М. М. Бондаренко

ОБЩАЯ ХИМИЯ В ТАБЛИЦАХ И СХЕМАХ ДЛЯ ИНОСТРАННЫХ СЛУШАТЕЛЕЙ =

GENERAL CHEMISTRY IN TABLES AND DIAGRAMS FOR INTERNATIONAL STUDENTS

Учреждение образования
«Международный государственный
экологический институт имени А. Д. Сахарова»
Белорусского государственного университета

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В издании содержатся теоретические материалы по общей химии для поступающих в высшие учебные заведения, представленные в виде таблиц и схем. Включены основные термины, формулы для расчетов, алгоритмы вычислений, которые помогут обучающимся овладеть знаниями в области общей химии.

Предназначается иностранным слушателям для подготовки к вступительным испытаниям по химии в учреждения высшего образования Республики Беларусь по следующим направлениям: биологические и смежные науки, окружающая среда, физические, математические и химические науки, науки о Земле; ветеринария; здравоохранение.

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Contents

Preface

The manual covers the main topics of the pre-university course of general chemistry and all the information is presented in the form of tables and diagrams, aiding clarity. Information in this form is convenient to use not only in preparation for final exams, but also for better assimilation of knowledge in the classroom, homework and self-preparation. Test questions and tasks for better assimilation of theoretical material are included after each topic for independent practice.

The materials were assembled in compliance with the training programme for international students studying at the preparatory courses of the Department of Continuing Education of the Faculty for Advanced Training and Re-training of ISEI BSU.

The study manual is intended for international students preparing for entrance examinations in chemistry at higher educational institutions of the Republic of Belarus in the following areas of specialisation: biological and related sciences; environmental sciences; physical, mathematical and chemical sciences, earth sciences; veterinary medicine; healthcare.

Введение

Издание включает основные темы довузовского курса общей химии. Теория представлена в виде таблиц и схем, наглядно демонстрируя основные законы, алгоритмы вычислений, формулы для решения задач. Представленный материал предназначен не только для подготовки к вступительным испытаниям, но также для использования на занятиях, выполнения домашних заданий и самоподготовки. После каждой темы представлены тестовые вопросы и задания для отработки теоретического материала.

Материалы пособия соответствуют учебной программе для иностранных слушателей, обучающихся на подготовительном отделении факультета повышения квалификации и переподготовки МГЭИ имени А. Д. Сахарова БГУ.

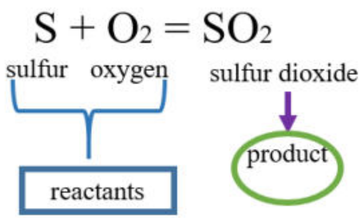
Пособие предназначено для подготовки иностранных слушателей к вступительным испытаниям по химии в учреждения высшего образования Республики Беларусь по следующим направлениям специальностей: биологические и смежные науки, окружающая среда, физические, математические и химические науки, науки о Земле, ветеринария, здравоохранение.

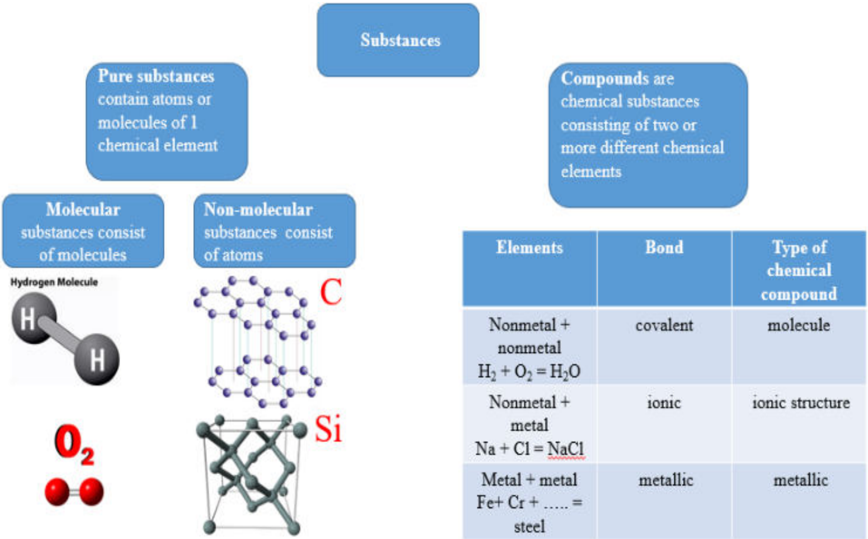
Lesson 1. Chemistry as a science. Properties of substance

Definitions:

1. **Chemistry** is a science that studies substances and their changes;
2. **Substance** is a particular kind of matter with uniform properties;
3. **Atom** is the smallest piece of an element;
4. **Molecule** is the smallest part of a substance that determines physical and chemical properties of that substance;

Physical and chemical properties of a substance — Table 1

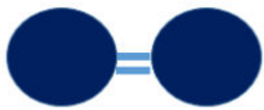
Physical properties of a substance: • state of matter (solid, liquid, gas and plasma) • density • colour • taste • solubility • boiling temperature • freezing temperature In physical processes, the substance changes at least one of its states of matter, while new substances are not formed. **Examples**: melting of ice and freezing of water, boiling of water and condensation of vapour.	**Chemical properties** of a substance are its abilities to form other substances in different conditions. **Chemical reaction** is a process that leads to the transformation of one chemical substance into another. **Example**:  Two pure matters (sulphur – solid, yellow substance with specific smell and oxygen – colourless, odourless gas) form a compound during a chemical reaction.



Elements	Bond	Type of chemical compound
Nonmetal + nonmetal $H_2 + O_2 = H_2O$	covalent	molecule
Nonmetal + metal $Na + Cl = NaCl$	ionic	ionic structure
Metal + metal Fe+ Cr + ... = steel	metallic	metallic

Scheme 1. Classification of substances

Atom and molecule in comparison — Table 2

Atom	Molecule
	
It does not have the composition and properties of those substances that it is part of. Chemically indivisible.	The smallest particle of a substance with its composition and chemical properties. Chemically divisible.

Questions after lesson:

1. What is chemistry as a science?
2. What is an atom?
3. What is a molecule?
4. What is a pure substance?
5. What is a chemical compound?
6. Give examples of molecular substances.
7. Give examples of non-molecular substances.

Test

1. The smallest particle of a chemical element, which is the carrier of its properties:
1) atom;
2) molecule;
3) electron;
4) proton.

2. Choose the chemical compound:
1) water;
2) nitrogen;
3) hydrogen;
4) chlorine gas.

3. Here is a line of pure substances: coal, diamond, graphite, oxygen, ozone. How many chemical elements are included in the composition of these substances:
1) 2;
2) 3;
3) 4;
4) 1.

4. Choose the chemical processes:
1) combustion of potassium;
2) pulling aluminum wire;
3) distillation of oil;
4) the dissolving of glucose in water.

5. Choose the physical processes:
1) evaporation of water from the body;
2) nail rusting;
3) the formation of nitrogen oxide in the atmosphere;
4) the burning of wood.

6. Choose the pure substance:
1) methane;
2) lime water;
3) caustic soda;
4) sodium.

7. Choose compounds which consist of three atoms:
1) H_2O;
2) H_2SO_4;
3) HCl;
4) HCN.

8. Choose compounds which are made from atoms of two chemical elements:
1) Cl_2;
2) N_2O_3;
3) P_4;
4) H_2S.

9. Choose oxygen-containing compounds:
1) O_3;
2) O_2;
3) H_3PO_4;
4) CH_3COOH.

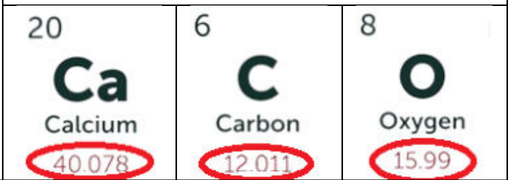
10. Choose nitrogen-containing compounds:
1) N_2;
2) NH_3;
3) HNO_3;
4) NO.

11. Choose molecular compounds:
1) CH_4;
2) NO_2;
3) NaCl;
4) CaO.

12. Choose non-molecular compounds:
1) SiO_2;
2) FeO;
3) H_2O;
4) CL_2.

Lesson 2. Atomic and molecular mass. Molar mass

Quantitative characteristics of a substance Table 3

The atomic mass unit (atomic mass constant) is defined as one-twelfth of the mass of a carbon-12 atom	$u = 1{,}66\ 10^{-24}$ g
The relative atomic mass is a physical quantity defined as the ratio of the mass of atoms of a chemical element in a given sample to the atomic mass constant	Ar 
The molecular mass is the sum of the relative atomic masses of all the atoms in a molecule	Mr Mr ($CaCO_3$) = 40+12+3(16)= 100
Mole is the number of particles of substance equal to the number of atoms in exactly 12 g of carbon-12	$n = \frac{N}{N_A}$ $n = \frac{m}{M}$ N – number of particles of substance; m – mass of substance (g).
Avogadro's number is a constant number of particles of a substance in 1 mol of substance	**1 mol** = $6{,}02 \cdot 10^{23}$ particles $\mathbf{N_A} = 6{,}02 \cdot 10^{23}$ mole^{-1}
Molar mass is the mass of 1 mole of a given substance	M, g/mole M = Mr M ($CaCO_3$) = Mr ($CaCO_3$) = 100 g/mole

Questions after lesson:

1. What is the meaning of mole?
2. What is relative atomic mass?
3. What is the meaning of molar mass?
4. What is an atomic mass constant?
5. What is the meaning of molecular mass?
6. What is an Avogadro's number?

Test

1. A set of atoms with the same charge of nucleus is called:
1) pure substance;
2) chemical compound;
3) molecule;
4) chemical element.

2. Choose substances with the same qualitative composition:
1) SO_2, CO_2;
2) NH_3, PH_3;
3) NO_2, NO_3;
4) HCl, HBr.

3. Choose substances with the same quantitative composition:
1) CO, CO_2;
2) NO, N_2O_3;
3) H_2SO_3; H2SiO_3;
4) $AgNO_3$; $AlPO_4$.

4. The same amounts of substance (in moles) of different substances also have the same:
1) mass;
2) volume;
3) number of structural units;
4) number of atoms.

5. The molecular mass of Na_2CO_3 equals:
1) 89;
2) 78;
3) 106;
4) 109.

6. The molecular mass of $Ba_3(PO_4)_2$ equals:
1) 599;
2) 600;
3) 587;
4) 601.

7. Find the mass of a sample of $ZnBr_2$, chemical amount 0,5 moles :
1) 123;
2) 124;
3) 111,9;
4) 112,5.

8. Find the mass of a sample of Ag_2SO_3, chemical amount 0,13 moles :
1) 38,7;
2) 39;
3) 38,5;
4) 38,4.

9. The number of molecules of 0,9 moles of ammonia equals:
1) $5{,}4 \cdot 10^{23}$;
2) $6{,}02 \cdot 10^{23}$;
3) $4{,}05\ 10^{22}$;
4) $3{,}85\ 10^{24}$.

10. The number of molecules of 0,55 moles of water vapor equals:
1) $0{,}3 \cdot 10^{24}$;
2) $3{,}33 \cdot 10^{23}$;
3) $3{,}3 \cdot 10^{23}$;
4) $33{,}1 \cdot 10^{24}$.

Tasks

1. Calculate the molecular mass of $Fe(OH)_3$.

____

2. Calculate the molecular mass of $Al_2(SO_4)_3$.

____

3. Find the number of moles for $44{,}6 \cdot 10^{23}$ molecules of CO_2.

____

4. Find the number of moles for $17{,}5 \cdot 10^{24}$ molecules of NH_3.

____

5. Calculate the number of moles in 25g of KCl.

____

____

6. Calculate the number of moles in 123g of $Ca_3(PO_4)_2$.

____

____

7. What is the mass of 1,25 moles of H_3PO_4?

8. What is the mass of 2,14 moles of Na_2SO_4?

9. Find the molar mass of a substance if 1,99 moles of it has a mass of 88,91g?

10. Find the molar mass of a substance if 3,1 moles of it has a mass of 56g?

11. How many moles of carbon atoms are there in 50g of CH_3COOH?

12. How many moles of nitrogen atoms are there in 73g of NH_4NO_3?

13. Calculate the moles of a sample of $Al(NO_3)_3$ if you know that there are 2,25 moles of oxygen atoms in that sample.

14. Calculate the moles of a sample of $(CH_3COOH)_2Ba$ if you know that there are 1,8 moles of carbon atoms in that sample.

15. Find the number of structural units of a sample of $ZnSO_4$ if you know that the mass of this sample equals 29g.

____

____

16. Find the number of structural units of a sample of $MgCl_2$ if you know that the mass of this sample equals 48g.

____

____

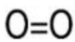
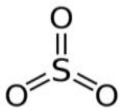
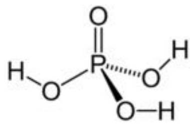
Lesson 3. Valence

Definitions:

1. **Valence** is the number of chemical bonds a given atom has formed in a given molecule. It is written in Roman numbers, which have no sign (no plus or minus).

Examples:

Structural formulas of substance and valence of atoms — Table 4

H – H Valence of pure substance oxygen equals I – single bond between atoms	 Valence of pure substance oxygen equals II – double bond between atoms	**SO_3** – sulphur trioxide Valence of : O = II S = VI 	**H_3PO_4** – phosphoric acid Valence of : H = I P = V O = II 

Valence of some elements — Table 5

Constant valence elements		Variable valence elements	
Elements	Valence	Elements	Valence
H, Li, Na, K, F	I	S	II, IV, VI
O, Mg, Ca, Ba, Zn	II	N	I, II, III, IV, V
Al, B	III	P	III, V
		Fe	II, III
		Cu	I, II
		C, Si	II, IV
		Cl, Br, I	I, III, V, VII

Valence calculation algorithm:

Step 1. Write a formula of compound and constant valences of elements	$\overset{I}{H_2}\,Se\,\overset{II}{O_4}$
Step 2. Set the unknown valence to be *x*	$\overset{I}{H_2}\,\overset{x}{Se}\,\overset{II}{O_4}$
Step 3. In compounds made from three elements including oxygen, the sum of valences of the other two elements is equal to the sum of valences of all the oxygen atoms	$2 * I + x = 4 * II$
Step 4. Solve the equation and find *x*	$2 + x = 8$ $x = 8-2$ $x = 6$ (VI)

Questions after lesson:

1. What is the meaning of valence?
2. What elements have a constant valence?
3. Why is it important to remember valences of elements?

Test

1. Choose the compound where the valence of carbon equals III:

1) H_2CO_3;
2) CO;
3) CH_4;
4) CO_2.

2. Find the group of compounds in which the valence of all nitrogen atoms equals III:

1) N_2, NO, N_2O_5;
2) HNO_3, HNO_2, NH_4OH;
3) N_2, N_2O_3, NH_3;
4) NH_3, NO_2, N_2O.

3. Choose the correct formula of P_xO_y if the valence of phosphorus equals V:

1) P_2O_5;
2) P_5O_2;
3) P_4O_{10};
4) $P_{10}O_4$;

4. What is the valence of sulphur in H_2SO_4?
1) IV;
2) V;
3) VI;
4) III.

5. Determine the valence of Na in the complex salt $Na_2[Zn(OH)]_4$:
1) I;
2) II;
3) III;
4) IV.

6. How many atoms are connected to the nitrogen atom in the HNO_3 molecule?
1) 2;
2) 3;
3) 4;
4) 5.

7. How many atoms are connected to the sulphur atom in the H_2SO_4 molecule?
1) 1;
2) 2;
3) 3;
4) 4.

8. Choose the correct formula of N_xO_y if the valence of nitrogen equals IV:
1) N_2O;
2) NO_2;
3) N_2O_3;
4) NO.

9. Determine the valence of nitrogen atoms in NH_4NO_3:
1) II, III;
2) III, IV;
3) IV, IV;
4) III, V.

10. Determine the valence of carbon atoms in CH_3COOH:
1) II, IV;
2) II, II;
3) III, IV;
4) IV, IV.

Tasks

1. Make formulas for the following compounds:

1) calcium with chlorine (I) ______________________;

2) magnesium with nitrogen (III)______________________;

3) potassium with oxygen______________________;

4) iron (III) with iodine (I)______________________;

5) silicon (IV) with chlorine (I)______________________;

6) magnesium with silicon (IV)______________________;

7) calcium with phosphorus (V)______________________;

8) silicon (IV) with oxygen______________________;

9) carbon (IV) with chlorine (I)______________________;

10) aluminum with bromine (I)______________________.

2. Determine the valences of all atoms in compounds:

1) SO_3______________________;

2) $FeSO_4$______________________;

3) $Cu(OH)_2$______________________;

4) $HClO_4$______________________;

5) $(NH_4)_2CO_3$______________________;

6) Na_2HPO_4______________________;

7) $CaOHCl$______________________;

8) $K_3[Al(OH)_6]$______________________;

9) $Na_2Cr_2O_7$______________________;

10) K_2NaPO_4______________________;

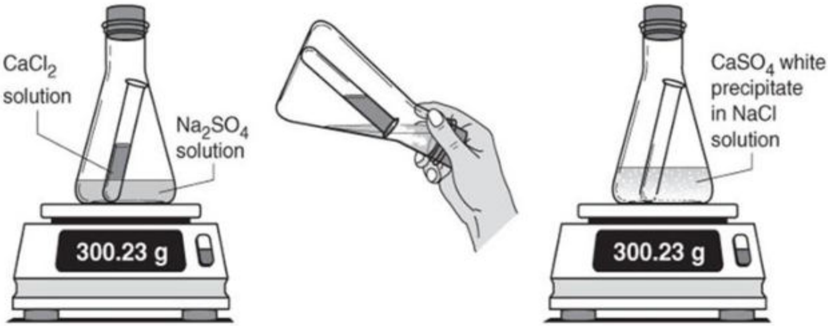
11) $Zn_3(PO_4)_2$______________________;

12) $K_3[Fe(CN)_6]$______________________.

Lesson 4. Chemical equations and their balancing

Definitions:

1. **The law of conservation of matter** is a scientific law which 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.



2. **Chemical equation** is the symbolic representation of a chemical reaction in the form of symbols and chemical formulas.
3. **Coefficient** is a number in a chemical equation indicating the number of molecules (or moles) of each substance. Therefore, you need to use coefficients so that the number of atoms of each element is the same for the reactants and products, so that the law of conservation of matter is observed.

How to balance chemical reactions?

Step 1. Count the number of atoms of each element in the left and right parts of the equation	$\overset{2}{H_2}\overset{1}{O} + \overset{2}{P_2}\overset{5}{O_5} = \overset{3}{H_3}\overset{1}{P}\overset{4}{O_4}$
Step 2. Determine which element has the number of atoms changing, find the least common multiple	$\overset{2}{H_2}\overset{1}{O} + \overset{2}{P_2}\overset{5}{O_5} = \overset{3}{H_3}\overset{1}{P}\overset{4}{O_4}$ (H)= 2*3 = 6 (P)= 2*1 = 2

Step 3. Divide the least common multiple into indices – get coefficients. Put coefficients before formulas	(H) = 6 / 2 = 3 (H) = 6 / 3 = 2 (P) = 2/ 2 = 1 (P) = 2/ 1 = 2 $3H_2O + P_2O_5 = 2H_3PO_4$
Step 4. Count the number of other atoms in the equation, repeat the steps if necessary	$3H_2O + P_2O_5 = 2H_3PO_4$ (O) = 3+5 = 8 (O) = 2*4= 8

Questions after lesson:
1. What is the main idea of the law of conservation of matter?
2. What do the chemical equations show?
3. Why is it important to arrange the coefficients in chemical equations?

Test

1. Choose the correct coefficient before the substance:

$$4P + \textbf{?}\ O_2 = 2P_2O_5$$

1) 2;
2) 3;
3) 4;
4) 5.

2. Choose the correct coefficient before the substance:

$$2Al(OH)_3 + 3H_2SO_4 = Al_2(SO_4)_3 + \textbf{?}\ H_2O$$

1) 3;
2) 6;
3) 12;
4) 14.

3. Choose the correct coefficients before the substances:

$$\textbf{?}\ Zn(NO_3)_2 + 2Al = 2Al(NO_3)_3 + \textbf{?}\ Zn\downarrow$$

1) 2, 2;
2) 3, 3;
3) 4, 4;
5) 1, 1.

4. Choose the correct coefficients before the substances:

$$\textbf{?}\ Cu(NO_3)_2 + 2Cr = \textbf{?}\ Cr(NO_3)_3 + 3Cu\downarrow$$

1) 3, 2;
2) 3, 3;

3) 2, 2;
4) 2, 3.

5. Calculate the sum of coefficients before reactants in the following chemical reaction:

$$Fe(OH)_3 + H_3PO_4 \rightarrow FePO_4\downarrow + H_2O$$

1) 4;
2) 5;
3) 2;
4) 1.

6. Calculate the sum of coefficients before reactants in the following chemical reaction:

$$Al + KOH + H_2O = K[Al(OH)_4] + H_2\uparrow$$

1) 14;
2) 12;
3) 11;
4) 10.

7. Calculate the sum of coefficients before products in the following chemical reaction:

$$H_2SO_4 + NaOH = Na_2SO_4 + H_2O$$

1) 2;
2) 3;
3) 4;
4) 5.

8. Calculate the sum of coefficients before products in the following chemical reaction:

$$K_3PO_4 + CaCl_2 = KCl + Ca_3(PO_4)_2\downarrow$$

1) 7;
2) 6;
3) 8;
4) 5.

9. Calculate the sum of all coefficients in the following chemical reaction:

$$Fe_2O_3 + H_2 = Fe\downarrow + H_2O$$

1) 7;
2) 8;
3) 9;
4) 6.

10. Calculate the sum of all coefficients in the following chemical reaction:

$$H_2SO_4 + Al_2O_3 = Al_2(SO_4)_3 + H_2O$$

1) 8;
2) 5;
3) 9;
4) 6.

Tasks

1. Balance the following chemical equations:

1) $Fe + H_2O = Fe_3O_4 + H_2\uparrow$;

2) $CaCl_2 + Na_2CO_3 = CaCO_3\downarrow + NaCl$;

3) $NaBr + Cl_2 = NaCl + Br_2$;

4) $NH_3 = N_2\uparrow + H_2\uparrow$;

5) $Na + H_2O = NaOH + H_2\uparrow$;

6) $Fe_2O_3 + Al = Fe\downarrow + Al_2O_3$;

7) $Sr(OH)_2 + HNO_3 = Sr(NO_3)_2 + H_2O$;

8) $Al_2O_3 + HCl = AlCl_3 + H_2O$;

9) $PbCl_2 + Al = Pb\downarrow + AlCl_3$;

10) $H_3PO_4 + NaOH = Na_3PO_4 + H_2O$;

11) $Na_2CO_3 + HCl = NaCl + CO_2\uparrow + H_2O$;

12) $AgNO_3 + BaCl_2 = Ba(NO_3)_2 + AgCl\downarrow$;

13) $NaHCO_3 = Na_2CO_3 + H_2O + CO_2\uparrow$;

14) $K_2SiO_3 + HCl = KCl + H_2SiO_3\downarrow$;

15) $AgNO_3 + Na_3PO4 = NaNO_3 + Ag_3PO_4\downarrow$;

16) $Na_2[Zn(OH)_4] + HCl = NaCl + ZnCl_2 + H_2O$.

17) $Mg(NO_3)_2 = MgO + NO_2\uparrow + O_2\uparrow$.

2. What is the mass of sulphur (II) oxide reacted with oxygen to produce 4,2 g of sulphur (III) oxide? Balance the equation:

$$SO_2 + O_2 = SO_3$$

____

____

____

3. What is the mass of oxygen reacted with zinc to produce 7,8 g of zinc (II) oxide? Balance the equation:

$$Zn + O_2 = ZnO$$

____

____

____

4. What is the mass of aluminum that can completely reduce 80 g of iron (III) oxide to pure iron? Balance the equation:

$$Al + Fe_2O_3 = Fe + Al_2O_3$$

5. What is the mass of barium sulphate produced in the reaction between 7 g of barium hydroxide and sodium sulfate? Balance the equation:

$$Ba(OH)_2 + Na_2SO_4 = BaSO_4\downarrow + NaOH$$

6. Find the mass of potassium sulfate formed in the reaction between a water solution containing 0,15 mol of sulphuric acid and the excess of potassium hydroxide. Balance the equation:

$$KOH + H_2SO_4 = K_2SO_4 + H_2O$$

7. How many moles of oxygen are needed to produce 1,5 g of calcium nitrate according to the following equation? Balance the equation:

$$Ca + N_2 + O_2 = Ca(NO_3)_2$$

8. What is the mass of iron (III) chloride formed in the reaction between iron and chlorine gas? The mass of iron is equal to 16,4 g, the mass of chlorine gas is equal to 14,2 g. Balance the equation:

$$Fe + Cl_2 = FeCl_3$$

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9. 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 34 g, the mass of carbon dioxide is equal to 12 g. Balance the equation:

$$CaO + CO_2 = CaCO_3\downarrow$$

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10. Find the number of moles of acetic acid that reacted with the excess of sodium hydroxide and produced 3,6 g of sodium acetate. Balance the equation:

$$NaOH + CH_3COOH = CH_3COONa + H_2O$$

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11. Find the number of moles and mass of new soluble salt that is formed by the reaction between silver nitrate and barium chloride. Mass of silver nitrate equals 75 g, mass of barium chloride equals 89 g. Balance the equation:

$$AgNO_3 + BaCl_2 = AgCl + Ba(NO_3)_2$$

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12. How many moles of oxygen are needed to produce 2,4 g of carbon dioxide during the combustion reaction of methane? Balance the equation:

$$CH_4 + O_2 = CO_2\uparrow + H_2O$$

Lesson 5. Avogadro's Law. Relative density of gases

Main characteristics of gases Table 6

Avogadro's Law is a statement that under the same conditions of temperature and pressure, equal volumes of different gases contain an equal number of molecules	1 mol = $6{,}02 \cdot 10^{23}$ molecules = 22,4 L (STP)
Molar volume is the volume occupied by one mole of any gas at STP	Vm = V / n Vm (molar volume) = 22, 4 L/mol
Volume is the space occupied by the gaseous particles at STP	V = n * Vm
Relative density of gases is a ratio between molar masses of these gases since their molar volumes are the same	ρ1 / ρ2 = (M1 / V1) / (M2 / V2) = (M1·Vm) / (M2·Vm) = M1 / M2 $D\,H_2(O_2) = M(O_2) / M(H_2) = 32 / 2 = 16$

Questions after lesson:
1. What is the main idea of Avogadro's Law?
2. What does the molar volume show?
3. What is the relative density of gases?

Test

1. Molar volume of 1 mole of CO_2 equals:
1) 22,4 L/mol;
2) 44,8 L/mol;
3) 11,2 L/mol;
4) 54 L/mol.

2. Molar mass of dry air equals:
1) 18 g/mol;
2) 36 g/mol;
3) 29 g/mol;
4) 26 g/mol.

3. Molar volume of 2 moles of H_2 equals:
1) 36 L/mol;
2) 18 L/mol;
3) 22,4 L/mol;
4) 44,8 L/mol.

4. Relative density of oxygen per dry air equals:
1) 2,15;
2) 1,1;
3) 3,05;
4) 1, 04.

5. How many moles are there in 10,5 L of nitrogen (STP)?
1) 1,1 moles;
2) 2,1 moles;
3) 0,48 moles;
4) 0,56 moles.

6. How many moles are there in 15 L of carbon monoxide (STP)?
1) 0,67 moles;
2) 0,57 moles;
4) 0,39 moles.

7. What volume is occupied by 2,4 moles of nitrogen (II) oxide (STP)?
1) 51,9 L;
2) 52,7 L;
3) 53,5 L;
4) 53,8 L.

8. What volume is occupied by 1,6 moles of dry air (STP)?
1) 35,84 L;
2) 34,84 L;
3) 36,82 L;
4) 35,82 L.

9. How many molecules are there in 7 moles of hydrogen (STP)?
1) $43{,}15 \cdot 10^{23}$;
2) $42{,}14 \cdot 10^{23}$;
3) $4{,}24 \cdot 10^{24}$;
4) $4{,}23 \cdot 10^{24}$.

10. How many molecules are there in 5 moles of chlorine gas (STP)?
1) $30{,}1 \cdot 10^{23}$;
2) $3 \cdot 10^{24}$;
3) $30{,}2 \cdot 10^{23}$;
4) $3{,}15 \cdot 10^{24}$.

Tasks

1. Calculate the relative density per helium for carbon dioxide.

2. Calculate the relative density per hydrogen for nitrogen (III) oxide.

3. What is the volume of unknown gas which has a mass equal to 3 g and a relative density per oxygen which is equal to 0,53?

4. What is the volume of unknown gas which has a mass equal to 4,5 g and a relative density per dry air which is equal to 1,51?

5. What is the relative density of unknown gas per hydrogen if its relative density per nitrogen is equal to 2,7?

6. What is the relative density of unknown gas per helium if its relative density per oxygen is equal to 0,73?

7. Find the volume of carbon dioxide (STP) produced in the reaction between 7,8 g of sodium carbonate and 6,5 g of hydrochloric acid.

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8. Find the volume of ammonia (STP) produced in the reaction between 4,5 g of ammonium nitrate and 5,9 g of sodium hydroxide.

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9. What is the volume of carbon monoxide produced from 30 L of carbon dioxide in its reaction with coal (STP)?

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10. What is the volume of nitrogen (IV) oxide produced from 11,6 L of nitrogen (II) oxide in its reaction with oxygen (STP)?

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11. Calculate the volume of ammonia (STP) if in the production reaction the volume of hydrogen equals 300 L and the volume of nitrogen equals 100 L.

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12. Calculate the volume of sulphur (IV) oxide (STP) if in the production reaction the volume of oxygen equals 15 L and mass of sulphur equals 18 g.

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13. What is the relative density of hydrochloride gas per dry air, which is produced during the reaction between hydrogen with the volume equal to 13,4 L and chlorine gas with the volume equal to 16,32 L?

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14. What is the relative density of hydrogen fluoride gas per ammonia, which is produced during the reaction between hydrogen with the volume equal to 15,75 L and fluorine with the volume equal to 12,48 L?

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Lesson 6. The periodic law. The periodic table

Definitions:

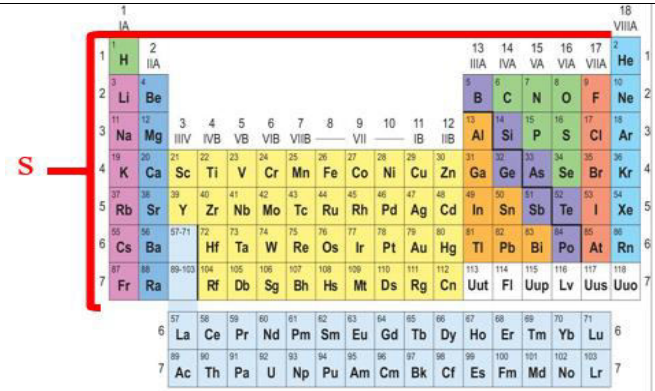
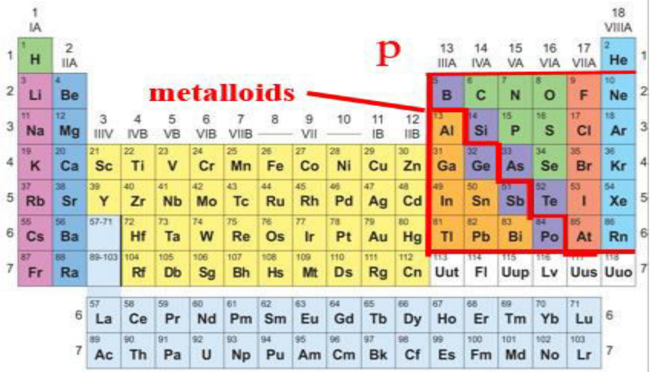
1. **The periodic law** is a law stating that the elements when arranged in the order of their atomic numbers show a periodic variation of atomic structure and of most of their properties.
2. **The periodic table** is a table of chemical elements arranged in order of atomic number, usually in rows, so that elements with similar atomic structure and similar chemical properties appears in vertical columns.
3. **Group** is a vertical column in the periodic table. Elements within the same group generally have the same electron configuration in their valence shell.
4. **Period** is a horizontal row in the periodic table. Metallic properties are decreasing from left to right, nonmetallic are increasing.
5. **Electron configuration of an atom** is the representation of the arrangement of electrons distributed among the orbital shells and subshells.
6. **Aufbau principle** states that an electron occupies the lowest energy orbital available.
7. **Pauli's exclusion principle** states that no two electrons in the same atom can have the same set of four quantum numbers. The two values of spin

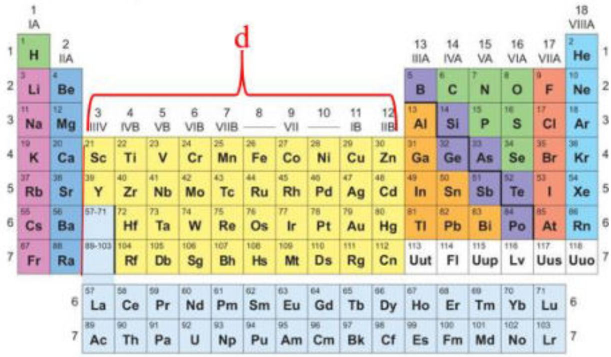
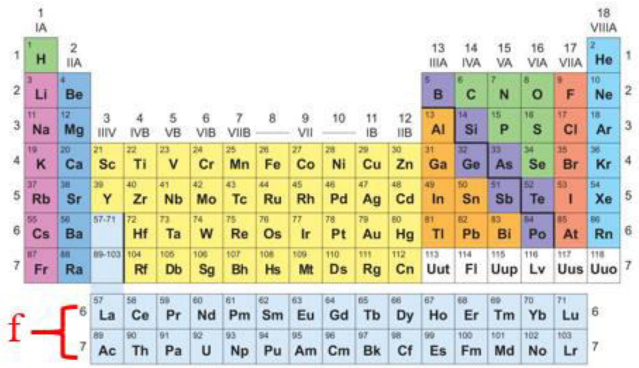
quantum number reflect the fact that for two electrons to occupy the same orbital they must have opposite spin states.

8. **Hund's rule** states that orbitals of equal energy are each occupied by one electron before any orbital is occupied by a second electron, and all electrons of a singly occupied orbital must have the same spin state.

There are four groups of elements according to the presence of electrons on the outer energy sublevel:

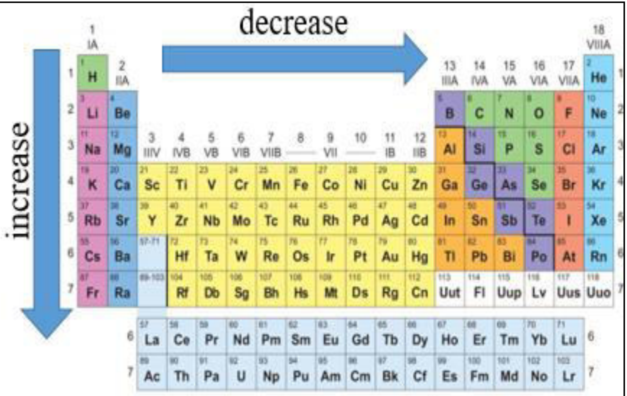
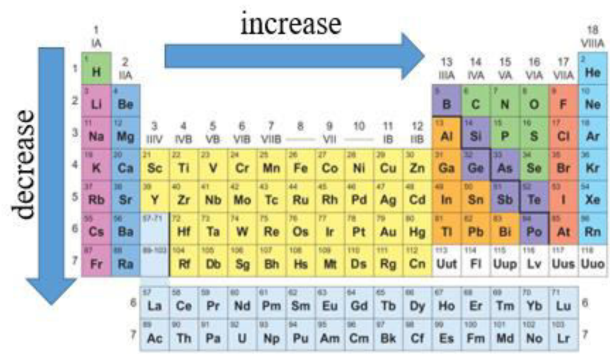
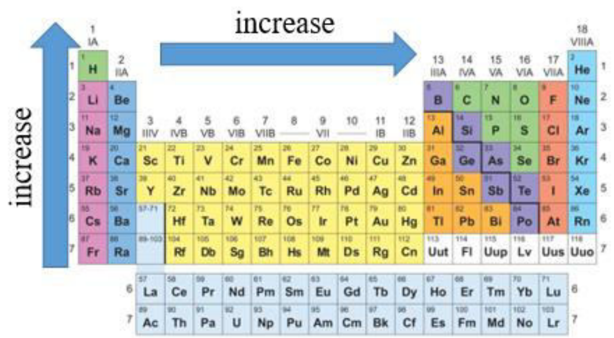
Blocks of elements — Table 7

s-elements I A group – alkali metals and hydrogen; II A group – alkaline-earth metals; Helium. These elements are active metals which is why it is difficult to find them in nature in their pure form (hydrogen and helium are exceptions).	
p-elements III A – VIII A groups except helium. The metalloids, or semiconducting elements, are located between nonmetals and metals in the p block.	

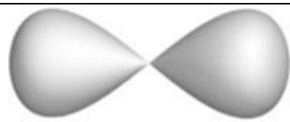
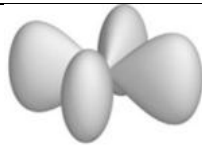
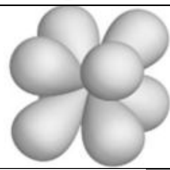
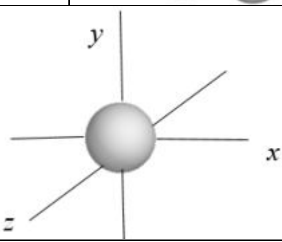
<table>
<tr>
<td>d-elements
Elements of B groups (side groups). Their position reflects the fact that they involve the filling of the 3d sublevel.</td>
<td>
</td>
</tr>
<tr>
<td>f-elements
In the periodic table, the f-block elements are wedged between Groups III and IV in the sixth and seventh periods. Their position reflects the fact that they involve the filling of the 4f sublevel.</td>
<td>
</td>
</tr>
</table>

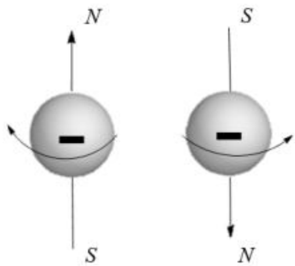
Some properties of atoms are also change according to the periodic law:

Changes in some properties of the atom Table 8

Atomic radius: • for metals, it is half the distance between adjacent nuclei in a crystal of the element; • for elements that occur as molecules, the atomic radius is half the distance between nuclei of identical atoms that are chemically bonded together.	
Electronegativity is the atom's ability to attract electrons in a chemical bond	
The first ionization energy is the energy required to remove one electron from a neutral gaseous atom of an element	

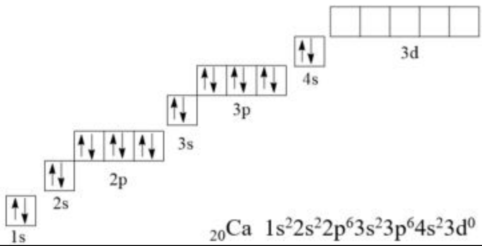
Quantum numbers – characteristics of an electron Table 9

1. **The principal quantum number** (**n**) determines the energy of the orbital and the number of the energy level. Electrons in the same atom that have the same principal quantum number occupy the same electron shell of the atom. The principal quantum number can be any nonzero positive integer. The number of energy sublevels at the energy level is equal to the principal quantum number.	First energy level n = **1** has only one sublevel 1s; Second energy level n = **2** has two sublevels 2s 2p; Third energy level n = **3** has three sublevels 3s 3p 3d; Fourth energy level n = **4** has four sublevels 4s 4p 4d 4f, etc.	
2. **The angular momentum quantum number** (**l**) determines the shape of the orbital, it can take values from zero to **n**-1.	l = 0 – s orbital	
	l = 1 – p orbital	
	l = 2 – d orbital	
	l = 3 – f orbital	
3. **Magnetic quantum number** (**m_l**) determines the orientation of an orbital around the nucleus, it can take values from **– l** to **+ l**.		

4. **Spin quantum number** (**m_s**) determines the orientation of electrons in the quantum cells, it has only two allowed values: + ½ ,- ½	

How to write the electron configuration of an atom?

Step 1. Find your atom's atomic number. This number equals the number of protons and electrons in a neutral atom	$_{20}Ca$
Step 2. The period shows us the number of energy levels.	The position of calcium in the fourth period means calcium has 4 energy levels.
Step 3. The group number shows us the number of electrons on the outer layer.	Its position in the II A group means calcium has 2 electrons on the outer layer
Step 4. Start to write the electron configuration. On the first energy level there is only one sublevel	$1s^2$
Step 5. On the second energy level there are 2 sublevels. In contrast to the s-orbital, p-orbitals have 3 quantum cells	$2s^22p^6$
Step 6. Starting from the third period of the periodic system, a 3d sublevel with 5 quantum cells appears. But 4s sublevel is lower than 3d in energy. Therefore, we add to the 4s 2 electrons then come back to the 3d until it is filled and missing only 2 electrons. (Aufbau principle)	$3s^23p^64s^23d^0$

Step 7. Write the electron configuration of the atom, you can also write a diagram with cells and arrows.	

Questions after lesson:

1. What is the main idea of the periodic law?
2. What is the structure of the periodic table?
3. What are the main characteristics of an electron?
4. What are the main blocks of elements, which can be found in the periodic law?
5. What rules are used to write the electron configuration of an atom?

Test

1. An eight-electron outer shell has an ion:
1) P^{3+};
2) S^{2-};
3) C^{4+};
4) Fe^{2+}.

2. The ion has a two-electron on the outer shell:
1) S^{6+};
2) S^{2-};
3) Br^{5+};
4) Sn^{4+}.

3. Number of electrons in an iron ion Fe^{2+} equals:
1) 54;
2) 28;
3) 58;
4) 24.

4. Electronic configuration $1s^22s^22p^63s^23p^6$ corresponds to the ion:
1) Sn^{2+};
2) S^{2-};
3) Cr^{3+};
4) Fe^{2+}.

5. In the normal state, an atom has three unpaired electrons:
1) Si;
2) P;
3) S;
4) Cl.

6. Element with electronic configuration of the outer level... $3s^2 3p^3$ forms a hydrogen compound:
1) EH_4;
2) EH;
3) EH_3;
4) EH_2.

7. Electronic configuration $1s^2 2s^2 2p^6 3s^2 3p^6$ corresponds to the ion:
1) Cl^-;
2) N^{3-};
3) Br^-;
4) O^{2-}.

8. Electronic configuration $1s^2 2s^2 2p^6$ corresponds to the ion:
1) Al^{3+};
2) Fe^{3+};
3) Zn^{2+};
4) Cr^{3+}.

9. The same electronic configuration of the outer level has Ca^{2+} and:
1) K^+;
2) Ar;
3) Ba;
4) F^-.

10. A metal atom whose highest oxide is Me_2O_3, has an electronic formula for the outer energy level:
1) $ns2np^1$;
2) ns^2np^2;
3) ns_2np_3;
4) ns^2np^4;

11. Element whose highest oxide is R_2O_7 has an electronic configuration of the external level:
1) ns^2np^3;
2) ns^2np^5;
3) ns^2np^1;
4) ns^2np^2.

12. The highest oxide E_2O_7 produced by the element in the atom of which energy levels with electrons corresponds to a line of numbers:
1) 2, 8, 1;
2) 2, 8, 7;
3) 2, 8, 8, 1;
4) 2, 5.

13. Electronic configuration $1s^2 2s^2 2p^6 3.s^2 3p^6 3d^1$ has ion:
1) Ca^{2+};
2) A^{3+};
3) K^+;
4) Sc^{2+}.

14. The sulphur atom has the number of electrons at the outer energy level and the charge of the nucleus are equal:
1) 4, +16;
2) 6, +32;
3) 6, +16;
4) 4, +32.

15. The number of valence electrons (electrons of outer sublevel) in manganese is:
1) 1;
2) 3;
3) 5;
4) 7.

16. Choose particles with the same electronic structure:
1) Na^0, Na^+;
2) Na^0, K^0;
3) Na^+, F^-;
4) Cr^{2+},Cr^{3+}.

17. The highest oxide of the EO_3 forms an element with the electronic configuration of the outer electron layer:
1) ns^2np^1;
2) ns^2np^3;
3) ns^2np^4;
4) ns^2np^6.

18. The number of energy layers and the number of electrons in the outer energy layer of arsenic atoms are equal:
1) 4, 6;
2) 2, 5;

3) 3, 7;
4) 4, 5.

19. What is the electronic configuration of the most active metal?
1) $1s^22s^22p^1$;
2) $1s^22s^22p^63s^1$;
3) $1s^22s^2$;
4) $1s^22s^22p^63s^23p^1$.

20. The number of electrons in the atom is determined by:
1) the number of protons;
2) the number of neutrons;
3) the number of energy levels;
4) the value of the relative atomic mass.

21. Nucleus ^{81}Br has:
1) 81p, 35n;
2) 35p, 46n;
3) 46p, 81n;
4) 46p, 35n.

22. Ion which contains 16 protons and 18 electrons, has a charge of:
1) +4;
2) -2;
3) +2;
4) -4.

23. Electronic configuration $1s^22s^22p^63s^23p^64s^1$ has an atom:
1) Li;
2) Na;
3) K;
4) Ca.

24. The number of protons and neutrons contained in the nucleus of an isotope ^{40}K equally:
1) 19, 40;
2) 21, 19;
3) 20, 40;
4) 19, 21.

25. Choose the chemical element, one of whose isotopes has a mass number of 44 and contains 24 neutrons in the nucleus:
1) Cr;
2) Ca;

3) Ru;
4) Sc.

Tasks

1. Find out numbers of period and group for Cu ______________________.
2. Find out numbers of period and group for Se ______________________.
3. Arrange these elements in the order of the increase of their metallic properties (S / Na / Cu / Si / P / C):

____.

4. Arrange these elements in the order of the increase of their nonmetallic properties (F / Ca / O / N / Li / Al):

____.

5. Write the formulas of the highest oxides of elements from the VA group of the periodic table starting from the 2ed period:

____.

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

____.

7. Calculate the number of neutrons in $_{37}Cl$ isotope.

____

8. Calculate the number of neutrons in $_{14}C$ isotope.

____

9. Calculate the number of protons in 17 g of hydrochloric acid.

____

10. Calculate the number of protons in 30 g of sodium hydroxide.

____

11. Calculate the atomic mass of an element that has three isotopes with the following atomic masses and corresponding abundances: 283 u (82 %), 278 u (2 %), 284 u (16 %).

____

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____

12. The percent of $_3H$ in the sample of H_2O is 3 %, the percent of $_1H$ is 97 %. Find the volume of hydrogen produced in the reaction between 18 g of that H_2O sample with sodium.

13. Write the complete electron configuration for calcium.

14. Write the complete electron configuration for bromine.

15. Write the short electron configuration for silver.

16. Write the short electron configuration for strontium.

17. Arrange the atoms from this line (He / Fe / Se / P / Na / Mg) in order of the increase of the number of unpaired electrons.

18. Draw the diagram with cells and arrows for the outer shell of manganese:

____

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____.

19. Draw the diagram with cells and arrows for the outer shell of vanadium:

____

____

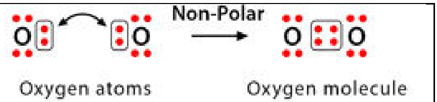
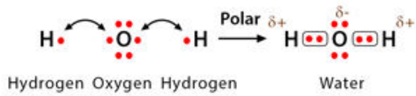
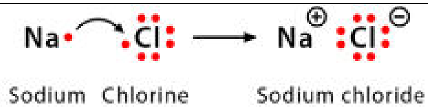
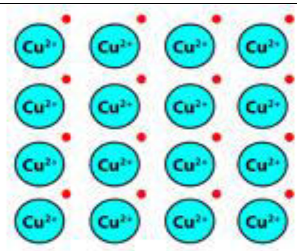
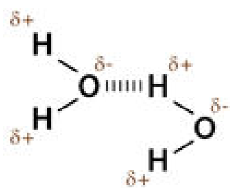
____.

Lesson 7. Chemical bonds

Definitions:

1. **Chemical bonds** are forces that hold the atoms together in a molecule. They are a result of strong intramolecular interactions among the atoms of a molecule.
2. **Ion** is an atom or group of atoms that has an electrical charge because it has added or lost one or more electrons.
3. **Electron gas** is the set of free electrons in the crystal lattice of a metal.

Classification of chemical bonds Table 10

<table>
<tr><td rowspan="2">Covalent</td><td>Nonpolar</td><td rowspan="2">Formed by sharing electron pairs between atoms of nonmetals</td><td></td></tr>
<tr><td>Polar</td><td></td></tr>
<tr><td colspan="2">Ionic</td><td>Formed by transferring electrons from metal to nonmetal</td><td></td></tr>
<tr><td>Metallic</td><td colspan="2">Formed by electron cloud around atoms of metals</td><td></td></tr>
<tr><td>Hydrogen</td><td colspan="2">Is electromagnetic attractive interaction of a polar hydrogen atom in a molecule or chemical group and an electronegative atom</td><td></td></tr>
</table>

Questions after lesson:

1. What is a chemical bond?
2. What types of particles are called ions?
3. What are the characteristics of a covalent bond?
4. What are the differences between metallic and ionic bonds?
5. What is the main characteristic of a hydrogen bond?

Test

1. Choose the type of chemical bond in the water molecule:
1) covalent nonpolar;
2) ionic;
3) covalent polar;
4) metal.

2. A covalent nonpolar bond is formed between:
1) atoms of different nonmetals;
2) atoms of the same element's nonmetals;
3) active metals and nonmetals;
4) metals.

3. Choose the correct statements:
1) by connection with each other, atoms strive to achieve a stable state;
2) the reason of the inertia of helium and neon is the gaseous state;
3) a covalent bond occurs between the atoms of active metals and nonmetals;
4) an ionic bond is formed between potassium and chlorine atoms.

4. Find compounds with an ionic bond:
1) $NaCl$;
2) SO_3;
3) $CaCO_3$;
4) NO.

5. Choose compounds with a covalent polar bond:
1) Fe_2O_3;
2) H_2O_2;
3) $AlPO_4$;
4) CO.

6. Find compounds with a metallic bond:
1) $AgCl$;
2) $Zn(OH)_2$;
3) Cr;
5) $CuAl_{11}Fe_4$.

7. Choose compounds containing a double bond:
1) O_2;
2) N_2;
3) C_2H_4;
4) C_2H_2.

8. Find a compound with the most polar covalent bond:
1) H_2O;
2) H_2S;
3) HCl;
4) HF.

9. Choose compounds with both ionic and covalent polar bonds:
1) Na_2SO_4;
2) K_2S;
3) $AgCl$;
4) $Ca(NO_3)_2$.

10. What is the type of chemical bond in steel?
1) ionic;
2) hydrogen;
3) metallic;
4) covalent nonpolar.

Tasks

1. Arrange substances in this line (NH_3 / CH_4 / H_2S / H_2O / HCl / HF) in order of the increase of the polarity of a covalent bond.

____

____

2. Describe all types of the chemical bonds in the following compounds:

1) Mn ____;

2) $Fe(OH)_3$____;

3) H_3PO_4____;

4) $Al_2(SO_4)_3$____;

5) CH3COOH____.

3. Find the mass of the ionic substance which was produced in the reaction between 10 g of NaOH and 8,9 g of hydrochloric acid.

4. Find the volume of the gas which was produced in the reaction between 12 L of O_2 and 24,47 g of the coal.

Lesson 8. Oxidation state

Definitions:

1. **Oxidation state** is an indicator of oxidation (loss of electrons) 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 matter, oxygen and hydrogen Table 11

Atom or ion	**Oxidation state**	**Example**
Pure matter	0	Na, C, O_2, H_2, Al
Hydrogen	+1	HCl, NaOH, NH_3
Hydrogen with metals	-1	NaH, KH, CuH_2
Oxygen	-2	CO_2, N_2O_5, H_2O
Oxygen in peroxides	-1	H_2O_2, Na_2O_2, K_2O_2

Oxidation state of some metals Table 12

IA group - Alkali metals	**IIA group – Alkaline-earth metals**	**Others**	
Li, Na, K, Rb, Cs, Fr	Be, Mg, Ca, Sr, Ba	Zn	Al
+1	+2	+2	+3

Maximal and minimal oxidation state of nonmetals Table 13

Element	Max O.S.	Min O.S.
P	VA group = +5 $\overset{+5}{H_3PO_4}$	8-V = 3 $\overset{-3}{PH_3}$
C	IVA group = +4 $\overset{+4}{CO_2}$	8-IV= 4 $\overset{-4}{CH_4}$

How to find the oxidation state of an atom in a chemical compound?

Step 1. Find the oxidation state of the atoms with constant value	$\overset{+1}{H}N\overset{-2}{O_3}$
Step 2. Set the unknown oxidation state to be x	$\overset{+1}{H}\overset{x}{N}\overset{-2}{O_3}$
Step 3. The sum of the oxidation states of atoms should be equal to zero because the molecule is neutral	$+1+x+(-2*(3)) = 0$
Step 4. Solve the equation and find x	$+1+x+(-6) = 0$ $x-5=0$ $x=+5$

Questions after lesson:

1. What shows the oxidation state?
2. Which elements have a constant oxidation state?
3. How can you find the maximal oxidation state of an element?
4. How can you find the minimal oxidation state of an element?

Test

1. The highest oxidation state of manganese is shown in the compound:
1) $KMnO_4$;
2) MnO_2;
3) K_2MnO_4;
4) $MnSO_4$.

2. Choose the compound in which the oxidation state of phosphorus equals -3:
1) PH_3;
2) P_2O_3;
3) NaH_2PO_4;
4) H_3PO_4.

3. Nitrogen has the same oxidation state in the substances:
1) N_2O_5, HNO_3, $NaNO_3$;
2) NO_2, HNO_3, KNO_3;
3) NO, NO_2, N_2O_3;
4) HNO_3, HNO_2, NO_2;

4. In the order of increasing electronegativity, the elements are arranged in a row:
1) O-N-C-B;
2) Si-Ge-Sn-Pb;
3) Li-Na-K-Rb;
4) Sb-P-S-Cl.

5. The oxidation state of chlorine in the compound $Ca(C1O)_2$ equals:
1) +1;
2) +3;
3) +5;
4) +7.

6. Make correlations between formulas of substance and the oxidation state of nitrogen:

Formulas:	Oxidation state of nitrogen:
A) NF_3;	1) -3;
B) $H_2N_2O_2$;	2) +1;
C) NH_4HCO_3;	3) +2;
D) $Ca(NO_2)_2$.	4) +3;
	5) +4;
	6) +5.

7. Make correlations between formulas of substance and the oxidation state of sulphur:

Formulas:	Oxidation state of sulphur:
A) $K_2S_2O_7$;	1) -2;
B) $NaHSO_3$;	2) -1;
C) SO_2Cl_2;	3) +1;
D) SO_2.	4) +4;
	5) +5;
	6) +6.

8. Make correlations between formulas of substance and the oxidation state of chrome:

Formulas:	Oxidation state of sulphur:
A) K_2CrO_4;	1) 0;
B) $CaCr_2O_7$;	2) +2;
C) CrO_2P_2;	3) +3;
D) $Ba_3[Cr(OH)_6]_2$.	4) +4;
	5) +5;
	6) +6.

9. Chlorine shows a positive oxidation state when combined with:
1) S;
2) H_2;
3) O_2;
4) Fe.

10. Choose the compound in which the oxidation state of nitrogen equals +3:
1) NH_4C1;
2) $NaNO_3$;
3) N_2O_4;
4) KNO_2.

11. Choose the compound with oxidation state +2, valence IV of carbon atom:
1) CO;
2) CO_2;
3) HCOOH;
4) CH_2Cl_2.

12. In which compounds do the oxidation states of chemical elements equal -3 and +1?

1) NF_3;
2) PH_3;
3) N_3O_3;
4) $AlCl_3$.

13. In which compound does the oxidation state of nitrogen equal +3:

1) Na_3N;
2) NH_3;
3) NH_3Cl;
4) HNO_2.

14. Choose the compound with oxidation state +4, valence IV of carbon atom:

1) CH_4;
2) CO;
3) H2CO_3;
4) Al_4C_3.

15. The nitrogen atom shows a valence other than III in the molecule:

1) HNO_3;
2) HNO_2;
3) NF_3;
4) NH_3.

16. Choose the compound with a positive oxidation state of the oxygen atom:

1) H_2O;
2) H_2O_2;
3) F_2O;
4) Fe_3O_4.

17. In which compound does the valence of chlorine equal VII?

1) HClO;
2) $ZnCl_2$;
3) $NaClO_3$;
4) $HClO_4$.

18. In which compound does the oxidation state of chlorine equal +1:

1) HClO;
2) $CaCl_2$;
3) CCl_4;
4) $Ca(ClO_2)_2$.

19. Nitrogen and carbon atoms have the same oxidation states in the compounds:

1) NH_3 / CO;
2) NO_2 / CCl_4;
3) N_2O_3 / CO_2;
4) Na_3N / CH_4.

Tasks

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

H_3PO_4 / $KMnO_4$ / $NaHCO_3$ / $Al_2(SO_4)_3$ / $NaBrO_4$ / HCOH

____
____
____.

2. Write oxidation states for all the elements in the following ions:

NO_3^- / ClO_4^- / $H_2PO_4^-$ / SO_4^{2-} / PO_4^{3-} / CH_3COO^-

____
____
____.

3. Arrange compounds in the line in order of the increase of the oxidation state of a metal:

Na_2O / $CaCl_2$ / CrO_3 / $Al_2(SO_4)_3$ / Mn_2O_7

____
____
____.

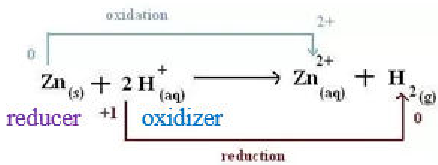
Lesson 9. Chemical reactions

Definitions:

1. **Chemical reaction** is a process in which one or more substances are converted to one or more different substances.

Classification of chemical reactions Table 14

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	2HCl (compound **1**) + Zn (element **1**) → $ZnCl_2$ (compound **2**) + H_2 (element **2**)
A **double-replacement** reaction occurs when parts of two compounds are exchanged, making two new compounds	$Cu\mathbf{Cl_2} + 2Ag\mathbf{NO_3} \rightarrow Cu(\mathbf{NO_3})_2 + 2Ag\mathbf{Cl}$
A **composition** reaction is a chemical reaction in which a single substance is produced from multiple reactants	$2H_2$ (reactant **1**) + O_2 (reactant **2**) → $2H_2O$ (a single substance)
A **decomposition** reaction starts from a single substance and produces more than one substance	$CaCO_3$ (a **single** substance) → CaO (product 1) + CO_2 (product 2)
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.	$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$
Exothermic reaction – is a chemical reaction that releases energy in the form of heat	$C_6H_{12}O_6 + 6\ O_2 \rightarrow 6\ CO_2 + 6\ H_2O$ +**ATP**

Endothermic reaction – is a chemical reaction in which the system absorbs energy from its surroundings in the form of heat	$CaCO_3 \rightarrow CaO + CO_2 - \mathbf{Q}$
A **reversible** reaction is a chemical reaction that results in an equilibrium mixture of reactants and products	$N_2 + H_2 \rightleftharpoons 2NH_3$
An irreversible reaction is one where the reactants react to form products that cannot revert back into reactants	$H_2 + O_2 \rightarrow H_2O$
Redox reaction – is a chemical reaction in which the atom changes its oxidation state	

Questions after lesson:

1. What types of chemical reactions do you know?
2. What are the differences between exothermic and endothermic chemical reactions?
3. What changes of atoms occur during the redox reaction?

Test

1. Choose the single-replacement chemical reaction:
1) $2H_2 + O_2 = 2H_2O$;
2) $CuSO_4 + Fe = FeSO_4 + Cu\downarrow$;
3) $NaOH + HCl = NaCl + H_2O$;
4) $C + O_2 = CO_2$.

2. Choose the double-replacement chemical reaction:
1) $2Na + 2H_2O = 2NaOH + H_2\uparrow$;
2) $2KNO_3 = 2KNO_2 + O_2\uparrow$;
3) $K_2O + SO_3 = K_2SO_4$;
4) $BaCl_2 + Na_2SO_4 = BaSO_4\downarrow + 2NaCl$.

3. Choose characteristics of the following reaction:

$$H_2S + O_2 = H_2O + SO_2\uparrow$$

1) redox reaction;
2) combustion reaction;
3) composition reaction;
4) endothermic reaction.

4. Make correlation between chemical reactions and their types:

Reactions:	Types:
A) $Si + O_2 = SiO_2$;	1) decomposition;
B) $H_2SO_4 + NaOH = Na_2SO_4 + H_2O$;	2) single-replacement;
C) $CuCl_2 + Zn = ZnCl_2 + Cu\downarrow$;	3) double-replacement;
D) $NH_4NO_2 = N_2\uparrow + 2H_2O$.	4) composition.

5. Choose characteristics of the following reaction:

$$N_2 + O_2 = 2NO - Q$$

1) exothermic reaction;
2) endothermic reaction;
3) combustion reaction;
4) decomposition reaction.

6. Which of the following reactions are reversible?
1) $CaCO_3 = CaO + CO_2\uparrow$;
2) $2SO_2 + O_2 = 2SO_3$;
3) $N_2 + 3H_2 = 2NH_3\uparrow$;
4) $Na_2O + H_2O = 2NaOH$.

7. Choose redox reactions:
1) $P_2O_5 + 3H_2O = 2H_3PO_4$;
2) $K_2O + CO_2 = K_2CO_3$;
3) $CO_2 + C = 2CO\uparrow$;
4) $2KI + Br_2 = 2KBr + I_2$.

8. Choose combustion reactions:
1) $CH_4 + 2O_2 = CO_2\uparrow + 2H_2O$;
2) $CaO + H_2O = Ca(OH)_2$;
3) $4NH_3 + 5O_2 = 4NO\uparrow + 6H_2O$;
4) $AgNO_3 + KCl = AgCl\downarrow + KNO_3$.

9. Choose characteristics of the following reaction:

$$Ca(HCO_3)_2 = CaO + 2CO_2\uparrow + H_2O$$

1) redox reaction;
2) decomposition reaction;
3) combustion reaction;

4) single-replacement.

10. Which of the following reactions are exothermic?
1) $H_2SO_4 + 2KOH = K_2SO_4 + 2H_2O$;
2) $CaCO_3 = CaO + CO_2\uparrow$;
3) $N_2 + O_2 = 2NO$;
4) $4Fe + 3O_2 = 2Fe_2O_3$.

Tasks

1. Write 3 samples of composition reaction for nitrides:

____

____

____.

2. Write 3 samples of single displacement reaction with H_2CO_3:

____

____

____.

3. Finish chemical reactions and classify them:

1) $Al_2O_3 + HCl$ = ____;

2) $Fe + H_3PO_4$ = ____;

3) $Cu(OH)_2 + H_2SO_4$ = ______________________________________;

4) $Na_2O + N_2O_5$ = ____;

5) $Cu_2S + HNO_3$ = ____;

6) $Ni(NO3)_2 + Zn$ = _______________________________________;

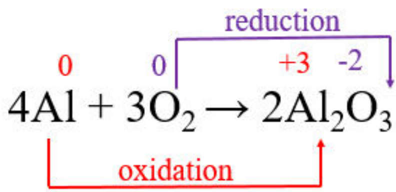
7) $C_6H_6 + H_2O$ = ____.

Lesson 10. Redox reaction

Definition

1. **Reduction** is decreasing the oxidation state of an atom through a chemical reaction.

2. **Oxidation** is increasing the oxidation state of an atom through a chemical reaction.

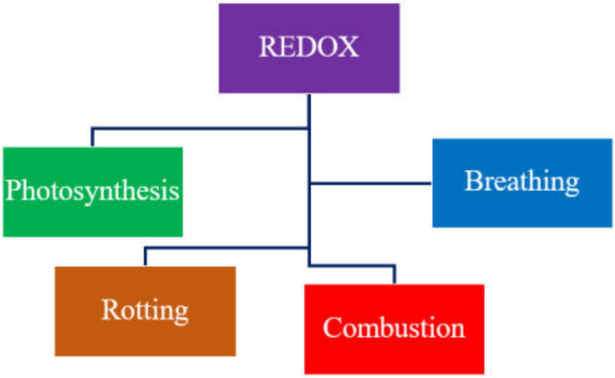


Scheme 2. Reduction and oxidation

How to balance redox reactions?

Step 1. Calculate oxidation states	$\overset{+1}{K}\overset{+7}{Mn}\overset{-2}{O_4} + \overset{+1}{H}\overset{-1}{Cl} \rightarrow \overset{+1}{K}\overset{-1}{Cl} + \overset{+2}{Mn}\overset{-1}{Cl_2} + \overset{0}{Cl_2} + \overset{+1}{H_2}\overset{-2}{O}$
Step 2. Notice those atoms for which the oxidation state has been changed	$\overset{+1}{K}\underline{\overset{+7}{Mn}}\overset{-2}{O_4} + \overset{+1}{H}\underline{\overset{-1}{Cl}} \rightarrow \overset{+1}{K}\overset{-1}{Cl} + \underline{\overset{+2}{Mn}}\overset{-1}{Cl_2} + \underline{\overset{0}{Cl_2}} + \overset{+1}{H_2}\overset{-2}{O}$
Step 3. Write down two half reactions	$\overset{+7}{Mn} \xrightarrow{+5\,e} \overset{+2}{Mn}$ reduction 5 e 2 oxidizer * 10 $2\overset{-1}{Cl} \xrightarrow{-2\,e} \overset{0}{Cl_2}$ oxidation 2 e 5 reducer
Step 4. Put coefficients before oxidizer, product of reduction and product of oxidation	$2\ KMnO_4 + HCl \rightarrow KCl + 2\ MnCl_2 + 5\ Cl_2 + H_2O$

Step 5. Find other coefficients	$2\ KMnO_4 + 16\ HCl \rightarrow 2\ KCl + 2\ MnCl_2 + 5\ Cl_2 + 8\ H_2O$
Step 6. Check by oxygen	$2\ KMnO_4 + 16\ HCl \rightarrow 2\ KCl + 2\ MnCl_2 + 5\ Cl_2 + 8\ H_2O$ 2*4 = 8 8 = 8 8*1 =8



Scheme 3. The main redox processes

Questions after lesson:

1. What is the point of the process of oxidation?
2. What is the point of the process of reduction?
3. Why is it important to use the half reaction method to balance redox reactions?

Tasks

1. Balance the following redox reactions:

1) $Si + HNO_3 + HF = H_2SiF_6 + NO + H_2O$

____

____

____;

2) $B + HNO_3 + HF = HBF_4 + NO_2 + H_2O$

____

____

____;

3) $PH_3 + AgNO_3 + H_2O = Ag + H_3PO_4 + HNO_3$

____

____

____;

4) $KNO_2 + HI + H_2SO_4 = I_2 + 2NO + K_2SO_4 + H_2O$

____

____

____;

5) $KMnO_4 + MnSO_4 + H_2O = MnO_2 + K_2SO_4 + H_2SO_4$

____

____

____;

6) $KMnO_4 + NH_3 = MnO_2 + N_2 + KOH + H_2O$

____

____

____;

7) $NO_2 + P_2O_3 + KOH \rightarrow NO + K_2HPO_4 + H_2O$

____

____

____;

8) $Fe(OH)_3 + Br_2 + KOH = K_2FeO_4 + KBr + H_2O$

____

____

____;

9) $KIO_3 + KI + H_2SO_4 \rightarrow I_2 + K_2SO_4 + H_2O$

____

____

____;

10) $KClO_3 + CrCl_3 + KOH \rightarrow K_2CrO_4 + KCl + H_2O$

Lesson 11. Chemical equilibrium

Definitions:

1. **Chemical equilibrium** is the state in which both reactants and products are present at concentrations which have no further tendency to change with time.

2. **Le Chatelier's principle** states that if a dynamic equilibrium is disturbed by changing the conditions, the position of equilibrium moves to counteract this change.

Factors affecting chemical equilibrium — Table 15

Factor	Changes
Concentration	$A + 2B \rightleftharpoons C + D$ ↑ concentration A – equilibrium moves to the right ↓ concentration A – equilibrium moves to the left ↑ concentration C – equilibrium moves to the left ↓ concentration C – equilibrium moves to the right
Pressure – only for gases	$\underbrace{A_{(g)} + 2B_{(g)}}_{3\ \text{moles}} \rightleftharpoons \underbrace{C_{(g)} + D_{(g)}}_{2\ \text{moles}}$ ↑ pressure – equilibrium moves to the right, the side with less volume of gases ↓ pressure – equilibrium moves to the left, the side with more volume of gases

Temperature	$A + 2B \rightleftharpoons C + 250\ kJ/mol$ $C \rightleftharpoons A + 2B - 250\ kJ/mol$ ↑ temperature – equilibrium moves to the left to the endothermic reaction  ↓ temperature – equilibrium moves to the right to the exothermic reaction 

Questions after lesson:

1. What is the chemical equilibrium?
2. What factors have an effect on the chemical equilibrium?
3. What is the main idea of Le Chatelier's principle?

Test

1. What effect can move the chemical equilibrium to the side of products in the following chemical equation:

$$FeO_{(s)} + H_{2(g)} \leftrightarrow Fe_{(s)} + H_2O_{(g)} - 23\ kJ/mol$$

1) increase in pressure;
2) increase in temperature;
3) decrease in temperature;
4) decrease in pressure.

2. Chemical equilibrium move to the side of products in the following chemical equation by:

$$2NO_{(g)} + O_{2(g)} \leftrightarrow 2NO_{2(g)} + Q$$

1) increase in pressure;
2) increase in temperature;
3) decrease in temperature;
4) decrease in pressure.

3. A change in pressure moves the chemical equilibrium in the system:

1) $3H_{2(g)} + N_{2(g)} \leftrightarrow 2NH3_{(g)}$;
2) $H_{2(g)} + S_{(s)} \leftrightarrow H_2S_{(g)}$;
3) $N_{2(g)} + O_{2(g)} \leftrightarrow 2NO_{(g)}$;
4) $H_{2(g)} + Cl_{2(g)} \leftrightarrow 2HCl_{(g)}$.

4. What effect can move the chemical equilibrium to the side of products in the following chemical equation:

$$CO_{(g)} + 2H_{2(g)} \leftrightarrow CH_3OH_{(g)} + Q$$

1) use catalyst;
2) decrease in temperature;
3) increase in temperature;
4) increase in pressure.

5. What effect can move the chemical equilibrium to the side of products in the following chemical equation:

$$2NOCl_{(g)} \leftrightarrow 2NO_{(g)} + Cl_{2(g)} + Q$$

1) decrease in temperature;
2) increase in pressure;
3) increase in pressure;
4) increase in temperature.

Lesson 12. Chemical reaction rate

Definitions:

1. **Chemical reaction rate** is the number of elementary acts between molecules or other particles of substance per unit time in a unit volume.
2. **Homogeneous chemical reaction** is the rate change of concentration of reactant or product per unit of time.
3. **Catalysts** are substances that affect the reaction rate but retain their chemical composition.

Chemical reaction rate for different types of system Table 16

V (r) - for a homogeneous reaction Reactants are in the same state of aggregation or phase.	**V(r) - for a heterogeneous reaction Reactants are in different state of aggregation or in different phases.**
$V(r)= \Delta C/\Delta t$ $\Delta C = C_2 - C_1$ (molar concentrations of reacting or forming substances) $t = t_2 - t_1$ (point in time) Reaction rate unit - mol / l · s	**$V(r)= \Delta C/\Delta t \cdot S$** S – contact area of reacting substances Reaction rate unit- mol / m^2 · s *using the above formulas, it is possible to calculate only a certain average rate of a given reaction in a selected time interval (after all, for most reactions, the rate decreases as they proceed)*

Factors affecting chemical reaction rate Table 17

State of matter	Gases react more readily than liquids, which react more readily than solids.
Concentration	The higher the concentration of reacting substances, the greater the rate of a chemical reaction. **The law of the masses** (N.I. Beketov) The speed of a chemical reaction is directly proportional to the product of the concentrations of the reactants. 2A + 3B = 2D $V(r) = k^* C[A]^2 {}^*C[B]^3$, k – rate constant.
Pressure (for gases)	Increasing pressure increases reaction rate.
Catalysts	Catalysts reduce the activation energy, which leads to an increase in active molecules, the reaction rate increases.
Temperature	With an increase in temperature for every **10** degrees, the reaction rate increases by **2-4** times (**Van Goff Rule**). A number showing how many times the reaction rate increases with a temperature increase of 10 °C is called the **temperature coefficient**: $$Q_{10} = \left(\frac{r_2}{r_1}\right)^{\frac{10}{(T_2 - T_1)}}$$ 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.

Questions after lesson:

1. What is the chemical reaction rate?
2. What are the differences between homogeneous and heterogeneous reactions?
3. What factors have an effect on the reaction rate?

Test

1. Find the substances that have the maximal chemical reaction rate between each other at room temperature:

1) Zn / H_2SO_4;

2) Na / H_2O;
3) Fe / O_2;
4) $CuSO_4$ (solution) / KOH (solution).

2. Find the substances that have the minimal chemical reaction rate between each other at room temperature:
1) Na / H_2O;
2) $CuSO_4$ (solution) / KOH (solution);
3) Fe / O_2;
4) Zn / H_2SO_4.

3. Choose the metal that reacts with HCl acid proceeding at the lowest chemical reaction rate at room temperature:
1) Zn;
2) Mg;
3) Pb;
4) Fe.

4. Choose the metal that reacts with HCl acid proceeding at the highest chemical reaction rate at room temperature:
1) Mg;
2) Pb;
3) Zn;
4) Fe.

5. What are the type of substances that delay chemical reaction rate?
1) enzymes;
2) catalysts;
3) inhibitors;
4) oxidants.

6. Choose factors affecting chemical reaction rate:
1) temperature;
2) using catalyst;
3) activation energy;
4) concentration of reagents.

7. Choose the main factor affecting chemical reaction rate of a heterogeneous system:
1) pressure;
2) temperature;
3) state of aggregation;
4) contact area of reacting substances.

8. Find compounds which have reaction with the highest chemical reaction rate at normal conditions:

1) Zn / HCl;
2) Na / H_2O;
3) Mg / H2O;
4) Pb / HCl.

9. Choose the factor which doesn't affect the chemical reaction rate between H_2SO_4 and iron:

1) H_2SO_4 concentration;
2) Fe grinding;
3) increase in temperature;
4) increase in pressure.

10. What is the reaction which proceeds at the lowest chemical reaction rate at normal conditions:

1) Fe / O_2;
2) Mg / HCl (10%);
3) Cu / O_2;
4) Zn / HCl (10%).

Tasks

1. How will the rate of the reaction change in case of a two times increase in reactants' concentration:

$$N_2(g) + O_2(g) = 2NO(g)$$

____

____

____.

2. How will the rate of the reaction change in case of a three times increase in pressure:

$$N_2(g) + 3H_2(g) = 2NH_3(g)$$

____

____

____.

3. How will the rates of chemical reaction change in case of a temperature increase from 20 °C to 80 °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 2. How will the rate of this reaction change in case of a 20 °C temperature increase?

____

____

____.

5. The Q_{10} coefficient for a certain reaction is equal to 4. At 10 °C that reaction lasts for 2 minutes. How long will that reaction last at 50 °C?

____

____

____.

Lesson 13. The main classes of inorganic compounds

Definitions:

1. **Oxide** is a binary chemical compound that contains oxygen with the oxidation state -2 and other chemical elements.
2. **Base** is an inorganic chemical compound that contains hydroxide group (-OH) and atoms of metal.
3. **Acid** is an inorganic chemical compound that gives just one type of a cation (H^+ cation), during dissociation in water solution.
4. **Salt** is an ionic compound that results from the neutralization reaction of an acid and a base.

Classification of inorganic compounds Table 18

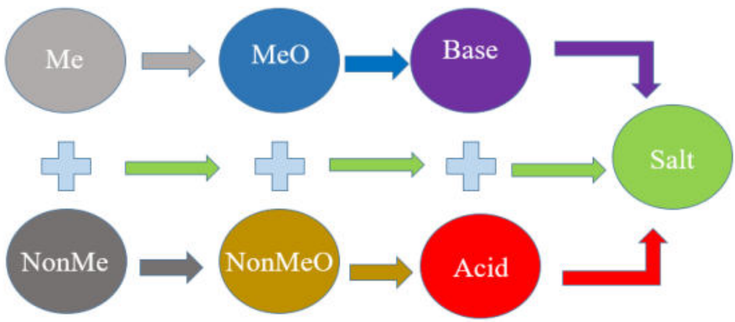
Class	Type			
Oxide	**Basic** Me + O_2 K_2O, Na_2O,CaO	**Amphoteric** Metalloids, d-block elements + O_2 BeO, ZnO, Al_2O_3	**Acidic** NonMe + O_2 SO_3,P_2O_5,Cl_2O_7	**Neutral** CO, SiO,N_2O, NO, S_2O (do not form salts)

Base	**Soluble** in water solutions (**Alkalis**) NaOH, KOH, $Ba(OH)_2$		**Insoluble** in water solutions $Cu(OH)_2$, $Zn(OH)_2$, $Al(OH)_3$	
Acid	**Containing** oxygen HNO_2, H_3PO_4, H_2SO_3	**Without** oxygen HCl, HF, HI	**Strong** H_2SO_4,HNO_3,HCl	**Weak** HNO_2, H_2SO_3, H_2CO_3
Salt	**Neutral** Na_2SO_4, KCl, $AlPO_4$	**Acidic** $KHSO_4$, $Ca(HCO_3)_2$, NaH_2PO_4	**Basic** CaOHCl, $BaOHNO_3$, $Fe(OH)_2Cl$	**Complex** $Na[Al(OH)_4]$, $K_3[Fe(CN)_6]$

Chemical properties of the main classes of inorganic compounds Table 19

Oxides			
Basic	Acidic	Amphoteric	
+Acid = Salt + H_2O $BaO + 2HCl \rightarrow BaCl_2 + H_2O$	**+Base = Salt+ H_2O** $P_2O_5 + 6KOH \rightarrow 2K_3PO4 + 3H_2O$	**+ Acid** **= Salt + H_2O/Complex** **+ Base** $Al_2O_3 + 6HCl \rightarrow 2AlCl_3 + 3H_2O$ $Al_2O_3 + 2NaOH + 3H_2O \rightarrow 2Na[Al(OH)_4]$	
IA, II A (exc. Mg,Be) + H_2O = Base $Na_2O + H_2O \rightarrow 2NaOH$	**+ H_2O = Acid** $N_2O_5 + H_2O \rightarrow 2HNO_3$		
MeO+NonMeO = Salt $3CaO + P_2O_5 \rightarrow Ca_3(PO_4)_2$			
Bases			
+ Acid = Salt + H_2O $2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$	**+ Acidic oxide = Salt+ H_2O** $2NaOH + SO_3 \rightarrow Na_2SO4 + H_2O$	**+ Salt = $Salt_2$ +gas/insoluble product** $2KOH + FeCl_2 \rightarrow 2KCl + Fe(OH)_2\downarrow$	**Amphoteric + Acid / Base = Salt + H_2O/ Complex** $Al(OH)_3 + 3HCl \rightarrow AlCl_3 + 3H_2O$ $Al(OH)_3 + NaOH \rightarrow Na[Al(OH)_4]$
Acids			

+ active Me = Salt+ H_2 $Zn + H_2SO4 \rightarrow ZnSO4 + H_2\uparrow$	**+MeO = Salt+ H_2O** $H_2SO_4 + CuO \rightarrow CuSO_4 + H_2O$	**+ Salt = $Salt_2$ + gas/insoluble product** $CaCO_3 + 2HCl \rightarrow CaCl_2 + CO2\uparrow + H_2O$	**+ MeOH = Salt+H_2O** $2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$
Salts			
Active Me + Salt = $ActMeSalt_2$ + NonactMe $CuSO_4 + Zn \rightarrow ZnSO_4 + Cu$			



Scheme 4. Genetic line of elements

Questions after lesson:
1. What are the main classes of inorganic compounds?
2. What are the main chemical properties of bases?
3. What is the classification of salts?
4. What are the main chemical properties of oxides?
5. What is the main difference of amphoteric oxides from the rest?

Test

1. Choose the formulas of basic oxides:
1) P_2O_5 / Na_2O / $Ca(OH)_2$;
2) SO_3 / K_2O / $Al(OH)_3$;
3) Cr_2O_3 / CO_2 / SiO_2;
4) Li_2O / K_2O / Na_2O.

2. Choose the formulas of acidic salts:
1) Na_2SO4 / $BaCl_2$ / $NaHCO_3$;
2) KHSO4 / Na_2HPO4 / $Mg(HCO_3)_2$;
3) CaOHCl / $KAlO_2$ / $AgNO_3$;
4) NaCl / KNO_3 / $CuSO_4$.

3. Find the reaction in which the product will be acidic oxide:
1) K + H2O =;
2) $Na + O_2$ =;
3) CO + C =;
4) $CaCO_3$ =.

4. Find the oxide which can react with bases and acids:
1) CO;
2) SiO;
3) ZnO;
4) FeO.

5. Choose bases which cannot be formed in the reaction between corresponding oxides and water:
1). KOH / NaOH / LiOH;
2) $Mg(OH)_2$ / $Ca(OH)_2$ / $Cu(OH)_2$;
3) $Al(OH)_3$ / $Zn(OH)_2$ / $Fe(OH)_2$;
4) $Ba(OH)_2$ / $Cr(OH)_3$ / $Fe(OH)_3$.

6. Find the formulas of neutral oxides:
1) CO / CO_2 / SO_2;
2) S_2O / N_2O / Na_2O;
3) N_2O5 / P_2O5 / SO_3;
4) SiO / CO / NO.

7. Choose strong acid:
1) H_2CO_3;
2) H_2SiO_3;
3) $HClO_3$;
4) H_3PO_4.

8. Find the reaction in which the product will be acidic salt:
1) K_2O + HCl =;
2) $K_2O + 2H_2SO_4$ =;
3) $K_2O + H_2SO_4$ =;
4) $K_2O + H_2CO_3$ =.

9. Choose salt which can react with alkali:
1) Na_2CO_3;
2) $BaCl_2$;
3) $ZnSO_4$;
4) KF.

10. Choose salts which can react with the acid containing the same anion:
1) NaCl;

2) K_2SO_3;
3) $AlPO_4$;
4) $Ca(NO_3)_2$.

Tasks

1. Write down formulas of the following compounds:

1) phosphorus (V) oxide ____________________;

2) iron (III) hydroxide ____________________;

3) manganese (VII) oxide ____________________;

4) aluminum sulphate ____________________;

5) chrome hydroxide (II) ____________________.

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

1) $Al(OH)_3$____________________;

2) $Fe(OH)_3$____________________;

3) $Zn(OH)_2$____________________;

4) KOH ____________________;

5) $Ca(OH)_2$____________________.

3. Arrange oxides in order of the increase of their acidic properties:

Cr_2O_3 / CrO_3 / CrO / CrO_2

____

____.

4. Write the products of the following reactions:

1) $NaOH + H_3PO_4 =$ ____________________;

2) $H_2SO_4 + Fe(OH)_3 =$ ____________________;

3) $KHSiO_3 + KOH =$ ____________________;

4) $Ba(OH)_2 + N_2O_5 =$ ____________________;

5) $AgNO_3 + CaCl_2 =$ ____________________;

6) $Zn + HNO_3 =$ ____________________;

7) $CaO + CO_2 =$ ____________________;

8) $Al_2O_3 + NaOH =$ ____;

9) $CuSO_4 + Fe =$ ____;

10) $K_2SO_4 + BaCl_2 =$ ____.

5. Solve the chemical equation chain:

1) $Al \rightarrow AlCl_3 \rightarrow Al(OH)_3 \rightarrow NaAlO_2 \rightarrow Al_2(SO_4)_3$

____

____

____

____;

2) $Mg \rightarrow MgO \rightarrow Mg(NO_3)_2 \rightarrow Mg(OH)_2 \rightarrow MgSO_4$

____

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____

____;

3) $P \rightarrow P_2O_5 \rightarrow Na_3PO_4 \rightarrow Na_2HPO_4 \rightarrow Ba_3(PO_4)_2$

____

____

____

____;

4) $Mg(OH)_2 \rightarrow MgO \rightarrow Mg(NO_3)_2 \rightarrow Mg(OH)_2 \rightarrow MgOHCl$

____

____

____

____;

5) $NaHCO_3 \rightarrow CO_2 \rightarrow CaCO_3 \rightarrow Ca(HCO_3)_2 \rightarrow CaO$

____

____

6. Calculate the mass of sulphuric acid produced from 11,6 g of sulphur (VI) oxide in its reaction with water.

7. What is the mass of CuSO4 produced from 12,8 g of sulphuric acid and 7,8 g of CuO?

8. Find what kind of salt(s) is formed in the reaction between 6 g of sodium hydroxide and 4,8 L (normal conditions) of carbon dioxide.

9. What is the mass of silver chloride produced in the reaction between 4,75 g of barium chloride and 34,9 g of silver nitrate?

10. What is the mass of barium sulphate produced in the reaction between 3,25 g of barium hydroxide and 2,4 g of sulphuric acid?

Lesson 14. Solutions

Definitions:

1. **Solution** is a homogenous mixture of two or more substances in relative amounts that can be varied.
2. **Solute** is the substance being dissolved.
3. **Solvent** is the substance that is present in greater amount.
4. **Dissolution** – is a physicochemical process of destruction of the crystals of the substance and the formation of separated particles.
5. **Solubility** is the maximum amount of a substance that will dissolve in a given amount of solvent at a specified temperature.

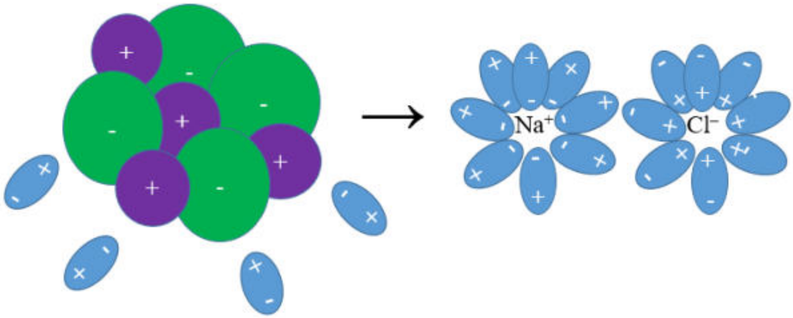
Common Types of solutions Table 20

Solution phase	Solute phase	Solvent phase	Example
Gaseous solutions	gas	gas	air (mostly N_2 and O_2)
Liquid solutions	gas liquid solid	liquid liquid liquid	soda (CO_2 in H_2O) vinegar (CH_3COOH in H_2O) seawater (NaCl in H_2O)
Solid solutions	solid	solid	brass (Zn in Cu)

Classification of solutions Table 21

Solutions	dilute	has a small amount of solute in a large amount of solvent
	concentrated	has a large amount of solute in a small amount of solvent
	unsaturated	more solute can be dissolved
	saturated	no more solute can be dissolved
	supersaturated	becomes unstable, crystals form

$$NaCl_{(s)} + H_2O \rightarrow Na^+_{(aq)} + Cl^-_{(aq)}$$



Scheme 5. Dissolution

Quantitative characteristics of solution — Table 22

Molarity	is the number of moles of solute (**n**) per liter (**L**) of solution: $C = \frac{n}{V}$; mole/L
Molality	is the number of moles of solute (**n**) per kilogram (**kg**) of solution: $C_m = \frac{n}{m}$; mole/kg
Mass percentage	Is the ratio of the mass of solute to the mass of the solution in % : $w = \frac{m\ (solute)}{m\ (solution)} \times 100\%$
Volume percentage	Is the ratio of the volume of solute to the volume of the solution in %: $\varphi = \frac{V\ (solute)}{V\ (solution)} \times 100\%$

Questions after lesson:

1. What type of mixture is called a solution?
2. What components does the mixture consist of?
3. What are the main types of mixtures?
4. What characteristics of solution are called quantitative?
5. What is the process of dissolution?
6. What does solubility show?

Test

1. Choose true statements about water:
1) boiling temperature equals 100 °C;
2) colourless, tasteless;
3) only liquid substance;
4) produce vapour after refrigeration.

2. Mass percentage is the ratio between:
1) number of moles of solution and moles of solute;
2) mass of solute and mass of solution;
3) volume of solution and volume of solvent;
4) number of moles of solute and volume of solution.

3. Molarity is the ration between:
1) number of moles of solute and volume of solution;
2) number of moles of solute and mass of solution;
3) mass of solute and mass of solution;
4) volume of solvent and volume of solution.

4. Molality is the ratio:
1) number of moles of solution and moles of solute;
2) mass of solute and mass of solution;
3) number of moles of solute and mass of solution;
4) number of moles of solute and volume of solution.

5. What is the dilute solution?
1) solution with the small amount of solute substance;
2) solution with the large amount of solute substance;
3) solution in which more solute can be dissolved;
4) solution in which no more solute can be dissolved.

6. Choose insoluble salts:
1) $NaCl$;
2) $BaSO_4$;
3) $AgNO_3$;
4) $CaCO_3$.

7. Choose soluble salts:
1) KNO_3;
2) $Na2CO_3$;
3) $BaCl_2$;
4) MgF_2.

8. Which substance demonstrates good solubility in water?
1) CH_4;

2) $Al_2(SO_4)_3$;
3) C_2H_2;
4) C_2H_5OH.

9. Which substance demonstrates good solubility in benzene?
1) CH_3OH;
2) KCL;
3) C_5H_{12};
4) H_3PO_4.

10. Choose the way of dissolving a precipitate in water solution:
1) decrease the temperature;
2) add more solvent;
3) increase the temperature;
4) add another substance which has reaction with precipitate.

Tasks

1. Calculate the mass of NaOH needed to make 300 ml of solution with mass percentage of 11 % and density equal to 0.848 g/ml.

____

____

____

2. Calculate the number of moles of KOH present in 750 ml of solution with the molarity equal to 0.025 mol/L.

____

____

____

3. Calculate the molarity of water solution made from 17 g of $MgSO_4$, if the final volume is 950 ml.

____

____

____

4. Find the mass of $CuSO_4 \cdot 5H_2O$ which is needed to prepare 550 ml of $CuSO_4$ water solution with the molarity equal to 0.05 mol/L.

5. Calculate the mass percentage of ammonia nitrate in water solution made from 19 g of NH_4NO_3 and 115 ml of pure water.

6. What volume of 4.28 M KCl is needed to obtain 1L of 0.75 KCl solution? Density is equal to 1 g/mL.

7. Calculate the mass of a precipitate formed after the mixing of 18 ml of 0.02 M sodium chloride and 17 ml of 0.01 M silver nitrate solutions.

8. Calculate the mass of a precipitate formed after the mixing of 37 ml of 0.01 M potassium sulfate and 25 ml of 0.02 M barium chloride solutions.

9. Calculate the mass percentage of nitrogen in $Al(NO_3)_3$.

10. Calculate the mass percentage of oxygen in $Na_2SO_4 \cdot 10H_2O$.

Lesson 15. Theory of electrolytic dissociation

Definitions:

1. **Electrolytic dissociation** is the process of decomposition of molecules into ions during its dissolution or melting.
2. **Electrolyte** is a substance that conducts electric current as a result of dissociation into ions. Only an ionic substance.

Electrolytic dissociation of different types of electrolytes Table 23

Electrolytes	Acid	1. Strong acids dissociate in one step: $H_2SO_4 \rightarrow 2H^+ + SO_4^{2-}$
		2. Weak acids dissociate in a stepwise manner: 1) $H_3PO_4 \rightleftarrows H_2PO_4^- + H^+$ 2) $H_2PO_4^- \rightleftarrows HPO_4^{2-} + H^+$ 3) $HPO_4^{2-} \rightleftarrows PO_4^{3-} + H^+$
	Base	1. Strong bases dissociate in one step: $NaOH \rightarrow Na^+ + OH^-$
		2. Weak bases dissociate in a stepwise manner: 1) $Cu(OH)_2 \rightleftarrows CuOH^+ + OH^-$ 2) $CuOH^+ \rightleftarrows Cu^{2+} + OH^-$
	Salt	1. Neutral salts dissociate in one step: $NaCl \rightarrow Na^+ + Cl^-$
		2. Acidic salts of strong acid dissociate in one step: $KHSO_4 \rightarrow K^+ + H^+ + SO_4^{2-}$
		3. Acidic salts of weak acid dissociate in a stepwise manner: 1) $KH_2PO_4 \rightarrow H_2PO_4^- + K^+$ 2) $H_2PO_4^- \rightleftarrows HPO4^{2-} + H^+$ 3) $HPO_4^{2-} \rightleftarrows PO_4^{3-} + H^+$
		4. Basic salts of strong bases dissociate in one step: $CaOHCl \rightarrow Ca^{2+} + OH^- + Cl^-$
		5. Basic salts of weak bases dissociate in a stepwise manner: 1) $ZnOHCl \rightarrow ZnOH^+ + Cl^-$ 2) $ZnOH^+ \rightleftarrows Zn^{2+} + OH^-$

Questions after task:

1. What are the types of chemical substances called electrolytes?
2. What are the processes that happen with substances during dissociation?
3. What kind of salts dissociate in a stepwise manner?

Tasks

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

1) H_2SO_3____

____;

2) HNO_2____

____;

3) $Zn(OH)_2$____

____;

4) $KHCO_3$____

____;

5) MgOHCl____

____.

2. Calculate the molar concentration of chloride ions in the water solution prepared from 8 g of chromium (III) chloride. The final volume equals 250 ml.

____

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3. Calculate the molar concentration of sulfate ions in the water solution prepared from 12 g of aluminum sulfate. The final volume equal 315 ml.

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____

____

4. Calculate the molar concentration of all ions in the water solution prepared from 18 g of calcium chloride. The final volume equal 1750 L.

____

____

____

5. Calculate the molar concentration of hydrogen ions in the water solution prepared from 4,6 L of hydrogen chloride (in normal conditions). The final volume is 370 ml.

Lesson 16. Ionic equations

Definitions:

1. **Ionic equation** is a chemical equation in which the electrolytes in water solution are expressed as dissociated ions.

How to write ionic equations?

Step 1. Write the molecular equation, put the coefficients in.	$AgNO_3 + NaCl \rightarrow NaNO_3 + AgCl$
Step 2. Check the solubility of reactants and products in the solubility table. Mark insoluble substance or gas.	$AgNO_3 + NaCl \rightarrow NaNO_3 + AgCl\downarrow$
Step 3. Write the reactants and products in the ionic form. Solid substance, gases, water are not electrolytes, therefore they do not have an ionic form. It will be a total ionic equation.	$Ag^+ + NO_3^- + Na^+ + Cl^- \rightarrow Na^+ +$ $NO_3^- + AgCl\downarrow$
Step 4. Cancel the same ions from right and left sides of the equation, and you will receive the short ionic equation.	$Ag^+ + \cancel{NO^{3-}} + \cancel{Na^+} + Cl^- \rightarrow \cancel{Na^+} +$ $\cancel{NO_3^-} + AgCl\downarrow$ $Ag^+ + Cl^- \rightarrow AgCl\downarrow$

Questions after lesson:

1. What do the ionic equations show?
2. What are the differences between ionic and molecular equations?

Tasks

1. Write the products of chemical reactions, balance and write complete and short ionic equations for them:

1) $SrCl_2 + Na_2SO_4 =$ ____

____

____;

2) $Ba(OH)_2 + H_2SO_3 =$ ____

____

____;

3) $NH_4NO_3 + NaOH =$ ____

____

____;

4) $Li_2S + HCl =$ ____

____

____;

5) $K_2SiO3 + HNO_3 =$ ____

____

____;

6) $NaHCO_3 + NaOH =$ ____

____

____;

7) $CaOHCl + HBr =$ ____

____

____;

8) $FeI_2 + LiOH =$ ____

____

____;

9) $Cr(OH)_3 + KOH =$ ____

____

____;

10) $CaCl_2 + Na_2CO_3 =$ ____

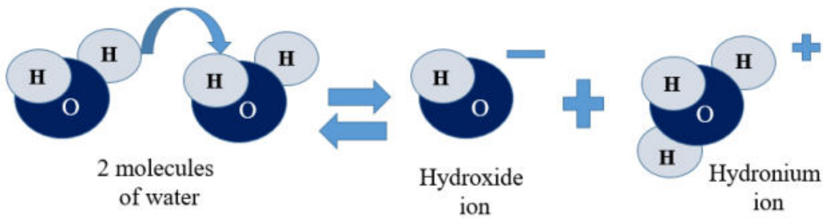
____

____.

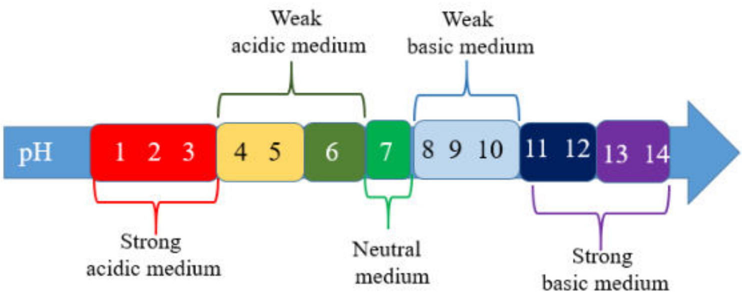
Lesson 17. Hydrolysis

Definitions:

1. **Hydrolysis** is a chemical process of decomposition involving the splitting of a bond and the addition of the hydrogen cation and hydroxide anion of water.
2. **Self-ionization of water** is the reaction in which water molecules produce ions.
3. **pH scale** is the quantitative measure of the acidity or basicity of water solutions.



Scheme 6. Self-ionization of water



$$pH = -\lg[H^+]; \quad [H^+] = 10^{-pH}$$

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

$$[H^+]\cdot[OH^-] = 10^{-14}$$

$$pH + pOH = 14$$

Scheme 7. pH scale

Hydrolysis Table 24

Type of salt	Mechanism of hydrolysis	Example
A salt formed between a strong acid and a strong base is a neutral salt.	No hydrolysis (pH = 7)	$NaCl + HOH \leftrightarrow NaOH + HCl$ $Na^{+} + Cl^{-} + HOH \leftrightarrow Na^{+} + OH^{-} + H^{+} + Cl^{-}$ $HOH \leftrightarrow OH^{-} + H^{+}$ Neutral medium
A salt formed between a strong acid and a weak base is an acid salt.	Cationic hydrolysis (pH< 7)	$NH_4Br + HOH \leftrightarrow NH_4OH + HBr$ $NH_4^{+} + Br^{-} + HOH \leftrightarrow NH_4OH + H^{+} + Br^{-}$ $NH_4^{+} + HOH \leftrightarrow NH_4OH + \boxed{H^{+}}$ Acidic medium
A salt formed between a weak acid and a strong base is a basic salt.	Anionic hydrolysis (pH>7)	$NaNO_2 + HOH \leftrightarrow NaOH + HNO_2$ $Na^{+} + NO_2^{-} + HOH \leftrightarrow Na^{+} + \boxed{OH^{-}} + HNO_2$ $NO_2^{-} + HOH \leftrightarrow OH^{-} + HNO_2$ Basic medium
A salt formed between a weak acid and a weak base can be neutral, acidic, or basic depending on the relative strengths of the acid and base.	Cationic-anionic hydrolysis (pH≈7)	$CH_3COONH_4 + HOH \leftrightarrow NH_4OH + CH_3COOH$ $CH3COO^{-} + NH_4^{+} + HOH \leftrightarrow NH_4OH + CH_3COOH$ Neutral medium

Questions after lesson:
1. What is the process of hydrolysis?
2. What does the pH scale show?
3. When is hydrolysis possible?

Tasks

1. Write the equation of hydrolysis for given substances, if it is possible:

1) $Mg(NO_3)_2 + H_2O =$;

2) $(NH_4)_2CO_3 + H_2O =$;

3) $AlCl_3 + H_2O =$;

4) (CH3COO)$_2$Ni $+ H_2O =$;

5) $NaH_2PO_4 + H_2O =$;

6) $FeCl_3 + H_2O =$;

7) $K_2S + H_2O =$;

8) $CuSO_4 + H_2O =$;

9) $Li_2SO_3 + H_2O =$;

10) $CaOHCl + H_2O =$.

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Name	Symbol	Unit of measurement
Atomic mass	A_r	–
Molecular mass	M_r	–
Mass	m	g
Molar mass	M	g/mole
Mole	n	mole
Volume	V	L
Molar volume	V_m	L/mole
Mass percentage	ω	%
Volume percentage	φ	%
Mole percentage	χ	%
Density	ρ	g/L
The relative density of the gas	D	–
Molarity	C	mole/L
Number of units	N	
Avogadro's Number	N_A	$mole^{-1}$

	Formula's		
	$N = N_A \cdot n$	$n = \frac{N}{N_A}$	$N_A = 6{,}02 \cdot 10^{23}\ mole^{-1}$
	$m = M \cdot n$	$n = \frac{m}{M}$	$M = \frac{m}{n}$
	$m = V \cdot \rho$	$V = \frac{m}{\rho}$	$\rho = \frac{m}{V}$
	$C = \frac{n}{V}$	$n = C \cdot V$	$V = \frac{n}{C}$
	$\omega(E) = \frac{A_r(E) \cdot x \cdot 100\%}{M_r}$	$x = \frac{\omega(E) \cdot M_r}{A_r(E) \cdot 100\ \%}$	$M_r = \frac{A_r(E) \cdot x \cdot 100\%}{\omega(E)}$
	$\omega(_{solvent}) = \frac{m(_{solvent}) \cdot 100\ \%}{m(_{solution})}$	$m(_{solvent}) = \frac{\omega \cdot m(_{solution})}{100\ \%}$	$m\ (_{solution}) = \frac{m(_{solvent}) \cdot 100\ \%}{\omega(_{solvent})}$

For gases			
	$V = V_m \cdot n$	$n = \frac{V}{V_m}$	$V = V_m \cdot n$ $M = V_m \cdot \rho(gas)$
	$M = V_m \cdot \rho(gas)$	$\rho(gas) = \frac{M}{V_m}$	
	$D_{gas\ 1}\ (gas\ 2) = \frac{M(gas\ 2)}{M(gas\ 1)}$		M_r (dry air) = 29
	$M_r(_{gas\ mixture}) = M_r(gas\ 1) \cdot \varphi_1 + M_r(gas\ 2) \cdot \varphi_2 + \ldots$		$\varphi(gas) = \frac{V_{(gas)} \cdot 100\ \%}{V_{(gas\ mixture)}}$
	$pV = nRT$, $R = 8{,}314$ – const.	$T_0 = 273$ °K (0 °C) $P_0 = 101{,}3$ кPa (= 1 atm) $\Big\}$ STP	

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