HW 9 – November 22 - review

1. The order of filling the energy sublevels by electrons is:

   1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, 5f, 6d, ...

   The maximum number of electrons on each s-orbital is 2, on p-orbitals is 6, on d-orbitals is 10.

   1.1. There are two gases with the charge of the nuclei 17 and 18. One is a poison, the other is not; it does not smell, does not burn. Write down electron configurations of both elements. Which one is poisonous?

2. The difference in electronegativities of atoms defines the nature of the bond between them. For mostly covalent bond the difference is < 0.4, for polar it is between 0.4 and 2, and for ionic bonds it is >2. The table below gives electronegativities of different atoms. E.g. the bond in O=O molecule is covalent: 3.44-3.44 = 0, the bond in H-F molecule is polar covalent: 3.98-2.2 = 1.78, and the bond K-O in K_2O is ionic: 3.44-0.82=2.62

Electronegativity:

Element	Electronegativity	Element	Electronegativity
Cs	0.79	H	2.20
K	0.82	C	2.55
Na	0.93	S	2.58
Li	0.98	I	2.66
Ca	1.00	Br	2.96
Mg	1.31	N	3.04
Be	1.57	Cl	3.16
Si	1.90	O	3.44
B	2.04	F	3.98
P	2.19		

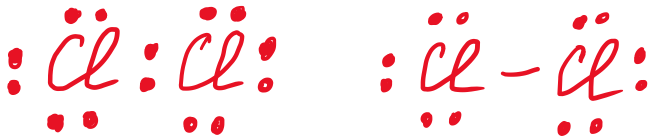
   2.1. Quartz and sand are silicon oxide SiO_2. What is the nature of the bond in this molecule – ionic or polar covalent? Use the table of electronegativities to answer the question.

   2.2. Which of the following two molecules have covalent bond and which one polar covalent: F-F, Cl-F?

3. **Lewis structures**, also known as Lewis dot diagrams, Lewis dot formulas, Lewis dot structures, electron dot structures, or Lewis electron dot structures (LEDS), are diagrams that show the bonding between atoms of a molecule and the lone (non-shared) pairs of electrons that may exist in the molecule.

   Lewis structures show each atom and its position in the structure of the molecule using its chemical symbol. Lines are drawn between atoms that are bonded to one another (pairs of dots can be used instead of lines). Excess electrons that form lone pairs are represented as pairs of dots, and are placed next to the atoms.

For example, the Lewis structures for chlorine molecule is:



Structural formulas will show locations of chemical bonds between the atoms of a molecule. They consist of symbols for the atoms connected by short line that represent chemical bonds (each short line represents 2 electrons):

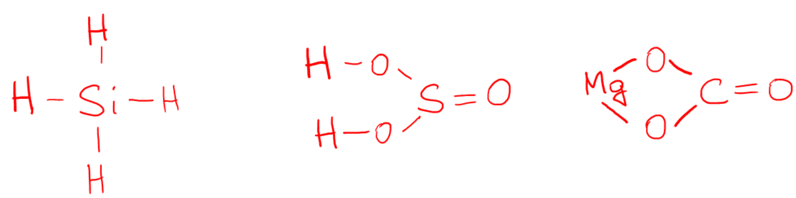
Cl-Cl, H-H, Cl-Mg-Cl

When detailed information is not needed chemical formulas are used. In chemical formulas the number on the right of the element symbol shows how many atoms of this element is present in a molecule.

E.g. for Cl-Cl the chemical formula is Cl_2, for Cl-Mg-Cl it is $MgCl_2$

3.1. Write down Lewis structures for molecules $BaCl_2$ and MgO.

3.2. Write down chemical formulas of the following compounds:



4. The **oxidation state**, which may be positive, negative or zero, is the hypothetical charge that an atom would have if all bonds to atoms of different elements were 100% ionic, with no covalent component. To determine the oxidation state of atoms in a molecule we assign the negative charge to the more electronegative atom and positive charge to the less electronegative (more electropositive) atom. The total charge in a molecule is 0. The total charge in an ion is equal to the charge of the ion. The oxidation state of H is almost always (+1, except in metal hydrides like NaH, where it is (-1)), the oxidation state of F is (-1), that of O is almost always (-2). E.g. in carbon dioxide CO_2 the oxidation states of C is (+4) since there are two oxygens with oxidation state of (-2) each and the molecule is not charged. We write: $C^{+4}O^{-2}{}_2$

   4.1. What is the oxidation state of each atom in the following molecules: Fe_2O_3, FeO, N_2, CaO, HNO_3?

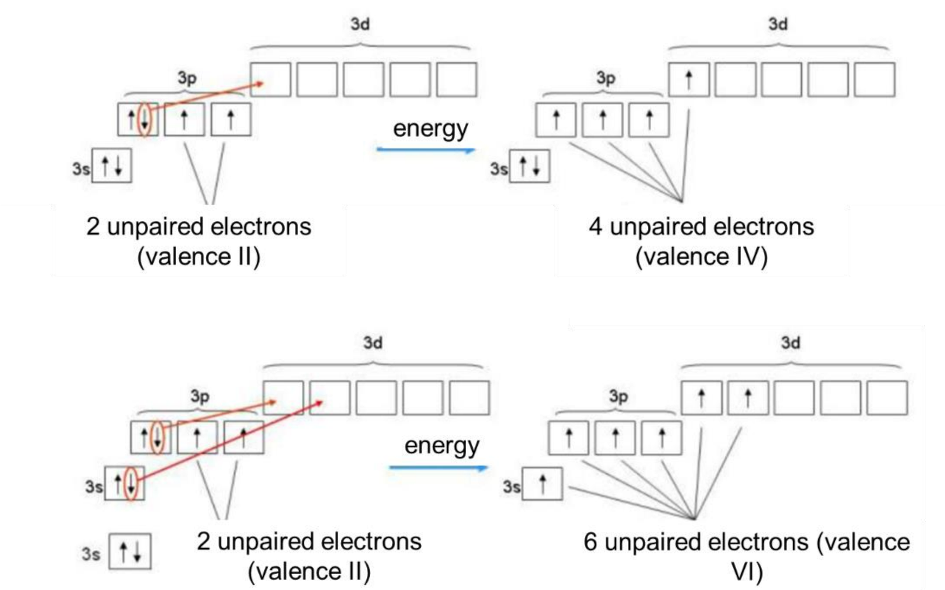
5. <u>**The valence**</u> is the number of electron pairs that binds the atom with other atoms. We learned how to determine the valence using “octet rule”. For some common elements it may be useful to remember their valences. The table below gives valences of some common elements. (The numbers in parentheses show possible valences for elements that may exhibit more than a single valence.)

The valence can be used to write down chemical formulas. E.g. if we want to write down the formula of Fe (III) compound with oxygen (iron oxide) we can write down the elements symbol with their valences on top and then move the valences to the opposite elements as their indexes:

III II

$Fe_2\ O_3$

If given an additional energy an atom can get into an excited state from the ground state where the energy is at its minimum. In the excited state electrons can unpair and move to different orbitals within the same shell increasing the valency of the atom. For example, sulfur ($_{16}S$) can have valences II, IV, and VI by transferring one or two electrons to 3d orbitals:



The energy necessary to unpair electrons and increase the valence may be compensated by formation of more molecular bonds with the excited atom.

Valences of some common elements

Element	Valence	Element	Valence
H	I	Ba	II
Na	I	O	II
K	I	Zn	II
Ag	I	Sn	II (IV)
F	I	Pb	II (IV)
Cl	I (III, V, VII)	Fe	II, III

Br	I (III, V, VII)	Cr	III, VI
I	I (III, V, VII)	S	II, IV, VI
Hg	I, II	Al	III
Cu	I, II	N	III (IV)
Be	II	P	III, V
Mg	II	C	IV
Ca	II	Si	IV (II)

5.1. Using valences of elements write down formulas of a) calcium with fluorine, b) magnesium with oxygen, c) aluminum with oxygen.

5.2. Draw structural formulas and determine the valence of each atom in a) HCl, b) $BeCl_2$, c) $AlBr_3$, d) PH_3, e) $TiCl_4$.

5.3. * Predict constant or variable valence of $_{34}Se$ using its valence (outer) shell configuration.