

# CLASSIFICATION OF ELEMENTS. PERIODICITY.

March 30, 2020

*This course is an expanded text used for teaching Chemistry semilabs in SigmaCamp.*

As we already know, all elements can be subdivided onto metals and nonmetals. Such a classification is intuitively clear: everyone understands what a metal is. Indeed, all metals (except mercury) are:

1. solids;
2. non-transparent;
3. lustrous. More precisely, they have metallic luster (“metallic luster” is a specific term; it is known in physics that reflection of light from metals is different from reflection of light from other lustrous surfaces<sup>1</sup>;
4. conductive. They have high electrical and thermal conductivity, and these two traits are not separable from each other;
5. usually, although not always, metals are malleable (i.e. they can be shaped by pounding with a hammer)<sup>2</sup>.

These properties forms an unseparable set. If some element possess *all* these properties, it is a metal. Accordingly, the elements that are devoid of at least one of these properties are not considered metals. However, this is a “physicist’s point of view” on metals. Is this classification arbitrary, or it has some chemical applications? To answer that, let’s look at metals and nonmetals from chemist’s point of view.

## 1 Metals. Chemist’s point of view.

Majority of chemical elements are metals. They all have some basic common physical properties: metallic luster, high thermal and electrical conductivity, non-transparency, and malleability. Such a similarity is not just a coincidence. There is an important fundamental reason for that, and this reason lies is the structure of the atoms of metals. One important

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<sup>1</sup>For example, a light reflected from water can be absorbed by so called polarisation light filters, a trick that is being widely used in photography. In contrast, a light reflected from a metal surface cannot be absorbed by these filters.

<sup>2</sup>Sometimes, metals can be fragile at low temperatures.

consequence of that is the following: **metals do not like to make chemical compounds with each other**. Of course, you can take, for example, 23 grams of sodium (monovalent; atomic weight 23 Da) and 20 grams of calcium (divalent; atomic weight 40 Da), and to melt them together. However, the resulting substance will be not a new compound with a formula  $\text{Na}_2\text{Ca}$ , but just an *alloy*, i.e. a mixture of two substances, and its properties would be similar to the properties of the starting metals. Definitely, if all elements in our Universe were metals, it would be a very boring and tiresome world.

What brings diversity to our world is the interaction of metals with *nonmetals*. Firstly, all metals are capable of forming binary compounds with almost every nonmetal. Thus, reaction of magnesium with oxygen, sulfur, or chlorine produce magnesium oxide, sulfide, and chloride, accordingly.



Interestingly, these compounds are either salts ( $\text{MgS}$  or  $\text{MgCl}_2$ ), or they can be converted into salts in a reaction with acids ( $\text{MgO}$ ). That is a fundamental property of metals that allows us to discriminate between metals and nonmetals.

**Metals are the elements that can produce salts in a reaction their oxides and an acid.**

However, as we already know, that does not mean all metal oxides are basic. They (for example, zinc or iron (III) oxides) can equally be amphoteric. What is important for us is the ability of such an oxide to react with acids.

In addition, one has to keep in mind the following. As we know, some metals exist in different valence state. For example, two oxides are possible for copper ( $\text{Cu}_2\text{O}$  and  $\text{CuO}$ ). As a rule, *lower valence oxides are more basic than the higher valence oxides*. For example, iron (II) oxide ( $\text{FeO}$ ) is basic, whereas iron (III) oxide ( $\text{Fe}_2\text{O}_3$ ) is amphoteric. Therefore, the above rule can be re-phrased as follows.

**If an element X forms at least one basic or amphoteric oxide, the element X is a metal, otherwise it is a nonmetal.**

## 2 Nonmetals.

In contrast to metals, the nonmetals alone are capable of forming a wide range of chemical compounds. They love to react with each other, and with metals. Almost every nonmetal is capable of forming binary compounds with other nonmetals, and almost every nonmetal (except fluorine, helium, neon and argon) can form oxides.

**Oxides formed by nonmetals are acidic.**

Does that mean *all* oxides of nonmetals are capable of generating acids in a reaction with water? No. One of the most important examples is *carbon monoxide* (CO), which does not react with water, and cannot be converted in a salt in a reaction with a base. Another example is a nitrogen (I) oxide (N<sub>2</sub>O), a gas that is being widely used as a propellant and for anesthesia. However, both carbon and nitrogen are nonmetals, because they form no basic or amphoteric oxides. That can be summarized as follows:

**Nonmetals are the elements that form neither basic nor amphoteric oxides.**

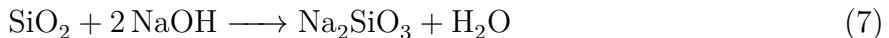
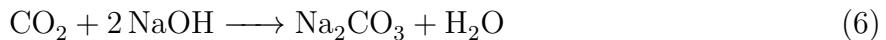
### 3 “Metalloids”: what is that?

The term “metalloid” has a bad karma. On the eve of inorganic chemistry, there was no term “nonmetal”, and the term “metalloid” was used instead. In other words, all elements were subdivided onto metals and *metalloids*. That was counter-intuitive and misleading. Indeed, the suffix “-oid” means “similar to”, or “resembling”, so the word “metalloid” literally means “resembling a metal”. Of course, that was not true: for example, can we seriously speak about any similarity between oxygen and magnesium? That is why this term was gradually dropped, and the term “nonmetal” was proposed instead.

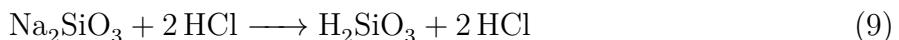
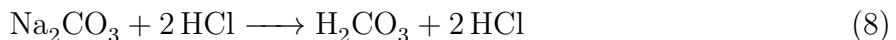
Nevertheless, some new textbooks started to use the term “metalloid”, although its modern meaning is different. It is used for some nonmetals that possess one or several traits of metals. For example, such elements as silicon or germanium are hard solids with a high melting point; they have a metal luster and moderate electrical conductivity, which is much lower than a conductivity of true metals. However, despite their visible similarity with metals, from chemist’s point of view they are nonmetals. Thus, silicon or germanium compounds are very similar to carbon compounds. A comparison of silicon and carbon shows that both elements form oxides:



These oxides are acidic:



The acids can be generated from their salts:



In other words, such "metalloids" as germanium (Ge) or antimony (Sb) are, from the chemist’s point of view, actually non-metals.

## 4 Groups of elements.

As we can see from the above equations, there is a striking similarity between chemical reactions of carbon and silicon compounds, which implies these two elements share some important common features. Such a parallelism is not merely a coincidence: all chemical elements can be split onto groups based on their chemical properties. That means, the two major classes of elements, metals and nonmetals, contain smaller sub-classes, and each element in a certain sub-class (or a group) have something in common with other members of that group. That observation was made in XIX century, and it made the life of chemists much easier: they realized that chemistry is not a random collection of facts about various reactions between various elements and their compounds, but some systematic discipline, which can be understood by humans. Starting from the next lesson, we will learn about the essential groups of elements, including halogens, chalcogens, and alkaline metals.

## 5 Hydrides

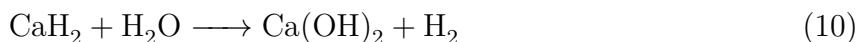
Our conclusion about similarities of carbon and silicon was based primarily upon a commonality in behaviour of their oxides and other oxygen containing compounds. Meanwhile, another type of compounds that may be useful for classification is so called *hydrides*. We already know an example of one hydride: a *oxygen hydride* (a. k. a. "hydrogen oxide", a. k. a. water). Indeed, if we define the term "hydride" as "a binary compound formed by some element and hydrogen", then water can be considered also an oxygen hydride. Can we really do that?

Sure, we can. Depending on a situation, we can call water "oxygen hydride", and, as we will see below, that may, sometimes, be useful.

Other examples of hydrides are methane ( $\text{CH}_4$ , "carbon hydride", a natural gas), ammonia ( $\text{NH}_3$ ), hydrochloric acid (an aqueous solution of  $\text{HCl}$ , or hydrogen chloride, which can also be considered as a chlorine hydride). Why we are speaking about all of that here? Because that will be useful for creation of classification of elements, as we will see below.

### 5.1 Hydrides of metals

All hydrides discussed above are formed by nonmetals. However, like nonmetals, most metals are capable of forming hydrides, and these hydrides have interesting properties. Let's take calcium hydride as an example. Calcium is always divalent, and hydrogen is monovalent, so the formula of calcium hydride should be  $\text{CaH}_2$ . This compound is a hard solid, thereby resembling calcium chloride or fluoride. However, its chemical properties are different from those of calcium halogenides or other salts. If we drop a piece of calcium hydride into water, a violent reaction begins, some gas evolves, and the liquid becomes turbid. The gas that forms in this reaction is hydrogen, and another product of is calcium hydroxide  $\text{Ca}(\text{OH})_2$  (if we add few drops of phenolphthaleine to the solution, it turns pink).



The same reaction can be observed if we take sodium hydride or other metal hydrides. In general, most metals form hydrides, and these hydrides react with water yielding a metal hydroxide and a hydrogen gas<sup>3</sup>.

With this knowledge in mind, let's try to identify some groups of elements based on the commonality of properties of their oxides and hydrides.

## 6 Groups of metals

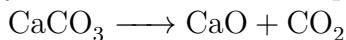
### 6.1 Alkaline metals

We already know that group metals are the most active and the most typical metals. In addition to sodium and potassium, the metals we are already familiar with, three more metals, **Lithium (Li)**, **Rubidium (Rb)**, and **Cesium, or caesium (Cs)** belong to this group<sup>4</sup>. These metals form oxides that react with water violently to form hydroxides, which are strong alkali (hence the name *alkaline metals*: “Al kali” is the Arabian word that means “a caustic substance”, “Al” is an article.) Another common feature of alkaline metals is high solubility of virtually all their salts<sup>5</sup>.

The important feature of alkaline metals is that **they are always monovalent**, so the common formula of their hydrides and oxides are MH and M<sub>2</sub>O, accordingly ("M" denotes a alkaline metal).

### 6.2 Alkaline earth metals

The typical representative of this group is calcium. Other members are magnesium, **Barium (Ba)**, **Strontium (Sr)**, and **Beryllium (Be)**. These metals also form highly alkaline hydroxides, but they are poorly soluble, hence the name “alkaline earth”. In Medieval times the term “earth” was used to denote any insoluble non-metallic solid obtained after heating some material for a long time; when alchemists heated calcium carbonate or magnesium carbonate in an oven they observed their conversion into some other solid. The reaction they observed was a decomposition of carbonates:



The new solid was also an “earth” (poorly soluble solid), but we know it is a basic oxide (or an “alkaline earth”, according to alchemical terminology). When other elements from this group (barium and strontium and berillium) were discovered their oxides also demonstrated the same behaviour, i.e. they were the “alkaline earthes”, hence the common name.

Alkaline earth metals **are always divalent** in their compounds, so the common formula of their hydrides and oxides are MH<sub>2</sub> and MO, accordingly ("M" denotes a alkaline earth metal).

<sup>3</sup>The situation is more complex for heavy metals, however, we do not consider these exceptions now.

<sup>4</sup>We are limit ourselves with non-radioactive elements only.

<sup>5</sup>Few lithium salts have low solubility. Other salts of lithium and all salts of other alkaline metals are soluble in water

## 7 A general approach towards the description of classes of elements

If we continue grouping of other elements, we can expect the following general formulas for other groups:

- a group 3 elements form hydrides  $XH_3$  and oxides  $X_2O_3$ ,
- a group 4 elements form hydrides  $XH_4$  and oxides  $XO_2$ ,
- a group 5 elements form hydrides  $XH_5$  and oxides  $X_2O_5$ , etc.

Unfortunately, a situation is not that simple. Whereas that logic works fine for oxides, the hydrides behave differently starting from the group 5. The actual scheme is as follows:

- a group 3 elements form hydrides  $XH_3$  and oxides  $X_2O_3$ ,
- a group 4 elements form hydrides  $XH_4$  and oxides  $XO_2$ ,
- a group 5 elements form hydrides  $XH_3$  and oxides  $X_2O_5$ ,
- a group 6 elements form hydrides  $XH_2$  and oxides  $XO_3$ ,
- a group 7 elements form hydrides  $XH$  and oxides  $X_2O_7$ ,
- a group 8 elements form **no** hydrides, and the general formula of their oxides is  $XO_4$ .

Let's identify these groups using the above scheme.

## 8 “Halos” means “salt”

We already know that addition of the sodium chloride solution to the solution of silver nitrate yields a precipitate of a salt called silver chloride:



Is such a reaction specific to the table salt ( $NaCl$ ), or some other compounds behave similarly? Let's look at the compounds formed by the elements we never discussed before, sodium bromide and sodium iodide.

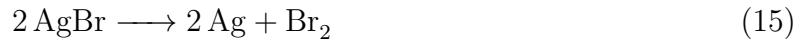
### Experiment 22.

*In three glass beakers, put approximately 1 gram of sodium chloride, sodium bromide, and sodium iodide. Add 50 mL of water to each beaker and stir. What do you see? Pour 2 mL of the each liquids into separate test tubes and add 1 mL of silver nitrate solutions to each of them. Describe what you see. Leave all three tubes under bright lamp or sunlight for a short period of time. Do you see any change?*

As we expected, sodium chloride produced a voluminous white  $AgCl$  precipitate, which is insoluble in water. Sodium bromide and iodide behave similarly.



When exposed to sunlight, all three salts behave similarly: all three white precipitates turn gray and then black. In all three cases, the same product, a fine silver powder is formed, which, as well as all finely divided metals, is black.<sup>6</sup> The reactions are as follows:



These three reactions are dramatically facilitated by light. It is the reaction 14 that allowed us to see the faces of our grand-grand parents when they were young, to see historical XX century events, to watch old movies, because that reaction is a core chemical reaction of the wet photography. For us, these three reactions are a good demonstration of the commonality between all three compounds.

However, what about other metals? Do bromides and iodides of other metals behave similarly to their chlorides too? Actually, yes. For example, we know calcium chloride is soluble in water, and its bromide and iodide are easily soluble too. The same can be said about magnesium, iron, copper, zinc chlorides, bromides and iodides. The striking similarity between those compounds allows us to formulate the following rule:

**If a chloride of some metal is easily soluble in water, its bromide and iodide are soluble too. The exceptions to this rule are extremely rare.**

Due to the ability of the three elements, chlorine, bromine, and iodine, to form salts with all metal, these three elements were combined into a separate group, which was named “halogens”, from two Greek words, “halos” (“salt”) and “gignomay” (“come to be”), literally “salt begetters”. Accordingly, the word “halogenides” is used as an umbrella term of all salts formed by halogens. For example, instead of writing: “potassium chloride, bromide, and iodide are white solids”, we can write “potassium halogenides are white solids”.

## 9 Hydrogen chloride and hydrochloric acid, or “acidum salis”.

What about the binary compounds formed by halogens and nonmetals? As we know, hydrogen chloride is a compound whose aqueous solution are strongly acidic (accordingly, its name is “hydrochloric acid”). Since hydrogen chloride is a gas, it can be prepared from sodium chloride and sulfuric acid: as we know, exchange reactions between a salt and an acid can lead to formation of a new salt and a new acid when the latter is being removed from the reaction mixture, by formation of either a precipitate (as the silicic acid,  $\text{H}_2\text{SiO}_3$ ), or a gas (like  $\text{HCl}$ ). The equation of this reaction is as follows:

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<sup>6</sup>Actually, finely divided metals are among the most black materials in the world.



This reaction was the first practical method for the preparation of the hydrochloric acid: gaseous HCl that forms in this reaction can be collected and bubbled through water. It dissolves in water vigorously to form *acidum salis*, or hydrochloric acid.

Like HCl, hydrogen bromide and chloride are gases that are highly soluble in water. Their solutions are called hydrobromic and hydroiodic acids, accordingly. These two acids are very strong and corrosive, even stronger than HCl is. Hydroiodic acid is among the strongest acids currently known.

## 10 Fluorine, the fourth halogen.

Besides the three halogens (Cl, Br, and I), one more element was discovered whose properties are similar to halogens. Its name is fluorine. Like other halogens, it forms salts with all metals, although their properties are somewhat different. Like HCl, hydrogen fluoride (HF) is an acid, although it is not as strong as other hydrogen halogenides are. By its properties, fluorine stays a little bit apart from chlorine, bromine and iodine, however, there are several important reasons to consider it a halogen. We will discuss them a little bit later.

## 11 Preparation of elementary halogens.

Pure halogens can be prepared from their binary compounds by taking away the second element. For example, chlorine can be prepared from hydrogen chloride and potassium permanganate:



In this reaction, potassium permanganate, which serves as a “depot form” of oxygen (as we can see, it is an oxygen rich molecule), donates oxygen to HCl. Since oxygen binds to hydrogen very actively, it takes it from chlorine, and free chlorine atoms form. These chlorine atoms bind together, and a diatomic chlorine molecule forms. The same reaction can be used to produce bromine and iodine, but not fluorine. Why? As we can see from the above reaction, to prepare a halogen in its free form, we need some chemical that binds hydrogen more avidly than the halogens themselves do. It is possible to find such a chemical for HCl. Hydrogen bromide and hydrogen iodide (in that order) give their hydrogen even easier. However, no chemicals were found that take hydrogen from HF to give F<sub>2</sub>. The reason is simple: fluorine is the most active element in the Universe. It reacts with other elements, both metals and nonmetals so vigorously, that such bonds are almost impossible to break. As a result, fluorine containing compounds prefer to participate in exchange reactions, but not in dissociation reactions.

**Fluorine is the most chemically reactive element in the Universe. It reacts vigorously with most known materials. Even oxygen burns in fluorine.**

In connection to that, every person who is capable of critical thinking must ask: “If fluorine binds so strongly to other elements than how did people manage to prepare it in a pure form?” This question deserves a separate chapter. We will discuss it during the next class.

Besides usage of oxygen rich compounds as  $\text{KMnO}_4$ ,  $\text{KClO}_3$  and others, another general way exists to generate halogens from their salts. This procedure is called **electrolysis** (literally “decomposition by electricity”).

## 11.1 Chlorine

### Experiment

*Put approximately 2-5-grams of table salt into a 250 mL glass beaker, and add 200 mL of water there. Stir the mixture until all salt dissolved. Attach insulated wires to a standard 9 V battery and mark positive and negative wires. Strip the opposite ends of each wire (about 3-5 cm) from the insulation. Put both ends of the wires into the  $\text{NaCl}$  solution and describe your observations.*

Before we continue, let's introduce three more terms. In chemistry and physics, positive wire end is called **anode**, and negative is called **cathode**. Remember that, because we will use these terms many times.

**An electric wire used to make contact with some nonmetallic media, be it a liquid, a solid, a human body, or a vacuum, is called “electrode”.**  
**A positive electrode is called “anode”, a negative electrode is called “cathode”.**

As we can see, bubbles of some gas are formed both at the anode and the cathode. Is it the same gas, or they are different? If we collect these gases, we can easily see that the gas formed at the cathode has no odor, no color, and it is combustible. It is easy to see that gas is hydrogen. It is easy to see that the gas formed at the anode is quite different, simply because of its unpleasant smell resembling the smell of bleach. That gas is an elementary chlorine. Moreover, it is easy to see the amount of chlorine formed on the anode is much smaller than the amount of hydrogen on the cathode. If you look carefully at the copper anode, you can see its color changing. Obviously, some reaction takes place at the anode that leads to its corrosion. To eliminate corrosion, let's replace the copper anode with a small piece of platinum wire. As we expected, platinum, which is among the most inert metals, does not corrode (its surface remains clear and shiny), and the amount of chlorine bubbles increases dramatically.

What is going on during this process? Obviously, when an electric current passes through sodium chloride, it decomposes onto chlorine, which we can see (the bubbles) and smell (a bleach smell). But what is the second product? Isn't it logical to conclude that the second product is a sodium metal? But, if that is the case, why we don't see it? To answer this question, let's repeat the experiment, but let's add few drops of phenolphthalein into the  $\text{NaCl}$  solution. We will see that immediately after connecting the electrodes to the battery a pink coloring starts to form near the cathode. That means the solution becomes basic. Now we can reconstruct all the processes in our apparatus.

Firstly, when an electric current passes through the NaCl solution, sodium chloride, initially, decomposes onto the elements:



An interesting feature of this reaction is that the products are separated in space: chlorine is formed at the anode, whereas sodium is formed at the cathode. Both substances are extremely reactive, so a major part of chlorine immediately reacts with the copper electrode to form copper chloride or other compounds:



If we need to obtain pure chlorine, the anode's copper must be replaced with some more inert material, for example, platinum. That allows us to see a considerable amount of chlorine bubbles, and to feel its smell. But what is going on at the cathode? why do we see hydrogen bubbles instead of sodium? The reason is obvious. As we already know, sodium metal is so reactive that it violently reacts with water to form sodium hydroxide and hydrogen. That is exactly what happens at the cathode: immediately after its formation every single sodium metal reacts with water to produce hydrogen (a gas we see) and NaOH (a purple phenolphthalein color). Therefore, the net result the **electrolysis** reaction is the formation of a chlorine gas at the anode, and sodium hydroxide and hydrogen at the cathode:



This process is general. All salts can be decomposed in such a way. In all cases, metals (or the products of their reaction with water) are formed at the cathode, and acidic residues are transformed at the anode. We cannot afford a luxury to discuss this process in details now, so let's return to its discussion in due time.

**Decomposition of salts by means of the electric current is called electrolysis.**

*Interestingly, liquid NaCl can be subjected to electrolysis too. Sodium chloride melts at 801 °C, and, when the electric current passes through the molten NaCl, sodium metal formed at the cathode has nothing to react with. As a result, we get pure chlorine gas at the anode, and pure sodium metal at the cathode. That is an industrial way to prepare large amount of these two elementary substances. Sir Humphry Davy, a brilliant British chemist, was a first person who isolated some alkaline metals in using electrolysis.*

## 11.2 Bromine and iodine

Both bromine and iodine can be prepared similarly. Since iodine is a solid, it forms a black film on the anode, which peels off easily.



Sir Humphry Davy (1778-1829),  
a person who first used electrolysis for isolation of some elements.



Henry Moissan (1852-1907),  
a discoverer of fluorine.

### 11.3 Fluorine

Since fluorine is much more reactive even than chlorine, it cannot be prepared by the electrolysis of the sodium fluoride solution, simply because it would immediately react with water to generate oxygen and hydrogen fluoride<sup>7</sup>



That means the result of such a reaction would be the formation of hydrogen at the cathode and *oxygen*, (not fluorine) at the anode. To obtain elementary fluorine, we must exclude water.

Fortunately, that is possible to do, because, in contrast to HCl or HBr, anhydrous HF is not a gas, but a liquid. Many salts, including potassium fluoride, are soluble in liquid HF, so its electrolysis can produce elementary fluorine. Unfortunately, that is not the only obstacle chemists had to overcome to prepare pure fluorine. The main difficulty was an extreme reactivity of fluorine: since it reacts with almost every material, its collection and storage is a formidable task. In addition, due to its high activity, fluorine is extremely dangerous. Several chemists, including Sir Humphry Davy, attempted to isolate it, and some of those attempts were almost fatal. It was probably a reason why Davy died so prematurely. A person who succeeded in isolation of pure fluorine was a French chemist Henri Moissan, who was awarded a Nobel Prize for that achievement in 1906.

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<sup>7</sup>Please, pay attention to that fact: fluorine is so reactive that it takes hydrogen even from oxygen. It is the only element capable of doing that.

## 12 Reaction of halogens with other elements

As we already know, halogens are very reactive substances. One of the most common halogen reaction is a formation of binary compounds. Thus, such metals as copper or iron burn in gaseous chlorine to produce chlorides:



(Obviously, such active metals as sodium react with chlorine even more vigorously, and the product of such a reaction is a metal chloride.)

Not only metals are capable of reacting with chlorine. Some moderately active nonmetals, such as phosphorus, ignite spontaneously in chlorine. The product of that reaction is phosphorus (III) chloride  $\text{PCl}_3$



Unfortunately, we cannot perform all these spectacular experiments in the class, because a special protective equipment is needed for that. However, you can watch some of those experiments on YouTube. In the Homework section, some links are provided.

## 13 Halogens and oxidation

As we can see, reaction of most metals and some nonmetals with halogens are vigorous, and they yield corresponding binary compounds (halogenides). In that sense, they have much in common with the reactions between oxygen and other elements (i.e. the combustion reaction, a.k.a. **oxidation reaction**). Therefore, the chemists realized the term “oxidation” is equally applicable to the reactions of halogens with metals and less active nonmetals.

**Reactions of halogens with metals and less active nonmetals are the oxidation reactions. Like oxygen, halogens serve as oxidizers in that type reactions. Fluorine is the strongest oxidizer among all elements.**

If we compare, for example, the reaction between magnesium and oxygen and the reaction between magnesium and chlorine, we will see that in both cases a zero valence magnesium becomes divalent.



That is a common trait of all oxidation reactions.

**A valence of the element that has been subjected to oxidation always increases. Conversely, when a valence of at least one atom has increased during some reaction, such a reaction is an oxidation reaction.**

In connection to that, it is interesting to look at the exchange reactions between acids, salts and bases we studied recently. As you have probably noticed, valences of all atoms in totally preserved for all reactants and products of those reactions. That observation deserves to be emphasized:

**Exchange reactions between acids, bases, or salts are not oxidation reactions.**

### 13.1 Chlorine or oxygen: what is stronger?

As we can see, chlorine can oxidize even the elements that seem not to react with oxygen. Does it mean it is a stronger oxidizer? Actually, no. There are two reasons why chlorine's apparent reactivity seems higher. Firstly, whereas chlorides of some elements are volatile or low melting solids, their oxides are hard and high melting compounds. For example, aluminium, which is a very active metal judging by its position in the reactivity series, does not ignite in air when heated. However, it ignites and violently burns in the chlorine atmosphere. The explanation is simple. Aluminium reacts with oxygen almost instantly, however, its product,  $\text{Al}_2\text{O}_3$  forms a very dense and hard film on the metal's surface, so the bulk aluminium becomes protected from further oxidation. In contrast, dry aluminium chloride is a low melting solid that provides no protection for the piece of aluminium metal. As a result, combustion of aluminium in chlorine lasts until all chlorine or all aluminium is consumed.

The second reason for high activity of halogens is in their monovalence. That means, in  $\text{F}_2$ ,  $\text{Cl}_2$ ,  $\text{Br}_2$  and  $\text{I}_2$ , halogen atoms are connected to each other just with a single bond, which is easy to break<sup>8</sup>. In contrast to halogens, oxygen is divalent, so the  $\text{O}_2$  molecule is much more stable. Therefore, much more energy is needed to break the oxygen molecule apart.

These two reasons make oxygen less active than chlorine or bromine are. However, oxygen is more active oxidizer. We can make that conclusion because, whereas it can form oxides with chlorine, bromine, or iodine (but not with fluorine), no chlorides, bromides or iodides of oxygen can be prepared.

## 14 High valence halogen oxides and oxo acids.

As we already know, overwhelming majority of elements form oxides. As a rule, their oxides have low stability, which is understandable, taking into account high halogen's activity: they do not "like" to be oxidized, when they are bound to oxygen, they donate an oxygen atom easily to other elements. In contrast to oxides, most oxygen containing acids generated from halogen oxides much more stable. A Table 1 summarize all acids formed by chlorine, bromine, and iodine.

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<sup>8</sup>Obviously, for oxidation to occur, oxidizer's molecule must fall apart first.

**Table 1. Halogen oxo acids**

Valence	Cl	Br	I	Generalized formula
I	HClO	HBrO	HIO	HXO
III	HClO <sub>2</sub>	HBrO <sub>2</sub>	-	HXO <sub>2</sub>
V	HClO <sub>3</sub>	HBrO <sub>3</sub>	HIO <sub>3</sub>	HXO <sub>3</sub>
VII	HClO <sub>4</sub>	HBrO <sub>4</sub>	HIO <sub>4</sub> (H <sub>5</sub> IO <sub>6</sub> )	HXO <sub>4</sub>

As we can see, in these acids, the valence of halogen atoms varies from 1 to 7, and it always takes only odd values. The maximal valence is 7 for all of them (fluorine is the sole exception).

## 15 General properties of halogens

Now we can summarize some common properties of halogens.

1. All halogens form binary compounds (“halides”, a.k.a “halogenides”) with metals. Halogen’s valence in halogenides is ***always equal to one***.

**Examples:** NaCl, AlBr<sub>3</sub>, CaF<sub>2</sub>.

2. Metal halides are salts. As a rule, if chloride of some metal is soluble, the same is true for its bromide or iodide. If one halogenide of some metal is insoluble, other halides are insoluble too.<sup>9</sup>

**Examples:** CuCl<sub>2</sub> is soluble in water; CuBr<sub>2</sub> and CuI<sub>2</sub> are soluble too. AgCl is insoluble in water; AgBr and AgI are also insoluble.

3. All halogens form binary compounds with hydrogen with a general formula HX, where X is halogen. All hydrogen halogenides are acids. ***Acidity of hydrogen halogenides increases with the increase of halogen’s atomic mass.***
4. All halogens form binary compounds with oxygen. Chlorine, bromine, and iodine (but not fluorine) form a series of oxygen containing acids. ***The maximal valence of halogen’s atom in those acids is 7; in other acids the valence assumes only odd values.***

## Homework

1. Read “Ask a Foolish Question” by Robert Sheckley. You can find it, for example, at the Project Gutenberg page <http://www.gutenberg.org/files/33854/33854-h/33854-h.htm><sup>10</sup>.

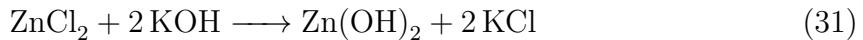
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<sup>9</sup>Sometimes, that doesn’t work for fluorine.

<sup>10</sup>This link, as well as all other urls are clickable.

2. Read the above Classwork materials. Answer the following questions:
1. Some solid ('A') has been prepared by heating an element X in an oxygen atmosphere. This solid is insoluble in water, however, it dissolves in aqueous hydrochloric acid or in aqueous sodium hydroxide. Is the element X a metal or a nonmetal?
  2. Evaporation of the solution prepared from hydrochloric acid and the solid 'A' (from the problem 1) produced some colorless crystalline solid. It was dissolved in water, and the solution obtained was poured into four test tubes. To each of those tubes, aqueous solution of one of the following chemicals were added: magnesium sulfate (to the tube No 1), sodium carbonate (to the tube No 2), sodium sulfide (to the tube No 3), and potassium phosphate. No precipitate was observed in the first test tube, whereas white precipitates were observed in the test tubes 2-4. We know for sure that the element X is one of the elements listed in the table 1 in a CW 3. Can you name the element X?
  3. How would you do the following transformations?
    - (a)
$$\text{ZnCl}_2 \longrightarrow \text{Zn(OH)}_2 \longrightarrow \text{ZnSO}_4 \longrightarrow \text{ZnS} \quad (28)$$
    - (b)
$$\text{Si} \longrightarrow \text{SiO}_2 \longrightarrow \text{CaSiO}_3 \quad (29)$$
    - (c)
$$\text{AlCl}_3 \longrightarrow \text{Al(OH)}_3 \longrightarrow \text{Al}_2\text{O}_3 \longrightarrow \text{Al}_2(\text{SO}_4)_3 \quad (30)$$

Each arrow corresponds to a separate chemical equation, for example, the first step in the scheme 28 can be written as:



Of course, you can choose any auxiliary chemicals you want. For example, instead of KOH you can use NaOH.

4. (a) Find the physical properties of all halogens on Internet<sup>11</sup>: their atomic masses, boiling and melting points. Arrange halogens in order where one of those properties increases.
- (b) Which halogen is the most chemically reactive? Arrange halogens according to its chemical activity.
- (c) Which conclusions can you draw from the results you obtained in (a) and (b)?
5. Glauber's salt, or *sal mirabilis* (Lat. "miraculous salt") was first prepared in 1625 by the German apothecary Johann Rudolf Glauber, who discovered its laxative properties. During the Industrial Revolution, when the need in Glauber's salt increased (it

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<sup>11</sup>The data from the current version of Wikipedia are correct.

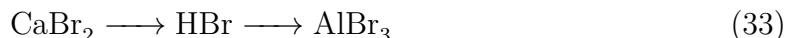
appeared to be a useful starting material for chemical industry), its large scale synthesis was performed by mixing an ordinary salt (table salt) with sulfuric acid in open tanks. Although this method was convenient, it was abandoned soon, because people from nearby houses started to complain that all their iron goods started to rust. Why did it happen? Draw chemical equations of the corresponding chemical reactions.

6. 25 grams of potassium chloride and 100 grams of sulfuric acid were placed into the flask. Immediately after that, the hose was attached to flask's neck, and the gas that formed during this reaction was bubbled through the solution of 12 grams of sodium hydroxide in 200 mL of water. All the gas has been absorbed by the NaOH solution (no bubbles escaped). When evolution of the gas has ceased, few drops of the phenolphthalein solution were added to the liquid. Will the solution turn pink or not?
7. How would you do the following transformations?

(a)



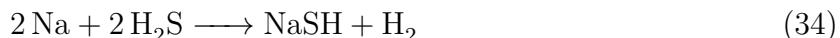
(b)



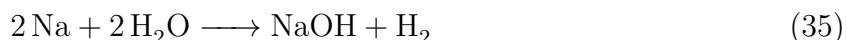
## 16 Sulfur. Oxygen group, a. k. a. Chalcogens

The major difference between oxygen and sulfur is that the former has an extremely low boiling temperature, and is a gas at normal conditions, whereas the latter is a yellow low melting solid insoluble in water. It is solid because, in contrast to oxygen, sulfur atoms do not like to form double bonds with each other<sup>12</sup>, and, since sulfur in a free form is divalent, like oxygen, it forms large circular molecules  $\text{S}_8$ , where each atom of sulfur is connected with other two, and they form an eight member ring.

If we compare chemical properties of oxygen and sulfur, we find some important similarities. Thus, both elements form hydrides ( $\text{H}_2\text{O}$  and  $\text{H}_2\text{S}$ ). Hydrogen sulfide is a gas with a very strong unpleasant odor. It is a weak acid, and it reacts with very active metals, for example, sodium:



in the same way water reacts with sodium:



The compound NaSH is called "sodium hydrosulfide", by analogy with sodium hydroxide (NaOH). Reactions of sulfur with sodium, calcium, zinc, and even iron powder yield sulfides, which look like oxides of the same metal.

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<sup>12</sup>This is a general rule: bigger and heavier atoms prefer to form single bonds with light atoms, whereas light atoms love to form double and even triple bonds when possible.



These reactions are very violent, and they remind similar reactions with oxygen, e.g.:



Obviously, in these reactions the metals are being oxidized, which means sulfur is an oxidizer (in other words, it plays the role of oxygen in these reactions). Sometimes, the reactions between sulfur and metals are even more violent than the reactions of the same metal and oxygen. That occurs for two reasons. Firstly, since sulfur is a solid (a liquid at elevated temperature) reaction is larger than the amount of oxygen atoms, which makes the reaction faster. Secondly, many sulfides do not form a solid film that protect a metal from further reaction. As a result, such a metal as aluminium vigorously reacts with sulfur:



whereas it seems inert towards oxygen (actually, it oxidizes instantly when it comes in a contact with oxygen, however, a thin film of aluminium oxide  $\text{Al}_2\text{O}_3$  is very dense and hard, so aluminium doesn't burn in oxygen).

Sulfides of some metals, such as sodium or calcium, are soluble in water, and, being salts, they participate in exchange reactions, such as:



Although sulfur is an oxidizer, it is a less strong oxidizer than oxygen. As we already know, oxygen oxidizes sulfur: it burns in oxygen yielding sulfur dioxide:



This oxide is a gas with a strong and unpleasant odor (an odor of burning matches). Sulfur dioxide is an acidic oxide, it reacts with water, and a moderate sulfuric acid forms as a result:



This reaction is reversible, which means it is impossible to make a concentrated sulfuric acid, because it slowly decomposes back onto the dioxide ( $\text{SO}_2$ ) and water.

Sulfur dioxide can be oxidized further at elevated temperature and in the presence of some catalysts. The product of this reaction is sulfur trioxide:



Sulfur trioxide is a liquid that reacts with water explosively. The product of this reaction is a sulfuric acid:

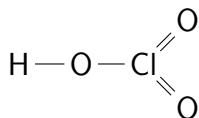


We are already familiar with this acid, which is one of the most common and most important products of chemical industry. It is a strong acid. 100% acid (anhydrous), is a viscous liquid high density liquid. It avidly absorbs water from atmosphere, so an open glass with sulfuric becomes heavier upon standing. Due to its high affinity towards water, sulfuric acid reacts with many organic materials, including sugar, paper, or wood and converts them into carbon. That makes sulfuric acid especially dangerous, because it is corrosive not only due to its acidity, but also because it destroys organic materials causing their dehydration. A video with this reaction is available online: [https://www.youtube.com/watch?v=xK4z\\_YhtTBM](https://www.youtube.com/watch?v=xK4z_YhtTBM).

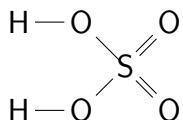
Salts of sulfuric acid are called sulfates.

## 16.1 Sulfates and acid salts

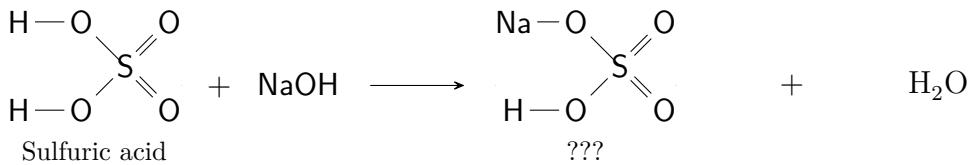
As we already know, acids are the compounds that can donate a hydrogen in a reaction with metals (left of hydrogen in a reactivity series) or with bases. Depending on the acid type, they can donate either one or several hydrogens. Thus, a chloric acid has just one hydrogen atom it can donate:



In contrast, sulfuric acid has two acidic hydrogens:



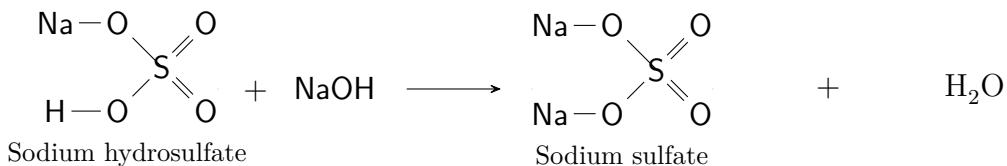
We know each hydrogen atom in sulfuric acid can be substituted with a metal atom, however, is it an “all-or-nothing” process, or such a substitution can be performed stepwise? Let’s imagine we took two water solutions. One solution (solution A) contains 40 grams of sodium hydroxide ( $\text{NaOH}$ ), another solution (solution B) contains 98 grams of sulfuric acid ( $\text{H}_2\text{SO}_4$ ). Note that the mass of sodium hydroxide in the solution A and the mass of sulfuric acid in a solution B are numerically equal to the masses of one  $\text{NaOH}$  and  $\text{H}_2\text{SO}_4$  molecules (40 Da and 98 Da, accordingly). That means the number of  $\text{NaOH}$  molecules in a solution A is equal to the number of  $\text{H}_2\text{SO}_4$  molecules in a solution B. What will happen when we mix these solutions together? Clearly, the amount of  $\text{NaOH}$  molecules is not sufficient to substitute all hydrogen atoms in the sulfuric acid. In our case, the reaction goes according to the following equation:



As we can see, the product of this reaction can simultaneously be considered a salt (it contains a metal atom and an acidic residue) and an acid (the molecule still contains a hydroxide atom that can be substituted in a reaction with a base or a metal). That type of compounds are called *acid salts*.

**Acid salts are the salts having at least one active (replaceable) hydrogen.**

The fact that active hydrogen is still present in the acid salts, a prefix “hydro” is added to its name. For example, the above acid salt ( $\text{NaHSO}_4$ ) is called “sodium hydrosulfate”.<sup>13</sup> Acid salts inherit some properties of acids. Thus, sodium hydrosulfate is capable of reacting with one more molecule of sodium hydroxide to produce sodium sulfate and water:



The acid salt everybody is familiar with is a baking soda (aka sodium hydrocarbonate, aka sodium bicarbonate). Its formula is  $\text{NaHCO}_3$ , and it contains one carbonate residue per one sodium atom (as opposed to 1/2 carbonate residue per one sodium in  $\text{Na}_2\text{CO}_3$ ).

**Important!** The term “acid salt” is somewhat misleading: acid salts are not necessarily acidic. Thus, whereas a sodium hydrosulfate solution is acidic (according to a pH paper), a baking soda solution is not. We will discuss that in more details next year.

## 16.2 Chalcogens, or oxygen family

Besides sulfur, there are two other elements that behave similarly. These elements (we are not familiar with them yet) are called **Selenium (Se)** and **Tellurium (Te)**. Since they have common properties, all of them are found in sulfur containing ores, especially in copper ores. The Greek name of copper is *chalkos*, hence the common name for this group **Chalcogens**.

Their common properties of chalcogen compounds can be summarized in the Table 2.

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<sup>13</sup>Sometimes, a prefix “bi” is used instead of the prefix “hydro”, so the alternative name of sodium hydrosulfate is “sodium bisulfate”, because such a salt contains two times more acidic residue per one metal atom as compared with ordinary salts.

**Table 2. Chalcogen hydrides and oxo acids**

Valence	S	Se	Te	Generalized formula
Hydride	H <sub>2</sub> S	H <sub>2</sub> Se	H <sub>2</sub> Te	H <sub>2</sub> X
IV	H <sub>2</sub> SO <sub>3</sub>	H <sub>2</sub> SeO <sub>3</sub>	H <sub>2</sub> TeO <sub>3</sub>	H <sub>2</sub> XO <sub>3</sub>
VI	H <sub>2</sub> SO <sub>4</sub>	H <sub>2</sub> SeO <sub>4</sub>	H <sub>2</sub> TeO <sub>4</sub> (H <sub>6</sub> TeO <sub>6</sub> )	H <sub>2</sub> XO <sub>4</sub>

## 17 General properties of chalcogens

The properties of chalcogens can be summarized using the same approach as we already used for halogens.

1. All chalcogens form binary compounds with metals. Chalcogen's valence in these compounds is ***always equal to two***.

**Examples:** Na<sub>2</sub>S, Al<sub>2</sub>Se<sub>3</sub>, CaO.

2. Metal chalcogenides are salts (except oxides). The same rule as the one we developed for halogenides works for chalcogenides too: if one chalcogenide (e.g. sodium sulfide) is soluble in water, sodium selenide or telluride will be soluble too. Accordingly, if silver selenide is insoluble, silver sulfide or telluride is insoluble too. **Examples:** CuO is insoluble in water; CuS and CuSe are insoluble too.
3. All chalcogens form binary compounds with hydrogen with a general formula H<sub>2</sub>X, where X is chalcogen. All these compounds (except H<sub>2</sub>O) are acids. Their acidity is low, but, similar to halogen hydrides, ***acidity of hydrogen chalcogenides increases with the increase of chalcogen's atomic mass***: thus, H<sub>2</sub>Se is more acidic than H<sub>2</sub>S.
4. All chalcogens (except oxygen) form binary compounds with oxygen (obviously, oxygen cannot form an oxide with itself). Sulfure, selenium and tellurium form two type oxides, where their valence is either four or six. Accordingly, these oxides produce two different acids (*sulfuric vs sulfurous, selenic vs selenous* etc).

## Homework

1. Compare the properties of chalcogens and halogens summarized during last two lessons. What is the general approach to the description of the properties of these two classes of elements? Do you think this approach can be expanded onto other classes?
2. A 10 g piece of copper ore (a mixture of copper sulfide and other rocks, mostly calcium silicate) was finely ground and mixed with an excess of 10 % hydrochloric acid. The gas formed in this reaction was collected and bubbled through the solution of 100g of silver nitrate in 5000 mL of water, and all the gas was absorbed. The precipitate was collected and dried. Its mass was 15 g. Find the content of copper in the ore.

## 18 Pnictogens, or nitrogen family

The element belonging to that group are nitrogen, phosphorus, as well as three other elements, **Arsenic (As)**, **Antimony (Sb)**, and **Bismuth (Bi)**.

The name of this family comes from Greek word *pnīgein*, meaning “to choke”, or “to suffocate”. The reason for that is obvious: a nitrogen gas does not support respiration. Nitrogen, as well as other pnictogens produce toxic gases (ammoina  $\text{H}_3\text{N}$ , phosphine  $\text{H}_3\text{P}$ , arsine  $\text{H}_3\text{As}$ , and stibine  $\text{H}_3\text{Sb}$ ) when combined to hydrogen. All of them are trivalent in their hydrogen compounds, and can be either tri or pentavalent in oxides or some other compounds with active nonmetals. The Table 3 summarizes the properties of their key compounds:

**Table 3. Pnictogen hydrides and oxo acids**

Valence	N	P	As	Sb	Generalized formula
Hydride	$\text{H}_3\text{N}$	$\text{H}_3\text{P}$	$\text{H}_3\text{As}$	$\text{H}_3\text{Sb}$	$\text{H}_3\text{X}$
Oxides:					
III	$\text{N}_2\text{O}_3$	$\text{P}_2\text{O}_3$	$\text{As}_2\text{O}_3$	$\text{Sb}_2\text{O}_3$	$\text{X}_2\text{O}_3$
V	$\text{N}_2\text{O}_5$	$\text{P}_2\text{O}_5$	$\text{As}_2\text{O}_5$	$\text{Sb}_2\text{O}_5$	$\text{X}_2\text{O}_5$
Acids:					
III	$\text{HNO}_2$	$\text{H}_3\text{PO}_3$	$\text{H}_3\text{AsO}_3$	$\text{H}_3\text{SbO}_3$	$\text{H}_3\text{XO}_3$
V	$\text{HNO}_3$	$\text{HPO}_3$ ( $\text{H}_3\text{PO}_4$ )	$\text{H}_3\text{AsO}_4$	$\text{H}_3\text{SbO}_4$	$\text{H}_3\text{XO}_4$

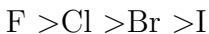
## Homework

1. Sodium sulfide is contaminated with some amount of sodium selenide. A sample of this mixture was dried and weighed. Its mass was 10.11 g. This sample was dissolved in 100 mL of water, a solution of 30 g of  $\text{AgNO}_3$  in 200 mL of water was added. The precipitate was collected by filtration and dried. Its mass was 21.3 grams. What was the content of  $\text{Na}_2\text{Se}$  in the sample?

We outlined six groups of elements based on their valence and on the chemical properties of their oxygen and hydrogen compounds. Now we are prepared for a wider generalization.

## 19 Atomic mass and the properties of elements.

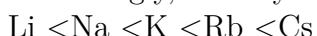
If we arrange halogens according to their nonmetallicity (which we provisionally define as the element's ability to act as an oxidizer) we will see that their nonmetallicity decreases in this order:



In other words, fluorine is the most active nonmetal, whereas iodine is the least active one. Interestingly, the masses of halogens increase in the same order:

Halogen	Atomic mass
F	19
Cl	35.5
Br	79
I	128

Accordingly, activity of alkaline metals (their *metallicity*) *increases* in that order:



That order corresponds to the increase of the element's mass

Alkaline metal	Atomic mass
Li	7
Na	23
K	39
Rb	85
Cs	133

In both cases, we observe some unambiguous dependence between the properties of elements and their masses.

One can conjecture such a dependence is universal, and it can be observed in other groups of elements: chalcogens, alkaline earth metals, and others. Later, we will check if that is the case, and, for a while, let's agree that the lightest element of the group can be used as a representative of the group as whole. In connection to that, let's take the lightest element from each group, and arrange all of them according to their chemical properties. Since the ability to oxidize other element, or the ability to be oxidized are hard to characterize quantitatively, let's choose some other, more quantitative criteria. Such criteria are the element's valence in their oxides and hydrides. Before we started, let's look at one more group of elements, the boron group.

## 20 Boron group

. This group is interesting because it is composed by both metals and nonmetals. The lightest element of this group, boron (chemical symbol B), is a nonmetal (sometimes it is called metalloid because of its metallic luster and hardness). Boron, as well as other representatives of this group (aluminium (Al), gallium (Ga), indium (In), and Thallium (Tl)) form oxides in reaction with oxygen, and their valence in the oxides is three:  $\text{B}_2\text{O}_3$ ,  $\text{Al}_2\text{O}_3$ ,  $\text{Ga}_2\text{O}_3$ , etc. Interestingly, whereas boron is nonmetal, other elements of the boron group are metals, although majority of them are amphoteric.

The boron group elements, as well as other elements are capable of forming hydrides:  $\text{BH}_3$ ,  $\text{AlH}_3$ , etc, where they are also trivalent.

## 21 Periodic law.

Now we know seven groups of elements (halogens, chalcogens, pnictogens, carbon group, boron group, alkaline metals and alkaline earth metals), and we are ready to start to arrange all of them together. Let's take the lightest element from each group and arrange them

according to their mass. In parallel, we will write the formula of the element's hydride and oxide.

Element	Li	Be	B	C	N	O	F
Hydride	LiH	BeH <sub>2</sub>	BH <sub>3</sub>	CH <sub>4</sub>	NH <sub>3</sub>	OH <sub>2</sub>	FH
Oxide	Li <sub>2</sub> O	BeO	B <sub>2</sub> O <sub>3</sub>	CO <sub>2</sub>	N <sub>2</sub> O <sub>5</sub>	-	-
Mass	7	9	10	12	14	16	19

We write the formula of water as OH<sub>2</sub> (not H<sub>2</sub>O), and the formula of hydrogen fluoride as FH, not HF. We do that for consistency only, that is not a correct style (although not a mistake either). In future, we will follow a standard style (H<sub>2</sub>O, HF).

From this table, it is clear that, when the mass of the listed elements increases the valence of the element in its hydride increases from one (in LiH) to four (in CH<sub>4</sub>), and then decreases back to one (in HF). In contrast, the valence of the elements in their oxides increases from one (in Li<sub>2</sub>O) to five (in N<sub>2</sub>O<sub>5</sub>), and never decreases back.<sup>14</sup> Interestingly, if we take the second element of each group, the valence of oxides will monotonously increase from left to right:

Element	Na	Mg	Al	Si	P	S	Cl
Hydride	NaH	MgH <sub>2</sub>	AlH <sub>3</sub>	SiH <sub>4</sub>	PH <sub>3</sub>	H <sub>2</sub> S	HCl
Oxide	Na <sub>2</sub> O	MgO	Al <sub>2</sub> O <sub>3</sub>	SiO <sub>2</sub>	P <sub>2</sub> O <sub>5</sub>	SO <sub>3</sub>	(Cl <sub>2</sub> O <sub>7</sub> )
Mass	23	24	27	28	31	32	35.5

Chlorine (VII) oxide (Cl<sub>2</sub>O<sub>7</sub>) is shown in parentheses because this compound is not stable a pure form. However, its corresponding acid, perchloric (HClO<sub>4</sub>) is a stable and well characterized compound.

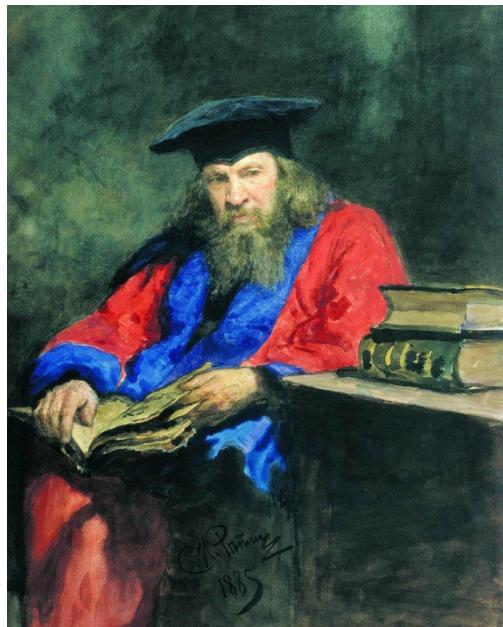
By combining the above rows into a single table, and by adding heavier elements from each group, we get the following table (we limit ourselves with four elements from each group, and we draw a generic formulas of hydrides and oxides):

**Table 1.**

Group	I	II	III	IV	V	VI	VII
Hydride	XH	XH <sub>2</sub>	XH <sub>3</sub>	XH <sub>4</sub>	XH <sub>3</sub>	XH <sub>2</sub>	XH
Oxide	X <sub>2</sub> O	XO	X <sub>2</sub> O <sub>3</sub>	XO <sub>2</sub>	X <sub>2</sub> O <sub>5</sub>	XO <sub>3</sub>	X <sub>2</sub> O <sub>7</sub>
I	Li 7	Be 9	B 10	C 12	N 14	O 16	F 19
II	Na 23	Mg 24	Al 27	Si 28	P 31	S 32	Cl 35.5
III	K 39	Ca 40	Ga 70	Ge 72.5	As 75	Se 89	Br 80
IV	Rb 85	Sr 88	In 115	Sn 119	Sb 122	Te 128	I 127

As we can see, all seven groups fit nicely into this table, and one important thing can be derived from it: atomic masses of elements increase when we move from the top to the bottom, and from the left side to the right. That means the periodic dependence exists between the masses of elements and their chemical properties. Indeed, if we unwrap the above table, we get:

<sup>14</sup>Obviously, oxygen cannot form oxides. Fluorine is more active than oxygen, so it is incorrect to speak about fluorine oxides.



Dmitry Mendeleev (1834-1907), a discoverer of the Periodic Law.

Li Be B C N O F Na Mg Al Si P S Cl .....

In this sequence, when we are going from left to right, the mass of the elements increase *monotonously*, but the properties change *periodically*: element's metallicity gradually decreases from lithium to fluorine, then we *jump* to sodium (an alkaline metal), and then the metallicity decreases again until we arrive to chlorine (a nonmetal). This periodicity was first formulated as a law by a Russian chemist Dmitry Mendeleev.

#### **Properties of elements are periodic functions of their atomic masses.**

Accordingly, each raw in the above table is called a *period*, and the table itself is called *Periodic Table*. Each table's column is called a *group*, and the number of each group is equal to the highest possible valence of the oxides of the elements from that group.

The above classification of elements looks clear and straightforward. However, it is not comprehensive. There is a serious problem with that table, and it required significant efforts to resolve that problem. We will learn about that during the next class.

## **Homework**

1. Watch the following YouTube videos<sup>15</sup>:
  - (a) <http://www.youtube.com/watch?v=D4pQz3TC0Jo>
  - (b) <http://www.youtube.com/watch?v=QoXaArRHfAw>
  - (c) <http://www.youtube.com/watch?v=vRKK6pliejs>
  - (d) <http://www.youtube.com/watch?v=vJslbQiYrYY>

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<sup>15</sup>If you open an electronic version of this document at our web site, you can click at the links directly



Sir Humphry Davy (1778-1829),  
a person who first used electrolysis for isolation of some elements.

Describe what do you see, and draw the equations of these reactions.

In which order can these metals be arranged based on their reactivity towards water?  
How does that order correlate with atomic properties of these metals?

2. Preparation of sodium and potassium metals was done for the first time by Sir Humphry Davy by decomposing sodium (or potassium) hydroxide using electrolysis. As we know from the previous lesson, it is impossible to prepare sodium by electrolysis in the presence of water. Davy solved that problem in a very elegant way: instead of dissolving the alkali in water he heated solid alkali until they melt, and performed electrolytic decomposition of liquid alkali. During those experiments he found that his standard corundum or quartz crucibles could not be used, because the alkali destroyed them easily. Please, draw the chemical equations of the reaction between the liquid alkali and the crucibles' material.

Which material, in your opinion, could Davy use for his experiments?

3. You decided to fill two balloons with hydrogen. To this end, you took 25 grams of zinc, placed it in the flask, added 200 g of 10% HCl, and attached a balloon to the neck of the flask. You also took another flask, and did the same using 20 grams of aluminium and 300 grams of 10% HI. When the reactions in both flasks have ceased, you disconnected the balloons and measured their lifting power. Which balloon will be capable of lifting the greater load?
4. What does the Periodic Law says?
5. How do metallic properties of element change when we move from the top of the Table 1 to the bottom?
6. Which element manifests stronger nonmetallic properties, arsenic or antimony?
7. Which sequence of elements is called a period?

8. List the elements from the third period that form basic oxides.
9. Using the Table 1, can you predict which acid is stronger, sulfuric acid ( $\text{H}_2\text{SO}_4$ ) or perchloric acid ( $\text{HClO}_4$ )? Which hydroxide is more basic, magnesium hydroxide, or strontium hydroxide?

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