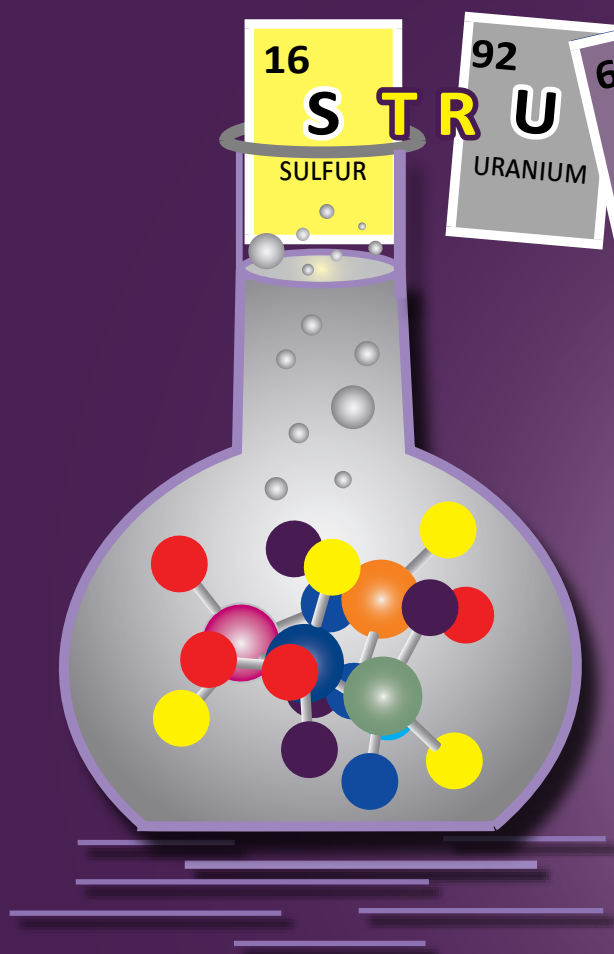




of MATTER

Electrons e^-

Protons p^+

Neutrons n^0

Electron shell Nucleus

Atoms Molecules

Chemical bonding



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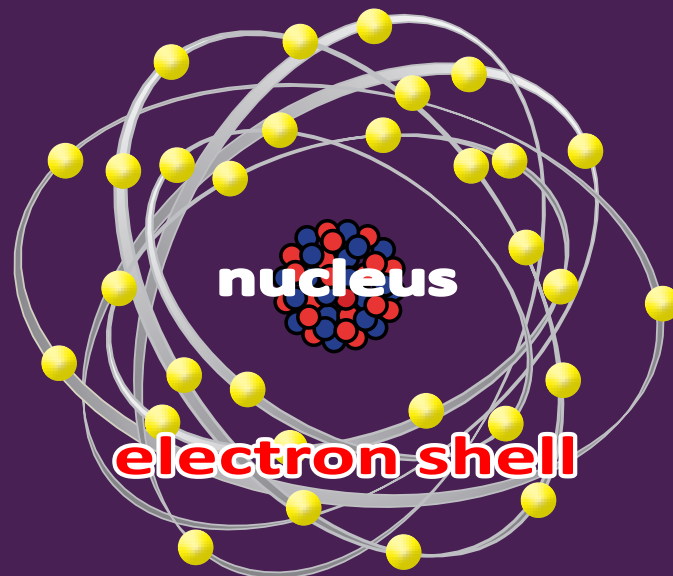
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1

ATOMIC STRUCTURE

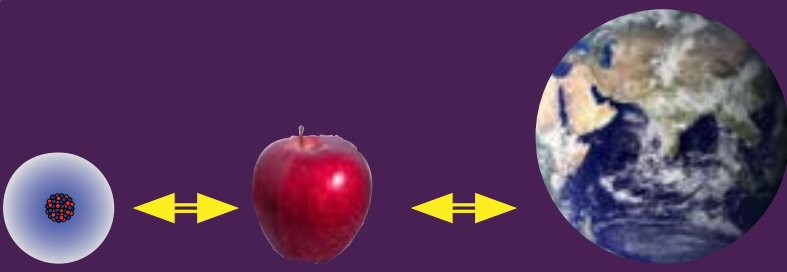
KEY TOPICS

1. What are atoms made of?
2. Why are atoms electrically neutral particles?
3. What are isotopes?
4. How do electrons move around the nucleus?



KEY WORDS

- Atom
- Nucleus
- Proton
- Neutron
- Electron
- Electron shell
- Mass number
- Isotopes



an atom **an apple** **the Earth**

Atoms are invisible as they are very, very small*. They are as small compared to an apple as the fruit is to the entire Earth. The atomic radii are about 50 - 500 pm.
(1 pm = 0,000 000 000 001 m)

All atoms are made up of ultrasmall particles – electrons, protons and neutrons.

The existence of electrons was established experimentally by Sir Joseph Thomson in 1897.

Ernest Rutherford, 1st Baron Rutherford of Nelson, with the collaboration of his students, established experimentally the existence of the atomic nucleus. He is also credited with the discovery of the proton in 1918. He realized the first nuclear reaction.

Sir James Chadwick discovered the neutron in 1932.

These prominent scientists were awarded the Nobel prize in chemistry or physics.

WHAT ARE ATOMS?

Science states that an **ATOM** is the smallest unit of matter that has the properties of a chemical element. But what does this mean?

Everything we know – our bodies, animals, plants, rocks, oceans, buildings, cars and all other man made objects, the planets and stars, etc. are made of the same material, the invisible for the human eye atoms of 98 chemical elements. In the Periodic table they occupy the first 98 cells. The first one is hydrogen H ($Z = 1$) and the last one is californium Cf ($Z = 98$). But only traces of astatine As ($Z = 85$), francium ($Z = 87$) and the elements with atomic numbers from 93 to 98 have been found in some ores. So it is considered that in nature there are atoms of 90 chemical elements (uranium U, $Z = 92$, is the last one), i.e. the atoms of 90 different elements build up all the Universe. There is nothing around us that is made of something different. The elements with

higher atomic numbers are radioactive, they are not found in nature and their atoms have been synthesized by man. Officially the *International Union of Pure and Applied Chemistry* (IUPAC) has approved (till december 2017) the synthesis of 26 new elements. Thus the known today elements are 118.

* It is worth watching the video „Have you ever seen an atom?“ on YouTube

Each chemical element, naturally occurring or synthesized by man, has its own „identity“, like the finger prints of any person. But in spite of that all atoms are made of the same „material“ – small particles called **subatomic** or **fundamental** – **electrons**, **protons** and **neutrons**. Here are their characteristics:

SUBATOMIC PARTICLE	SYMBOL	ELECTRIC CHARGE (in elementary charges 1 el. ch. = $1,60218 \cdot 10^{-19}$ C)	MASS in kg	MASS in amu* (1 amu = $1,661 \cdot 10^{-27}$ kg)
Electron	e^-	-1	$0,911 \cdot 10^{-30}$	0,000548
Proton	p^+	+1	$1,673 \cdot 10^{-27}$	1,0073
Neutron	n^0	0	$1,675 \cdot 10^{-27}$	1,0084

According to the modern view, atoms consist of two main parts – a **nucleus** and an **electron shell**.

The **nucleus** is built up of **protons** and **neutrons** (sometimes named **nucleons**), held tightly together by strong nuclear forces. Almost the whole atomic mass is concentrated in the nucleus. It is positively charged. The number of protons in the nuclei of the atoms of a chemical element is equal to its atomic number Z. For example all carbon atoms have 6 protons in their nuclei and the atomic number of carbon is $Z = 6$. All uranium atoms have 92 protons in their nuclei and the atomic number of uranium is $Z = 92$. In chemical reactions the nucleus remains unchanged.

ALL ATOMS OF A CHEMICAL ELEMENT HAVE THE SAME NUMBER OF PROTONS IN THEIR NUCLEI.

* **1 amu** – **A**tomic **M**ass **U**nit, the basic mass unit in chemistry

1 amu = 0,000 000 000 000 000 000 000 001 661 kg.

This is 1/12th of the mass of an unbound carbon atom with 6 protons and 6 neutrons in its nucleus.

The **electron shell** consists of all **electrons** moving around the nucleus. Almost the whole atomic volume is occupied by the electrons in the electron shell.

Atoms are electrically neutral particles as the number of protons in their nuclei is always equal to the number of protons in the electron shells. For example all oxygen ($Z = 8$) atoms have 8 protons in their nuclei and 8 electrons in their electron shell. Osmium ($Z = 76$) has 76 protons in its nucleus and 76 electrons move around it.

ATOMIC STRUCTURE KEY POINTS:

Number of protons **$N_{p^+} = Z$** (atomic number), identifies the chemical element

Number of protons $N_{p^+} =$ number of electrons N_{e^-}

ATOMS ARE ELECTRICALLY NEUTRAL PARTICLES.

The sum of the number of protons and the number of neutrons in the nucleus is called **mass number A**. For example a carbon atom with 6 protons and 6 neutrons in its nucleus has a mass number of 12. The number of protons in the nuclei of the atoms of a chemical element is always the same but the number of neutrons may be different. Such atoms are **isotopes** of the chemical element.

ISOTOPES are atoms of a chemical element with atomic number Z but different mass number A.

Isotopes are identical in their chemical behavior but differ in some physical properties. The chemical elements exist in nature as mixtures of isotopes in different proportions. For that reason the atomic masses of the elements in the periodic table are not an integer.

PROBLEM!

In the atmosphere there are three isotopes of neon Ne. The first with mass number $A_1 = 20$ is 89,97 % of the neon atoms, the second with $A_2 = 22$ is 9,73 % of the neon atoms and the third with $A_3 = 21$ is 0,3 % of the neon atoms. Calculate the relative atomic mass of neon.

SOLUTION: The mass of the electrons in the electron shell is negligible compared to the mass of the nucleus. So the mass of the electrons can be neglected and the relative atomic mass may be calculated from the mass numbers of all isotopes taking into consideration their abundance, i.e.:

$$A_r(\text{Ne}) = 0,8997 \cdot A_1 + 0,0973 \cdot A_2 + 0,003 \cdot A_3 = 0,8997 \cdot 20 + 0,0973 \cdot 22 + 0,003 \cdot 21 = 20,2$$

ANSWER: The relative atomic mass of neon Ne is $A_r(\text{Ne}) = 20,2$

Isotopes are used in science and technologies. The carbon-14 (carbon atoms with 6 protons and 8 neutrons) radioactive decay is used for dating of objects, iodine-125 or 131 are used to monitor the activity of the thyroid gland, cobalt-60 produces γ -rays used in the radiation treatment of cancer.

DO ELECTRONS MOVE FREELY AROUND THE NUCLEUS?

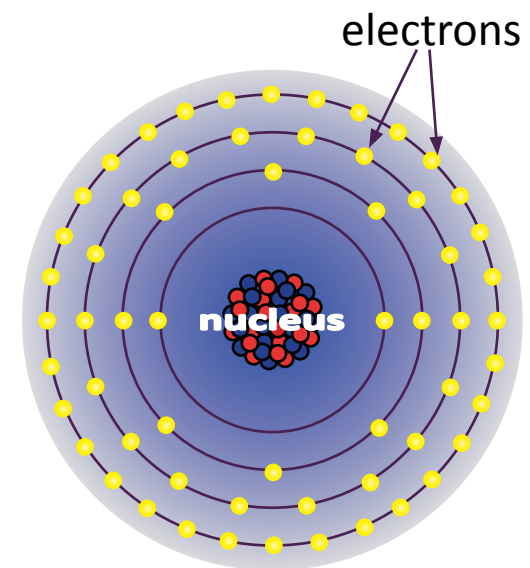
The movement of the electrons around the nucleus is not random. The energy of each electron is strictly determined and it spins very quickly inside a particular volume of the space around the nucleus. This movement does not obey to the rules of the visible for man world. The world of the subatomic particles is governed by the rules of the quantum mechanics and the 3D shapes of the electrons trajectories are not simple at all*. Only for simplicity we depict them as circles around the nucleus (similarly to the movement of the planets around the sun). But remember that this is not so simple!

The contemporary science considers that the electrons in atoms are grouped in **electron shells** (called also energy levels or layers), the same term as for the whole assemblage of electrons in an atom. The shells are denoted with Arabic numbers (1, 2, 3 ... 7) or with Latin letters (K, L, M, N, O, P and Q). Till now atoms with more than 7 electron shells are not known.

The energy of the electrons in an electron shell is approximately the same. But electrons in different shells have quite different energies. The electrons of the nearest to the nucleus shell have the lowest energy. The energy of electrons increases with the distance away from the nucleus.

The maximum number of electrons in each shell is strictly determined – if the number of the shell is ***n*** it is **$2n^2$** . Thus there are maximum only two electrons on the nearest to the nucleus K shell. On the second L shell may be located up to 8 electrons, on the third M shell – 18 electrons and so on.

*It is impossible to know the velocity and the location of the electron at any point and time. Only the probability of where the electron is likely to be found is known.

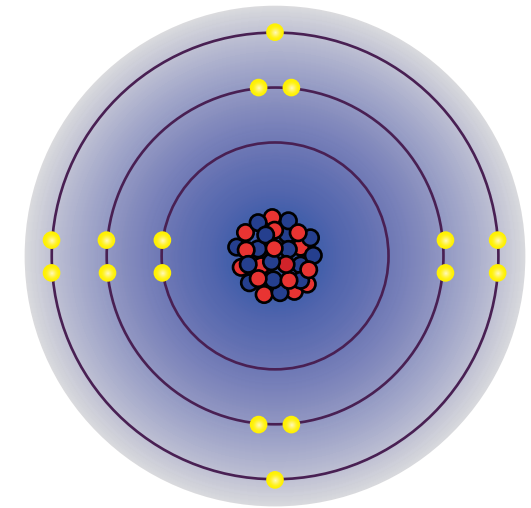


The energy of the electrons in an atom is strictly determined. Electrons are grouped in electron shells. The maximum number of electrons in each shell is also strictly determined.

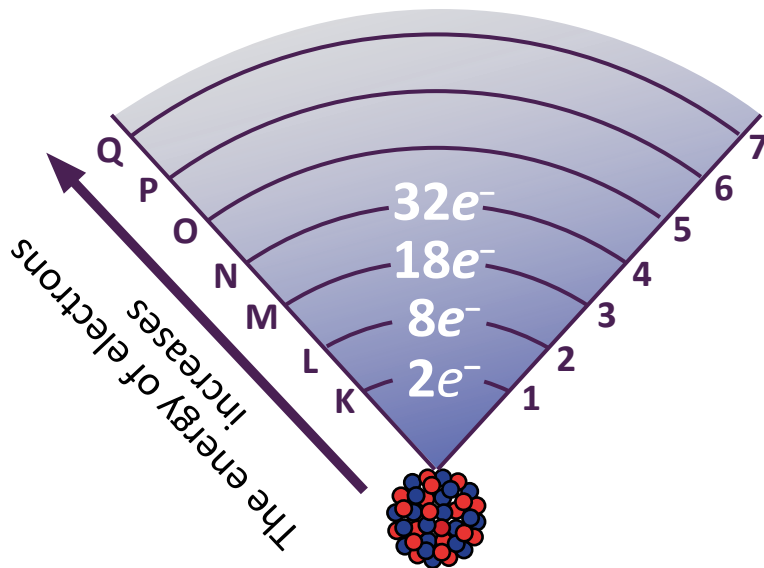
The atoms of an element have a definite number of protons in their nuclei that is equal to the number of electrons around it and to the atomic number Z :

$$Np^+ = Ne^- = Z$$

For example the atomic number of sulfur S is $Z = 16$ and all sulfuric atoms have 16 electrons. But is it possible that all electrons occupy the third electron shell as it can hold up to 18 electrons? The answer is „no“. The electrons occupy the electron shells with the lowest possible energy. So the first 2 electrons go on the K shell. The next 8 electrons go on the second L shell. Only the remaining 6 electrons go on the third M shell.



ELECTRON SHELL DIAGRAM
OF SULFUR ATOMS S
 $Z = 16$

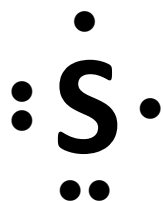


ELECTRON SHELLS IN ATOMS

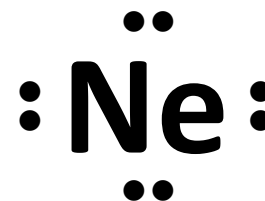
It is interesting to know:

The fifth O shell can hold up to 50 electrons, the P shell – 72 electrons and the Q shell – 98 electrons. But the outer three shells are not filled with the maximum electrons they can hold in any atom we know today. The reason is that some electrons of the 4th, 5th, 6th and 7th shells have lower energy than some of the electrons in the preceding shells. Thus for example potassium has 19 electrons which occupy four shells – the K and the L shells are filled with 2 and 8 electrons, but in the M shell there are only 8 electrons and one electron is in the N shell.

The energy of the electrons of the outermost electron shell (called **valence shell**) is high compared to the electrons occupying the inner shells. They are called **valence electrons** and may be shared with other atoms thus forming chemical bonds. These electrons may be unpaired (single electrons) or paired (electron pairs) – two electrons move together around the nucleus. If the electron shell is filled with the maximum electrons it can hold, all electrons are paired. (See the sulfur diagram on the previous page.) A common way of presenting the electrons of the valence shell is by the **Lewis-dot formulae** – each valence electron is presented with a dot. Single dots • represent single electrons and double dots •• represent electron pairs. Both single and double dots are positioned around the chemical symbol of the element:



LEWIS-DOT FORMULA
OF THE SULFUR ATOM
Z = 16



LEWIS-DOT FORMULA
OF THE NEON ATOM
Z = 10

As a rule of thumb the valence shell may be filled with up to 8 electrons (as in the atoms of noble gases like neon Ne). This statement is known as the **octet rule**. Atoms with less than 8 valence electrons are more or less reactive. If in an atom the valence electrons are 8 it is practically not reactive. Thus sulfur easily reacts with many substances, while neon is not a chemically active element. Exceptions of the octet rules are the elements of the first period – hydrogen H (Z = 1) and helium He (Z = 2). In their atoms there is only one electron shell that is filled with up to 2 electrons.

Why is matter organized in this way? We don't know, but it is like that!

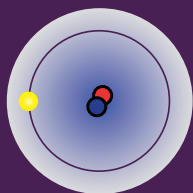
CAN YOU ANSWER THESE QUESTIONS ?

1. Why are atoms invisible?
2. What are protons, neutrons and electrons?
3. How many protons are there in the nuclei of nitrogen ($Z = 7$)?
4. How many electrons do the atoms of chlorine ($Z = 17$) have?
5. Why is atom electrically neutral?
6. How do electrons move around the nucleus?
7. Why are there four electron shells in the atoms of calcium ($Z = 20$)?
8. How many are the valence electrons of silicon ($Z = 14$)?
9. What are Lewis-dot formulae used for?
10. What does the octet rule state?
11. Why is sodium a highly reactive metal while Helium is an inert gas?

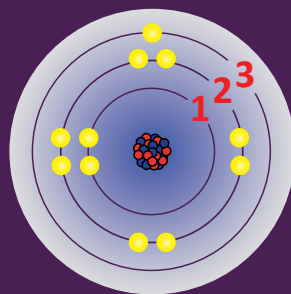
2 ATOMIC STRUCTURE AND THE PERIODIC TABLE

KEY TOPICS

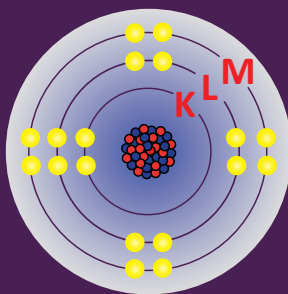
1. How is the modern periodic table organized?
2. Do the period number and the main group number depend on the atomic structure?
3. How do the properties of simple substances change across a period and in a group?
4. What is electronegativity?



Hydrogen **H**
 $Z = 1$



Sodium **Na**
 $Z = 11$



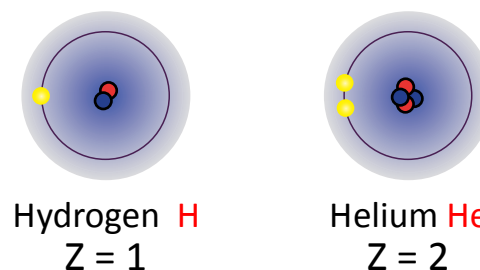
Argon **Ar**
 $Z = 18$

KEY WORDS

- Electron shell diagram
- Electron configuration
- Electronegativity

DOES THE PLACE OF ELEMENTS IN THE PERIODIC TABLE DEPEND ON THE ATOMIC STRUCTURE?

The periodic table summarizes all known chemical elements. About one and a half century ago Dmitrii Mendeleev classified the elements according to their atomic masses. The modern science, based on the knowledge of the atomic structure, classifies the elements according to the number of protons in the nuclei of their atoms. Each element has in its atoms 1 proton more compared to the previous element in the periodic table and 1 proton less than the following element. Thus the first element in the periodic table hydrogen H has only 1 proton in the nuclei of its atoms. Helium He, the second one, has 2 protons and lithium Li (atomic number $Z = 3$) – three protons. But why are hydrogen and helium in the first period of the periodic table while lithium is in the second period? The reason is that hydrogen and helium have only one electron shell in their atoms. It can hold up to two electrons and these two elements form the first period in the periodic table.

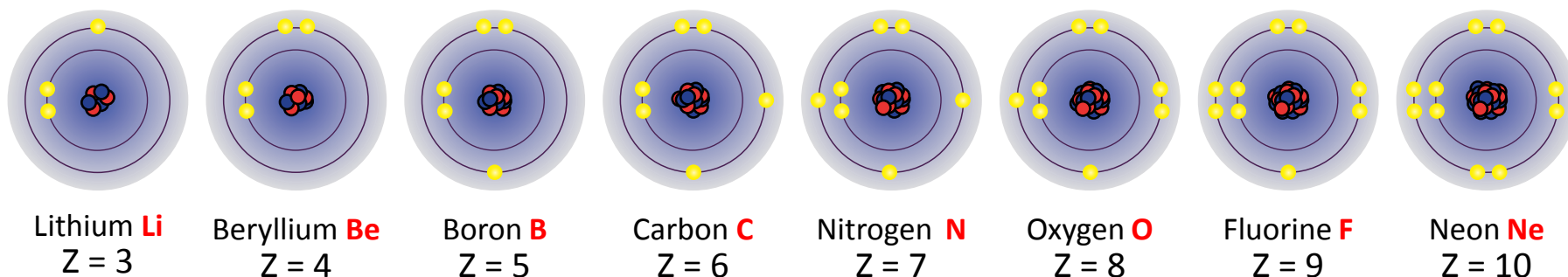


ELECTRON SHELL DIAGRAM
OF THE PERIOD 1 ELEMENTS

AS E RULE:

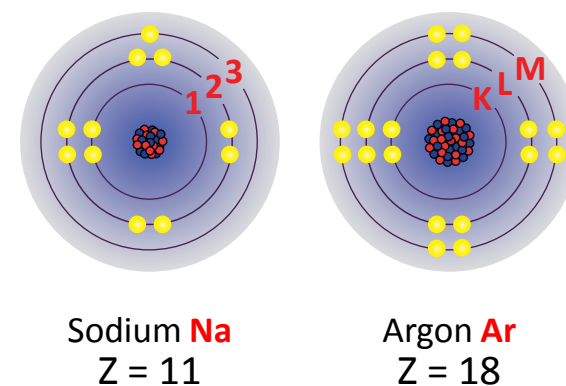
**THE PERIOD NUMBER IS EQUAL TO THE NUMBER OF ELECTRON SHELLS
OF THE ATOMS OF ELEMENTS.**

Lithium has two electron shells. The first K shell is filled with two electrons. The third electron is in the second L shell and lithium is the first element in the second period. With the gradually increase of the atomic numbers by 1 the number of protons in the nuclei of atoms also increases by 1 thus filling the second electron shell with up to 8 electrons:



ELECTRON SHELL DIAGRAMS OF THE PERIOD 2 ELEMENTS

The next chemical element is sodium Na (Z = 11). Its atoms have 11 protons in their nuclei. As the valence shell of the previous element (the noble gas neon Ne) is filled with the maximum possible number of electrons (according to the octet rule they are 8) the eleventh electron of sodium occupies the next M shell. Thus sodium has three electron shells and it is the first element of the third period. Along the period the valence shell of the atoms of the elements is gradually filled with electrons till the noble gas argon Ar. It is not a reactive element as its valence shell has 8 electrons.



ELECTRON SHELL DIAGRAMS OF THE FIRST AND THE LAST ELEMENT OF THE THIRD PERIOD

Each period begins with an element with 1 electron in its outer shell and ends with a noble gas with the stable configuration of 8 valence electrons (except He in the first period).

The elements are grouped in periods according to the number of electron shells in their atoms. But why are the chemical properties of the elements in a group similar? Lithium Li and sodium Na are alkali metals. They are group 1 (IA) elements and both have only 1 valence electron in their valence shells. Potassium has four electron shells but in the outermost shell it has only 1 valence electron too. Rubidium Rb (with five electron shells) and caesium Cs (with six electron shells) both also have only 1 valence electron in their valence shells. Magnesium Mg and the alkali earth metals (calcium Ca, strontium Sr and barium Ba) are group 2 (IIA) elements. All they have 2 valence electrons in their valence shells. Aluminum is a 13 (IIIA) group element and has 3 valence electrons.

AS A RULE:

**THE MAIN GROUP NUMBER (WITH ROMAN LETTERS IA, IIA, ... VIIIA)
IS EQUAL TO THE NUMBER OF VALENCE ELECTRONS OF THE ATOMS OF ELEMENTS.**

PROBLEM!

How many electron shells and how many valence electrons do the atoms of chlorine Cl and barium Ba have? Use the periodic table.

SOLUTION: Chlorine Cl is located in the **third** period and has **3 electron shells** in its atoms. It is a **VIIA group** element and has **7 valence electrons** in its M shell (the outermost shell).

Barium Ba is located in the **sixth period** and has **6 electron shells** in its atoms. It is a **IIA group** element and it has **2 valence electrons** in its outermost P shell.

PROBLEM!

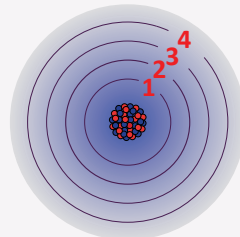
Draw the electron shell diagram of calcium Ca.
Use the periodic table.

SOLUTION: 1. Determine the number of electron shells, the total number of electrons and the number of valence electrons in calcium atoms:

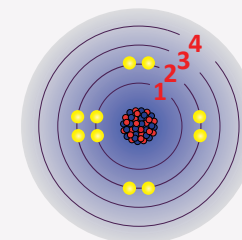
Calcium Ca is the 20th element in the periodic table. So it has 20 protons in the nuclei of its atoms and **20 electrons** in their electron shells.

Calcium is a period 4 and group IIA element. So its atoms have **4 electron shells** and **2 valence electrons**.

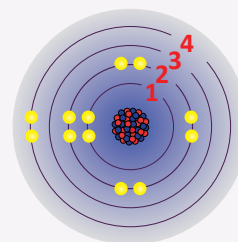
2. Draw the electron shells of calcium atoms:



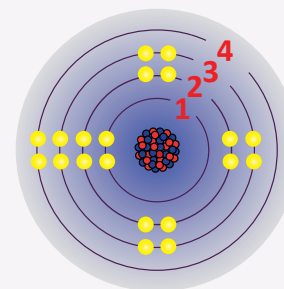
3. Fill the first two electron shells following the rule that each shell may hold up to $2n^2$ electrons (n is the number of the shell). If the shell is filled with the maximum number of electrons it can hold, all electrons are paired!



4. Put 2 electrons in the outer valence shell:
(If in a shell there are 2 electrons, they are paired.)



5. Dispose the remaining 8 electrons in the third M shell. As this is an octet all electrons are paired. Note that this shell is not filled with 18 electrons (the maximum number of electrons it can hold).



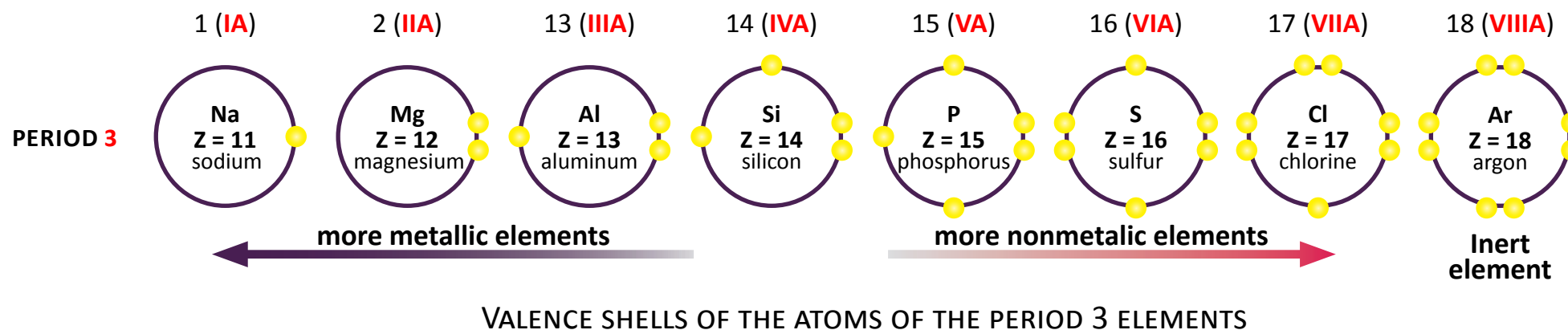
Electron shell diagram
of calcium Ca ($Z = 20$)

TRENDS IN THE PERIODIC TABLE

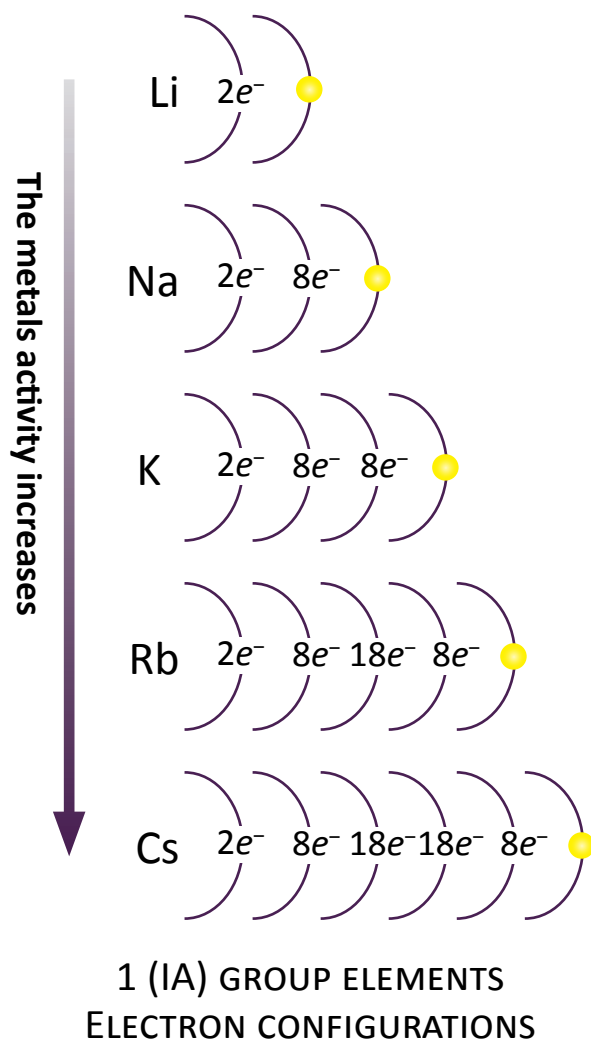
The structure of the atomic electron shell (called **electron configuration**) determines the place of the elements in the periodic table. Metals and nonmetals occupy different parts of the periodic table as their electron configurations are quite different. But along a period and in a group the properties of elements (their character) change gradually.

TRENDS ALONG A PERIOD

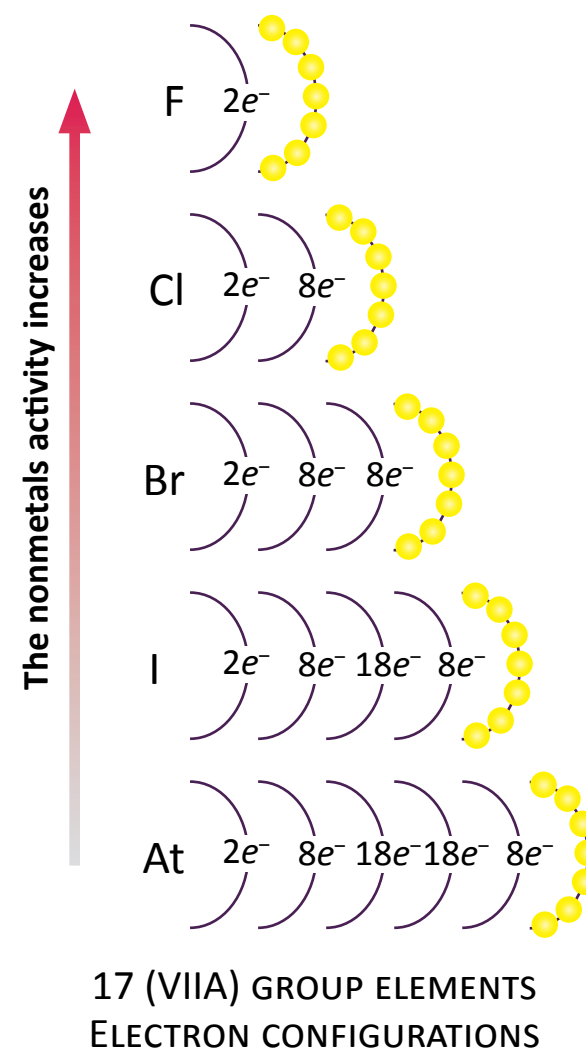
Each period starts with an alkali metal. All alkali metals have only one valence electron in their outer shells and they are the most active metals. Along a period with the gradually increase of the number of valence electrons elements become less metallic and more nonmetallic. The 17 (VIIA) group elements are typical nonmetals with 7 valence electrons in their atoms. For example the third period starts with sodium Na, a very reactive metal. Chlorine Cl is the last active element (a typical nonmetal). The period ends with argon Ar (an inert element with 8 valence electrons in its atoms):



TRENDS IN A MAIN GROUP



The atoms of the elements in a main group have the same number of valence electrons but different numbers of electron shells. The identical structure of the valence shells determines the similarity in their properties. But down a group, with the increase of the number of electron shells, elements become more metallic. Thus the most metallic element in 1 (IA) group (and the most active metal) is caesium Cs (francium Fr, just below it, is radioactive). The most nonmetallic element in 17 (VIIA) group (and the most active nonmetal) is fluorine F.

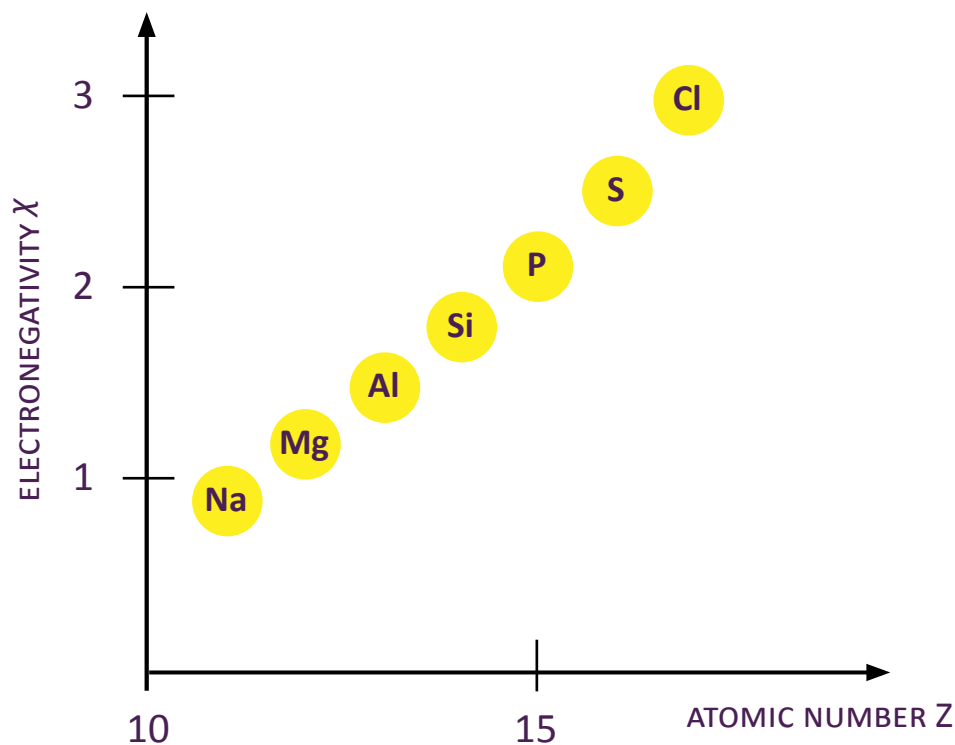


ATOMIC STRUCTURE AND PERIODIC TABLE KEY POINTS:

- ▶ The properties of elements depend on the electron configurations of their atoms.
- ▶ The periodic recurrence of similar properties of the chemical elements with the gradually increase of their atomic numbers is due to the periodical recurrence of identical outer shell electron configurations of their atoms.
- ▶ Each period starts with an element with 1 electron in its outer shell (an alkali metal) and ends with an inert element with a stable octet of valence electrons (a noble gas) (except helium He in the first period – its outer shell is stable with 2 electrons).
- ▶ The period number is equal to the number of electron shells of the atoms of chemical elements.
- ▶ The main group number is equal to the number of the electrons in the outer shell of the atoms of chemical elements.

ELECTRONEGATIVITY

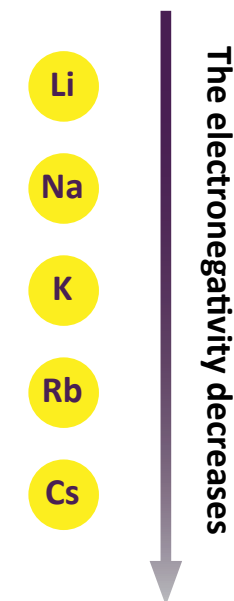
Atoms of different elements bond to one another by losing or gaining valence electrons thus forming shared electron pairs. The capability of atoms to attract the bonding pairs is called **electronegativity χ** . Like, for example the temperature, electronegativity is evaluated using specific scales. The Pauling scale of electronegativity is the most commonly used. The most electronegative element is fluorine F ($\chi = 4,0$). Bonded to atoms of other elements it always attracts the bonding pairs. The electronegativity of caesium Cs is only 0,7. So it is not capable to attract strongly the bonding pairs.



TREND IN ELECTRONEGATIVITY IN THE 3RD PERIOD

The electronegativity of elements changes periodically. It generally increases across a period from left to right and decreases down a group. The elements in the upper right-hand corner of the periodic table are therefore the most electronegative.

The difference in electronegativity of atoms is essential for the type of chemical bonds, as will be discussed in the next chapter.



TREND IN ELECTRONEGATIVITY IN THE 1 (IA) GROUP

The oxides and the hydroxides of typical metals differ greatly in their properties from the oxides and the hydroxides of typical nonmetals. The reason is the difference in the atomic electron configurations (and consequently the electronegativity) of the metallic and the nonmetallic elements. Thus the first element in a period always forms the typical for metals basic oxide and hydroxide (for example Li_2O , Na_2O , K_2O , LiOH , NaOH , KOH , etc.). Along a period with the increase of the atomic numbers the properties of the oxides and hydroxides gradually change and the elements at the end of the periods (except the inert gases) form the typical for nonmetals acidic oxide and hydroxide (for example dichlorine heptaoxide Cl_2O_7 and perchloric acid HClO_4). Down a main group the properties of oxides and hydroxides also change gradually as the elements become more metallic. You will become more familiar with these trends after you study the chapters about metals and nonmetals.

CAN YOU ANSWER THESE QUESTIONS ?

1. Why are elements arranged in periods and groups in the periodic table?
2. Why is there a periodic recurrence of similar properties of the elements?
3. Why do metals and nonmetals differ greatly in their properties?
4. Why are calcium Ca and bromine Br period 4 elements?
5. Why are nitrogen N and phosphorus P group 15 (VA) elements?
6. Why is calcium Ca less metallic than barium Ba?
7. Why is oxygen O more nonmetallic than sulfur S?
8. Why are noble gases inert elements?
9. Why is fluorine F the most electronegative element?
10. Why is metals electronegativity low compared to nonmetals electronegativity?

3

CHEMICAL BONDING

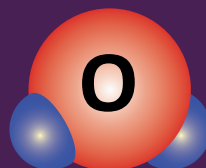
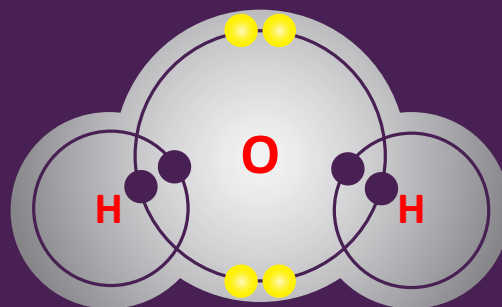
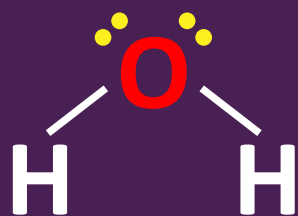
Covalent bonding. Simple and multiple bonds

KEY TOPICS

1. How do atoms bond to other atoms?
2. What is a covalent bond?
3. Why are there simple and multiple covalent bonds?
4. How to draw Lewis-dot formula of molecules?
5. What are atomic and molecular lattices?

KEY WORDS

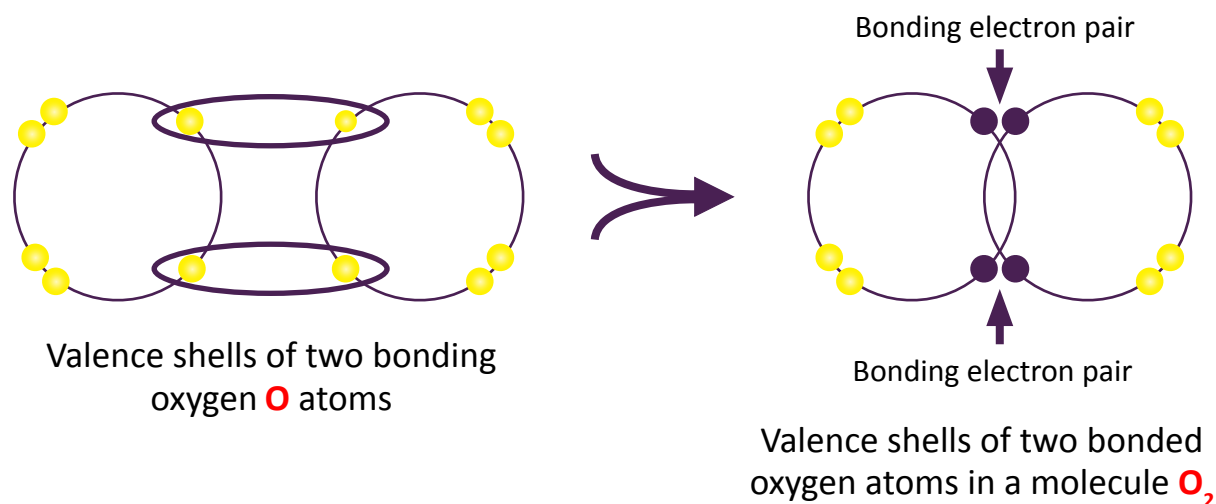
- Bonding electron pair
- Lone electron pair
- Covalent bond
- Single bond
- Multiple bond
- Double bond
- Triple bond
- Atomic lattice
- Molecular lattice



HOW DO ATOMS BOND TO OTHER ATOMS?

Simple substances are built up of identical atoms. Oxygen molecules O_2 are built up of two oxygen atoms, nitrogen molecules N_2 of two nitrogen atoms, metals are built up of many metallic atoms of an element (sodium Na, calcium Ca, etc.), carbon C exists in several forms (amorphous, graphite, diamond, fullerenes, graphene) all built up of carbon atoms. The great variety of substances, naturally occurring and man made, is due to the famous property of the atoms of an element to bond to atoms of other elements thus forming molecules of chemical compounds. But how does all this work?

Atoms always use their valence electrons (not necessary all valence electrons) to bond to other atoms – valence electrons of the two bonding atoms form electron pairs shared between them thus generally acquiring a stable octet configuration. For example oxygen atoms have 6 valence electrons – two electron pairs and two single electrons. Each oxygen atom needs two more electrons to form a stable octet. When two oxygen atoms bond to one another the four single valence electrons form two bonding electron pairs shared between atoms. Thus both atoms achieve a stable octet configuration of the outermost electron shell and an oxygen molecule O_2 is formed:

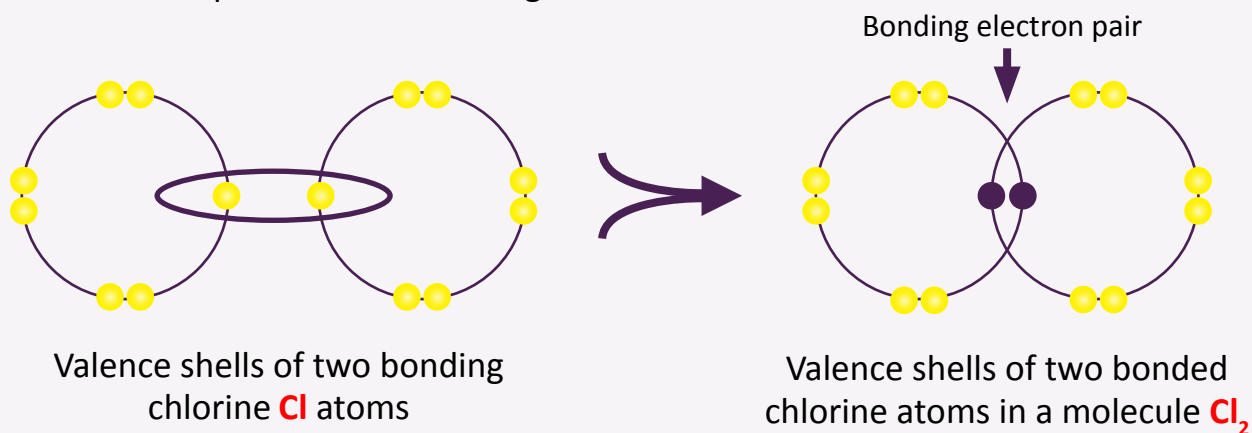


2 OXYGEN ATOMS BOND TO FORM AN OXYGEN MOLECULE O_2

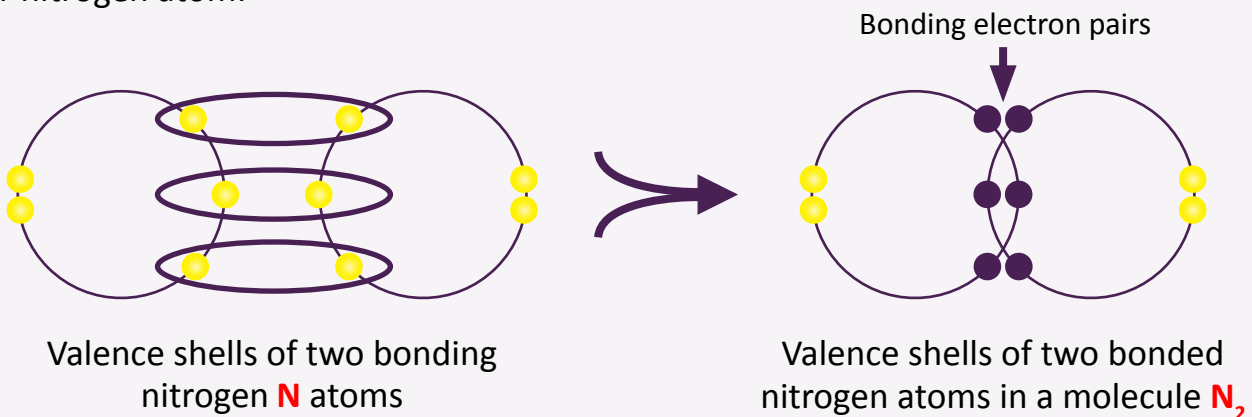
PROBLEM!

How do chlorine Cl and nitrogen N bond to form the molecules of their simple substances – Cl_2 and N_2 ?

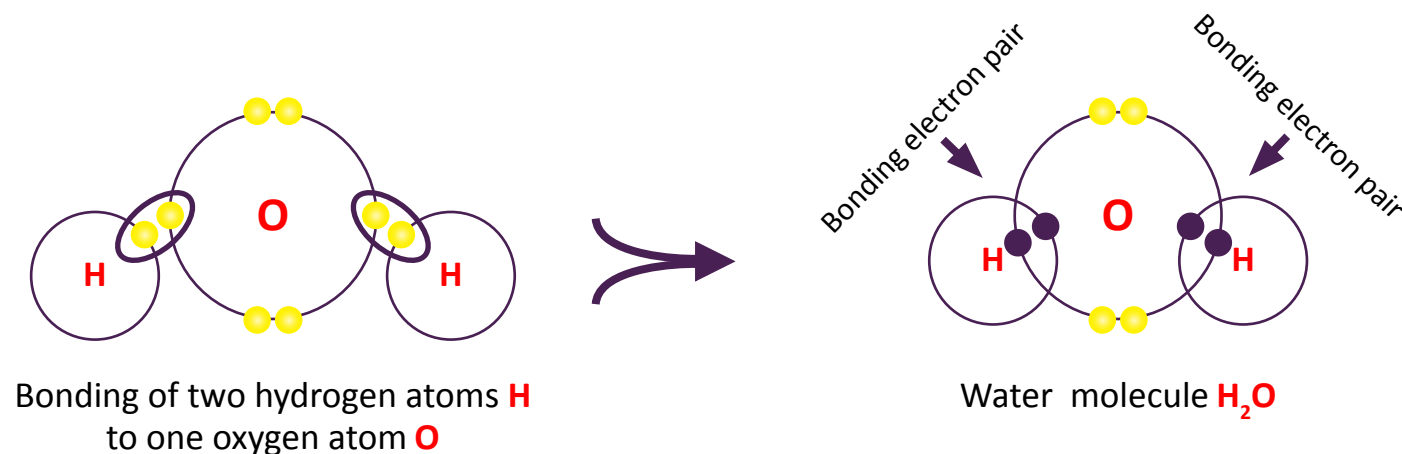
SOLUTION: Chlorine Cl is a 17 (VIIA) group element and has 7 valence electrons (3 pairs of electrons and one single electron). Its atoms need only 1 electron to form a stable octet of electrons. If two chlorine atoms bond using their single electrons each atom acquires the octet configuration:



Nitrogen is a 15 (VA) group element. It has 5 valence electrons (one electron pair and 3 single electrons). To acquire a stable 8-electrons configuration of the valence shell it forms three shared electron pairs with the single electrons of another nitrogen atom:

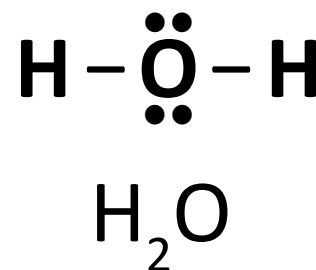
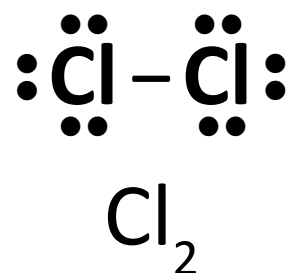
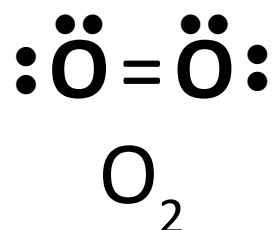
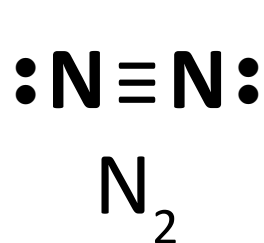


The bonding of atoms of different elements follows the same rule – generally, bonded atoms acquire a stable octet configuration of their valence shells (only hydrogen atoms need 2 electrons). For example in the water molecule H_2O the oxygen atom uses its two single electrons to form two bonding electron pairs with the single electrons of two hydrogen atoms:



2 HYDROGEN ATOMS BOND TO 1 OXYGEN ATOM
TO FORM A WATER MOLECULE H_2O

Lewis-dot formula are often used to represent the bonds formation. The shared electron pairs are represented with small lines „–“ linking the chemical symbols of the bonding atoms. One line corresponds to one bonding electron pair. The electron pairs that are not used in chemical bonding (called „lone“ pairs) are represented with double dots near the chemical symbol of the element. Thus the Lewis-dot formulae of N_2 , O_2 , Cl_2 , H_2O look like that:



PROBLEM!

Draw the Lewis dot formula of hydrogen chloride HCl.

SOLUTION: 1. Calculate the total number of valence electrons:

Hydrogen H is the first element of the periodic table and it has only 1 electron in its atoms. **Chlorine Cl** is a 17 (VIIA) group element and has 7 valence electrons. The total number of valence electrons of the HCl molecule is $1 + 7 = 8$.

2. Draw one bonding electron pair: $\text{H} - \ddot{\text{Cl}}\text{:}$

Calculate the number of electrons in the valence shells of both atoms – hydrogen atom has 2 electrons and chlorine atom has 8 electrons. Both atoms acquire a stable configurations of their valence shells – hydrogen atoms with 2 electrons and chlorine atoms with an octet of valence electrons.

In the HCl molecules the hydrogen atom is bonded to the chlorine atom by one bonding electron pair.

PROBLEM!

Draw the Lewis-dot formula of carbon monoxide CO.

SOLUTION: 1. Calculate the total number of valence electrons:

Carbon C is a 14 (IVA) group element and has 4 valence electrons. **Oxygen O** is a 16 (VIA) group element and has 6 valence electrons. The total number of valence electrons in the CO molecule is $4 + 6 = 10$.

2. Draw one bonding electron pair: $:\dot{\text{C}} - \dot{\text{O}}:$

Calculate the number of electrons in the valence shells of both atoms – carbon atom has 5 electrons and oxygen atom has 7 electrons. Both atom do not acquire the octet configuration.

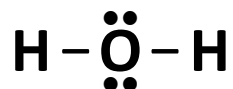
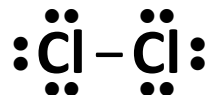
3. Draw two bonding electron pairs: $:\text{C} = \ddot{\text{O}}:$

Calculate the number of electrons in the valence shells of both atoms – carbon atom has 6 electrons and oxygen atom has 8 electrons. Now oxygen is with a stable octet in its valence shell but carbon atom needs two more electrons.

4. Draw three bonding electron pairs: $:\text{C} \equiv \text{O}:$

Calculate the number of electrons in the valence shells of both atoms – carbon atom has 8 electrons and oxygen atom has 8 electrons. Now both atoms have a stable octet configuration.

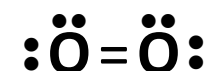
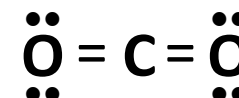
In the CO molecules a carbon atom is bonded to an oxygen atom by three bonding electron pairs.



SINGLE
COVALENT BONDS

In the examples above atoms in molecules are held together thanks to one, two or three bonding electron pairs. This type of chemical bonding is known as **covalent bonding**. If two atoms bond by means of one bonding electron pair the bond is called **simple** or **single covalent bond**. For example the Cl – Cl bond in the Cl₂ molecule, also the O – H bonds in the water molecule and the H – Cl bond in the HCl molecule. If there are two or three bonding electron pairs the bond is called **multiple** (**double or triple**) **covalent bond**. Thus the oxygen atoms in the O₂ molecule are bonded by a double covalent bond O = O. The nitrogen atoms in the N₂ molecule are bonded by a triple covalent bond N ≡ N. The bond in the CO molecule is also a triple covalent bond C ≡ O.

COVALENT BONDING is a chemical link of two atoms by means of one, two or three bonding electron pairs shared by atoms.



DOUBLE
COVALENT BONDS



TRIPLE
COVALENT BONDS

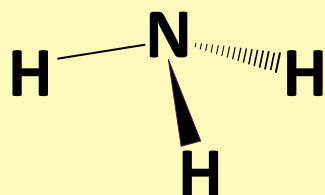
HOW TO DRAW MOLECULES?

Molecules of chemical compounds and simple substances are more often represented with their **molecular formulae** (for example H_2O , Cl_2 , H_2 , NaOH , etc.). The only information we get is about the elements building molecules and the exact number of each type of atoms. But although molecules are very, very small, they are three dimensional structures and their properties may depend strongly on the 3D arrangement of the atoms (especially if molecules are built up of many atoms like most of the organic compounds). That is why in chemistry, in addition to the molecular formulae, scientists have introduced the so called **structural formulae**. They represent not only the type and quantity of atoms in a molecule but also the molecular structure. **Lewis-dot formulae** are **2D structural formulae**. They illustrate the type of bonding in a molecule but do not give any idea about the 3D arrangement of the atoms.

It is interesting to know:

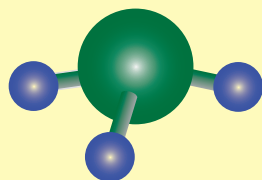
Three types of 3D structural formula are widely used:

- **line-wedge-dash model** – atoms are represented with the chemical symbols of the elements. Chemical bonds are represented with a dash line showing that an atom is away from you, a wedge showing that an atom is toward you and a straight line for bonds on the plane of the sheet;



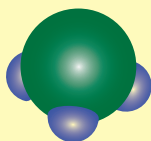
ammonia NH_3

- **ball-and-stick model** – atoms are represented as spheres. Chemical bonds are represented as cylinders. The size of spheres does not correspond to the true dimensions of atoms. Different colours are used for different types of atoms;



ammonia NH_3

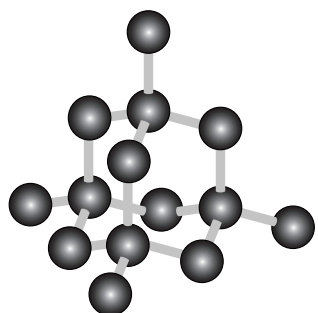
- **space-filling model** – atoms are represented as spheres (different colours are used for different types of atoms). They penetrate into one another to show the bonding of atoms.



ammonia NH_3

WHAT IS AN ATOMIC AND A MOLECULAR LATTICE OF A COVALENT SUBSTANCE?

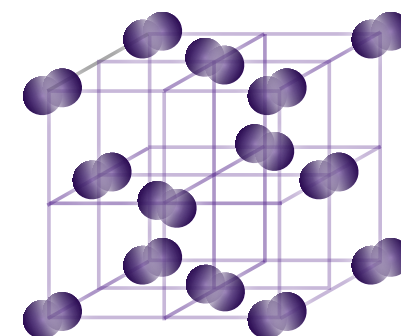
At **STP*** most of the covalent substances are gases (nitrogen N_2 , oxygen O_2 , carbon oxide CO , carbon dioxide CO_2 , hydrogen chloride HCl , etc.) or liquids (bromine Br_2 , water H_2O , sulfuric acid H_2SO_4 , nitric acid HNO_3 , many organic compounds). But covalent bonding occurs also in some solids. When molecules (or atoms if a simple substance is built up of many covalently bonded identical atoms) are randomly positioned to one another solids are amorphous. For example coal is an amorphous simple covalent substance built up of carbon atoms. But if covalent molecules or covalently bonded identical atoms are arranged in some specific pattern and form 3D structures with a definite geometry, covalent solids are crystalline substances with atomic or molecular crystal lattices.



ATOMIC LATTICE
OF DIAMOND
Carbon atom 

The **atomic lattice** represents the regular arrangement of covalently bonded identical atoms in space. Well known is the lattice of diamond. It is formed of many tetrahedrons built up of five carbon atoms – one atom in the center and four at the vertices. Thus each carbon atom is bonded to four carbon atoms. The strong covalent carbon-carbon bonds hold carbon atoms together thus forming a giant 3D crystal structure (a giant molecule).

Iodine I_2 is also a simple solid covalent substance but it is built up of diatomic molecules I_2 . It is a crystalline solid and its structure is represented by a **molecular lattice**. Inside the molecule the iodine atoms are linked by a strong covalent bond. But the crystal is held together by intermolecular forces of attraction which are much weaker. That is why iodine easily sublimates when heated. Ice and solid CO_2 are also crystalline covalent substances with molecular lattices.



MOLECULAR LATTICE OF I_2
 I_2 molecule 

***STP** – **S**tandard **T**emperature 273,15 K (0 °C) and **P**ressure 101 325 Pa

COVALENT BONDING KEY POINTS:

- ▶ **Covalent bond** is a chemical link between two atoms realized thanks to pairs of shared by the two atoms electrons. These electron pairs are called „bonding“ (or „shared“) electron pairs. Generally, bonding atoms acquire an octet configuration of the outer electron shell.
- ▶ Covalent bonds with one bonding pair are called **simple** or **single bonds**.
- ▶ Covalent bonds with more than one bonding electron pair are called **multiple bonds** – **double** (with 2 bonding electron pairs) and **triple** (with 3 bonding electron pairs) bonds.
- ▶ Covalent substances are most often gases or liquids. Solid covalent substances are amorphous or crystalline with atomic or molecular lattices.

CAN YOU ANSWER THESE QUESTIONS ?

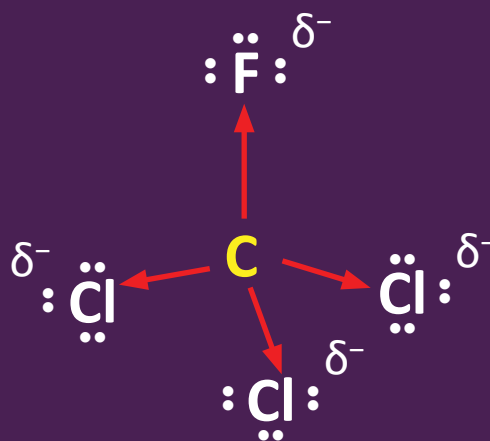
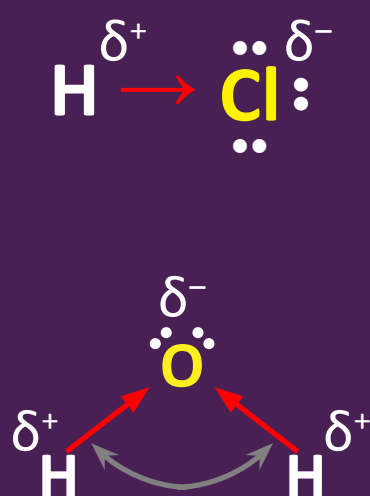
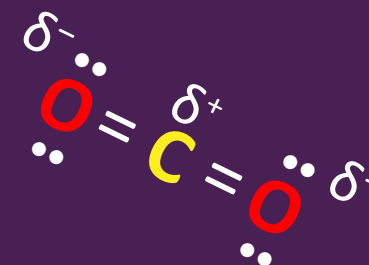
1. Do atoms use all their electrons to bond to other atoms?
2. Do atomic nuclei change when atoms bond to each other?
3. What is a bonding (shared) electron pair?
4. What is a single covalent bond?
5. What is a multiple covalent bond?
6. Is the bond between the two H atoms in the hydrogen molecule H_2 single or multiple? Why?
7. Why are the two nitrogen atoms of the N_2 molecule linked by a triple covalent bond?
8. Can you give examples of molecules with single covalent bonds?

4

Polar and nonpolar covalent bonds Polar and nonpolar molecules

KEY TOPICS

1. Why is electronegativity so important?
2. What are polar and nonpolar bonds?
3. What are polar and nonpolar molecules?
4. How does polarity affect the properties of substances?



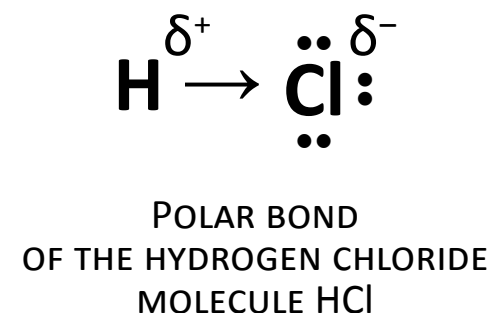
KEY WORDS

- Polar bond
- Nonpolar bond
- Polar molecule
- Nonpolar molecule

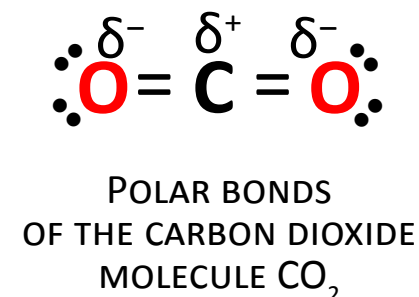
WHY IS ELECTRONEGATIVITY A CHEMICAL BONDING KEY FACTOR?

When atoms bond to each other they use their valence electrons to form shared electron pairs. These bonding pairs are attracted by both nuclei. If two identical atoms bond (for example the molecules of H_2 , O_2 , N_2) the shared electron pairs (one, two or three) are equally attracted by two identical nuclei. These covalent bonds are **nonpolar** – none of the bonded atoms has electric charge. But when different atoms bond the more electronegative atom attracts to itself the bonding pair thus acquiring a partial negative charge denoted by the Greek letter δ^- . The other atom acquires a partial positive charge denoted by the Greek letter δ^+ . As a result of the unequal sharing of the bonding electron pair an electric dipole* is constituted. These covalent bonds are called **polar** bonds.

For example in the diatomic molecules of hydrogen chloride HCl a hydrogen atom is linked to a chlorine atom by a single covalent bond. The chlorine atom is more electronegative than the hydrogen atom and it attracts to itself the bonding pair. Thus a dipole is formed – chlorine atom has a partial negative charge δ^- while hydrogen atom has a partial positive charge δ^+ . The polar covalent bond is often denoted by an arrow „ $\rightarrow$ “ in the direction of the shifting of the shared electron pair.



In the molecule of carbon dioxide CO_2 the less electronegative carbon atom is linked by two double covalent bonds to two oxygen atoms. Each oxygen atom attracts to itself two bonding electron pairs. As a result oxygen atoms acquire partial negative charges δ^- . The carbon atom has a partial positive charge δ^+ .



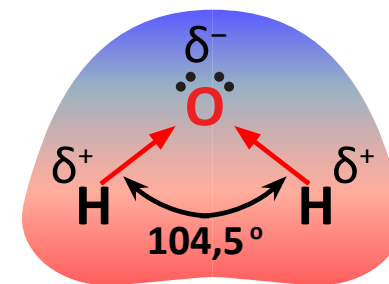
* **Electric dipole** is a pair of equal and opposite electric charges separated over a short distance.

POLAR AND NONPOLAR MOLECULES

As covalent bonds may be polar or nonpolar, molecules built up of covalently bonded atoms also may be polar or nonpolar. Molecules built up of two identical atoms are always **nonpolar**, i.e. they are not dipoles because the bonding pairs are evenly shared between atoms. For example the molecules of some simple substance – O_2 , N_2 , F_2 , Cl_2 , Br_2 , I_2 . Generally, diatomic molecules built up of covalently bonded different atoms are **polar**, i.e. they are dipoles. For example the hydrogen halides (HCl, HBr, etc.), carbon monoxide CO.

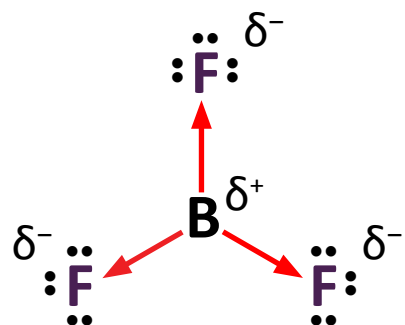
But it is not so simple if molecules are built up of three, four or more atoms (for example CO_2 , H_2O , H_2S , NH_3 , CH_4 , H_2SO_4 , etc.). Generally, these molecules may be polar or nonpolar depending on the atoms electronegativity and molecules shape. Triatomic molecules may be linear or bent. For example carbon dioxide CO_2 molecules are linear, while water molecules H_2O are bent (the angle between the two O–H bonds in the molecule is $104,5^\circ$). Look at the picture of CO_2 on the previous page. The molecule of carbon dioxide CO_2 is linear and symmetrical. The oxygen atoms are evenly distributed around the central carbon atom. As a result the partial positive charge of the carbon atom cancels the partial negative charges of the oxygen atoms and the CO_2 molecule is nonpolar.

The water molecule H_2O is polar. Because of its bent shape the partial electric charges are unevenly distributed. One side of the molecule is more positive (coloured in red), while the other side of the molecule is more negative (coloured in blue). As a result H_2O molecules are dipoles. But remember that in spite of their polarity molecules are always electrically neutral units (in the water molecule there are 10 protons and 10 electrons).



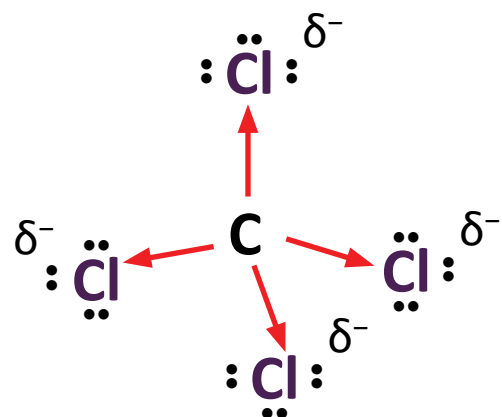
THE WATER MOLECULE

H_2O IS POLAR



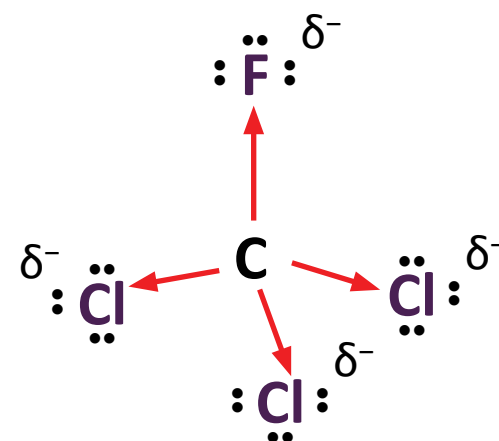
BF_3 (BORON TRIFLUORIDE)
MOLECULES ARE PLANAR AND
NONPOLAR

Trigonal planar molecules built up of one central atom and three other identical atoms are nonpolar (for example BF_3). Boron **B** has a partial positive charge while the more electronegative fluorine atoms **F** have partial negative charges. As a whole molecules are nonpolar as the outside atoms are evenly distributed around the central atom.



CCl_4 (TETRACHLOROMETHANE)
MOLECULES ARE TETRAHEDRAL
AND NONPOLAR

Tetrahedral molecules built up of one central atom and four identical atoms are also nonpolar. For example tetrachloromethane CCl_4 (one of the simplest organic compounds) – the bonds $\text{C}-\text{Cl}$ are polar but the partial positive charge of the central carbon atom **C** cancels the partial negative charges of the symmetrically distributed around it more electronegative chlorine atoms **Cl**. But if one of the outside chlorine atoms is substituted by a fluorine atom **F** the molecule becomes a dipole, i.e. it is polar. The reason is the high electronegativity of fluorine atoms. The bond $\text{C}-\text{F}$ is more polar than the three $\text{C}-\text{Cl}$ bonds and the positive charge of the central atom cannot cancel the negative charges of the outside atoms. The compound CCl_3F is known as freon-11. It can cause the degradation of the ozone layer.

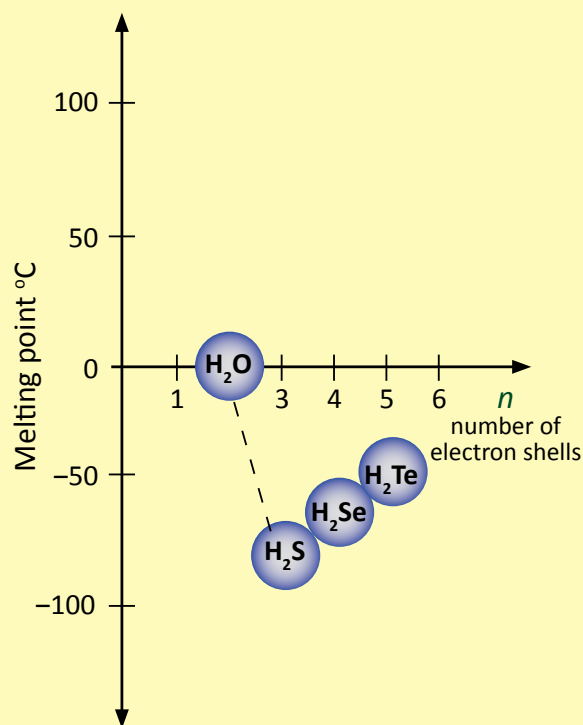


CCl_3F (TRICHLOROFLUOROMETHANE)
MOLECULES ARE TETRAHEDRAL AND
POLAR

It is interesting to know:

WHY IS BONDS AND MOLECULES POLARITY SO IMPORTANT?

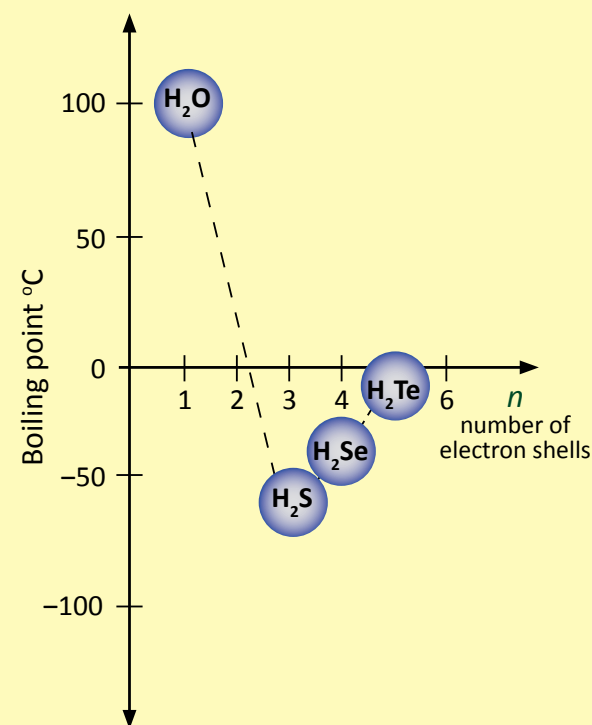
The bonds and molecules polarity is a key factor for the chemical activity of both simple substances and chemical compounds. Nonpolar bonds are stable and it is difficult to break them down in chemical reactions. Thus at STP nitrogen nonpolar molecules N_2 are practically nonreactive. (That is why the Earth atmosphere consists of more than 70 % of nitrogen N_2).



MELTING POINTS OF THE 16 (VIA) GROUP HYDROGEN COMPOUNDS

Polar bonds are more easily disrupted in chemical reactions and substances are more reactive. Thus HCl molecules dissociate in water forming H^+ ions and Cl^- ions and the hydrochloric acid is a very reactive acid.

Molecules polarity may also influence some physical properties of the substances. For example water boils at $100\text{ }^{\circ}C$ and melts at $0\text{ }^{\circ}C$. But these temperatures are too high if compared to the boiling and melting points of the other hydrogen compounds of the 16 (VIA) group elements.

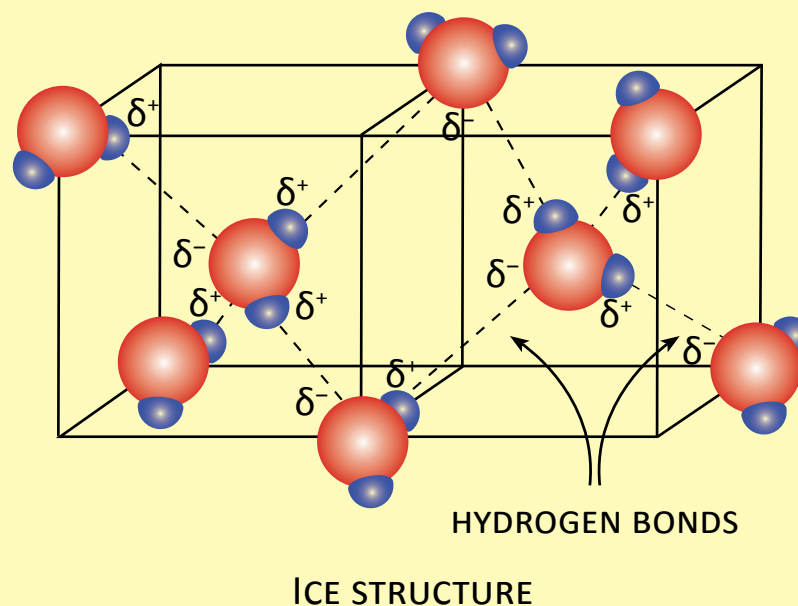


BOILING POINTS OF THE (16) VIA GROUP HYDROGEN COMPOUNDS

The reason is the high polarity of the O–H bonds in the water molecule. The electronegativity of sulfur **S**, selenium **Se** and **Te** is much less than the electronegativity of oxygen **O** and the bonds in their hydrogen compounds (H_2S , H_2Se and H_2Te) are not so polar as the O – H bonds in the water molecule H_2O . The water molecules dipoles are attracted by electrostatic forces – the partial positive charge of a hydrogen atom in a water molecule is electrostatically attracted by the partial negative charge of the oxygen atom of another water molecule (more precisely by one of the lone pairs of the oxygen atom).

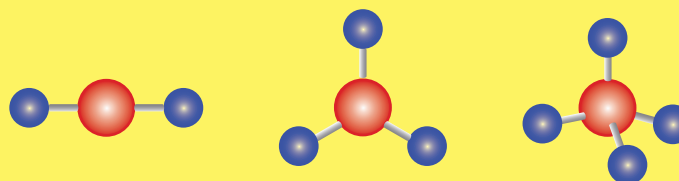
These intermolecular interactions are considered as an intermolecular bond, called **hydrogen bond**. This bond is much weaker than the O – H bonds in the water molecules but an additional amount of energy is needed to overcome it when heating water to boiling temperature or ice to melting temperature. Hydrogen bonding is also the reason for the low density of ice as compared to water. Generally, hydrogen bonding occurs between polar molecules with H-atoms covalently bonded to electronegative atoms with lone (unshared) electron pairs (F, N, O).

Hydrogen bonds play an important role in organic biological molecules. They are essential for the shape of the protein molecules, for holding the double helix in DNA, etc.



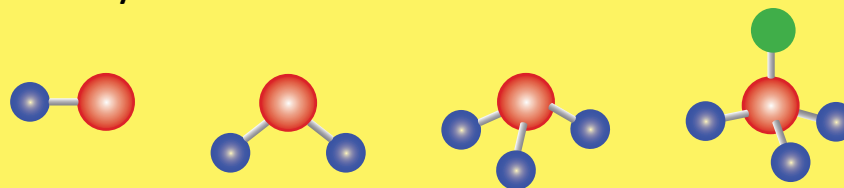
POLAR AND NONPOLAR COVALENT BONDS AND MOLECULES KEY POINTS:

- ▶ Identical atoms of nonmetals link to each other by covalent nonpolar bonds.
- ▶ Bonds are polar if there is a difference in the electronegativity of the two bonding atoms. The greater the electronegativity difference between bonding atoms the more polar the bond is.
- ▶ Covalently bonded atoms form polar or nonpolar molecules.
- ▶ Molecules built up of identical atoms and molecules built up of a central atom and symmetrically distributed around it identical atoms are nonpolar. The partial negative charges cancel the partial positive charges:



SYMMETRICAL MOLECULES

- ▶ Polar molecules are not symmetrical:



NONSYMMETRICAL MOLECULES

CAN YOU ANSWER THESE QUESTIONS ?

1. What is a polar bond?
2. Why are some covalent bonds nonpolar and others – polar?
3. Why is electronegativity a key parameter for the covalent bonds polarity?
4. Hydrogen is covalently bonded to fluorine in the molecules of hydrogen fluoride HF. Is the bond H – F polar or nonpolar? Why?
5. What is a polar molecule?
6. Is polar or nonpolar the molecule of carbon monoxide CO? Why?
7. Why are tetrahedral molecules with identical outside atoms always nonpolar?
8. Why are nonsymmetrical molecules with covalently bonded atoms polar?
9. Does polarity affect the physical and chemical properties of the chemical compounds?
10. Can you explain why ice floats over the oceans?

5 IONIC BONDING

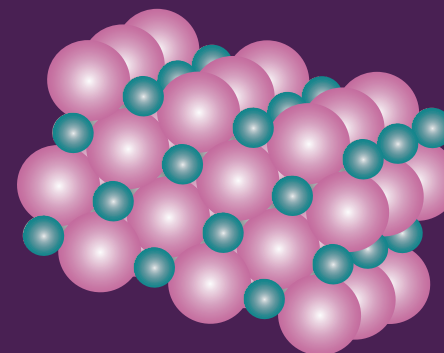
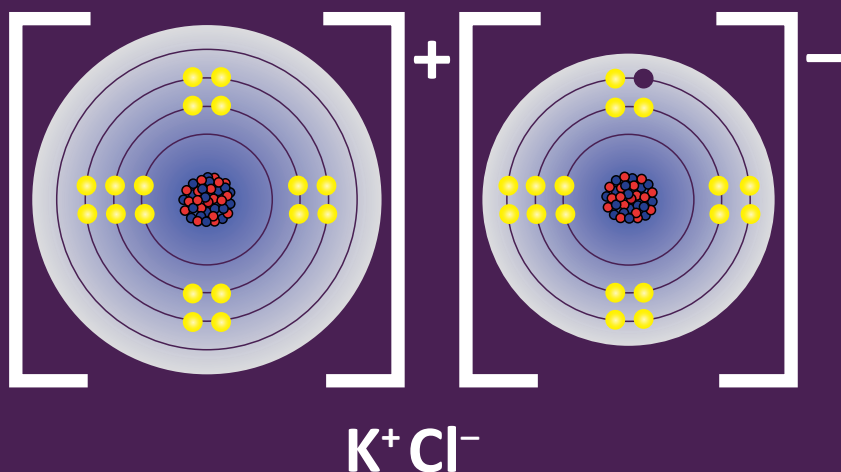
METAL - NONMETAL

KEY TOPICS

1. What is an ionic bond?
2. What is an ionic compound?
3. Why is table salt an ionic compound?
4. What is an ionic lattice?

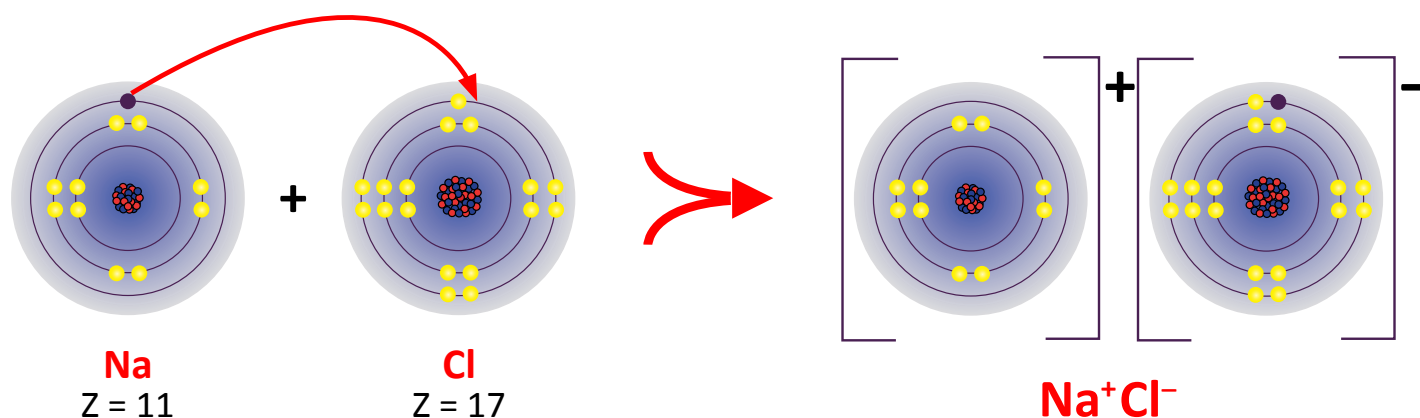
KEY WORDS

- Ionic bond
- Ionic crystal lattice



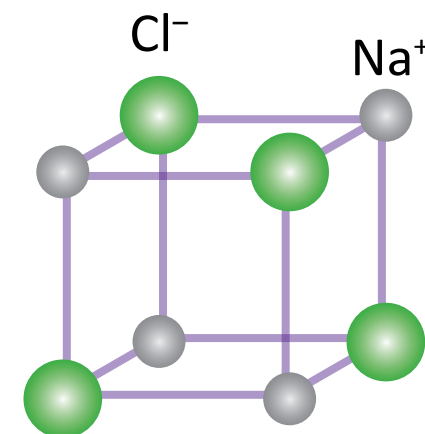
WHAT DOES AN IONIC BOND STAND FOR?

When two nonmetals bond they are linked to one another by a simple or a multiple covalent bond. Depending on the atoms electronegativity covalent bonds are nonpolar (bonding pairs are equally attracted by both nuclei) or polar (bonding pairs are attracted by the atom with higher electronegativity). But what happens if a metallic atom bonds to a nonmetallic atom? For example in the reaction of sodium Na with chlorine Cl_2 the well known compound sodium chloride NaCl is produced. The sodium electronegativity is only 0,9 while the chlorine electronegativity is much higher – 3,0. So the difference in electronegativity of the bonding atoms is 2,1. This means that the bonding pair is strongly attracted by the chlorine atom. It may be considered that one electron from a neutral sodium atom is transferred to a neutral chlorine atom:

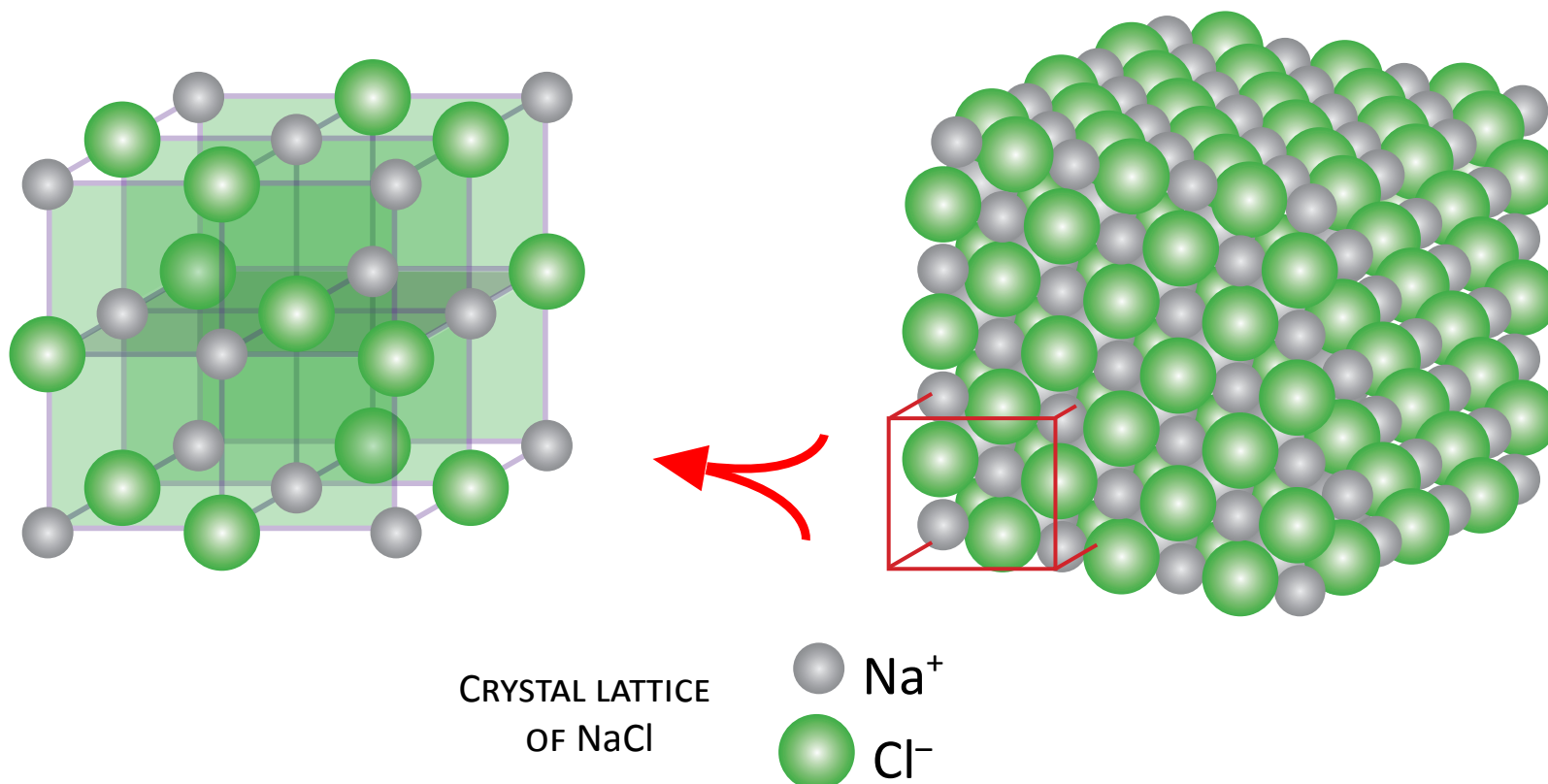


The neutral sodium atom Na loses its single valence electron and transforms into a positive ion Na^+ with an octet of electrons in the outer **L** shell. The neutral chlorine atom gains one electron and transforms into a negative ion Cl^- thus also acquiring an octet electron configuration of its outer **M** shell. The Na^+ and Cl^- ions are held together by an electrostatic force of attraction – the oppositely charged ions are linked by an **ionic bond**. So the table salt NaCl is an ionic crystalline chemical compound.

Ionic compounds consist of regular three-dimensional arrangements of positive and negative ions held together by electrostatic forces and forming the so called **ionic crystal lattices**. Their chemical formulae represent the simplest ratio between the numbers of the oppositely charged neighbour ions. For example the NaCl crystal lattice is built of many identical cubes (unit cells) arranged side by side, one above the other and one below the other. In the ionic lattice each positive ion Na^+ is surrounded by 6 symmetrically positioned Cl^- ions and each negative ion Cl^- is surrounded by 6 symmetrically positioned Na^+ ions. So the ratio is 6:6. The simplest is 1:1 and the ionic compound is represented by its chemical formula NaCl.



UNIT CELL OF THE NaCl
IONIC LATTICE.



Most of the ionic compounds are solids with relatively high melting and boiling points. Most of the salts are ionic compounds – for example KI, CaCl_2 , NaI, CsF, the oxides of metals as Na_2O , CaO, also NaOH, KOH, Ca(OH)_2 , etc. If soluble in water they dissociate into the corresponding ions. Ionic compounds in solution or in molten state conduct electricity.

IONIC BONDING KEY POINTS:

- ▶ Ionic bond is a chemical link between two atoms caused by the electrostatic attraction between oppositely charged ions forming an ionic compound.
- ▶ Ionic bond is formed between atoms with a great difference in electronegativity – usually between metals and nonmetals. The metallic atom loses electrons and transforms into a positively charged ion. The nonmetallic atom gains electrons and transforms into a negatively charged ion.
- ▶ An ionic bond is considered to be an extremely highly polar covalent bond. It is assumed that the bond is ionic if the difference in the electronegativity of the two bonded atoms exceeds 1,7 ($\Delta\chi > 1,7$).
- ▶ Ionic compounds are solid crystalline substances. Oppositely charged ions form ionic crystal lattices.

CAN YOU ANSWER THESE QUESTIONS ?

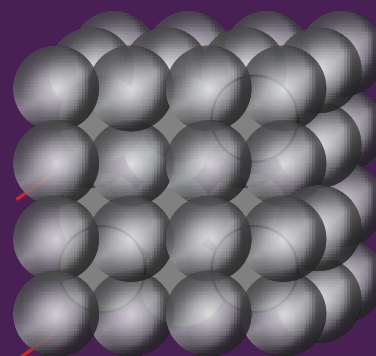
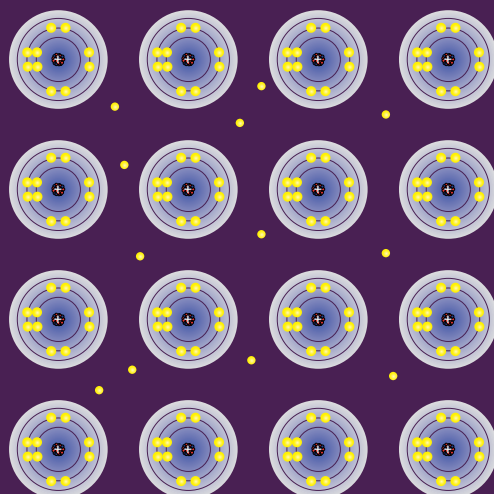
1. What is the difference between ionic and covalent bonding?
2. Why are usually metallic atoms linked to nonmetallic atoms by an ionic bond?
3. What must be the most „ionic“ diatomic salt?
4. Which of the substances are ionic compounds: Li_2O , KCl , CaO , HCl , Na_2S , Cl_2 , KOH ?
5. What is an ionic crystal lattice?
6. Why do molten and dissolved in water ionic compounds conduct electricity?
7. Why are ionic compounds solid at STP?
8. Why do ionic compounds melt and boil at a relatively high temperature?

6

METALLIC BONDING

KEY TOPICS

1. What is a metallic bond?
2. Why do metals conduct heat and electricity?
3. Why can metals be formed in different shapes and drawn out into thin wires?
4. What is a metallic lattice?
5. Why do metals melt at a high temperature?



Na

Cu Fe
Zn Li Al
Ag Au Ba
K
Ca Mg

KEY WORDS

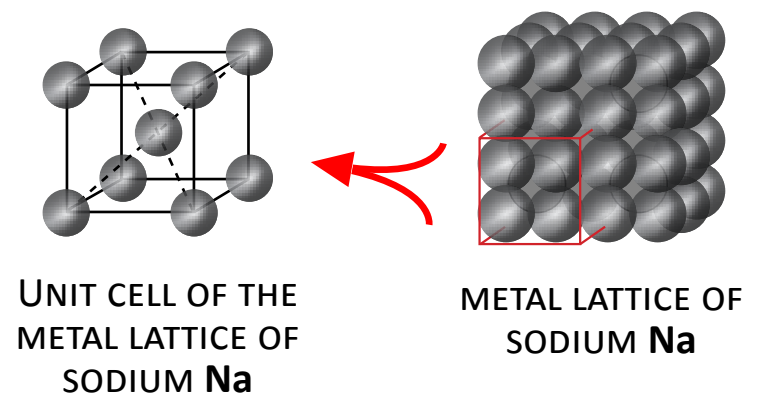
- Metallic bond
- Metallic crystal lattice
- Delocalized electrons (Sea of electrons)
- Malleability
- Ductility
- Metallic lustre

WHY ARE METALS MALLEABLE AND DUCTILE?

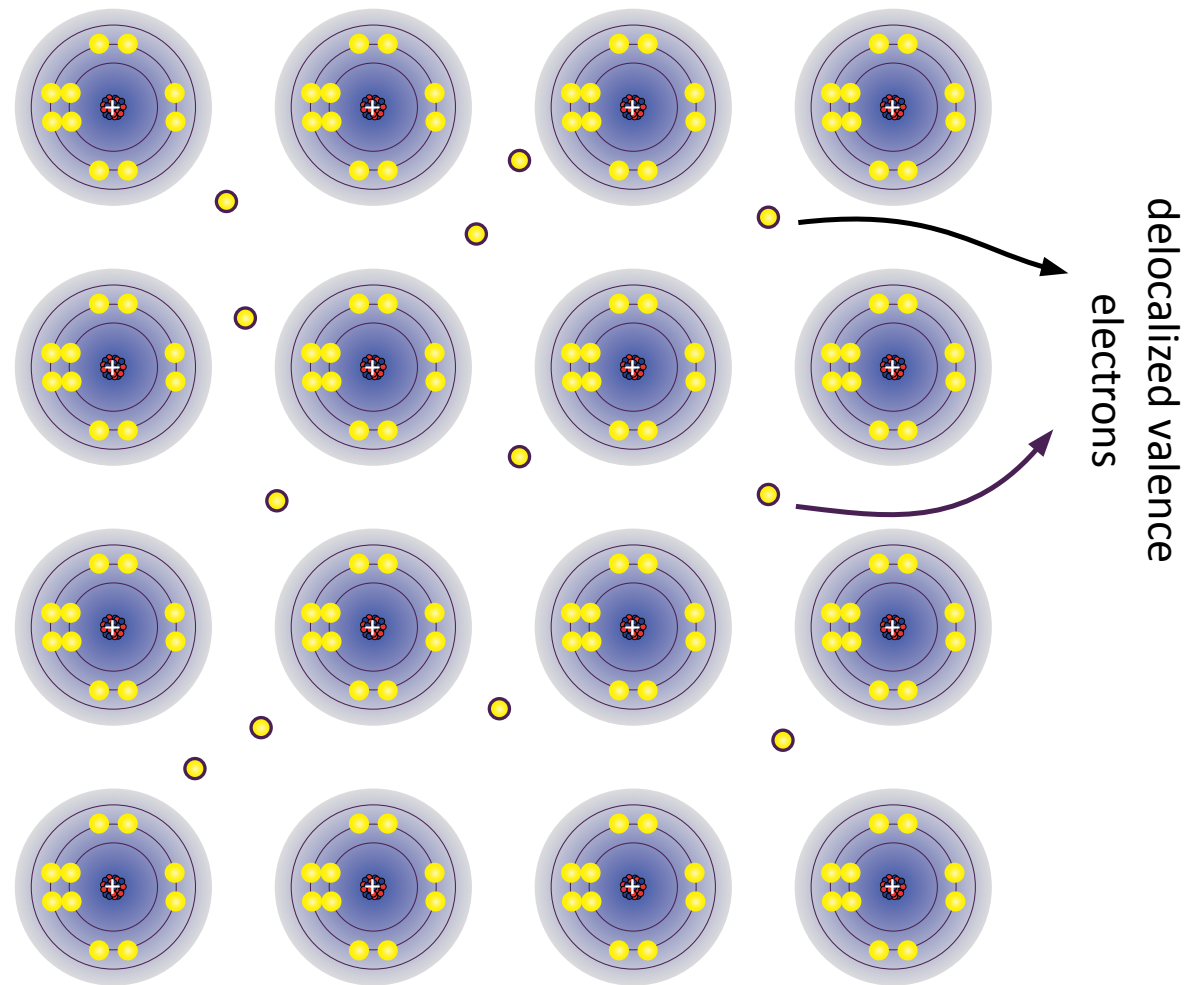
Metals conduct electricity and heat, they can be shaped into different forms and drawn out into thin wires, they reflect light (metals are shiny substances with a typical „metallic“ lustre). These famous physical properties, important both for the industry and the everyday life, are due to the specific structure of these simple substances.

It is obvious that metallic atoms are held together by some strong force. In covalent bonding atoms link to each other thanks to one, two or three bonding pairs, shared between the two atoms. Ionic crystals are held together by strong electrostatic forces acting between oppositely charged ions. But there is not just one ionic bond (for example between one ion Na^+ and one ion Cl^-). In the ionic lattice each positively charged ion is attracted simultaneously by all neighbour negatively charged ions and each negatively charged ion is attracted by all neighbour positively charged ions. That is why ionic crystals are stable structures. In a way it is similar for metals – identical metallic atoms form a **metallic lattice**.

In metallic lattices atoms are strictly arranged thus forming a 3D structure – identical unit cells built up of identical metallic atoms are arranged side by side, one above another and one below another. For example the metallic lattice of sodium Na is built up of many identical cubes with one sodium atom at each of their vertices and one more atom in the centre of the cube. Calcium Ca, magnesium Mg and aluminum Al are more tightly packed metals.



But why do these identical atoms form a stable structure? The modern science states that in metallic lattices all valence electrons are detached from the atoms. They are mobile and move freely in the space between the atoms, i.e. they are **delocalized**. Strong electrostatic forces acting between the positive nuclei of the atoms and these „free“ electrons hold the metal crystal together. So in metal lattices there are not many metal – metal bonds. All atoms building the lattice share their valence electrons thus forming the so called „sea of electrons“*. But it is not appropriate to consider metals as built up of ions as the electrons stay somewhere between atoms. They are not detached from one atom to be attached to another atom.



„SEA OF ELECTRONS“ IN THE METALLIC LATTICE OF SODIUM Na.

* In Bulgarian language the assemblage of delocalized electrons is called „electron gas“.

The famous physical properties of metals – their electrical and thermal conductivity, their plasticity (metals are malleable and ductile) and also their metallic lustre, are due to the mobility of the delocalized valence electrons.

Electrical current is a flow of electric charges. In a metal wire these are the freely moving in the metal lattice valence electrons. When a metal is heated the energy of these delocalized electrons increases thus making metals conductors of heat.



GOLD FOIL

Metals **malleability** is their ability to be shaped by hammering or pressing in useful forms even into thin sheets (foils). **Ductility** is their ability to be drawn out into thin wires. Metals can be deformed without breaking down because under pressure or when stretched metal atoms shift relative to each other

but remain linked one to another as the „bonding“ valence electrons move freely inside the lattice. The most malleable and ductile metal is gold. From 1 g of gold a wire of 2 km of length can be drawn.



COPPER WIRES

Metals are shiny as the delocalized electrons easily absorb and re-emit visible light. This makes possible the fabrication of reflecting mirrors and coatings.

The low volatility and the high melting points of metals are also due to the metal bonding as metal atoms are strongly held together.

METALLIC BONDING KEY POINTS:

- ▶ Metallic bond is formed between many identical atoms forming metallic lattices.
- ▶ All metal atoms in a metal lattice share their valence electrons. These electrons are free to move inside the lattice, i.e. they are delocalized. Metal atoms are held together by strong electrostatic forces of attraction acting between the positively charged atomic nuclei and the negatively charged delocalized electrons.
- ▶ Metals are shiny crystalline substances.
- ▶ Metals conduct electricity and heat.
- ▶ Metals are low volatile solid crystalline substances with high melting points.

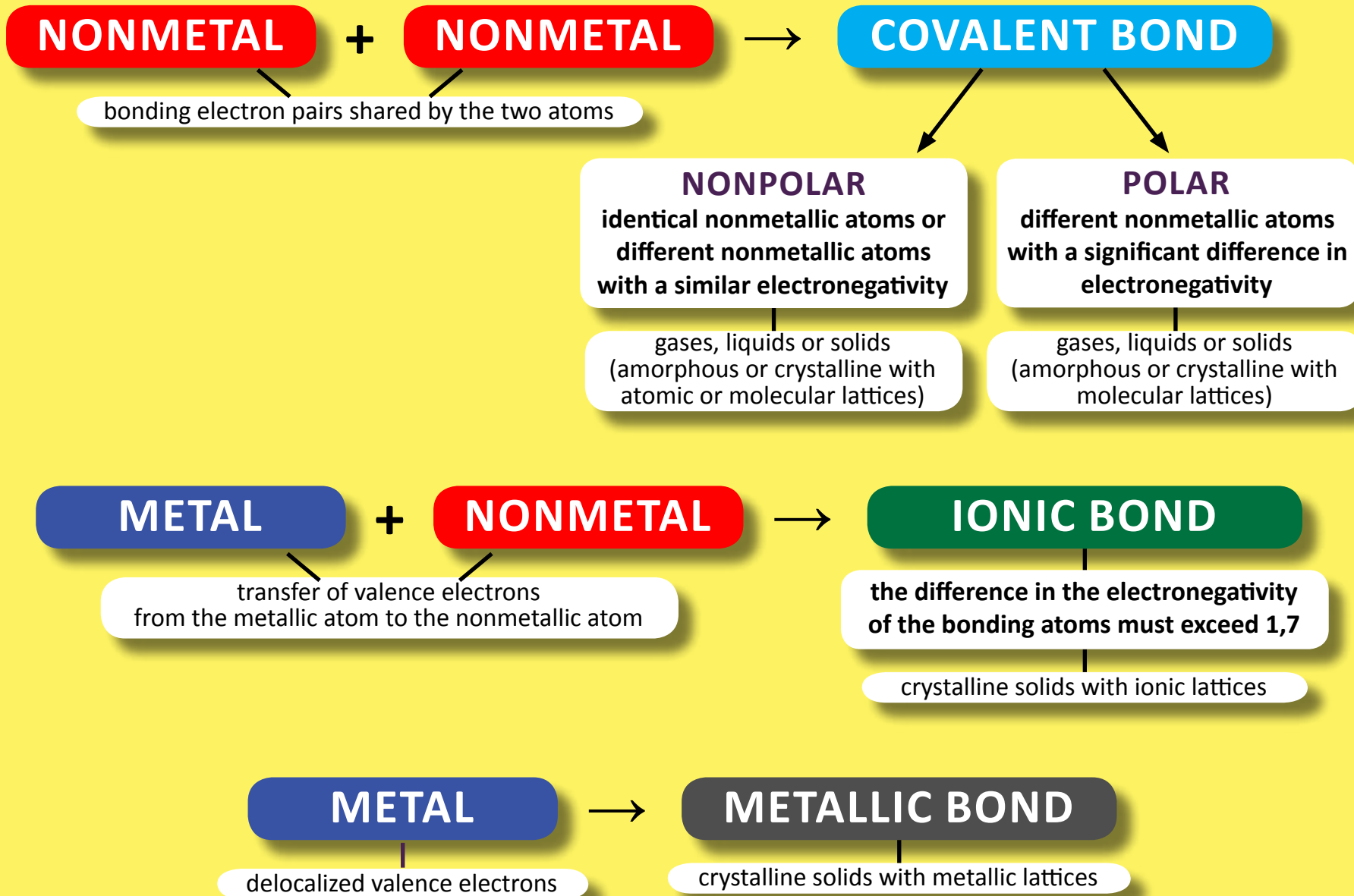
CAN YOU ANSWER THESE QUESTIONS ?

1. Why are metals crystalline substances?
2. How do metal atoms bond to form a metallic crystal lattice?
3. Why do we consider that the valence electrons in a metal lattice are delocalized?
4. What is a „sea of electrons“?
5. What properties of metals make possible the production of nails?
6. Why are not ionic substances malleable and ductile?
7. Sodium ($Z=11$) has only one electron more than neon Ne ($Z=10$). But neon is a gas while sodium is a solid substance with a high melting point. Why?
8. Can you give examples of metallic objects of different shapes used in the everyday life?

CHEMICAL BONDING SUMMARY:

- ▶ **Atoms (same or different) can bond to each other to form simple substances or chemical compounds.**
- ▶ **The type of bonding depends on the nature of the bonding atoms.**
- ▶ **Usually the atoms of two (same or different) nonmetals are covalently bonded. Bonds are polar or nonpolar depending on the electronegativity of the bonding atoms and the molecules shape.**
- ▶ **The structure of the covalent simple substances is atomic or molecular. If solid they are amorphous or crystalline with atomic or molecular lattices.**
- ▶ **The structure of the covalent compounds is molecular. If solid they are amorphous or crystalline with molecular lattices.**
- ▶ **If the difference in the electronegativity of the two bonded atoms is greater than 1,7 the bond is ionic (usually when a metal atom bonds to a nonmetal atom). Ionic compounds are solid substances with ionic crystal lattices. Ionic compounds are brittle and are insulators.**
- ▶ **Metal atoms link one to another by metallic bond and form metallic crystal lattices. Metals are plastic (malleable and ductile). They are shiny substances with metallic lustre and conduct heat and electricity.**

TO REMEMBER:



PROJECTS

1. ATOMIC STRUCTURE – HISTORY

Read and found out more about contributions made to the atomic theory by *Dalton, Thomson, Milican, Rutherford, Chadwick, Becquerel*.

2. USES OF ISOTOPES

Search for more information about some isotopes, used in science, technologies and medicine.

3. CHEMICAL BONDING AND PROPERTIES OF SUBSTANCES

Explain why and how the type of bonding determines the properties of substances. Discuss the properties of some well known substances, for example nitrogen, fluorine, helium, sodium, water and ice, sodium base, hydrogen chloride and others.

4. IS SiO_2 A MOLECULAR CRYSTAL?

Learn about the structure of the solid silicon dioxide SiO_2 . Explain how silicon and oxygen atoms bond to form a giant network structure. Find out information about the uses of this compound.