO}_2_{(g)}$ reaction in case of 3 times increase in CO concentration:
 - 3 times increase
 - 9 times increase
 - 3 times decrease
 - 9 times decrease
- The velocity of endothermic reaction increases in case of:
 - increase in temperature
 - increase in pressure
 - decrease in temperature
 - decrease in pressure
- The velocity of the reaction has become 4 times higher due to the growth of temperature from 30 to 50°C . Find out the Q_{10} coefficient.
 - 2
 - 3
 - 4
 - 5
- How the velocity of the forward reaction $\text{C}_2\text{H}_2_{(g)} + 2\text{H}_2_{(g)} \rightarrow \text{C}_2\text{H}_6_{(g)}$ will change in case of 2 times decrease in reactants concentration:
 - decrease 2 times
 - increase 4 times
 - increase 16 times
 - decrease 8 times
- How the velocity of the forward reaction $\text{N}_2_{(g)} + 6\text{Li}_{(s)} \rightarrow 2\text{Li}_3\text{N}_{(s)}$ will change in case of 3 times increase in pressure:
 - increase 3 times
 - decrease 3 times
 - increase 2187 times
 - decrease 2187 times

10. How the velocity of the forward reaction $2P_{(s)} + 3Cl_{2(g)} \rightarrow 2PCl_{3(g)}$ will change in case of 3 times increase of the volume of the gas container?
- a. increase 3 times c. increase 27 times
b. decrease 3 times d. decrease 27 times

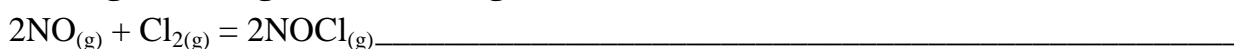
TEST FOR HOMEWORK

1. Which actions can shift the equilibrium of the following process towards products?
 $2H_2S_{(g)} + O_{2(g)} \leftrightarrow 2S_{(s)} + 2H_2O_{(g)}$
- a. pressure decrease c. addition of H_2S
b. addition of O_2 d. volume decrease
2. The decrease of pressure will shift the equilibrium of $2N_2O_{(g)} + O_{2(g)} \leftrightarrow 4NO_{(g)}$ reaction:
- a. towards reactants
b. towards products
c. it will not affect the equilibrium
3. The decrease of the volume of the gas container will shift the equilibrium of $P_{4(l)} + 6Cl_{2(g)} \leftrightarrow 4PCl_{3(l)}$ process:
- a. towards reactants
b. towards products
c. it will not affect the equilibrium
4. How many times the velocity of $PCl_{3(g)} + Cl_{2(g)} \leftrightarrow PCl_{5(g)}$ will grow in case of 3 times increase in Cl_2 concentration?
- a. 2 b. 3 c. 4 d. 9
5. Indicate the change of the velocity of $2P_{(s)} + 5O_{2(g)} \leftrightarrow 2P_2O_{5(s)}$ reaction in case of 2 times decrease in O_2 concentration:
- a. 4 times increase c. 32 times increase
b. 4 times decrease d. 32 times decrease
6. The velocity of exothermic reaction increases in case of:
- a. increase in temperature c. decrease in temperature
b. increase in pressure d. decrease in pressure
7. The velocity of the reaction has become 9 times higher due to the growth of temperature from 37 to 57 °C. Find out the Q_{10} coefficient.
- a. 2 b. 3 c. 4 d. 5
8. How the velocity of the forward reaction $2SO_{2(g)} + O_{2(g)} \rightarrow 2SO_{3(g)}$ will change in case of 3 times increase in reactants concentration:
- a. decrease 8 times c. increase 27 times
b. increase 9 times d. decrease 4 times

9. How the velocity of the forward reaction $F_{2(g)} + H_{2(g)} \rightarrow 2HF_{(g)}$ will change in case of 2 times increase in pressure:
- a. increase 2 times c. increase 4 times
b. decrease 2 times d. decrease 4 times
10. How the velocity of the forward reaction $CO_{(g)} + Cl_{2(g)} \rightarrow COCl_{2(g)}$ will change in case of 2 times increase of the volume of the gas container?
- a. increase 4 times c. increase 2 times
b. decrease 4 times d. decrease 2 times

EXERCISES FOR CLASSWORK

1. Write down the kinetic equations for the following reactions according to the original Waage and Guldberg law of mass action:



2. How the rate of the reaction $N_{2(g)} + 3H_{2(g)} = 2NH_{3(g)}$ will change in case of:
the two times increase of reactants concentration

the three times increase of the pressure

the four times increase of the volume of gas mixture

3. How will the rates of chemical reaction change in case of the temperature increase from 40 °C to 100 °C? The rate of that reaction become two times faster with the increase of temperature equal to 10 °C.

4. The Q_{10} coefficient is equal to 4. How will the rate of this reaction change in case of:
20 °C temperature increase

30° temperature decrease

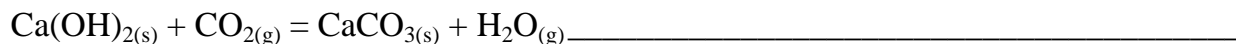
5. The Q_{10} coefficient for a certain reaction is equal to 3. At 20 °C that reaction lasts for 2 minutes. How long that reaction lasts:

at 40 °C

at 0 °C

EXERCISES FOR HOMEWORK

1. Write down the kinetic equations for the following reactions according to the original Waage and Guldberg law of mass action:



2. How the rates of the reaction $2\text{SO}_{2(g)} + \text{O}_{2(g)} = 2\text{SO}_{3(l)}$ will change in case of:
the two times decrease of reactants concentration

the three times decrease of the pressure

the four times decrease of the volume of gas mixture

3. How will the rates of chemical reaction change in case of the temperature decrease from 40 °C to 10 °C? The rate of that reaction become three times slower with the decrease of temperature equal to 10 °C.

4. The Q_{10} coefficient is equal to 3. How will the rate of this reaction change in case of:

50 °C temperature increase

20° temperature decrease

5. The Q_{10} coefficient for a certain reaction is equal to 3. At 20 °C that reaction lasts for 2 minutes. How long that reaction lasts:

at 40 °C

at 0 °C

LESSON 15

SAMPLE TICKET FOR CONTROL TASK #2

ON THE PERIODIC LAW AND CHEMICAL KINETICS

1. Determine the oxidation state for each atom in the following compounds:



2. Write down complete and short electron configurations for:

Na: _____

S: _____

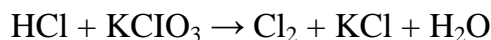
P: _____

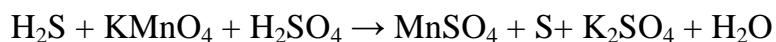
Cu: _____

Ca: _____

Cr: _____

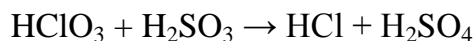
3. Balance chemical reactions using electron balance method:









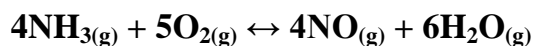




4. Q_{10} coefficient for a given chemical reaction is equal to 3. How will the rates of this reaction change in case of 40 °C temperature increase?

5. Q_{10} coefficient for a given chemical reaction is equal to 3. At 20 °C reaction lasts for 7 minutes. Determine how long that reaction lasts at 60 °C.

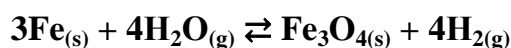
6. Given this equilibrium, predict the direction of shift for each stress.



removal of NO: _____

decrease of the volume of a gas container: _____

the increase of the pressure: _____



removal of H_2 : _____

addition of H_2O : _____

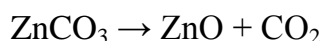
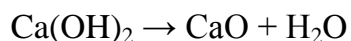
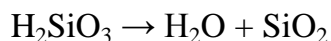
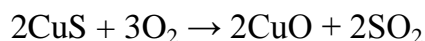
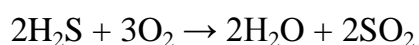
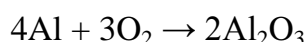
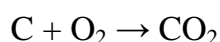
decrease in pressure: _____

LESSON 16

16.1 OXIDES

An **oxide** is a binary chemical compound that contains oxygen in the oxidation state of -2 and other chemical element. Different oxides of the same element are distinguished by Roman numerals denoting their oxidation number (rather than valence): iron (II) oxide (FeO) versus iron (III) oxide (Fe_2O_3).

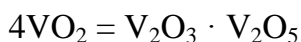
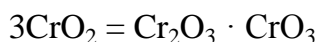
Oxides may be produced either by the way of combustion (of both pure chemical elements and compounds), or by decomposition of certain acids, bases and salts.



Due to its electronegativity, oxygen forms stable chemical bonds with almost all elements to give the corresponding oxides. Noble metals (such as gold or platinum) are prized because they resist direct chemical combination with oxygen, and substances like gold (III) oxide must be generated by indirect routes.

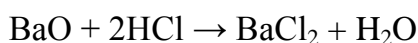
Oxides of most metals adopt polymeric structures with M-O-M crosslinks. Because these crosslinks are strong, the solids tend to be insoluble in solvents, though they are

attacked by acids and bases. The formulas are often deceptively simple, while many of them are nonstoichiometric compounds.

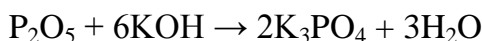


Metal oxides are substances of ionic crystal structure, while the most of nonmetal oxides are molecules. For example, carbon dioxide (CO_2) and carbon monoxide (CO) are molecular oxides. However, silicon oxide (SiO_2) is a substance with atomic crystal structure. Phosphorus pentoxide is a complex molecular oxide with a deceptive name, the formula being P_4O_{10} ($\text{P}_2\text{O}_5 \cdot \text{P}_2\text{O}_5$). Some polymeric oxides when heated depolymerize to give molecules, examples being selenium dioxide (SeO_2) and sulfur trioxide (SO_3).

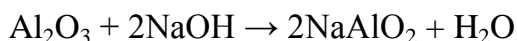
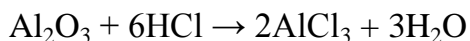
Oxides can be attacked by acids and bases. Those attacked only by acids are *basic oxides*.



Those attacked only by bases are *acidic oxides*.



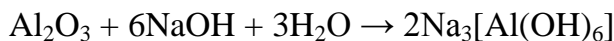
Oxides that react with both acids and bases are *amphoteric*.



In water solution complex salt is formed.



In the excess of concentrated NaOH another complex is formed.



Metals tend to form basic oxides, non-metals tend to form acidic oxides, and amphoteric oxides are formed by elements near the boundary between metals and non-metals (metalloids) and by elements from d-block.

Behavior of oxides for d-elements depends on their oxidation state.

Mostly basic oxides: MnO and CrO

Amphoteric oxides: MnO_2 and Cr_2O_3

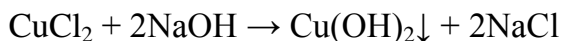
Mostly acidic oxides: Mn_2O_7 and CrO_3

16.2 BASIC AND ACIDIC ANHYDRIDES

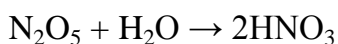
Oxides of more electropositive elements are called "*basic anhydrides*". Exposed to water, oxides of alkali (subgroup IA) and earth-alkali metals (subgroup IIA, except Be and Mg) form basic hydroxides. For example, sodium oxide is basic — when hydrated, it forms sodium hydroxide.



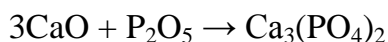
Oxides of other metals cannot react with water, though their insoluble hydroxides can be produced in reactions between their soluble salts and alkali.



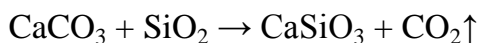
Oxides of more electronegative elements are called “*acid anhydrides*”; adding water, they form oxoacids (oxygen containing acids).



Basic oxides react with acidic oxides and form salts.



Less volatile acidic oxides react with salts of more volatile oxides. Volatility is the tendency of a substance to vaporize. In other words, the more volatile oxide simply flies away and cannot participate in the reverse reaction anymore.



Some oxides do not show behavior as either acid or basic anhydrides (for example: CO, SiO, N₂O, NO). They are known under the name “neutral oxides”. In other words, those oxides do not form salts.

Questions:

- Give examples of basic, acidic and amphoteric oxides.
- How do the properties of oxides change if we go from left to right in a period of the Periodic table?
- How do the properties of oxides change if we go from top to bottom in an A subgroup of the Periodic table?
- What is anhydride?
- What is volatility?

EXERCISES FOR CLASSWORK

1. Write down formulas of the following compounds:

sulfur (IV) oxide _____

iron (II) oxide _____

iron (III) oxide _____

manganese (II) oxide _____

manganese (IV) oxide _____

manganese (VII) oxide _____

2. Write the formulas of oxides corresponding to the following hydroxides:

Mg(OH)₂ _____

LiOH _____

Fe(OH)₂ _____

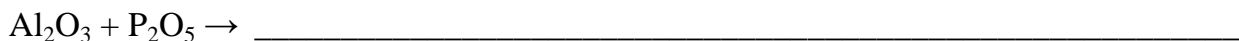
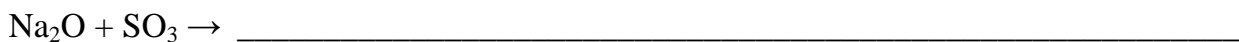
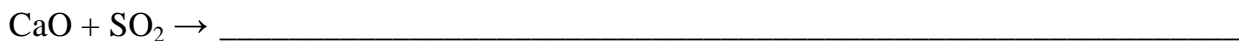
3. Write 5 samples of absolutely basic oxides:

4. Write 5 samples of absolutely acidic oxides:

5. Write 5 samples of amphoteric oxides:

6. Arrange these oxides (MnO / Mn_2O_7 / MnO_2 / MnO_3 / Mn_2O_3) in the order of the increase of their acidic properties:

7. Finish chemical reactions and balance them:



calcium oxide + sulfur (VI) oxide

iron (III) oxide + phosphorus (V) oxide

zinc oxide + nitrogen (V) oxide

sodium oxide + carbon dioxide

calcium oxide + sulfuric acid

iron (II) oxide + nitric acid

8. Calculate the mass of a salt produced in the reaction between 10 g of calcium oxide and 4 L of carbon dioxide (in normal conditions).

9. Find the mass of a salt formed in the reaction between 5 L of sulfur (IV) oxide (in normal conditions) and 2.3 g of barium oxide.

10. Calculate the mass of iron in 500 g of:



Fe_2O_3 _____

Fe_3O_4 _____

EXERCISES FOR HOMEWORK

1. Write down formulas of the following compounds:

potassium oxide _____

sodium oxide _____

phosphorus (III) oxide _____

phosphorus (V) oxide _____

nitrogen (III) oxide _____

nitrogen (V) oxide _____

2. Write the formulas of oxides corresponding to the following hydroxides:

$\text{Fe}(\text{OH})_3$ _____

$\text{Cu}(\text{OH})_2$ _____

$\text{Cr}(\text{OH})_3$ _____

3. Write 5 samples of basic anhydrides:

4. Write 5 samples of acidic anhydrides:

5. Write 5 samples of neutral oxides:

6. Arrange these oxides (CrO / CrO_3 / Cr_2O_3) in the order of the increase of their basic properties:

7. Finish chemical reactions and balance them:

$\text{BaO} + \text{CO}_2 \rightarrow$ _____

$\text{K}_2\text{O} + \text{SiO}_2 \rightarrow$ _____

$\text{MgO} + \text{P}_2\text{O}_5 \rightarrow$ _____

$\text{Na}_2\text{O} + \text{HNO}_3 \rightarrow$ _____

$\text{LiOH} + \text{N}_2\text{O}_3 \rightarrow$ _____

$\text{CaO} + \text{H}_2\text{O} \rightarrow$ _____

$\text{SO}_2 + \text{H}_2\text{O} \rightarrow$ _____

chromium (III) oxide + sulfuric acid

aluminum oxide + phosphoric acid

sodium hydroxide + carbon dioxide

potassium hydroxide + nitrogen (III) oxide

barium hydroxide + hydrochloric acid

calcium oxide + hydrogen sulfide

-
- 8. Calculate the mass of a salt produced in the reaction between 8 g of strontium oxide and 9 g of silicon (IV) oxide.**

-
- 9. Find the mass of a salt formed in the reaction between 3 g of sulfur (VI) oxide and 1.4 g of lithium oxide.**

-
- 10. Calculate the mass of oxygen in 5 L of:**

NO _____

NO₂ _____

N₂O _____

LESSON 17

17.1 BASES

The Brønsted-Lowry theory defines bases as proton (hydrogen ion) acceptors, while the more general Lewis theory defines bases as electron pair donors, allowing other Lewis acids than protons to be included. The oldest Arrhenius theory defines bases as substances which produce just a single type of anions — hydroxide anions (OH^-) — in water solutions. By altering the autoionization equilibrium ($\text{H}_2\text{O} = \text{H}^+ + \text{OH}^-$), bases give solutions with a hydrogen ion activity lower than that of pure water ($\text{pH} > 7.0$).

Bases can be thought of as the chemical opposite of acids. A reaction between an acid and base is called neutralization — aqueous solutions of bases react with aqueous solutions of acids to produce water and salts.

A strong base is a basic chemical compound that is able to deprotonate very weak acids in an acid-base reaction. Common examples of strong bases are the hydroxides of alkali metals and alkaline earth metals like NaOH and Ca(OH)₂. Here is a list of strong bases:

- Lithium hydroxide (LiOH)
- Sodium hydroxide (NaOH)
- Potassium hydroxide (KOH)
- Rubidium hydroxide (RbOH)
- Cesium hydroxide (CsOH)
- Calcium hydroxide (Ca(OH)₂)
- Strontium hydroxide (Sr(OH)₂)
- Barium hydroxide (Ba(OH)₂)

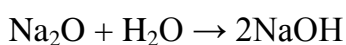
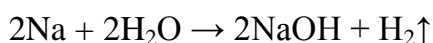
17.2 ALKALIS

Alkali is a strong base that dissolves in water.

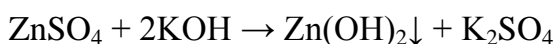
The word “alkali” is derived from Arabic *al qalīy* (or *alkali*), meaning *the calcined ashes*, referring to the original source of alkaline substances. A water-extract of burned plant ashes, called *potash* and composed mostly of potassium carbonate (K₂CO₃), was mildly basic (K₂CO₃ + H₂O = KHCO₃ + KOH). After heating this substance with calcium hydroxide (*slaked lime*), a far more strongly basic substance known as *caustic potash* (potassium hydroxide) was produced (K₂CO₃ + Ca(OH)₂ → CaCO₃↓ + 2KOH). Caustic potash was traditionally used in conjunction with animal fats to produce soft soaps, one of the caustic processes that rendered soaps from fats in the process of saponification, known since antiquity. Plant potash lent the name to the element potassium, which was first derived from caustic potash, and also gave potassium its chemical symbol **K** (Kalium), which ultimately derives from **alkali**.

Soluble hydroxides of alkali metals and alkaline earth metals (i. e. alkalis) are often called “alkali salts”. There are commonly used names of some alkali salts. “Caustic soda” is sodium hydroxide (NaOH). “Caustic potash” is potassium hydroxide (KOH). “Limewater” is calcium hydroxide (Ca(OH)₂). Lime is a general term for calcium-containing inorganic materials, in which carbonates, oxides and hydroxides predominate. The word “lime” originates with its earliest use as building mortar and has the sense of “sticking or adhering”.

Alkali can be produced in the reaction between active metals (or their oxides) and water. Active metals are: Cs, Rb, K, Na, Li, Ba, Sr, Ca. Magnesium, aluminum and zinc can react with water, but the reaction is usually *very slow* unless the metal samples are specially prepared to remove the surface layer of oxide which protects the rest of the metal.



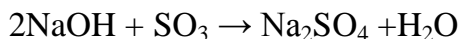
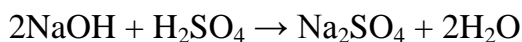
Insoluble bases can be produced in reactions between soluble salts and alkalis.



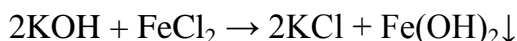
In case of the excess of alkali complex soluble salts can be formed.



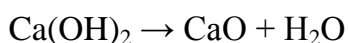
Chemical properties. Alkalis react with both acids and acidic oxides. The products of those reactions are salts and water.



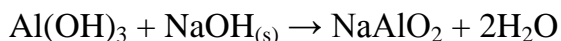
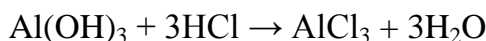
Alkalis react with certain salts (in case if one of the products of the reaction is insoluble or volatile).



All bases, except NaOH and KOH, are decomposed at high temperatures before the melting. The products of such decomposition are oxide and water.



Amphoteric hydroxides react with both acids and alkalis.



The last reaction is possible in the absence of water at high temperature (during the melting of solid aluminum and sodium hydroxides together). In water solution complex salt is formed.



Questions:

- Give a definition of base.
- Give a definition of alkali.
- Give a definition of hydroxide.
- What metals are active?

EXERCISES FOR CLASSWORK

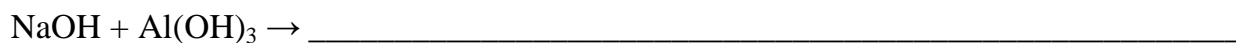
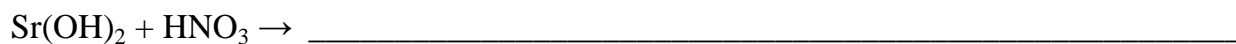
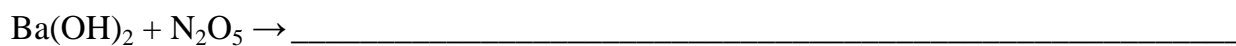
1. Write 5 samples of absolutely basic hydroxides:

2. Write 5 samples of alkalis:

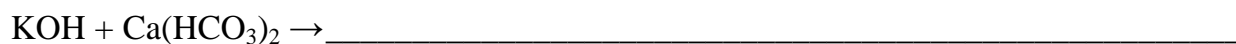
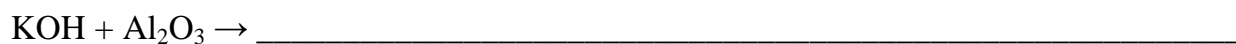
3. Write 5 samples of amphoteric hydroxides:

4. Finish chemical reactions and balance them:





- 5. What substances react with KOH? Write down equations of possible reactions.**



- 6. How can Zn(OH)_2 be produced from the set of the following substances: Na; H_2SO_4 ; ZnO ; H_2O ? Write chemical reactions.**

- 7. Calculate the mass of an alkali produced in the reaction between 4 g of sodium and 50 g of water.**

- 8. Find the maximal mass of oxide that can be produced due to thermal decomposition of 200 g of magnesium hydroxide.**

- 9. What is the mass of a salt formed in the reaction between 5 g of solid potassium hydroxide and 8 g of aluminum oxide at high temperature?**

10. Calculate the absolute and relative decrease of mass of a solid substance after the complete thermal decomposition of 215 g of calcium hydroxide.

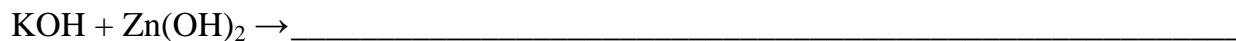
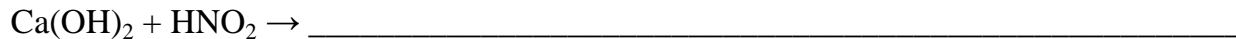
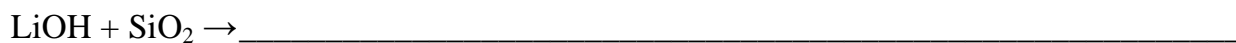
EXERCISES FOR HOMEWORK

1. Write 5 samples of soluble metal hydroxides:

2. Write 5 samples of metal hydroxides which can be thermally decomposed:

3. Write 5 samples of diacidic metal hydroxides:

4. Finish chemical reactions and balance them:



5. What substances react with $\text{Zn}(\text{OH})_2$? Write down equations of possible reactions.



6. How can $\text{Fe}(\text{OH})_3$ be produced from the set of the following substances: K; HCl; Fe_2O_3 ; H_2O ? Write chemical reactions.

7. Calculate the mass of an alkali produced in the reaction between 7 g of calcium oxide and 20 ml of water (density of water is 1 g/ml).

8. Find the maximal mass of water that can be produced due to thermal decomposition of 300 g of lithium hydroxide.

9. What is the mass of a salt formed in the reaction between 4 g of solid sodium hydroxide and 7 g of zinc oxide at high temperature?

10. Calculate the absolute and relative decrease of mass of a solid substance after the complete thermal decomposition of 440 g of magnesium hydroxide.

LESSON 18

18.1 ACIDS

An Arrhenius acid is a compound that gives just one type of a cation — H^+ cation — during dissociation in water solution. The H^+ ion is just a bare proton, and it is rather clear that bare protons are not floating around in an aqueous solution. Instead, chemistry has defined the hydronium ion (H_3O^+) as the actual chemical species that represents an H^+ ion. Classic Arrhenius acids are molecules that are ionized in water solution producing H^+ cations.

If an acid is composed of only hydrogen and one other element, the name is *hydro-* + the stem of the other element + *-ic acid*. For example, the compound $\text{HCl}(\text{aq})$ is hydrochloric acid, while $\text{H}_2\text{S}(\text{aq})$ is hydrosulfuric acid. If these acids were not dissolved in water, the compounds would be called hydrogen chloride and hydrogen sulfide,

respectively. Both of these substances are well known as molecular compounds; when dissolved in water, however, they are treated as acids.

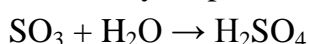
Table 18.1

Formulas and names of some acids

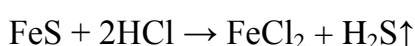
Formula	Name
CH ₃ COOH	acetic acid
HCl	hydrochloric acid
HClO ₃	chloric acid
HClO ₄	perchloric acid
HBr	hydrobromic acid
HI	hydroiodic acid
HF	hydrofluoric acid
HNO ₂	nitrous acid
HNO ₃	nitric acid
H ₂ C ₂ O ₄	oxalic acid
H ₃ PO ₄	phosphoric acid
H ₂ SO ₄	sulfuric acid
H ₂ SO ₃	sulfurous acid

If a compound is composed of hydrogen ions and a polyatomic anion, then the name of the acid is derived from the stem of the polyatomic ion's name. Typically, if the anion name ends in *-ate*, the name of the acid is the stem of the anion name plus *-ic acid*; if the related anion's name ends in *-ite*, the name of the corresponding acid is the stem of the anion name plus *-ous acid*.

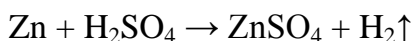
Acids may be produced in reactions between acidic oxides and water.



Less volatile acids react with salts of more volatile acids.



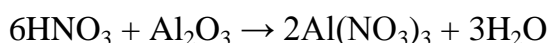
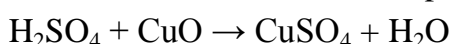
Chemical properties. Acids have some properties in common. They react with metals situated before hydrogen in the electrochemical series of metals to give off H₂ gas.



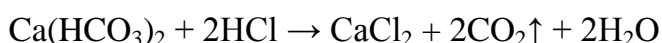
Electrochemical series of metals:

Li > K > Sr > Ca > Na > Mg > Al > Zn > Cr > Fe > Cd > Co > Ni > Sn > > Pb > H > Cu > Ag > Hg > Pt > Au

Acids react with basic and amphoteric oxides.



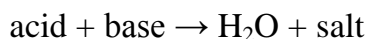
Acids react with salts in case if insoluble substances or gases are formed. For example, they react with carbonate and hydrogen carbonate salts to give off CO₂ gas.



Acids that are ingested typically have a sour, sharp taste. The name *acid* comes from the Latin word *acidus*, meaning “sour”.

18.2 NEUTRALIZATION REACTION

Acids and bases react with each other to make water and an ionic compound called a salt. A salt, in chemistry, is any ionic compound made by combining an acid with a base. A reaction between an acid and a base is called a neutralization reaction and can be represented as follows:



The stoichiometry of the balanced chemical equation depends on the number of H^+ ions in the acid and the number of OH^- ions in the base.

Questions:

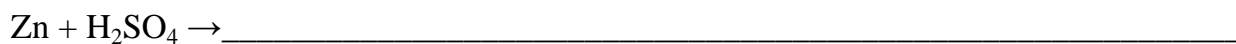
- a. Give the formula for each acid.
 1. perchloric acid
 2. hydroiodic acid
 3. hydrosulfuric acid
 4. phosphorous acid
- b. Name each acid.
 1. HF
 2. HNO_3
 3. $\text{H}_2\text{C}_2\text{O}_4$
 4. H_2SO_4
 5. H_3PO_4
 6. HCl
- c. Name the properties that acids have in common.
- d. What metals among the given list react with hydrochloric acid: Li; Ba; Cu; Mg; Al; Au; Ag?

EXERCISES FOR CLASSWORK

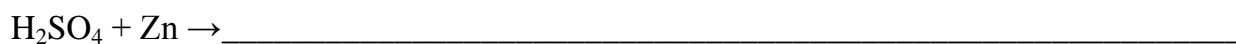
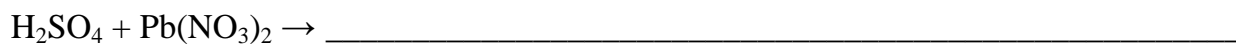
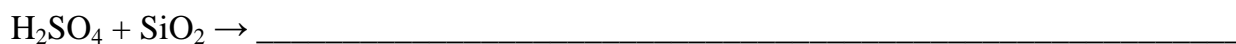
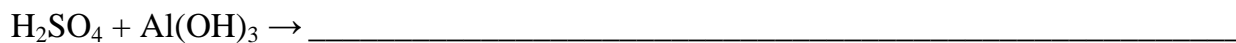
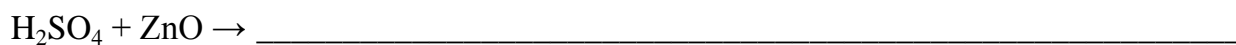
1. Write 7 strong acids:

2. Write 5 samples of monoprotic acids:

3. Finish chemical reactions and balance them:



4. What substances react with sulfuric acid? Write down equations of possible reactions.



5. Calculate the volume of hydrogen (in normal conditions) released in the reaction between 6 g of zinc and 10 g of sulfuric acid.

6. Find the volume of carbon dioxide (in normal conditions) produced in the reaction between 5 g of sodium carbonate and 17 g of hydrochloric acid.

7. What is the mass of a salt formed in the reaction between 13 g of sodium hydroxide and 8 g of hydrobromic acid?

8. Find the mass of sulfuric acid produced from 13 g of sulfur (VI) oxide in its reaction with water (in normal conditions).

9. What is the mass of $ZnSO_4$ produced from 9.8 g of sulfuric acid and 8.1 g of ZnO ?

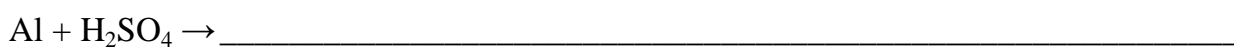
10. Sulfuric acid reacted with zinc and produced 0.25 mol of hydrogen. Calculate the masses of both reactants participated in this reaction.

EXERCISES FOR HOMEWORK

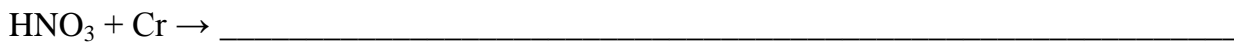
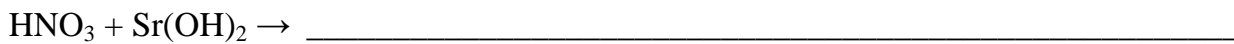
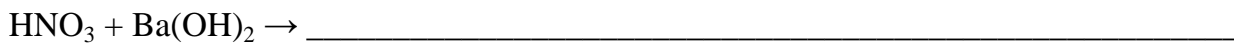
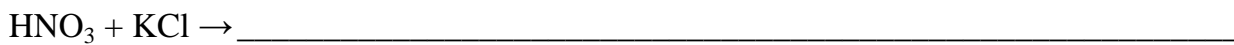
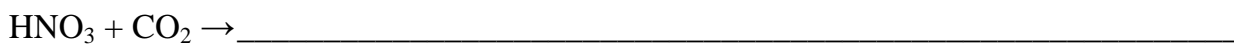
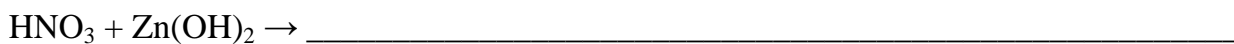
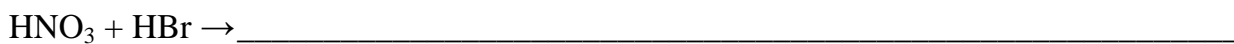
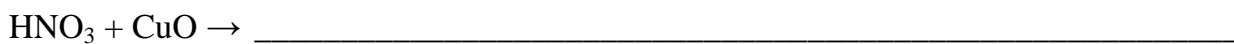
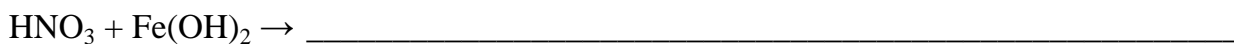
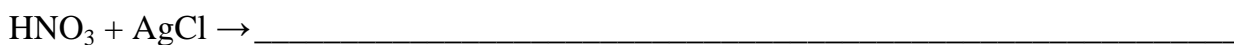
1. Write 5 samples of weak acids:

2. Write 5 samples of diprotic acids:

3. Finish chemical reactions and balance them:



4. What substances react with nitric acid? Write down equations of possible reactions.



5. Calculate the volume of hydrogen (in normal conditions) released in the reaction between 3 g of aluminum and 6 g of hydrochloric acid.

6. Find the mass of a salt produced in the reaction between 4 g of potassium carbonate and 10 g of sulfuric acid.

7. What is the mass of a salt formed in the reaction between 11 g of lithium hydroxide and 28 g of hydroiodic acid?

8. Find the mass of orthophosphoric acid produced from 6 g of phosphorus V oxide in its reaction with water.

9. What is the mass of K_3PO_4 produced from 49 g of phosphoric acid and 380 g of KOH?

10. Hydrochloric acid reacted with iron and produced 0.16 mol of hydrogen. Calculate the masses of both reactants participated in this reaction.

LESSON 19

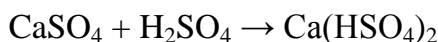
19.1 SALTS

Salts are ionic compounds that result from the neutralization reaction of an acid and a base. They are composed of such numbers of cations (positively charged ions) and anions (negative ions) that the product is electrically neutral (without a net charge). These component ions can be inorganic such as chloride (Cl^-), as well as organic such as acetate (CH_3COO^-) and monoatomic ions such as fluoride (F^-), as well as polyatomic ions such as sulfate (SO_4^{2-}).

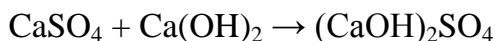
There are several varieties of salts. Salts that dissociate to produce hydroxide ions along with other anions when dissolved in water are *basic salts* ($Fe(OH)_2Cl$) and salts that dissociate to produce hydronium ions along with other cations in water are *acidic salts* ($NaHSO_4$). *Neutral salts* are those that are neither acidic nor basic salts: they are not

producing neither OH^- nor H^+ during dissociation in water. Zwitterions contain an anionic center and a cationic center in the same molecule but are not considered to be salts. Examples include amino acids, many metabolites, peptides, and proteins.

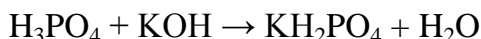
Acidic salts can be produced from neutral salts after the addition of acid.



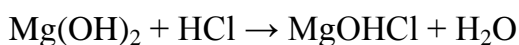
Basic salts can be produced from neutral salts after the addition of base.



Another way to produce an acidic salt is to perform a partial neutralization of a polyprotic acid. Monoprotic acid cannot give an acidic salt.



To produce a basic salt one may perform a partial neutralization of a base containing multiple OH groups. Bases with a single OH group cannot give a basic salt.



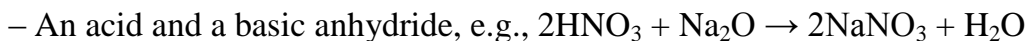
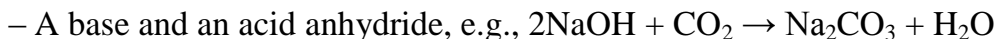
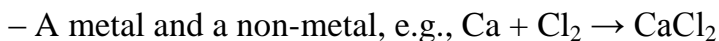
19.2 SOLUBILITY CHART OF SALTS

Many ionic compounds can be dissolved in water. The exact combination of ions involved makes each compound have a unique solubility in any solvent. The solubility is dependent upon how well each ion interacts with the solvent, and how the ions interact with each other, so there are certain patterns.

For example, the most of the salts of sodium, potassium and ammonium are soluble in water, as are all nitrates and many sulfate salts except barium sulfate, calcium sulfate (sparingly soluble) and lead (II) sulfate.

However, ions that bind tightly to each other and form highly stable lattices are less soluble, because it is harder for these structures to break apart for the compounds to dissolve. For example, most carbonate salts are not soluble in water, such as lead carbonate and barium carbonate. Soluble carbonate salts are: sodium carbonate, potassium carbonate and ammonium carbonate.

Salts are formed by chemical reactions between:



– Salts can also be formed if solutions of different salts are mixed with each other or with acid or alkali solutions. Their ions recombine, and in some cases new salt (base or even acid) precipitates (see: the solubility chart below): $\text{Pb}(\text{NO}_3)_2 + \text{Na}_2\text{SO}_4 \rightarrow \text{PbSO}_4\downarrow + 2\text{NaNO}_3$. The same thing can be said about reactions in which gases are formed: $\text{NH}_4\text{Cl} + \text{KOH} \rightarrow \text{KCl} + \text{NH}_3\uparrow + \text{H}_2\text{O}$

– Metal is able to substitute another metal in a salt in case if it is situated before the second one in the reactivity series: $\text{CuSO}_4 + \text{Zn} \rightarrow \text{ZnSO}_4 + \text{Cu}$

Table 19.1

Solubility chart

S soluble (more than 1 g per 100 g of water) **M** marginally soluble (0,1 g - 1 g per 100 g of water) **I** insoluble (less than 0,1 g per 100 g of water) **D** decomposes in water **U** compound doesn't exist or is unstable

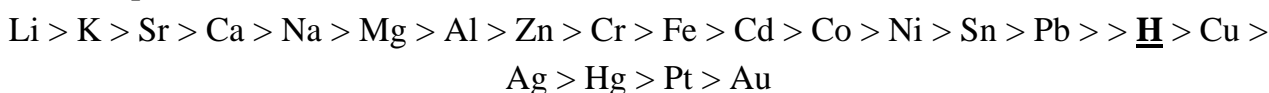
cation anion	H ⁺	Li ⁺	K ⁺	Na ⁺	NH ₄ ⁺	Ba ²⁺	Ca ²⁺	Mg ²⁺	Sr ²⁺	Al ³⁺	Cr ³⁺	Fe ²⁺	Fe ³⁺	Ni ²⁺	Co ²⁺	Mn ²⁺	Zn ²⁺	Ag ⁺	Hg ²⁺	Pb ²⁺	Sn ²⁺	Cu ²⁺	
OH ⁻		S	S	S	S	S	M	I	M	I	I	I	I	I	I	I	I	D	D	I	I	I	
F ⁻	S	S	S	S	S	M	I	I	M	S	I	I	I	S	S	M	S	S	M	I	S	S	
Cl ⁻	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	I	S	M	S	S
Br ⁻	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	I	M	M	S	S
I ⁻	S	S	S	S	S	S	S	S	S	S	U	S	U	S	S	S	S	S	I	I	I	M	S
S ²⁻	S	S	S	S	S	S	D	D	S	D	D	I	D	I	I	I	I	I	I	I	I	I	I
HS ⁻	S	S	S	S	S	S	S	S	S	U	U	U	U	U	I	U	U	U	U	U	U	U	U
SO ₃ ²⁻	S	S	S	S	S	M	M	M	I	U	D	M	U	I	I	U	M	I	I	I	U	U	U
HSO ₃ ⁻	S	U	S	S	S	S	S	S	S	U	U	U	U	U	U	U	U	U	U	U	U	U	U
SO ₄ ²⁻	S	S	S	S	S	I	M	S	I	S	S	S	S	S	S	S	S	M	D	I	S	S	S
HSO ₄ ⁻	S	S	S	S	S	U	U	U	D	U	U	U	U	U	U	U	U	U	U	U	I	U	U
NO ₃ ⁻	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	D	S
NO ₂ ⁻	S	S	S	S	S	S	S	S	S	U	U	U	U	S	M	U	U	M	U	U	U	U	U
PO ₄ ³⁻	S	I	S	S	D	I	I	I	I	I	I	I	I	I	I	I	I	I	I	I	I	I	I
HPO ₄ ²⁻	S	U	S	S	S	I	I	M	I	U	U	I	U	U	U	I	U	U	U	M	I	U	U
H ₂ PO ₄ ⁻	S	S	S	S	S	S	S	S	S	U	U	S	U	U	U	S	S	S	U	D	U	U	U
CO ₃ ²⁻	S	S	S	S	S	I	I	I	I	U	U	I	U	I	I	I	I	I	I	U	I	U	I
HCO ₃ ⁻	S	S	S	S	S	S	S	S	S	U	U	S	U	U	U	U	U	U	U	S	U	U	U
CH ₃ COO ⁻	S	S	S	S	S	S	S	S	S	D	S	S	D	S	S	S	S	S	S	S	D	S	S
SiO ₃ ²⁻	I	S	S	S	U	I	I	I	I	U	U	I	U	U	U	I	I	U	U	I	U	U	U

Table 19.2

Reactivity series of metals

Active metals — those which react with water (at normal t°) and acids											
Cs	Rb	K	Na	Li	Ba	Sr	Ca				
Metals which react with acids and produce salts and H ₂											
Mg	Al	Mn	Zn	Cr	Fe	Cd	Co	Ni	Sn	Pb	
Metals which react with H ₂ SO ₄ conc. and HNO ₃ and don't produce H ₂											
Sb	Bi	Cu	Hg	Ag							
Metals which cannot react even with H ₂ SO ₄ conc. and HNO ₃											
Au	Pt										

The reactivity series is sometimes quoted in the strict reverse order of standard electrode potentials, when it is also known as the “electrochemical series”:



The positions of lithium and sodium are changed on such a series; gold and platinum are also inverted, although this has little practical significance as both metals are highly unreactive. Hydrogen is included in standard electrode potentials order because the power of a reducing agent is measured relatively to the standard hydrogen electrode.

Concentrated sulfuric acid (H_2SO_4) and nitric acid (HNO_3) of all concentrations are not producing hydrogen in their reactions with metals. However, even these two acids cannot react with gold and platinum.

Questions:

- Give a definition of the term "salt".
- What is the difference between normal, acidic and basic salts?
- How can the reactivity series of metals be used?
- How can the solubility chart be used?

TEST FOR CLASSWORK

- Choose basic oxides:
a. SiO_2 b. K_2O c. ZnO d. SrO
- Choose acidic oxides:
a. CO b. CO_2 c. BaO d. SO_2
- Choose amphoteric oxides:
a. Al_2O_3 b. ZnO c. SiO d. Cr_2O_3
- Choose neutral oxides (those which cannot form salts):
a. NO b. SiO_2 c. MgO d. N_2O
- Choose strong acids:
a. HCl b. HBr c. HI d. HF
- Choose strong bases:
a. $\text{Be}(\text{OH})_2$ b. $\text{Sr}(\text{OH})_2$ c. KOH d. LiOH
- Choose acidic salts:
a. NH_4NO_3 b. $\text{NH}_4\text{H}_2\text{PO}_4$ c. KHCO_3 d. $(\text{MgOH})_2\text{CO}_3$
- Choose bases which cannot be formed in the reaction between corresponding oxide and water:
a. $\text{Al}(\text{OH})_3$ b. NaOH c. KOH d. $\text{Zn}(\text{OH})_2$
- Choose salts which can react with the acid containing the same anion:
a. KNO_3 b. Na_2SO_3 c. K_3PO_4 d. KH_2PO_4
- Choose salts which can react with alkali:
a. NaCl b. KBr c. MgCl_2 d. AlPO_4

TEST FOR HOMEWORK

- Choose basic oxides:
a. CaO b. Mn_2O_7 c. CrO_3 d. Cs_2O
- Choose acidic oxides:
a. SO_2 b. BeO c. N_2O_5 d. H_2O_2
- Choose amphoteric oxides:
a. Fe_2O_3 b. MnO_2 c. P_2O_3 d. Na_2O
- Choose neutral oxides (those which cannot form salts):
a. CO b. SiO c. CaO d. SrO
- Choose weak acids:
a. H_2SO_4 b. HNO_3 c. HNO_2 d. H_2SO_3
- Choose weak bases:
a. $Fe(OH)_2$ b. $Cu(OH)_2$ c. NaOH d. NH_4OH
- Choose basic salts:
a. CaOHCl b. $(MgOH)_2SO_4$ c. $K[Al(OH)_4]$ d. $FeCl_3$
- Choose bases which cannot be formed in the reaction between corresponding oxide and water:
a. $Fe(OH)_3$ b. LiOH c. CsOH d. $Fe(OH)_2$
- Choose salts which can react with the acid containing the same anion:
a. KCl b. $NaHCO_3$ c. K_3PO_4 d. K_2HPO_4
- Choose salts which can react with alkali:
a. NH_4Cl b. $ZnBr_2$ c. NaCl d. $BaCl_2$

EXERCISES FOR CLASSWORK

- Write 5 samples of acidic salts:

- Write 5 samples of basic salts:

- Write down the formulas of the following compounds:

sodium sulfate _____

magnesium hydrogen carbonate _____

potassium sulfide _____

potassium hydrogen sulfate _____

aluminum dihydroxy bromide _____

calcium hydrogen sulfide _____

4. Which metals from this line (Na / Fe / Zn / Ag / Ba / Ni) can substitute copper in CuCl_2 in water solution?

5. What substances react with AgNO_3 ? Write down equations of possible reactions.

$\text{AgNO}_3 + \text{HCl} \rightarrow$ _____

$\text{AgNO}_3 + \text{FeSO}_4 \rightarrow$ _____

$\text{AgNO}_3 + \text{CaCl}_2 \rightarrow$ _____

$\text{AgNO}_3 + \text{BaCl}_2 \rightarrow$ _____

$\text{AgNO}_3 + \text{H}_2\text{SO}_4 \rightarrow$ _____

$\text{AgNO}_3 + \text{NaI} \rightarrow$ _____

$\text{AgNO}_3 + \text{KBr} \rightarrow$ _____

$\text{AgNO}_3 + \text{K}_3\text{PO}_4 \rightarrow$ _____

$\text{AgNO}_3 + \text{Na}_2\text{CO}_3 \rightarrow$ _____

$\text{AgNO}_3 + \text{Ba}(\text{NO}_3)_2 \rightarrow$ _____

6. Finish chemical reactions and balance them, notice the coefficients:

$2\text{NaOH} + 1\text{H}_2\text{SO}_3 \rightarrow$ _____

$1\text{NaOH} + 1\text{H}_2\text{SO}_3 \rightarrow$ _____

$1\text{Zn}(\text{OH})_2 + 2\text{HCl} \rightarrow$ _____

$1\text{Zn}(\text{OH})_2 + 1\text{HCl} \rightarrow$ _____

$1\text{KHSiO}_3 + 1\text{KOH} \rightarrow$ _____

$1\text{KHSiO}_3 + 1\text{HCl} \rightarrow$ _____

$\text{Al}(\text{OH})_3 + 1\text{KOH} \rightarrow$ _____

7. What is the mass of silver chloride produced in the reaction between 5.85 g of sodium chloride and 33.8 g of silver nitrate?

8. Determine the mass of a salt produced in the reaction between 10.5 g of calcium oxide and 13.7 g of sulfur (IV) oxide.

9. Determine what kind of salt(s) is formed in the reaction between 2 g of sodium hydroxide and 25 g of phosphoric acid, find its mass.

10. Find what kind of salt(s) is formed in the reaction between 4 g of potassium hydroxide and 3.5 L (normal conditions) of carbon dioxide.

EXERCISES FOR HOMEWORK

1. Write 5 samples of acids which can form acidic salts:

2. Write 5 samples of bases which can form basic salts:

3. Write down the formulas of the following compounds:

iron (III) sulfate _____

barium dihydrogen phosphate _____

magnesium hydroxy chloride _____

potassium hydrogen sulfide _____

iron (III) dihydroxy chloride _____

calcium hydrogen phosphate _____

4. Which metals from this line (K / Cr / Zn / Ag / Sn / Ni) can substitute cobalt in CoCl_2 in water solution?

5. What substances react with Na_2CO_3 ? Write down equations of possible reactions.

$\text{Na}_2\text{CO}_3 + \text{HCl} \rightarrow$ _____

$\text{Na}_2\text{CO}_3 + \text{H}_2\text{SO}_4 \rightarrow$ _____

$\text{Na}_2\text{CO}_3 + \text{CaCl}_2 \rightarrow$ _____

$\text{Na}_2\text{CO}_3 + \text{BaCl}_2 \rightarrow$ _____

$\text{Na}_2\text{CO}_3 + \text{K}_2\text{SO}_4 \rightarrow$ _____

$\text{Na}_2\text{CO}_3 + \text{NaI} \rightarrow$ _____

$\text{Na}_2\text{CO}_3 + \text{CO}_2 + \text{H}_2\text{O} \rightarrow$ _____

$\text{Na}_2\text{CO}_3 + \text{SiO}_2 \rightarrow$ _____

$\text{Na}_2\text{CO}_3 + \text{NaOH} \rightarrow$ _____

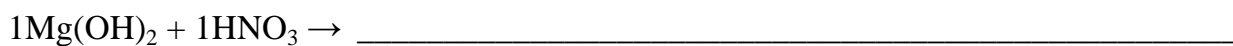
$\text{Na}_2\text{CO}_3 + \text{CH}_3\text{COOH} \rightarrow$ _____

6. Finish chemical reactions and balance them, notice the coefficients:

$2\text{NaOH} + 1\text{CO}_2 \rightarrow$ _____

$1\text{NaOH} + 1\text{CO}_2 \rightarrow$ _____

$1\text{Mg}(\text{OH})_2 + 2\text{HNO}_3 \rightarrow$ _____



7. What is the mass of barium sulfate produced in the reaction between 9.8 g of sulfuric acid and 41.6 g of barium chloride?

8. Determine the masses of the products of the reaction between 12.3 g of sodium hydroxide and 26 g of copper chloride?

9. Determine what kind of salt(s) is formed in the reaction between 3 g of potassium hydroxide and 4 g of phosphoric acid, find its mass.

10. Find what kind of salt(s) is formed in the reaction between 7 g of sodium hydroxide and 1.2 g of zinc hydroxide, find its mass.

LESSON 20

20.1 CLASSIC CHAINS OF CHEMICAL REACTIONS

Chain of chemical reactions is one of the most commonly used types of tasks in chemistry. In the classic type of that kind of task student has to write down all the reactions from each chain. Each next substance must be somehow produced from the previous substance. Student has a kind of freedom to choose additional reactants and conditions to make reactions possible. In case if the one-step reaction is impossible (for example, it is impossible to produce insoluble base directly from the metal), two or even more reactions should be written to complete one step in the chain.

20.2 MODERN CHAINS OF CHEMICAL REACTIONS

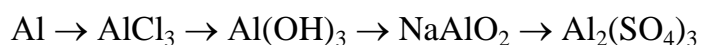
Modern type of the chain of chemical reactions includes two tasks in one. At first, student must write down all the reactions. All the reactants, conditions and by-products are given (if they are not clearly obvious), while main products are hidden behind the letters. At second, student must find out molar mass of the last product, or the sum of molar masses for several substances (for example, for substance hidden behind the letter “B” and substance hidden behind the letter “D”).

EXERCISES FOR CLASSWORK

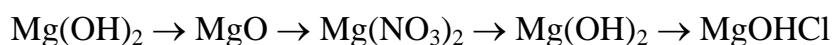
1. Write four reactions according to the following classic chains of chemical reactions and balance them:



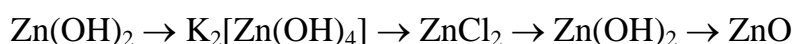
1. _____
2. _____
3. _____
4. _____



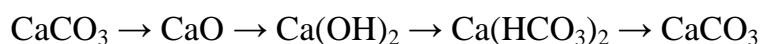
1. _____
2. _____
3. _____
4. _____



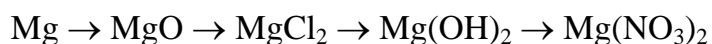
1. _____
2. _____
3. _____
4. _____



1. _____
2. _____
3. _____
4. _____



1. _____
2. _____
3. _____
4. _____



1. _____
2. _____

3. _____

4. _____



1. _____

2. _____

3. _____

4. _____



1. _____

2. _____

3. _____

4. _____



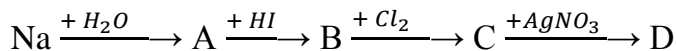
1. _____

2. _____

3. _____

4. _____

2. Write four reactions with sodium containing compounds according to the chain of chemical reactions



1. _____

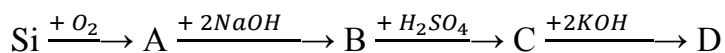
2. _____

3. _____

4. _____

Calculate the sum of molar masses for substances B and D:

3. Write four reactions with silicon containing compounds according to the chain of chemical reactions



1. _____

2. _____

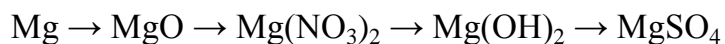
3. _____

4. _____

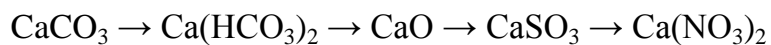
Calculate the sum of molar masses for substances B and D:

EXERCISES FOR HOMEWORK

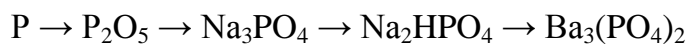
1. Write four reactions according to the following classic chains of chemical reactions and balance them:



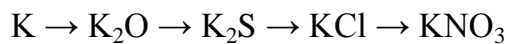
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1. _____
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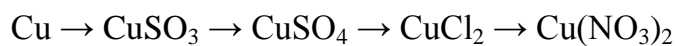
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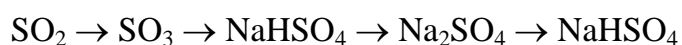
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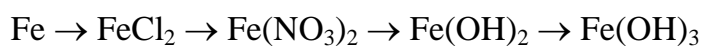
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1. _____
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1. _____

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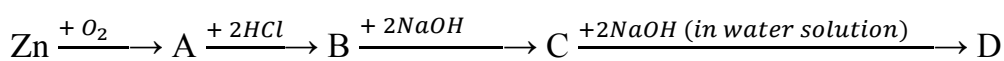
1. _____

2. _____

3. _____

4. _____

2. Write four reactions with zinc containing compounds according to the chain of chemical reactions



1. _____

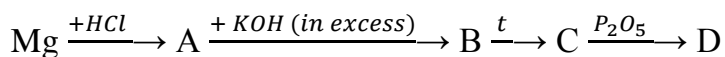
2. _____

3. _____

4. _____

Calculate the sum of molar masses for substances C and D:

3. Write four reactions with magnesium containing compounds according to the chain of chemical reactions



1. _____

2. _____

3. _____

4. _____

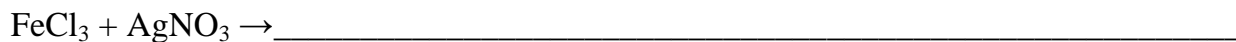
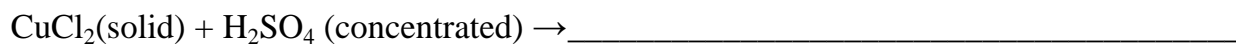
Calculate the sum of molar masses for substances C and D:

LESSON 21

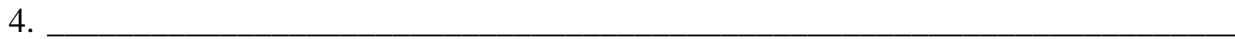
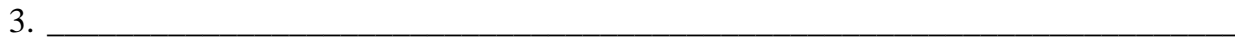
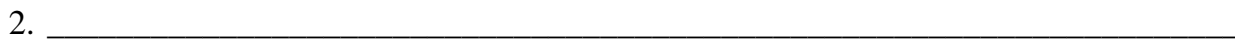
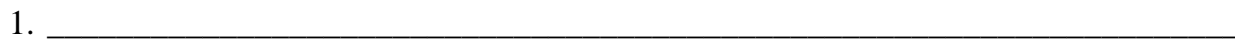
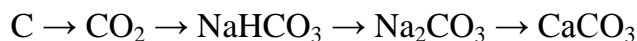
SAMPLE TICKET FOR CONTROL TASK #2

ON MAIN TYPES OF INORGANIC COMPOUNDS

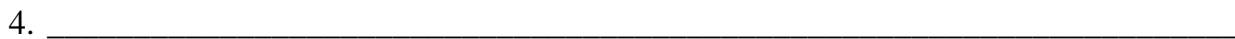
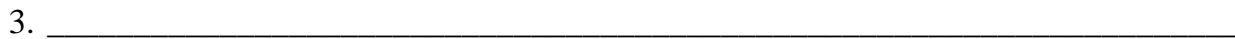
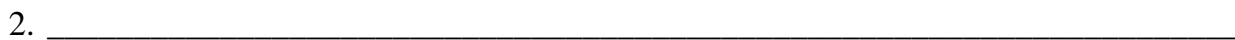
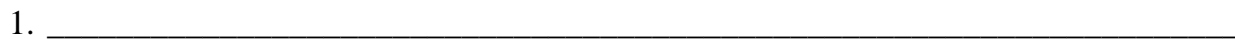
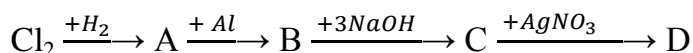
1. Write down equations of chemical reactions between the following substances (in case if they are possible) and balance them:



2. Write four reactions according to the following classic chain of chemical reactions and balance them:



3. Write four reactions with chlorine containing compounds according to the chain of chemical reactions



4. What salt (potassium sulfite, potassium hydrogen sulfite, or the mixture of them) will be formed from 16.8 g of potassium hydroxide and 6.7 L of sulfurous acid anhydride? Calculate the mass of that salt.

LESSON 22

22.1 QUALITATIVE DESCRIPTION OF SOLUTIONS

The **major component** of a solution is called the **solvent**. The **minor component** of a solution is called the **solute**. By major and minor we mean whichever component has the greater presence by mass or by moles. Sometimes this becomes confusing, especially with substances with very different molar masses. For example, if the mass percentage of ethanol solution in water is equal to 70 %, the mole percentage of ethanol is still equal to 47.6 % (because molar mass of water is lower than that for ethanol). Obviously, such expression as “98 % ethanol” widely used for a solution in which there are 98 % of ethanol and just 2 % of water is not correct, because ethanol is still mistakenly considered to be solute and not solvent.

Salt water is a solution of solid NaCl in liquid water; soda water is a solution of gaseous CO₂ in liquid water, while air is a solution of a gaseous solute (O₂) in a gaseous solvent (N₂). In all cases, however, the overall phase of the solution is the same phase as the solvent. So, if we deal with a mixture of a solid substance and a liquid, liquid is always the solvent.

One important concept of solutions is in defining how much solute is dissolved in a given amount of solvent. **Dilute** describes a solution that has very little solute, while **concentrated** describes a solution that has a lot of solute. One problem is that these terms are qualitative; they describe more or less but not exactly how much.

22.2 SOLUBILITY OF IONIC COMPOUNDS

In most cases, only a certain maximum amount of solute can be dissolved in a given amount of solvent. This maximum amount is called the **solubility** of the solute. It is usually expressed in terms of the amount of solute that can dissolve in 100 g of the solvent at a given temperature.

When the maximum amount of solute has been dissolved in a given amount of solvent, we say that the solution is **saturated** with solute. When less than the maximum amount of solute is dissolved in a given amount of solute, the solution is **unsaturated**. A solution of 0.00019 g of AgCl per 100 g of H₂O may be saturated, but with so little solute dissolved, it

is also rather dilute. A solution of 36.1 g of NaCl in 100 g of H₂O is also saturated but rather concentrated.

Table 22.1

Solubilities of Some Ionic Compounds

Solute	Solubility (g per 100 g of H ₂ O at 25 °C)
AgCl	0.00019
CaCO ₃	0.0006
NaCl	36.1
KBr	70.7
NaNO ₃	94.6

In some circumstances, it is possible to dissolve more than the maximum amount of a solute in a solution. Usually, this happens by heating the solvent, dissolving more solute than would normally dissolve at regular temperatures, and letting the solution cool down slowly and carefully. Such solutions are called *supersaturated* solutions and they are not stable; given an opportunity (such as dropping a crystal of solute in the solution), the excess solute will precipitate from the solution.

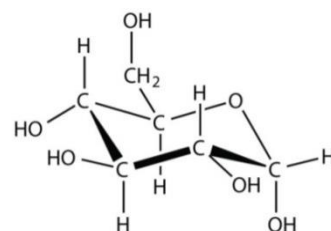
It should be obvious that some solutes dissolve in certain solvents but not in others. NaCl, for example, dissolves in water but not in vegetable oil. Beeswax dissolves in liquid hexane but not in water.

From experimental studies, it has been determined that if molecules of a solute experience the same intermolecular forces that the solvent does, the solute will likely dissolve in that solvent. So, NaCl — a very polar substance because it is composed of ions — dissolves in water, which is very polar, but not in oil, which is generally nonpolar. Nonpolar wax dissolves in nonpolar hexane but not in polar water. This concept leads to the general ancient alchemic rule that “like dissolves like” for predicting whether a solute is soluble in a given solvent. However, this is a general rule, not an absolute statement, so it must be applied with care.

Questions:

- Define *solute* and *solvent*.
- Define *saturated*, *unsaturated*, and *supersaturated* solutions.
- Differentiate between polar and nonpolar solvents.
- Which solvent is Br₂ more likely soluble in – CH₃OH or C₆H₆?
- Which solvent is NaOH more likely soluble in – CH₃OH or C₆H₆?
- Compounds with the formula C_nH_{2n + 1}OH are soluble in H₂O when *n* is small but not when *n* is large. Suggest an explanation for this phenomenon.
- Glucose has the following structure:

What parts of the molecule indicate that this substance is soluble in water?



TEST FOR CLASSWORK

- Choose insoluble (solubility < 0.1 g per 100 g of H_2O) salts:
a. KCl b. Na_2SO_4 c. CaCO_3 d. BaSO_4
- Choose soluble (solubility > 1 g per 100 g of H_2O) salts:
a. NaI b. $\text{Zn}(\text{NO}_3)_2$ c. AgCl d. $\text{Zn}_3(\text{PO}_4)_2$
- Slightly soluble salt (with solubility between 0.1 and 1 g per 100 g of H_2O) is considered to be insoluble in the written form of chemical reaction:
a. if it is a reactant
b. if it is a product
c. always
d. never
- Will there be a precipitate if we put 0.5 mg or 0.05 mg of CaCO_3 in 200 g of water (CaCO_3 solubility is 0.0006 g per 100 g of H_2O)?
a. Yes / Yes b. Yes / No c. No / Yes d. No / No
- Which substances demonstrate good solubility in water?
a. O_2 b. $\text{C}_2\text{H}_5\text{OH}$ c. N_2 d. HCl
- Which substances demonstrate good solubility in benzene?
a. CH_4 b. C_7H_8 c. H_2O d. C_6H_{14}
- How can we dissolve a precipitate in water solution?
a. increase the temperature
b. decrease the temperature
c. add more water
d. add a substance which reacts with that precipitate
- Molarity is the ratio between:
a. the mass of a solute and the mass of a solution
b. the number of moles of a solute and the mass of a solvent
c. the number of moles of a solute and the volume of a solution
d. the volume of a solute and the volume of a solution
- Mass percentage is the ratio between:
a. the mass of a solute and the mass of a solvent
b. the number of moles of a solute and the mass of a solution
c. the mass of a solute and the mass of a solution
d. the number of moles of a solute and the volume of a solution
- Choose true statements about saturated solution:
a. saturated solution exists upon the precipitate of a solute
b. saturated solution cannot dissolve more solute
c. saturated solution is always considered as concentrated solution
d. saturated solution may have rather low concentration

TEST FOR HOMEWORK

- Choose insoluble (solubility < 0.1 g per 100 g of H_2O) salts:
a. MgSO_4 b. $(\text{NH}_4)_2\text{SO}_4$ c. SrCO_3 d. K_2SiO_3
- Choose soluble (solubility > 1 g per 100 g of H_2O) salts:
a. BaCl_2 b. HNO_3 c. NiCl_2 d. KOH
- Slightly soluble salt (with solubility between 0.1 and 1 g per 100 g of H_2O) is considered to be soluble in the written form of chemical reaction:
a. if it is a reactant c. always
b. if it is a product d. never
- Will there be a precipitate if we put 100 g or 10 g of KBr in 100 g of water (KBr solubility is 70.7 g per 100 g of H_2O)?
a. Yes / Yes b. Yes / No c. No / Yes d. No / No
- Which substances demonstrate low solubility in water?
a. H_2 b. CH_3COOH c. HCOOH d. H_2SiO_3
- Which substances demonstrate low solubility in benzene?
a. C_2H_2 b. NaOH c. $\text{Mg}(\text{OH})_2$ d. C_4H_8
- How can we produce a precipitate in water solution?
a. increase the temperature, dissolve high amount of solute, then cool down the solution
b. increase the temperature and wait until sufficient amount of water will be evaporated, then cool down the solution
c. add more water
d. add a substance which produces precipitate in reaction with a given solute
- Molality is the ratio between:
a. the mass of a solute and the mass of a solvent
b. the number of moles of a solute and the mass of a solvent
c. the number of moles of a solute and the mass of a solution
d. the mass of a solute and the volume of a solution
- Mole fraction is the ratio between:
a. the mass of a solute and the mass of a solvent
b. the number of moles of a solute and the number of moles of all components of a solution
c. the volume of a solute and the mass of a solution
d. the number of moles of a solute and the mass of a solution
- Choose true statements about unsaturated solution:
a. unsaturated solution always have rather low concentration
b. unsaturated solution can dissolve more solute
c. unsaturated solution can dissolve more solvent
d. unsaturated solution may become saturated at lower temperature

EXERCISES FOR CLASSWORK

1. A solution is prepared by combining 2.1 g of CO₂ (gas) and 35.5 g of H₂O (liquid).
Identify the solute and solvent _____
Explain your answer _____
2. A solution is prepared by combining 10.5 g of CH₃OH (liquid) and 18.5 g of H₂O (liquid).
Identify the solute and solvent _____
Explain your answer _____
3. Decide if a solution containing 45.0 g of NaCl per 100 g of H₂O is unsaturated, saturated, or supersaturated.

Is this solution rather dilute or concentrated? _____
Explain your answer _____

EXERCISES FOR HOMEWORK

1. A solution is prepared by combining 10.3 g of KOH (solid) and 50.0 g of H₂O (liquid).
Identify the solute and solvent _____
Explain your answer _____
2. A solution is prepared by combining 15.3 g of Hg (liquid) and 17.5 g of Ag (liquid).
Identify the solute and solvent _____
Explain your answer _____
3. Decide if a solution containing 0.000092 g of AgCl per 100 g of H₂O is unsaturated, saturated, or supersaturated. _____
Is this solution dilute or concentrated? _____
Explain your answer _____

LESSON 23

23.1 MOLARITY AND MOLALITY

Molarity (molar concentration) is defined as the number of moles of a solute divided by the number of liters of a solution:

$$\text{molarity} = \text{moles of solute} / \text{liters of solution}$$

$$C = n(\text{solute}) / V(\text{solution})$$

Molarity is expressed in mol/L which can be simplified as just big letter "M".

A similar in spelling but different in meaning unit of concentration is molality, which is defined as the number of moles of a solute per kilogram of a solvent, not per liter of a solution:

$$\text{molality} = \text{moles of solute} / \text{kilograms of solvent}$$

$$C_m = n(\text{solute}) / m(\text{solvent})$$

23.2 MASS PERCENTAGE

Another way to specify an amount of solute is percentage composition by mass (or *mass percentage*, % m/m). It is defined as follows:

$$\% \text{ m/m} = (\text{mass of solute} / \text{mass of entire sample}) \times 100 \%$$

$$\omega = m(\text{solute}) / m(\text{solution})$$

Mass percentage has a wider sense than just a fraction of solute in a solvent multiplied by 100. The same index is used to describe the mass content of a compound. For example, the mass percentage of potassium (K) in potassium oxide (K₂O) is equal to the ratio between molar mass of potassium multiplied by 2 (imagine that potassium is a solute) and the molar mass of the whole compound (imagine that oxygen is a solvent). In more complicated compounds “solvent” is everything else except atoms for which the mass percentage has to be calculated.

In a similar manner one can calculate volume percentage (percent by volume). It is equal to the ratio between the volume of a solute and the volume of a solution.

$$\% \text{ V/V} = (\text{volume of solute} / \text{volume of entire sample}) \times 100 \%$$

$$\varphi = V(\text{solute}) / V(\text{solution})$$

In the same manner we can compare the numbers of moles for a solute and for all the substances in a solution. That is how a mole percentage is calculated.

$$\% \text{ n}/\Sigma\text{n} = (\text{amount of solute} / \text{sum of amounts for all substances in a sample}) \times 100 \%$$

$$\chi = n(\text{solute}) / \Sigma n(\text{all substances in a solution})$$

In case if three above mentioned indices are calculated without the multiplication by 100, they are called: mass fraction, volume fraction, and mole fraction, respectively.

In the most of the exercises a solution is made just from a solvent and a single solute. For the conversion of mass fraction into molarity (and vice versa) you can take any volume or any mass of a given solution. In many exercises you need to use the density of a solution, that is the ratio between mass and volume. It is usually measured in g/ml or kg/L.

$$\rho = m(\text{solution}) / V(\text{solution})$$

EXERCISES FOR CLASSWORK

1. What is the molarity of a solution made by dissolving 13.4 g of NaNO_3 in water? The final volume of that solution is equal to 345 mL.

2. How many moles of MgCl_2 are present in 0.0331 L of a 2.55 M water solution? Density is 1.2 g/mL. What is the mass percentage of MgCl_2 ?

3. Convert the molarity (0.001 mol/L) of the H_2SO_4 solution into the mass percentage. The density is approximately equal to 1 g/ml.

4. Calculate the molarity of $\text{Ba}(\text{OH})_2$ in the final volume of 750 ml, if the initial mass of $\text{Ba}(\text{OH})_2$ was equal to 17 g.

5. Calculate the mass percentage of KOH solution made from 30 g of KOH dissolved in water until the final volume of 300 ml. The density of the solution is equal to 0.92 g/ml.

6. Calculate the mass of KOH needed to make 400 ml of solution with mass percentage of 15 % and density equal to 0.85 g/ml.

7. Calculate the number of moles of $\text{Ba}(\text{OH})_2$ present in 950 ml of solution with the molarity equal to 0.03 mol/L

8. Calculate the molarity of the water solution made from 40 g of Na_2SO_4 , if the final volume is 1200 ml.

9. Calculate the mass percentage of sodium hydroxide in the water solution made from 12 g of NaOH and 68 ml of pure water.

10. Find the mass of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ which is needed to prepare 1 L of CuSO_4 water solution with the molarity equal to 0.02 mol/L.

11. What volume of 5.56 M NaCl is needed to obtain 2L of 0.85 % NaCl solution? Density is equal to 1 g/mL.

12. Calculate the mass of a precipitate formed after the mixing of 20 ml of 0.02 M potassium chloride and 15 ml of 0.01 M silver nitrate solutions.

13. Calculate the mass of a precipitate formed after the mixing of 40 ml of 0.01 M sodium sulfate and 35 ml of 0.02 M strontium chloride solutions.

14. Calculate the mass percentage of nitrogen in NH_4NO_3 .

15. Calculate the mass percentage of oxygen in $\text{CaO} \cdot 2\text{K}_2\text{O} \cdot 6\text{SiO}_2$.

EXERCISES FOR HOMEWORK

1. What is the molarity of a solution made by dissolving 332 g of $C_6H_{12}O_6$ in 4.66 kg of water? The density of the final solution is 1.12 g/ml.

2. How many moles of NH_4Br are present in 88.9 mL of a 0.228 M water solution? Density is 1.1 g/mL. What is the mass percentage of NH_4Br ?

3. Convert the mass percentage (0.85 %) of the $NaCl$ water solution into the molarity. The density is approximately equal to 1 g/ml.

4. Calculate the molarity of $LiOH$ in the final volume of 300 ml, if the initial mass of $LiOH$ was equal of 7 g.

5. Calculate the mass percentage of $NaOH$ solution made from 23 g of $NaOH$ dissolved in water until the final volume of 200 ml. The density of the solution is equal to 0.91 g/ml.

6. Calculate the mass of KCl needed to make 200 ml of solution with mass percentage of 5 % and density equal to 0.95 g/ml.

7. Calculate the number of moles of $Sr(OH)_2$ present in 650 ml of solution with the molarity equal to 0.09 mol/L

8. Calculate the molarity of the water solution made from 10 g of H_2SO_4 , if the final volume is 2400 ml.

9. Calculate the mass percentage of potassium hydroxide in the water solution made from 15 g of KOH and 135 ml of pure water.

10. Find the mass of $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$ which is needed to prepare 1 L of Na_2SO_4 water solution with the mass percentage equal to 0.5 %. The density is 1 g/ml.

11. What mass of 96% $\text{C}_2\text{H}_5\text{OH}$ is needed to obtain 3L of 40 % $\text{C}_2\text{H}_5\text{OH}$ solution? Density is equal to 0.94 g/ml.

12. Calculate the mass of a precipitate formed after the mixing of 30 ml of 0.03 M barium chloride and 20 ml of 0.02 M lithium sulfate solutions.

13. Calculate the mass of a precipitate formed after the mixing of 70 ml of 0.02 M potassium silicate and 10 ml of 0.1 M hydrochloric acid solutions.

14. Calculate the mass percentage of hydrogen in $\text{H}_2\text{C}_2\text{O}_4 \cdot 2\text{H}_2\text{O}$.

15. Calculate the mass percentage of oxygen in $\text{CaSO}_4 \cdot \frac{1}{2}\text{H}_2\text{O}$.

LESSON 24

24.1 THEORY OF ELECTROLYTIC DISSOCIATION

Dissociation in chemistry is a general process in which ionic compounds separate or split into smaller charged particles — ions. For example, when a hydrogen chloride is put in water, a covalent bond between an electronegative chlorine atom and a hydrogen atom is broken by heterolytic fission, which gives a proton and a negatively charged ion. Dissociation process is frequently confused with ionization. Although it may seem as a case of ionization, in reality the ions of any salt already exist within the crystal lattice. When salt is dissociated, the bonds between cations and anions are destroyed and they become surrounded by water molecules (i. e. solvation of ions happens) and their effects become visible (e. g. the solution becomes electrolytic). However, no transfer or displacement of electrons occurs. Actually, the chemical synthesis of salt from metals and nonmetals involves ionization.

Solvation, also sometimes called dissolution, is the process of attraction and association of molecules of a solvent with molecules or ions of a solute. As ions dissolve in a solvent they spread out and become surrounded by solvent molecules.

Polar solvents are those with a molecular structure that contains dipoles. The polar molecules of these solvents can solvate ions because they can orient the appropriate partially charged portion of the molecule towards the ion in response to electrostatic attraction. This stabilizes the system and creates a solvation shell (or hydration shell in the case of water), as it is shown in Figure 24.1.

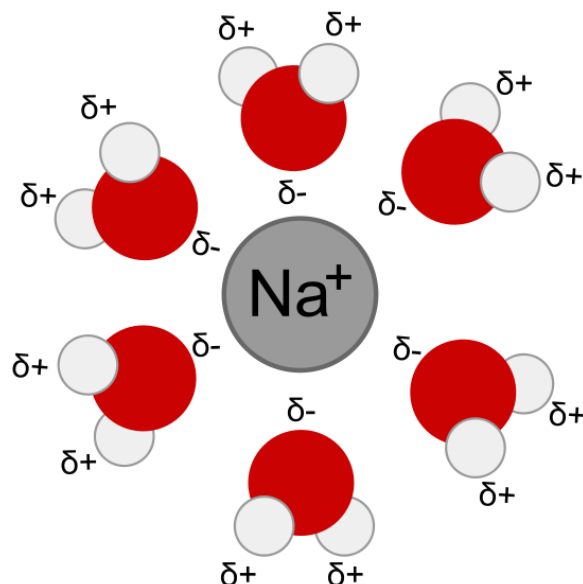
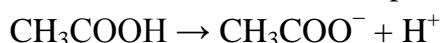


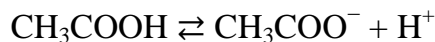
Figure 24.1. A scheme of hydrated sodium cation

Any acid that dissociates 100 % into ions is called a strong acid. If it does not dissociate 100 %, it is a weak acid. CH_3COOH is an example of a weak acid:



The ratio between the number of dissociated compounds and the number of dissolved compounds is called dissociation degree (α). It can be expressed in percent as well. Namely, for 1M acetic acid dissociation degree in water in normal conditions is equal to 0.4 %.

Because the process of dissociation for weak electrolytes does not go 100 % to completion, it is more appropriate to write it as an equilibrium:



As it turns out, there are very few strong acids, which are given in the Table 24.1 below. If an acid is not listed here, it is a weak acid. It may be 1 % ionized or 99 % ionized, but it is still classified as a weak acid.

The issue is similar with bases: a strong base is a base that is 100 % ionized in water solution. If it is less than 100 % ionized in solution, it is a weak base. There are very few strong bases; any base not listed is a weak base.

Table 24.1

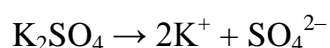
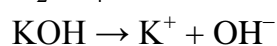
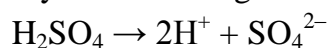
Strong Acids and Bases

Acids	Bases
HNO ₃	LiOH
H ₂ SO ₄	NaOH
HI	KOH
HBr	RbOH
HCl	CsOH
HClO ₃	Ca(OH) ₂
HClO ₄	Sr(OH) ₂
	Ba(OH) ₂

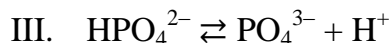
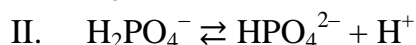
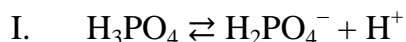
The definitions of strong acid and base given above come from Arrhenius dissociation theory. It is known that the higher the concentration of a weak electrolyte, the lower the dissociation degree. The lower the concentration, the higher the percent of dissociated compounds. So, in very dilute solutions weak acids become completely dissociated. In concentrated solutions strong acids partially exist as nondissociated molecules. The information from the table 24.1 is reliable for average concentrations of acids and bases usually used in chemistry.

24.2 DISSOCIATION EQUATIONS

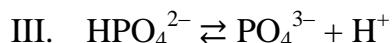
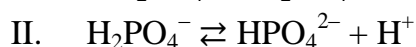
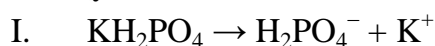
To write the equation of electrolytic dissociation for a strong electrolyte one needs to separate cations from anions, prescribe then charges and balance the whole equation. Strong electrolytes are: strong acids, strong bases and all the soluble salts.



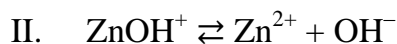
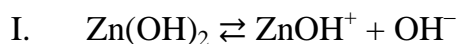
Weak acids containing several hydrogen atoms (polyprotic acids) which can be dissociated in water solutions dissociate in a stepwise manner, just like phosphoric acid does.



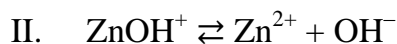
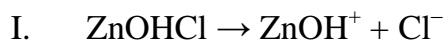
Acidic salts of weak acids dissociate in a similar stepwise manner. However, there is a difference from the previous example for a weak acid. The first step is the dissociation of a metal cation from the remaining part of a salt (a protonated anion). This step is irreversible (you need to draw a one way arrow to the right). Other steps (deprotonation of an anion) are reversible (you need to draw two arrows: to the right and to the left).



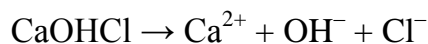
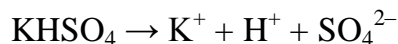
Weak bases containing several OH^- groups dissociate in several reversible steps.



Yet another example is for the dissociation of a basic salt made from a weak base. The first step is irreversible, unlike the second one.



Acidic salts of strong acids and basic salts of strong bases dissociate irreversibly in one step.



Molar mass of any ion can be calculated as a sum of molar masses for all of its atoms. Masses of electrons are usually ignored.

Number of protons in the whole ion equals to the sum of protons in each of its atoms, while the number of electrons is equal to the sum of protons in each of the atoms *minus total charge of the ion*. It means that for negatively charged ions the number of electrons is higher than the number of protons (minus and minus give plus). For positively charged ions the number of electrons is lower than the number of protons.

Questions:

- Describe the process of electrolytic dissociation.
- What is solvation?
- What is dissociation degree?
- What is the difference between weak and strong acids, weak and strong bases?
- How do acidic compounds containing several hydrogen atoms which can become H^+ (H_3O^+) ions dissociate?

f. How do basic compounds containing several OH groups which can become OH⁻ ions dissociate?

EXERCISES FOR CLASSWORK

1. Write 5 samples of strong electrolytes:

2. Write 5 samples of weak electrolytes:

3. Write equations of electrolytic dissociation for the following substances:

HCl: _____

H₂SO₄: _____

LiOH: _____

Ca(OH)₂: _____

4. Write equations of *stepwise* dissociation for the following substances:

H₂SO₃: _____

H₃PO₄: _____

NaHCO₃: _____

ZnOHCl: _____

5. How many types of ions are formed in the process of electrolytic dissociation of the following substances:

K₃PO₄: _____

K₂HPO₄: _____

KH₂PO₄: _____

MgCl₂: _____

MgOHCl: _____

6. What is the number of moles of ions formed after the dissociation of 1 mole of the following substances:

NaCl: _____

MgCl₂: _____

AlCl₃: _____

Al₂(SO₄)₃: _____

Ba(OH)₂: _____

7. Calculate the molar concentration of chloride ions in the water solution prepared from 10 g of aluminum chloride. The final volume is 300 ml.

8. Calculate the molar concentration of sulfate ions in the water solution prepared from 15 g of potassium sulfate. The final volume is 400 ml.

9. Calculate the molar concentration of all ions in the water solution prepared from 20 g of barium chloride. The final volume is 2 L.

10. Calculate the molar concentration of hydrogen ions in the water solution prepared from 5 L of hydrogen chloride (in normal conditions). The final volume is 500 ml.

EXERCISES FOR HOMEWORK

1. Write 5 samples of nonelectrolytes:

2. Write 5 samples of substances which have two steps of dissociation:

3. Write equations of electrolytic dissociation for the following substances:

NaOH: _____

Ba(OH)₂: _____

HI: _____

HClO₄: _____

4. Write equations of stepwise dissociation for the following substances:

H₂S: _____

Al(OH)₃: _____

KHSO₃: _____

CuOHBr: _____

5. How many types of ions are formed in the process of electrolytic dissociation of the following substances:

K_2SO_4 : _____

$KHSO_4$: _____

$KHSiO_3$: _____

$AlCl_3$: _____

$Al(OH)_2Cl$: _____

6. What is the number of moles of ions formed after the dissociation of 1 mole of the following substances:

KI: _____

Na_2SO_4 : _____

CaI_2 : _____

CH_3COOK : _____

$Sr(OH)_2$: _____

7. Calculate the molar concentration of sodium ions in the water solution prepared from 13 g of sodium sulfate. The final volume is 200 ml.

8. Calculate the molar concentration of potassium ions in the water solution prepared from 3 g of potassium. The final volume is 100 ml.

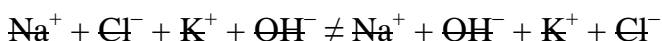
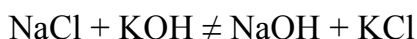
9. Calculate the molar concentration of hydrogen ions in the water solution prepared from 4 g of potassium hydrogen sulfate. The final volume is 200 ml.

10. Calculate the total molar concentration of all phosphate, hydrogen phosphate and dihydrogen phosphate anions in the water solution prepared from 14 g of sodium phosphate. The final volume is 400 ml.

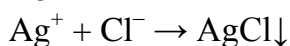
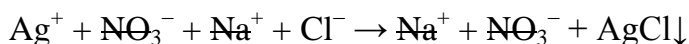
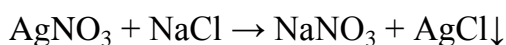
LESSON 25

25.1 IONIC EQUATIONS

Ionic equations are used to display what chemical processes really take place in the water solution. Imagine that two solutions are mixed together. As a result, another solution can be formed in which all the ions from two solutions are present, while those ions don't react with each other. For example, chemical reaction does not take place in the mixture of NaCl and KOH solutions:



However, in case if insoluble substance is formed, one may write down the short ionic equation.

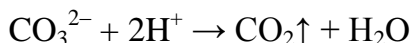
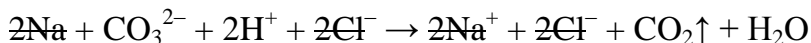
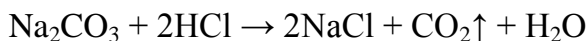


The first so-called "complete" form of ionic equation contains identical ions in both parts. These ions do not participate in the chemical reaction and must be crossed out from the equation. The final so-called "short" ionic equation contains no repeating particles, and in a given case it is common for all the reactions between soluble salts of silver and soluble chlorides.

According to the rules, in ionic reactions strong electrolytes are decomposed into ions (cations and anions are written separately). All the soluble salts are considered to be strong electrolytes (together with strong acids and alkali). Weak electrolytes are written as they are (they are not separated into cations and anions), as well as nonelectrolytes. Weak electrolytes are weak acids and bases.

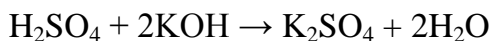
25.2 EXAMPLES OF IONIC EQUATIONS

In case if volatile substance (gas) is formed after the double replacement reaction, then one may also write down a short ionic equation. Remember that NH_4OH decomposes into NH_3 (gas) and H_2O , H_2CO_3 decomposes into CO_2 and H_2O , H_2SO_3 decomposes into SO_2 and H_2O , while H_2S is volatile itself (without any decomposition).



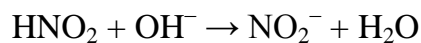
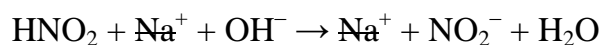
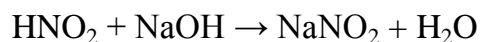
That ionic equation is a common one for all the reactions between soluble carbonates and strong acids.

Neutralization reactions between strong acids and strong bases can be described by the common short ionic equation as well.



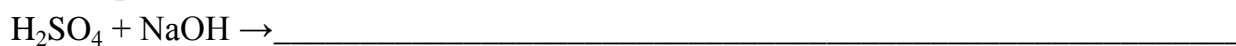


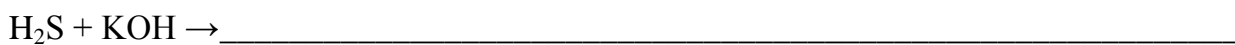
All weak electrolytes are not written as ions in ionic equations.

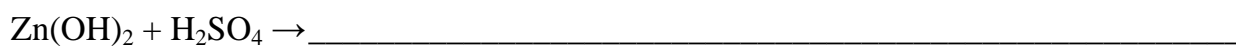


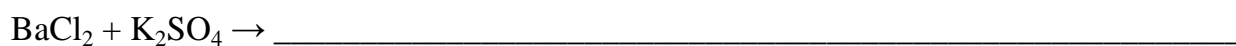
EXERCISES FOR CLASSWORK

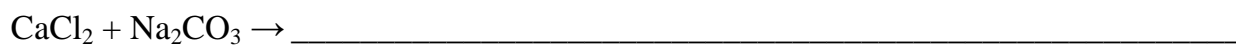
1. Continue equations of chemical reactions, balance and write complete and short ionic equations for them:

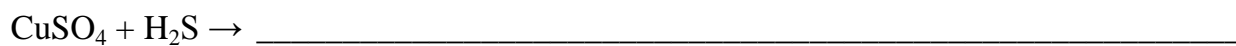


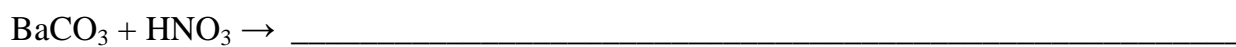


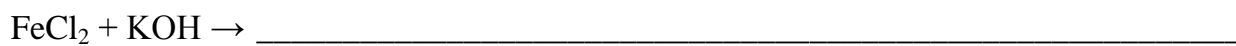


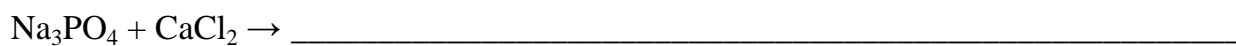




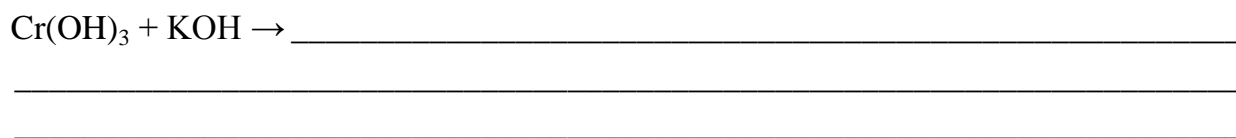




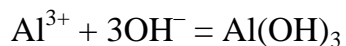
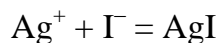






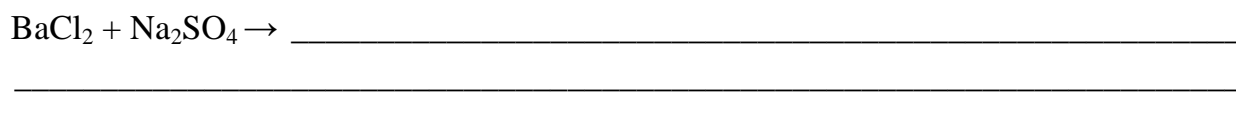
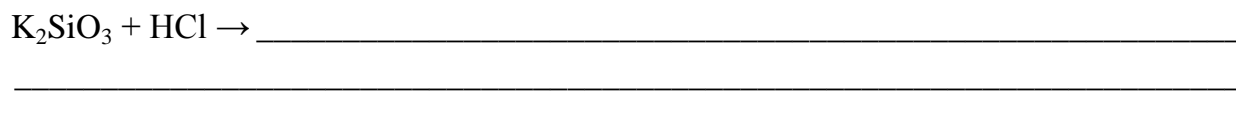
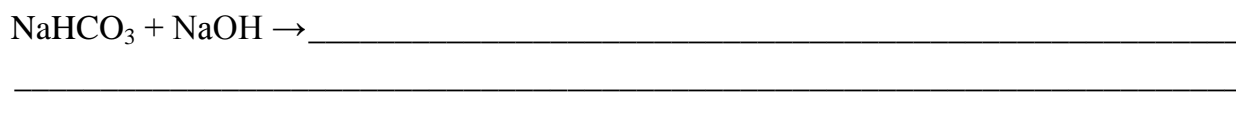
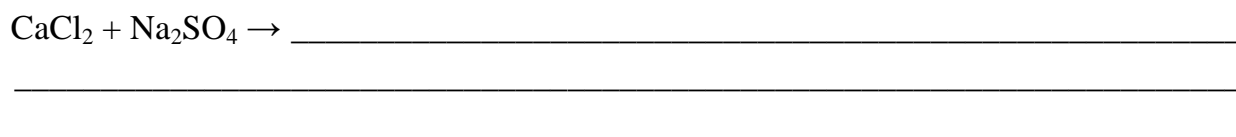
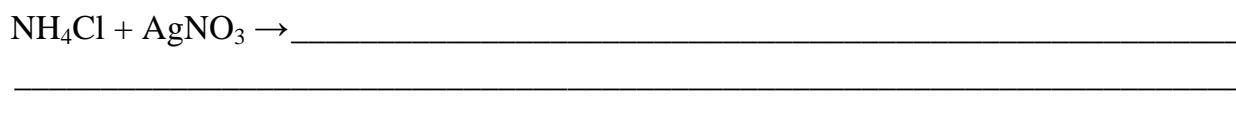
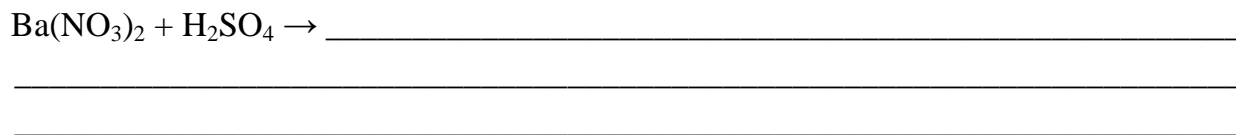
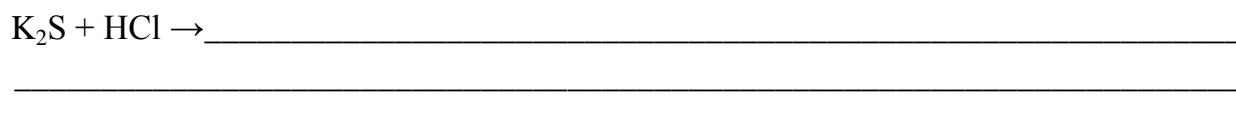
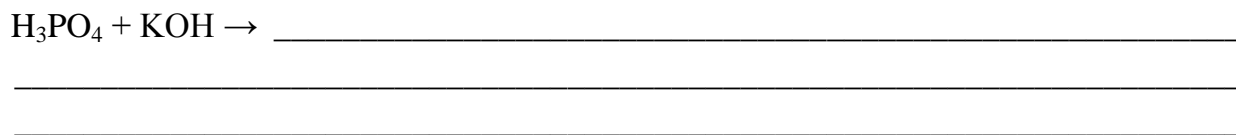


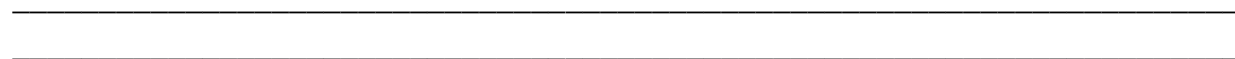
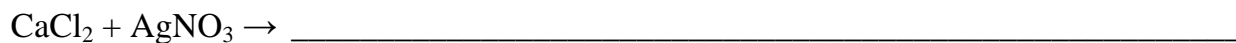
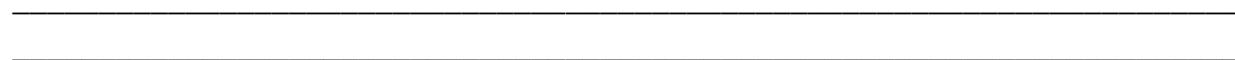
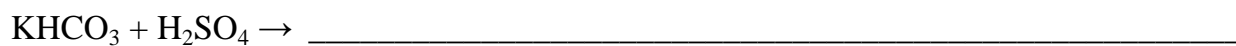
2. Suggest a chemical reaction which can be expressed by the given short ionic equation and write it down:



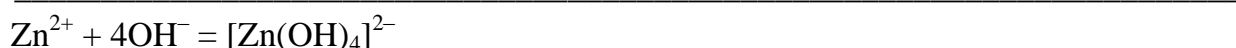
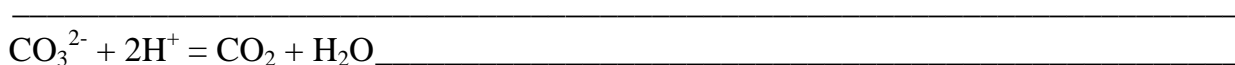
EXERCISES FOR HOMEWORK

1. Continue equations of chemical reactions, balance and write complete and short ionic equations for them:





2. Suggest a chemical reaction which can be expressed by the given short ionic equation and write it down:



LESSON 26

26.1 WHEN HYDROLYSIS IS IMPOSSIBLE

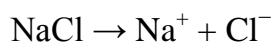
Hydrolysis of a salt is the process that is a backward one to the neutralization. Corresponding acid and base may be formed in the reaction between salt and water. One can say that cations may reunite with OH^- anions to restore a base, while anions may reunite with H^+ cations to restore an acid. Some salts cannot exist in water at all. They are hydrolyzed completely.



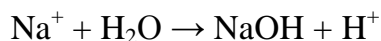
However, the most of the salts are either hydrolyzed partially, or cannot be hydrolyzed at all.

Certain salts affect the acidity or basicity of aqueous solutions because some of the ions are hydrolyzed. The general rule is that salts with ions that are **parts of strong acids or bases will not hydrolyze**, while salts with ions that are **parts of weak acids or bases will hydrolyze**.

Consider NaCl . When it dissolves in an aqueous solution, it separates into Na^+ ions and Cl^- ions:



Will the Na^+ ion be hydrolyzed? If it does, it will interact with the OH^- ion to make NaOH :



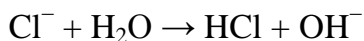
However, NaOH is a strong base, which means that it is 100 % ionized in solution:



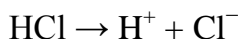
The free OH^- ion reacts with the H^+ ion to remake a water molecule:



The net result? There is no change, so there is no effect on the acidity or basicity of the solution from the Na^+ ion. What about the Cl^- ion? Will it be hydrolyzed? If it does, it will take an H^+ ion from a water molecule:



However, HCl is a strong acid, which means that it is 100 % ionized in solution:



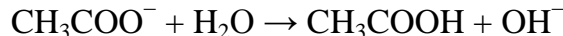
The free H^+ ion reacts with the OH^- ion to remake a water molecule:



The net result? There is no change, so there is no effect on the acidity or basicity of the solution from the Cl^- ion. Because neither ion in NaCl affects the acidity or basicity of the solution, NaCl is an example of salt which is not hydrolyzed at all.

26.2 WHEN HYDROLYSIS IS POSSIBLE

Things change, however, when we consider a salt like CH_3COONa . We already know that the Na^+ ion won't affect the acidity of the solution. What about the acetate ion? If it hydrolyzes, it will take an H^+ from a water molecule:



Does this happen? Yes, it does. Why? *Because CH_3COOH is a weak acid.* Any chance a weak acid has to form, it will (the same with a weak base). As some CH_3COO^- ions hydrolyze with H_2O to make the molecular weak acid, OH^- ions are produced. OH^- ions make solutions basic. Thus CH_3COONa solutions are slightly basic.

There are also salts which aqueous solutions are slightly acidic. NH_4Cl is an example. When NH_4Cl is dissolved in H_2O , it separates into NH_4^+ ions and Cl^- ions. We have already seen that the Cl^- ion does not hydrolyze. However, the NH_4^+ ion will:



H_3O^+ ion is the hydronium ion, the more chemically proper way to represent the H^+ ion. This is the classic acid species in solution, so a solution of NH_4^+ ions is slightly acidic. The molecule NH_3 is a weak base, and it will be formed when it can, just like a weak acid will be formed when it can.

So there are two general rules: 1) if an ion derives from a strong acid or base, it will not affect the acidity of the solution; 2) if an ion derives from a weak acid, it will make the solution basic; if an ion derives from a weak base, it will make the solution acidic.

Follow the Table 26.1 to predict the acidity of a solution of a salt.

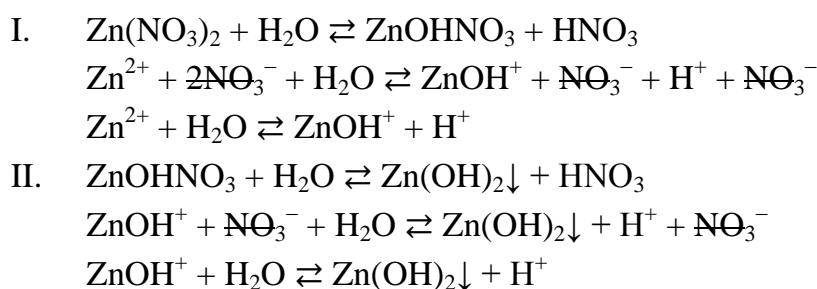
pH in water solutions of salts

Cation / Anion	Strong acid	Weak acid
Strong base	pH = 7	pH > 7
Weak base	pH < 7	pH ≈ 7

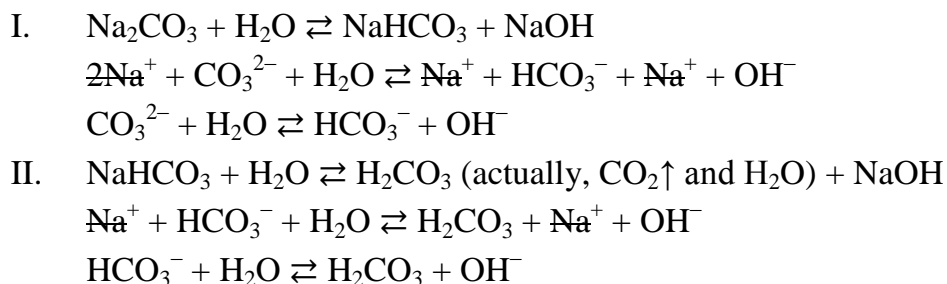
If pH is equal to 7, the medium is neutral, if pH is lower than 7, the medium is acidic, if pH is higher than 7, the medium is basic.

The most popular indicator of pH is litmus. This substance is red in acidic solutions, violet (or purple) in neutral solutions, and blue in basic solutions.

Hydrolysis is stepwise for compounds consisting from cations of weak bases with charge higher than +1 and anions of weak acids with charge lower than -1.



Another example is for salt containing anion with 2- charge.



Each next step of reversible hydrolysis is much less productive than the previous one.

TEST FOR CLASSWORK

- Which ions are present in NaH_2PO_4 solution?
 - Na^+
 - H_2PO_4^-
 - HPO_4^{2-}
 - NaH_2^{3+}
- Choose strong electrolytes:
 - NaCl
 - KNO_2
 - $\text{C}_6\text{H}_{12}\text{O}_6$
 - HNO_2
- Choose weak electrolytes:
 - $\text{C}_2\text{H}_5\text{OH}$
 - CH_3COOH
 - AgNO_3
 - $\text{Zn}(\text{OH})_2$
- Which reactions can be expressed by the same ionic equation?
 - $2\text{NaOH} + \text{H}_2\text{SO}_4 \rightarrow \text{Na}_2\text{SO}_4 + 2\text{H}_2\text{O}$
 - $\text{Na}_2\text{O} + \text{H}_2\text{SO}_4 \rightarrow \text{Na}_2\text{SO}_4 + \text{H}_2\text{O}$
 - $2\text{Na}_3\text{PO}_4 + 3\text{Li}_2\text{SO}_4 \rightarrow 3\text{Na}_2\text{SO}_4 + 2\text{Li}_3\text{PO}_4$
 - $\text{KOH} + \text{HNO}_3 \rightarrow \text{KNO}_3 + \text{H}_2\text{O}$

5. Choose the correct ionic equation for the following reaction:
 $\text{Na}_2\text{CO}_3 + \text{Ca}(\text{OH})_2 \rightarrow \text{CaCO}_3 + 2\text{NaOH}$
- $\text{CO}_3^{2-} + \text{Ca}(\text{OH})_2 \rightarrow \text{CaCO}_3 + 2\text{OH}^-$
 - $\text{CO}_3^{2-} + \text{Ca}^{2+} \rightarrow \text{CaCO}_3$
 - $\text{Na}^+ + \text{OH}^- \rightarrow \text{NaOH}$
 - $\text{Na}_2\text{CO}_3 + \text{Ca}^{2+} \rightarrow \text{CaCO}_3 + 2\text{Na}^+$
6. In water solutions of which substances there is acidic medium ($\text{pH} < 7$)?
- HCl
 - H_2S
 - ZnCl_2
 - KCl
7. In water solutions of which substances there is basic medium ($\text{pH} > 7$)?
- H_2SO_4
 - K_3PO_4
 - NH_3
 - NH_4Cl
8. What substances will be formed if we mix solutions of K_2S and AlCl_3 together?
- H_2S
 - Cl_2
 - Al_2S_3
 - $\text{Al}(\text{OH})_3$
9. Dissolving HCl in water includes such steps, as:
- ionization and dissociation
 - just dissociation
 - just ionization
 - neither ionization, nor dissociation
10. Dissolving ZnCl_2 in water includes such steps, as:
- ionization and dissociation
 - dissociation and partial hydrolysis
 - just dissociation
 - just complete hydrolysis

TEST FOR HOMEWORK

- Which ions are present in NH_4Cl solution?
 - N^{3+}
 - NH_4^+
 - H^+
 - Cl^-
- Choose strong electrolytes:
 - NaOH
 - NO_2
 - HClO_4
 - H_2O_2
- Choose weak electrolytes:
 - $\text{C}_3\text{H}_7\text{OH}$
 - CH_3NH_2
 - HI
 - $\text{Ca}(\text{OH})_2$
- Which reactions can be expressed by the same ionic equation?
 - $\text{Ba}(\text{OH})_2 + \text{H}_2\text{SO}_4 \rightarrow \text{BaSO}_4 + 2\text{H}_2\text{O}$
 - $\text{Ba} + \text{H}_2\text{SO}_4 \rightarrow \text{BaSO}_4 + \text{H}_2$
 - $\text{BaCl}_2 + \text{Na}_2\text{SO}_4 \rightarrow \text{BaSO}_4 + 2\text{NaCl}$
 - $\text{Ba}(\text{NO}_3)_2 + \text{K}_2\text{SO}_4 \rightarrow 2\text{KNO}_3 + \text{BaSO}_4$
- Choose the correct ionic equation for the following reaction:
 $\text{HCl} + \text{KOH} \rightarrow \text{KCl} + \text{H}_2\text{O}$
 - $\text{HCl} + \text{OH}^- \rightarrow \text{Cl}^- + \text{H}_2\text{O}$

- b. $K^+ + Cl^- \rightarrow KCl$
 c. $H^+ + OH^- \rightarrow H_2O$
 d. $KOH + H^+ \rightarrow K^+ + H_2O$
6. In water solutions of which substances there is acidic medium ($pH < 7$)?
 a. CO_2 b. $AlCl_3$ c. $FeBr_2$ d. K_2SO_3
7. In water solutions of which substances there is basic medium ($pH > 7$)?
 a. NO_2 b. KNO_2 c. Na_2SiO_3 d. $NaCl$
8. What substances will be formed if we mix solutions of Na_2SO_3 and $CrCl_3$ together?
 a. SO_2 b. Cr c. $Cr(OH)_3$ d. $NaCl$
9. Dissolving $NaCl$ in water includes such steps, as:
 a. ionization and dissociation
 b. just dissociation
 c. just ionization
 d. neither ionization, nor dissociation
10. Dissolving Na_3PO_4 in water includes such steps, as:
 a. ionization and dissociation
 b. dissociation and partial hydrolysis
 c. just dissociation
 d. just complete hydrolysis

EXERCISES FOR CLASSWORK

1. Write the equation of hydrolysis for given substances, if it is possible, provide its complete and short ionic forms.

$KNO_2 + H_2O$ _____

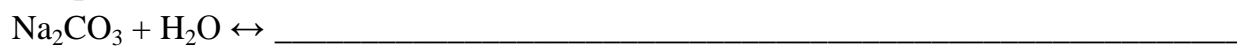
$NH_4NO_3 + H_2O$ _____

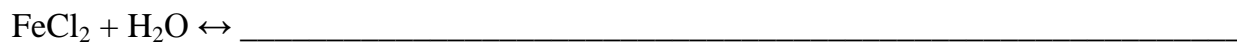
$CH_3COONa + H_2O$ _____

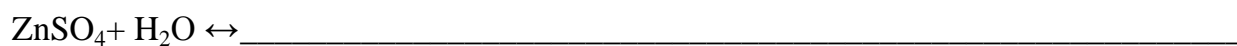
$KCl + H_2O$ _____

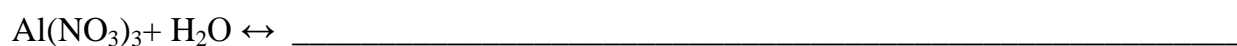
$NH_4Br + H_2O$ _____

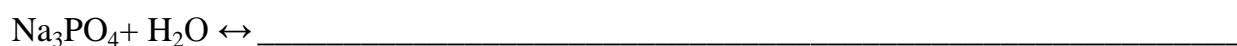
2. Write the equation of stepwise hydrolysis for given substances, provide its complete and short ionic forms.









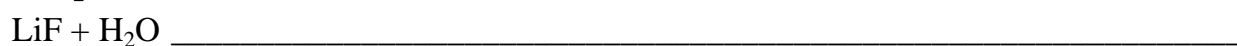






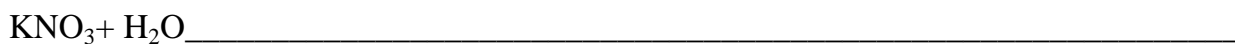
EXERCISES FOR HOMEWORK

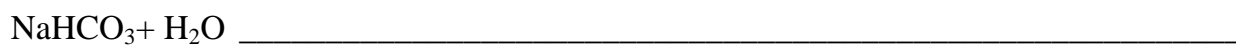
1. Write the equation of hydrolysis for given substances, if it is possible, provide its complete and short ionic forms.



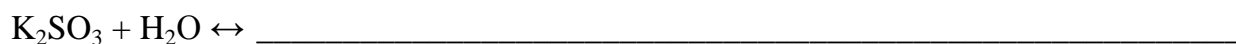


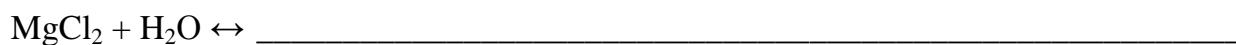


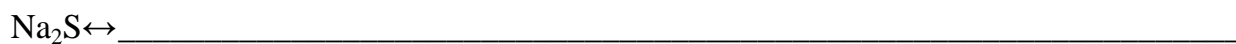


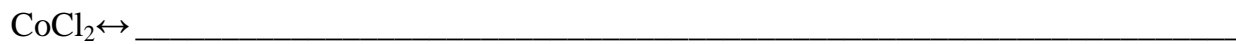


- 2. Write the equation of stepwise hydrolysis for given substances, provide its complete and short ionic forms.**

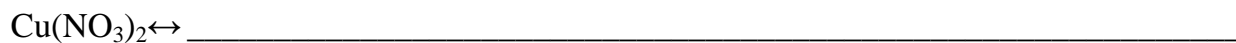


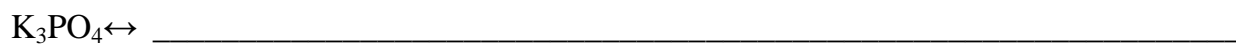










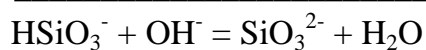
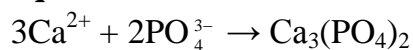


LESSON 27

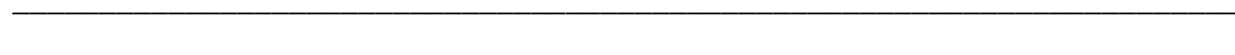
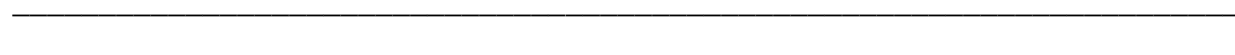
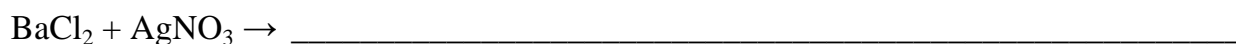
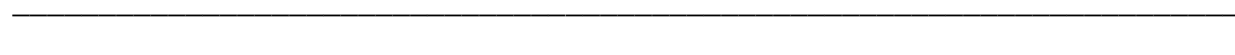
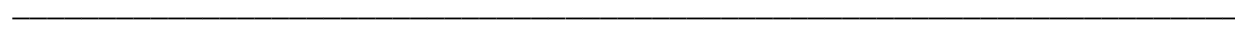
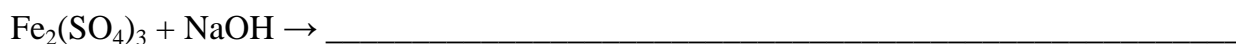
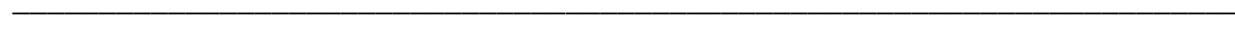
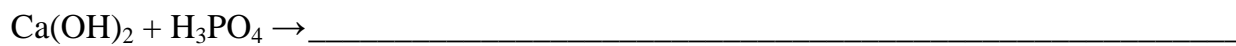
SAMPLE TICKET FOR CONTROL TASK #4

ON ELECTROLYTIC DISSOCIATION

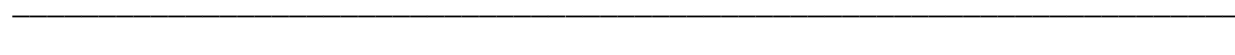
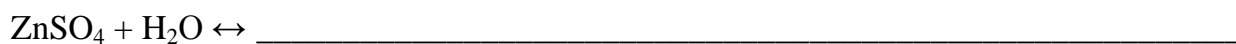
1. Suggest a chemical reaction which can be expressed by the given short ionic equation and write it down:

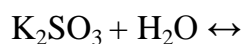


2. Continue equations of chemical reactions, balance and write complete and short ionic equations for them:



3. Write the equation of stepwise hydrolysis for given substances, provide its complete and short ionic forms.





4. Indicate the pH level in water solutions of the following substances:



5. Calculate the number of protons and the number of electrons for the following ions:



6. Calculate the mass of sulfate ions formed after the addition of 3L of sulfur (VI) oxide (density = 1.92 g/ml) into 2L of liquid water (density = 1 g/ml).

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PERIODIC TABLE OF THE ELEMENTS

Period	GROUPS OF ELEMENTS																	
	IA											VIIA	VIIIA					
1	1 2,1 H Hydrogen 1,00794											2 - He Helium 4,0026						
2	3 0,97 Li Lithium 6,941	4 1,47 Be Beryllium 9,01218											5 2,01 B Boron 10,811	6 2,50 C Carbon 12,011	7 3,07 N Nitrogen 14,007	8 3,50 O Oxygen 15,9994	9 4,10 F Fluorine 18,9984	10 - Ne Neon 20,1797
3	11 1,01 Na Sodium 22,9898	12 1,23 Mg Magnesium 24,305											13 1,47 Al Aluminum 26,9815	14 1,74 Si Silicon 28,086	15 2,10 P Phosphorus 30,9738	16 2,60 S Sulfur 32,066	17 2,83 Cl Chlorine 35,452	18 - Ar Argon 39,948
4	19 0,91 K Potassium 39,0983	20 1,04 Ca Calcium 40,078	21 1,20 Sc Scandium 44,956	22 1,32 Ti Titanium 47,87	23 1,45 V Vanadium 50,942	24 1,56 Cr Chromium 51,996	25 1,60 Mn Manganese 54,938	26 1,64 Fe Iron 55,845	27 1,70 Co Cobalt 58,933	28 1,75 Ni Nickel 58,693	29 1,75 Cu Copper 63,546	30 1,66 Zn Zinc 65,39	31 1,82 Ga Gallium 69,723	32 2,02 Ge Germanium 72,61	33 2,20 As Arsenic 74,922	34 2,48 Se Selenium 78,96	35 2,74 Br Bromine 79,904	36 - Kr Krypton 83,80
5	37 0,89 Rb Rubidium 85,468	38 0,99 Sr Strontium 87,62	39 1,11 Y Yttrium 88,906	40 1,22 Zr Zirconium 91,224	41 1,23 Nb Niobium 92,906	42 1,30 Mo Molybdenum 95,94	43 1,36 Tc Technetium [98]	44 1,42 Ru Ruthenium 101,07	45 1,45 Rh Rhodium 102,905	46 1,35 Pd Palladium 106,42	47 1,9 Ag Silver 107,868	48 1,7 Cd Cadmium 112,411	49 1,49 In Indium 114,82	50 1,72 Sn Tin 118,71	51 1,82 Sb Antimony 121,76	52 2,01 Te Tellurium 127,60	53 2,21 I Iodine 126,904	54 - Xe Xenon 131,29
6	55 0,86 Cs Cesium 132,905	56 0,97 Ba Barium 137,327	57 1,08 *La Lanthanum 138,906	72 1,23 Hf Hafnium 178,49	73 1,33 Ta Tantalum 180,948	74 1,40 W Tungsten 183,84	75 1,46 Re Rhenium 186,207	76 1,52 Os Osmium 190,23	77 1,55 Ir Iridium 192,22	78 1,44 Pt Platinum 195,08	79 2,4 Au Gold 196,967	80 1,9 Hg Mercury 200,59	81 1,44 Tl Thallium 204,383	82 1,55 Pb Lead 207,2	83 1,67 Bi Bismuth 208,980	84 1,76 Po Polonium [209]	85 1,90 At Astatine [210]	86 - Rn Radon [222]
7	87 0,86 Fr Francium [223]	88 0,97 Ra Radium 226,025	89 1,00 ♦Ac Actinium [227]	104 - Rf Rutherfordium [261]	105 - Db Dubnium [262]	106 - Sg Seaborgium [266]	107 - Bh Bohrium [264]	108 - Hs Hassium [265]	109 - Mt Meitnerium [268]	110 - [Uun] [271]	111 - [Uuu] [272]	112 - [Uub] [277]	113 - [Uut] [282]	114 - [Uug] [285]	115 - [Uuh] [289]	116 - [Uuq] [293]	117 - [Uuq] [293]	118 - [Uuo] [293]
*Lanthanide Series				58 1,08 Ce Cerium 140,115	59 1,07 Pr Praseodymium 140,907	60 1,07 Nd Neodymium 144,24	61 1,07 Pm Promethium 144,913	62 1,07 Sm Samarium 150,36	63 1,01 Eu Europium 151,965	64 1,11 Gd Gadolinium 157,25	65 1,10 Tb Terbium 158,925	66 1,10 Dy Dysprosium 162,5	67 1,10 Ho Holmium 164,93	68 1,11 Er Erbium 167,26	69 1,11 Tm Thulium 168,934	70 1,06 Yb Ytterbium 173,04	71 1,14 Lu Lutetium 174,967	
♦Actinide Series				90 1,11 Th Thorium 232,038	91 1,14 Pa Protactinium 231,035	92 1,22 U Uranium 238,028	93 1,22 Np Neptunium 237,048	94 1,22 Pu Plutonium 244,064	95 1,20 Am Americium 243,061	96 1,20 Cm Curium 247,07	97 1,20 Bk Berkelium 247,07	98 1,20 Cf Californium 251,079	99 1,20 Es Einsteinium 252,083	100 1,20 Fm Fermium 257,095	101 1,20 Md Mendelevium 258,099	102 1,20 No Nobelium 18,9984	103 1,20 Lr Lawrencium 260,105	

SOLUBILITY OF INORGANIC SUBSTANCES IN WATER

	Na ⁺	K ⁺	Li ⁺	Ca ²⁺	Ba ²⁺	Sr ²⁺	NH ₄ ⁺	Ag ⁺	Mg ²⁺	Mn ²⁺	Zn ²⁺	Cu ²⁺	Fe ²⁺	Hg ²⁺	Pb ²⁺	Fe ³⁺	Al ³⁺	Cr ³⁺	H ⁺
Cl ⁻	S	S	S	S	S	S	S	I	S	S	S	S	S	S	M	S	S	S	S
Br ⁻	S	S	S	S	S	S	S	I	S	S	S	S	S	M	M	S	S	S	S
I ⁻	S	S	S	S	S	S	S	I	S	S	S	—	S	I	I	—	S	S	S
NO ₃ ⁻	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S
NCS ⁻	S	S	S	S	S	S	S	I	S	S	S	—	S	I	I	S	S	S	S
SO ₄ ²⁻	S	S	S	M	I	I	S	M	S	S	S	S	S	S	I	S	S	S	S
F ⁻	S	S	M	I	M	S	S	S	I	S	S	S	M	—	I	M	M	S	S
CH ₃ COO ⁻	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S
S ²⁻	S	S	S	M	I	I	S	I	I	I	I	I	I	I	I	—	—	—	S
SO ₃ ²⁻	S	S	S	I	I	I	S	I	M	I	I	I	I	—	I	—	—	—	S
CO ₃ ²⁻	S	S	M	I	I	I	S	I	I	I	I	I	I	I	I	—	—	—	S
SiO ₃ ²⁻	S	S	S	I	I	I	—	I	I	I	I	I	I	—	I	—	—	—	I
PO ₄ ³⁻	S	S	I	I	I	I	S	I	I	I	I	I	I	I	I	I	I	I	S
NO ₂ ⁻	S	S	S	S	S	S	S	M	S	—	—	—	—	—	—	—	—	—	S
HS ⁻	S	S	S	S	S	S	S	—	S	—	—	—	—	—	—	—	—	—	S
HSO ₃ ⁻	S	S	S	S	S	S	S	—	S	—	—	—	—	—	—	—	—	—	S
HPO ₄ ²⁻	S	S	—	I	I	I	S	S	M	I	—	—	I	—	M	—	—	—	S
H ₂ PO ₄ ⁻	S	S	S	S	S	S	S	S	S	S	S	—	S	—	—	—	—	—	S
OH ⁻	S	S	S	M	S	S	S	—	I	I	I	I	I	—	I	I	I	I	

S — Soluble; I — Insoluble; M — Marginally soluble

ELECTROCHEMICAL SERIES OF METALS

$\frac{Me^{n+}}{Me^0}$	$\frac{Li^+}{Li^0}$	$\frac{K^+}{K^0}$	$\frac{Sr^{2+}}{Sr^0}$	$\frac{Ba^{2+}}{Ba^0}$	$\frac{Ca^{2+}}{Ca^0}$	$\frac{Na^+}{Na^0}$	$\frac{Mg^{2+}}{Mg^0}$	$\frac{Al^{3+}}{Al^0}$	$\frac{Zn^{2+}}{Zn^0}$	$\frac{Cr^{2+}}{Cr^0}$	$\frac{Fe^{2+}}{Fe^0}$	$\frac{Ni^{2+}}{Ni^0}$	$\frac{Sn^{2+}}{Sn^0}$	$\frac{Pb^{2+}}{Pb^0}$	$\frac{2H^+}{H_2^0}$	$\frac{Bi^{3+}}{Bi^0}$	$\frac{Cu^{2+}}{Cu^0}$	$\frac{Ag^+}{Ag^0}$	$\frac{Hg^{2+}}{Hg^0}$
e ⁰ ,V	-3.04	-2.92	-2.89	-2.90	-2.87	-2.71	-2.37	-1.70	-0.76	-0.74	-0.44	-0.25	-0.14	-0.13	0	+0.21	+0.34	+0.80	+0.85

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