CEEA

Chemistry Experiments

A EUROPEAN APPROACH

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coordinator

ŞIMLEU SILVANIEI
2017

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FOREWORD

Chemistry is by default an experimental science. The laws of chemistry, the way the chemical processes undergo, interactions between matter in different contexts can be better understood if experimental work is combined with theory. The chemistry experiment has to be a routine procedure in the laboratory.

We live in a world of instant communications, we are surrounded by electronic devices helping us to communicate with our peers. The available online information is diverse, in a huge amount and virtually accessible by anyone. Nowadays the information amount is doubling every year (can you make a comparison with the 1900's where the information was doubling every 100 years and in 1945 it was doubling every 25 years?!?). But this power of information and information society comes with a cost: there is a huge amount of information but the knowledge of humans how to manage it, is limited. Knowledge is needed in order to ensure a proper understanding and processing of the available information.

If we take, for instance, the case of a huge number of chemical demonstrations available online as chemistry videos. Many of them lack a truly pedagogical approach making them accessible to a wider public. Unexpected phenomena and transformations, presented in a very showy way, which leave the public speechless are available in a huge number online. But what remain in the public's mind after the explosion is gone, after the light has disappeared and the colour has faded … when the movie is over?!?!

This question found an answer in an Erasmus+ strategic partnership between six school across Europe who proposed to apply a pedagogical approach to some of the chemistry demonstrations presented in the classroom or online to the wide public. It is not sufficient to show what is happening in a chemical process ... the public has to be prepared, the process has to be explained and demonstrated, we have to deal with hazards of the used reagents and waste management, the chemical equation have to be shown ... in other words we have to escape the illustrative phase and go on a higher level of fully understanding the whole process and its consequences.

Trying to answer this question:

Colegiul Național ”Simion Bărnuțiu” Șimleu Silvaniei; ROMANIA

together with five schools accepting to join this idea:

- Höhere Technische Bundes- Lehr- und Versuchsanstalt, Dornbirn, AUSTRIA
- Ekpedeftiria Vassiliadi, Thessaloniki, GREECE
- Scholengemeenschap Helinium, Hellevoetsluis, THE NETHERLANDS
- Christelijke Scholengemeenschap Comenius College, Rotterdam, THE NETHERLANDS
- Ozel Samsun Egitim Sentezi, Samsun, TURKEY

shake hands in a two years international collaboration (2015-2017), organized within the Erasmus+ actions, through a strategic partnership called ”Chemistry Experiments a European Approach”, project number 2015-1-RO01-KA219-015157, where the main objective was to make chemistry demonstrations more intelligible and illustrative, for a wider public.

The results of this partnership are:

- a collection of 190 CHEMISTRY VIDEOS, on a basis of a previously presented list, videos of a length of 5-15 minutes and containing the stated parts for a pedagogically correct approach (preamble, chemical hazards, the experiment recorded on video, the chemical equation, the explanation of phenomena and waste management).

- a YouTube account with all the videos uploaded on-line, open to the users who are searching the online database for related subjects.

- a HANDBOOK OF CHEMISTRY EXPERIMENTS joining the results of the team during two years of cooperation as a learning device and as a dissemination tool for the project. The book is written in English and it is translated into Romanian, Greek and Turkish.

- a CHEMISTRY SOFTWARE developed by the GR team, containing all videos, recorded on a USB stick, in a form of an off-line learning tool. This software will allow users to access and use the video database without on-line connection. The software will allow the user, to browse off line specific subjects related to Chemistry (recorded in the database joined to the software) to access and present chemical subjects related to Chemistry experiments and/or practical demonstrations.

Through this three pedagogical auxiliaries we address the chemistry teachers offering them a theoretical basis and a methodological one to setup experimental work in chemistry lessons, to illustrate, during different learning sequences, different techniques and phenomena which might be difficult to do in their school, although very relevant or to extend the number of the experimental works they apply in the school.

We address also the secondary students, whose native curiosity, doubled by passion for chemistry will allow them to deepen their chemical knowledge, to extend their practical skills, learning new methods, or maybe different approaches, to help them see what is behind different chemical reactions and better understand the secrets of life.

The three end products were realized with a huge amount of work, done by each partner team. The activities were a challenge for everyone for they required resourceful people, capable to cope with various tasks: not only chemistry and experimental setup but also video recording and editing, voice recording and editing, photography and image editing, documents creating, editing and book assembly, language knowledge and communication skills. Many times people involved in the project "learned by doing" and it was not an easy task to keep the scientific rigour, to focus on the essence of the chemical process while addressing the secondary students who will be the final beneficiaries of these end products. The multiple and different tasks proposed for three years had to be accomplished at the same quality level within two years. This lead to a huge burden pushing the teams in order to be able to finish in time the assumed job and to harvest now the fruits of the partnership, after all the videos are online, after the chemistry software is accomplished and the video database joined to it and after the book is printed out.

At the end of the project, which is in fact the beginning of this publication's life, we acknowledge and thank the involvement of all the participants, I express my gratitude and consideration for the team members which were steering in the right direction the work and the activity complex done in the project, a say many thanks to the team leaders which kept close and, with a supportive attitude and good risk management, ensured the accomplishments of all assumed end products.

I raise my hat in front of the those colleagues, both from the local team and from the partners, who succeeded to do quality work, to respect the deadlines and to provide the requested materials, who succeeded to give a personal touch to the end products due to their passion, perseverance and scientific rigour. This was a project in which "thinking differently" was made synonymous with the words "collaboration" and "partnership". For all their efforts I express my appreciation, gratefulness and many thanks for the good work and collaboration.

Mircea BRĂBAN, PhD
project coordinator

DEAR READERS,

We are glad we were able to arouse your curiosity to open the book!

The contents is primarily intended for secondary school pupils. At the age of about 14, they make a first acquaintance with chemistry. We teachers are supposed to be familiar with the necessary scientific backgrounds but for the pupils´ sake we should reduce complexity to a minimum. Therefore, we decided to keep reaction equations as simple as possible. Wherever possible, interactions with solvents are ignored, so dissociation into ions and formation of hydration layers around ions are omitted. These effects can be included in a later stage yet.

Please take in account that the experiments contained herein should only be conducted by qualified chemists who are aware of proper chemical and apparatus usage.

Safety precautions such as the use of goggles, aprons, gloves, ... are assumed. It is also assumed that all chemicals used will be disposed in a proper way.

Although the authors elaborated the experiments in an accurate way, they reserve the right not to be responsible for the topicality, correctness, completeness, update and quality of the information provided.

HOMOGENEOUS AND HETEROGENEOUS MIXTURES

OVERVIEW

A mixture is a system of materials made up of two or more different mixed substances but not combined chemically. A mixture refers to the physical combination of two or more substances on which the identities are retained. Many common materials of daily life are mixtures. Mixtures can either be homogeneous or heterogeneous. A homogeneous mixture is a type of mixture in which the composition is uniform and every part of the solution has the same properties. A heterogeneous mixture is a type of mixture in which the components can be seen, as there are two or more phases present. In this experiment we will study the differences between homogeneous and heterogeneous mixtures [1].

MATERIALS AND METHODS

Use water as a solvent, sand, ink, wine, salt, sugar and oil as solutes and six glass test tubes which are mounted in a test tube rack. Add a small quantity of water approximately 3 mL in each test tube. Then add a small amount of each substance in each test tube [2]. The differences between homogenous and heterogeneous mixtures are obvious.

HAZARDS

No hazards posed.

CONCLUSIONS

We can observe that in mixtures of water with ink, wine, sugar and salt the components are not visible because they are homogeneous mixtures or solutions. However, in mixtures of water with oil and water with sand the components are visible, because they are heterogeneous mixtures. Water, then, although it is the most common solvent, cannot dissolve all the substances. Thus, we notice the differences between homogeneous and heterogeneous mixtures.

WASTE MANAGEMENT

All mixtures can be poured down the drain. Mixture of sand and water should be first filtrated.

PHOTOS

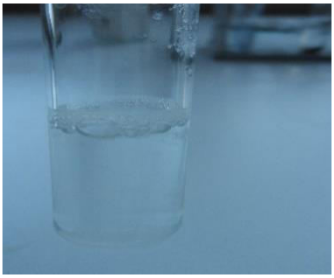

Fig. 1 Homogenous mixture of water and salt.

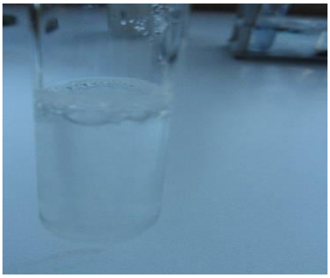

Fig. 2 Homogenous mixture of water and sugar.

Fig. 3 Homogenous mixture of water and wine.

Fig. 4 Heterogeneous mixture of water and oil.

Fig. 5 Heterogeneous mixture of water and sand.

Fig. 6 Homogenous mixture of water and ink.

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SEPARATING MIXTURES THROUGH EVAPORATION

OVERVIEW

In this experiment we study the way of separating mixtures by evaporation. Evaporation is great for separating a mixture (solution) of a soluble solid and a solvent. The process involves heating the solution until the solvent evaporates leaving behind the solid residue.

MATERIALS AND METHODS

Place 50 mL of water in a beaker and add a small amount (approximately 10 grams) of salt. Then stir the mixture thoroughly with a glass stirring rod until it becomes completely homogenous. Finally, place the beaker on a hot plate and note that in a few minutes the mixture boils and water vapor escapes from the beaker while salt remains at the bottom of the beaker [1].

HAZARDS

No hazards posed.

CONCLUSIONS

When the boiling process is completed, the entire amount of water has escaped in the form of water vapor, while the salt remains at the bottom of the beaker so the mixture has been separated. After weighing the beaker, the entire amount of salt remains in the beaker. Evaporation is a common way to collect from the sea fine salt for the kitchen. The salt of the sea water is taken up as follows: The water is collected in shallow pools, which are called salt marshes. There, with the aid of solar heat and wind it evaporates. In the Nordic countries, where no solar heat is used, the operation is as follows: The water freezes at a low temperature, so the pure ice is separated and a solution of salt is left behind. This solution is then evaporated to obtain the pure product.

WASTE MANAGEMENT

The salt can be poured down the drain.

PHOTOS

Fig. 1 Solution of water and salt.

Fig. 2 Boiling of the solution and water evaporation.

Fig. 3 The remaining salt.

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SEPARATING MIXTURES - CHROMATOGRAPHY

OVERVIEW

In this experiment we study the way of separating mixtures of chromatography, which is a physical separation process. Paper chromatography is an analytical method used to separate colored chemicals or substances. It is primarily used as a teaching tool, having been replaced by other chromatography methods, such as thin-layer chromatography. Paper chromatography is a useful technique because it is relatively quick and requires small quantities of material. In paper chromatography, substances are distributed between a stationary phase and a mobile phase. The stationary phase takes place when the water is trapped between the cellulose fibers of the paper. The mobile phase is a developing solution that travels up the stationary phase, carrying the samples with it. Components of the sample will separate readily according to how strongly they adsorb onto the stationary phase versus how readily they dissolve in the mobile phase.

When a colored chemical sample is placed on a filter paper, the colors separate from the sample by placing one end of the paper in a solvent. The solvent diffuses on the paper, dissolving the various molecules in the sample according to the polarities of the molecules and the solvent. If the sample contains more than one color, that means it must have more than one kind of molecules. Because of the different chemical structures of each kind of molecule, the chances are very high that each molecule will have at least a slightly different polarity, giving each molecule a different solubility in the solvent. The unequal solubilities cause the various color molecules to leave solution at different places as the solvent continues to move up the paper. The more soluble a molecule is, the higher it will migrate up the paper. If a chemical is very non-polar it will not dissolve at all in a very polar solvent. This is the same for a very polar chemical and a very nonpolar solvent.

It is very important to note that when using water (a very polar substance) as a solvent, the more polar the color, the higher it will rise on the paper. [1].

MATERIALS AND METHODS

Cut four filter paper strips about 1.5 cm thick. In each strip form a line with a different color which are mixtures of several pigments. In a beaker place a small amount of water of about twenty ml. Then, place the strips vertically in the beaker so that their bottoms are wetted by the solvent. After some minutes, the solvent wets the strips and drifts the pigments upward with different speed and splits.

HAZARDS

No hazards posed.

WASTE MANAGEMENT

All the filter paper strips can be placed in a trash bin while the water can be poured down the drain.

PHOTOS

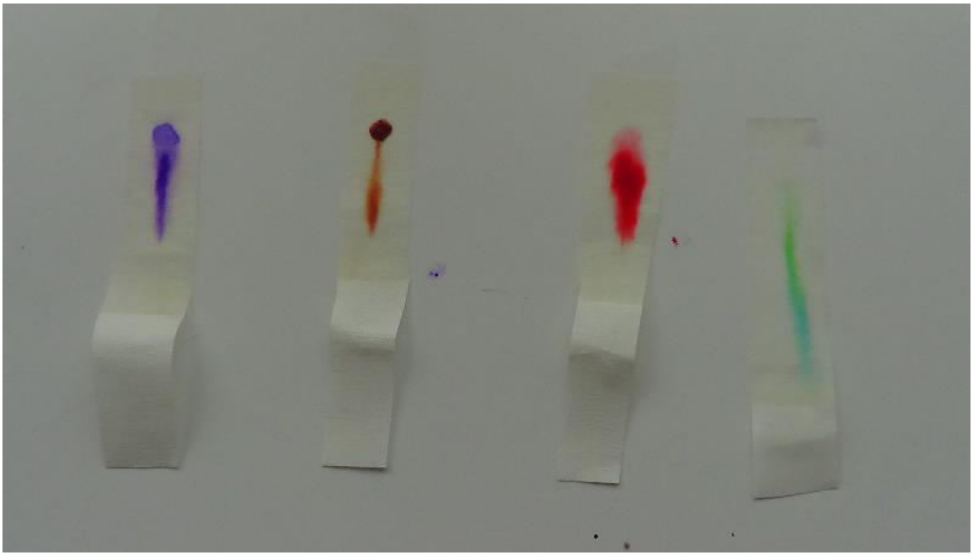

Fig. 1 The pigments have been separated

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SEPARATING MIXTURES USING EXTRACTION

OVERVIEW

In this experiment we study the way of separating mixtures using extraction, which is a physical separation process. Beverages cover a vast variety of addictive drinks out of which tea and coffee are the most popular drinks. Tea powder is extracted from tea leaves which contain tannins, which are acidic in nature, a number of colored compounds and a small amount of unrecompensed chlorophyll and an important stimulant called caffeine. Caffeine can stimulate the nervous system and can cause relaxation of respiratory and cardiac muscles. Caffeine is well known to increase both the alertness level and attention span. But like all other addictives, tea also shows withdrawal symptoms like headache, nervousness and insomnia for a regular consuming person [1].

MATERIALS AND METHODS

Initially, place tea leaves in a beaker and mix with about 100 mL of water. Then the beaker is transferred to an electric eye. Within minutes the mixture starts boiling [2].

In the chemistry laboratory you have another option: to do an extract using a Soxhlet apparatus. This apparatus allows you to do the extraction and to collect the liquid which is not mixed with the leaves from which we extract the components.

HAZARDS

No hazards posed. Just pay attention when handling hot liquids as they can cause severe burns if not paying attention when manipulating them.

CONCLUSIONS

We notice that during boiling, the pigments and the aroma compounds are transferred from the tea leaves in the solvent, so they are extracted. This method is used at home to do tea, or coffee and sometimes other extracts used in the kitchen.

WASTE MANAGEMENT

The tea leaves can be thrown in a bucket while the tea extract in the sink.

PHOTOS

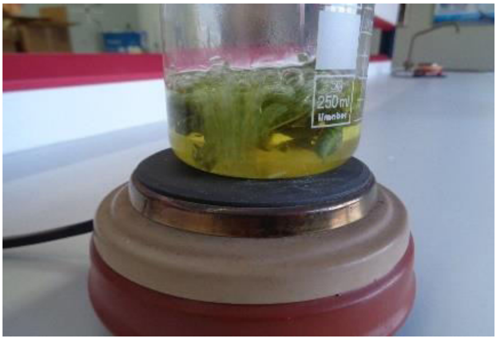

Fig. 1 The mixture boils.

Fig. 2 The pigments and the aroma compounds are transferred from the tea leaves in the solvent.

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SEPARATING MIXTURES USING FILTER PAPERS

OVERVIEW

In this experiment we attempt to separate mixtures using paper filters. Filtration is a technique used to separate an insoluble solid from a liquid using a filter paper. The fluid that passes through is called the filtrate.

MATERIALS AND METHODS

Initially, place 100ml of water in a beaker and add 15 grams sand. Then use a polyethylene or glass funnel with stem and filter paper. The paper is folded in half, then folded in quarters and the tip of one corner is torn off to allow for a snug fit in the funnel cone. The funnel is supported with a ring and a clean beaker is placed beneath the funnel. Then, the filter paper is wetted completely with the solvent. This step adheres the filter paper to the funnel walls preventing solid from escaping. Then the supernatant liquid is poured through the filter first. This allows the initial part of the filtration to proceed faster and may prevent clogging of the filter by the solid. The solid must be scraped onto the filter with a spatula [1]. The mixture of water and sand has been separated.

HAZARDS

No hazards posed.

CONCLUSIONS

By filtration we separate the mixture of water and sand. The solid material remains on the paper filter, while in the beaker the filtrate passes. The filter paper has pores which allow the particles of the fluid to penetrate but not the particles of solid material because they are larger. Filtration techniques are often used at home. Filtering of coffee, water and tea are some corresponding examples of filtration techniques at home. Also, filtration techniques are very common in industry.

WASTE MANAGEMENT

The filtrate can be poured down the drain while the filter paper can be placed in a bin.

PHOTOS

Fig. 1 The solid material remains on the paper filter.

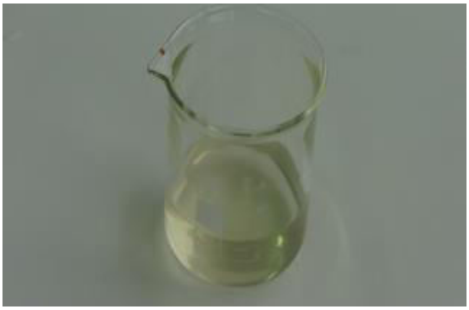

Fig. 2 The filtrate after filtration. Some sand particles penetrated the pores of the filter paper for this reason, the filtrate is not completely clear.

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THE USAGE OF A BUNSEN BURNER IN A LABORATORY

OVERVIEW

The aim of this experiment is to learn how to use a Bunsen burner in a laboratory. One of the most used equipment in a laboratory is a Bunsen burner. A Bunsen burner is equipment that helps propane, butane gases which are pure gases or producer gas that is the mixture of different types of flammable gas mixtures to burn. According to the gas types the burner gives off 400-500oc to 900-100oc heat. Assuming perfect combustion conditions, such as complete combustion under adiabatic conditions (i.e., no heat loss or gain), the adiabatic combustion temperature can be determined. The formula that yields this temperature is based on the first law of thermodynamics and takes note of the fact that the heat of combustion is used entirely for heating the fuel, the combustion air or oxygen, and the combustion product gases.

MATERIALS AND METHODS

The main points that should be known when using a Bunsen burner are;

1. When the burner is being lit, a match should be in waiting position at the barrel of the burner before the gas valve is slowly turned.
2. When a lit burner is not provided enough oxygen it will burn as a sooty or yellow flame. The sooty flames heat is relatively quite low to that of the flame without soot. In a situation like this the collar is opened in order for the flame to be burning without soot.
3. When the air holes are opened too much the flame might extinguish. In this situation the gas valve must be turned off straight away.
4. When the gas is not provided enough the burner might burn from the inside. A situation like this if not paid attention to, may not be noticed. If the burner burns from the inside, the gas valve must be turned off immediately, and the heated burner should be put aside to cool down.
5. The burners' entire flame is not at the same heat. The heat at the base of the barrel is the lowest, whereas the top of the barrel is the highest. In other words, the heat in the outer surrounding of the blue flame is relatively higher than the inner surrounding. According to this it can be stated that the outer surrounding of the blue flame, which is one third above the barrel is of highest heat.

6. The burners flame is separated into sections, oxidizing and reducing. The oxidizing section is where the heat is the highest (the blue flame), as for the reducing section the heat is the lowest (the colourless flame closest to the barrel).

HAZARDS

Methane

CHEMICAL REACTION EQUATION

$CH_{4(g)}$ + 2 $O_{2(g)}$ ---> $CO_{2(g)}$ + 2 $H_2O_{(l)}$

CONCLUSION

In the oxidizing section, due to there being enough oxygen the oxidizing process is seen, whereas in the reducing section from the lack of oxygen the reducing process is seen.

PHOTOS

Fig. 1 Lightning up the gas burner

Fig. 2 Adjusting the volume of the burned gas.

Fig. 3 A yellow flame means a "cold" flame whereas a "blue" one is hot

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THE LAVOISIER’S LAW APPLIED TO THE REACTION OF HCl WITH SODIUM CARBONATE

OVERVIEW

Lavoisier’s Law (the law of conservation of mass) is generally applicable to all chemical reactions. The total mass of all reaction products is always equal to the total mass of the reactants. In some situations gaseous products are generated, in the case they escape, it might appear as if mass has disappeared. When the experiment is carried out in an enclosed space, so the gas can’t escape, a conservation of mass can be observed. During the reaction in this experiment, carbon dioxide is generated but when a rubber plug is placed on the Erlenmeyer in time this will stay inside the Erlenmeyer.

MATERIALS AND METHODS

- Sodium Carbonate
- 10 mL 0.1 M HCl
- Scale
- Test tube
- Erlenmeyer 100 mL
- Erlenmeyer 50 mL
- Rubber plug

Prepare all materials for the experiment. Place the empty Erlenmeyers and the rubber plug on the scales and set the mass to 0 grams. Measure approximately 1 gram of sodium carbonate in the 100 mL Erlenmeyer and place the test tube with 10 mL of hydrochloric acid in the 50 mL Erlenmeyer.

Write down the total mass of the whole. Afterwards pour the hydrochloric in the 100 mL Erlenmeyer and place the rubber plug on it as soon as possible.

Place the test tube back in the 50 mL Erlenmeyer and write down the mass of the whole.

Repeat the experiment but this time don’t place the rubber plug on the100 mL Erlenmeyer.

HAZARDS

HCl		**Na_2CO_3**	

CHEMICAL REACTION EQUATION

$2H^{+}_{(aq)} + CO_3^{2-}_{(aq)} \rightarrow H_2O_{(l)} + CO_{2\,(g)}$

CONCLUSIONS

When diluted hydrochloric acid reacts with sodium carbonate you will not only generate water but also carbon dioxide. When the carbon dioxide is free to escape it is clear that the mass of the products on the scale after the reaction is lower than before. It might look as if Lavoisier's Law is not applicable to this reaction. However, when the process is carried out in a closed Erlenmeyer the total mass after the reaction is equal to the total mass before the reaction.

WASTE MANAGEMENT

The generated reaction product can be flushed down the drain.

PHOTOS

Fig. 1 Scale with Erlenmeyer with sodium carbonate and test tube with hydrochloric acid

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2. http://www.cdc.gov/niosh/ipcsneng/neng0163.html
3. https://en.wikipedia.org/wiki/Conservation_of_mass

THE LAVOISIER'S LAW APPLIED TO THE REACTION OF Fe WITH S

OVERVIEW

It is one of the basic laws of chemistry. Since this law was defined in 1789 by French scientist Antonie Lavoisier, it is thus known as **Lavoisier Law**.

According to this law: The total mass in a chemical reaction equals the mass of the reactants.

According to this;

- In **X + Y → Z + T** reaction X and Y are reagents, Y and T are products and they are equal in terms of mass.

(Besides the conservation of mass in a chemical reaction, the amount of atom, type of atom, electrical load, proton, neutron, electron amount and the nucleic structure of the atoms are also conserved. The molecule amount and the physical state of the matters are not conserved in all reactions.)

- Atoms of an object cannot be created or destroyed, but can be moved around and be changed into different particles.
- Atoms in a chemical reaction are rearranged and since no atom will be lost in the system the total mass in a reaction remains constant.

In physical states and chemical reactions mass is conserved.

If the elements in a chemical reaction are completely finished, then it is called an un-remaining reaction, if at least one of them is finished then it is called a 100% full effective reaction.

MATERIALS AND METHODS

An empty glass clock is taken and 4g of sulphur powder and 7g of iron powder are weighted and stirred. Then a beaker is put on the scale with an empty test tube and its weight is reset. Sulphur powder and iron powder are put into the test tube and the weight is measured. The spirit stove is it. The test tube is then taken with a tube pincer and heated on a spirit stove. The heat is given until the mixture turns to liquid. The test tube is left to cool down and weighted.

CHEMICAL REACTION EQUATION

$\mathbf{Fe_{(s)} + S_{(s)} \text{ ---> } FeS_{(s)}}$

CONCLUSIONS

Even though the iron powder is grey and the zinc powder is yellow, the new substance is black in colour. When this substance is examined with a magnifier the zinc and iron particles will not be seen. In the beginning the mass of the iron and sulphur powder mixture is used as a reactive. When they are compared with the results after the reaction, the total amount of the mass is the same. No change will be seen.

WASTE MANAGEMENT

If the obtained iron sulphate compound is in contact with acid a toxic gas will form that is why the test tube should be broken and the iron sulphate compound should be put in a sealed bowl.

PHOTOS

Fig. 1 After the watch glass is tarred, sulphur and iron powder are weighed in a row and mixed.

Fig. 2 A test tube is put in a beaker and tarred.

Fig. 3 The sulphur and iron powder in the watch glass are emptied into the test tube that was tarred, and the weighing is taken note of.

Fig. 4 The test tube is held by pincers and heated over the spirit stove until smoke starts and observe until the mixture turns into a darker liquid colour.

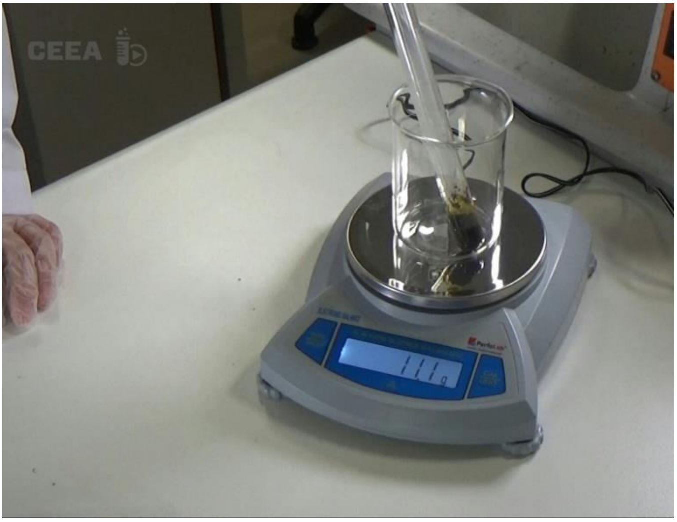

Fig. 5 The obtained dark coloured component is iron sulphur (FeS) and when the tube is cooled the weighing is repeated and when compared with the first weighing it is observed that no mass is lost.

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ALCOHOL CONTENT IN BEVERAGES

OVERVIEW

In chemistry, an alcohol is any organic compound in which the hydroxyl functional group (–OH) is bound to a saturated carbon atom. The term alcohol originally referred to the primary alcohol ethanol (ethyl alcohol), the predominant alcohol in alcoholic beverages.

The word alcohol appears in English as a term for a very fine powder in the sixteenth century. It was borrowed from French, which took it from medical Latin.

Ultimately the word is from the Arabic كحل (al-kuḥl, "kohl, a powder used as an eyeliner"). Al- is the Arabic definite article, equivalent to "the" in English, alcohol was originally used for the very fine powder produced by the sublimation of the natural mineral stibnite to form antimony sulfide Sb_2S_3 (hence the essence or "spirit" of the substance), which was used as an antiseptic, eyeliner, and cosmetic (see kohl (cosmetics)). According to this theory, the meaning of alkuhul would have been first extended to distilled substances in general, and then narrowed to ethanol.

The term ethanol was invented in 1838, modeled on the German word äthyl (Liebig), which is in turn based on Greek aither "ether" and hyle "stuff" [1].

Ethanol is the highest in importance and in quantity component of alcoholic drinks. It is formed during fermentation, i.e. the anaerobic biological degradation of the sugars - mostly glucose and fructose musts of grapes-by baker. It is used in alcoholic beverages, as antifreeze, antiseptic, fuel, preservative and solvent. Measure of the ethanol content of a beverage is the alcoholic strength defined as cubic centimeters of dry ethanol contained in hundred cubic centimeters of solution at 20°C [2].

Tsipouro is a Greek beverage which started its course about seven centuries ago at Mount Athos monasteries. The raki from Crete is something similar, but the main difference is that the raki is a product of single distillate. In other countries, similar drinks are the Italian Grappa, the Arak from Middle East and Zivania from Cyprus.

Tsipouro has about 36-45%v/v content in ethanol. It is produced by distillation of marc ie from rags (residues) of grapes left after pressing and extraction of the must for wine production. In case there is distillation of whole grapes, the yields are higher. Both white and red grapes are appropriate.

When distillation is sometimes added to the retort, other than the marc, various flavorings such as anise, fennel, etc. or in Crete walnut leaves. The anise is the cause of the "whitening" of tsipouro when water or ice is added.

Often tsipouro is distilled a second time, because it improves its quality.

MATERIALS AND METHODS

In a graduated cylinder add ethanol (70%v/v) to 2/3. The alcoholmeter is immersed in the graduated cylinder and expected to balance. Then read the evidence of the alcoholmeter which shows us the content percentage from volume to volume. Then repeat the measurements for tsipouro in which is added water in a ratio of 1/1 in volume. The content of alcohol in tsipouro is 20%v/v.

HAZARDS

Wear eye shields and gloves. Avoid the contact with open flame.

CH_3CH_2OH

WASTE MANAGEMENT

The used ethanol should be disposed as hazardous waste in the appropriate waste container.

PHOTOS

Fig. 1 Alcohol content in tsipouro with water.

Fig. 2 Alcohol content in tsipouro.

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RED WINE DISTILLATION

OVERVIEW

Yeasting of sugar during the production of wine normally results in a maximum volume of 15% alcohol. Situations in which we would like a higher percentage of alcohol can be envisioned. In order to realise this, the wine can be distilled. In this the alcohol will be evaporated and later condensed, which means it can be collected nearly pure.

MATERIALS AND METHODS

- Micro distillation set
- Heating block
- Thermometer
- Pipette
- Stand

Prepare the micro distillation set for use and place the heating block under the round bottom flask. Use the pipette to insert 5 mL of red wine into the round bottom flask and place the thermometer into the flask, resting above the wine. Turn the heating block on and wait for the alcohol to evaporate. After a while you will see the condensed alcohol trickle into the Erlenmeyer flask.

HAZARDS

Ethanol

CHEMICAL REACTION EQUATION

$\mathbf{C_2H_5OH_{(l)} \rightarrow C_2H_5OH_{(g)} \rightarrow C_2H_5OH_{(l)}}$

WASTE MANAGEMENT

The remaining red wine and alcohol should be disposed with organic waste.

PHOTOS

Fig. 1 Heat block with sand bath.

Fig. 2 Overview of the set up.

REFERENCES

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IDENTIFICATION OF CO_2 AND O_2 WITH BURNING WOODEN STICK

OVERVIEW

In this experiment we study the production and identification of both carbon dioxide and oxygen. Sodium bicarbonate is a chemical compound with the formula $NaHCO_3$. It is a salt composed of sodium ions and bicarbonate ions. Sodium bicarbonate is a white solid that is crystalline but often appears as a fine powder. It has a slightly salty, alkaline taste resembling that of washing soda (sodium carbonate). The natural mineral form is nahcolite. It is a component of the mineral natron and is found dissolved in many mineral springs. It is one of the food additives encoded by European Union, identified as E 500. Since it has long been known and is widely used, the salt has many related names such as baking soda, bread soda, cooking soda, and bicarbonate of soda [1].

The hydrogen and oxygen, except for water, form hydrogen peroxide with formula H_2O_2 but with much more different properties than those of water. In pure form it is a liquid like syrup with pale blue color. It is used as an antiseptic, in hair bleaching, in textile fibers and in pulp bleaching. It can be used even in the sewage treatment. In most reactions H_2O_2 acts as an oxidizing agent but in the presence of strong oxidants it acts as a reducing agent to yield oxygen and hydrogen. Potassium permanganate is one of these powerful oxidizing agents. Violet solutions of KMnO4 reduced to weak pink manganese (II) ions in the presence of H_2O_2 [3]. Potassium permanganate is an inorganic chemical compound. More specifically it is a salt consisting of K^+ and MnO^{-4} ions. Almost all applications of potassium permanganate exploit its oxidizing properties. As a strong oxidant that does not generate toxic by-products, $KMnO_4$ has many niche uses in water treatment and disinfection, synthesis of organic compounds, analysis, fruit preservation, survival kits and in fire service [4].

In this experiment we will use the combustion reaction which is accompanied by heat release and perhaps light, often combined with a flame appearance by hot gaseous products, or flash. Combustion requires oxygen in order to take place.

MATERIALS AND METHODS

In a glass test tube place about 15 grams of sodium bicarbonate and pour a few drops of citric acid (lemon juice). Then place the tube in a test tube rack. In another test tube place five mL aqueous solution of potassium permanganate (10%w/v) and 5 mL solution of hydrogen peroxide (30%v/v) and place it in the test tube rack. Then light a wooden stick and approach the first test tube. Then move the wooden stick closer to the second test tube and observe. Two different gases are produced.

HAZARDS

Attention should be given to the fact that $C_6H_8O_7$ and O_2 are flammable and they should not come into contact with an open flame. Wear protective glasses and gloves. Use a respirator filter.

$C_6H_8O_7$		**$KMnO_4$**	
CO_2		**H_2O_2**	
MnO_2		**KOH**	
O_2			

CHEMICAL REACTION EQUATION

$$CH_3COOH_{(aq)} + NaHCO_{3(s)} \longrightarrow CH_3COONa_{(aq)} + CO_{2(g)} + H_2O_{(l)}$$

$$2KMnO_{4(aq)} + 3H_2O_{2(aq)} \longrightarrow 2MnO_{2(aq)} + 2KOH_{s} + 3O_{2(g)} + 2H_2O_{(l)}$$

CONCLUSIONS

Above the first tube, the stick stops burning due to the production of CO_2 and the acidic solution effect in carbonates. Above the second tube, the wooden stick rekindled due to the catalytic decomposition of H_2O_2 and the simultaneous production of O_2.

WASTE MANAGEMENT

The $C_6H_7O_7Na$ can be placed in the bucket. The second solution should be neutralized. Add extra hydrogen peroxide to the solution to convert the Mn^{2+} ions to MnO_2, a brown solid. Separate the resulting mixture by filtration and dispose of the solid MnO_2 in an appropriate container.

PHOTOS

Fig. 1 The stick stops burning due to the production of CO_2.

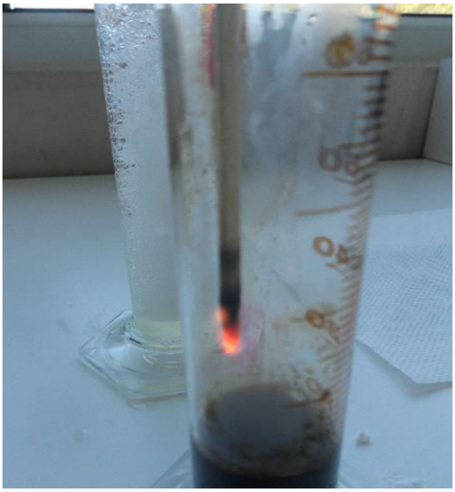

Fig. 2 The wooden stick rekindles due to the catalytic decomposition of H_2O_2 and the simultaneous production of O_2.

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FABRICATION OF SOAP

OVERVIEW

Both animal fat and vegetable oils are quite complex esters formed by fatty acids with glycerine. These esters are called glycerides or triglycerides. Soaps are carboxylic acid salts obtained through the reaction of vegetable oils and animal fat with alkaline bases. They are materials used for cleaning. By adding sodium chlorine into the sodium salt solutions of fatty acids with high carbon, it is possible to settle the sodium salts. The material obtained in this way is called soap. The fats or oils which have saponified are quite tender and are easy to put into the desired shape. When KOH is used instead of NaOH, 'soft soap' is obtained. The chain length of the hydrocarbon and the double bonds in the acids determine the properties of soap. Soap is ineffective in hard water. Hard water contains Ca^{2+}, magnesium Mg^{2+}, and iron Fe^{2+} and Fe^{3+}. Soap forms a calcium soap in hard water. This calcium salt of fatty acids cannot dissolve in water and the soap becomes insoluble. This being the case, soap is ineffective in hard water.

MATERIALS AND METHODS

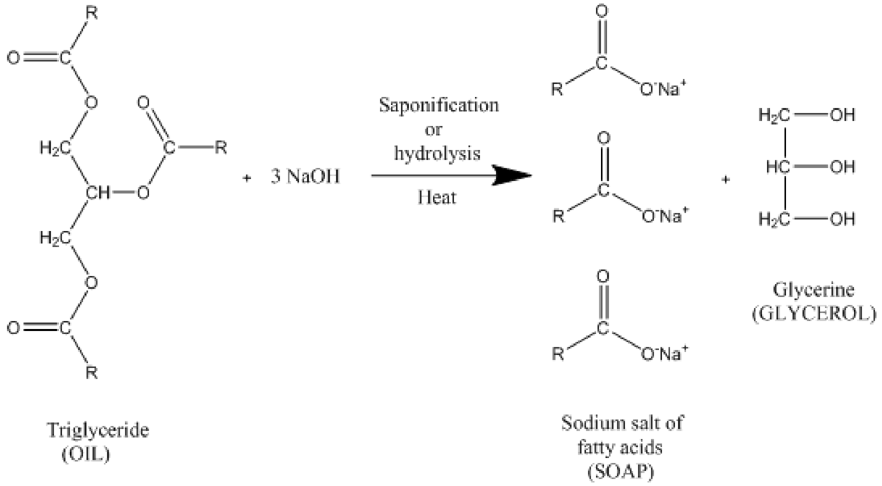
Put 20 g vegetable oil into the beaker. Heat it over the burner. Add slowly the ethyl alcohol and NaOH solution. Boil the mixture under low heat, until the oil drops disappear in the mixture and form a whitish homogeneous paste. After boiling it about for a while, add 50 mL water into the solution and continue to boil it. Prepare a solutıon of 20g sodium chlorine with 30 mL water and add it into the boiling mixture. Let it cool for almost one hour until a precipitation is obtained. Filter this precipitation wash it with distilled water and 50 mL thinned sodium chlorine. Then, if desired, adding dye and perfume, put it into molds and leave it to dry. To test the synthesized soap, take a small piece of soap into the beaker, add water and rinse it. Soap is obtained if foam forms.

HAZARDS

NaOH
(Sodium hydroxide)

C_2H_5OH
(Ethyl Alcohol)

CHEMICAL REACTION EQUATION



WASTE MANAGEMENT

The material obtained is diluted and discharged into a sink.

CONCLUSIONS

Oils, in general esters, when boiled for a long time with concentrated sodium hydroxide or sodium solution are transformed into salt or alcohol and acid, that is, they are hydrolysed. This process is called the saponification of fats. When the fats are saponified, they transform into sodium salts that dissolve in water of glycerine and fatty acids.

PHOTOS

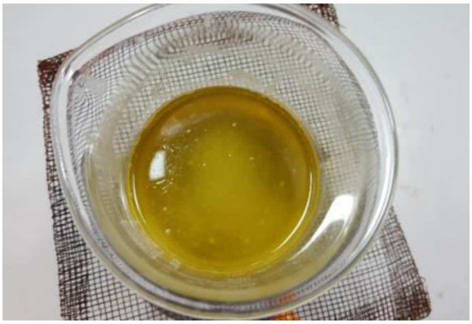


Fig. 1 Put ethyl alcohol and NaOH solutions into the vegetable oil.

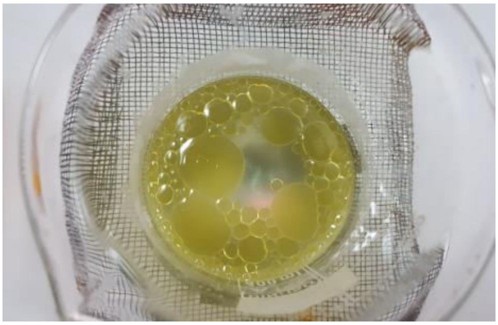


Fig. 2 Boil it until oil drops completely disappear.

Fig. 3 Obtain a whitish homogenous paste.

Fig. 4 Wash this paste with thinned NaCl solution and distilled water. Add food coloring and strain it.

Fig. 5 The soap obtained after straining is poured into a mold.

Fig. 6 The dried soap is taken out of the mold.

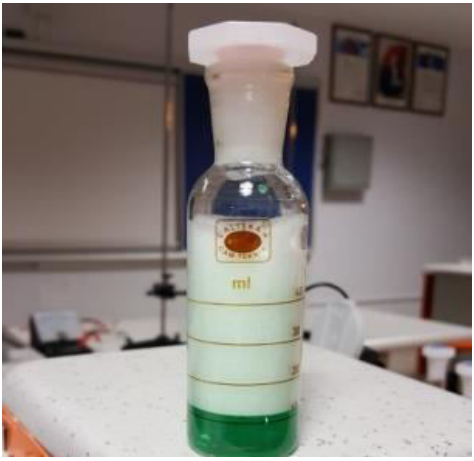


Fig. 7 Checking whether the soap has washing power.

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ALUM CRYSTALIZATION

OVERVIEW

If the saturated solutions of two different salts in stoichiometric ratio are crystallized together, the state produced is called double salts. Although, they are specific substances, in crystallized form, they are compounds acting as associate solutions of two simple salts in their liquid solutions. The double salt property is valid only in a crystallized phase. Namely, although metals give reactions special to themselves, common anions give reactions special to themselves too. These products consist of compound salt structures. Double salts are crystallized in crystal bond so as to consist excessive amount of crystal liquids in their structures. Contrary to coordinated compounds, double salts lose their properties in the solution and they separate into ions, which form them, when dissolved in water. In terms of chemical and physical properties, double salts carry the properties of the compounds that constitute them. When the chemical properties of double salts are studied carefully, they have a sweet taste, and are soluble in water. The double salts which crystallize in octahedral state show an acidic property towards litmus paper. When double salts are heated, they firstly become liquid, and if heating continues the crystal fluid recedes by bubbling. Finally, a powder in the amorph state remains. Double salts are generally defined as alum and they are mostly called by the names of their 3+ loaded cations, rarely by +1 loaded cations. It is for this reason that 3+ loaded cations (Al^{3+}) name is given to the potassium aluminium sulphate crystals so it is named as Aluminium alum.

General Formulas: **Me I. Me II $(SO)_4$. $12H_2O$**
Me I = K^{+}, $NH^{4\,+}$, Cs^{+}, Rb^{+}, Tl^{+} Me II = Al^{3+}, Fe^{3+}, Cr^{3+}, Tl^{3+}

MATERIALS AND METHODS

Prepare water solutions separately of solid Potassium sulphate and solid aluminium sulphate with equal moles. Add $Al_2(SO_4)_3$ solution into K_2SO_4 solution, heat and stir it slowly until it becomes half in volume, continue stirring and heating it. Take the beaker off the burner when the desired volume is obtained and leave it to cool down. Observe the crystal formation in the mixture and after the crystallization is completed, filter the solution. Concentrate it by heating it several times, reducing it into half of its existing volume and cool it. This way it is possible to obtain several fractions of crystal.

HAZARDS

$KAl(SO)_4\ 12H_2O$

WASTE MANAGEMENT

The product so as not to get in contact with a heavy oxide is put in a waste container.

PHOTOS

Fig. 1 Make the solution half its total volume by mixing two salt solutions.

Fig. 2 Leave the solution to cool down, but if cooling down is fast, then the crystallization will not occur.

Fig. 3 It is observed that in the solution left to cool down crystals have formed.

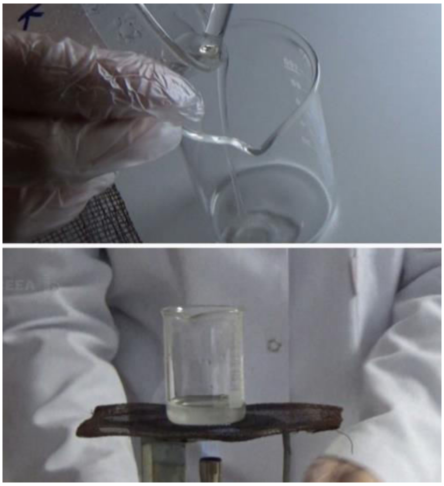

Fig. 4 If the solution is cooled down straight away crystals will not form, but a paste form will be observed.

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OIL SOLUBILITY

OVERVIEW

Oil and grease are generally considered together they are composed of the same element. Although grease has the same oil and other additives, extra materials (components) are added partially to increase its viscosity to resist heat and similar corrosion. Namely, grease oil is oil that shows a structural change between a solid and half-fluid of fluid oil with a thickening substance. The term grease is used for all the organic substances (oil, grease, vax, oil, acid, etc.) that can be extracted from a liquid solution or suspension through hexane or Freon.

If a molecule is formed from different atoms, the relation of each atom to electrons will be different. For this reason, excess of electron in a molecule shows partial negative load; whereas lack of electron in some shows a partial positive load. These molecules are called polar molecules. The molecules, which do not show any polarization as in the electron distribution in polar molecules, are called apolar molecules. Polar solvents dissolve polar solutions, apolar solvents dissolve apolar solutions. The attracting force between molecules in polar compounds is extremely strong. The negatively loaded part of the molecule is attracted by the positively loaded part of the molecule.

Thus, there is a network structure among all the molecules. An apolar molecule cannot dissolve this network structure in a polar molecule. Hexane is a straight chained alkane. It is used in labs as an oil solvent, and as an organic solvent in industry. Its formula is C6H14, so it is heavier than air. It causes burnings by spreading on the ground.

MATERIALS AND METHODS

Put a small amount of grease oil into water, rinse them. Observe that the grease oil does not dissolve in water. Pour the grease oil water mixture into the funnel with hexane two separate phases from. Stirring the funnel, by using the extractıon method, get the grease oil in water to the hexane phase. The phases formed are, the grease oil solution dissolved in water and hexane separately.

HAZARDS

C_6H_{14}

CONCLUSIONS

Oils hold nonpolar hydrocarbon rings in their structures. Since water, as a solvent, is in polar state, apolar structured oil does not dissolve in water hexane is an organic solvent with an apolar structure, so it can dissolve grease oil with an apolar structure. This shows that similar things dissolve similar things.

WASTE MANAGEMENT

Separate water and hexane using a separative funnel and store them in different containers. Discharge water through the tap. Collect the hexane solution in a container from halogen.

PHOTOS

Fig. 1 The grease is taken with a spatula and its degree of dissolubility is observed.

Fig. 2 The water with grease is put into a test balloon with a funnel containing hexane.

Fig. 3 With the mixture consisting of water, hexane and grease are shook to see its dissolubility.

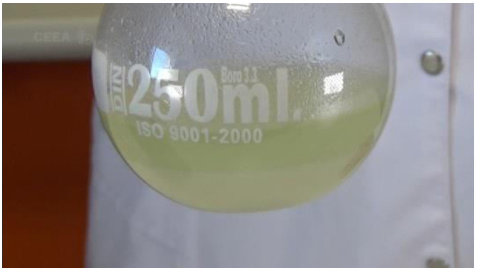

Fig. 4 It is observed that due to the hexanes volume being low, the hexane forms two phases, with hexane on top and that the grease dissolves in the hexane phase.

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SUBLIMATION AND DISSOLUTION OF IODINE

OVERVIEW

In this experiment the sublimation of Iodine and the dissolving in different solvents is shown. Sublimation is a mechanism, in which a solid substance directly moves to gaseous state. The reversed reaction is called condensation. Iodine does not change to a liquid state when it is heated up, it always sublimes. Water in form of snow and ice shows similar reaction. This effect is used in freeze-drying. [1, 2, 3]

In the video this experiment is shown with Tetrachloromethane. For safety reasons, in this description the less toxic Hexane is used.

MATERIALS AND METHODS

Substances: Iodine (I_2), ice cubes, distilled water, Hexane or CCl_4, Sodiumthiosulfate solution $C(Na_2S_2O_3)$ =1 mol/L

Material: flask, watch glass, heating plate, test tubes

Clean thoroughly all devices you use. Take only as much material as you really use to protect the environment and your health.

Sublimation reaction: Add a spoon of Iodine into the flask, which is shown in Fig. 1. Put a watch glass on the neck of the flask and add an ice cube. Heat up until you see purple steam inside the flask. The Iodine sublimates on the bottom and resublimes at the watch glass.

Dissolving: Take a test tube and fill in about 5 mL of Hexane Add a small amount of iodine. The Iodine dissolves in Hexane, but not in water.

HAZARDS

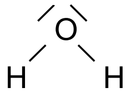
Iodine is harmful to health when contacts with skin. Use protective equipment. Hexane is highly flammable and may be toxic when swallowed Keep away from hot surfaces and fire.

I_2

C_6H_{14}

CONCLUSION

Iodine is a compound of two identical iodine atoms forming biatomic molecules.



Water is a compound of atoms with different electronegativity, oxygen and hydrogen. This causes that the molecule is bipolar, it has a positive area (hydrogen) and a negative area (oxygen).

Hexane or Tetrachloromethane are like Iodine balanced compounds because of their symmetric charge distribution. To dissolve a substance, you need a polar solvent (water) or a non polar solvent (Hexane). In general, the polarity must be similar for solvent and dissolved compound.

WASTE MANAGEMENT

Dissolve the remaining Iodine in Sodium thiosulfate-solution and pour it to the sink, dilute with water.

Hexane is collected in special containers for non-water soluble solvents.

PHOTOS

Fig. 1 Sublimation and desulblimation experiment of iodine

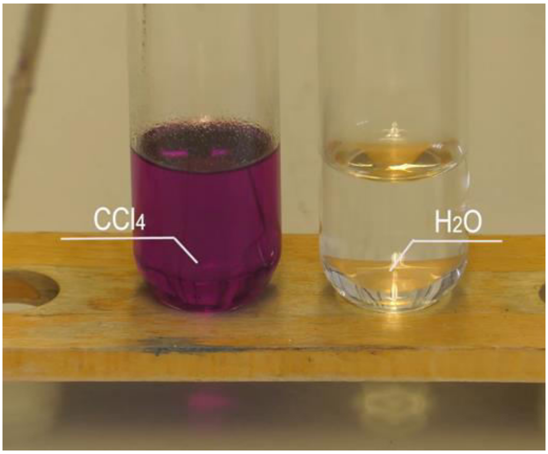


Fig. 2 Dissolving iodine in polar/non polar solvent

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DETECTION OF WATER BY COPPER(II)-SULPHATE

OVERVIEW

Aim of the experiment: detection of water vapour.

Theoretical background: anhydrous copper(ii)sulphate crystals are white, while hydrated copper(II)sulphate is bright-blue.

Like for most metals heavier than calcium, the colouring of ions can be ascribed to the annulation of the degeneration of the 3d - orbitals upon the approach of ligand molecules. In the case of Cu^{2+} the partly negatively charged oxygen atoms of the water molecules cause the 5 3d - orbitals to split into two energy levels. Electrons can resonate between the two levels by absorbing energy quanta with a frequency corresponding to orange light, leaving the complementary colour for us to observe.

The reaction is specific which makes white copper(ii)sulphate an agent for the detection of water.

MATERIALS AND METHODS

Requirements:

White and blue copper(II)sulphate, Erlenmeyer flask with stopper, water, glass tube, heater, water jet pump.

Procedure:

Use in practice: if one is interested in the presence of water in a gas flow, just lead the flow, or a part of it, through a simple device containing white copper(II)sulphate. If it remains white, one can be sure there is no detectable amount of water present in the gas flow. If it turns blue, there is….. It should be mentioned that the principle can also be used to detect minor quantities of water in non-aqueous liquids.

HAZARDS

$CuSO_4$

CHEMICAL REACTION EQUATION

$CuSO_{4(s)}$ + 5 $H_2O_{(l)}$ ---> $CuSO_4 \cdot 5\,H_2O_{(s)}$

CONCLUSIONS

Many of us know copper(II)sulphate pentahydrate as large, bright blue and clear crystals; see photo 1. Anhydrous copper(II)sulphate is much more powder like; the reaction with water vapour occurs only at the surface area so a vaguely-blue powder results; see photos 1 and 2.

WASTE MANAGEMENT

The reaction is reversible,the copper(ii)sulphate can be used repeatedly.

PHOTOS

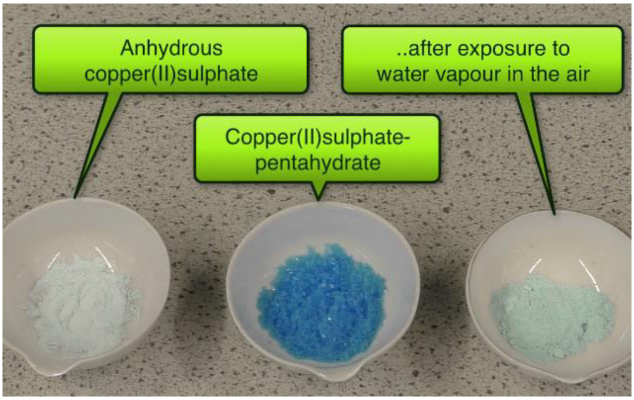

Fig. 1 Left: anhydrous copper(II)sulphate; middle: Copper(II)sulphate pentahydrate; right: anhydrous copper(II)sulphate powder after exposure to water vapour.

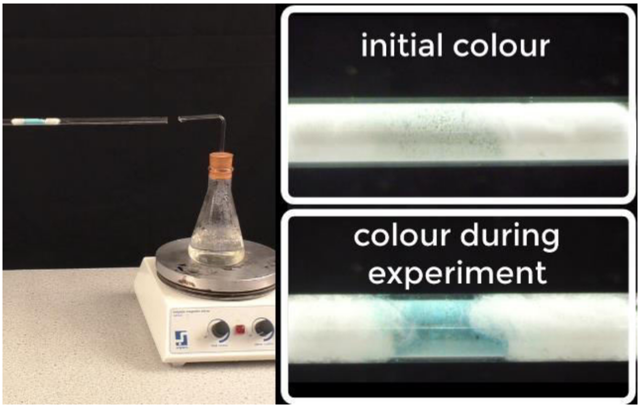

Fig. 2 Tube filled with anhydrous copper(II)sulphate, before and after exposure to water vapour.

REFERENCES

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AN EMPTY JAR... HOW TO MAKE THE SALT?

OVERVIEW

In the fourth year of the course, secondary students have to study various aspects of salts, one of them being solubility in water. They have to get familiar with the fact that precipitation reactions can be used for qualitative analyses, but also for removal of undesired ions or synthesis of other salts. This experiment deals with the latter case.

MATERIALS AND METHODS

Requirements:

Copper(II) sulphate, copper(II)chloride, calcium chloride, beakers, balance, filter, water jet pump

Procedure:

Suppose, for some reason we need a certain salt. Open the chemicals cupboard, which of course fulfils all safety demands, take out the jar bearing the name of the salt you are looking for and ...... surprise, surprise ...... it appears to be empty!!!! Apparently, somebody used the last bit of the material and forgot to order new stock. Now we have a problem ..... or perhaps not?

Thanks to different solubility behaviour of salts, we can produce salts ourselves by smartly making use of the solubility charts. The basic thought is to combine two salts to yield a soluble pair of ions and an insoluble one. When mixing, the insoluble pair will precipitate, allowing us to simply separate the salt mixture by filtration. The soluble pair remains in the filtrate and can be solidified by evaporating the water.

Let's illustrate this idea with some examples now.

1. We need to make calcium sulphate. Calcium sulphate is hardly soluble. When mixing a soluble calcium salt with a soluble sulphate, calcium sulphate will precipitate. (Reaction 1).

All sodium, potassium and ammonium salts are soluble as well as nitrates and acetates. So we are relatively free to choose our initial salts.

The suspension can be separated by filtration. Note that contamination will occur to some extent because of adhering filtrate. The filtrate contains the soluble pair: e.g. sodium and nitrate.

2. Now we need copper(II)chloride. In this case we are dealing with a soluble salt which means that the second pair of ions has to be insoluble. In contrast to the first case sodium, potassium and ammonium salts as well as nitrates and acetates are excluded as initial salts.

Let copper(ii) be provided by copper(ii)sulphate and chlorine by calcium chloride. When mixing, reaction (1) proceeds. Again, the precipitate can be removed by filtration. If stoichiometric quantities are used the filtrate is a solution of copper(II)chloride. It is important to realize that the solution is saturated with calcium sulphate.

HAZARDS

$CuSO_4$

$CuCl_2$

$CaCl_2$

CHEMICAL REACTION EQUATION

$$\mathbf{Ca^{2+}_{(aq)} + SO_4^{2-}{}_{(aq)} \quad {-}{-}{-}{>} \quad CaSO_{4\,(s)} \qquad (1)}$$

CONCLUSIONS

Salts can be prepared by making use of differences in solubility behaviour. Two initial salts have to be chosen in such a way that mixing them results in a soluble and an insoluble pair of ions. If the insoluble pair is the required salt, it can be separated by filtration. When the soluble pair is the required one, the filtrate has to be evaporated.

WASTE MANAGEMENT

The chemicals involved do not harm the environment when used in small quantities. Therefore, they can be treated as waste water.

PHOTOS

Fig. 1 Surprise….empty jar!

Fig. 2 Don´t give up….

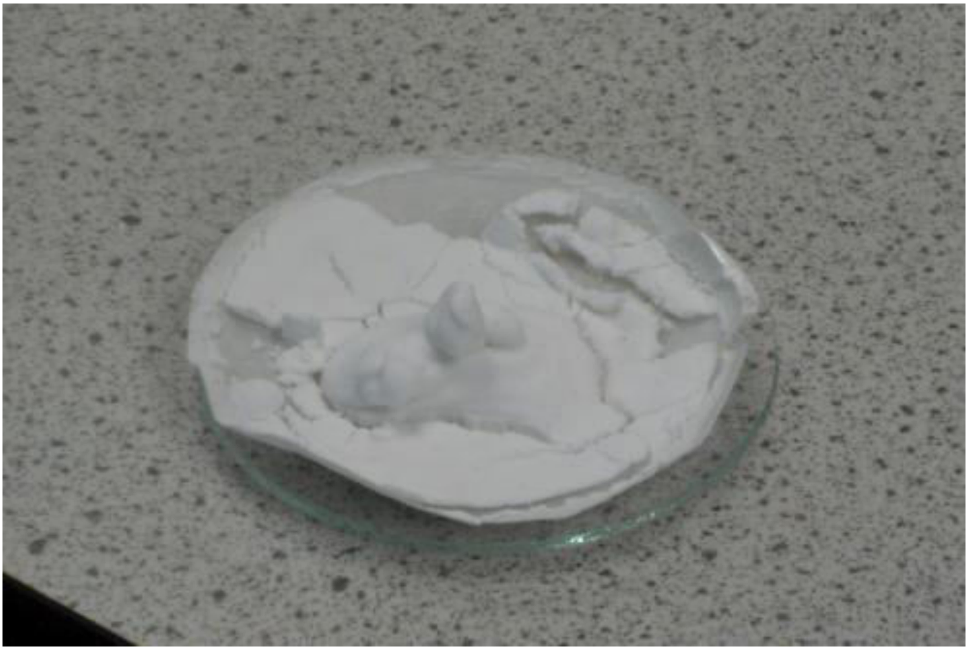

Fig. 3 Here is our salt!

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2. wikipedia.org/wiki.copper(II)chloride;
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RELATION BETWEEN $[H_3O^+]$ and pH

OVERVIEW

Neutral water has a pH of 7. When we add a small amount of HCl solution the pH value decreases steeply; the decrease becoming more gradual as addition of HCl proceeds. This phenomenon can be explained by the fact that the pH is a logarithmic function of the concentration of hydronium ions

($pH = -10\log [H_3O^+]$). As a consequence, a decrease of the pH by one, means that the solution becomes 10 times more acidic. The relationship between the amount of HCl added and the corresponding value of the pH will be demonstrated by adding drops of a 0.1 M solution of HCl to water while measuring the pH with an electrode.

Water molecules are strongly bound by hydrogen bonds. These bonds are responsible for dissociation of a small part of the water molecules according to the following equilibrium:

$$H_2O(l) \leftrightarrow H_3O^+(aq) + OH^-(aq) \qquad (1)$$

with the value of the equilibrium constant:

$$K_w = [H_3O^+] \times [OH^-] \qquad (2)$$

being 10-14 at room temperature. In pure water, $[H3O^+] = [OH^-]$ so pH = 7. Upon addition of HCl solution, $[H3O^+]$ increases and because the equilibrium condition has to remain fulfilled, [OH−] has to decrease. So, the total amount of $[H3O^+]$ consists of two contributions: one due to the addition of HCl solution, and one due to the dissociation of water molecules.

Realizing that the mixture is electrically neutral throughout the experiment, we have two equations to describe the system: (i) the equilibrium condition (2), and (ii) the amount of positive charge equals the amount of negative charge according to:

$$[H_3O^+] = [OH^-] + [Cl^-] \qquad (3)$$

Eqs. (2) and (3) can be combined to yield a quadratic equation in $[H3O^{+]}$ with the following solution:

$$[\mathrm{H_3O^+}] = \frac{[\mathrm{Cl^-}] + \sqrt{[\mathrm{Cl^-}]^2 + 4 \cdot \mathrm{K_w}}}{2} \qquad (4)$$

$[Cl^-]$ follows directly from the amount of acid added. From eq. 4 it can be seen that already upon addition of 10-6 mole HCl/liter, the contribution of the dissociation of water molecules becomes negligible.

MATERIALS AND METHODS

Requirements: Beaker, stirring device, pH electrode and measuring device, biurette, hydrochloric acid, desalted water.

Procedure: A beaker containing 500 mL of desalted water is put onto an electrical stirring device. The pH is measured with a pH electrode. At the start, the pH can be somewhat below 7 due to dissolution of carbon dioxide. Therefore, the water has to be boiled first. Now, the HCl solution is added dropwise; after each droplet, the pH has to stabilize. The volume of a drop can be determined separately by measuring the volume of a known number of droplets. By proceeding this way, a curve is obtained representing the relation between pH and the volume of HCl-solution added.

HAZARDS

HCl

CHEMICAL REACTION EQUATION

$H_2O(l) \leftrightarrow H_3O^+(aq) + OH^-(aq)$

WASTE MANAGEMENT

Diluted hydrochloric acid can be treated as waste water.

PHOTOS

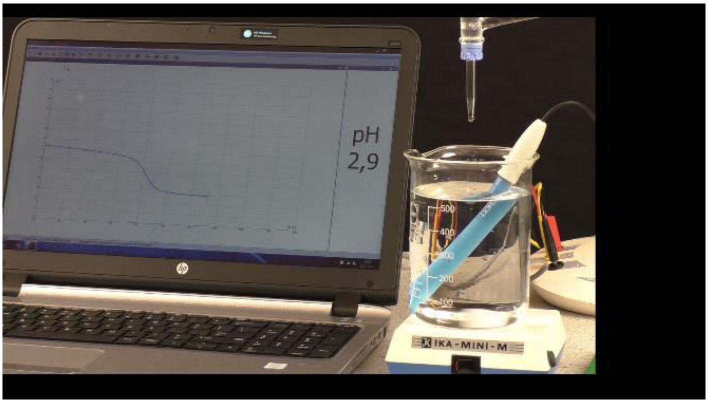

Fig. 1 pH-curve. Obviously the water was buffered with CO_2 / HCO_3^- .

REFERENCES

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RELATION BETWEEN [OH$^-$] AND pH

OVERVIEW

Neutral water has a pH of 7. When we add a small amount of NaOH solution the pH value increases steeply; the increase becoming more gradual as addition of NaOH proceeds. This phenomenon can be explained by the fact that the pH is a logarithmic function of the concentration of hydronium ions

(pH = – $^{10}\log [H_3O^+]$). As a consequence, an increase of the pH by one, means that the solution becomes 10 times more basic. The relationship between the amount of NaOH added and the corresponding value of the pH will be demonstrated by adding drops of a 0.1 M solution of NaOH to water while measuring the pH with an electrode.

Water molecules are strongly bound by hydrogen bonds. These bonds are responsible for dissociation of a small part of the water molecules according to the following equilibrium:

$$2\ H_2O_{(l)} \leftrightarrow H_3O^+_{(aq)} + OH^-_{(aq)} \qquad (1)$$

with the value of the equilibrium constant:

$$Kw = [H_3O^+] \times [OH^-] \qquad (2)$$

being 10-14 at room temperature. In pure water, $[H3O^+] = [OH^-]$ so pH = 7. Upon addition of NaOH solution, $[OH^-]$ increases and because the equilibrium condition has to remain fulfilled $[H_3O^+]$ has to decrease.

Realizing that the mixture is electrically neutral throughout the experiment, we have two equations to describe the system: (i) the equilibrium condition (2), and (ii) the amount of positive charge equals the amount of negative charge according to:

$$[H_3O^+] + [Na^+] = [OH^-] \qquad (3)$$

Eqs. (2) and (3) can be combined to yield a quadratic equation in $[H_3O^+]$ with the following solution:

$$[H_3O^+] = \frac{\sqrt{[Na^+]^2 + 4 \cdot K_w} - [Na^+]}{2} \qquad (4)$$

$[Na^+]$ follows directly from the amount of NaOH added. Eq. 4 shows that $[H_3O+]$ is fully determined by dissociation of water molecules, the number of them becoming smaller as the amount of NaOH-solution increases.

MATERIALS AND METHODS

Requirements: Beaker, stirring device, pH electrode and measuring device, biurette, NaOH, desalted water.

Procedure: A beaker containing 500 mL of desalted water is put onto an electrical stirring device. The pH is measured with a pH electrode. At the start, the pH can be somewhat below 7 due to dissolution of carbon dioxide. Therefore, the water has to be boiled first. Now, the NaOH solution is added dropwise; after each droplet, the pH has to stabilize. The volume of a drop can be determined separately by measuring the volume of a known number of droplets. By proceeding this way, a curve is obtained representing the relation between pH and the volume of NaOH-solution added.

HAZARDS

NaOH

CHEMICAL REACTION EQUATION

$$\mathbf{2\,H_2O_{(l)} \leftrightarrow H_3O^+_{(aq)} + OH^-_{(aq)}}$$

WASTE MANAGEMENT

Diluted NaOH solution can be treated as waste water.

PHOTOS

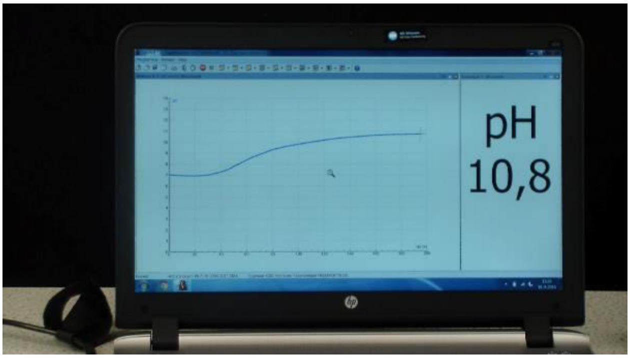

Fig. 1 pH curve in alkaline solution.

REFERENCES

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SEPARATION OF MIXTURES COCA-COLA (COKE)®

OVERVIEW

The goal of this experiment is to separate the beverage Cola (Coke)® into its components: sugar, pigment and the remaining liquid residues. Additionally, qualitative tests for aldehydes, caramel and phosphorous, which are commonly present in carbonated drinks, may be performed. [1, 2, 3]

MATERIALS AND METHODS

Chemicals: Fehling's reagent I (3.5 g Copper(II) sulphate in 50.0 mL H_2O) and II (17.5 g Sodium-potassium-tartrate and 6.0 g Sodium hydroxide in 50.0 mL H_2O), Nitric acid (2.0 M), Ammonium molybdate (or Ammonium molybdate-tetrahydrate), activated carbon filter powder

Materials: Two round-bottomed flasks with a volume of 100.0 ml, a heat-resistant cover (fitting to the round-bottomed flasks), heating apparatus, analytical funnel, thermometer (20°C - 120°C), Liebig condenser (fitting to the round-bottomed-flask), 4 test tubes (heat resistant), filter paper

Procedure: In the first step the pigment is removed by adsorption through an active carbon distillation. Put the filter paper into the funnel and fill with carbon(powder). Pour the Cola (Coke®) sample into the apparatus, preferably with a funnel (Fig. 1). [4]

In the next step the glucose/ sucrose portion is separated by a distillation process. A part of the filtrate is transferred to the distillation apparatus (Fig. 2), the round-bottomed flask should be attached and secured with a safety clip. Start the distillation process by turning on the heating for approximately 15 minutes or until the liquid components have evaporated. The distillate consists of water with some low boiling substances, the remaining (solid) component is the sucrose/glucose portion.

In a third step, qualitative tests of caramel, aldehydes and phosphorus are performed. For indication of caramel, the remaining filtrate is mixed with Fehling-reagent and heated. Caramel is present if the analyst loses its brown colour after the filtration. Phosphorus is detected by mixing the filtrate with ammonium-molybdate and adding nitric acid carefully until the solution turns yellow. Aldehydes are detected by a brick-red precipitate after the addition of the Fehling-reagent.

HAZARDS

Ammonium molybdate is dangerous if swallowed, by skin and eye contact and if inhaled. Do not inhale it in any form(!), wear safety glasses [if the chemical gets into contact with the eye, rinse the affected eye with clear water], the chemical has to be kept in a closed-off container. Diluted nitric acid is dangerous by skin and eye contact, contact with fire and corrosive by contact with metals. Wear protection gloves and safety glasses. Fehling reagent (I) is dangerous if swallowed or inhaled. Wear safety glasses, handle the chemicals in a vent. Fehling (II) reagent is dangerous by skin contact or if swallowed. Wear protection gloves and safety glasses.

$(NH_4)_6Mo_7O_{24} \cdot 4\ H_2O$

HNO_3 2.0 M:

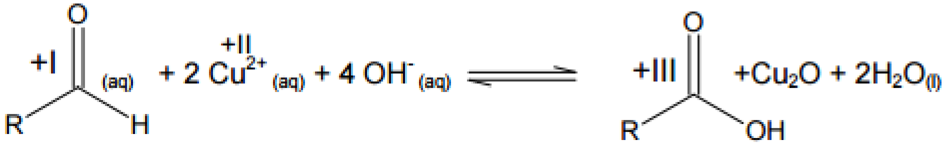
Fehling reagent I

Fehling reagent II

CHEMICAL REACTION EQUATION

Fehling-test (sucrose):

$RCOH + 2\ Cu^{2+} + 4\ OH^- \rightarrow RCOOH + Cu_2O + 2\ H_2O$



CONCLUSION

Caramel is present if the analyst loses its brown colour after the filtration. Caramel is commonly used as a colouring agent in carbonated beverages. Phosphorus is present if the pH-value of the filtrate is <3 due to phosphoric acid (less than carbonic acid). Phosphoric acid is commonly used as an acidifier in the production of carbonated beverages, like cola. Aldehydes, which are indicated by the Fehling test, are commonly used as flavouring agents.

WASTE MANAGEMENT

Removal of the pigment: The residue can be flushed into the drain, but has to be diluted with water.

Removal of the water portion: The residue has to be scraped out of the round-bottomed flask and flushed out with hydrochloric acid (0.1M). The remaining solutions (Fehling reagent) contain heavy metal residue. Disposal in the container for heavy metal waste is necessary.

PHOTOS

Fig. 1 Filtration procedure

Fig. 2 Distillation procedure

Fig. 3 The sample (left), liquid portion (middle), solid portion (right), the sucrose in the sample caramelized during the distillation, hence the brown colour.

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REDUCTION AND SUBSEQUENT OXIDATION OF METHYLENE BLUE (BLUE BOTTLE II)

OVERVIEW

The blue bottle experiment is based on a redox reaction. An aqueous solution containing glucose, sodium hydroxide, methylene blue is shaken in a closed bottle. The colour turns from colourless to blue and then decolorizes again after a while. With further shaking, the cycle can be repeated several times. In this experiment the change of the redox potential of the solution during the decolourization and the restoration of the solution is demonstrated.[1,2]

MATERIALS AND METHODS

Take 10 g of Sodium hydroxide, 10 g of Glucose and 250mL deionized water. Dissolve in a 250mL round-bottomed flask with a lab stirrer in order to get a homogenous solution. Prepare a Methylene blue (MB) solution (w=1%) in Ethanol). Add 1 mL of the MB-solution with a lab pipette. The effect of decolouration can be watched immediately. To measure the redox potential of the solution use a pH/mV-metering system with platinum electrode (Fig.1, 2). You can hold the electrode with your hand or optional fix it on a stand. The first 4-5 cm of the electrode need to be in touch with the solution. Activating the stirrer leads to an increased amount of dissolved oxygen in the solution and therefore the redox potential changes to a more positive level. When the solution is allowed to rest, the Glucose reduction of the dye gains the upper hand on the reaction. This is observed with the changing level of the redox potential and the disappearance of the colour. This redox cycle can be repeated many times.

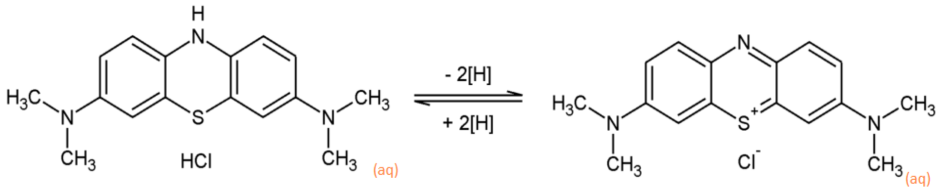
HAZARDS

Like other corrosive acids and alkalines, drops of sodium hydroxide solutions can readily decompose proteins and lipids in living tissues via amide hydrolysis and ester hydrolysis, which consequently cause chemical burns and may induce permanent blindness upon contact with eyes. Protective equipment, like rubber gloves, safety clothing and eye protection, should always be used when handling this chemical. [3] The liquid and vapour methylene blue is flammable, harmful if inhaled and causes damage to organs. [4]

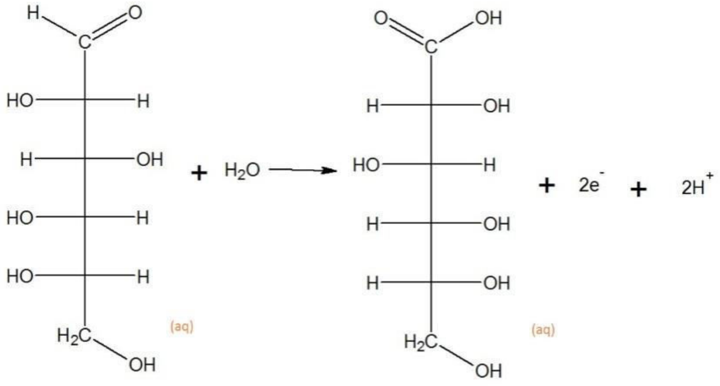
NaOH

$C_{16}H_{18}ClN_3S$
(Methylene blue)

CHEMICAL REACTION EQUATION



Colourless blue



Glucose Gluconic acid

CONCLUSION

The aqueous solution in the classical reaction contains glucose, sodium hydroxide and methylene blue. In the first step the enolate of glucose is formed. The next step is a redox reaction of the enolate with methylene blue. The glucose is oxidized to gluconic acid which, in alkaline solution is in the sodium gluconate form. Methylene blue is reduced to colourless leucomethylene blue. The dissolved Glucose, a strong reducing agent, is driving the redox potential of the solution to a much lower negative value level thus the dyestuff gets reduced to a colourless form.

WASTE MANAGEMENT

The remaining chemicals may be diluted with water and poured into the sink.

PHOTOS

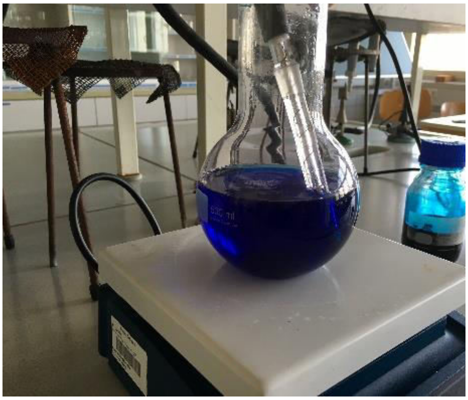


Fig. 1 oxidized dye = blue

Fig.2 Electrochemical potential of -176.9mV

Fig. 3 reduced dye = clear

Fig. 4 potential of -352.8mV

REFERENCES

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3. http://en.wikipedia.org/wiki/Sodium_hydroxide#Safety
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SECRET INKS

OVERVIEW

Secret inks are substances which can be used to write messages which are invisible until the substance is heated. They are written on a sheet of paper using different substances like onion juice, lemon juice or milk to write like a cotton swab or a toothpick. There are some more methods to make the secret ink visible. Some messages are developed by spraying or wiping the paper with a second chemical, for example phenolphthalein, further processed by ammonia fumes or sodium carbonate. [1]

MATERIALS AND METHOD

Take a toothpick and soak it in onion juice. You can buy the onion juice in a supermarket or you can squeeze an onion. Then write on the paper. Put the sheet of paper in a drying cabinet or use a blow-dryer. After about 15 minutes you will be able to see what you wrote.

CONCLUSION

After you have written with the onion juice on the paper, the carbon-based compounds are absorbed into the paper's fibres. When the paper is heated, the heat causes some of the chemical bonds to break down, setting free the carbon. Once the carbon comes in contact with the air, it goes through a process called oxidation.

WASTE MANAGEMENT

The remaining chemicals can be poured in the sink with water.

PHOTOS

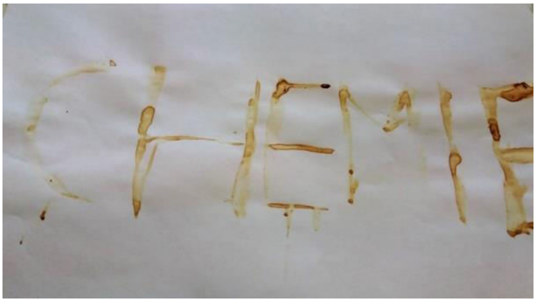

Fig. 1 The result of the "Secret Ink"

REFERENCES

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IRON (III) COMPLEXES AS SECRET INKS

OVERVIEW

Iron (Fe) is a very important metal for the living cells. It is present in many biochemical processes showing a wide range of properties. One of the most important processes it manages, in living organisms, is the gases transportation in warm blooded animals, participating to the oxygen and carbon dioxide transportation to/from the cells. This feature is possible due to the Fe^{2+} capacity to form a special complex with various ligands (heme, oxygen, histidine) called hemoglobin (Hbg) whose stability depends on the partial pressures of both O_2 and CO_2 [1]. In the same way, but showing other roles, Fe^{3+} can form various biochemically important complexes.

The present series of experiments studies the ability of the ferric ion to bind other ligands generating various complexes. It is a series of three experiments having in common the presence of the same metallic ion Fe(III).

EXPERIMENT 1 The reaction of iron (III) chloride $FeCl_3$ with potassium thiocyanate KSCN is very sensitive and gives a specific color ensuring its applicability in analytical chemistry as an iron (III) recognition reaction. Due to its blood like color it was used in the film industry to simulate wounds and blood spills. [2]

EXPERIMENT 2 Iron (III) chloride $FeCl_3$ reacting with potassium hexacyanoferate (II) $K_4[Fe(CN)_6] * 3H_2O$ gives another historically famous complex (the Prusian Blue or Berlin Blue known by the chemists as iron (III) hexacyanoferate (II)). This substance is the first synthetic modern pigment. Despite its cyanide content the complex itself is considered safe for humans and is used as an antidote for Cs^+, Tl^+ and other heavy metal poisoning or even anticaking agent for both road and table salt [3]. Being quite safe for humans who can eat up to 10g/day without any harm, and due to its low solubility in water, this complex is a byproduct of the iron removal process from vines. As it is very sensitive to small amounts of iron Fe(III) in solution it is used as identifying reagent for iron in laboratory. Also the Lassaigne Reaction which is a qualitative elemental analysis of elements in organic compounds (N, C, X, S) makes use of this complex when identifying carbon and nitrogen in a form of CN^- ion.

EXPERIMENT 3 is also of historical importance as iron containing solutions were reacted with gallic acid from tannins giving a dark (almost black) dye used as ink [4] not only from 12th to 19th century but going back to the times when the Dead Sea Scrolls were written

[5]. It is a well-known reaction used in the late centuries in fairs when people, with iron impurities on their hands, were stained with a black dye after damping their hands with towels (secretly powdered previously with tannin extracted from gallnuts).

The reason why these three experiments were chosen lies in their property to transform three different colorless or slightly colored compounds in a colored complex when reacting with iron. This property was used, in the last centuries, to transmit "invisible" messages to friends/intimates/subordinates avoiding the interception of the letter or its content by the enemies. Performing these experiments, with the students, we don't just speak about complexes, chemical equilibrium or precipitation reactions with analytical applications but we make our students part in a "Chemistry related conspiracy" and we introduce elements of history of science.

MATERIALS AND METHODS

Prepare four different aqueous solutions:

SOLUTION 1 as reagent for EXPERIMENT 1; 2% potassium thiocyanate (might be sodium or ammonium as well) (KSCN).

SOLUTION 2 as reagent for EXPERIMENT 2; 3% potassium hexacyanoferate (II) $K_4[Fe(CN)_6] * 3H_2O$.

SOLUTION 3 is a reagent for EXPERIMENT 3; 2% aqueous gallic acid.

As you will notice all three solutions are either colorless or light yellowish. The three experiments consist in writing with these three solutions, using a quill or nib, on a large piece of paper or a cardboard. Allow it to dry.

In the mean time prepare **SOLUTION 4**, which is a 5% aqueous solution of ferric chloride ($FeCl_3$). This solution is pale yellowish. Place this solution in a sprayer. (You can use even a discarded and thoroughly washed perfume bottle).

Spray, over the previously written cardboard, the SOLUTION 4 and record your observations.

HAZARDS

Ferric chloride is irritant for skin and harmful if swallowed. Always use protecting clothing/eye protection. Potassium thiocyanate is harmful if swallowed or inhaled, if in contact with skin. It can release very toxic gases in contact with acids (e.g. hydrogen cyanide and sulfur dioxide) [6].

Surprisingly some legislation are very strict concerning the potassium hexacyanoferate (II) and potassium hexacyanoferate (III) but, according to the European legislation and rues, no GHS symbols are allocated only GHS statements: harmful for aquatic environment and avoid release in the environment [6].

Gallic acid can cause skin, respiratory and eye irritation.

$FeCl_3$	🔥 ⚠	**KSCN**	⚠
$K_4[Fe(CN)_6]$ * $3H_2O$	Harmful for aquatic life	**Gallic acid**	⚠

CHEMICAL REACTION EQUATION

for the experiment nr. 1 for the iron thiocyanate complex

$$FeCl_{3\,(aq)} + KSCN_{(aq)} \longrightarrow Fe(SCN)Cl_{2\,(aq)} + KCl_{(aq)}$$

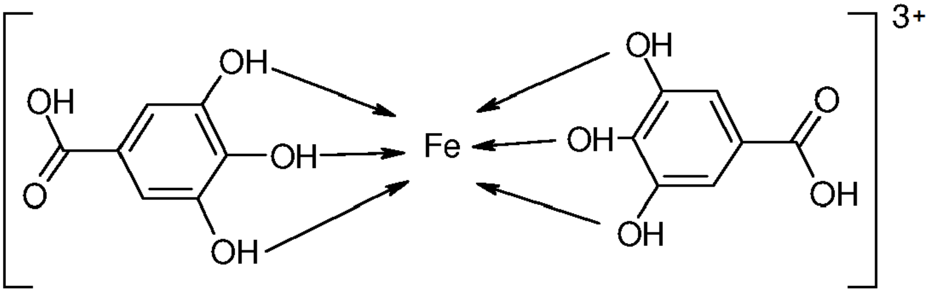
$$FeCl_{3\,(aq)} + 3\,KSCN_{(aq)} \longrightarrow Fe(SCN)_{3\,(aq)} + 3\,KCl_{(aq)}$$

$$FeCl_{3\,(aq)} + 6\,KSCN_{(aq)} \longrightarrow K_3[Fe(SCN)_6]_{(s)} + 3\,KCl_{(aq)}$$

for the experiment nr. 2 formation of iron (III) hexacyanoferate (II))

$$4\,FeCl_{3\,(aq)} + 3K_4[Fe(CN)_6]_{(aq)} \longrightarrow Fe_4[Fe(CN)_6]_{3\,(s)} + 12\,KCl_{(aq)}$$

for the iron (III) complex with gallic acid



CONCLUSIONS

Observe that after drying the writing is hidden

EXPERIMENT 1 The potassium thiocyanate is a white solid reagent generating colorless solution when dissolved in water while Fe(III) ions are hydrated in aqueous solutions forming the yellowish hexahydrate ion $Fe(H_2O)_6^{3+}$. When mixing the two solutions the red iron thiocyanate complex ion is formed $[Fe(SCN)(H_2O)_5]^{2+}$. If more concentrated solutions are used an abundant dark red precipitate is formed due to further reactions of the iron ions, forming iron tris-thiocyaniate

$Fe(SCN)_3$ respectively the potassium hexathiocianato iron (III) complex. Thus, spraying the iron solution over the colorless thiocyanate, a red color is observed due to the iron thiocyanate complex formation which is similar to the developing techniques when making photos.

EXPERIMENT 2 reveals the blue color of the Berlin Blue or iron(III) hexacyanoferate(II) complex when sprayed with $FeCl_3$ while **EXPERIMENT 3** shows the black color of the complex of Fe^{3+} with galic acid.

WASTE MANAGEMENT

The solutions can be poured down the drain after considerable dilution and the iron chloride solution can be collected in containers for less toxic metal waste. The potassium ferrocyanide solutions have to be first reacted with Fe(II) or Fe(III) solutions and then the blue complex can be safely discarded in the sink.

PHOTOS

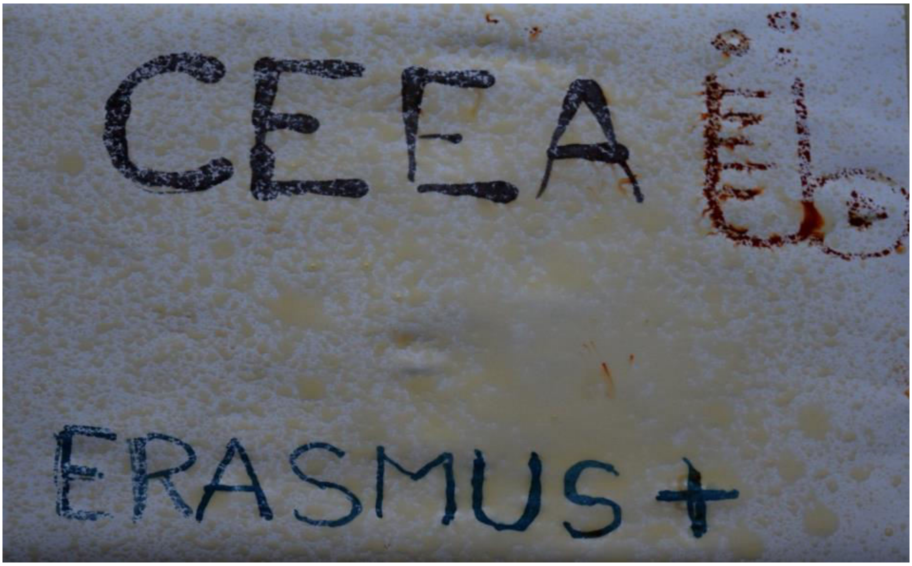


Fig. 1 Tree different colors revealed with the same reagent - the iron(III) chloride (*CEEA* is EXP 3, *Erasmus+* is EXP. 2 and the *logo* is EXP. 1)

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5. https://en.wikipedia.org/wiki/Gallic_acid

BLUE BOTTLE EXPERIMENT

OVERVIEW

The blue bottle experiment is based on a redox reaction. An aqueous solution containing glucose, sodium hydroxide and methylene blue is shaken in a closed bottle in the presence of oxygen from the air. It turns from colourless to blue and then decolorizes again after a while. With further shaking, the cycle can be repeated several times. The reaction will work with other reducing sugars besides glucose and also with other reducing dyes. Methylene blue is a dyestuff frequently used in chemistry, medicine and dyeing technology. [1,2,3]

MATERIALS AND METHODS

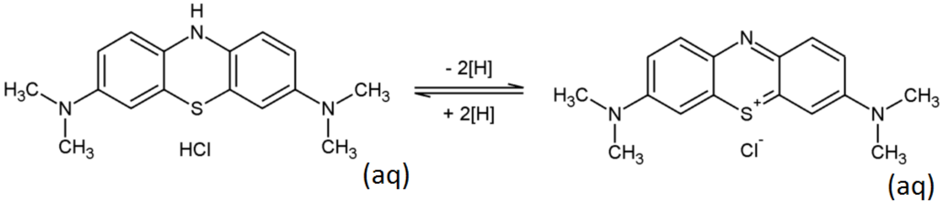
Take 10 g of Sodium hydroxide, 10 g of Glucose and 250 mL of deionized water. Dissolve in a 250mL round-bottomed flask with a lab stirrer in order to get a homogenous solution. Prepare a methylene blue (MB) solution (1% w/w in Ethanol). Add 1 mL of the MB solution with a lab pipette. The effect of decolouration can be watched immediately. Shaking the flask leads to an increased amount of dissolved oxygen in the flask, and the blue colour of the solution is appearing again. When the solution is allowed to rest, the colour disappears again. This redox cycle can be repeated many times.

HAZARDS

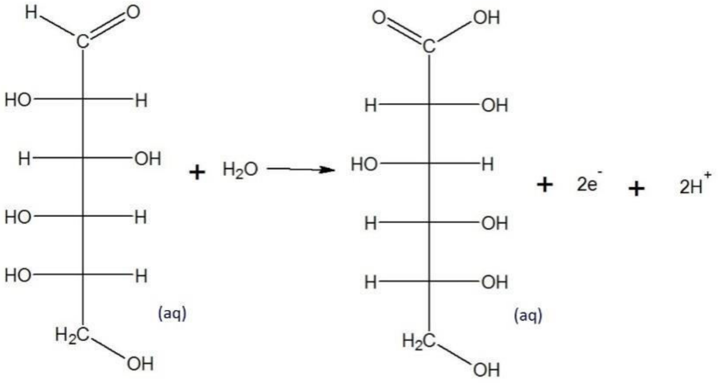
Sodium hydroxide causes severe burns, therefore the use of protective gloves and glasses is mandatory. Methylene blue is harmful if swallowed.

NaOH		**$C_6H_{12}O_6$**	none
$C_{16}H_{18}ClN_3S$ **(Methylene blue)**			

CHEMICAL REACTION EQUATION



colourless blue



Glucose Gluconic acid

CONCLUSIONS

The effect of decolouration can be watched immediately, because glucose is reducing the indicator to a colourless form. The aqueous solution in the classical reaction contains glucose, sodium hydroxide and methylene blue. In the first step the enolate of glucose is formed. The next step is a redox reaction of the enolate with methylene blue. The glucose is oxidized to gluconic acid which, in alkaline solution is in the sodium gluconate form. Methylene blue is reduced to colourless leucomethylene blue. Shaking the flask leads to an increased amount of dissolved oxygen in the deionized water and therefore the indicator is re-oxidized and the blue colour of the solution is appearing again.

WASTE MANAGEMENT

The remaining chemicals may be poured into the sink, diluted with water.

PHOTOS

Fig. 1 Discoloration of Methylene blue

REFERENCES

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“TRAFFIC LIGHT”: RED – YELLOW - GREEN

OVERVIEW

The goal of this experiment is to induce a finely tuned reversible redox reaction with glucose under basic conditions, which gets visible in three different colours. An alkaline glucose solution is prepared and coloured with indigo carmine dye. After a short waiting time the colour changes to yellow. Different intensities of shaking cause colour changes to red and green. [1,2,3]

MATERIAL AND METHODS

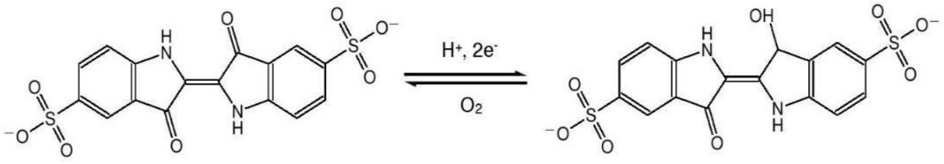
Dissolve 2 g of glucose in 50 mL of hot water in a small beaker. Add 10 mL Sodium Hydroxide solution [2 M] to the glucose solution. Fill a 250 mL Erlenmeyer flask with 50 mL of hot water and add a small amount (as much as the tip of a spatula) of indigo carmine. The colour changes to a deep blue. Now pour the alkaline solution of glucose into the flask and observe the change of colour, wait for a while for further colour changes. This reaction can be reversed simply by shaking the solution.

HAZARDS

Sodium hydroxide is very corrosive and can cause severe burns. It is highly recommended to wear safety googles and gloves.

NaOH

CHEMICAL REACTION EQUATION



CONCLUSIONS

Several changes of colour are observed. The starting colour is blue, upon addition of alkaline glucose solution it changes to green, after a while without shaking it changes to red and yellow. When shaking vigorously, red (yellow) changes to green. The colour change observed is a result of the reversible redox-reaction of the dye present in the solution. By shaking, the dye is oxidized by the atmospheric oxygen. Oxygen dissolves in the solution and oxidizes the indigo carmine. The colour of the solution depends on the oxidation grade of the indigo carmine. In the reduced form it is yellow, partly oxidized it is red, in the fully oxidized form it is green.

WASTE MANAGEMENT

Dispose the contents of the flask in the sink and rinse with sufficient amount of water.

PHOTOS

Fig. 1-4 Changing colours resulting from different states of oxidation

REFERENCES

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2. https://projects.ncsu.edu/project/chemistrydemos/Kinetics/Blue%20Bottle.pdf
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PHARAOH SNAKE

OVERVIEW

Emser® Pastille tablets contain natural salts as a mineral source, including the two key ingredients for the experiment, sodium and potassium bicarbonate. Admixed to the whole is the third key ingredient, sugar. The two bicarbonates perish under heat; a chemical reaction starts, forming carbon dioxide, which now expands the remains of burning sugar. The history behind this experiment according to the bible is as follows: Aaron threw his rod at Pharaoh's feet and his rod became a serpent. Then the pharaoh took his chemists, who also had to make snakes. But Aaron's snake was better than that from the magicians. Several other macabre miracles followed until Pharaoh let the people of Israel go. [1,2,3,4,5]

MATERIALS AND METHODS

Add some sand to a porcelain bowl. Then immerse 3-4 pastille tablets with a pipette in ethanol and stick them into the sand close to each other. Now add some more ethanol on the tablets again. Put away the flaks with ethanol to prevent accidents. Put the tablets on fire with a lighter and watch the "black snake" called "Pharaoh Snake" growing out of the mixture.

HAZARDS

Ethanol is highly flammable and must be kept away from an open flame.

C_2H_5OH

CHEMICAL REACTION EQUATION

$$2\ NaHCO_{3\,(s)} \longrightarrow Na_2CO_{3\,(s)} + H_2O_{(g)} + CO_{2\,(g)}$$

$$C_6H_{12}O_{6\,(s)} \longrightarrow 6\ C_{(s)} + 6\ H_2O_{(g)}$$

CONCLUSIONS

When the ethanol is ignited, sodium bicarbonate decays, which is present in the pastille, whereby, inter alia, carbon dioxide is released. Furthermore, the sugar contained in the pastilles is heated. However, this sugar can only burn to a small extent on account of the carbon dioxide formed, so that a large portion of the sugar is charred due to the oxygen deficiency. The carbon produced during the case of charring is inflated by the resulting carbon dioxide, whereby the so-called "Pharaoh Snake" is generated.

WASTE MANAGEMENT

The remaining chemicals may be poured in the sink.

PHOTO

Fig. 1 Reaction of Pharaoh snake

REFERENCES

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5. WEKA Fachverlag für Behörden und Institutionen Chem

EATABLE CANDLES

OVERVIEW

This experiment describes how to produce a candle with materials which are eatable. For combustion of the candle three components are necessary: fuel, ignition energy and oxygen. Fuel is taken from the oil of the almond or from olive oil, the ignition energy is delivered by a lighter or by matches. Oxygen is drawn from the air. [1, 2]

MATERIALS AND METHODS

Materials: Orange, Apple/Potato, Almond (unshelled), Olive oil

Devices: Knife and fire source (lighter/matches)

Apple candle:

Cut the apple or potato into a candle like form. Take a piece of an almond and put it into the apple or potato. Now light it with a lighter or matches.

Orange candle:

The orange is placed on the table with the stalk facing upwards or downwards. Cut the skin in the middle of the fruit with a knife, so that the top and bottom can be taken away. The half part with the stalk is filled up with olive oil.

CONCLUSIONS

Apple Candle: During the heating of the oil, it starts to evaporate, the steam burns above the almond. Almond has got a lot of fat, causing the flame to burn for a longer time. The Apple and potato are just for stabilisation and are not relevant for the actual combustion.

Orange Candle: By capillary effect of the stalk, olive oil is sucked into the upper end and is vaporized by the fire source, which burns above the stalk.

WASTE MANAGEMENT

No hazardous substances or chemicals are involved; therefore it is no problem to put them into the household waste or on the compost pile.

PHOTOS

Fig.1-3: preparing the orange candle

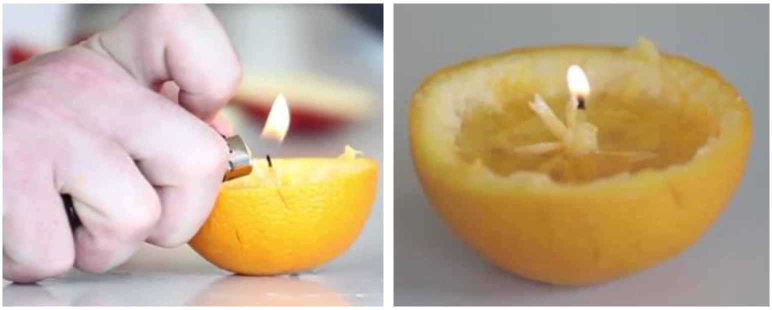

Fig.4-5: Lighting the stalk

Fig. 6-8: preparing the apple and lighting the almond

REFERENCES

1. https://www.youtube.com/watch?v=KKeGi-MYc3o 10.01.2017
2. https://www.youtube.com/watch?v=833FgExj-I0 10.01.2017

SEVEN COLOURS FROM ONE SOLUTION

OVERVIEW

This experiment shows the influence of different densities and polarities of substances and the behaviour, when added together in a container. Because of the different density and polarity, the chemicals are not mixing, but lay on top of each other. [1, 2]

MATERIALS AND METHODS

Take the following materials and chemicals in the order mentioned below and pour them into a measuring cylinder.

Devices	Chemicals
Measuring Cylinder	Honey
Peleus Ball	Karo Syrup
Pipette	Dish Soap
Spoon	Water
	Vegetable Oil
	Alcohol
	Lamp Oil
	Food Colour

HAZARDS

Lamp oil is highly flammable. Do not work near open flame.

Lamp oil 

WASTE MANAGEMENT

Because none of the listed chemicals are hazardous, it is no problem to pour it in the sink or the toilet. Lamp oil and vegetable oil should be collected separately in a tank.

PHOTOS

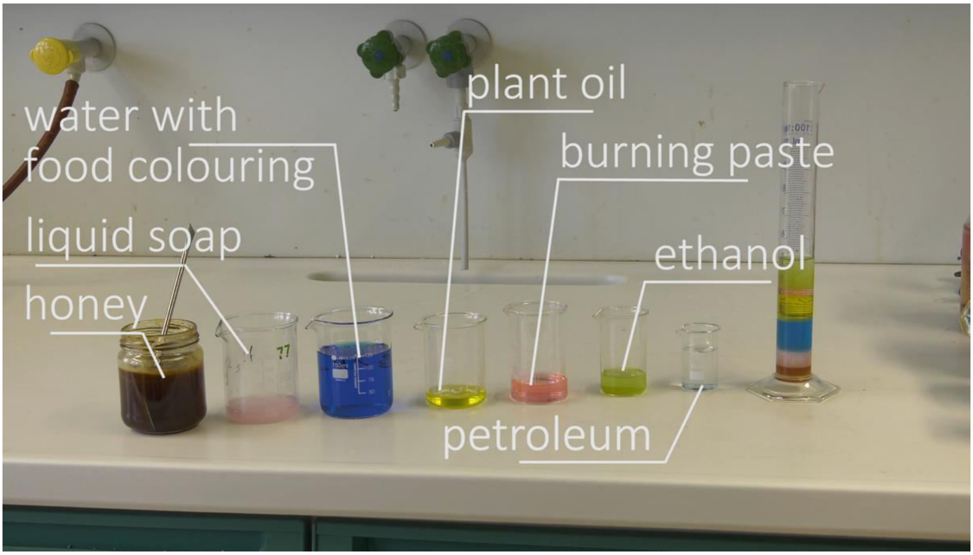

Fig. 1 Compounds for creating seven coloured layers

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2. http://www.lehrmittelboutique.net/home-mainmenu-1/751-chemie-im-alltag, 10.01.17

WATER DETECTION IN SOLID MATERIALS $CuSO_4 \cdot 5H_2O$

OVERVIEW

Salts are compounds consisting of a single metal or polyatomic ion which is positively charged and a non-metallic single or a polyatomic ion which is negatively charged. Hydrated salts are those containing water in a plexus. The number of water molecules contained in the crystal of a salt plexus is always an integer. Removal of water from a hydrated salt often is accompanied by a change of structure. This shows that between the metal or the anion and the water molecules, a weak bond is formed which upon heating can be cleaved [1].

Copper (II) sulphate, also known as cupric sulphate or copper sulphate, is the inorganic compound with the chemical formula $CuSO_4$. This salt exists as a series of compounds that differ in their degree of hydration. The anhydrous is in its pure form a white powder, whereas the pentahydrate ($CuSO_4 \cdot 5H_2O$), the most commonly encountered salt, is bright blue, known also as bluestone. The anhydrous form occurs as a rare mineral known as chalcocyanate. The hydrated copper sulfate occurs in nature as chalcanthite (pentahydrate), and two more rare ones: bonattite (trihydrate) and boothite (heptahydrate)**.** $CuSO_4 \cdot 5H_2O$ is usually used in various crops as fertilizer, reinforcing the ground e.g. wheat before sowing and spraying of fruit trees, especially in viticulture [2].

MATERIALS AND METHODS

In a glass test tube place about 10 grams of hydrated copper (II) sulphate. Then approach the test tube on a naked flame taking care not to have turned the test tube to us or to any other person. Observe that during heating the hydrated copper (II) sulphate changes colour from blue to white, while drops of water appearing on the walls of the test tube.

HAZARDS

Wear gloves, eye shields and a dust mask.

$CuSO_4 \cdot 5H_2O$

$CuSO_4$

CHEMICAL REACTION EQUATION

$$CuSO_4 \cdot 5H_2O_{(s)} \rightleftharpoons CuSO_{4(s)} + 5H_2O_{(g)}$$

CONCLUSIONS

The observed changes take place because the water present in hydrated copper (II) sulphate is evaporated due to the heating process. Then remove the test tube from the flame and in a few minutes we observe the color of hydrated copper (II) sulphate turn blue. The reason for this is the drops of water falling into the copper (II) sulphate.

WASTE MANAGEMENT

Copper dusts or mist or copper compounds may be disposed in Group III sealed containers in a secure sanitary landfill.

PHOTOS

Fig. 1 The $CuSO_4·5H_2O$ before heating.

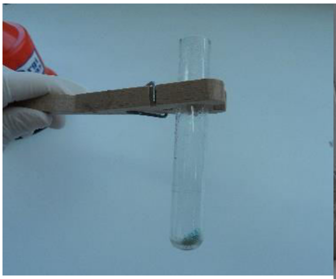

Fig. 2 $CuSO_4·5H_2O$ after heating turns to white while water drops appear on the tube.

Fig. 3 After a few minutes the $CuSO_4·5H_2O$ turns into blue again.

REFERENCES

1. P. Aslanidis, A. Giouri, C. Bolos; *Synthetic Inorganic Chemistry*; Pigasos Publishers, Thessaloniki 1995
2. https://en.wikipedia.org/wiki/Copper(II)_sulfate

DETERMINING THE CRYSTALWATER CONTENT OF SALTS

OVERVIEW

Having a look on the label of different salts provided by the chemical companies we often observe that the chemical formula of the salt include a various numbers of water molecules. These salts are called hydrated salts and their molar mass includes the mass of the water molecules included as these water molecules are present in the crystal lattice being called *crystal water*. We see for instance labels of $CaSO_4*2H_2O$, $2CH_3$-$COOH*3H_2O$, $CuSO_4*5H_2O$, $NiCl_2*6H_2O$, $NiSO_4*7H_2O$, $Na_2CO_3*10H_2O$. There is no rule concerning the number of water molecules for a salt. Knowing the number of these molecules is important as you have to take in account the mass of crystal water when preparing solutions of exact concentrations as for instance the content of salt in the hydrated soda, $Na_2CO_3*10H_2O$, is only 37%.

The following experimental setup describes the way of determining the crystal water of a salt. The procedure is general despite the fact that it is described for a certain salt, the principle being applicable for any hydrated salt.

MATERIALS AND METHODS

To determine the crystal water content of a salt you will need to heat a certain amount of hydrated salt to allow the removal (evaporation) of the water. Recording the initial and final mass of the salt (prior to heating and after being heated) you can perform calculations to determine the content of water.

Take an porcelain evaporation capsule with a lid and weigh it. Place in it 24.64 g of $MgSO_4*7H_2O$ and weigh the total mass. [1] Place the capsule on a heater and heat it to 250-300°C in order to allow the water molecules to leave the crystal lattice and to evaporate. If the mixture starts spitting, cover the capsule with the lid to avoid the salt to leave the system. When the salt is dry weigh it with the lid. Record the mass loss. Place again the capsule on the heater and continue heating. After a few minutes weigh again the vessels and the dehydrated salt. Record again the mass loss. Do these procedures until you get the same mass for two consecutive weighings. Make the calculation what is the amount of the salt in moles. Calculate the weight loss considering it's due only by water who leaves the system, calculate the number of moles of water and associate the two results in order to write the correct formula of the salt.

HAZARDS [2]

There are no hazards related to the work with magnesium sulfate as it is a salt used in medicine and agriculture.

CHEMICAL REACTION EQUATION

$$MgSO_4{}^{*}7\ H_2O_{(s)} \ ---> MgSO_{4(s)} \ + \ 7\ H_2O_{(l)}$$

CONCLUSIONS

Applying the described procedure we obtained the following results: 24.69 g of hydrated salt, 12.06 g of weight loss = 12.06 g of water = 0.67 moles of water, 12.63 g of dehydrated salt = 0.104 moles of salt.

The calculated formula for the salt is $MgSO_4{*}7H_2O$ which is in accord with the formula stated by the label of the jar.

WASTE MANAGEMENT

The magnesium sulfate is a naturally found dissolved in both surface and deep waters. The anhydrous salt can be further used for other experiments (determination of $SO_4{}^{2-}$) or either it might be discarded in the bin or dissolved with water and poured down the drain.

PHOTOS

Fig. 1 Hydrated salt of $MgSO_4{*}7H_2O$

Fig. 2 White solid salt resulted after dehydration of $MgSO_4{*}7H_2O$.

REFERENCES

1. Earl B.; Wilford D.; *Cambridge IGCSE Chemistry Laboratory Practical Book*; Hodder Education, London, 2015, p. 103-105
2. https://en.wikipedia.org/wiki/Magnesium_sulfate

BASIC REAGENTS FOR WATER, CARBON DIOXIDE AND SULFUR DIOXIDE

OVERVIEW

The combustion of many materials leads to the emission of water, carbon dioxide and possibly sulfur dioxide. The compounds will only be emitted if the starting material consists carbon, hydrogen and sulfur atoms. This experiment demonstrates a method for the indication of water, carbon dioxide and sulfur dioxide.

MATERIALS AND METHODS

Equipments" Bubbler (2), Watch glass, Vacuum pump, Combustion tube, Water, Spatula, Carbon powder, Hoses to connect the bubblers, Sulfur powder, Matches, Bunsen burner, Stand (2), Beaker, Pipette

Reagents Lime water, app. 200 ml; Iodine solution in water, app. 200 ml, Anhydrous copper sulphate, 1.0 gram

To indicate the presence of water, first inject a couple of mL in a beaker. Put the anhydrous copper sulphate on the watch glass and add a few drops of water to it with the pipette. Mark your observations.

To indicate carbon dioxide, use the spatula to add a bit of carbon in the combustion tube and fix it on the stand. Add the lime water to the bubbler and connect the bubbler and the combustion tube in a way that the gas leaving the tube will travel through the lime water. Then connect the bubbler to the vacuum pump and turn it on. Now light the Bunsen burner and heat the carbon in the combustion tube with a roaring blue flame. Mark your observations, both about the carbon as well as the lime water.

To indicate sulfur dioxide, use the same setup as with the indication of carbon dioxide. In this case, use sulfur instead of carbon and use the iodine solution instead of lime water.

HAZARDS

C		**I_2**	
S		**$Ca(OH)_2$**	
SO_2			

CHEMICAL REACTION EQUATION

$C_{(s)} + O_{2\,(g)} \rightarrow CO_{2\,(g)}$

$S (s) + O_2 (g) \rightarrow SO_2 (g)$

$CO_{2\,(g)} + H_2O_{(l)} \rightarrow 2\,H^+_{(aq)} + CO_3^{2-}{}_{(aq)}$

$H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$

$CO_3^{2-}{}_{(aq)} + Ca^{2+}_{(aq)} \rightarrow CaCO_{3\,(s)}$

$I_{2\,(aq)} + SO_{2\,(g)} + 2\,H_2O_{(l)} \rightarrow 2\,I^-_{(aq)} + SO_4^{2-}{}_{(aq)} + 4\,H^+_{(aq)}$

CONCLUSIONS

The identification of water can be done by adding it to anhydrous copper sulphate. It can be observed that the colour of the copper sulphate changes from white to blue when water is added.

The identification of carbon dioxide can be done by bubbling it through lime water. It can be observed that the lime water changes from clear and colourless to cloudy and white.

The identification of sulfur dioxide can be done by bubbling it through an iodine solution. It can be observed that the iodine solution changes from clear and brown to clear and colourless.

WASTE MANAGEMENT

The watch glass with blue copper sulphate must be discarded in the inorganic waste barrel. The used lime water can be discarded in the sink and the used iodine solution must be discarded with the halogen waste barrel.

PHOTOS

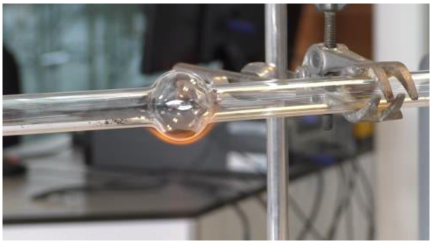

Fig. 1 Combustion tube with carbon

Fig. 2 Combustion tube with sulfur

REFERENCES

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2. Chemiekaarten 15e editie, 2000
3. http://www.labchem.com/tools/msds/msds/LC15590.pdf

FLASH IN A PAN

OVERVIEW

When oil is heated in a pan it can catch fire. In order to extinguish this fire, it is best not to use water as is demonstrated in this experiment. A fire blanket should be used instead.

MATERIALS AND METHODS

- Furnace crucible
- Bunsen burner
- Matches
- Crucible tongs
- Stand with triangle for crucible
- Organic oil, such as sunflower oil
- Water
- Pipette

Install the stand and the Bunsen burner in a fume hood and put 10 mL of oil in the crucible. Light the burner and heat the oil with a roaring blue flame until the oil starts to fume. Carefully take the burner and light the fume coming off the oil, so there is a flame above the oil. Now set the burner aside and carefully drip a few drops of water in the burning oil. Make sure you step back a bit, since a big flame will arise.

HAZARDS

None

CHEMICAL REACTION EQUATION

We will assume sunflower oil is an ester of glycerol and oleic acid, so its formula will be $C_{57}H_{104}O_6$.

$$\mathbf{C_{57}H_{104}O_{6\ (l)} + 80\ O_{2\ (g)} \text{ ---> } 57\ CO_{2\ (g)} + 52\ H_2O_{\ (g)}}$$

CONCLUSIONS

When you have a flash in the pan, it is discouraged to try to extinguish this with water since a blowtorch can arise. The advice is to use a fire blanket or simply use the cover of the pan to make sure oxygen will run out.

WASTE MANAGEMENT

The remaining oil must be thrown away in the organic waste bin.

PHOTOS

Fig 1. Flash in a pan

Fig. 2 Extinguishing the flame

REFERENCES

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IS MAGNESIUM COMBUSTIBLE?

OVERVIEW

Magnesium in its metal form can react with oxygen from the air to form magnesium oxide. In this experiment, a flame will be used as activation energy so the reaction between magnesium and oxygen takes place. The activation energy is the minimum amount of energy that is needed to make the reaction occur. In this process light and warmth will be produced. After the flame extinguishes, a white powder is formed (magnesium oxide). By doing this experiment we will investigate under which conditions magnesium is combustible.

MATERIALS AND METHODS

- Magnesium powder
- Magnesium ribbon
- Burner
- Matches
- Small spoon/spatula
- Crucible tongs
- Stand
- Scissors
- Petri dish

Start with gathering all of the materials. Then connect the burner to the gas outlet and light it with a match. Make sure the flame is a roaring blue flame. Now place the burner horizontally in the clip of the stand. Be careful of the flame, since it will bend upwards. Place the petri dish under the flame. Cut a piece of 2cm magnesium ribbon and use the crucible tongs to hold it in the flame. Remember to avoid looking directly into the blinding light! When the white flame has extinguished, place the remnants of the ribbon in the petri dish. After that, put a small amount of magnesium powder on the small spoon or spatula. Carefully drop this in the flame, while making sure that the powder falls into the petri dish.

HAZARDS

Mg

CHEMICAL REACTION EQUATION

$\mathbf{2\,Mg_{(s)} + O_{2\,(g)} \rightarrow 2\,MgO_{(s)}}$

CONCLUSIONS

Magnesium is combustible when using a roaring blue flame. It creates warmth and a white light. When the magnesium has a lot of surface and is more distributed, it is most combustible. For instance when using magnesium ribbon or powder.

WASTE MANAGEMENT

Magnesium oxide, the reaction product, is a solid chemical which may be thrown away in the trash can.

PHOTOS

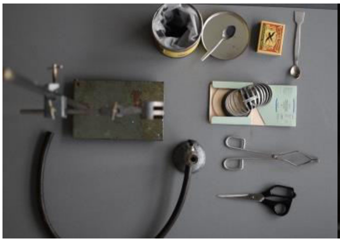

Fig. 1 Materials to be used

Fig. 2 Setup for the experiment with magnesium ribbon

Fig. 3 Setup for the experiment with magnesium powder

REFERENCES

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2. https://www.uwgb.edu/safety/envpolicies/

VARIATION OF METALLIC CHARACTER IN GROUP 1 -Li, Na, K REACTIVITY

OVERVIEW

Secondary students learn about metallic and nonmetallic character of the elements. From the beginning they learn that every element capable to donate electrons in a chemical process is a METAL. The metallic character is associated with the electron donor capacity. Then is taught the variation of the metallic character in group and in period. An experimental approach based on two sets of experiments, applied to three successive metals of the same group (Li, Na, K: Group 1 in the Periodic Table) can easily demonstrate how the metallic character varies in group. [1]

MATERIALS AND METHODS

1. Reaction of alkali metals with OXYGEN The first step in determining which alkali metal has a more evident metallic character is the reaction of three alkali metals with air (oxygen). Take one piece of each of the following metals: Li Lithium, Na Sodium, K Potassium, not bigger than a green pea size (0.1-0.2 g). Dab on a filter paper the taken pieces to absorb the paraffin oil or kerosene in which they were kept. Take a knife and cut in two a metal piece measuring, with a watch or chronometer, the time needed for the fresh surface, exposed to the oxygen from air, to get tarnished. Repeat twice the operation to ensure uniformity for the measured times. Try to recognize the same grade of oxidation for the recently exposed surface to ensure consistency of the measurements. Repeat the operation, successively, for the other two metals. Record in a table the measured data and try to conclude which metal is the most active in its reaction with oxygen.

2. Reaction with WATER After the study of the alkali metals behavior in air (oxygen) the next step is a comparative study of the three metals' reactivity with H_2O. Because the reaction is very energetic it's advisable to use large glass vessels, preferably crystallizers.

Fill in three crystallizers or Petri dishes with water and add a few drops of phenolphthalein solution. Add a small piece of lithium -Li- (0.2 g) in the first vessel. Observe where the metal stays during the reaction, notice its form, observe the evolved gas, notice the color of water after adding the metal in it and record the time needed to the metal piece to get consumed. Repeat the experiment in a second crystallizer with a piece (0.2 g) of sodium -Na-. Make the same observations and notice the differences and similarities. The third crystallizer will host the reaction of potassium -K- with water. Don't use more than 0.1 g as it reacts violently. There are similarities and differences, as well.

Some users cover the reaction vessel with a funnel, in order to avoid the spreading of the molten metal (Sodium and Potassium) or of the hydroxide solution if the reaction goes too vigorous. This technique is not advisable because it allows the evolved hydrogen to accumulate in the reaction site and a spark can ignite it generating a small explosion which can lead to a much bigger damage.

To avoid accidents during the reaction the best safety measure is to use small amounts of metal e.g. 0.1 g will do.

The presence of hydrogen can safely be demonstrated by collecting the evolved gas in a test tube kept over the reacting metal, test tube which is then put over an open flame (of a candle, gas burner or even a burning match) the hydrogen burning quickly emitting a characteristic sound.

HAZARDS

Attention has to be paid to Li, Na and K not to enter in contact with skin and eyes because they are extremely corrosive. When reacting with water, H_2 gas is released which can lead to explosions in a hot reaction site (especially for Na and K) [2,3,4]. Therefore only small amounts of metal should be used and do the reaction in a ventilated hood or behind a protective screen.

The filter papers used to dry the taken pieces of alkali metals, prior to their disposal, should be washed first in the drain to ensure the complete neutralization of metallic remains on them. Otherwise they can be ignited by the metallic remains and lead to fire in the laboratory.

If the reaction is performed in a test tube take care to orient it toward a neutral zone as melted sodium can be thrown out during the reaction! Do not do the experiment with potassium in a test tube nor with a piece larger than 0.1-0.15 g. Use always a protective screen.

Li, Na, K

H_2

CHEMICAL REACTION EQUATION

$$4\,Li_{(s)} + O_{2(g)} \longrightarrow 2\,Li_2O_{(s)}$$

$$4\,Na_{(s)} + O_{2(g)} \longrightarrow 2\,Na_2O_{(s)}$$

$$4\,K_{(s)} + O_{2(g)} \longrightarrow 2\,K_2O_{(s)}$$

$$2\,Li_{(s)} + 2\,H_2O_{(l)} \longrightarrow 2\,LiOH_{(aq)} + H_{2(g)}$$

$2\ Na_{(s)} + 2\ H_2O_{(l)} \dashrightarrow 2\ NaOH_{(aq)} + H_{2(g)}$

$2\ K_{(s)} + 2\ H_2O_{(l)} \dashrightarrow 2\ KOH_{(aq)} + H_{2(g)}$

CONCLUSIONS

Alkaline metals have a very low ionization energy which determine their high reactivity with a large range of elements and chemicals.

1. Alkaline metals keep their metallic shine as far as they don't react with oxygen and humidity from air. As water and oxygen are not soluble in nonpolar liquids, such as mineral oil or kerosene, the alkali metals are kept in these liquids which seal the metal from the environment. Once taken out and dried on a filter paper, the soft metal which can be cut with a knife is exposed to the oxygen from air and the characteristic shine gradually disappears due to the formation of an oxide layer on the fresh cut. The quickest disappearance is observed for K and the longest shine for Li as metallic character is increasing in group from Li to K. Potassium reacts almost instantly with oxygen covering itself with a layer of bluish oxide. Sodium covers itself with a layer of white and yellow oxides whereas lithium quickly forms a layer of a black oxide.

2. Even more obvious increase in the metallic character from $_{3}$Li to $_{11}$Na and further to $_{19}$K is observed via their behavior when reacting with water:

Lithium (Li) is a low density metal and floats on the surface of water. The observed effervescence is due to the evolved hydrogen gas which moves the piece of Li in all directions as it is consumed. The water becomes fuchsia-red due to a synthesis of a base: lithium hydroxide.

Sodium (Na) also floats on the surface of water but in addition the metal melts due to the highly exothermic reaction. Because the effervescence is more vigorous it generates a much active movement of the molten metallic sphere on the surface of water. The piece of metal is consumed more rapidly and the solution becomes intense fuchsia-red. When the metal is almost consumed sparks are emerging from the reaction site... this is the moment when accumulated hydrogen can explode. If the evolved hydrogen is ignited at the beginning of the reaction it burns with a yellow flame.

Potassium (K) is even more energetic as it has a stronger metallic character. The evolved hydrogen is self-igniting from the beginning due to the exothermic reaction, a reddish-violet flame is observed and in many cases molted potassium is thrown out from the reaction vessel (if using >0.1 g).

WASTE MANAGEMENT

The produced hydroxides are not dangerous for the environment because they are formed in small quantities. The content of the reaction vessels has to be diluted with large amounts of water and poured down the drain.

PHOTOS

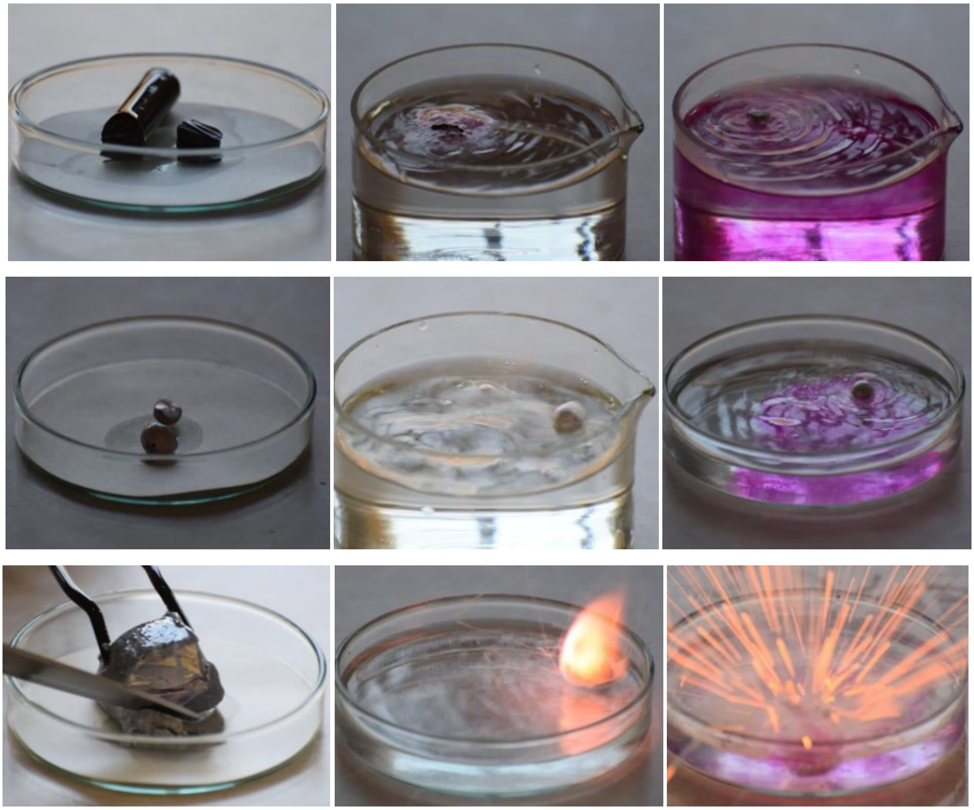

Fig. 1 - Fig. 9 Comparative study for Li (top), Na (middle) and K reactivity with water

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VARIATION OF METALLIC CHARACTER IN PERIOD 3 - Na, Mg, Al REACTIVITY

OVERVIEW

Secondary students are taught about metallic and nonmetallic character. Apart from the fact that they have to recognize the metals based on their electron donor capacity they have to make also semi-quantitative analyses by integrating their knowledge in comparing different metals and predict their strength in hypothetic situations. An experimental approach based on two sets of experiments, applied to three successive metals of the same period (Na, Mg and Al) can easily demonstrate how the metallic character varies in Period 3.

The reaction of Mg with oxygen has a historical significance in photography as it was replacing the electrical flash in the early age of the photo technology. Large quantities of Mg are used in fireworks (combined with different alkaline or earth-alkaline salts which color the sparks).

Aluminum on the other hand gives an interesting discussion subject, when analyzed, since its variation of free enthalpy is negative for the reaction with water but no reaction occurs. This is due to the superficial layer of aluminum oxide formed at the metal's surface by passive corrosion, layer which is transparent, very compact and seals the metal from action of both the oxygen and water [1,2].

MATERIALS AND METHODS

1. Reaction of period 3 metals with OXYGEN The first step in determining which metal (of the group Na, Mg, Al) has a more evident metallic character is their reaction with air (oxygen). When taking all three metals for the study we notice that Na is kept under mineral oil or kerosene. This is obviously the most reactive metal with oxygen whilst Mg and Al are just kept in simple plastic bottles and they can be handled freely in air. If taking a small piece of Na (a green pea size weighing 0.2 g) and we damp it on a filter paper we notice that if cut with knife it shows a metallic luster which tarnishes in seconds after the surface was exposed to air. So its reactivity with oxygen is far higher than the other two metals in discussion.

Mg and Al on the other hand keep their shine for a long period. To compare their reactivity with oxygen, fine powdered Mg and Al is taken with a spatula (around 0.2 g) and their reaction with oxygen is observed only when the metal is poured gently over a gas burner's flame.

The comparison of the flame with Mg and respectively Al burned in it, gives an obvious measurement of the metallic character of these metals.

2. Reaction with WATER After the study of the three metals behavior in air (oxygen) the next step is a comparative study of the three metals' reactivity with H_2O. Because the reaction is very energetic for Na it's advisable to use large glass vessels, as for example crystallizers or Petri dishes.

Put in three crystallizers 50 mL of water and add a few drops of phenolphthalein solution. Add a small piece of Na (approx. 0.1-0.2 g), in the first crystallizer. Observe where the metal stays during the reaction, notice its form, observe the evolved gas, observe the movement of the metallic piece, notice the color of water after adding the metal in it and record the time needed for the metallic piece to get consumed.

Repeat the experiment in a second crystallizer with 0.2 g Mg turnings. Make the same observations and notice the differences and similarities. Observe the reactivity of Mg when reacted with hot water (70 °C).The third crystallizer will host the interaction of Al Aluminum with water. Observe if there are similarities and notice the differences.

The presence of hydrogen can be safely demonstrated only for Na, by collecting the evolved gas in a test tube kept over the reacting metal, test tube which is then put over an open flame (of a candle, gas burner or even a burning match). You will hear the characteristic sound of H_2 burning in a tube.

HAZARDS

General safety rules were presented, for the alkali metals, in the previous experiment: "*Variation of the metallic character in Group 1*". The same rules apply and have to be obeyed for Na in this experiment too. To avoid accidents during the reaction of Na with water the best safety measure is to use small amounts of metal e.g. 0.1 g. Wear goggles, apron and safety gloves when handling all chemicals.

Na

H_2

Mg, Al

Attention has to be paid when working with Magnesium (Mg) or Aluminum (Al). When finely powdered Mg or Al is used, avoid the close proximity of the container to an open flame as it can lead to the ignition of the entire content causing severe fire. DO NOT STARE at the Mg flame when Mg is "burned" in open fire to avoid eye damage!

CHEMICAL REACTION EQUATION

$4\ Na_{(s)} + O_{2(g)} \text{--->} 2\ Na_2O_{(s)}$

$2\ Mg_{(s)} + O_{2(g)} \text{--->} 2\ MgO_{(s)}$

$4\ Al_{(s)} + 3O_{2(g)} \text{--->} 2\ Al_2O_{3(s)}$

$2\ Na_{(s)} + 2\ H_2O_{(l)} \text{--->} 2\ NaOH_{(aq)} + H_{2(g)}$

$Mg_{(s)} + 2\ H_2O_{(l)} \text{--->} Mg(OH)_{2(aq)} + H_{2(g)}$

$Al_{(s)} + H_2O_{(l)}$ **---> no reaction**

CONCLUSIONS

Alkaline metals have a very low ionization energy which determine their high reactivity with a large range of elements and chemicals.

1. REACTION WITH OXYGEN Alkaline metals (in this case Na) keep their metallic shine as far as they don't react with oxygen and humidity from air. As water and oxygen are not soluble in nonpolar liquids, such as mineral oil or kerosene, sodium is kept in these liquids which seal the metal from the environment. Once taken out and dried on a filter paper, the soft metal which can be cut with a knife is exposed to the oxygen from air and the characteristic shine gradually disappears due to the formation of an oxide layer on the fresh cut.

Aluminum powder when poured over a flame generates small shiny sparks which can be associated with the oxidation process of the metallic dust generating Al_2O_3. The reaction is more vigorous with Mg the metallic dust generating a very bright flame (similar to a flash) while producing MgO.

If the same experiment is conducted with Mg band respectively Al band we will observe an auto ignition for Mg and a very bright reaction spot until the entire band is consumed whilst the Al band goes red but it doesn't "ignite".

2. REACTION WITH WATER Sodium (Na) is a low density metal and floats on the surface of water. The observed effervescence is due to the evolved hydrogen gas which moves the piece of Na in all directions as it is consumed. The water becomes fuchsia-red due to a synthesis of a base: sodium hydroxide which turns red the indicator present originally in the water: the phenolphthalein. Once dropped in water the metal melts due to the highly exothermic reaction. When the metal is almost consumed sparks are emerging from the reaction site... this is the moment when accumulated hydrogen can explode. If the evolved hydrogen is ignited at the beginning of the reaction it burns with a yellow flame.

Magnesium (Mg) does not react visibly with water. If the water is heated a pale pink coloration of it suggests a hydroxide formation, accompanied by few gas bubbles. This means Mg reacts with hot water leading to magnesium hydroxide. [1]

Aluminum (Al) requires special surface cleaning techniques in order to make visible the reaction (e.g. cleaning the metal surface with mercury chloride).

Taking in account all these observations it's obvious that the metallic character is increasing in Groups from the right to left (from Al to Mg and Na).

A very interesting experiment can be carried out in a test tube with sodium in a two layer environment: water and kerosene (or gasoline). Na has a density higher than the gasoline and lower than the water. So it sinks in gasoline to the water, meeting the water reacts with it, generates hydrogen and with the H_2 bubbles attached to it floats back to the surface of gasoline, looses the hydrogen "swimming belt" when becomes again denser than gasoline and the process repeats several times giving the impression of a dancing piece of Na.

WASTE MANAGEMENT

The produced hydroxides and oxides are not dangerous for the environment as they are formed in small quantities. The content of the reaction vessels has to be diluted with large amounts of water and afterwards can be poured down the drain.

PHOTOS

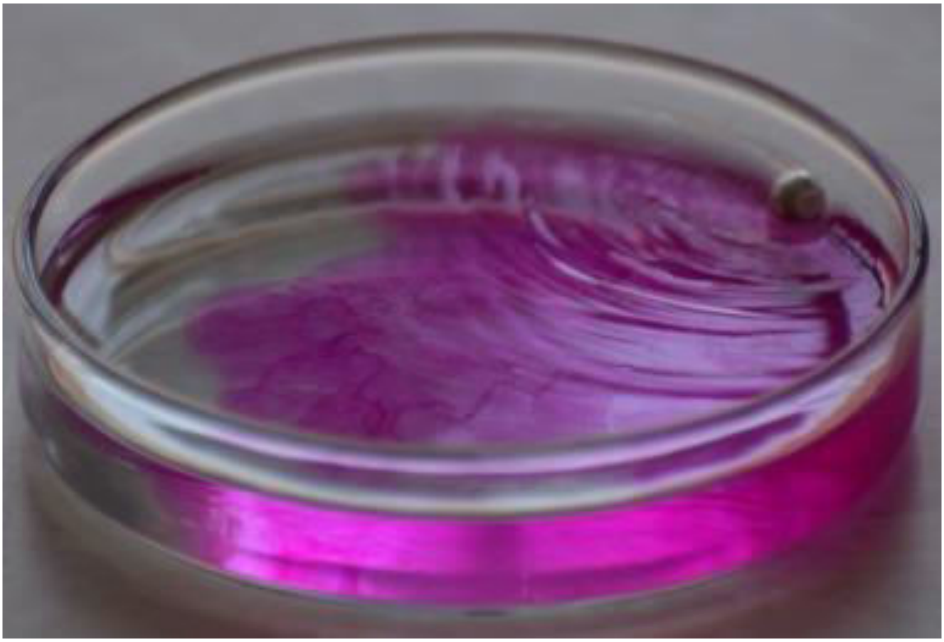

Fig. 1 Period 3 metal: Na and its interaction with water

Fig. 2-3 Period 3 metals Mg (left) and Al (right) and thier interaction with water

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METALS REACTION WITH WATER - COMPARISON

OVERVIEW

Metals react more or less quickly and violently with water. This depends on the position of the element in a group in the periodic system. Less noble metals have a faster and more intensive reaction with water. A base metal is sodium or potassium. For example, noble metals are copper, gold, silver. By the reaction between water and an earth metal, the water gets basic. All the alkali metals react vigorously with cold water. In each reaction, hydrogen gas is given off and the metal hydroxide is produced. The speed and violence of the reaction increases within the element group of the periodic system. [1, 2]

MATERIALS AND METHODS

The following materials are tested: Copper (or if available gold, silver), Magnesium, Sodium, Lithium as metals. Take a small amount of a metal (around 0,1 g) and put it into a bowl full of water. As evidence for the emerging hydrogen, observe the development of a gas and perhaps an audible pop from the burning reaction. The generated alkaline solution is detected by means of pH paper. For this test take a small strip of pH paper and dip it into the water. When the strip showing a pH > 7, the water is basic.

HAZARDS

Attention has to be paid to Li and Na. They should not get in contact with skin and eyes because they are extremely corrosive. When reacting with water, H_2 gas is released which can lead to explosions. Only use small amounts of metal and perform the reaction in a ventilated hood or behind a screen.

Na

Li

H_2

CHEMICAL REACTION EQUATION

$Na_{(s)} + H_2O_{(l)} \text{--->} H_{2(g)} + NaOH_{(aq)}$

CONCLUSIONS

The reaction between ignoble metals and water produces hydrogen and lye. The reaction usually runs hard and fast.

There is almost no reaction between noble metals and water. The reaction usually runs slowly and can take days, weeks, years, for example the discoloration of copper on roofs.

WASTE MANAGEMENT

The waste solutions should be neutralized with a weak acid.

PHOTOS

Fig. 1 No reaction of copper with water

Fig. 2 Sodium reacting with water

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REACTIONS OF DIFFERENT METALS WITH OXIGEN

OVERVIEW

As showed in the two previous experiments (experiments 33 and 34) metals react with oxygen from air providing metallic oxides. This reaction can be used for secondary students to introduce general properties of metals, the basic character of metallic oxides not to mention that these simple processes are most useful when discussing the combination (or synthesis) reactions in early years of studying Chemistry.

Also through these reactions different approaches can be used to speak, in the classroom, about corrosion, other electrochemical aspects and/or redox reactions. Comparing the way different metals react with oxygen and observing the results, qualitative conclusions can be drawn concerning the reactivity of metals, their reductive capacity thus their metallic character.

Whereas the above mentioned experiments were dealing with the reaction between metals and oxygen on the purpose to determine the variation of the metallic character in a series (in group or in period) the present approach will simply highlight the differences between metals when reacting with the same nonmetal: the oxygen.

MATERIALS AND METHODS

The present series of chemistry experiment uses a general approach: that of sprinkling a small quantity of metallic dust over an open flame and observing the sparks emitted by the metallic powder when hot and reacting with oxygen. Only safe metals are chosen in order to avoid exposure to dangerous chemicals and obviously alkali metals cannot be used as they are provided in chunks under mineral oil.

Adjust a gas burner's flame to be oxidative and as colorless as possible. Take alternatively the following metals: Cu, Sn, Ni, Zn, Cr, Fe, Mn, Al, Mg, Pb and sprinkle a pinch of each one over the flame. Observe the reaction for each metal and record the observation in a table. After doing the whole series repeat the experiment in order to ensure proper observations and recordings. Try to interpret the observed phenomena and range the metals according to their brightness in the flame. Can you arrange them in the same order as they are in the electrochemical series?

HAZARDS

As a general rule finely powdered metals are dangerous being very reactive therefore they are considered flammable and capable to generate hydrogen (explosive) when reacting with water (Mg, Al, Fe, Mn). Some metals are irritant for skin, respiratory tract or eyes (Cu, Ni, Sn). Others are harmful (Ni). Interaction with the environment has to be a concern when using Zn, Cr. As the element Sb is carcinogenic if inhaled its use is not recommended. Same for Co which is very toxic [1]. If possible perform the reaction in a fume hood or a well ventilated place. Wear goggles, apron and safety gloves when handling all chemicals.

Mg, Al, Fe, Mn		**Ni**	
Cu, Sn		**Zn, Cr**	

Attention has to be paid when working with Magnesium (Mg) or Aluminium (Al). When finely powdered Mg or Al is used, avoid the close proximity of the container to an open flame as it can lead to the ignition of the entire content causing severe fire. DO NOT STARE at the Mg flame when Mg is "burned" in open fire to avoid eye damage!

CHEMICAL REACTION EQUATION

$$2\,Cu_{(s)} + O_{2(g)} \text{ ---> } 2\,CuO_{(s)}$$

$$2\,Zn_{(s)} + O_{2(g)} \text{ ---> } 2\,ZnO_{(s)}$$

$$3\,Fe_{(s)} + 2\,O_{2(g)} \text{ ---> } Fe_3O_{4(s)}$$

$$2\,Mg_{(s)} + O_{2(g)} \text{ ---> } 2\,MgO_{(s)}$$

$$4\,Al_{(s)} + 3\,O_{2(g)} \text{ ---> } 2\,Al_2O_{3(s)}$$

CONCLUSIONS

Copper (Cu) dust in flame oxidizes to CuO showing red glowing sparks and the fame is coloured in green. Zinc (Zn) shows a greater reactivity with white sparks, not very abbundants. If iron (Fe) is used the sparks are brighter, turning to yellow, the process leading to a mixture of oxides of which the most common is the hematite Fe_3O_4 [2]. Aluminum (Al) powder, when sprinkled over a flame, generates small but abbundant white shiny sparks which can be associated with the oxidation process of the metallic dust generating Al_2O_3. The reaction is far more vigorous with magnesium (Mg) the metallic dust generating a very bright flame (similar to a flash) while MgO is produced.

WASTE MANAGEMENT

The produced oxides are not dangerous for the environment as they are formed in small quantities. The reaction site has to be cleaned from any metallic rest. Dispose the produced oxides in the trash bin or in a container for storing inorganic waste.

PHOTOS

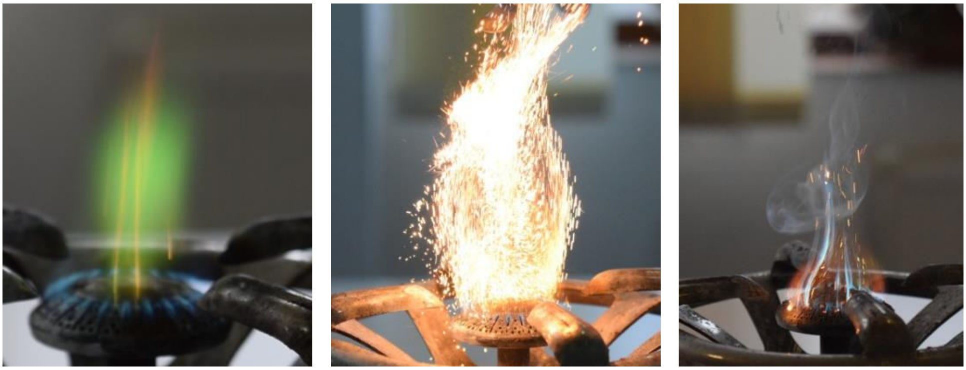

Fig. 1-3 As the sparks are whiter the reactivity of the metal reacting with oxygen is higher. Copper (Cu) left; Iron (Fe) middle; Zinc (Zn) right.

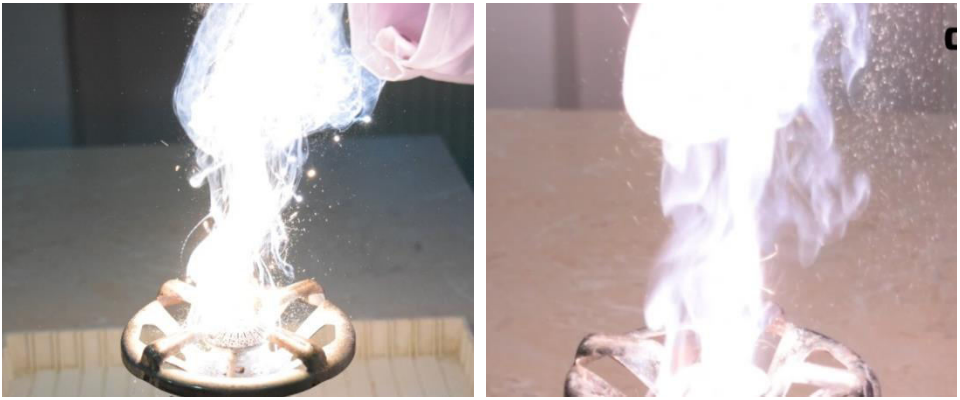

Fig. 4-5 High reactivity (and electrochemical potential) shown by bright sparks formed in the reaction with oxygen shown by Aluminium (**Al - left**) and Magnesium (**Mg - right**)

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GALVANIC ELEMENTS AND ELECTROCHEMICAL POTENTIAL

OVERVIEW

In this experiment the difference between the potentials of two electrodes that dip into the two solutions is measured. Each electrode-solution pair builds up a half-cell, the sum of the two half cell potentials is measured. The whole arrangement is called a galvanic cell. A typical cell consists of two pieces of metal, each immersed each in a solution containing a dissolved salt of the corresponding metal. The two solutions are separated by a porous barrier which prevents them from rapidly mixing but allows ions to diffuse.

In the arrangement of the experiment different metals are combined, and the different voltages between the electrodes can be measured. The values may be compared to the combined values for the elements according to the table of electrochemical potential. [1,2,3,4]

MATERIALS AND METHODS

The following materials are used: Zinc-plate, Copper-plate, Aluminium-plate, Iron-plate, Nickel-plate, Lead-plate. Fill the six beakers with 100 mL acetic acid [C=0,1 mol/L] each. In each of the beakers place one of the metal plates. Connect two of the elements to the voltage meter and measure the electric voltage for each combination.

HAZARDS

Acetic Acid is a weak acid. It attacks some forms of plastic, rubber and coatings. Heavy metals like Pb, Ni are harmful to health and environment. Wear gloves and avoid remaining or rubbed-off material. [3]

Pb

Ni

CH_3COOH (0,1 M)

CHEMICAL REACTION EQUATION

$$2\,Al_{(s)} + 3\,Cu^{2+}_{(aq)} \longrightarrow 2\,Al^{3+}_{(aq)} + 3\,Cu_{(s)}$$

$$Zn_{(s)} + Cu^{2+}_{(aq)} \longrightarrow Zn^{2+}_{(aq)} + Cu_{(s)}$$

$$Fe_{(s)} + Cu^{2+}_{(aq)} \longrightarrow Fe^{2+}_{(aq)} + Cu_{(s)}$$

$$Ni_{(s)} + Cu^{2+}_{(aq)} \longrightarrow Ni^{2+}_{(aq)} + Cu_{(s)}$$

$$\mathrm{Pb_{(s)} + Cu^{2+}_{(aq)} \longrightarrow Pb^{2+}_{(aq)} + Cu_{(s)}}$$
$$\mathrm{Zn_{(s)} + Pb^{2+}_{(aq)} \longrightarrow Zn^{2+}_{(aq)} + Pb_{(s)}}$$
$$\mathrm{Fe_{(s)} + Pb^{2+}_{(aq)} \longrightarrow Fe^{2+}_{(aq)} + Pb_{(s)}}$$
$$\mathrm{2\,Al_{(s)} + 3\,Zn^{2+}_{(aq)} \longrightarrow 2\,Al^{3+}_{(aq)} + 3\,Zn_{(s)}}$$

CONCLUSIONS

The different combination of metals show different voltages. Compare the measured values with the difference of the voltages in the table of the electrochemical potentials. There may be differences because the documented values refer to standard conditions (concentration of electrolyte 1 M).

WASTE MANAGEMENT

The remaining solutions and rubbed-off heavy metal material are collected separately in containers.

PHOTOS

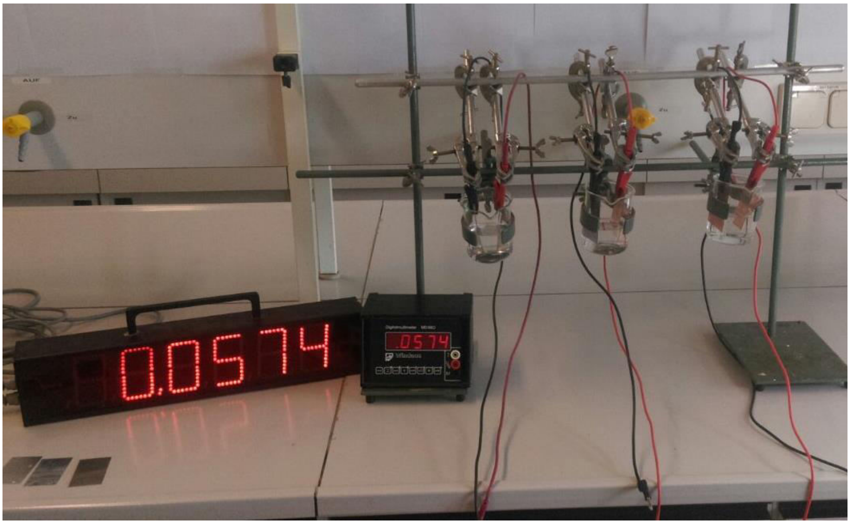

Fig. 1 Construction of 3 different galvanic cells to change the element combinations rapidly

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COMPARISON OF COPPER AND SILVER REACTIVITY

OVERVIEW

Copper is a soft, malleable and ductile metal with very high thermal and electrical conductivity. A freshly exposed surface of pure copper has a reddish-orange colour. Copper was the first metal used by humans to replace the stone. The copper name (cuprum) and the symbol Cu comes from the Latin aes Cuprium which means "from Cyprus", because the Romans brought the first copper from Cyprus [1]. Presently, copper has many different uses in building construction, power generation and transmission, electronic product manufacturing, and production of industrial machinery and transportation vehicles.

Silver nitrate compound is a versatile precursor to many other silver compounds, such as those used in photography. It is far less sensitive to light than the halides [2].

In this experiment we study the reactivity series of metals by means of a type of oxidation-reduction chemical reaction, which is the reaction in which electrons are transferred between atoms, or else the reaction in which the atoms change the oxidation state. In this experiment a simple displacement reaction is used in which a component reacts with a compound and replaces one of the compounds' elements [3].

MATERIALS AND METHODS

Place in a beaker 10 mL of $AgNO_3$ solution (20%w/v) and add a copper strip. Observe the color formed. The copper strip quickly becomes coated with a bunch of elemental silver. At the same time, the copper solution turns into a characteristic blue.

HAZARDS

Due to the toxicity of the reagent and the product, avoid skin and eye contact. Wear protective glasses and gloves. $Cu(NO_3)_2$ is harmful when swallowed.

Cu

$AgNO_3$

$Cu(NO_3)_2$

CHEMICAL REACTION EQUATION

$Cu_{(s)} + 2\ AgNO_{3\ (aq)}$ ---> $Cu(NO_3)_{2\ (aq)} + 2\ Ag_{(s)}$

CONCLUSIONS

The silver crystallizes on the surface of the copper and the copper strip quickly becomes coated with a bunch of elemental silver. At the same time, the copper ions go into solution and the colourless solution turns into a characteristic blue as the concentration of copper ions is growing [4]. This is due to the higher electrochemical potential of Cu comparing to the electrochemical potential of the silver Ag. Due to this fact, the Cu atoms suffer an oxidation process reducing at the same time the silver Ag^+ ions. Thus, silver atoms are deposed onto the copper strip. Cu is more reactive than Ag.

WASTE MANAGEMENT

The solution from the reaction is not poured down the drain because of the toxicity of the $Cu(NO_3)_2$. This solution should be collected in a container used for collecting heavy metals. Ag, once separated from the solution, can be disposed in a container for heavy metals because it is highly toxic for microorganisms. Otherwise, Ag formed during the experiment should be recycled for qualitative use as silver nitrate.

PHOTOS

Fig. 1 $AgNO_3$ solution and Copper strip

Fig 2. The copper strip quickly becomes coated with a bunch of elemental silver.

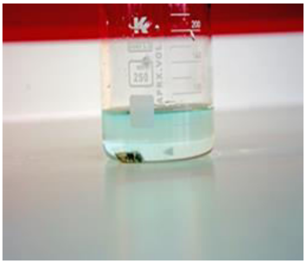

Fig. 3 The colourless solution of $AgNO_3$ turns into a characteristic blue as the concentration of copper ions builds

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COMPARISON OF COPPER AND IRON REACTIVITY

OVERVIEW

Iron is a metal in the first transition series. It is by mass the most common element on Earth, forming much of Earth's outer and inner core. It is the fourth most common element in the Earth's crust [1]. The use of iron dates back to prehistoric times. The Iron Age is the technological and cultural step that followed the Stone and Copper Age. The Hittites were the first that manufactured iron tools. Biologically iron is a component of all organisms because it plays an important role in the transport of oxygen carried by hemoglobin, as and in oxygens storage. The deficiency causes anemia. The symbol Fe comes from the Latin word ferrum.

Copper (II) sulfate, also known as cupric sulfate or copper sulphate, is an inorganic compound. Copper sulfate pentahydrate is used as an herbicide, fungicide and pesticide. It has also niche uses. It has been used as analytical reagent, in organic synthesis, in art and in dyeing [2].

In this experiment we study the reactivity series of metals by means of a type of an oxidation-reduction chemical reaction, which is the reaction in which electrons are transferred between atoms, or else the reaction in which the atoms change oxidation state. In this experiment is used a simple displacement reaction in which a component is reacted with a compound and replaces one of the compounds elements [4].

MATERIALS AND METHODS

Put in a beaker an iron nail and add about 100 mL of an aqueous solution of $CuSO_4$ (20%w/v). Observe that the iron nail obtains a brown coating.

HAZARDS

Due to the toxicity of the reagents and the products avoid the skin and eye contact. Wear protective glasses and gloves.

$CuSO_4 \cdot 5H_2O$

$FeSO_4$

Cu

CHEMICAL REACTION EQUATION

$Fe_{(s)} + CuSO_{4(aq)} \longrightarrow FeSO_{4(aq)} + Cu_{(s)}$

CONCLUSIONS

After several minutes we observe the color of the iron nail change due to the deposition of Cu onto the iron nail. Specifically, there is a brown coating on the iron nail dipped in the $CuSO_4$ solution. This is due to the higher electrochemical potential of Fe comparing to the electrochemical potential of the Cu. Due to this fact the Fe atoms suffer an oxidation process reducing in the same time the silver Cu^{2+} ions. Thus copper atoms are deposed onto the iron nail. Fe is more reactive than Cu [5].

WASTE MANAGEMENT

The solution from the reaction is not poured down the drain because of the toxicity of the $FeSO_4$. This solution should be collected in a container used for collecting heavy metals. Cu once separated from the solution can be disposed in a container used for collecting heavy metals.

PHOTOS

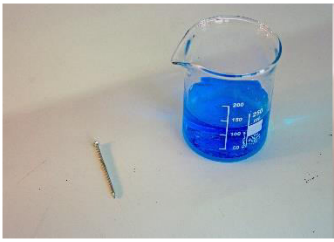

Fig. 1 Iron nail and $CuSO_4$ solution.

Fig. 2 It starts the reaction between Fe and $CuSO_4$ solution.

Fig. 3 We observe the color change of the iron nail due to the deposition of Cu onto the iron nail.

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ACID – BASE REACTIONS

OVERVIEW

Following Bronsted's definition, acids are proton-donors and bases are proton-acceptors. When equilibrium conditions are favourable, reactions will proceed spontaneously and almost instantaneously because the reaction mechanisms are simple in the sense that no complicated transition states with corresponding high activation energies have to be overcome.

Acid-base reactions have a wide range of applicability. In living organisms the functioning of proteins depends largely on the value of the pH and stomach acid is nothing more than diluted hydrochloric acid. Also in chemical practice acid-base reactions are widely used, e.g. for synthesis of salts and analytical purposes.

In our direct environment acid-base reactions are frequently met as well. Detergents we use can be acid as well as basic and everybody has heard of acid rain. For an introductory experiment, we choose to take an example from common daily household practice: the deposition of calcium carbonate. Especially in regions where water contains huge amounts of Ca^{2+} and / or Mg^{2+} ions equipment like washing machines, water boilers and coffee makers can suffer seriously from scaling of carbonates.

At moderate temperatures, CO_2 is, to some extent, soluble in water causing the pH-value to drop somewhat below 7 due to the following equilibrium:

$$CO_{2(aq)} + 2\ H_2O_{(l)} \leftrightarrow HCO_3{}^-{}_{(aq)} + H_3O^+{}_{(aq)} \qquad (1)$$

When we increase the temperature, the solubility of carbon dioxide decreases rapidly. When Ca^{2+} is present, the following reaction will occur:

$$Ca^{2+}{}_{(aq)} + 2\ HCO_3{}^-{}_{(aq)} \rightarrow CaCO_{3(s)} + CO_{2(aq)} + H_2O_{(l)} \qquad (2)$$

The solid $CaCO_3$ will be deposited at the heat exchanger surfaces, leading to malfunctioning of the apparatus. Therefore, the scaling layer has to be removed periodically. This can be done with an acid – base reaction. The scaling layer will dissolve upon addition of acid:

$$CaCO_{3(s)} + 2\ H_3O^+{}_{(aq)} \rightarrow Ca^{2+}{}_{(aq)} + CO_{2(aq)} + H_2O_{(l)} \qquad (3)$$

Of course we can expect reaction (3) to proceed rapidly upon addition of strong and concentrated acid. However, we have to restrict ourselves to weak acids in order not to damage our equipment while conducting the cleaning operation. In this experiment we will compare dissolution properties of calcium carbonate in different acid solutions.

MATERIALS AND METHODS

Requirements: A beaker 250 mL, a stirrer / heater, a thermometer, 0.1 M hydrochloric acid, 0.1 M acetic acid, $CaCO_3$.

Procedure: Fill the beaker with 100 mL of acid solution. Add about 0.5 grams of calcium carbonate while stirring, and measure the dissolution time. The experiment should be conducted for both acid solutions at room temperature and elevated temperature (70 – 80^0C.)

HAZARDS

CH_3COOH

HCl

CHEMICAL REACTION EQUATIONS

$$\mathbf{CO_2(aq) + 2\ H_2O(l) \leftrightarrow HCO_3^{-}(aq) + H_3O^{+}(aq)}$$

$$\mathbf{Ca^{2+}(aq) + 2\ HCO_3^{-}(aq) \rightarrow CaCO_3(s) + CO_2(aq) + H_2O(l)}$$

$$\mathbf{CaCO_3(s) + 2\ H_3O^{+}(aq) \rightarrow Ca^{2+}(aq) + CO_2(aq) + H_2O(l)}$$

CONCLUSIONS

The reaction time for dissolution of calcium carbonate in hydrochloric acid is only slightly less than in acetic acid. Thus, due to much more “friendly” pH-values, acetic acid is recommended.

WASTE MANAGEMENT

Diluted solutions of HCl and acetic acid can be treated as waste water.

PHOTOS

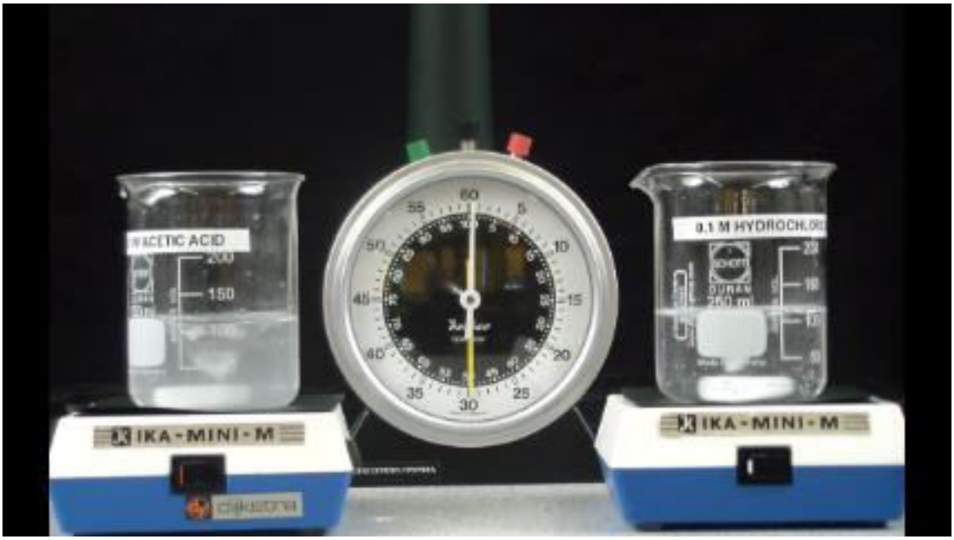

Fig. 1 After 30 seconds the solution containing HCl is clear. In acetic acid, dissolution is not complete yet.

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THE ACIDIC SOLUTION EFFECT IN METALS

OVERVIEW

Zinc is the first element of group 12 of the periodic table. Ancient people used the alloys of zinc, because they faced many difficulties in the production of pure metal from the mineral. Zinc dazzles from moisture and air. Also, this is used to protect other metals, particularly iron and steel against atmospheric corrosion by a process called galvanization. Its alloys are of great industrial interest [1].

Copper is a soft, malleable and ductile metal with very high thermal and electrical conductivity. A freshly exposed surface of pure copper has a reddish-orange colour. Copper was the first metal used by humans to replace the stone. The copper name (cuprum) and the symbol Cu comes from the Latin aes Cuprium which means “from Cyprus”, because the Romans took the first copper from Cyprus [1]. Presently, copper has many different uses among building construction, power generation and transmission, electronic product manufacturing, and production of industrial machinery and transportation vehicles [1].

Aluminum is a silvery-white, soft, nonmagnetic, ductile metal. It is the third most abundant element in the Earth's crust and the most abundant metal. Aluminium is a metal with important applications because it has high thermal and electrical conductivity with high corrosion resistance. When it participates in the alloys, it makes them more hard and durable [1].

Hydrochloric acid is a clear, colorless, highly pungent solution of hydrogen chloride in water. It is a highly corrosive, strong mineral acid with many industrial uses. Hydrochloric acid is found naturally in gastric acid [2].

In this experiment we study the reactivity series of metals by means of a type of oxidation-reduction chemical reaction, which is the reaction in which electrons are transferred between atoms, or else the reaction in which the atoms change oxidation state. In this experiment a simple displacement reaction is used in which a component is reacted with a compound and replaces one of the compounds’ elements [3].

MATERIALS AND METHODS

In three glass test tubes place a few ml of HCl (38%w/v). In the first tube add Cu powder. In the next tube add powder of Zn. In the last tube add a sheet of Al. Observe the reactions that take place.

HAZARDS

The materials are corrosive, very toxic and hazardous. Avoid inhalation. Wear gloves and protective glasses. Perform the experiment under a gas absorbent.

HCl

Cu

Al

$ZnCl_2$

$AlCl_3$

Zn

CHEMICAL REACTION EQUATION

$Cu_{(s)} + HCl_{(aq)} \rightarrow X$

$Zn_{(s)} + 2HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$

$2Al_{(s)} + 6HCl_{(aq)} \rightarrow 2AlCl_{3(aq)} + 3H_{2(g)}$

CONCLUSIONS

In the tube with Cu no change is visible, because Cu is less reactive than H_2 so single displacement reaction is not performed. However, in the second tube we can release hydrogen gas because Zn is more reactive than H_2 so single displacement reaction is performed. In the third tube the reaction is very intense, because Al is more reactive than both H_2 and Zn. The strong reaction of Al with hydrochloric acid is not quite visible in this video because aluminum is in sheet form which does not have a large contact surface. Unlike the Zn, which is in the form of powder, the occurring reaction with hydrochloric acid is very intense. Furthermore, the temperature in the two last tubes increases because the reaction is exothermic. In conclusion, the less reactive metal is Cu and the most reactive one is Al.

WASTE MANAGEMENT

The solution of $ZnCl_2$ is not poured down the drain because of the toxicity. This solution should be collected in a container used for collecting heavy metals. The mixture between HCl and Cu should be separated and the Cu can be disposed in a container used for collecting heavy metals, while the HCl solution can be disposed in an appropriate container. The solution of $AlCl_3$ is not poured down the drain because of the toxicity. This solution should be collected in a container used for collecting heavy metals.

PHOTOS

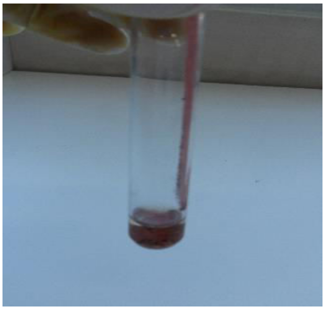

Fig. 1 No reaction occurs because Cu is less reactive than hydrogen.

Fig. 2 The intense reaction between Zn and HCl.

Fig. 3 The intense reaction between Al and HCl.

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THE ACIDIC SOLUTION EFFECT IN CARBONATES

OVERVIEW

In this experiment we observe the acidic solution effect in carbonates. A carbonate is a salt of carbonic acid (H_2CO_3) which dissociates in H_3O^+ and CO_3^{2-} [1]. Sodium bicarbonate is a chemical compound with the formula $NaHCO_3$. It is a salt composed of sodium ions and bicarbonate ions. Sodium bicarbonate is a white solid that is crystalline but often appears as a fine powder. It has a slightly salty, alkaline taste resembling that of washing soda (sodium carbonate). The natural mineral form is nahcolite. It is a component of the mineral natron and is found dissolved in many mineral springs. It is one of the food additives encoded by European Union, identified as E 500. Since it has long been known and is widely used, the salt has many related names such as baking soda, bread soda, cooking soda, and bicarbonate of soda [2].

A carboxylic is an organic compound that contains a carboxyl group –COOH. The general formula of a carboxylic acid is R–COOH, with R referring to the rest of the (possibly quite large) molecule. In nature we find many carboxylic acids such as acetic acid which is the main organic component of vinegar, butanoic acid which is responsible for the unpleasant smell of the spoiled butter and hexanoic acid which is responsible for the smell of goats and sweaty feet. Other examples are cholic acid, the main component of human bile and oleic and linoleic acid which are precursors of fats and other lipids [3]. Acetic acid is a colourless liquid organic compound with the chemical formula CH3COOH. Vinegar is roughly 3–9% acetic acid by volume, making acetic acid the main component of vinegar apart from water. Acetic acid has a distinctive sour taste and pungent smell. Although it is classified as a weak acid, concentrated acetic acid is corrosive and can attack the skin [4].

MATERIALS AND METHODS

In a test tube place about 5mL of vinegar. Then place 5 grams of $NaHCO_3$ in a balloon. Immediately place the balloon on the top of the tube. Observe the balloon inflating.

HAZARDS

Attention should be given to the fact that CH_3COOH is flammable and it should not come in contact with an open flame. Wear protective glasses and gloves.

CH_3COOH **CO_2**

CHEMICAL REACTION EQUATION

$$CH_3COOH_{(aq)} + NaHCO_{3(S)} \longrightarrow CH_3COONa_{(s)} + CO_{2(g)} + H_2O_{(l)}$$

CONCLUSIONS

We observe that the balloon is inflated due to the production of gas CO_2 from the reaction between acid and carbonate.

WASTE MANAGEMENT

The precipitate can be placed in the trash bin.

PHOTOS

Fig. 1 The balloon is inflated due to the production of gas CO_2

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THE FLAME REACTION

OVERVIEW

Some alkaline and earth-alkaline metals, in their ionic form, are currently found in each living cell and participating to many biochemical processes. Despite of their common occurrence in the living cells it is quite difficult to identify them in different solutions for they don't have very specific precipitation reaction or they don't form specific coloured complexes. Despite this there are techniques which allow a very quick and easy determination of these cations. Using tiny amounts of alkaline and earth-alkaline salts and a colourless flame there is always possible to identify them. [1]

MATERIALS AND METHODS

Alkaline and earth-alkaline metals are very easily put in evidence with flame spectroscopy. This experiment aims only to highlight the presence or absence of a certain element so a colourless flame will do very well and no need to have a flame photometer as it is necessary for quantitative measurements. The experimental method is very simple and it requires small amounts for chemicals therefore is also a safe reaction from ecological point of view. Just introduce a small amount of metallic compound, on a chemically inert support, directly in a colourless flame and notice the colour appearance. Each alkaline and earth-alkaline metal are colouring the flame in specific shades (see Table 1). There are also other metals (d- and p-block) which are colouring the flame of a gas burner but they are not very specific (except Thallium which is described elsewhere) [2].

Set the burner's flame to an oxidative state with a pale blue colour.

As chemically inert support you can use either a platinum loop of wire or even a 0.7-0.9 mm pencil lead. Both supports have to be thoroughly cleaned before use and between two consecutive analyses. If a pencil lead is used, prior to the first analysis it has to be burned in the gas burner until the flame is colourless. Warning! The pencil lead is made of graphite which conducts the heat so take care not to burn your fingers. Also take care as some leads are breaking in pieces when burned...so use protective goggles.

Once cleaned or burned, the support is wetted in double distilled water and it's immersed in a probe containing a metallic salt or a solution of the salt e.g. [LiCl, NaCl, K_2SO_4, KNO_3 (avoid KCl for it's often unpurified with small amounts of NaCl which alters the observations), $CaCl_2$, $Ca(NO_3)_2$, $Ba(OH)_2$, $Ba(NO_3)_2$]. The few crystals adhered on the wet support, once introduced in a naked flame, give a characteristic colour of the flame according to the metal present in the analysed salt.

Flame colours of different metals and their wavelength (nm) [4]

E	Colour	λ	E	Colour	λ
Li	Deep carmine red	**670**	**Co**	Blue	**475**
Na	Bright yellow	**590**	**Ca**	Orange red	**660**
K	Light purple	**435**	**Sr**	Crimson	**740**
Rb	Purple reddish		**Ba**	Pale yellow green	**620**

Another approach uses an aqueous solution of the salt sprayed over a colourless flame. The drawback of this second method is that you have to thoroughly clean the spray after each use of a metal salt or you need several sprays, as many as the metal salts are to be analysed.

HAZARDS

Li salts

Li, Na, K, Mg, Ca, Ba nitrates

Ba salts (except sulphate)

CONCLUSIONS

The heat of the flame vaporizes the solution over the flame (if sprayed), breaks the ionic bonds in the salt and creates free atoms. The created atoms are excited into excited electronic states, due to the thermal energy of the burner flame and when returning to ground electronic state they emit the previously absorbed energy in the form of light. Each element emits light at characteristic wavelengths [4] (e.g. for Na from 3p back to 3s causing an emission at 589.3 nm [5]). Potassium K^+ salts give similar colour to the burner flame's (435 nm) so a blue cobalt glass shall be used to evidence the purple/violet colour. Also some advanced techniques use the green filter to better observe the colour emerging. These properties of different metals are used to make the fireworks more coloured when exploding in pyrotechnical display.

WASTE MANAGEMENT

As we are using tiny amounts of alkaline and earth-alkaline salts, which are analysed in flame, there is no special procedure for waste management. Just obey general safety rules not to spill the reagents on the working place, to avoid skin and eye contact with the salts used, taking in account also their possible skin absorbance and toxicity (e.g. Barium salts).

PHOTOS

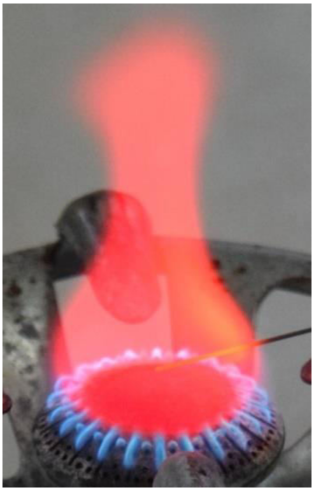

Fig. 1 Li -deep carmine red

Fig. 2 Na - bright yellow

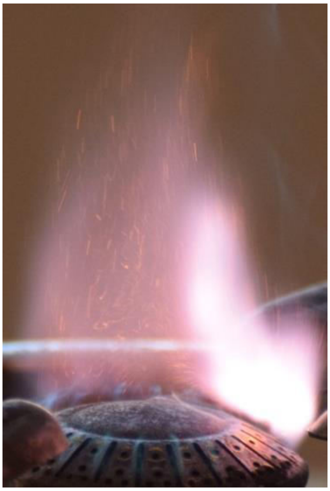

Fig. 3 K- light purple

Fig. 4 Ca - orange red

Fig. 5 Ba -pale yellow green

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WHY IS THALLIUM CALLED SO?!?

OVERVIEW

Thallium is a metal from the "p" block of elements with atomic number Z=81 and its most stable isotopes ^{203}Tl and ^{205}Tl. It's an extremely soft metal malleable and soft enough to be cut with a knife. If kept under protective atmosphere it has a gray color with metallic luster. Exposed to the air it rapidly tarnishes and covers the metallic surface with a grey layer [1].

MATERIALS AND METHODS

The discovery of the element Thallium (Tl) is linked to the names of William Crookes (British researcher) and Claude-Auguste Lamy (French researcher) who independently discovered a specific green line, in 1861, when analyzing different substances using flame spectroscopy. The notably green spectral line observed for Tl containing chemicals gave the name of this element *THALLIUM* from the Greek ***thallos*** (ταλλοξ) *green shoot or twig*. Both researchers showed Tl containing products at the International Exhibition in 1962 [1] and it's an interesting story to see the parentage of this element.

Thallium is very easily put in evidence when using flame spectroscopy even if a simple gas burner is used. This experiment aims only to highlight the presence or absence of a certain element so a colorless flame will do very well and no need to have a flame photometer.

As reagent support you can use either a platinum loop of wire or even a 0.7 mm pencil lead. Both supports have to be thoroughly cleaned before use. If a pencil lead is used, prior to the first analysis it has to be burned in the gas burner until the flame is colorless. Warning! The pencil lead is made of graphite which conducts the heat. Also take care as some leads are breaking in pieces when burned. Normally a 0.7 or 0.9 mm lead will do.

Once cleaned or burned the support is wetted in double distilled water and it's immersed in a probe containing a Thallium salt for example ($TlNO_3$, Tl_2SO_4 or $Tl(NO_3)_3*3H_2O$). The few crystals adhered on the wet support, once introduced in a naked flame, give a characteristic, bright, green color of the flame. This color motivates the name of the element ***thallium (green)***.

HAZARDS

The human body absorbs thallium very effectively, especially through the skin, the breathing organs and the digestive tract. [2] Thallium (I) salts are highly toxic because of their solubility in water (most of them) and because the ionic radii Tl^+ is similar to K^+ and it can enter the body via potassium uptake pathways. Tl salts are suspected human carcinogens [3].

Even if the online literature is inconsistent concerning the GHS symbols for the Tl salts, T sign appear in all materials and the O or N are quite frequent. Thallium salts were used as pesticides, mainly as rodenticides during the last two centuries, now it's banned because of its low selectivity [4].

$TlNO_3$		**$Tl(NO_3)_3*3H_2O$**	
Tl_2SO_4		**TlCl**	

CONCLUSIONS

A sample of a solid thallium salt is brought in the flame by use of a pencil lead (or platinum wire). The heat of the flame breaks the ionic bonds in the thallium salt and creates free atoms/ions. The created atoms are excited into excited electronic states, due to the thermal energy of the burner flame and when returning to ground electronic state they emit the previously absorbed energy in the form of light. Each element emits light at characteristic wavelengths, for Thallium the abundant green light being very characteristic.

WASTE MANAGEMENT

As the experimental processes described in this material are using tiny amounts of Tl^+ or Tl^{3+} salts, which are analyzed in flame, there is no special procedure for waste management. Just obey general safety rules not to spill reagent on the working place, to avoid skin and eye contact with the salts used taking in account also its cumulative toxicity.

PHOTOS

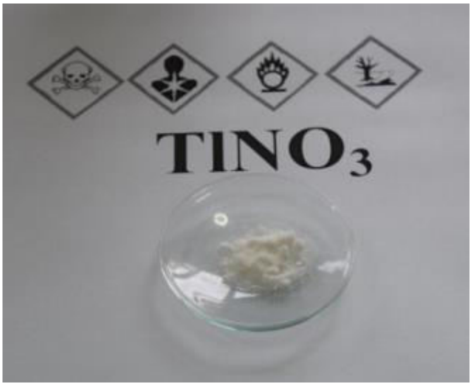

Fig. 1 Tl (I) salts looks like and behave chemically as alkaline metal salts but they are highly toxic.

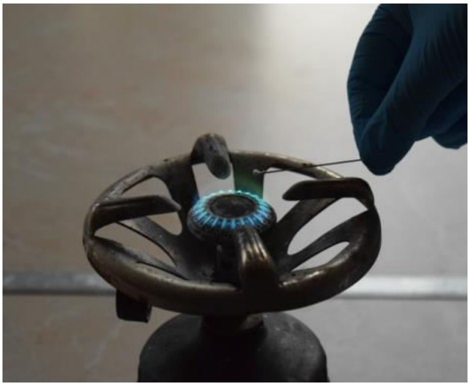

Fig. 2 Tl(I) sample analysis in flame using a previously cleaned pencil lead.

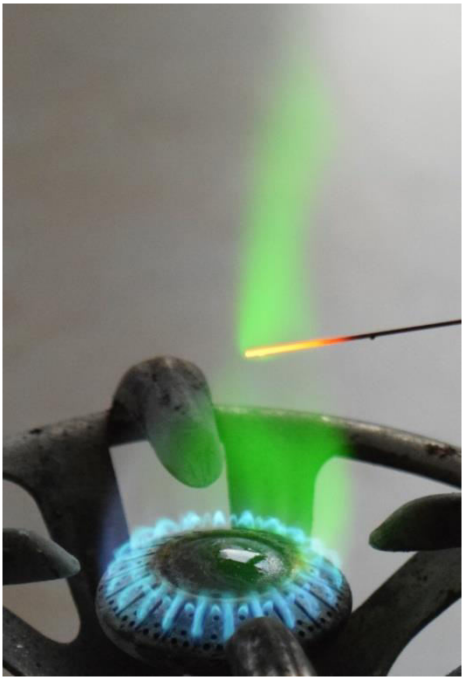

Fig. 3 Green flame of the Tl(I) sample

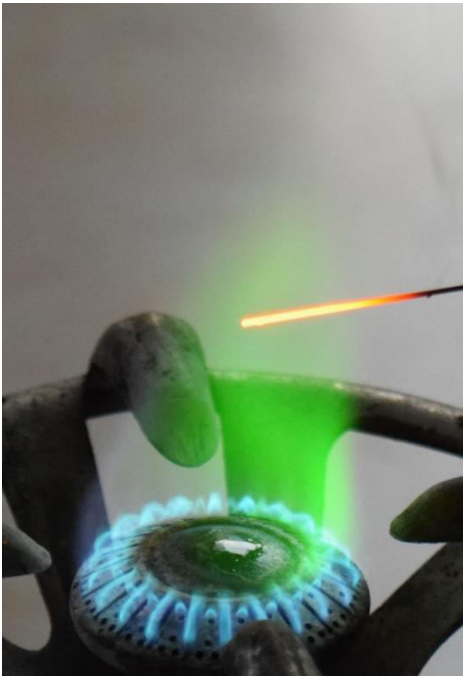

Fig. 4 Thallium (Tl) comes from the Greek ***thallos*** (ταλλοξ) *green shoot or twig*

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REACTION BETWEEN SODIUM AND SULFURIC ACID

OVERVIEW

Sodium is a very reactive, soft, silvery metal. It easily reacts with a large range of elements and of different chemicals [1]. This is the reason why metallic sodium is kept under paraffin oil or kerosene to avoid the contact with oxygen or moisture. On the other hand, metallic sodium is a fine candidate to study different types of chemical reactions as it is relatively cheap and it's a very versatile metal from chemical point of view.

Sodium is a strong metal reacting violently with water. Thus a non aqueous environment has to be found when studying the reaction of sodium with acids. A good candidate for this can be the pure sulfuric acid. If the reaction is performed in an aqueous solution the results are distorted due to the parallel reaction of Na with water [2].

MATERIALS AND METHODS

The reaction is to be performed on small scale, in a test tube. Place a 160 mm test tube with a few mL of pure H_2SO_4 on a stand clamp, behind a protective screen.

Take a small piece of metallic sodium 0.1 g (of a half green pea size ...or even smaller), dab it on a filter paper to absorb the paraffin oil and then place it, with the help of tweezers, into the previously setup test tube.

HAZARDS

Attention has to be paid to Na and H_2SO_4 not to enter in contact with skin and eyes because both are extremely corrosive. When reacting with acids and water, H_2 gas is released and this can lead to explosions in a hot reaction site. Therefore only small amount of Na should be used and do the reaction in a ventilated hood or behind a protective screen. There are complex side reactions occurring, some of them generating SO_2 which is toxic. Do not inhale the evolved vapors and gas.

Take care to orient the test tube toward a neutral zone as molten sodium and concentrated sulfuric acid can be thrown out during the reaction!

Na		**H_2SO_4**	
H_2		**SO_2**	

CHEMICAL REACTION EQUATION

$2\ Na_{(s)} + H_2SO_{4(l)} \text{--->} Na_2SO_4 + H_{2(g)}$

CONCLUSIONS

Sodium is an active metal capable to remove hydrogen from acidic compounds. When reacting with H_2SO_4, which is a strong proton donor, a very vigorous reaction occurs. A huge amount of heat is released which makes incandescent the reaction site. The evolved hydrogen is self ignited due to the heat released and sometimes explodes throwing out pieces of melted sodium. Due to the violence of the reaction side reactions also occur, the most notable being the redox process in which the sulfuric acid is reduced to sulfur dioxide SO_2.

WASTE MANAGEMENT

The produced sodium sulfate is not dangerous for the environment. The remained acid has to be neutralized using classical neutralizing approaches, by adding the resulted (or remained) chemicals to a beaker containing milk of lime. After neutralization the content has to be diluted with a large amount of water and afterwards can be poured down the drain [2].

PHOTOS

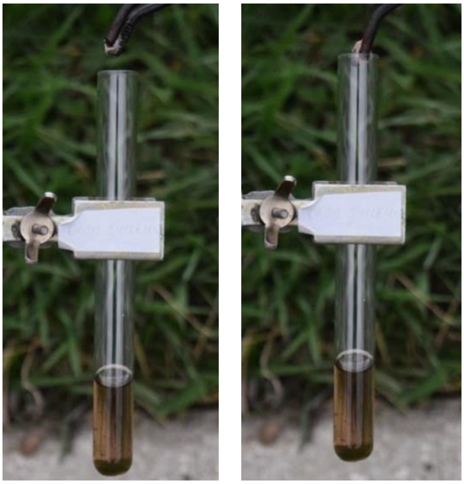

Fig. 1 and **2** A 0.1 g piece of sodium is dropped in technical H_2SO_4 98% solution

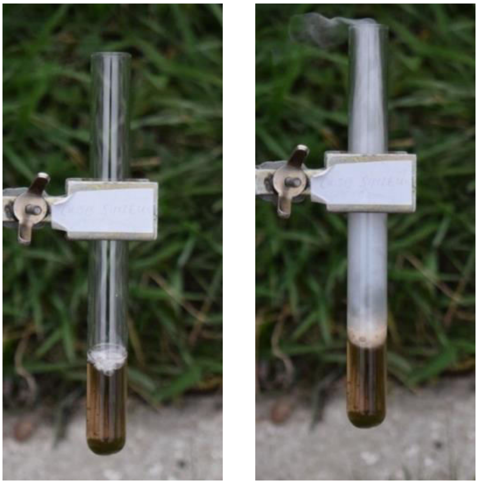

Fig. 3 and **4** The reaction is instantaneous, vigorous and the sulfuric acid starts to oxidize the metal evolving hydrogen and sulfur dioxide SO_2.

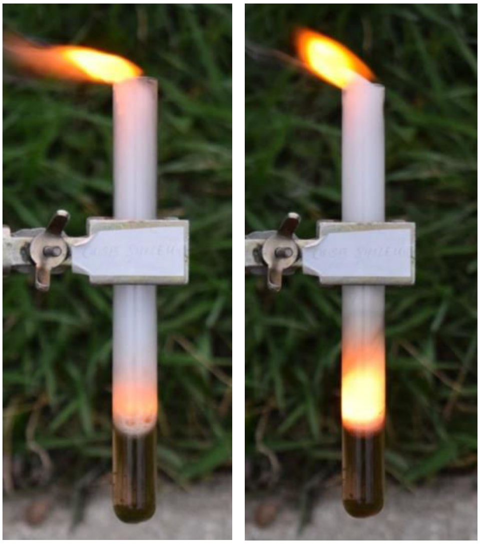

Fig. 5 and **6** The reaction zone becomes hot which ignites the evolved hydrogen.

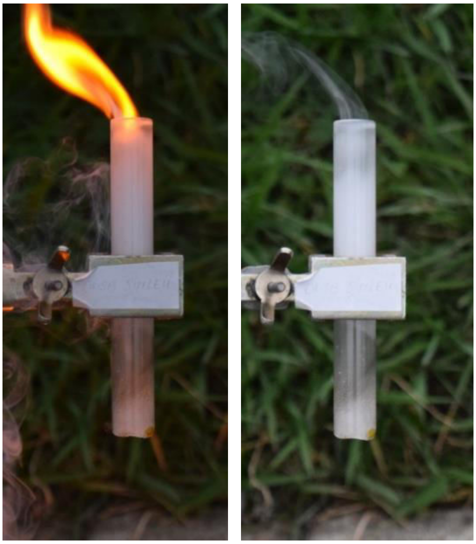

Fig. 7 and **8** Due to the strong exothermic character of the reaction sometimes the glass test tube breaks.

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SYNTHESIS OF SODIUM CHLORIDE FROM ELEMENTS

OVERVIEW

The synthesis of sodium chloride is a good example to demonstrate a redox reaction in the laboratory. Sodium Chloride is known as table salt in the kitchen. It may be synthesized from the metallic sodium and gaseous chlorine. In this experiment, gaseous chlorine is produced from a redox reaction of potassium permanganate and hydrochloric acid. The corresponding sodium and chloride ions are created by an electron transition between sodium and chlorine, forming an ionic bond between the elements. Sodium metal and chlorine gas are particular forms of matter. In burning, they undergo a chemical change- a chemical reaction- in which these forms of matter change to a form of matter with different chemical and physical properties. [1,2]

MATERIALS AND METHODS

Chemicals:

Potassium permanganate ($KMnO_4$), concentrated hydrochloric acid (HCl, 32% w/w), metallic sodium (Na)

Materials:

Bunsen burner, two stands, test tube, cannula, two syringes (with about 5 mL and 20 mL volume)

Procedure:

Fill a test tube with potassium permanganate (about 0.5 cm high). Attach a plug with two syringes to the test tube. The bigger syringe contains about 20 mL of concentrated hydrochloric acid. Once a few drops of the acid are added, the other syringes filling up with chlorine gas. Whilst adding the acid, a small piece of sodium is put into another test tube and heated with the Bunsen burner. As the sodium begins to melt and glow, the syringe with the chlorine gas is transferred to the test tube with the molten sodium. The immediate, highly exothermic reaction between chlorine gas and sodium is observed.

HAZARDS

Keep metallic sodium away from combustible materials and water. Concentrated hydrochloric acid is highly corrosive. Only use the required quantity of potassium permanganate. Follow the hazard and safety instructions on the label as well as the instructions for use/dosage. For sodium chloride are no special protective measures required. Chlorine gas is toxic, strongly irritant and harmful tom the environment. Always wear protective goggles and gloves. Avoid contact with your skin.

Na **HCl**

$KMnO_4$ **Cl_2**

CHEMICAL REACTION EQUATION

$16\ HCl_{(l)} + 2\ KMnO_{4(s)} \rightarrow 5\ Cl_{2(g)} + 2\ MnCl_{2(aq)} + 2\ KCl_{(aq)} + 8\ H_2O_{(l)}$

$2\ Na_{(s)} + Cl_{2(g)} \rightarrow 2\ NaCl_{(s)}$

CONCLUSIONS

Chlorine is a yellow-greenish gas. Heated sodium burns with a bright flame. The two substances react in an exothermic reaction to produce sodium chloride. After the reaction, the test tube is covered with a white layer of crystallized sodium chloride.

WASTE MANAGEMENT

The remaining chemicals may be diluted with water and poured into the sink. Remaining or rubbed- off heavy metal material is collected separately in containers.

PHOTOS

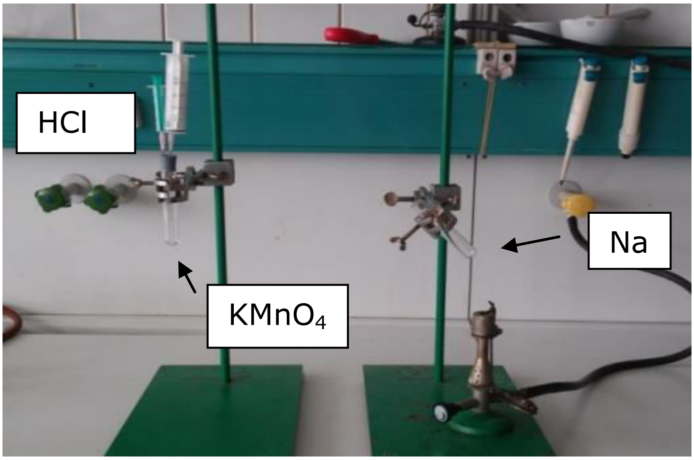

Fig. 1 Synthesis of sodium chloride

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THE REACTION BETWEEN SODIUM AND SULFUR

OVERVIEW

The reaction of sodium with sulfur is a simple combination reaction leading to a single reaction product the sodium sulfide Na_2S. The reaction has pedagogical value not only for the study of different reaction types but also when analyzing the chemical behavior of the metal (Na), the reactivity of the nonmetal and/or when studying the exothermic reactions. Also discussions can be done on the subject: *sodium-sulfur battery*, which accounts for the study of this reaction [1].

Sodium is a very reactive, soft, silvery metal. It easily reacts with a large range of elements and of different chemicals [2]. This is the reason why metallic sodium is kept under paraffin oil or kerosene to avoid the contact with oxygen or moisture. On the other hand, metallic sodium is a fine candidate to study different types of chemical reactions as it is a very versatile metal from chemical point of view.

MATERIALS AND METHODS

The reaction is to be performed on small scale, in a porcelain crucible placed over a burner's flame. Take a small piece of metallic sodium (0.4 g), dab it on a filter paper to absorb the paraffin oil and then place it, with the help of tweezers, into the crucible (or watch glass) containing 0.3 g of sulfur (powder). Mix it carefully and place the crucible on a fire resistant support over a flame to ensure the melting of the Na. As soon as the sodium is melted the reaction begins and soon it turns very energetic.

HAZARDS

Attention has to be paid to Na not to enter in contact with skin and eyes because it's extremely corrosive. When reacting with sulfur, SO_2 gas is released, which is toxic, due a side reaction occurring because of the high exothermic character of the main reaction. Therefore only small amounts of Na should be used and do the reaction in a ventilated hood or behind a protective screen.

Take care during the experiment as molten sodium is often thrown out during the reaction!

Na

SO_2

Na_2S

CHEMICAL REACTION EQUATION

$16\ Na_{(l)} + S_{8(l)} \ \text{--->}\ 8\ Na_2S_{(s)}$

$S_{8(l)} + 8\ O_{2(g)} \ \text{--->}\ 8\ SO_{2(g)}$

CONCLUSIONS

Sodium is an active metal capable to give out easily the outmost electron which is very easily lost [3]. When reacting with sulfur a very vigorous reaction occurs. A huge amount of heat is released which makes incandescent the reaction site. Observe the yellow flame characteristic to the Na thermal excitation. Observe the yellowish color of the produced sodium sulfide (other polysulfide are equally produced).

Due to the violence of the reaction side reactions also occur, the most notable being the sulfur reacting with oxygen. Sulfur is burning in oxygen with blue flame and evolves sulfur dioxide SO_2 which is toxic.

WASTE MANAGEMENT

The produced sodium sulfide is corrosive and dangerous for the environment. Small quantities of sodium sulfide dissolved in water can be neutralized with sodium hydroxide solution 1% and sodium hypochlorite solution 5% in a 1:1.75 ratio. Let the mixture stand overnight, than it can be safely poured down the drain as Na_2S was transformed in Na_2SO_4 [4].

PHOTOS

Fig. 1 Place the mixture of Na and S on a disposable watch glass and heat it.

Fig. 2 The sulfur and the sodium start melting.

Fig. 3 Once melted the reaction begins and rapidly becomes very energetic.

Fig. 4 Energetic reaction between sulfur and sodium. The dense fume is the sulfur dioxide generated by the side reaction of sulfur burning in oxygen. The yellow color of the reaction site is specific for Na atom excited in flame (see experiment: *The flame reaction*).

Fig. 5 Hot sodium sulfide generated in the reaction. Sometimes the used watch glass or the crucible breaks into pieces due to the generated hot.

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REACTION BETWEEN SULFUR AND IRON

OVERVIEW

This experiment shows the exothermic reaction of two elements, iron and sulfur, to form a compound, iron sulfide. The two solids are mixed and heated in a test tube. The two solids are mixed and heated in a test tube. In this reaction, the properties of the substances are changed. Iron is magnetic, but iron sulfide is not. This reaction is exothermic, energy is released in form of heat and light. Also, the colour is changing showing clearly the reaction process. [1, 2, 3, 4, 5]

MATERIALS AND METHODS

Method 1: Construct a stand with a test tube holder. Then add approximately 0.5 cm of fine iron powder into a test tube. Subsequently add sulfur purissimum in excess. If the mixture is heated by a Teclu burner, the iron reacts with the sulfur to form iron sulfide.

Method 2: Put 3 g of fine iron powder on a marble slab. Subsequently add sulfur purissimum in excess. Mix the two substances. Heat a glass rod for 40 seconds with a Teclu burner and hold it into the mixture, the iron reacts with the sulfur to iron sulfide.

HAZARDS

Sulfur causes skin irritation, so protective gloves are recommended. Iron sulfide is very toxic to aquatic organisms and develops toxic gases if it gets in contact with acids.

Fe	none	**S**	
FeS			

CHEMICAL REACTION EQUATION

$Fe_{(s)}$ + $S_{(s)}$ ---> $FeS_{(s)}$

CONCLUSIONS

This experiment shows the exothermic reaction of iron and sulfur. The two solids are mixed and heated in a test tube. The result is iron sulfide, a black and gold coloured cluster. It is also known as Pyrite, Marcasite or fool’s gold.

By heating, iron is oxidized to the cation Fe^{2+} and sulfur is reduced to the anion S^{2-}. The cation and the anion react to golden shining iron sulfide FeS. In this reaction heat is released.

WASTE MANAGEMENT

The product is collected separately for reuse in other experiments.

PHOTOS

Fig. 1 Method 2 with a glass rod

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THE REACTION BETWEEN ZINC AND SULFUR

OVERVIEW

Zinc is an important metal both for technical purposes and for life. Copper/Zinc alloys, commonly named *brass,* are widely used in our households. The galvanization method finds its applications for protecting iron from rusting. Its biological importance is acknowledged as it is an essential mineral. Its combinations are forming different enzymes (e.g. alcohol dehydrogenase) or even hormones (insulin) and some are used as dietary supplement, in deodorants, anti-dandruff shampoo. [1] Being such a common metal, reactivity studies are relevant in the secondary level.

The reaction of zinc with sulfur is another simple combination reaction leading to a single reaction product: the zinc sulfide ZnS (see also the experiment describing the reaction of sodium (Na) with sulfur . The chemical process is an example for the study of different reaction types, for analyzing the chemical behavior of the metal (Zn) or of the nonmetal and it's a good example either when studying the exothermic reactions or when studying the reaction rate as an example of fast occurring reaction.

MATERIALS AND METHODS

The reaction is to be performed on small scale, in a porcelain evaporating dish, porcelain crucible or on a disposable watch glass in a ventilated hood or outside, behind a protective screen. Mix carefully 1.0 g of Zinc (Zn) powder and 2.0 g of sulfur . Place it in the crucible and either initiate the reaction with a flame or with a red hot iron rod (both long enough to avoid burning during the reaction). A big flame occurs and instantly the entire reaction mass is transformed in a burning ball of fire. The process can be initiated also by heating the mixture of Zn-S the reaction starting once the sulfur is molten.

HAZARDS

Zinc powder is a highly reactive metal. It can spontaneously catch fire when exposed to air. In contact with water can release hydrogen which is highly flammable. Also is very toxic to aquatic life [2].

For the reaction product: the ZnS, avoid dust inhalation and skin contact. The reaction is very quick and it generates a large amount of heat. Therefore only small amount of Zn should be used to do the reaction. When reacting with sulfur, SO_2 gas is released, which is toxic, due a side reaction occurring because of the high exothermicity of the main reaction.

Zn  **SO_2** 

CHEMICAL REACTION EQUATION

$$8\ Zn_{(s)} + S_{8(l)} \longrightarrow 8\ ZnS_{(l)}$$

$$S_{8(l)} + 8\ O_{2(g)} \longrightarrow 8\ SO_{2(g)}$$

CONCLUSIONS

Zinc is a very reactive metal, especially when is finely powdered. This is the reason why it reacts violently with sulfur, another reactive element, in a combination reaction. Zinc is losing its 4s electrons in favor of the nonmetallic sulfur generating an ionic combination: the zinc sulfide ZnS which is left as a yellow residue in the crucible. The reaction is highly exothermic and generates a large flame and a huge amount of smoke along with the previously mentioned side compounds. The violence of the reaction increases with the decreasing size of the zinc and sulfur particles. Due to this energetic process this reaction was formerly used as model rocket propellant [3]. If larger quantities are used, in the same mass ratio, but placed in several successive cones, linked together by narrow stripes of mixture, once ignited at one end the reaction will continue successively through all cones [4].

WASTE MANAGEMENT

Zinc sulfide is collected in containers for less toxic metals. Avoid the exposure of ZnS to acids or oxidizers as it can generate H_2S, a very toxic gas [5].

PHOTOS

Fig. 1 The Zn-S mixture (1.0g/2.0g) before initiating the reaction.

Fig. 2 Initiation of the exothermic reaction between Zn and S with SO_2 evolvement

Fig. 3 Incandescent reaction site for a process lasting a few seconds.

Fig. 4 The white end product, ZnS known also as white pigment is yellow when still hot.

Fig. 5 The reaction between zinc and sulfur

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THE REACTION BETWEEN ALUMINIUM AND IODINE

OVERVIEW

The reaction between aluminium and iodine is a typical combination reaction with multiple pedagogical uses. A very versatile redox reaction it can demonstrate the strong reducing character of the powdered aluminium, the strong oxidant character of the iodine, or the catalytic properties of water used to initiate the reaction. It's spectacular due to the abundant violet clouds of iodine generated during the process leading to discussions about sublimation and exothermic chemical processes.

MATERIALS AND METHODS

The reaction has to be performed in a fume cupboard on a fire resistant support. The reaction is not done with stoichiometrically weighed reagents as for 0.5 g of aluminium 7.0 g of iodine are necessary. Instead take equal amounts (in volume) of aluminium and iodine (e.g. 0.5 g Al powder and 2.0 g of grinded I_2) mix them carefully, place the mixture, as a cone, on a fire resistant support, ensure the ventilation of the fume cupboard is on and start the reaction by placing a few drops of hot water in the middle of the cone.

HAZARDS

Aluminium powder is a highly reactive chemical and being a finely powdered product is flammable and in contact with water releases hydrogen which is also very flammable (explosive) [2].

Iodine is a dangerous chemical, it's harmful if inhaled or swallowed, it causes skin, respiratory and eye irritation, is dangerous for the environment, toxic to aquatic life and may be corrosive to metals [3].

During the reaction abundant iodine vapors are released therefore the reaction has to be done in a fume cupboard or in open space. Dispose properly the formed aluminium iodide.

Al

I_2

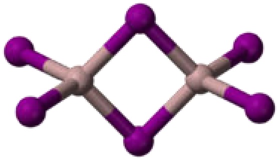
AlI_3 it can generate hydroiodic acid fumes HI when interacting with water

CHEMICAL REACTION EQUATION

$2\,Al_{(s)} + 3\,I_{2(s)} \text{ ---> } 2\,AlI_{3(s)}$ (a dimer is formed)

CONCLUSIONS

Even if not as reactive as bromine and chlorine, iodine acts also as an oxidant. When mixing with aluminium (especially when finely powdered) it reacts violently when a few drops of hot water are added. The reaction is strongly exothermic which leads to the sublimation of iodine generating the abundant violet clouds at the beginning of the process. Later the mixture burst in flame and the white smoke appearing is due to the formation of aluminium iodide which is also left on the reaction site as a glowing solid [4]. This is a typical combination redox reaction, electrons being transferred from aluminium (as a metal it's an electron donor) to the nonmetallic iodine which accepts them leading to aluminium iodide which often is represented as a chemical compound with a strong covalent character [5].



WASTE MANAGEMENT

The anhydrous aluminium iodide reacts with water releasing hydroiodic acid which is a strong acid. Therefore the reaction product has to be handled with care. To dispose the waste mix the reaction product, in several steps, in small amounts, with a 1M solution of sodium carbonate (Na_2CO_3). The resulted slurry can be disposed in the drain with plenty of water[4] or filtered off, the filtrate can be washed in the drain and the solid kept in a container for less toxic inorganic waste [6].

PHOTOS

Fig. 1-3 The initiation with hot water of the chemical reaction between 0.5 g powdered Al and 2.0 g I_2. After a short period of time the reaction starts and due to the evolved heat, abundant violet iodine vapors are formed.

Fig. 4-6 The reaction site becomes incandescent especially when excess of metal is used.

Fig. 7 Reaction between 1.0 g of Al and 4.0 g of I_2.

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THE REACTION BETWEEN MAGNESIUM AND IODINE

OVERVIEW

The reaction between magnesium and iodine is a typical combination reaction as it is the reaction presented in the previous experiment. It's a redox reaction, it can be used to demonstrate a highly exothermic reaction, catalysis can be taken in account if speaking about the role of water, the sublimation process can be easily demonstrated in a spectacular way.

This is another example demonstrating the strong reducing tendency of magnesium and the strong oxidant character of a halogen, in this case the iodine.

MATERIALS AND METHODS

Mix carefully equal volumes of magnesium powder and grinded iodine (e.g. 2.0 g of Mg and 4.0 g of I_2) and form with the mixture a cone in a ventilated hood. "Ignite" the reaction mixture using hot water, dropped from a Pasteur pipette right in the middle of the mixture. Other authors suggest the use of $Mg:I_2$ în the ratio 1:6. [1]

As this reaction is strongly exothermic, like the similar reaction described in experiment *The reaction between aluminium and iodine*, it has to be performed on a fire resistant support, associated with a good ventilation ensuring the toxic vapors elimination during the process.

HAZARDS

Magnesium powder is a highly reactive chemical and as a fine powder, is highly flammable and in contact with water can release hydrogen which is an explosive gas [2].

Iodine is a dangerous chemical, it's harmful if inhaled or swallowed, it causes skin, respiratory and eye irritation, is dangerous for the environment, toxic to aquatic life and may be corrosive to metals [3]. During the reaction abundant iodine vapors are released therefore the reaction has to be done in a fume cupboard. Dispose properly the formed magnesium iodide.

Mg **I_2**

MgI_2 is not classified as dangerous according to GHS; is irritant to eye and skin , causes skin sensitization and eye damage.

CHEMICAL REACTION EQUATION

$$Mg_{(s)} + I_{2(s)} \text{--->} MgI_{2(s)}$$

CONCLUSIONS

Both elements are highly reactive and they combine violently as the reaction is initiated. Magnesium is giving its 2 outermost electrons being transformed into a 2+ ion. The donated electrons are accepted by the iodine which turns into a negative -1 ion. Thus a ionic compound is formed: the magnesium iodide MgI_2. The water dropped on the mixture, at the beginning, acts as a catalyst allowing the reaction to start.

The abundant violet cloud, formed during the reaction demonstrates, once again, the strong tendency of the iodine to sublimate, generating violet vapors.

Slight excess of metal generates a more violent and exothermic reaction.

WASTE MANAGEMENT

Magnesium iodide is not a dangerous substance according to GHS. Despite this the reaction product is not disposable in the drain [4]. Dispose it in a container for collecting less toxic inorganic waste.

PHOTOS

Fig. 1 Mixture of Mg and I_2 prepared carefully by mixing both reagents on a sheet of paper.

Fig. 2 Rapid reaction initiation using few drops of hot water. The volume of the iodine cloud is directly proportional with the amount of water used.

Fig. 3 The I_2 sublimates due to the heat evolved on the reaction site.

Fig. 4 Reaction core still glowing at the end of the reaction.

Fig. 5 The reaction of magnesium with iodine

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COMBUSTION OF HYDROGEN (SYRINGE, MICROSCALE)

OVERVIEW

This experiment demonstrates that some reactions need energy to start called activation energy. In a first step, gaseous oxygen and hydrogen are produced separately and then mixed as oxyhydrogen. This gas will combust when brought to the autoignition temperature. The lighter (igniter) starts the violent reaction. From this reaction, water results as a more stable product. [1]

MATERIALS AND METHODS

For the experiment 3 stands, 2 small syringes, 4 cannulas, one strong/bigger syringe, 2 test tubes, 2 press-in test tube adapters have to be prepared. The synthesis of hydrogen is carried according to Fig. 1: Place a piece of granular zinc into a test tube. Close with a plug and insert two cannulas. Add 1 mL of hydrochloric acid (37 % w/w) using a small syringe with plunger. Wait for a minute, then draw 10 mL of hydrogen from the test tube using the second cannula and syringe. For the synthesis of oxygen, proceed as shown in Fig. 2: Place a small piece of granular MnO_2 into the second test tube. Close with a plug and insert two cannulas. Add 1 mL of hydrogen peroxide (30 % w/w) using the second syringe with plunger. Wait for a minute, then draw 5 mL of oxygen out of the tube using the second cannula and syringe. The reaction of hydrogen and oxygen is carried out according to Fig. 3: Push the contents of both small syringes in one strong syringe with connected ignition adapter. Start the reaction by applying heat to the oxygen and hydrogen in the syringe, wait for water filling up the syringe.

HAZARDS

Hydrogen and zinc are flammable. Oxygen, MnO_2 and hydrogen peroxide are strong oxidizing agents. Hydrogen peroxide and hydrochloride acid are highly corrosive. Wear safety gloves and glasses.

H_2		**O_2**	
Zn		**HCl_2**	
MnO_2		**H_2O_2**	

CHEMICAL REACTION EQUATION

1. Synthesis of Hydrogen:
   $Zn_{(s)} + 2\,HCl_{(l)} \text{--->} ZnCl_{2(s)} + H_{2(g)}$
2. Synthesis of Oxygen:
   $H_2O_{2(l)} \text{--->} H_2O_{(l)} + O_{(g)}$
3 Final Reaction:
   $2\,H_{2(g)} + O_{2(s)} \text{--->} 2\,H_2O_{(l)}$

CONCLUSIONS

With the help of activation energy the non polar atomic bonds in the hydrogen and oxygen molecules are split up to their atoms and are allowed to form new polar bonds in the water molecule.

WASTE MANAGEMENT

The products can be disposed of via normal waste water system.

PHOTOS

Fig. 1 Synthesis of hydrogen

Fig. 2 Synthesis of oxygen

Fig. 3 Apparatus for the reaction (H_2/O_2)

REFERENCES

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REACTION OF CHLORINE AND HYDROGEN (PHOTOCHEMICAL)

OVERVIEW

Some reactions need an activation energy as a starting condition. This experiment shows, how a typical activator, the photons of UV-light, can start a reaction. It is remarkable not only because there is a visual/audible effect, it also gives an example of a free radical chain reaction initiated photochemically by the homolytic cleavage of chlorine molecules to give chlorine atoms. A low cost version of the experiment is described below [1, 2].

MATERIALS AND METHODS

Construct a stand with three sleeves, three clamps and three test tubes and place them in an exhaustion hood (apparatus also purchasable as "Gasentwickler-Set" from [3]). Prepare a small syringe (2 ml), two big syringes (20 ml) with easily moving piston and four cannulas. Fill potassium permanganate (about 1 cm high) into the first tube and and zinc into the second. Take two plugs, stick two cannulas into each plug, and close the test tubes. Attach the big syringes to one of the cannulas of each tube.

Producing Chlorine: Fill 2 mL of hydrochloride acid (36 % w/w) into a small syringe, attach it to the test tube filled with potassium permanganate, push it with caution, so that HCl drops slowly into the tube. The escaping chlorine fills the big syringe. When the syringe is filled with 20 mL of gas, let it escape from the syringe (this fraction is still mixed with oxygen), fill once again with 20 ml, let escape again. The remaining gas volume in the tube is filled with chlorine, which will be used later.

Producing Hydrogen: Once again fill 2 mL of hydrochloride acid (36 % w/w) into the small syringe, attach it to the test tube filled with zinc, push it with caution, so that HCl drops slowly into the tube. When the syringe is filled with 20 mL of gas, let it escape from the syringe (this fraction is still mixed with hydrochloric acid), fill once again with 20 ml, let escape again, and finally keep the third fraction with 20 mL of hydrogen gas in the syringe. Cover big syringe with aluminium foil.

Reaction of chlorine and hydrogen: Let 10 mL of hydrogen escape from the second big syringe by pushing the piston downwards(!) and keep exactly 10 ml. Attach the big syringe with hydrogen to the free cannula of the chlorine reaction tube and let it fill with exactly 10 mL of chlorine gas by cautiously drawing the piston and remove it. Now attach a blunt cannula to the syringe and stick it vertically into a cleaning sponge which acts as a mount for the syringe. Place the sponge in this way that above the piston there is free space. Warn the auditorium about hearing a loud plop, then start the reaction by illuminating the syringe with UV light (UV lamp or flash). The piston is catapulted heavily to the ceiling. In the syringe body the generated HCl may be detected with an indicator paper. If the reaction does not work, cover the syringe with aluminium foil. Now it should work.

HAZARDS

This experiment is dangerous and should be attempted only by a well-trained person. The mixture of chlorine and hydrogen must be prepared in a darkened room. The explosion of a hydrogen-chlorine mixture can be triggered not only by bright light but also by a catalyst. Dust, bits of rubber in a rubber tube, *etc.* can sometime act as such catalyst. Hydrogen and zinc are flammable. Chlorine is highly toxic, and can cause lung damage (bronchitis) on inhalation even at low concentrations. Hydrochloric acid is highly corrosive. Potassium permanganate and zinc are harmful to the environment. Use safety glasses and gloves.

H_2

Cl_2

$KMnO_4$

Zn

HCl $_{conc}$

CHEMICAL REACTION EQUATION

$$16\ HCl_{(l)} + 2\ KMnO_{4(s)} \text{--->} 2\ KCl_{(s)} + 2\ MnCl_{2(s)} + 2\ Cl_{2(g)} + 8\ H_2O_{(l)}$$

$$2\ HCl_{(l)} + 2\ Zn_{(s)} \text{--->} 2\ ZnCl_{(s)} + H_{2(l)}$$

$$H_{2(g)} + Cl_{2(g)} \text{--->} 2\ HCl_{(l)}$$

CONCLUSIONS

Hydrogen and Chlorine react only with the activation energy from electromagnetic radiation. Visual light with colours of red, yellow, green does not give a reaction, blue sometimes, and violet and UV light causes a sudden starting.

WASTE MANAGEMENT

There are no solid or liquid reaction products. The remaining chemicals may be diluted with water and poured to the sink.

PHOTOS

Fig.1 Synthesis of hydrogen

Fig.2 Synthesis of chlorine

Fig.3 Syringe for the reaction (H_2/Cl_2)

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HALOGENURE IONS PRECIPITATED WITH SILVER NITRATE

OVERVIEW

In this experiment, the presence of Cl^-, Br^- and I^- in solutions is proved by a precipitation reaction. Depending on the salt the addition of NH_3 will trigger a different reaction. This is a very important method for the qualitative verification of anions, which is used for the analysis of inorganic compounds like minerals and natural water. [1,2,3,4,5]

MATERIALS AND METHODS

Dissolve one spatula tip of NaCl, NaI and NaBr with water in a test tube. Prepare the $AgNO_3$ solution (0,1 M) and add it dropwise to the salt solutions. Due to this, a precipitation is formed. When Ammonia solution (2 M) is added dropwise to the solution, there will be a different reaction for each salt. The precipitate of the NaCl solution will disappear, NaI and NaBr will not disappear.

HAZARDS

Ammonia solution causes irritation of skin and eye damages. Silver nitrate is very fire promoting, corrosive and harmful for the environment. Natrium iodide is harmful for the environment. Wear safety glasses.

NH_3 (2 M)

NaI

$AgNO_3$

CHEMICAL REACTION EQUATION

$$NaCl_{(aq)} + AgNO_{3(aq)} \text{ ---> } AgCl_{(s)} + NaNO_{3(aq)}$$

$$NaBr_{(aq)} + AgNO_{3(aq)} \text{ ---> } AgBr_{(s)} + NaNO_{3(aq)}$$

$$NaI_{(aq)} + AgNO_{3(aq)} \text{ ---> } AgI_{(s)} + NaNO_{3(aq)}$$

CONCLUSIONS

Due to the addition of $AgNO_3$, a precipitation is formed which is different for each salt. AgCl is cheesy white, AgBr is yellowish and AgBr yellow. With the addition of ammonia, the precipitation of the AgCl disappears and the precipitation of AgBr and AgI stays the same. This proves the presence of the ions in the respective salt.

WASTE MANAGEMENT

Dilute the ammonia solution with much water and pour it into the sink. For the disposal of silver nitrate add iron shavings or heat it with glucose. The silver nitrate reduces to silver which is collected in a container for heavy metal wastes. The remaining chemicals can be poured into the sink.

PHOTOS

Fig. 1 AgBr precipitation

Fig. 2 AgI precipitation

Fig. 3 AgCl precipitation

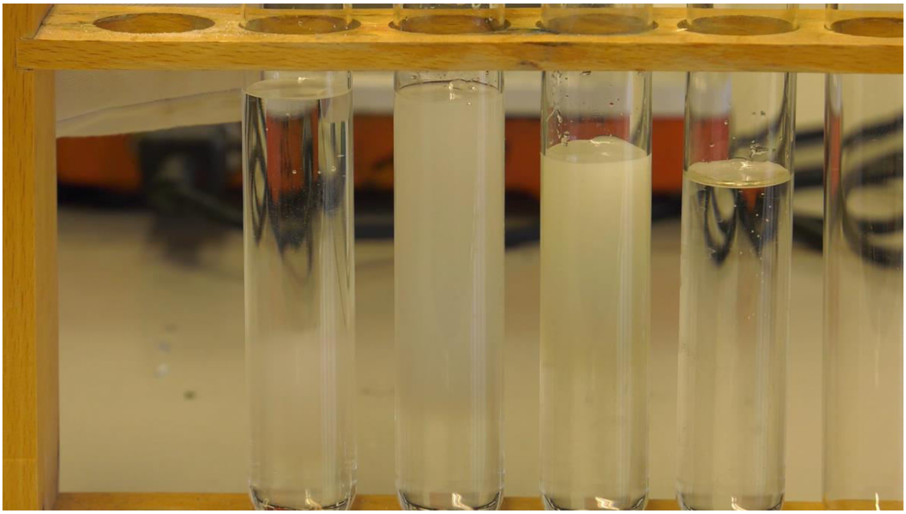

Fig. 4 Adding ammonia to the precipitations (from the left to the right): AgCl is dissolved by ammonia, AgBr and AgI stay the same, comparison with the clear solution of NaCl

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EARTH ALKALINE IONS PRECIPITATED WITH SODIUM SULPHATE

OVERVIEW

This experiment is carried out to observe the precipitation reaction of Calcium, Strontium and Barium Ions with Sodium Sulphate. These reactions are important for the identification of sulphate in a sample such as mineral water, Glauber’s salt etc. [1,2]

MATERIALS AND METHODS

Prepare testing solutions (about 1% w/w) of $BaCl_2$, $CaCl_2$ and $Sr(NO_3)_2$. Put a small amount of the Na_2SO_4 into three test tubes and solve with deionized water. Swirl the mixture and heat it for better solution of the solid compound. Add the testing solutions to the test tube. Observe the reaction.

HAZARDS

Barium chloride is toxic and can be fatal if swallowed or inhaled. It is also fatal when in contact with skin. If swallowed call a poison centre immediately. Calcium chloride causes serious eye irritation, if it comes in contact with the eyes wash thoroughly with water. Strontium nitrate is a strong oxidizing agent.

$BaCl_2$

$CaCl_2$

$Sr(NO_3)_2$

CHEMICAL REACTION EQUATION

$$Na_2SO_{4(aq)} + BaCl_{2(aq)} \text{ ---> } 2\ NaCl_{(aq)} + BaSO_{4(l)}$$

$$Na_2SO_{4(aq)} + CaCl_{2(aq)} \text{ ---> } 2\ NaCl_{(aq)} + CaSO_{4(l)}$$

$$Na_2SO_{4(aq)} + Sr(NO_3)_{2(aq)} \text{ ---> } 2\ NaNO_{3(aq)} + SrSO_{4(l)}$$

CONCLUSION

This experiment shows the precipitation reaction between sodium sulphate and different chlorides and nitrates of the earth-alkaline metals. The results can be seen immediately as the two components mix together and a cloudy precipitation will appear. The reaction between the two components is independent of the sequence they are put together.

WASTE MANAGEMENT

The contents of the test tube with Barium must be disposed of by pouring into the heavy metal waste collecting container. The other reaction products can be poured to the sink.

PHOTOS

Fig. 1 Na_2SO_4 solution in the test tubes

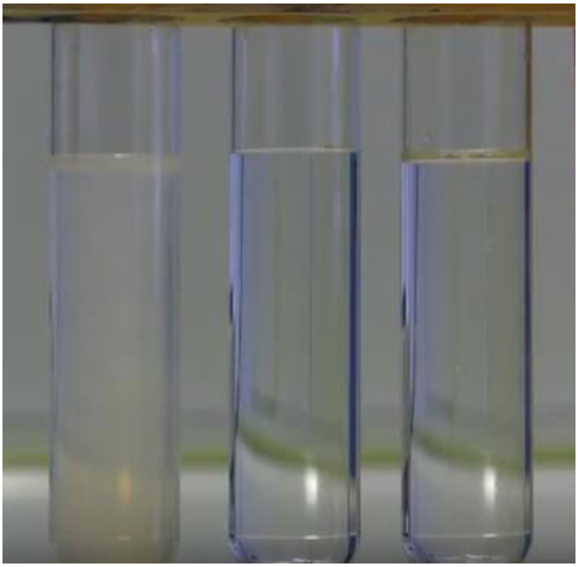

Fig. 2 The result of the precipitation reaction of Na_2SO_4 and the testing solutions (in the following order: Ba/Ca/Sr)

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THE GOLDEN RAIN

OVERVIEW

The Golden Rain experiment is a very versatile experiment for different approaches in the chemistry laboratory. It is a demonstration of a precipitation reaction, of the ionic bonding formation, fosters solubility approaches, it has great analytical value and it might demonstrate the diffusion in water. Basically it deals with a precipitation reaction between a soluble lead salt (nitrate or acetate) $Pb(NO_3)_2$ and potassium iodide KI.

The beauty of this experiment lies in the capacity of the warm saturated lead iodide solution to give spectacular hexagonal shaped yellow crystals as it gets cooled from 80-90°C to room temperature. The crystals formed are giving a "*golden rain*" effect which is best viewed when the cooling flask is placed in the light of an overhead projector. [1]

MATERIALS AND METHODS

First a lead iodide solution is prepared. Take two 250 mL beakers or Erlenmeyer flasks with 150 mL of distilled water each. Dissolve in one beaker 0.5 g of $Pb(NO_3)_2$ ($Pb(CH_3\text{-}COO)_2$ can be also used [2]) and in the second beaker 0.6 g of KI (20% more than stoichiometrically needed). Add a few drops of 1M HNO_3 in the lead salt solution in order to avoid the inherent cloudiness due to hydrolysis or carbonation. Then add carefully the lead nitrate solution to the potassium iodide solution and observe the yellow swirls formed by the amorphous precipitate when the solutions are mixing. In order to observe only the yellow precipitate a much concentrated solution can be used, leading to an abundant precipitate formation, only when adding drops of one solution to the other.

***Golden rain* formation** is the process started with the heating of the yellow suspension formed after mixing both solutions. The mixture is heated to 80-90°C, the yellow color fades and a pale yellow solution is obtained. If this solution shows a certain degree of turbidity few more drops of acid are needed and if there is still some yellow precipitate, quick filtration is needed. Then the hot solution is allowed to cool down giving birth to shiny golden, hexagonal, yellow crystals which generate the impression of an increasingly abundant "*golden rain*". The process can take more than an hour, depending on the cooling rate of the flask so a good timing is needed if the experiment is setup in a Chemistry lesson that lasts less than an hour.

HAZARDS

Attention has to be paid to the lead compounds which are toxic and harmful for the environment. Avoid the lead salts to come in contact with the skin as they are cumullative poisons.

$Pb(NO_3)_2$

PbI_2

CHEMICAL REACTION EQUATION [3]

$Pb(NO_3)_{2(aq)}$ + $2\ KI_{(aq)}$ ---> $PbI_{2(s)}$ + $2\ KNO_{3(aq)}$

CONCLUSIONS

The mixing of a soluble lead nitrate solution and very soluble potassium iodide solution brings together the lead and iodide ions in the same environment. Whereas the lead iodide PbI_2 compound's solubility constant is very low, $K_{SP} = 9.8 \times 10^{-9}$ [4], an abundant golden yellow precipitate is formed in the mixture giving very attractive glittering swirls. This precipitate dissolves and forms a solution at 80-90°C as the PbI_2 solubility is 5 times higher (S=0.294 g/100g H_2O) than at room temperature (S=0.0756 g/100g H_2O) [5]. Because at 90°C we had a saturated solution of PbI_2 crystal formation start if the hot saturated solution is allowed to cool, crystals which are more abundant as the temperature decreases.

WASTE MANAGEMENT

An excess of KI is used during the experiment in order to maximize the chances of the lead ions to precipitate out of the solution in a form of a less soluble lead iodide. After filtration of the PbI_2, the resulted solution is treated with K_2CO_3 or Na_2CO_3 to allow a more complete precipitation of the solvated lead ions. The $PbCO_3$ formed, with a lower solubility constant $K_{SP} = 7.4 \times 10^{-14}$ than the lead iodide PbI_2, $K_{SP} = 9.8 \times 10^{-9}$, ensures an almost complete elimination of the lead ions from the solution. After a second filtration phase the solution can be poured down the drain. The filtered precipitates are treated as toxic solid waste.

PHOTOS

Fig 1. PbI_2 precipitate formation when reacting $Pb(NO_3)_2$ with KI.

Fig. 2 Recrystallizing from a hot, saturated solution of PbI_2.

Fig. 3 Hexagonal PbI_2 crystals falling as *GOLDEN RAIN*

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NH_3 FOUNTAIN

OVERVIEW

Ammonia or azane is a compound of nitrogen and hydrogen with the formula NH_3. The simplest pnictogen hydride, ammonia is a colourless gas with a characteristic pungent smell. When ammonia reacts with water, it forms OH^- and produces a solution in which only a small percentage of ammonia molecules react. NH_3 is used as fertilizer, precursor in nitrogenous compounds, cleaner, in the fermentation industry, as an antimicrobial agent in food products, etc [1], [2].

In this experiment we study the acid-base reaction between ammonia and water.

MATERIALS AND METHODS

In a large beaker place about 500 mL of water and a few drops of phenolphthalein indicator. Place the beaker under a ring stand in which a support ring is already adapted. Put about 30 mL of NH_3 (1 M) in a cone flask and place the flask on a hot plate. Above the cone flask place a round bottom flask which rests on a metal support to collect the vapors of ammonia resulting from heating. When enough ammonia vapors have been collected remove the round bottom flask from the heating. Without delay, place a cap capillary tube to the flask and then place it upside down on the support ring, so that the tube will be immersed in the water and phenolphthalein mixture. The produced ammonia colors the mixture fuchsia. So much gas dissolves in the water by rising up the tube that a partial vacuum is created in the flask. This draws up water into the flask so fast that it sprays like a fountain. Observe the rise of the mixture.

HAZARDS

Wear protective gloves and eye shields. Perform the experiment under an absorber.

NH_3

phenolphthalein

CHEMICAL REACTION EQUATION

$$NH_{3(aq)} + H_2O_{(l)} \rightleftharpoons NH_4^+{}_{(aq)} + OH^-{}_{(g)}$$

CONCLUSIONS

We can observe the ammonia fountain. The gas ammonia has big solubility in the water because of the high polarity, the high dielectric constant and the forming of hydrogen bonds with water molecules. The produced gas ammonia removes a large amount of atmospheric air from the flask. Then the produced ammonia is dissolved in the water and the pressure in the flask is reduced thereby the water enters in the flask as a fountain. The mixture turns into fuchsia because of the basicity of ammonia and the indicator phenolphthalein.

WASTE MANAGEMENT

Neutralize the ammonia solution using hydrochloric acid or diluted sulphuric acid to a pH-value of between 7 and 8 and pour it down the drain.

PHOTOS

Fig. 1 The NH_3 fountain.

Fig. 2 The mixture turns into fuchsia because of the basicity of ammonia and the indicator phenolphthalein.

REFERENCES

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HCl FOUNTAIN

OVERVIEW

Hydrochloric acid is a colourless, highly pungent solution of hydrogen chloride (HCl) in water. It is a corrosive, strong mineral acid with many industrial uses. Hydrochloric acid is found naturally in gastric acid. When it reacts with an organic base it forms a hydrochloride salt. It was historically called acidum salis, muriatic acid, and spirits of salt because it was produced from rock salt and green vitriol (by Basilius Valentinus in the 15th century) and later from the chemically similar common salt and sulfuric acid (by Johann Rudolph Glauber in the 17th century). Free hydrochloric acid was first formally described in the 16th century by Libavius. Later, it was used by chemists such as Glauber, Priestley, and Davy in their scientific research [1].

It is used as solvent in the food industry, steel purification, synthesis of organic and inorganic compounds in the regeneration of ion exchangers etc.

Gastric acid is one of the main secretions of the stomach. It consists mainly of hydrochloric acid and acidifies the stomach content to a pH of 1 to 2. Chloride (Cl^-) and hydrogen (H^+) ions are secreted separately in the stomach fundus region at the top of the stomach by parietal cells of the gastric mucosa into a secretory network called canaliculi before it enters the stomach lumen. Gastric acid acts as a barrier against microorganisms to prevent infections and is important for the digestion of food. Its low pH denatures proteins and thereby makes them susceptible to degradation by digestive enzymes such as pepsin. The low pH also activates the enzyme precursor pepsinogen into the active enzyme pepsin by self-cleavage. After leaving the stomach, the hydrochloric acid of the chyme is neutralized in the duodenum by sodium bicarbonate [1].

In this experiment we study the acid-base reaction between hydrochloric acid and water.

MATERIALS AND METHODS

In a large beaker place about 500ml of water and a few grams of blue bromothymol indicator. Place the beaker under a ring stand in which a support ring has already been adapted. Put about 30 mL of HCl (1 M) in a cone flask and place the flask on a hot plate. Above the cone flask place a round bottom flask which rests on a metal support to collect the vapors of HCl resulting from heating. When enough HCl vapors have been collected remove the round bottom flask from the heating. Without delay, place a capillary tube to the flask and then place it upside down on the support ring, so that the tube will be immersed in the water and blue bromothymol mixture. The HCl colors the mixture yellow. So much gas dissolves in the water by rising up the tube that a partial vacuum is created in the flask. This draws up water into the flask so fast that it sprays like a fountain. Observe the rise of the mixture.

HAZARDS

Wear protective gloves and eye shields. Perform the experiment under an absorber.

HCl

blue bromothymol

CHEMICAL REACTION EQUATION

$HCl_{(aq)} + H_2O_{(l)} \rightarrow Cl^{-}_{(aq)} + H_3O^{+}_{(aq)}$

CONCLUSIONS

The produced hydrochloric gas stains the mixture yellow. So much gas dissolves in the water rising up the tube that a partial vacuum is created in the flask. This draws up water into the flask so fast that it sprays like a fountain. The gas hydrochloric acid is very soluble in the water because of the high polarity. The produced gas hydrochloric acid removes a large amount of atmospheric air from the flask. Then the produced hydrochloric acid is dissolved in the water and the pressure in the flask is reduced thereby the water enters in the flask as a fountain. The mixture turns yellow because of the acidity of hydrochloric acid and the indicator bromothymol blue.

WASTE MANAGEMENT

By diluting hydrochloric acid to 5% v/v or less, and neutralizing it with $NaHCO_3$ to a neutral pH, the solution is no longer corrosive and can, in some jurisdictions, be disposed down the drain.

PHOTOS

Fig. 1 The HCl fountain.

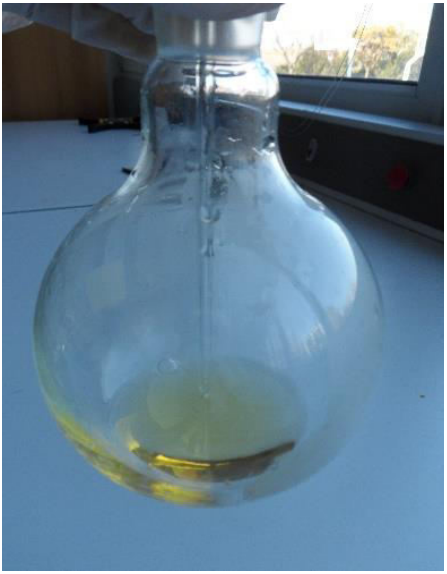

Fig. 2 The mixture turns yellow because of the acidity of HCl and the indicator blue bromothymol.

REFERENCES

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WATER QUALITY TESTING BY USING NITRATE/NITRITE TEST STRIPS

OVERVIEW

Water is an important utility in every household not to mention different economic sectors like agriculture and industry or energy generation processes. Out of the total of water consumed in Europe 20% is coming from public water supply system and form this only 1% is used as drinking water and 6% is used in food preparation [1]. The sources of water supply for municipal waters are different and their ratio in the total water supply varies from country to country. Many countries rely on groundwater sources combined with surface water, springs and bank water. Generally in the European countries' ground water covers around 60-65% of the public water sources [2] but there are still many household relying on their private water sources as wells or springs [3].

One of the biggest concerns of the European Union related to water quality is the presence of nitrates NO_3^- (and worse of nitrites NO_2^-) in drinking water. This ion is present in the ground water as a byproduct of biological decay of plants and animals, and it can increase in water due to extensive use of manure or due to excessive use of fertilizers. Nitrates in excess in the drinking water can have adverse effect especially in infants less than six months of age causing the so called cyanosis or "*methaemoglobinemia*". Therefore it is very important to know the levels of nitrates and nitrites in drinking waters, if they are coming from not verified sources and to avoid consumption if the level is higher than 45 mg/l. A simple, quick and effective method to measure the levels of nitrates and nitrites is to rely on commercially available test strips which come in a great variety of combinations allowing multiple parameter check alongside nitrate levels measurements.

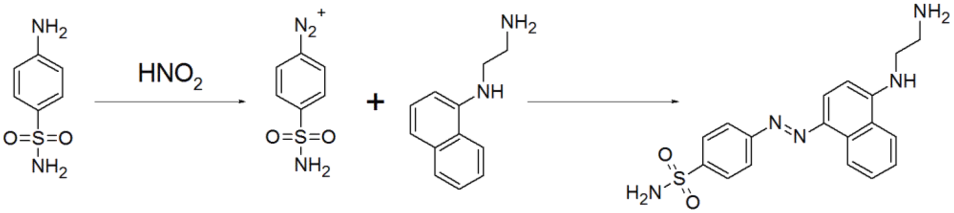
MATERIALS AND METHODS

Take the water sample to be checked for its nitrate/nitrite levels and pour 20 mL in a clean beaker. Take one test strip form the container and close it immediately. Immerse the test strip in the sample solution for 1 second, remove it and gently shake off the excess liquid. Wait 1 minute and compare the color of the reaction zone with the color scale provided on the container.

HAZARDS

No specific hazards are linked to the use of commercially available test strips.

CHEMICAL REACTION EQUATION [4]



CONCLUSIONS

The test strips contain a reaction zone in which different chemicals are absorbed. Once immersed in water a reducing agent reduces the nitrate ions NO_3^- to nitrite NO_2^-. In the presence of an acid buffer, the nitrite is converted to nitrous acid, which diazotizes an aromatic amine. Coupling with N-(1-naphthyl) ethylenediamine produces a red-violet azo dye, the concentration of which can be determined by comparing the colored reaction zone with the provided color scale [5]. One minute waiting time is necessary prior to the final reading. The second reaction zone doesn't contain the reducing agent in the test zone thus allowing the direct determination of nitrite NO_2^- levels.

WASTE MANAGEMENT

The used test strips can be safely discarded in the waste bin.

PHOTOS

Fig. 1 Experimental setup to determine NO^{3-} and NO^{2-} in drinking water.

Fig. 2 Comparing the reactive zone with the scale provided allows semiquantitative measurements

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IDENTIFICATION OF COMBUSTION GASES

OVERVIEW

From the third grade on, chemistry is being taught as a separate discipline. The students are being introduced to basic concepts like atoms and molecules being the building blocks of matter and how chemical reactions proceed, regarding them as a rearrangement of atoms present in the reactants into new molecules that form the reaction products. Furthermore they become aware of the accompanying energy effects and environmental aspects.

Combustion processes provide a proper context to illustrate these phenomena. Liquid and/or solid fuels are converted to gaseous products while releasing huge amounts of energy. Every atom in the fuel reacts with oxygen to yield its corresponding flue gas. So, the presence of a certain compound in the flue gas is a direct proof of the corresponding atom being present in the fuel. The composition of the flue gas can be monitored with so-called reagents; i.e. substances with a specific reaction with one of the flue gas compounds.

MATERIALS AND METHODS

Requirements: Gas washing bottles, air pump, hoses, funnel, lime water, bromine water, copper(II)sulphate.

Procedure: The core of the experimental equipment to be used is the "bubbler" as shown in figure 1.

The gas flow to be investigated is inserted in the left tube. The bubbler is filled with reactant to a level that allows intensive contact between the flue gas bubbles and the reagent. The liquid level however should not reach the exhaust tube so that after passing through the reactant the gas can leave the bubbler. More bubblers with different reagents can be arranged in series if required.

Most common gases to be detected in this way are:

(i) Carbon dioxide; a solution of calcium hydroxide will get turbid due to precipitation of calcium carbonate;

(ii) Sulphur dioxide; a brownish solution of bromine water will decolorize upon reaction with sulphur dioxide.

(iii) Water vapour can be detected with copper(II) sulphate.

HAZARDS [1,2]

Br_2 **$Ca(OH)_2$**

CHEMICAL REACTION EQUATIONS

$Ca^{2+}{}_{(aq)} + CO_{2(aq)} + 2OH^{-}{}_{(aq)} \rightarrow CaCO_{3(s)} + H_2O_{(l)}$

$SO_{2(aq)} + Br_{2(aq)} + 2\ H_2O_{(l)} \rightarrow SO_4^{2-}{}_{(aq)} + 2Br^{-}{}_{(aq)} + 4H^{+}{}_{(aq)}$

$CuSO_{4(s)} + 5\ H_2O_{(l)} \rightarrow CuSO_4.\ 5\ H_2O_{(s)}$

CONCLUSIONS

In this experiment qualitative analytical techniques to detect flue gases are presented. Students are made familiar with reagents and how to use them. Furthermore they learn how their observations can be translated to conclusions with respect to the composition of the fuel and on environmental impact.

WASTE MANAGEMENT

Lime water and bromine solution can be treated as waste water. Flue gases are emitted in the environmental air.

PHOTOS

Fig. 1 Bubbler.

Fig. 2 The set up

Fig. 3 SO_2 and CO_2 detection.

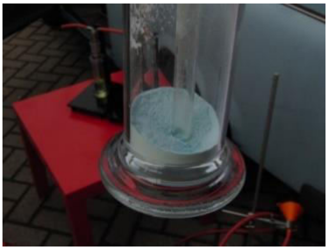

Fig. 4 Water vapour detection.

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WHAT PRODUCTS DO WE BREATHE OUT?

OVERVIEW

Already in the first grade students learn that our energy source is glucoses. Glucoses are being processed with oxygen in a complex scheme of reactions, known as the Krebs- or citric acid cycle to yield carbon dioxide and water. Our daily beverage provides the glucose and the reaction product leaves our body through our lungs (reaction 1). In this experiment students will notice that indeed our breath contains much more water vapour and carbon dioxide than environmental air does.

MATERIALS AND METHODS

Requirements:

Gas washing bottles, lime water, bromine water, copper(II)sulphate.

Procedure: The core of the experimental equipment to be used is the "bubbler" as shown in figure 1. We use three bubblers in series:

(i) The first one, with some with dry white copper(II) sulphate at the bottom in order to detect water vapour (reaction 2);

(ii) A second one with lime water, in order to detect CO_2 (reaction 3);

(iii) A third one with bromine water, to detect SO_2 (reaction 4).

The sequence is important because the air should not be saturated with water vapour prior to subjection to copper(II)sulphate.

We start the experiment by leading air through the bubblers. Because environmental air contains only small amounts of water vapour and carbon dioxide, it will take quite some time to absorb detectable amounts of the indicated substances. The air flow is inserted in the left tube. The bubblers are filled with reagent to a level that allows intensive contact between the flue gas bubbles and the reagent. The liquid level however should not reach the exhaust tube so that after passing through the reagent the gas can leave the bubbler.

Secondly we subject our breath to analysis. Breathe in deeply and breathe out gently in a small hose connected to the left tube. It may be necessary to repeat a couple of times. The blue colouring of the copper(II) sulphate and turbidity of the calcine solution however will be quickly noticeable. We expect no reaction with the bromine water to occur because our breath should not contain sulphur dioxide.

HAZARDS [1,2,3]

$CuSO_4$

$Ca(OH)_2$

Br_2

CHEMICAL REACTION EQUATIONS

$$C_6H_{12}O_{6(aq)} + 6O_{2(aq)} \rightarrow 6CO_{2(aq)} + 6H_2O_{(l)} \qquad (1)$$

$$CuSO_{4(s)} + 5H_2O_{(l)} \rightarrow CuSO_4.5H_2O_{(s)} \qquad (2)$$

$$Ca^{2+}{}_{(aq)} + CO_{2(aq)} + 2OH^{-}{}_{(aq)} \rightarrow CaCO_{3(s)} + H_2O_{(l)} \qquad (3)$$

$$SO_{2(aq)} + Br_{2(aq)} + 2\,H_2O_{(l)} \rightarrow SO_4^{2-}{}_{(aq)} + 2Br^{-}{}_{(aq)} + 4H^{+}{}_{(aq)} \qquad (4)$$

CONCLUSIONS

In this experiment we have shown that the air we breathe out contains detectable amounts of water vapour and carbon dioxide. As was to be expected, no sulphur dioxide was detectable.

WASTE MANAGEMENT

Lime water can be treated as waste water. Flue gases are emitted in the environmental air. Copper(II)sulphate and bromine solution have to be collected and handed over for professional waste treatment.

PHOTOS

Fig. 1 Bubbler

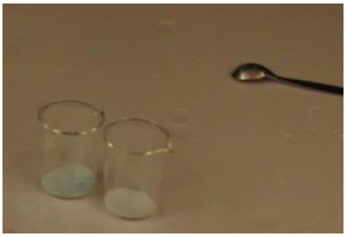

Fig. 2 Water detection

Fig. 3 CO_2 detection and no SO_2

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COOLING MIXTURE

OVERVIEW

The following experiment shows the reaction between barium hydroxide and ammonium thiocyanate. The observed endothermic process in this reaction makes the experiment quite exciting. By draining energy from the surrounding, an enormous temperature reduction of about 30-40 °C is achieved. The described reaction is a good example for showing the properties of cooling mixtures used for industrial and health uses. [1,2,3]

MATERIALS AND METHODS

40 g barium hydroxide and 21 g ammonium thiocyanate are generously mixed together in a 250 mL Flask. To illustrate the reaction, take a wet wooden plate and place the flask on top of it. The bottom part of the beaker will freeze together with the plate. With a thermometer, the temperature reduction (from room temperature to -20 °C) during the reaction can be trailed

HAZARDS

Barium hydroxide is corrosive. Wear safety goggles while working with this substance. The ascending gases irritate eyes and the respiratory system. [1,2]

$Ba(OH)_2$

NH_4SCN

CHEMICAL REACTION EQUATION

$$Ba(OH)_2 \cdot 8H_2O_{(s)} + 2NH_4SCN_{(s)} \text{--->} Ba(SCN)_{2(s)} + 2NH_{3(g)} + H_2O_{(l)}$$

CONCLUSION

Barium hydroxide and ammonium thiocyanate react to barium thiocyanate and gaseous ammonia. As a by-product water is generated from the reaction. The generation of the gas can be smelled, and the formation of a liquid from the solids may be observed. The extreme endothermic reaction occurs, because the splitting of the ionic bonds consumes much energy. Nevertheless, the reaction is driven by a strong increase in entropy, as two ionic solid compounds are converted to a liquid and gaseous component.

WASTE MANAGEMENT

Do not pour the produced slury into the sink, but collect it in a container for heavy metal waste.

PHOTOS

Fig. 1 Reaction

Fig. 2 Solid NH_4SCN

Fig. 3 Reaching -10°C

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ENDOTHERMIC REACTIONS

OVERVIEW

One of the most common questions asked by the teacher when introducing "*Thermochemistry*" is to have some examples of **exothermic reactions**. The examples are flowing according to the students' personal experience. But when examples are asked for **endothermic reactions** there aren't so many answers. This is mostly due to the fact that they aren't that common and there are few examples known when a chemical reaction is performed to lower the temperature.

The most common answers emerging from this question are: synthesis of glucose from CO_2 and H_2O in the leaves; ATP formation in the cell; decomposition of carbonates etc. Obviously there are many more examples.

A very versatile chemical reaction, used sometimes with spectacular effects, is the reaction between barium hydroxide and ammonium salts, the most used being the ammonium thiocyanate. If the reagents are fresh the temperature drop is under 0°C which can be demonstrated in several ways.

MATERIALS AND METHODS

Place a 150 mL beaker on a light wooden plate. Smear some water on the surface of the wood where you are going to place the beaker. Weigh separately 32 g of barium hydroxide octahydrate $Ba(OH)_2*8H_2O$ and 10 g of ammonium chloride NH_4Cl.[1] Measure with a digital thermometer the temperature of the reagents. Place the two solids in the beaker and mix them with a glass rod. Notice that the mixture becomes wet and slushy and observe the evolvement of ammonia. Because of this the reaction should be performed in a hood or in a ventilated laboratory. The temperature drops dramatically to about -20°C (depending also on the granulation of both reactants). As it reaches a negative value raise the beaker from the table. Observe that the wooden plate is frozen to the beaker.

HAZARDS [2]

Barium hydroxide is caustic and irritant whereas the evolved ammonia is a poisonous gas.

NH_4Cl

$Ba(OH)_2$

NH_3

CHEMICAL REACTION EQUATION

$Ba(OH)_2 \cdot 8H_2O_{(s)} + 2NH_4Cl_{(s)}$ ---> $BaCl_{2(s)} + 2NH_{3(g)} + H_2O_{(l)}$

CONCLUSIONS

For the reaction: barium hydroxide and ammonium chloride an enthalpy change of $\Delta H = +164\ kJ \cdot mol^{-1}$ is obtained if the product is assumed to be $BaCl_2$. The entropy change for the system has a positive value because a gas NH_3 and a liquid H_2O are formed from the two solids. The actual entropy change for the system and the surroundings is $\Delta S_{total} = +591\ J \cdot mol^{-1} \cdot K^{-1}$ leading to a negative ΔG calculation, hence the reaction is spontaneous.

$\Delta G = \Delta H - T * \Delta S = 164000 - 293*591 = -9163\ J \cdot mol^{-1}$.[1] If we envisage the formation of $BaCl_2*2H_2O$ at 25°C we have $\Delta H = +63.6$ kJ/mol, $\Delta S = 368$ J/K respectively $\Delta G = -47.7$ k/j/mol showing again a spontaneous process. [3]

WASTE MANAGEMENT

Discard the reaction products in a container used to collect toxic inorganic waste due to the presence of barium chloride.

PHOTOS

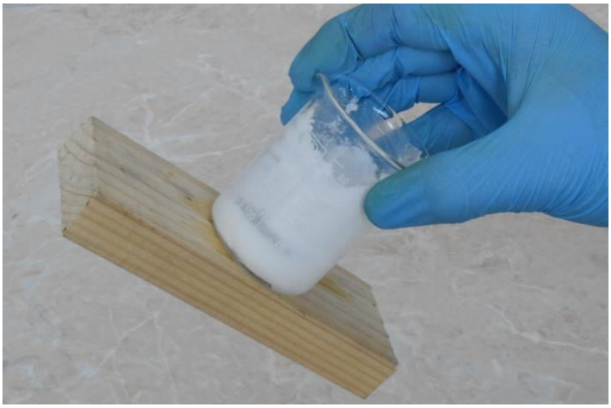

Fig. 1 Barium hydroxide mixed with ammonium chloride is a classical example of endothermic reaction. Spectacular effect is given by the beaker frozen to a moist wooden plate.

Fig. 2 Temperature drop of 37°C from the room temperature 20°C to -17.1°C when barium hydroxide $Ba(OH)_2$ is mixed with ammonium choride NH_4Cl.

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SYNTHESIS OF SCHWEIZER'S REAGENT

OVERVIEW

A much loved reaction by the secondary students is the copper(II) ammine complex synthesis when generally tetraammine diaquacopper(II) dihidroxyde is formed $[Cu(NH_3)_4(H_2O)_2]*(OH)_2$ known under the name of Schweizer's Reagent. The reagent is named after the Swiss chemist Matthias Eduard Schweizer (1818-1860) who studied the properties of this complex.[1]

The synthesis of this reagent can be studied in the class while learning about complexes, while studying the properties of cellulose and sometimes as identification method for copper(II) ions. As it is easy to prepare, except the necessity of using concentrated ammonia when caution is needed, the experiment proves its versatility having in the same time unexpected properties and applications.

MATERIALS AND METHODS

Weigh 3 g of $CuSO_4*5H_2O$ place it in a 150 mL beaker and add water until the solution reaches 50 g. Hence you will obtain a solution having 3.84% concentration. Add 25 mL of 1.0M NaOH solution. The abundant sky blue precipitate either filtered off or is treated in the mother liquor with concentrated ammonia solution (NH_3 25%) until complete dissolution of the precipitate and obtaining a clear deep blue solution, the Schweizer's Reagent.

HAZARDS

Strong bases are caustic. Handle them with care! Concentrated ammonia solutions are toxic if swallowed and in contact with skin and fatal when inhaled. Copper sulfate causes irritations and is very toxic for aquatic life. [2]

$CuSO_4$

NaOH

NH_3

CHEMICAL REACTION EQUATION

$$CuSO_{4(aq)} + 2\ NaOH_{(aq)} \text{--->} Cu(OH)_{2(s)} + Na_2SO_{4(aq)}$$

$$Cu(OH)_{2(s)} + 4\ NH_{3(aq)} \text{--->} [Cu(NH_3)_4(H_2O)_2]*(OH)_{2(aq)}$$

CONCLUSIONS

The copper sulfate solution precipitates when a strong base is added generating a light blue copper(II) hydroxide. If this precipitate is filtered off we remove the sulfate ions (which remain in solution during the precipitation process). A further addition of concentrated ammonia to the precipitate dissolves it, the copper hydroxide becoming a water soluble complex, the tetraammine diaquacopper(II) dihidroxyde $[Cu(NH_3)_4(H_2O)_2]*(OH)_2$ which is a deep blue complex known also under the name of Schweizer's Reagent. The sulfate ions should be discarded to avoid the formation of another deep blue complex, the tetraamminecopper(II) sulfate $[Cu(NH_3)_4]SO_4$ [3].

WASTE MANAGEMENT

The Schweizer's reagent can't be stored for long time because the volatile ammonia will leave the complex regenerating the original light blue precipitate which is copper hydroxide $Cu(OH)_2$. If Schweizer's Reagent or copper hydroxide have to be discarded they have to be collected in a container used for less toxic inorganic waste, due to copper toxicity for aquatic life.

PHOTOS

Fig. 1 A 3.84% solution of $CuSO_4$ is the starting point for the Schweizer's Reagent synthesis.

Fig. 2 Abundant precipitate of $Cu(OH)_2$ obtained by the addition of 1M NaOH solution to the copper salt.

Fig. 3 The precipitate dissolves as concentrated ammonia is added to the copper hydroxide.

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COMPLEX FORMATION BALANCE

OVERVIEW

From a solution containing copper(II)ions insoluble basic salts can be precipitated by the action of ammonium hydroxide. The light blue precipitate dissolves in the excess of NH_4OH, forming dark blue $[Cu(NH_3)_4]^{2+}$ complex ions. Reacting copper (II) sulphate with excess ammonium hydroxide results in a dark blue solution, from which $[Cu(NH_3)_4]SO_4 . H_2O$ can be crystallized.

MATERIALS AND METHODS

We will learn how the complex formation in solutions change solubility. The materials used in this experiment are; bottle with a dropper (1) test tube (100mL,1), 0,1M 25mL $CuSO_4$ solution, 1M 50mL NH_3 solution, pipette.

The steps used to conduct this experiment are;

1. Put 10mL of 0,1M $CuSO_4$ solution with the help of a pipette into a test tube.
2. 1ml (20 drops) NH_3 solution is dropped in the test tube with a dropper into the $CuSO_4$ solution carefully.
3. Observe the change and color forming in the solution in the test tube.
4. Continue to drop, bit by bit, NH_3 into the $CuSO_4$ solution in the test tube. For this, nearly 60 drops (3mL) is enough.
5. At this step observe the change and color in the solution in the test tube.

HAZARDS

Ammoniac (NH_3)

CONCLUSIONS

When excessive amount of NH_3 solution is added, $[Cu(NH_3)_4]^{2+}$ complex is formed, thus increasing the solubility of $Cu(OH)_2$.

WASTE MANAGEMENT

Due to $CuSO_4$ being a type of salt it can be washed down the sink. Whereas, ammoniac is to be sealed in a waste container.

PHOTOS

Fig. 1 Put 10mL of 0,1M $CuSO_4$ solution with the help of a pipette into a test tube.

Fig. 2 1mL (20 drops) of NH_3 solution is dropped in the test tube with a dropper into the $CuSO_4$ solution carefully. Continue to drop NH3 bit by bit, into the $CuSO_4$ solution in the test tube.

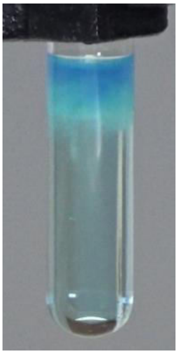

Fig. 3 Finally, observe the change of colour in the solution in the test tube.

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SYNTHESIS OF THE NESSLER'S REAGENT AND NH_4^+ ION DETERMINATION

OVERVIEW

The mercury salts had various applications in the past from curing diseases to photography. The mercury (II) chloride, $HgCl_2$ named also corrosive sublimate was used last centuries to cure syphilis, to wounds disinfection and even to preserve biological specimens [1]. The mercury (II) iodide, HgI_2 found its utility in analytical chemistry being used to prepare the Nessler reagent and it is used to demonstrate thermochromism [2,3]. Nowadays, due to their toxicity, mercuric and mercurous salts are used with care, mostly for analytical purposes (e.g. Nessler solution as ammonia reagent).

The synthesis of Nessler's Reagent is made in 3 steps. The first step includes the insoluble mercury iodide preparation and it's a practical demonstration for precipitation reactions. The second step can demonstrate the high solubility in water of some complexes as potassium tetraioido mercurate (II), obtained from insoluble mercury (II) iodide, being at the same time a spectacular demonstration for the students obtaining different results when mixing two solutions in different amounts. The third step, leading to the Nessler's Reagent, is an alkalinisation step performed with either NaOH or KOH [4].

MATERIALS AND METHODS

Prepare in a 200 mL beaker a saturated solution of $HgCl_2$ at 20°C – aprox 0.23M (**SOLUTION1**) by dissolving 4.4 g $HgCl_2$ in 70 mL of distilled water (the solubility of mercuric chloride is about 6g/100g H_2O at 20°C and increases with temperature to 36 g/100g H_2O at 100°C). Personal observations didn’t confirm the solubility at 20°C. Instead the solution was heated to allow the dissolution of mercuric chloride and then cooled down in order to be able to observe the orange precipitate when adding KI. Otherwise thermochromism of HgI_2 could be observed when precipitating HgI_2 from hot solutions. The **SOLUTION2** to prepare is a saturated KI solution; 11 g KI in the minimum amount of water necessary to dissolve (approx 11 mL of H_2O needed). A third solution, **SOLUTION3,** necessary is 40 mL of a 5N NaOH, obtained by dissolving 20 g of NaOH in 100 mL of solution.

Add drop wise, under constant stirring half of the SOLUTION2 to the SOLUTION1. Observe the changes occurred during the addition. Continue the process by adding the rest of SOLUTION2. Observe the dissolution of the primarily formed precipitate. Stop the addition when a clear solution is obtained (see figures 2-12).

To finish the Nessler's Reagent preparation, add the alkaline solution to the mixture and then dilute it with distilled water to a volume of 200 mL. [5]

Take a 4-5 mL sample of this solution in a test tube and add a drop of ammonium containing solution. Observe the color change.

HAZARDS

Mercury compounds are highly toxic and harmful and all are cumulative poisons. $HgCl_2$ is corrosive and easily sublimates (thus his name as corrosive sublimate). Sodium hydroxide is a highly caustic reagent.[6]

$HgCl_2$ 

HgI_2 

Nessler reagent 

NaOH 

CHEMICAL REACTION EQUATION

For the preparation steps of the Nessler reagent we have the following chemical processes occur.

$$HgCl_{2(aq)} + 2KI_{(aq)} \longrightarrow HgI_{2(s)} + 2KCl_{(aq)}$$

$$HgI_{2(s)} + 2KI_{(aq)} \longrightarrow K_2[HgI_4]_{(aq)}$$

As for the use of the Nessler reagent as an analytical tool to determine the presence of the ammonium ion, the following reaction takes place.

$$2K_2[HgI_4]_{(aq)} + 4KOH_{(aq)} + NH_{4\,(aq)}^{+} \longrightarrow$$

$$\longrightarrow HgO*Hg(NH_2)I_{(s)} + 3H_2O + 7KI_{(aq)} + K^{+}_{(aq)}$$

in higher concentration a BROWN PRECIPATE is formed

$$3HgO*Hg(NH_3)_2I_2 \quad \text{and} \quad NH_2*Hg_2I_3$$

CONCLUSIONS

Once the slightly soluble mercuric chloride (**SOLUTION1**) is treated with potassium iodide (**SOLUTION2**) a precipitation reaction occurs giving an abundant orange precipitate of mercury(II) iodide. This compound is far less soluble than the chloride (<100ppm meaning <0.01g/100mL H_2O).

Continuing the addition of the potassium iodide a new complex is formed: the potassium tetraioido mercurate (II) $K_2[HgI_4]$, very soluble in aqueous solutions so the reaction vessel start to become clear and colorless (yellowish if working with higher concentrations). When no more orange precipitate is noticed the tetraiodo mercurate complex generation is accomplished. At this moment any addition of SOLUTION1 generates the orange mercuric iodide precipitate which can be dissolved by adding the SOLUTION2 to the reaction vessel.

The clear colorless/yellowish solution of tetraioido mercurate (II) $K_2[HgI_4]$ turns to Nessler's Reagent with the addition of the alkaline solution. Once prepared, a sample of this solution reacted with any solution containing ammonium salt gives a yellow-brownish precipitate which is a positive test for the NH_4^+ ion. Being very sensitive this reaction can be used as a spot test using only 0.3 μg NH_3/2μL [6].

WASTE MANAGEMENT

As mercury (II) compounds are very toxic and harmful for the environment all the solutions or suspensions containing mercury ions have to be neutralized with special reagents or collected in containers for toxic inorganic waste.

PHOTOS

Fig. 1 Clear colorless solution of $HgCl_2$ having the volume ~70 mL, the starting reagent of the Nessler solution.

Fig. 2 Insoluble orange HgI_2 start to precipitate when adding to the $HgCl_2$ a saturated solution of KI.

Fig. 3 Complete precipitation is obtained when $HgCl_2$:KI reaches the 1:2 molar ratio

Fig. 4 Further addition of KI dissolves the HgI_2 precipitate giving the water soluble complex $K_2[HgI_4]$

Fig. 5 The solution transferred in a 200 mL beaker after the addition of the 5M NaOH solution (40 mL) is the con-centrated Nessler Reagent.

Fig. 6 The Nessler Reagent ready to be used in the identification process of the ammonium ions (after diluted to 200mL with distilled water).

Fig. 7a and 7b Ammonium chloride NH_4Cl reacted with Nessler Reagent

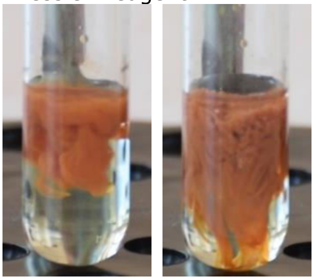

Fig. 8a and 8b Ammonium acetate CH_3COONH_4 reacted with Nessler Reagent

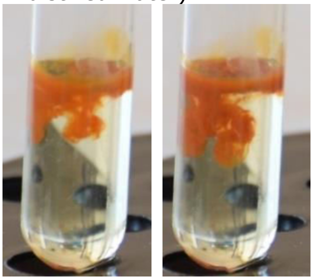

Fig. 9a and 9b Ammonium hydroxide NH_4OH reacted with Nessler Regent

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THERMOCHROMISM OF MERCURY (II) COMPOUNDS

OVERVIEW

Mood rings, aquarium wall thermometers, "*hypercolor*" T-shirts, baby bottles and spoons, battery testers, "*ice cold*" message on beer cans are all developed due to different substances which change color at different temperatures. This property is called ***thermochromism*** [1] and even if this property is not studied at secondary level it has enough practical applications to worth "a word" during Chemistry lessons.

The practical work presented here develops the previous experiment 64 called "*Synthesis of the Nessler Reagent and ammonium ion determination*" oriented toward a special subject: the thermochromism.

MATERIALS AND METHODS

The Nessler's Reagent preparation method presented in experiment nr. 64 is adapted in order to ensure a larger quantity of chemicals isolated and studied for their thermochromism properties.

Insoluble, orange mercury (II) iodide HgI_2 is prepared from a saturated solution of mercury (II) chloride $HgCl_2$ obtained by dissolving 9 g of $HgCl_2$ in 150 mL of distilled water. The resulted solution is then treated with a saturated solution of potassium iodide KI (11 g in the necessary amount of water to be dissolved). An abundant precipitate of mercuric iodide HgI_2 is formed. Split in two parts the resulted suspension and treat it as follows:

- one half is filtered off in a beaker and the orange precipitate HgI_2 is allowed to dry. You will have a certain amount of orange powder of HgI_2.

- the second half is treated with a new amount of KI (5.5 g) until the orange precipitate of HgI_2 is completely dissolved obtaining a clear yellowish solution. Add to this solution 110 mL $AgNO_3$ 0,3N. Observe the abundant, yellow, very fine precipitate formed. Filter off this precipitate and allow it to dry [2]. If the Nessler's Reagent is used, due to its high alkalinity, when we add the silver nitrate we observe the formation of a grey precipitate of Ag_2O, (in parallel with the abundant yellow precipitate) which eventually turns yellow when mixing the mixture. There you will have a certain amount of yellow $Ag_2[HgI_4]$.

Place the two filter papers containing the dried precipitates on a hot plate adjusted to a temperature of 150°C. Observe the changes occurring. Let the two compounds to cool down. Observe again what changes appear.

HAZARDS

Mercury compounds are highly toxic and harmful and all are cumulative poisons. $HgCl_2$, NaOH and $AgNO_3$ are corrosive. The silver salts are very harmful for the environment [3].

$HgCl_2$

HgI_2

Nessler's reagent

NaOH

$AgNO_3$

CHEMICAL REACTION EQUATION

$$HgCl_{2(aq)} + 2KI_{(aq)} \longrightarrow HgI_{2(s)} + 2KCl_{(aq)}$$

red orange $\alpha\,HgI_{2(s)} \xrightarrow{126\ ^{o}C} \beta HgI_{2(s)}$ yellow

$HgI_{2(s)} + 2KI_{(aq)} \longrightarrow K_2[HgI_4]_{(aq)}$ Nessler Reagent

$$K_2[HgI_4]_{(aq)} + 2AgNO_{3(aq)} \longrightarrow Ag_2[HgI_4]_{(s)} + 2KNO_{3(aq)}$$

yellow precipitate

CONCLUSIONS

The mercury (II) iodide HgI_2 is an orange solid at room temperature. When heated over 126°C it becomes bright yellow. If cooled again it regains its orange color.

The silver tetraiodomercurate obtained from potassium tetraiodo mercurate and and silver nitrate, $Ag_2[HgI_4]$ is a yellow precipitate which turns orange-red if heated over 50°C. The color change is reversible and changes back to yellow if cooled back [4].

WASTE MANAGEMENT

As mercury (II) compounds are very toxic and harmful for the environment (as Ag is) all the solutions or suspensions containing mercury (II) and silver (I) ions have to be neutralized with special reagents and collected in containers for toxic inorganic waste.

PHOTOS

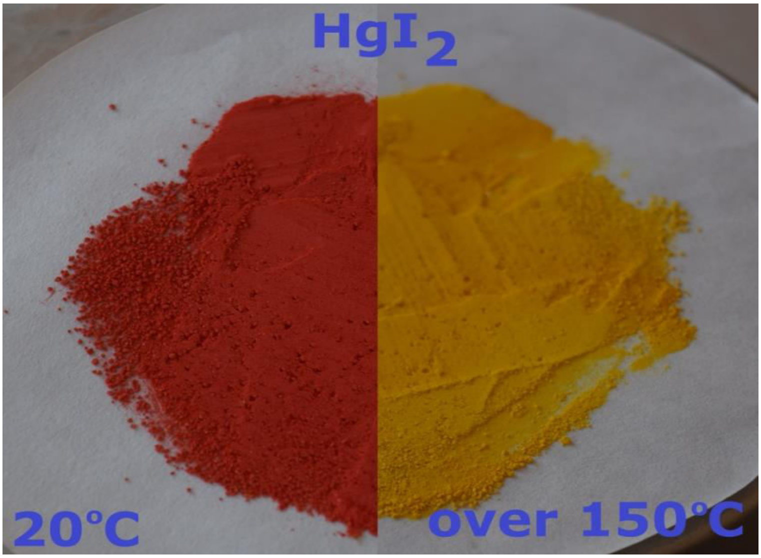

Fig. 1 Thermochromism shown for HgI_2

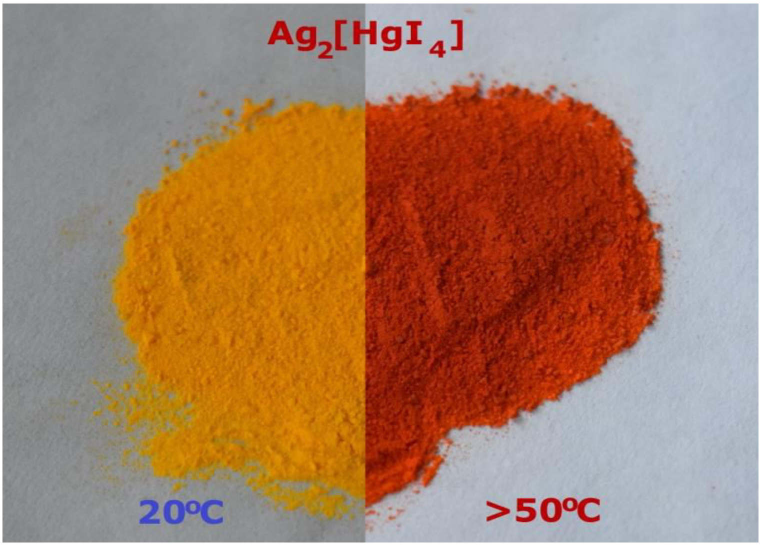

Fig. 2 Thermochromism shown for $Ag_2[HgI_4]$

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ALUMINIUM HYDROXIDE SOLUBILIZATION

OVERVIEW

Aluminium hydroxide $Al(OH)_3$ is a common chemical found in nature, as ores. The most common ore, the bauxite contains $Al(OH)_3$, mixed with $AlO(OH)$ and Al_2O_3 [1]. In chemistry aluminium hydroxide $Al(OH)_3$ is a very good example of an amphoteric chemical combination reacting with both acids and bases. From this property two very important practical applications of this compound emerge:

- in medicine/pharmacy aluminium hydroxide $Al(OH)_3$ is used as a main ingredient in many antacids (drugs used to relieve stomach ache) [2]. Due to its reactivity with acids (like HCl in the stomach) and to its very low alkalinity aluminium hydroxide can be used as medicine.

- the reactivity of aluminium hydroxide $Al(OH)_3$, a low solubility gel like precipitate, with alkaline bases is an example of reduced usability of aluminium dishes in cooking. The water does not attack aluminium, despite its very high oxidation electrochemical potential because aluminium is covered by a resistant layer of aluminium oxide which seals the metal from the water. But if an alkaline base is present (for instance NaOH), the oxide layer is removed, the metal is attacked by water and reacts further, giving at the end a water soluble complex of sodium tetrahydroxo aluminate $Na[Al(OH)_4]$. This is why Al dishes, washed in a very alkaline environment, get spoiled very soon due to holes appeared in the metallic structure.

The capacity of this precipitate to get solubilized when adding sodium hydroxide (NaOH) has also analytical importance. The 3rd analytical group of cations includes Al^{3+} ions which forms insoluble hydroxides even al low concentrations of the hydroxide ion. The addition of sodium hydroxide in excess causes the dissolution of the gelatinous white $Al(OH)_3$ precipitate [3].

The capacity of aluminium hydroxide $Al(OH)_3$ to further react with NaOH giving a soluble compound can be discussed in the classroom either when preparing and discussing coordination compounds, or when studying the properties of aluminium. Also this reaction may be used when having in the classroom analytical approaches discussing the precipitation reactions.

MATERIALS AND METHODS

To demonstrate in the classroom this capacity we need to prepare aqueous Al^{3+} solutions (0,1M). For this we may use solid $Al(NO_3)_3*9H_2O$, $Al_2(SO_4)_3$ or even alum $KAl(SO_4)_2*12H_2O$ or $AlCl_3$. To prepare 50 mL of a 0,1M solution of $Al(NO_3)_3$ you have to weigh on a watch glass, or on a weighing boat, 1.875 g of $Al(NO_3)_3*9H_2O$ (M=375.13), transfer it to a 50 mL volumetric flask, add 30-40 mL of distilled water and agitate the mixture in order to let the salt dissolved. Then slowly add the rest of the water (in the end add drop wise using a Pasteur pipette) until you reach the mark. Proceed accordingly with the other salts to obtain other Al^{3+} solutions.

The precipitation reagent is a 0.2M solution of NaOH. To prepare 100 mL of 0,2M solution weigh on another watch glass or on a weighing boat 0.800 g NaOH (when using NaOH pellets it will be difficult to reach exactly the 0.800 g value because one pellet can weigh between 0.050-0.100 g). As we don't need to create a solution of an exact concentration accept any value around 0.800 g. Try to weigh rapidly in order to keep open the flask as short as possible to avoid interaction of the NaOH with either moisture or CO_2 from air. Put the NaOH in a 100 mL volumetric flask, dissolve it in 70 mL of distilled water and, after complete dissolution, fill in the bottle to the mark with water.

Take 2 mL of 0.1M solution of $Al(NO_3)_3$ in a 16mm diameter test tube and add drop wise with a Pasteur pipette up to 3.0 mL of 0.2M NaOH solution. Shake from time to time the mixture. Observe what happens by adding the basic solution. Continue to add a new amount of NaOH solution and observe the changes occurred. Try again the procedure with a 0.2M solution of $Al(NO_3)_3$ and observe the even more abundant precipitate formed.

HAZARDS

Aluminium salts are irritants, the nitrate is an oxidizer. Sodium hydroxide is corrosive. Pay attention when handling these solutions.

$Al(NO_3)_3*9H_2O$

NaOH

CHEMICAL REACTION EQUATION

$$Al(NO_3)_{3(aq)} + 3\ NaOH_{(aq)} \text{--->} Al(OH)_{3(s)} + 3\ NaNO_{3(aq)}$$

$$Al(OH)_{3(s)} + NaOH_{(aq)} \text{--->} Na[Al(OH)_4]_{(aq)}$$

CONCLUSIONS

An Al^{3+} containing solution is clear and colorless but it's very sensitive to the presence of the hydroxide ions. Having a very low solubility product, $3x10^{-34}$ [3] the $Al(OH)_3$ formed appears as a white gelatinous precipitate. Continuing the addition of the base, until stoechimetric ratio is achieved, more and more precipitate appear. If the NaOH:Al^{3+} 3:1 molar ratio is reached the precipitation formation process stops, further addition of the alkali base causing the precipitate to dissolve and the solution to become clear again, due to a water soluble complex formation: sodium tetrahydroxo aluminate $Na[Al(OH)_4]$.

WASTE MANAGEMENT

After dilution the reaction mixture was considered safe to be poured down the drain [4]. New recommendations require aluminium solutions to be collected in containers for less toxic metallic waste.

PHOTOS

Fig. 1 Clear solution of $Al(NO_3)_3$ 0,1M

Fig. 2 $Al(OH)_3$ precipitate after addition of NaOH

Fig. 3 The precipitate dissolves after further addition of base.

Fig. 4 $Na[Al(OH)_4]$ as soluble complex formed when excess base is used.

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SOLUBILIZATION OF THE $Zn(OH)_2$ PRECIPITATE WITH EXCESS NaOH

OVERVIEW

Most of the zinc compounds are in an ionic state; however, in their organic compounds a covalent property is seen. Zinc ions are generally not simple ions, they are in the form of complex ions.

Some of the metals and semimetals in the periodic table show amphoteric properties. The oxides of these elements carry amphoteric properties as well.

Amphoter: are, those which carry the acidic and basic properties together.

The zinc metal, dissolving in strong bases, releases hydrogen gas and forms the zinc ion.

$\mathbf{Zn + 2\ NaOH + 2\ H_2O \rightarrow Na_2Zn(OH)_4 + H_2}$

When a base is added into zinc salt solutions, whitish and jelly-like zinc hydroxide (is deposited in the bottom). Zinc hydroxide has also an amphoteric property. They release zinc salts in acids, and zinc sulphur in bases.

CHEMICAL REACTION EQUATION

$\mathbf{Zn(OH)_2 + 2\ HCl \ \text{--->}\ ZnCl_2 + 2\ H_2O}$

$\mathbf{Zn(OH)_2 + 2\ NaOH \ \text{--->}\ Na_2ZnO_2 + 2\ H_2O}$

$\mathbf{Zn(OH)_2 + 2\ OH^- \ \text{--->}\ [ZnO_2]^{2-} + 2\ H_2O}$

MATERIALS AND METHODS

Using the glass pipette, a solution of Zinc Nitrate [$Zn(NO_3)_2$] is dropped into the test tube. Sodium hydroxide solution is added drop by drop. We can observe a white jelly-like Zinc Hydroxide(amphoteric). When excessive amount of sodium hydroxide solution is added, we observe that this sediment is dissolved.

HAZARDS

NaOH

$Zn(NO_3)_2$

CONCLUSIONS

Whitish, jelly-like hydroxyl deposit has occurred through the reaction of zinc with bases. The more the base is added, the more the solution dissolved. This shows the occurrence of zinc ion ($[ZnO_2]_2^-$) in the excessive addition of bases. Besides, before copper, silver or chromium covering on alloys, a special cleaning process is applied. For this reason, by immersing/dipping the material into the (Na_2ZnO_2) solution, the covering is made.

WASTE MANAGEMENT

The solution container sodium zinc sulphur salt. Thinning, this solution can be discharged through the tap.

PHOTOS

Fig. 1 The zinc nitrate put in the test tube is a transparent solution.

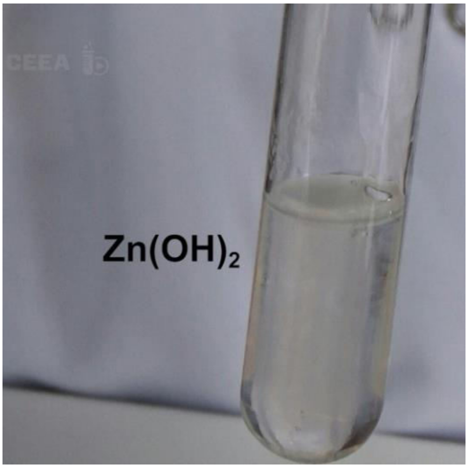

Fig. 2 When drops of sodium hydroxide are added to the test tube a white precipitate formation is observed. This precipitate is zinc hydroxide.

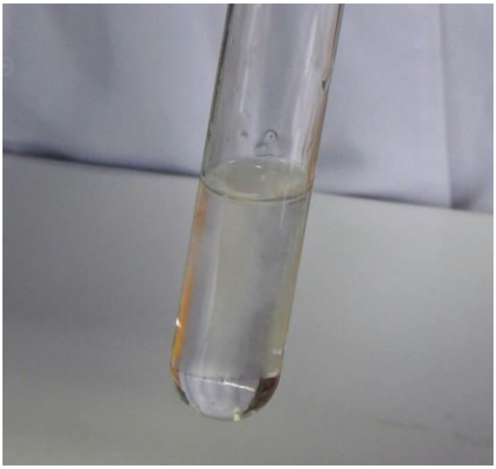

Fig. 3 Continuing the addition of sodium hydroxide the white precipitate dissolves.

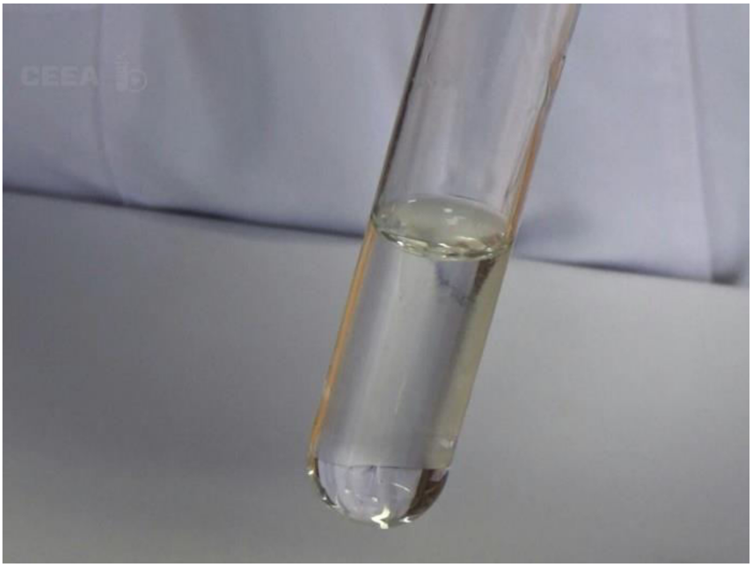

Fig. 4 When excess of sodium hydroxide is added, the solution formed that allows the paste to dissolve is zinc ion $[ZnO_2]_2^-$.

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QUENCHING A FLAME WITH CO_2

OVERVIEW

In this experiment the extinction of a flame with carbon dioxide is shown. CO_2 is produced from a reaction of limestone and hydrochloric acid. Carbon dioxide is being used to extinguish flames by replacing oxygen. Carbon dioxide extinguishers work well on small flammable liquid and electrical fires because Carbon dioxide is heavier than air. It will not cause damage to electrical equipment. [1, 2]

MATERIALS AND METHODS

A little amount of limestone is put in an Erlenmeyer flask which is closed with a plug with two holes inside. A dropping funnel with hydrochloric acid (2 M) is attached to the first hole, curved glass tube to the second. When opening the valve of the dropping funnel, hydrochloric acid drops down and reacts with limestone forming carbon dioxide. Yuo may envisage also the use of a Kipp's apparatus to generate carbon dioxide from the mentioned reagents.

Via the glass tube the gas may flow into a beaker with a lighted candle inside. After some seconds the candle will not burn anymore.

HAZARDS

HCl 

CHEMICAL REACTION EQUATION

$CaCO_{3\,(S)} + 2\,HCl_{\,(l)} \rightarrow CaCl_{2\,(l)} + H_2O_{\,(l)} + CO_{2\,(g)}$

CONCLUSIONS

Carbon dioxide has a higher density than air, so it sinks down to the ground and replaces the breathable air. In this way, the candle does not burn any more because of a lack of oxygen.

WASTE MANAGEMENT

Remaining materials may be diluted with water and poured in the sink after a previous separation of the non reacted solid (lime).

PHOTOS

Fig. 1 Model of a fire extinguisher

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THERMITE EXPERIMENT

OVERVIEW

This experiment demonstrates the exothermic reaction of Thermite (a mixture of aluminium with iron(III)oxide. During the reaction Iron(III)-oxide is reduced to Iron, Aluminium is oxidized. Due to the high temperatures arising during the reaction, iron melts and may be poured out of the reaction vessel. This process is also used today for track construction, e.g. for welding the track construction together. Thermite needs an activating energy of nearly 1500°C to start the reaction. The heat can come from a gas-burner with Propane and Oxygen or a Thermite combustive-stick. For safety precautions, the thermite and the lighter are ignited separately.

Thermite should not be extinguished with water or wet items. Otherwise there is a further redox reaction, in which the water is reduced by the ignoble metal to metal oxide and hydrogen. Because of that water can lead to an explosive reaction where liquid can be expulsed. Also, it could happen that a mixture of hydrogen and oxygen (Oxyhydrogen gas) leads to an explosive reaction. Because of that Thermite must be put on a dry area. [1, 2]

MATERIALS AND METHODS

Materials

- Thermite-mixture purchasable from [3]
- Flower pot of clay (Height: 8,5 cm; upper total width: 8,7cm),
- Standing plate
- Stand rod
- Two stand rings (One has ring a smaller diameter than the aperture of the clay-pot)
- Two stand clips
- Galvanized tin bucket (Minimum capacity of 12 Liter, precisely 30 cm opening diameter, half-filled with dry sand)
- Thermite combustive-stick purchasable from [3]
- Test-tube-holder
- Micro burner
- Coin of aluminium or copper aluminium or copper coin
- Safety glasses

Procedure

Place the stand on a flat-ground. A radius of 2 meters should be cleared of burnable items and materials. The two stand rings are getting attached with the stand in the way that the pot is above the galvanized tin bucket with sand. The stand ring below the pot should be smaller than the bottom of the flower pot, in this way the pot can be placed in the stand ring. The second stand ring surrounds the flower pot on the upper margin and prohibits overturning or a total burst. The bottom of the pot shouldn’t be positioned higher than 40 cm above the galvanized tin bucket. This bucket is filled with sand, and a cavity should be made in the sand for the arising raw iron pouring down. The hole in the ground of the flower pot must be shut with a coin of aluminium or Cu/Al. Fill the pot with the Thermit mixture: 3 cm or maximum 4 cm, about 150 g – 180 g. Regard the safety distance for flammable items and materials (also the Thermite-bin and the Thermite combustive-sticks) around. Set on fire the combustive-stick with a micro or gas burner. The burning stick is thrown into the flower pot. At the same moment, drag back the hand and return back to the safety area.

HAZARDS

There are hazards from burning as well as from hot sputtering material. Gloves are not necessarily needed, protection glasses are necessary. Keep away from the pot till the contents of the galvanized tin bucket does not sputter any more.

Fe_2O_3

Al

CHEMICAL REACTION EQUATION

$\mathbf{Fe_2O_3\,_{(s)} + 2\,Al\,_{(s)} \rightarrow 2\,Fe\,_{(s)} + Al_2O_3\,_{(s)}}$

CONCLUSIONS

After admiring the strongly exothermic reaction, let the reaction mixture in the bucket cool down and observe the still glowing iron material. The raw iron product may be analysed for properties like metallic lustre, magnetizability, density, etc.

WASTE MANAGEMENT

The reaction products may be disposed in the household waste.

PHOTOS

Fig. 1: Building the stand

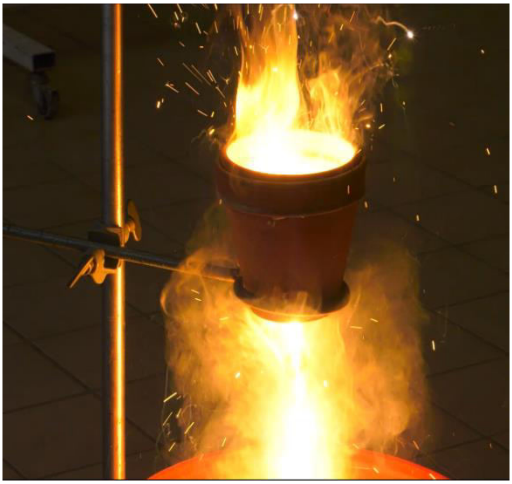

Fig. 2: Thermite reaction

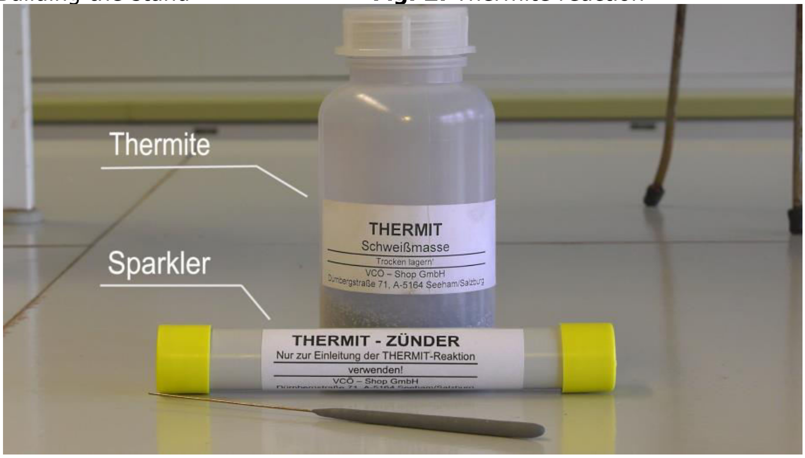

Fig. 3: Thermite reaction mixture and combustive sticks (sparklers)

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SYNTHESIS OF "GOLDEN COINS"

OVERVIEW

Production of gold from other substances was one of the main goals of alchemistry. Many attempts have been made to produce gold from inexpensive materials. Today we know that it is impossible to produce gold from other substances with a chemical reaction. One of the easiest produced material that looks like gold is brass. [1, 2, 3]

MATERIALS AND METHODS

Dissolve 4 g of sodium hydroxide in 50 mL of water in a 100 mL beaker. Pour the sodium hydroxide solution and the 25 g of zinc powder into a porcelain bowl. Place the porcelain bowl onto the wire net above the gas burner. Light the burner, heat the suspension thoroughly and stir well until nearly boiling. Insert a copper coin into the suspension using tongs. After 30 seconds the coin is coated in a silver-colored layer. Lift the bowl with the tongs off the wire net, then lift the coin out of the suspension using the tongs and purify the coin under running water. Dry the coin with a cotton cloth. Heat the Silvery coin in the flame of the Bunsen burner, the colour changes from silver to gold.

HAZARDS

Concentrated sodium hydroxide solution is highly corrosive. Safety glasses and rubber gloves should be worn at all times during the experiment and while handling the solution.

Careful handling is necessary when heating the zinc powder, toxic fumes may be formed. The substance is a strong reducing agent and reacts violently with oxidants. It reacts with water and reacts violently with acids and bases forming flammable/explosive gas.

NaOH **Zinc**

CHEMICAL REACTION EQUATION

$Zn_{(s)} + 2\,H_2O_{(l)} + 2\,NaOH_{(aq)} \rightarrow Na_2Zn(OH)_{4(s)} + H_{2(g)}$

CONCLUSIONS

In strongly alkaline media, zinc reacts to produce hydrogen to sodium tetrahydroxozincate. In the further course of the reaction, zinc deposits on the copper surface. After polishing with a soft cloth, a silvery zinc surface remains on the copper coin. By heating copper and zinc, an alloy is formed (brass), thus a gold like look is achieved.

WASTE MANAGEMENT

Dilute the suspension with water and pour it down to a sink.

PHOTOS

Fig. 1 copper coin, coated coin with Zinc, burned coin to brass

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EXPERIMENTS WITH LIQUID NITROGEN

OVERVIEW

Liquid nitrogen (LN or LN2) is nitrogen in liquid aggregate state, which boils under normal pressure at 77 K (-196 ° C). The dinitrogen molecule consists of two nitrogen atoms connected by a triple bond. It is chemically inert and therefore frequently used as a protective gas filling preventing oxidation processes. Liquid nitrogen is industrially produced in large quantities together with liquid oxygen by distillation of liquid air. In the liquid state it shows some astonishing properties like change of colour of materials treated with liquid nitrogen, volume change, deformability, brittleness etc. [1,2,3]

MATERIALS AND METHODS

For the test of **Change in colour** fill copper sulfate and sulfur 3 cm high in sample tubes. Put the sample tubes in a beaker filled with liquid nitrogen. After a short time, the colors are very bright. For the observation of **volume change** inflate the balloon, tie it up and put the balloon in an empty plastic bowl. Release the liquid nitrogen into the bowl and observe the volume reduction. For experiencing the **brittleness effect**, put a flower in a beaker with liquid nitrogen, take it out and smash with a hammer. Put a sausage in liquid nitrogen, *it can be easily broken*. **Changes in deformability** may be explored by striking with the hammer on a rubber tube at normal atmospheric conditions and after putting the rubber tube into liquid nitrogen.

HAZARDS

Liquid nitrogen is very cold (boiling point -196 °C), there is danger of frostbite. Handle with insulating gloves and protecting glasses.

CONCLUSION

Liquid nitrogen is very cold, so everything that comes into contact with it freezes in a few seconds. Liquids need less space than gas, which is an advantage for storing. *Flexible materials show brittleness and can thus be split up mechanically.* In medicine and biological sciences, *bloodless cell destruction or preservation of cells are areas of application.*

WASTE MANAGEMENT

Do not pour back the liquid nitrogen in the reservoir because of the oxygen enrichment, let it rather evaporate slowly.

PHOTOS

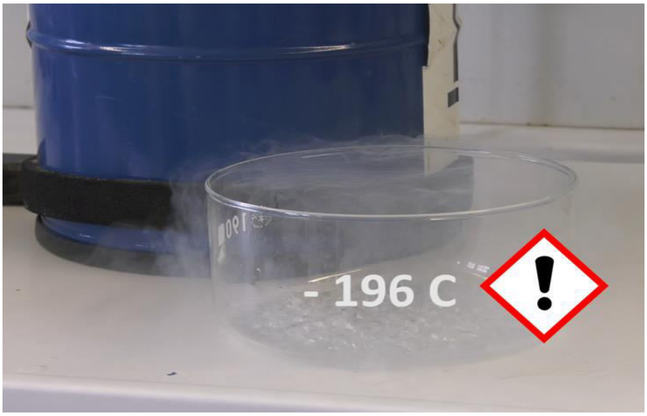

Fig. 1: Evaporation of liquid nitrogen

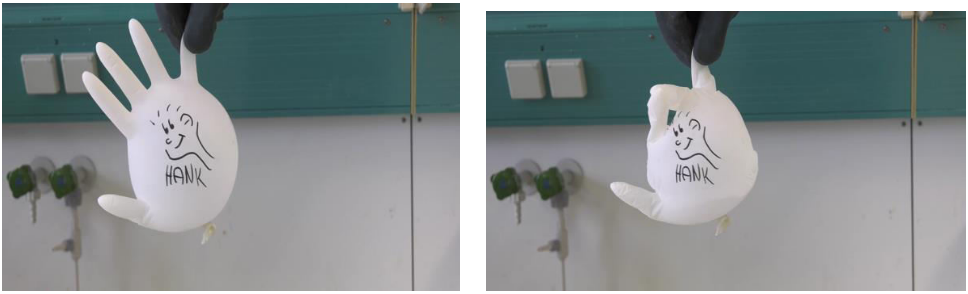

Fig. 2 and **3:** Volume change experiment

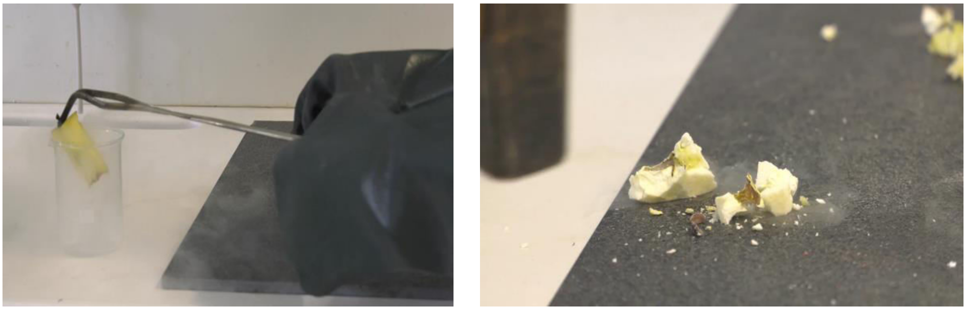

Fig. 4 and **5:** Brittleness effect shown with an apple

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DECOMPOSITION OF NITROGEN TRIIODIDE

OVERVIEW

Nitrogen triiodide is a well-known compound due to its high instability being very sensitive to touch or to an open flame. Due to its low stability it decomposes with a very loud bang, when dry, even if it's touched by a feather.

Even if it was discovered two centuries ago (first reported by Bernard Courtois 1812)[1] it has a very poor commercial value due to its extreme shock sensitivity. Despite this the reaction which leads to nitrogen triiodide is often done in laboratory for showy demonstrations. As a matter of fact chemistry is, for many people not studying it in depth, a science where you can make loud BOOOMS. ☺

MATERIALS AND METHODS

Several procedures are available for the synthesis of nitrogen triiodide. Most of them rely on only small amounts of finely grinded iodine (a pinch [3] or 0.5 g [4]) which is allowed to react, under constant stirring, in about 10 mL of concentrated ammonia solution (30% NH_3). The black reaction product is filtered off after 5-10 minutes of stirring, the solid can be washed with ethanol in order to eliminate all the non reacted iodine and let to get dry in a fume cupboard. While wet, the filter paper containing the nitrogen triiodide, is taken out of the funnel, open up and placed on a heat resistant support. Allow it to get dry while taking care not to have strong air movements in the room and to prevent any other persons to get near the drying filter paper. Some experimental approaches make use of diethyl ether (Caution!!! extremely flammable) [3] to wash the filtered precipitate, in order to allow a much rapid drying of it.

Approach the filter paper only wearing protective cloths and eye protection.

Once dry, touch the filter paper containing the nitrogen triiodide with a feather attached to a 1 m stick. You will notice a big explosion accompanied by an abundant violet cloud.

HAZARDS

Iodine is a dangerous chemical, it's harmful if inhaled or swallowed, it causes skin, respiratory and eye irritation, is dangerous for the environment, toxic to aquatic life and may be corrosive to metals [2].

Concentrated ammonia solutions are toxic if swallowed and in contact with skin and fatal when inhaled.

NH_3

NI_3*NH_3

I_2

Caution The reaction product is a contact explosive. Take all the necessary steps to protect the participants and to prevent other persons to get close to the experiment.

CHEMICAL REACTION EQUATION

The equation of the chemical reaction is not as simple as presented in some sources [5] but it can be simplified as follows;

$$2\ NH_{3(aq)} + 3\ I_{2(s)} \text{--->} NI_3*NH_{3(s)} + 3\ HI_{(aq)}$$

CONCLUSIONS

The nitrogen triiodide is a so called contact explosive. If dry it is decomposing easily, through an explosion, setting free a large amount of gases. It's so unstable when dry that even the touch of a feather can determine a blast. There are several approaches for the chemical reaction of the decomposition process. Of course if abundant violet iodine vapors are formed the first reaction (see below) has to be taken in account. If the cloud has an obvious composition of white dust, providing a more white-reddish cloud, the second process is more likely to happen when white ammonium iodide is present in the fume. Probably the conditions of the reaction, the amount of nitrogen triiodide used, the temperature of the environment are influencing one or the other of decomposition routes.

$$2\ NH_3*NH_{3(s)} \text{--->} 3\ I_{2\,(g)} + 2\ NH_{3(g)} + N_{2(g)}$$

$$8\ NH_3*NH_{3(s)} \text{--->} I_{2\,(g)} + 6\ NH_4I_{(s)} + 5\ N_{2(g)}$$

WASTE MANAGEMENT

Switch on the fume cupboard, right after you set on the explosion (not before), in order to allow the removal of the produced iodine vapors. Rinse the remains of the filter paper with 5M NaOH to ensure the safe eliminations of any nitrogen triiodide. If the reaction place is stained with iodine remove it by washing the place with a 1M solution of $Na_2S_2O_3$ (sodium thiosulfate).

PHOTOS

Fig. 1 Wet precipitate of nitrogen triiodide

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GLASS ENGRAVING WITH HF

OVERVIEW

Hydrofluoric acid is the name given to the solution of hydrogen fluoride a compound in water. It is an acid used in refining high octane petrol and production of synthetic cryolite ($NaAlF_6$), which is especially used in the glass industry. Hydrogen fluorine is a transparent liquid and defined as HF. Hydrofluoric acid is a solution of hydrogen fluoride (HF) in water. It is a precursor to almost all fluorine compounds, including pharmaceuticals such as fluoxetine (Prozac), diverse materials such as PTFE (Teflon), and elemental fluorine itself.

Glass is a liquid formed by the excessively heated alkali, soil alkali and metal oxides with its main substance being silicium (SiO_2). The silicium dioxide in the structure of glass (SiO_2) is the substance giving the feature of glass, forming a web like layer. The Hydrofluoric acid (HF) etches the glass which has silicium dioxide (SiO_2) by the fluorine part of acid to silicium and the hydrogen part to the hydrogen atom.

Due to hydrofluoric acid having a feature of etching a glass surface, it cannot be stored in glass containers. It can only be kept safely in polyethylene plastic containers. When in contact with the human body it does not give pain, but gives harm to the bones through penetrating into the body.

MATERIALS AND METHODS

The paraffin that is to be spread on a piece of retangular glass is pot in a beaker and is melted with the Bain Marie method (which is putting a container beaker within hot water for the substance to melt). The glass is then heated and the melted paraffin is poured and spread all over the surface. After covering the entire surface with paraffin, take the glass away from the burner. Give time for the paraffin to freeze. After its frozen using the steel pencil, anything can be written or drawn on the paraffin. After this add hydrofluorıc acid in porcelain kroze. Using a syringe, drip the hydrofluorıc acid on the scratched drawing or writing, wait for a while and clear all the paraffin off the glass.

HAZARDS

Hydrofluoric acid (HF)

CHEMICAL REACTION EQUATION

$SiO_2 + 4HF \rightarrow SiF_4 + 2H_2O$

$SiF_4 + 2HF \rightarrow H_2SiF_6$

CONCLUSIONS

The paraffin layer hinders the reaction of HF acid with glass. The section of glass that the hydrofluoric acid is in contact with reacts with the silicium dioxide in glass and erodes the web like structure of it.

WASTE MANAGEMENT

The hydrogen fluorine, treated with 1 mole of calcium carbonate ($CaCO_3$) is precipitated as Calcium fluorine (CaF_2). The sediments formed are collected in an inorganic solid container.

PHOTOS

Fig. 1 HF is lighter than air and diffuses in air easily, so the experiment must be conducted in a fume hood.

Fig. 2 The section of glass that the hydrofluoric acid is in contact with reacts with the silicium dioxide in glass and erodes the web like structure of it.

Fig. 3 After covering the surface of the glass with paraffin, a figure is drawn with a metal pencil. Then HF acid is applied on the figure covered with paraffin.

Fig. 4 Clean by scraping all the paraffin on the glass.

Fig. 5 As a result of the glass and HF reaction, the composed sandy like waste is cleaned with a metal pen.

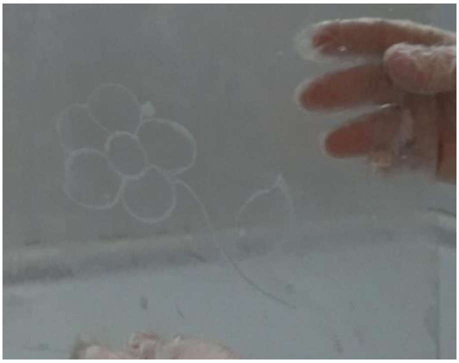

Fig. 6 Finally the drawing is visible.

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THE CHEMICAL VOLCANO

OVERVIEW

The *Chemical Volcano* is one of many volcano experiments suitable for secondary level, but due to the generation of an abundant reaction product and of its exothermic character is one of most spectacular ones. It can be applied when teaching the exothermic reactions, when discussing the oxidation numbers and redox reactions, when different reaction types are presented (in this case the decomposing reactions). As dramatic changes occur, from orange crystals as reactant to a green solid as product the reaction is an example of changes occurred during a chemical transformation.

MATERIALS AND METHODS

A few grams (not more than 10 g [1]) of amonium dichromate $(NH_4)_2Cr_2O_7$ are placed in a crucible in a form of a cone, in a well ventilated hood, on a fire resistant support. The orange salt is ignited with the help of a Bunsen burner or even a match or lighter. For a rapid ignition you may add 1 mL of acetone in the center of the cone and set on fire [2].

To ensure a spectacular decomposition reaction, a large crucible is preferred which rapidly will get filled with "ashes" enhancing the volcano aspect of the reaction site. Preferably place under the fire resistant support a large sheet of aluminum foil or paper which can be wrapped up after the experiment with all the evolved chromium(III) oxide.

HAZARDS

Like all Chromium (VI) salts ammonium dichromate is highly toxic, irritant and a proven carcinogen. Do not breathe in chromium compounds dust [3] Preferably the reaction is to be performed in a ventilated hood.

$(NH_4)_2Cr_2O_7$

CHEMICAL REACTION EQUATION

$$(NH_4)Cr_2O_{7(s)} \text{ ---> } Cr_2O_{3(s)} + N_{2(g)} + 4\ H_2O_{(l)}$$

CONCLUSIONS

Once ignited the amonium dichromate $(NH_4)_2Cr_2O_7$ is self decomposing setting free a copious amount of a green fluffy chromium(III) oxide, Cr_2O_3 which comes out in waves due to the abundant generation of gases (N_2 and H_2O). Due to the exothermic effect of the reaction the evolved water is also in gaseous form and is accompanied by many sparks which give an overall aspect of a volcano. A specific aspect of this chemical reaction is that both the oxidizer and the reducer are present in the same chemical compound.

WASTE MANAGEMENT

As for the hazards concerned a very toxic, carcinogenic, irritant, caustic compound, the ammonium dichromate generates only harmless compounds as Cr_2O_3, nitrogen (a component of air) and steam, which is hot water. Collect the produced oxide in a container for less toxic metallic waste.

PHOTOS

Fig. 1 The chemical volcano

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ETCHING OF THE METAL SURFACE

OVERVIEW

Metals can react with several acidic solutions. The process of reacting a metal surface with an acidic solution is called etching.

MATERIALS AND METHODS

- Hydrochloric acid (HCl, 2 M)
- Small plates of zinc (around 2,5 x 4,0 cm and not more than 1 cm thick)
- Pincers
- Matches
- Etching tool
- Small brush
- Petri dish
- Paraffin wax
- Paper towel
- Gas burner + Tripod + Gauze

Heat up the paraffin wax above the burner until it melts. Take a piece of zinc and apply a layer of wax to it using the small brush. Make sure you cover the entire piece of zinc. Take an etching tool and scrape away the wax on places of the zinc you want to be etched. Place the piece of zinc in a glass container with hydrochloric acid using pincers, make sure the zinc is fully submerged. After 10 minutes take out the piece of zinc using pincers and dry it off with some paper towel. Now you can remove the wax revealing your final etched piece of zinc.

HAZARDS

Hydrochloric acid

CHEMICAL REACTION EQUATION

$\mathbf{Zn_{(s)} + 2HCl_{(l)} \rightarrow ZnCl_{2(aq)} + H_{2(aq)}}$

CONCLUSIONS

When you place zinc inside a glass dish with hydrochloric acid, places not covered with wax will turn to zinc ions and dissolve in the solution.

WASTE MANAGEMENT

Hydrochloric acid	Can be flushed down the drain with water
Zinc	Discard with solid materials
End product	Discard with solid materials

PHOTOS

Fig. 1 Paraffin wax

Fig. 2 Zinc plates

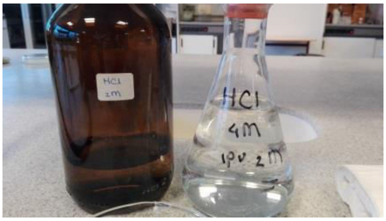

Fig. 3 Hydrochloric acid

REFERENCES

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COPPER OXIDE REDUCTION WITH CARBON

OVERVIEW

Copper is a well-known metal which is used for many things. It is formed in many ways. In this experiment we will reduce copper oxide to copper using carbon.

MATERIALS AND METHODS

- Copper oxide (CuO)
- Carbon (C)
- Test tube
- 2 spatulas
- Test tube clamp
- Petri dish
- Mortar and pestle
- Matches
- Bunsen burner
- Funnel

Put some copper oxide and carbon in the mortar (10:1) using spatulas. Mix the copper oxide and the carbon together with the pestle. Pour the mixture into the test tube with a funnel. Hold your test tube with the test tube clamp and heat it with the Bunsen burner until it glows. Eventually red brown metal can be seen on the surface. Pour the mixture into the petri dish to show the formed copper.

HAZARDS

CuO

Cu

C

CO_2

CHEMICAL REACTION EQUATION

$$2\ CuO_{(s)} + C_{(s)} \ \text{--->}\ 2\ Cu_{(s)} + CO_{2(g)}$$

CONCLUSIONS

Copper oxide can be reduced to the metal copper by using carbon which is oxidised to become carbon dioxide.

WASTE MANAGEMENT

The mixture in the petri dish should be disposed of in an organic waste bin.

PHOTOS

Fig. 1 Materials needed for the experiment.

Fig. 2 Materials.

Fig. 3 Copper.

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ACIDS EFFECT ON THE CALCIUM CARBONATE EMERY

OVERVIEW

Egg shells contain water, some minerals and calcium carbonate ($CaCO_3$) in its structure. Vinegar contains acetic acid (CH_3COOH) in its structure. Now, let's see the general properties of acids before testing the experiment.

General Properties of acids:

1. Acids are sour in taste. For example, the acetic acid in vinegar, malic acid in sour apple, citric acid in lemon and lactic acid in water in yoghurt form, and carbonic acid in cola and beverages. However we cannot taste all the acids, as some of them are toxic.
2. When acids dissolve in water the more the ions occur, the more its conductivity is. In strong acids the conductivity is high, whereas in weak acids it is low.
3. Acids turn blue litmus paper into red.
4. Acids have corrosive property. This property is not the same in all acids.

MATERIALS AND METHODS

The materials used for this experiment are;

A pad to keep the egg still (a plastic bottle can be used here) whereas, in this experiment a beaker is used, egg and vinegar.

1. Put the egg into the container you are to use.
2. Pour vinegar into the container until half of the egg is in it.

There is then a simple procedure, of observation in which these several questions will be answered. What are the changes within 3 hours? Is the change visible? Only some air bubbles will be observed on the shell of the egg. These bubbles are CO_2 particles due to the reaction.

What are the changes at the 6th hour?

The air bubbles multiply. The colour of the vinegar becomes lighter. The water of the reaction leads to this fading in the colour of the vinegar.

12th Hour

Taking the egg out of the container, we will see that the egg is somehow much softer than that of the 6th hour. The shell has thinned, and in some parts the membrane will be seen.

24th Hour

The colour of the vinegar is lighter and its level is lower. When cleaning the air bubbles accumulated on the egg shell, the shell layers is almost gone and the membrane has become visible.

48th Hour

The accumulation of the air bubbles is almost less. The reason to this is the reduction of $CaCO_3$ which will react with the vinegar. The colour of the vinegar is almost gone. When we hold the egg, we will see that its upper part, which has not entered the reaction, is hard.

CHEMICAL REACTION EQUATION

$$2CH_3COOH + CaCO_3 \text{ ---> } CO_2 + H_2O + Ca(CH_3COO)_2$$

CONCLUSIONS

The general changes after 48 hours are as follows;

1. CO_2 has formed in the reaction and as the reaction advanced, CO_2 has accumulated on the egg's surface.
2. The $CaCO_3$ in the shell has dissolved and the shell has lost its hardness.
3. During the reaction the level of the vinegar has reduced and its colour has faded.
4. The less the $CaCO_3$ in the shell, the less the CO_2 is.

PHOTOS

Fig. 1 Vinegar is added to the beaker.

Fig. 2 The vinegar is added, but not completely covering the egg.

Fig. **3** 12 hours after vinegar was added.

Fig. 4 24 hours after vinegar was added.

Fig. 5 48 hours after vinegar was added.

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THE CRACKING OF PARAFFINS

OVERVIEW

In this experiment paraffin (a mixture of saturated hydrocarbons) vapour is cracked by leading it over a heated (500 – 600°C) catalyst. Cracking occurs due to pyrolysis; the high temperatures result in such agitation of the molecules that they break into smaller pieces. This experiment models the industrial cracking process which is important due to increasing demand for fuel.

The hydrocarbons are boiled and the hydrocarbon gases are passed over a bed of hot pumice stones acting as a catalyst. The catalyst provides the hydrocarbon gases with an extended surface area for the cracking to proceed. At least half of the 'cracked' molecules will become alkenes. For example, n-decane (an unbranched alkane with 10 carbons) can be cracked to produce n-octane and ethene to start with (see reaction equation). Subsequently, octane can be cracked further.

The mixture of gaseous short-chain hydrocarbons produced is collected and tested for (i) odour, (ii) flammability and (iii) unsaturation with bromine water as well as with a potassium permanganate solution.

MATERIALS AND METHOD

Materials needed:

A heating tube, a pipette, a delivery tube fitted with a rubber stopper that fits the boiling tube, a trough, 6 test tubes fitted with rubber stoppers, clamp, stand and Bunsen burner.

Reagents:

Pumice stones (or crushed unglazed pottery), liquid paraffin, bromine water, potassium permanganate, concentrated (sulphuric) acid.

Procedure: See figure 1

1. Put about 2 cm of mineral wool at the bottom of the heating tube. Drop about 2 cm^3 of liquid paraffin on the mineral wool, using a dropping pipette. Use just enough paraffin to completely soak the mineral wool; no paraffin should run along the inside of the tube when it is tilted.
2. Clamp the boiling tube near its mouth so that it is tilted slightly upwards, as shown in the figure below. Place a heap of catalyst (crushed pottery or pumice stones) in the centre of the tube and fit the delivery tube.

3. Fill the basin about two-thirds with water and position the apparatus in such a way that the end of the delivery tube is well immersed in the water.
4. Fill six test-tubes with water and immerse them in the trough. The test-tube rubber stoppers have to be immersed as well.
5. Heat the catalyst intensively in the middle of the tube for a few minutes, until the glass is up to a dull red heat. Avoid heating the tube too close to the rubber stopper.
6. While keeping the catalyst at elevated temperature, flick the flame from time to time to the end of the tube for a few seconds to vaporize some of the liquid paraffin. Try to produce a steady flow of bubbles from the delivery tube. Be careful not to heat the liquid paraffin too strongly or let the catalyst cool down. To avoid suck-back, do not remove the flame from heating the tube while gas is being collected. If it looks like suck-back is about to occur, lift the whole apparatus by lifting the clamp stand.
7. When a steady flow of gas bubbles is established, collect six tubes full of gas by holding them over the delivery tube end. Take care not to lift the water-filled tubes out of the water when moving them to avoid leakage of air. Seal the full tubes by pressing them down on the rubber stoppers, then place them in a rack.
8. When gas collection is complete, first remove the delivery tube from the water. Only then stop heating.
9. Then the tubes have to be analysed as follows:
   The first two tubes can be discarded because they will contain air.
   The third tube will be used to investigate the flammability of the gas. Once the gas is lit, invert the test-tube to allow the heavier-than-air gas to flow out and burn.
   Add one mL of bromine water to the fourth tube of gas, stopper and shake well. Decolourization reveals the presence of ~C=C~ bonds.
   Add one mL of acidified potassium permanganate solution to the fifth tube, stopper and shake well. Also here unsaturated bonds cause the purple colour to disappear and turn yellow/brownish.
   The last tube is used to get an impression of the odour. Do we smell a typical light hydrocarbon odour? Like camping gas or LPG??

HAZARDS

C_8H_{18}		**C_2H_4**	
Br_2		**$KMnO_4$**	
$C_{10}H_{22}$		**H_2SO_4**	

Eye protection: use goggles. During the experiment the students and teacher should be protected by safety screens in case of unexpected suck-back causing the hot tube to shatter.

CHEMICAL REACTION EQUATION

$$C_{10}H_{22(g)} \dashrightarrow C_8H_{18(g)} + C_2H_{4(g)}$$

WASTE MANAGEMENT

Dispose of all liquid residues into a sink and wash away with copious amounts of tap water. Dispose of all solid residues into a normal dustbin.

PHOTOS

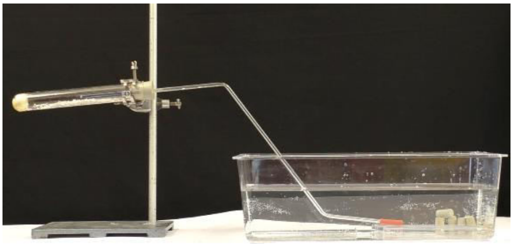

Fig. 1 Set-up for the cracking of paraffine.

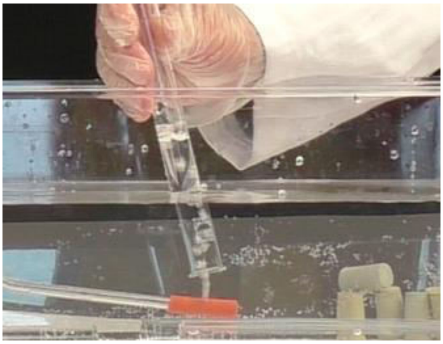

Fig. 2 Collecting the gas

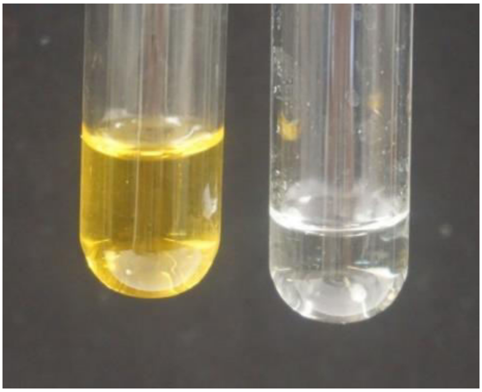

Fig. 3 Test with bromine water

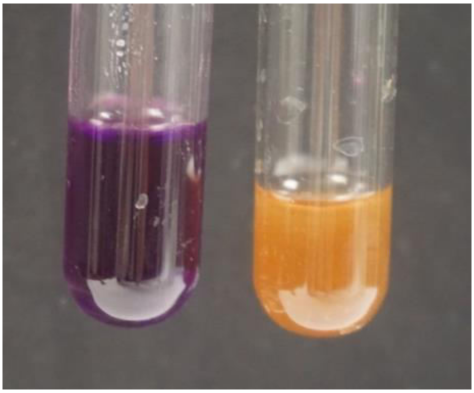

Fig. 4 Test with potassium permanganate

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EXPLOSION OF NATURAL GAS IN A TIN

OVERVIEW

Natural gas is a widely used fuel. For domestic purposes it is used for central heating and as a heat source for cooking. Some technical applications are: driving engines and turbines.

The composition of natural gas depends on the geographical location where it is found. The main component, however, is always methane, CH_4.

The way we use natural gas is by first mixing it with air; when this is ignited with e.g. a match, the combustion process starts (see reaction equation). When sufficient oxygen is present, the combustion process proceeds with blue flames, indicating complete combustion.

This experiment will show that the combustion process does not necessarily proceed in a controlled way. When a certain ratio of natural gas to air is achieved, the reaction looks like it has stopped. Then, unexpectedly, the mixture explodes…………

MATERIALS AND METHODS

Requirements:

A paint can with lid, a tripod, a rubber hose, matches, natural gas.

Procedure:

First we make two holes in the can: one in the lid and the other in the bottom. We fill the can with natural gas through the bottom hole using the rubber hose. After filling, the can is placed on the tripod; the hole in the bottom has to remain unobstructed. Then we ignite the gas flowing out through the hole in the lid. The gas will burn for some time until it is exhausted. Meanwhile, air flows into the tin through the hole in the bottom. The flame will become smaller and smaller until there is no more sign of a reaction at all… then, the opportune explosion limit is reached and the lid is blown off with a loud BANG!!

HAZARDS

CH_4

CHEMICAL REACTION EQUATION

$\mathbf{CH_{4(g)} + 2O_{2(g)} \text{--->} CO_{2(g)} + 2H_2O_{(g)}}$

CONCLUSIONS

Mixing natural gas and air is restricted to certain limits. With too small quantities of air, yellow flames indicate incomplete combustion. With a too high air to gas ratios, however, the mixture becomes explosive.

WASTE MANAGEMENT

There is no waste to handle.

PHOTO

Fig. 1 The moment of explosion.

REFERENCES

1. https://en.wikipedia.org/wiki/Flammability_limit

SYNTHESIS AND COMBUSTION OF ETHYNE (ACETYLENE)

OVERVIEW

In this experiment the synthesis of ethyne (acetylene) from calcium carbide and water and the following combustion reaction is shown. In the industrial use ethyne is gained by the high temperature pyrolysis of mineral oil or petroleum gas. Carbide lamps using the reaction of calcium carbide with water were frequently used in mining and are still used in caves because the consumption of oxygen is low. [1,2]

MATERIALS AND METHODS

Put a little calcium carbide in a tin can with a small hole in the bottom and close it with the lid. Then inject a few millilitres of water through the hole into the can; ethyne is formed immediately. After a few seconds, keep the flame of the lighter to the hole and experience the explosive combustion of ethyne in the can.

HAZARDS

Ethyne is highly flammable and in many proportions mixed with air explosive. Calcium carbide is highly flammable and corrosive. It causes skin irritation and eye damage. Calcium hydroxide is very corrosive and is harmful for skin. A fire extinguisher should be available. Protective gloves and protection goggles should be worn.

C_2H_2

CaC_2

$Ca(OH)_2$

O_2

CHEMICAL REACTION EQUATION

$$CaC_{2(s)} + 2\,H_2O_{(l)} \text{ ---> } C_2H_{2(g)} + Ca(OH)_{2(aq)}$$

$$2\,C_2H_{2(g)} + 5\,O_{2(g)} \text{ ---> } 4\,CO_{2(g)} + 2\,H_2O_{(g)}$$

CONCLUSION

The two reactions are both exothermic, which means that much energy is released to the surrounding area in forms of heat and light. The advantage of the carbide lamp is that it also works in areas with high humidity. The formed $Ca(OH)_2$ solution has a high pH value as it gives an alkaline reaction with water.

WASTE MANAGEMENT

The $Ca(OH)_2$ suspension may be diluted with water and poured in the sink.

PHOTOS

Fig. 1 combustion of ethyne

REFERENCES

1. http://www.youtube.com/watch?v=KGfE-geH5qo(12.01.17)
2. http://en.wikipedia.org/wiki/Acetylene

ETHYNE EXPLOSION IN CAN

OVERVIEW

In this experiment we will demonstrate the energy released in an ethyne explosion using basic household tools and some chemical compounds.

MATERIALS AND METHODS

- Matches/lighter
- Dish soap
- Distilled water
- Calcium carbide
- Tin can, 0.5 L
- Spatula + spoon
- Tennis ball
- Tape (optional)
- Scissors

Take the tin can and make a small hole in in the side of the can approximately 1cm above the bottom. Add 1 gram of calcium carbide into the can. Close the can with the tennis ball and push firmly to secure it in place. Make sure it closes up the top of the can tightly. Add 3ml of dish soap in the can, through the small hole. Add roughly 11ml of the distilled water to the can through the hole. Hold a match/lighter close to the small hole. Step back.

HAZARDS

CaC_2

$Ca(OH_2)$

C_2H_2

CHEMICAL REACTION EQUATION

$$CaC_{2\,(s)} + 2H_2O_{\,(l)} \rightarrow C_2H_{2\,(g)} + Ca(OH)_{2\,(s)}$$

$$C_2H_{2\,(g)} + 2O_{2\,(g)} \rightarrow H_2O_{\,(l)} + 2CO_{2\,(g)}$$

WASTE MANAGEMENT

The remaining substance in the can has to be mixed with hot water and can be thrown in the sink.

PHOTOS

Fig. 1 Overview of the setup

Fig. 2 Addition of soap

REFERENCES

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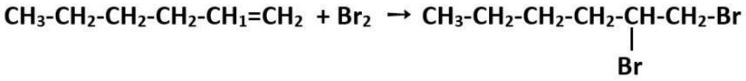
2. https://en.wikipedia.org/wiki/Calcium_carbide
3. BINAS HAVO/VWO, informationbook havo/vwo for the education in naturalsciences, version 6, 2013
4. https://en.wikipedia.org/wiki/Carbide

BROMINATION OF ALKENES

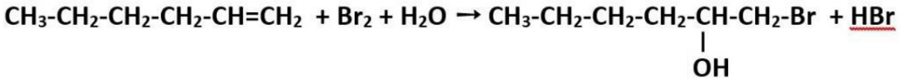
OVERVIEW

When an alkene such as 1-hexene is allowed to react with bromine water the brownish/red solution becomes colourless. Alkenes decolourize bromine water because bromine is added to the double ~C=C~ - bond yielding 1,2-dibromohexane:



$$CH_3\text{-}CH_2\text{-}CH_2\text{-}CH_2\text{-}CH_1{=}CH_2 + Br_2 \rightarrow CH_3\text{-}CH_2\text{-}CH_2\text{-}CH_2\text{-}\underset{\substack{|\\ Br}}{CH}\text{-}CH_2\text{-}Br$$

In case we use a solvent which can potentially act as a nucleophile, such as water or an alcohol, the intermediate halonium ion can be "trapped" by the solvent, resulting in "halohydrin" products in the case where the solvent is water. The water molecule attacks the most substituted carbon atom because this carbon atom is best able to stabilize the (intermediate) positive charge:



$$CH_3\text{-}CH_2\text{-}CH_2\text{-}CH_2\text{-}CH{=}CH_2 + Br_2 + H_2O \rightarrow CH_3\text{-}CH_2\text{-}CH_2\text{-}CH_2\text{-}\underset{\substack{|\\ OH}}{CH}\text{-}CH_2\text{-}Br + HBr$$

Now, the product is 1-bromo-hexanol-2. Both reactions are likely to compete.

MATERIALS AND METHODS

Materials and reagents needed:

Gloves, goggles, aprons, test tubes, bromine water, 1-hexene or any other alkene), potassium bromide, sulphuric acid, hydrogen peroxide

Procedure:

Making bromine water: Add 0.1 gram of a bromide salt (such as potassium bromide) to 5-10 mL of 35% H2O2 solution in a test tube and acidify with 1 mL of concentrated sulphuric acid. Place a rubber stopper onto the test tube and shake until all the bromide salt is dissolved.

Pour 1 mL of cyclohexene in a clean test tube. Add 10-15 drops of bromine water, shake the tube and observe the result. The brownish/red colour of the bromine water will disappear.

HAZARDS

Br_2

KBr

H_2SO_4

C_6H_{12}

H_2O_2

$C_6H_{12}Br_2$
1,2-dibromohexane

$C_6H_{13}BrO$
1-bromo-2-hexanol

CHEMICAL REACTION EQUATIONS

$C_6H_{12(aq)} + Br_{2(aq)} + H_2O_{(l)} \rightarrow C_6H_{13}BrO_{(aq)} + H_3O^+_{(aq)} + Br^-_{(aq)}$ (1)

$C_6H_{12}(aq) + Br_2(aq) \rightarrow C_6H_{12}Br_2(aq)$ (2)

CONCLUSIONS

In case a nucleophilic solvent is used bromine molecules can be added to double ~C=C~ bonds in two different ways. However, in both cases bromine molecules are consumed leading to decolourization of the bromine water. Hence it can be concluded that bromine water is a reliable agent to detect the presence of double ~C=C~ bonds in aqueous solutions.

WASTE MANAGEMENT

Dispose of all residues by pouring into a drum designated for waste collection.

PHOTO

Fig. 1 and **2** Decolourization of bromine water

REFERENCES

http://www2.ucdsb.on.ca/tiss/stretton/chem2/orglab1.htm

ADDITION OF BROMINE AND POTASSIUM MANGANATE (VII) TO HEXENE

OVERVIEW

The most common chemical transformation of a carbon-carbon double bond is the addition reaction. A large number of reagents, both inorganic and organic, have been found to add to this functional group.[1]The carbon-carbon double bonds in alkenes such as ethene react with potassium manganate(VII) solution (potassium permanganate solution). Alkenes react with potassium manganate(VII) solution in the cold. If the potassium manganate(VII) solution is slightly alkaline, the purple solution first becomes dark green and then produces a dark brown precipitate.[2] In this experiment the electrophile addition of bromine with hexene and the oxidation of $KMnO_4$ solution with hexene is shown. Both times the double bound splits apart but the results are different. Once a clear single phase solution is produced, the second time a solution with two phases is formed, a clear one on the top and a brown one at the bottom.

MATERIALS AND METHODS

Two test tubes are safely fixed on a stand. Both need to be filled with 2 mL of hexene. As next step 4-6 drops of bromine water (saturated solution) are added to the first tube and 4-6 drops of $KMnO_4$ solution (0.1 N) are added to the second tube. ATTENTION: The whole experiment needs to take place in a laboratory hood. The bromine is a high graded volatile chemical and reacts very quickly with hexene.

HAZARDS

Hexene irritates the eyes and respiratory tracts to a lesser extent than do the vapours. It is easily inflammable. In high doses, hexene acts narcotically. [3] $KMnO_4$-solution is a strong oxidizer, thus should be kept separated from oxidizable substances such as skin, paper, and clothing. [4] Breathing bromine gas could cause you to cough, have trouble breathing, get a headache, have irritation of your mucous membranes, be dizzy, or have watery eyes. Getting bromine liquid or gas on your skin could cause skin irritation and burns. Liquid bromine that touches your skin may first cause a cooling sensation that is closely followed by a burning feeling. [5] 1,2-Dibromohexane is irritant to skin and eyes.

C_6H_{12}

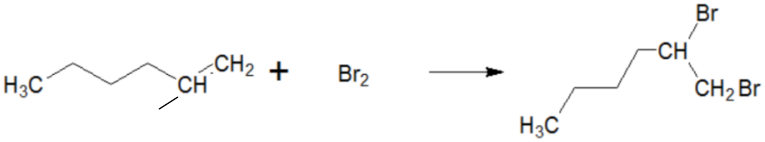
Br_2

$KMnO_4$

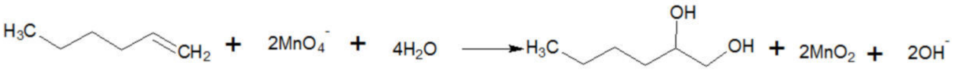
1,2-dibromohexane

CHEMICAL REACTIONS EQUATION

Addition of Bromine



Oxidation of $KMnO_4$



CONCLUSION

The final result of this experiment is one clear solution from the addition with bromine and one solution with 2 phases, one clear and one brown (MnO_2) from the oxidation with $KMnO_4$-solution. In the second reaction, an alcohol compound is produced.

WASTE MANAGEMENT

The remaining chemicals are disposed in an organic waste collector.

PHOTOS

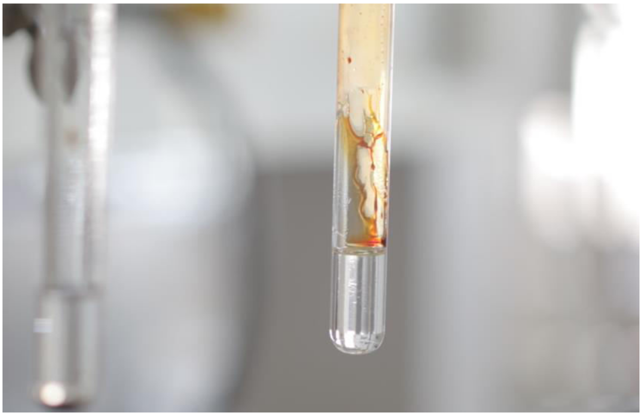


Fig. 1: Starting the reaction with bromine

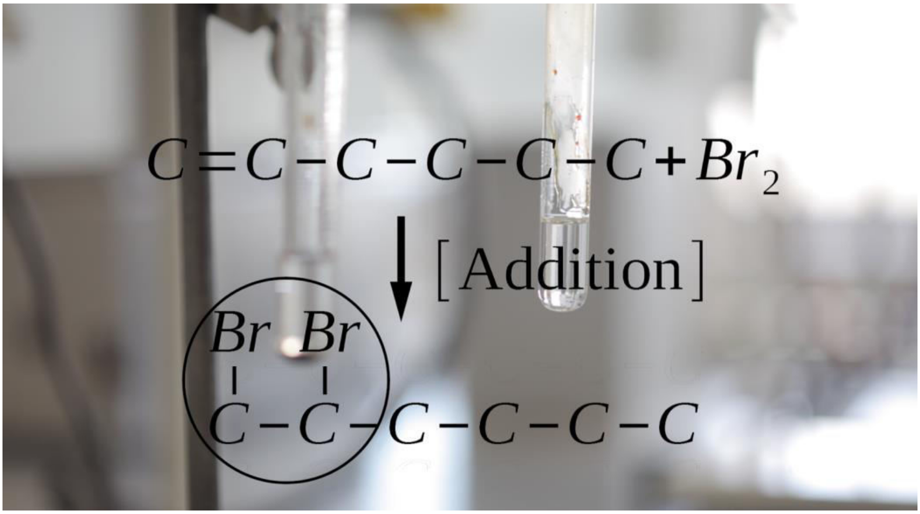


Fig. 2 Decolouring of the liquid after reaction with bromine

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2. http://chem.libretexts.org/Core/Organic_Chemistry/Alkenes/Reactivity_of_Alkenes/Oxidation_of_Alkenes_with_Potassium_Manganate
3. http://de.wikipedia.org/wiki/1-Hexen
4. http://en.wikipedia.org/wiki/Potassium_permanganate
5. http://emergency.cdc.gov/agent/bromine/basics/facts.asp

LIGHT INDUCED RADICAL BROMINATION REACTION OF HEXANE

OVERVIEW

In this experiment the radical substitution of hydrogen atoms by bromine atoms is performed with hexane. Radical reactions are often induced by light, photons of high energy are necessary to split the C-H bond and replace with a C-Br bond. Elemental bromine is present in liquid form under standard conditions. In nature, bromine does not occur elementally, but only in different compounds. The most important compounds are the bromides. Hexane (n-hexane is a chemical compound associated with the alkanes. It is a colorless liquid with the chemical formula C_6H_{14}. [1,2]

MATERIALS AND METHODS

Put approximately 100 mL n-hexane in an Erlenmeyer flask, then place the flask in the exhausting hood and add a few drops of bromine so that *a distinct coloration is visible*. Then place the flask on an overhead projector and observe the disappearing of the colour when switched off light and switched on light. Put a moistened indicator paper across the opening of the Erlenmeyer flask and measure the pH value.

HAZARDS

Caution, n-hexane is highly flammable, harmful to health and must not be released to the environment. Bromine is corrosive, very toxic and harmful of health and must not be released to the environment Attention by working with bromine. Sodium hydroxide solutions are corrosive. Wear protection gloves and glasses and always work under an exhaustion hood.

C_6H_{14}

NaOH

Br_2

CHEMICAL REACTION EQUATION

$C_6H_{14(l)} + Br_{2(l)}$ ---> $C_6H_{13}Br_{(l)} + HBr_{(g)}$

CONCLUSIONS

No change is visible when the light is off. In the case of illumination, a gradual discoloration of the reddish solution is observed from orange and yellow to almost colorless. When the stopper is removed, white fumes rise and the indicator paper turns red.

WASTE MANAGEMENT

To the used solution, saturated sodium thiosulfate solution is added to remove excess bromine. Then neutralized with sodium hydroxide and the solution is then disposed of in the organic solvent waste.

PHOTOS

Fig. 1 Photo induced bromination of n-hexane

REFERENCES

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2. http://www.chids.de/dachs/praktikumsprotokolle/PP0077Bromierung_von_Hexan.pdf

UNDERWATER REACTION OF ETHYNE AND CHLORINE

OVERVIEW

This experiment shows the exothermic underwater reaction of two gases, ethyne (acetylene) and chlorine, to form dichloroethene. The two gases are brought to reaction with a low-cost gas developer. An exothermic reaction means that energy is released in form of heat or light, in the present reaction as lightnings under water. The reaction of ethyne with chlorine produces small amounts of chlorinated hydrocarbons (dichloroethene, tetrachloroethane), which are relatively dangerous. [1, 2]

MATERIALS AND METHODS

Prepare two low cost gas developers to produce ethine and chlorine consisting of a test tube, soft rubber stopper with 2 cannulas, 2 mL syringe (running roughly), 30 mL syringe (oiled, running smoothly). Place a spatula of potassium permanganate and calcium carbide respectively in a test tube. Close the test tube with a soft rubber stopper with two long cannulas. Put the 30-mL syringe on a cannula and the 2-mL syringe on the other. Fill the 2-mL syringe with concentrated hydrochloric acid (for potassium permanganate tube) and with water (for Calciumcarbide tube) respectively. Inject the hydrochloric acid and the water respectively dropwise into the test tube. Chlorine as well as ethane gas is formed in the 30 mL syringes. The first fraction of gas must be removed because it contains too much air. The second fraction can be used. Fill the Erlenmeyer flask with approximately 80 mL of water and seal it with the rubber stopper with the two long cannulas. The ends of the two cannulas must be closely adjacent. The rubber stopper has a hole to allow the developing pressure to escape. Mount the two large syringes on the cannulas of the Erlenmeyer flask. If the two gases are pressed simultaneously into the Erlenmeyer flask, the ethyne reacts with the chlorine to dichloroethene. This reaction is strongly exothermic.

HAZARDS

When calcium carbide gets in contact with water, flammable gases are produced, which can ignite spontaneously. Therefore, the use of protective gloves is mandatory (hand sweat). CaC_2 causes skin irritation, grave eye damage and can irritate the respiratory tract.

Hydrochloric acid causes severe skin burns and serious eye damage and may irritate the respiratory tract. HCl can be corrosive to metals. The use of goggles is mandatory.

Potassium permanganate can enhance fires because it is an oxidizing agent. Furthermore, $KMnO_4$ is harmful if swallowed, causes severe skin burns and serious eye damage and is very toxic to aquatic organisms, with long – term effects. Therefore, gloves are mandatory.

Chlorine can cause or intensify fires because it is an oxidizing agent. Chlorine is a gas under pressure which can explode when heated. There is a danger to life when inhaled, as it can irritate the respiratory tract. Cl_2 causes skin irritation, severe eye irritation and is very toxic to aquatic organisms. Therefore, protective gloves are mandatory.

Ethyne is an extremely flammable gas and can react explosively even in the absence of air.

The liquid and steam of hydrocarbons are highly flammable. Furthermore, they are harmful by inhalation and harmful to aquatic organisms, with long - term effects.

CaC_2		**HCl conc.**	
$KMnO_4$		**Cl_2**	
C_2H_2			

CHEMICAL REACTION EQUATION

$$CaC_{2\,(s)} + 2H_2O_{(l)} \rightarrow Ca(OH)_{2\,(s)} + C_2H_{2\,(g)}$$

$$2KMnO_{4\,(s)} + 16HCl_{(l)} \rightarrow 5Cl_{2\,(g)} + 2KCl_{(s)} + 2MnCl_{2(s)} + 8H_2O_{(l)}$$

$$C_2H_{2\,(g)} + Cl_{2\,(g)} \rightarrow 2C_{(s)} + 2HCl_{(aq)}$$

$$C_2H_{2\,(g)} + Cl_{2\,(g)} \rightarrow C_2H_2Cl_{2\,(g)}$$

In addition to larger amounts of soot, halogenated hydrocarbons are also formed.

CONCLUSIONS

This experiment shows the exothermic reaction of acetylene and chlorine. The result is an immediate and strong exothermic reaction of the two gases. Lightnings under water are observed. When the equipment is dismantled, the characteristic smells of the gases can be noted.

WASTE MANAGEMENT

Calcium carbide may be collected for other experiments. The other remaining chemicals may be diluted with water and poured into the sink.

PHOTOS

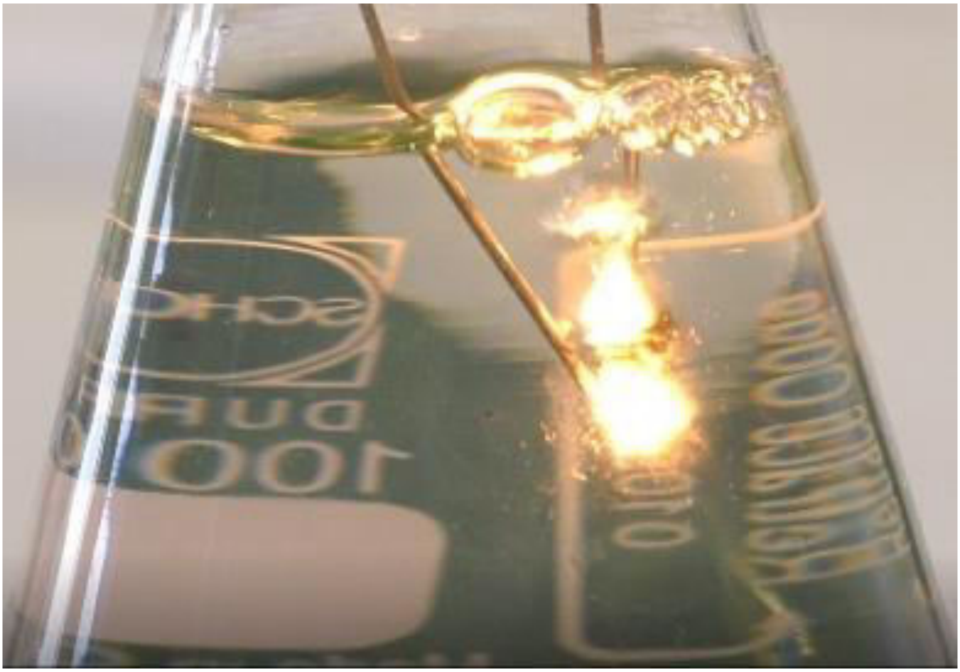

Fig. 1 Lightnings under water

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2. http://www.chids.de/dachs/praktikumsprotokolle/PP0115Blitze_unter_Wasser.pdf, 19.12.16

SYNTHESIS OF SILVER ACETYLIDE

OVERVIEW

Transitional metal acetylides like silver acetylides are insoluble in water therefore these compounds can be synthesized in watery solutions as precipitates. Due to this property, marginal alkynes can be easily identified in a gaseous mixture by bubbling the mixture in a silver complex solution (Tollens' Reagent) and precipitating the silver acetylide, this reaction having analytical value. Moreover silver acetylides were used as a mean of isolating silver form different discarded solutions containing this metal. Although this reaction is a simple one the silver acetylide has to be handled with extreme precaution as this compound is highly explosive when dry [1]. The synthesis and properties of silver acetylide are presented in the following steps.

MATERIALS AND METHODS

The preparation of the Tollens' Reagent Silver (I) diammine hydroxide $[Ag(NH_3)_2]OH$ is prepared in two steps: First 20 mL of a 0.3M $AgNO_3$ is treated with a few drops of concentrated 3-4M NaOH solution. The abundant, brown silver oxide precipitate is dissolved by adding, in small amounts, a concentrated solution of ammonia NH_3 (35%) until the solution becomes clear and colorless. The Tollens' Reagent is not suitable for storage; it has to be used only when fresh. Other sources [2] allow the use only of the concentrated NH_3 solution skipping the alkalizing step with NaOH.

Preparation of acetylene Acetylene is generated using the reaction between calcium carbide and water. For this place in a 250 mL flask 3 g of calcium carbide (a few pieces of a size of a large peanut) and add water to these. As the acetylene has to be bubbled through the Tollens' reagent solution use any glassware which can help you to conduct the evolved gas from a flask to the Tollens' reagent containing beaker (glass retort, Buchner funnel, Schlenk or any other flask with side arm and stopper, a Kipps apparatus or a simple flask with a one hole rubber stopper of which a tube is conducted to the beaker).

The precipitation reaction demonstrates, once again if necessary, the acidic properties of acetylene (and by extrapolating of the marginal alkynes) and their tendency to form substituted acetylenes with complexes of transitional metals. As acetylene is bubbled through the Tollens' reagent solution a grayish precipitate is formed. If necessary add more calcium carbide to the gas generator glassware and try to precipitate all the silver from the Tollens' reagent beaker. When finished, filter off the precipitate and let it dry in a dark place. **When dry, handle with care as it is explosive!**

HAZARDS [3]

$AgNO_3$		**NaOH**	
NH_3		**Tollens'**	
CaC_2		**C_2Ag_2**	

CHEMICAL REACTION EQUATION

$$2\,AgNO_{3\,(aq)} + 2\,NaOH_{(aq)} \longrightarrow Ag_2O_{(s)} + 2\,H_2O + 2\,NaNO_{3\,(aq)}$$

$$Ag_2O_{(s)} + H_2O + 4\,NH_{3\,(aq)} \longrightarrow 2\,[Ag(NH_3)_2]OH_{(aq)}\ \text{Tollens}$$

$$CaC_{2\,(s)} + 2H_2O \longrightarrow HC{\equiv}CH_{(g)} + Ca(OH)_{2\,(s)}$$

$$2\,[Ag(NH_3)_2]OH_{(aq)} + HC{\equiv}CH_{(g)} \longrightarrow AgC{\equiv}CAg_{(s)} + 4\,NH_{3\,(aq)} + 2\,H_2O$$

CONCLUSIONS

The calcium carbide reacts with water generating acetylene which is a gas. The liquid phase becomes cloudy as insoluble lime is formed. The evolved acetylene is conducted to the silver amino complex solution where is bubbled generating insoluble silver acetylide in a form of a grey precipitate. Filtered off this precipitate is suitable to decomposition in light therefore it has to be dried in darkness. Once dried, if the precipitate is exposed to a flame, it explodes. Caution!!! Use a long stick to generate the explosion and only for quantities less than 0.3 g.

WASTE MANAGEMENT

Silver containing solutions are very dangerous for the environment. They destroy the microorganisms used in the wastewater treatment plants therefore these solutions are not poured down the drain. Usually a double precipitation reaction is performed with HCl followed by filtration. The filtered solid containing AgCl is disposed in containers for collecting Ag salts. The lime remained after acetylene generation is diluted and poured down the drain.

Caution is needed when handling the silver acetylide. As it is explosive the accumulation and storage of this highly instable compound is not advisable. Some procedures envisage the storage of the acetylide under water, acetone or ethanol, far away from acidic solutions to avoid decomposition.

PHOTOS

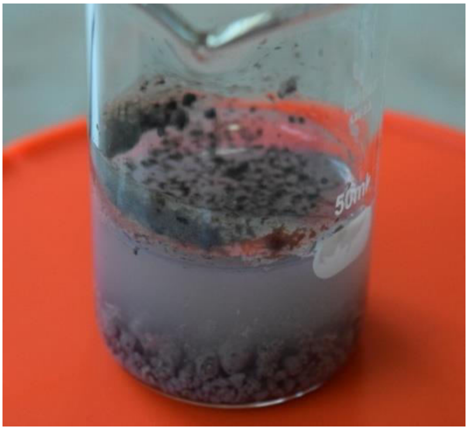

Fig. 1 The result of interaction between acetylene and the Tollens' Reagent.

Fig. 2 Silver acetylide freshly filtrated.

Fig. 3 A small amount of silver acetylide (three times the size of a match head) is capable to produce a loud bang when ignited from a safe distance.

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SYNTHESIS OF COPPER (I) ACETYLIDE

OVERVIEW

We presented in the previous experiment the low solubility of transitional acetylides by presenting the silver compounds. In the same manner the copper acetylides can be synthesized by bubbling acetylene in a copper(I) complex solution. Copper(I) acetylide has to be handled also with extreme precaution as this compound is highly explosive when dry [1]. The synthesis and properties of cuprous acetylide are presented in the following steps.

This experiment is useful in the secondary chemistry classroom when discussing the coordination compounds and the synthesis and properties of acetylene.

MATERIALS AND METHODS

The preparation of the Copper(I) complex 1.0 g of copper(II) sulfate $CuSO_4*5H_2O$ is dissolved in the minimum required amount of water. Concentrated ammonia (25%) solution is added until the liquid is clear deep blue. Dilute the solution to 50 mL and then add 3.0 g of hydroxylamine hydrochloride $NH_2OH*HCl$. A copper(I) complex is formed $[Cu(NH_3)_2]Cl$ which will be used in the copper(I) acetylide synthesis [2]. **Preparation of acetylene** Acetylene is generated using the reaction between calcium carbide and water. Use the same amount of reagents and the same apparatus to generate the alkyne as in the previous experiment. **Cuprous acetylide generation** occurs when the acetylene is bubbled in the previously prepared cuprous complex solution. The clear solution becomes dark red when a fine precipitate is formed. When you reach the end of the precipitation process filter off the solid and let it dry. **When dry, handle with care as it is explosive!**

HAZARDS

Do not overcome the recipe presented in this section. Large quantities of copper acetylide could lead to accidents when handling inappropriate.

$CuSO_4$

NH_3

CaC_2

$NH_4OH*HCl$

C_2Cu_2

CHEMICAL REACTION EQUATION

$$CaC_{2\,(s)} + 2H_2O \longrightarrow HC{\equiv}CH_{(g)} + Ca(OH)_{2\,(s)}$$

$$2\,[Cu(NH_3)_2]Cl_{(aq)} + HC{\equiv}CH_{(g)} \longrightarrow CuC{\equiv}CCu_{(s)} + 2\,NH_{3\,(aq)} + 2\,NH_4Cl_{(aq)}$$

CONCLUSIONS

The copper sulfate solution reacts with concentrated ammonia generating the precipitation of copper hydroxide in the first step. Further addition of ammonia lead to a deep blue complex known under the name: Schweizer Reagent $[Cu(NH_3)_4](OH)_2$. The addition of hydroxylamine hydrochloride $NH_2OH*HCl$ reduces the Cu(II) ions to Cu(I) transforming the complex into another one $[Cu(NH_3)_2]Cl$. The calcium carbide reacts with water generating acetylene which is a gas. The solution becomes cloudy as insoluble lime is formed. The evolved acetylene is conducted to the copper(I) amino complex solution where insoluble cuprous acetylide is precipitated in a form of a dark red precipitate. Once dried, if the red precipitate is exposed to a flame or to friction it explodes. Caution!!! Use a long stick holding a flame at the end to generate the explosion and only for small quantities (0.5 g).

WASTE MANAGEMENT

The cuprous solutions are very harmful for the environment. Dispose these solutions in a container used for collecting less toxic inorganic compounds.

PHOTOS

Fig. 1 The copper (I) acetylide synthesis

Fig. 2 Copper (I) acetylide decomposition

REFERENCES

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REACTION OF ALCOHOLS WITH SODIUM

OVERVIEW

Alcohols are R-OH structured compounds carrying single hydroxyl (-OH) group. The naming of alcohols according to traditional system, is made by adding ALCOHOL to the alkyl which belong to the hydroxyl group.

The names sec-, tert-, which exist in naming and the samples showing the branching state of the attaching carbons exist in the names of alcohols. The primer alcohols in the carbon which the hydroxyl groups belong to contain two hydrogen. İn the tertiary alcohols no hydrogen is found on the carbon.

According to the number of atoms tread directly to the carbon atom of the functional group. The alkyl groups are classified as primer(pir-), secondary(sec-) and tertiary(tert-). Alcohols exhibit weak acidic properties. The acidic properties of alcohol can be shown by adding the sodium metal into alcohol. The alcohols are weak acids when they react with sodium metal. The hydroxyl group can act as a proton donor to form an alkoxide ion. Alkoxide ions dissolved in alcohol are strong bases which can be prepared by the reaction of an alcohol with sodium metal.

Because the hydroxyl hydrogen is acidic, alcohols, reaction alkaline metals such as Na and K, just like water, enter in a exothermic reaction resulting with a hydrogen output, thus forming the alkolats. The hydrogen gas can be collected and tested by using a burning wooden splinter. A pop sound will be produced. These alcohols to not release H_2 gas as Mg, Zn an Ca.

The alkolate formation is:

$$\mathbf{2ROH + 2\,Na \quad \text{--->} \quad 2RONa + H_2}$$

Alkolates: the salts that from as a result of a metal ion replacing the proton in an OH group of alcohol. Alkolates also form alcohol and hydroxide by water hydrolysis.

$$\mathbf{RONa + 2\,H_2O \quad \text{--->} \quad ROH + NaOH}$$

If the alcohol is in a branched state, the basic property of the alcolate is more (in tertiary it is much higher)

Metallic sodium blows up (explodes) when in water and for the elimination of the remainders, it is necessary to change them into alcolite.

MATERIALS AND METHODS

We put ethyl alcohol, primary butyl alcohol and secondary butyl alcohol into the tubes in succession. Then we add sodium into the tubes. We fix balloons on the ends of the tubes. A couple of minutes later we observe that all the balloons inflated, except for the green balloon that took more time to inflate.

HAZARDS

Na metal

C_2H_5OH

C_4H_9OH

CHEMICAL REACTION EQUATION

$$2\ C_2H_5OH + 2\ Na \longrightarrow 2\ C_2H_5ONa + H_{2(g)}$$

$$2\ C_4H_9OH + 2\ Na \longrightarrow 2\ C_4H_9ONa + H_{2(g)}$$

CONCLUSIONS

The alcohols used in the experiment are ethyl alcohol, primary butyl alcohol and secondary butyl alcohol. When looked at their formulas, it can easily be seen that a methyl($-CH_3$) group is attached to the C atom of hydroxyl(-OH)group in ethyl alcohol, a propyl($-C_3H_7$) group is attached to the C atom of the OH group in the primary alcohol, a methyl($-CH_3$) group and an ethyl($-C_2H_5$) group are attached to the C atom of the OH group in the secondary alcohol. These alkyl groups direct the electron intensity to the carbon atom. As the number of the methyl group or the alkyl chain length increases, the electron which will be directed towards the carbon atom will increase too. Then the acidic property of alcohol will weaken, that is, the breaking of hydroxyl hydrogen will be hard. When the structures of such alcohols are observed, it will be seen that the fastest gas release exists in ethyl alcohol-Na interaction, and the slowest gas release exists in secondary alcohol-Na interaction. İn the experimental study, a fast H_2 gas release is observed in the reaction of ethyl alcohol and primary alcohol with Na.

WASTE MANAGEMENT

Alkoxide / sodium alkoholates (RONa) compound forming as the result of the reaction of alcohols with metallic Na, by hydrolysing with water, forms alcohol and NaOH. While alcohols produce alkonium salts with strong acids, strong bases produce the acids and salts-making use of this property of the compounds, when strong base is added to the alkoxide, thinned with water, medium salt is produced and this material is kept in the container where salt solutions are.

PHOTOS

Fig. 1 With the help of a pipette 2ml of Ethyl alcohol, primer butyl alcohol and secondary butyl alcohol are added in test tubes in a row.

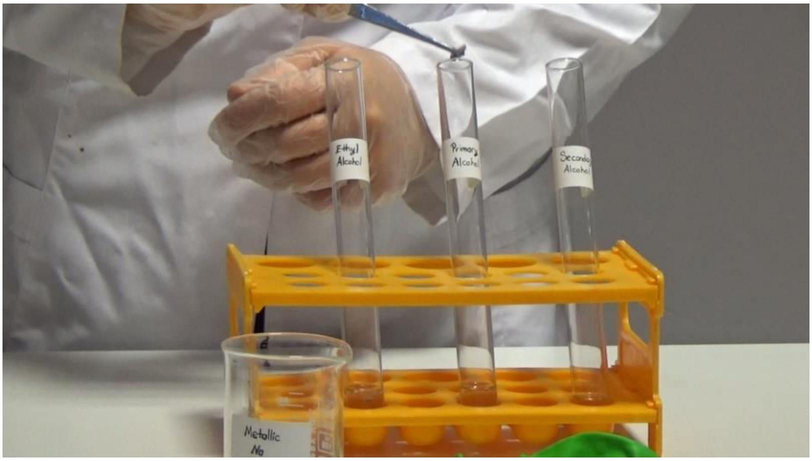

Fig. 2 A piece of sodium metal is added with a spatula into the alcohol in the test tubes and its gas outlet is observed. The gas outlet as a result of the formation of alkoxide from the tube is **hydrogen gas**.

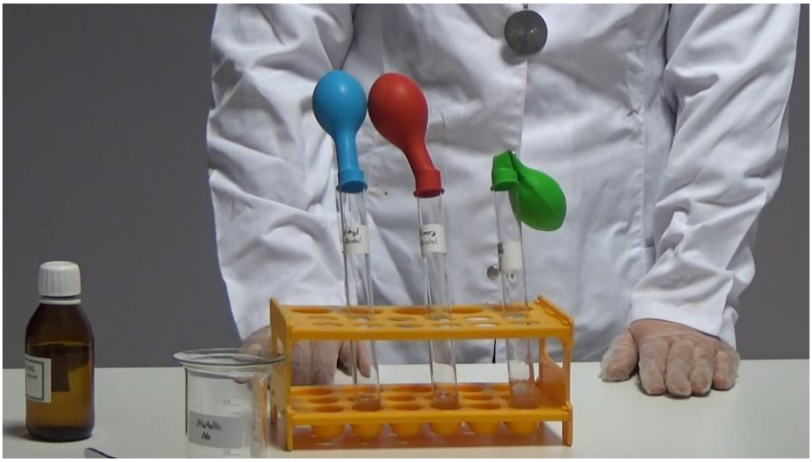

Fig. 3 As a result the alcohol with more carbon atoms that is connected to the OH group which has the most H atom, gives a more rapid reaction to Na and gives off more hydrogen gas.

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PREPARATION OF HOME-MADE WINE

OVERVIEW

Wine…… no consumable liquid has been more written about than wine….. drinking wine is almost synonym of enjoying life! Red wine and romance are called in one breath and no meal is complete without an excellent glass of wine…….

In this experiment however we will regard the production of wine as a mean to produce alcohol [1], just as hydration of ethylene on industrial scale is [2]. Sugars in fruit can be fermented with yeast to yield ethanol. The process is limited to 12 vol% because at higher values the yeast cells become deactivated. Contact with oxygen has to be avoided because ethanol can be oxidized to acetic acid.

We use an apple as an example here, but almost every kind of fruit or mixtures can be used.

MATERIALS AND METHODS

Materials: 1 apple, 50 grams of sugar, 250 mL of water, 250 mL of apple juice, pinch of yeast, 1000 mL Erlenmeyer flask, mortar, heater, water lock, grater, potato peeler.

Procedure: Peel the apple, free it from its seeds and shred it with the grater. Grind 50 grams of sugar in the mortar and mix sugar and apple rasp. Then bring the mixture into the Erlenmeyer flask. Add 250 mL of water, 250 mL of apple juice and heat at 70°C for five minutes. Let the mixture cool down to 25 °C and add the yeast. Close the flask with a water lock in order to prevent air from leaking in. Now let it rest for about three weeks.

HAZARDS

C_2H_5OH

CHEMICAL REACTION EQUATION

$$C_6H_{12}O_6(aq) \rightarrow 2\ C_2H_5OH(aq) + 2\ CO_2(aq)$$

CONCLUSIONS

In this experiment we produced ethanol by fermentation of glucose in an apple. After allowing the reaction to proceed for a couple of weeks, a substantial amount of ethanol was detectable.

WASTE MANAGEMENT

There is no waste to handle.

PHOTOS

Fig. 1 Exposition of our initial equipment and materials.

Fig. 2 Ready!

Fig. 3 Measuring the alcohol content with a vinometer

REFERENCES

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2. http://chemguide.co.uk/organicprops/alcohols/manufacture.html

ALCOHOLIC FERMENTATION WITH YEAST AND DISTILLATION

OVERVIEW

Biochemical reactions are very important in nearly every part of daily life. They are used to produce and conserve food and drinks, produce drugs (e.g. Insulin) and they are responsible for every life on earth. The reaction done in this experiment is a classical reaction. It is easy to perform, and it has a practical application: some drinks, like beer or wine, are produced in this way. Sugar is fermented to ethanol by using a biochemical reaction. Humans have known this process since antiquity. The aim of this experiment is to get an idea of how biochemistry works and to perform a simple reaction.

In this experiment a classical biochemistry task is shown. Sugar is fermented with yeast. The product of this reaction is alcohol (Ethanol). By distillation, the Ethanol is separated from the sugar-solution. A positive evidence is shown by inflaming the alcohol. [1]

MATERIALS AND METHODS

Chemicals: Yeast (*Saccharomyces cerevisiae*), sugar (saccharose), water
Equipment: Sealable container, magnetic stirrer, 2 round flasks, thermometer, distillation gadgetry, round flask heater, lifting platform, flask ring, watch glass
Procedure: Dissolve 80 g of sugar in 1000 ml of water. Add 1 g of yeast. Pour this mixture in the sealable container. Stir it for 48 hours. Construct the distillation gadgetry as shown in Fig. 1. Fill 500 mL of fermented sugar solution into a round flask. Put it into the gadgetry. Boil it until the temperature of the vapour reaches 78°C and collect a certain amount of distillate. Fill 5 ml of the distillate into the watch glass and try to inflame it. Be careful: The flame is not clearly visible.

HAZARDS

In this experiment no hazardous chemicals are used.

CHEMICAL REACTION EQUATION

$$\mathbf{C_{12}H_{22}O_{11(aq)} \xrightarrow{\text{Invertase}} 2\ C_6H_{12}O_{6(aq)}}$$

$$\mathbf{C_6H_{12}O_{6(aq)} \xrightarrow{\text{Glucose Oxidase}} 2\ CO_{2(g)} + 2\ C_2H_5OH_{(aq)}}$$

CONCLUSION

Saccharose ($C_{12}H_{22}O_{11}$) is a carbohydrate from glucose and fructose. An enzyme called "invertase" decomposes Saccharose to its monomers, glucose and fructose. They are oxidized by glucose oxidase to two molecules of carbon dioxide and two molecules of ethanol.

WASTE MANAGEMENT

The chemicals used in this experiment are not hazardous, pour them to the sink.

PHOTOS

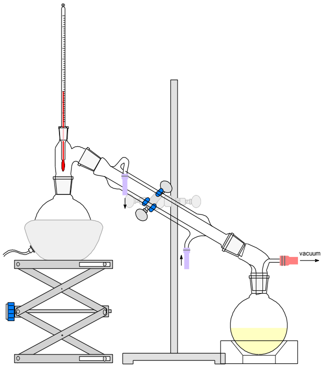

Fig. 1 distillation-gadgetry

Fig. 2 Used chemicals: sugar and yeast

Fig. 3 Burning test for ethanol

REFERENCES

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SYNTHESIS OF METHYLATED BORIC ESTERS

OVERVIEW

A spectacular experiment, much loved by the students is suitable both to study the esterification process in general and more than that, to show students that esterification can be done with inorganic acids as well. As a simple apparatus has to be setup in order to allow a reflux of the reactants and ester distillation this process is also suitable to access practical skills of the students when setting up the experimental work.

MATERIALS AND METHODS

Take a 250 mL one neck round-bottomed glass flask and fix it over an electric heater or gas burner. Add in the flask 2.5g of boric acid (H_3BO_3), pour over 30 mL of methanol and 1-2 mL of concentrated sulfuric acid (H_2SO_4). Close the system with a one hole rubber stopper. Fix in the rubber hole a 10-20 mL glass pipette with its point oriented upwards. This pipette will act as an air cooled refrigerant and will allow in the same time the escape of vapors from the system when overpressure occurs. Start the heating up to 50°C. Observe when the boiling starts. If reflux is observed on the flask, light the vapors escaping the pipette. Observe how the vapors are catching fire and the color of the flame.

HAZARDS

Attention is needed when handling concentrated sulfuric acid as it is very corrosive and if cotton cloths are stained with it, degradation will occur in a few hours. Therefore wearing the apron is compulsory. Due to methanol toxicity and its low lethal dose special attention has to be paid when handling this alcohol. Good ventilation is needed and performing the reaction in a hood is recommended.

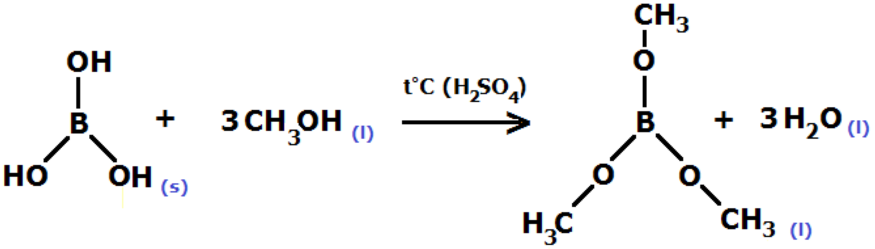
H_3BO_3

H_2SO_4

CH_3OH

$C_3H_9BO_3$

CHEMICAL REACTION EQUATION



$$2(CH_3)_3BO_{3(l)} + 9O_{2(g)} \longrightarrow B_2O_{3(s)} + 6\,CO_{2(g)} + 9H_2O_{(l)}$$

$$2\,CH_3OH_{(l)} + 3O_{2(g)} \longrightarrow 2\,CO_{2(g)} + 4H_2O_{(l)}$$

CONCLUSIONS

The concentrated sulfuric acid catalyzes the esterification process between boric acid and methanol and in the same time is quenching the water evolved from the esterification reaction acting as a dehydrating agent. An ester of an inorganic acid is formed: trimethyl borate, if the reaction mixture is heated and let to reflux 30 minutes. As the methyl esters of this acid are volatile they are easily vaporized, create an azeotrope with methanol which is in excess, and due to the created suprapresure are leaving the system through the tube and catch fire when ignited. These esters are burning in a specific way coloring the flame in green which is visible once the generated esters are light up at the end of the tube. Depending on the tube height and diameter the length of the flame can be varied. The same is observed if we vary the rate of the heating, varying in this case the ester formation rate [1].

The liquid phase containing both non reacted methanol mixed with the trimethyl borate, after cooling down, can be used to produce a green flame as well. Just place 2-3 mL of this liquid in a glass mortar, evaporating dish or even a watch glass and set it to fire with a match. (Don’t forget to ensure proper ventilation!!!). A spectacular effect is obtained if, in a darkened room, 2 mL of this liquid phase is dropped in a PET flask, spread over the sides of the bottle to allow vaporization and light up with a burning match inserted in the flask.

WASTE MANAGEMENT

The trimethyl borate is flammable and irritant. Because vapors can cause explosions it has to be avoided to reach confined spaces [2]. Water causes its decomposition. Generally the reaction is not performed with large quantities of reagents but still caution is needed. Transfer the residue in a container used for collecting flammable organic solvents.

PHOTOS

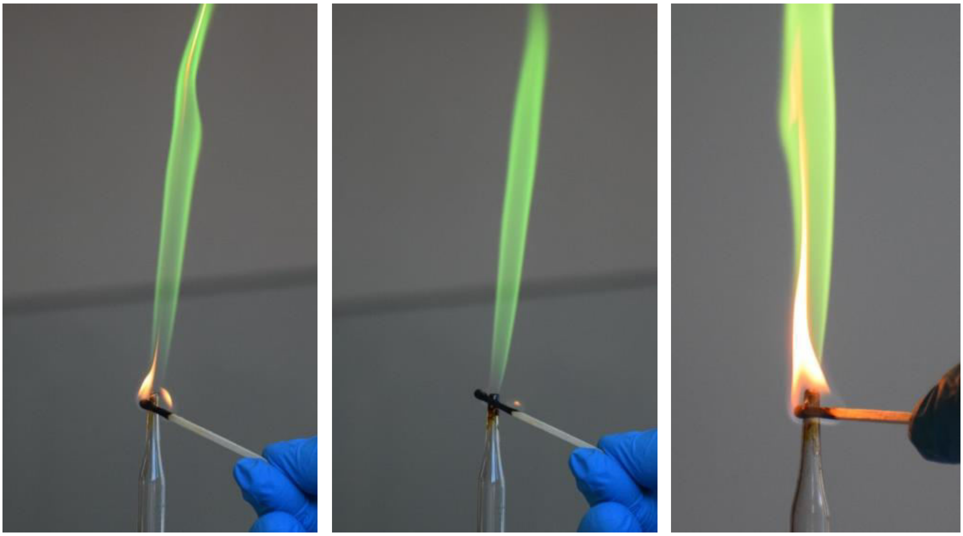


Fig. 1-3 The methyl esters of the boric acid burning with green flame once ignited.

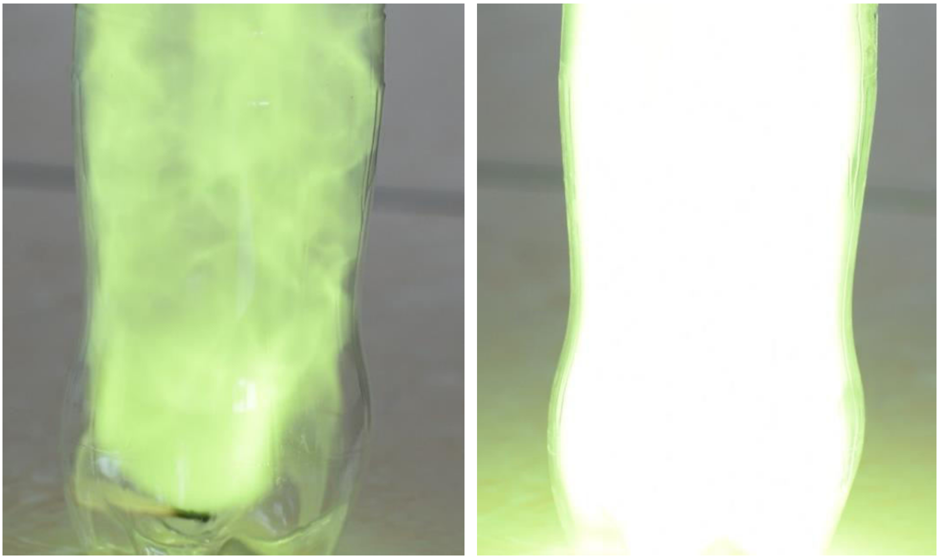


Fig. 4 Two mL of the reaction product, spread in the interior of a 1.5 or 2L PET bottle, burns rapidly (sometimes with an explosion) once ignited.

REFERENCES

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CHEMICAL PROPERTIES OF CARBOXYLIC ACIDS

OVERVIEW

Compounds containing a carboxylic group (-COOH), generally shown with the formula $C_nH_{2n}O_2$ are called carboxylic acids. There are various carboxylic acids and they show differences in their structures and properties. The mostly used and known one is acetic acid. It is known as vinegar acid among the public. When classified with the number/amount of the carboxyl group in their structures; those with only one group of carboxylic acid are called mono carboxylic acids. The reactions of carboxylic acids are :

1. Carboxylic acids, reacting with active metals, release H2 gas, such as Na, K, Ca, Mg. During the release of gas, the product formed is a colorless compound of crystals that can dissolve in water but that cannot dissolve in alcohol and ether. These are used in the drug industry and analytical chemistry.
2. Carboxylic acids give a neutralization reaction with bases, that is, they form salts.
3. Reacting with bases, carboxylic acids constitute carboxylic salts. The hydrogen in the hydroxyl change places with a metal cation. Thus, the acetic acid in vinegar, reacting with sodium bicarbonate, form sodium acetate, carbon dioxide and water.

MATERIALS AND METHODS

Take four test balloons and fill them with water. Put solid sodium carbonate (Na_2CO_3) in the first balloon, solid sodium bicarbonate ($NaHCO_3$) in the second, sodium hydroxide (NaOH) in the third and solid magnesium (Mg) metal in the fourth one. Insert a glass tube through a hole in a cork and fix a hose to its edge. Put the other end of the hose into a beaker which is full of water. Put a small amount of acetic acid solution into the first and second balloons with sodium carbonate (Na_2CO_3) and sodium bicarbonate ($NaHCO_3$). Close the test balloon with the cork and try to observe if there is any gas coming out of the other end of the hose in the water. Do the same experiment with the third and fourth balloon adding formic acid to them.

HAZARDS

HCOOH

CH3COOH

NaOH

CHEMICAL REACTION EQUATION

$$Na_2CO_{3(aq)} + 2CH_3COOH_{(aq)} \rightarrow 2CH_3COONa_{(aq)} + CO_{2(g)} + H_2O_{(l)}$$

$$NaHCO_{3(aq)} + CH_3COOH_{(aq)} \rightarrow CH_3COONa_{(aq)} + CO_{2(g)} + H_2O_{(l)}$$

$$NaOH_{(aq)} + HCOOH_{(aq)} \rightarrow HCOONa_{(aq)} + H_2O_{(l)}$$

$$Mg_{(s)} + HCOOH_{(aq)} \rightarrow Mg(HCO_2) + \tfrac{1}{2}H_{2(g)}$$

CONCLUSIONS

Gas release will be observed as the result of the reaction of sodium bicarbonate and sodium bicarbonate with acetic acid. The gas formed is carbon dioxide (CO_2) formic acid is a carboxylic acid, sodium hydroxide is a base. Salts are formed as a reaction of acids with bases. When formic acid and sodium hydroxide take part in the reaction, sodium format (HCOONa), which is the sodium salt of formic acid forms. The gas released through the reaction of acetic acid with Mg, an active metal, is a flammable gas, H_2.

WASTE MANAGEMENT

The products formed during the release of gas are all salts, so when thinned, they can be discharged through a sink.

PHOTOS

Fig. 1 Na_2CO_3, $NaHCO_3$, NaOH and Mg is added in balloons containing water in a row and stirred until dissolved.

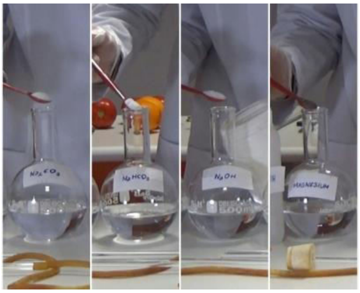

Fig. 2 When acetic acid is added to Na_2CO_3 gas release is observed. This gas is CO_2.

Fig. 3 When acetic acid is added to $NaHCO_3$ gas release is observed. This gas is CO_2.

Fig. 4 When formic acid is added to the solution consisting of NaOH, it is observed that no gas release happens.

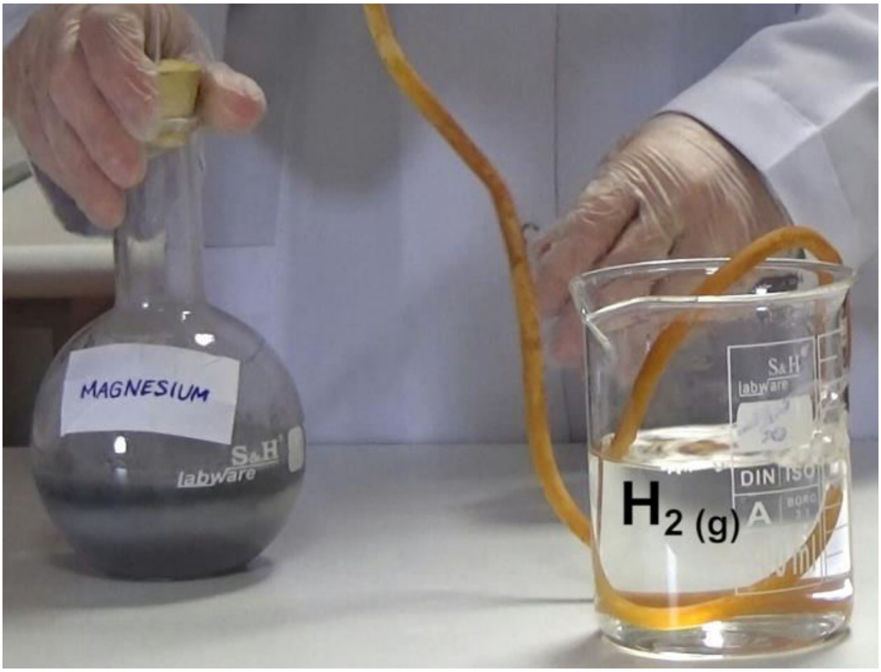

Fig. 5 As a result of the reaction of carboxylic acid with active metals, when formic acid is added, it is observed that H_2 gas will be released.

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SUGAR DEHYDRATION BY H_2SO_4

OVERVIEW

Concentrated sulfuric acid is a common reagent in the chemistry laboratory. Special precautions are needed when handling this chemical as any spill can cause severe burns of the skin and when dropped on cotton fabrics in few hours it creates holes … (on trousers on shirts, on towels). This is due to the high affinity for water showing a strong dehydrating action especially on organic substrates.

Saccharides are taught in secondary years, during chemistry lessons, as a class of compounds consisting of carbon (C), hydrogen (H) and oxygen (O) usually with a H:O ratio of 2:1. Due to this H:O ratio the saccharides are commonly known as carbohydrates. Discussing the structure of different monosaccharides (glucose, fructose) it's obvious that the name carbohydrates is wrong as it suggests a combination of C with water H_2O (not true). [1]

However the following experiment shows the great capacity of concentrated sulfuric acid to take out water from saccharides as it shows a strong dehydrating action. The experiment is suitable both during the lessons when analyzing the structure of saccharides or during the lessons when acids are taught.

MATERIALS AND METHODS

Weigh 70 g of powdered sugar (sucrose $C_{12}H_{22}O_{11}$) and place it in a tall 250 mL beaker. Measure 70 mL of concentrated sulfuric acid (H_2SO_4 98%) and pour it over the sucrose. Place the beaker on a fire resistant support. Mix the two reagents and observe the evolution of the mixture over a period of 10-15 minutes. [2] You will notice at a certain point that steam is evolving from the beaker and you can smell the caramelized sugar odor.

HAZARDS

Sulfuric acid is very corrosive and as shown in the experiment destroys organic substances, tissues. Handle with caution vessels containing concentrated sulfuric acid.

H_2SO_4

CHEMICAL REACTION EQUATION

$$C_{12}H_{22}O_{11(s)} + H_2SO_{4(l)} \longrightarrow 12C_{(s)} + \text{solution } H_2SO_{4(aq)}$$

CONCLUSIONS

Sulfuric acid is a very powerful dehydrating agent. As sucrose is mixed with concentrated sulfuric acid a white yellowish paste is formed which gradually darkens through brown to pitch black. This is due to carbon formation from the organic substance (sucrose) as water is removed from the molecules. At a certain moment the black residue starts to expand when steam evolvement is noticed. The residue grows double the height of the beaker and sometimes it's advisable to support it with a glass stirring rod.[3,4]

WASTE MANAGEMENT

Wash with plenty of water the black residue (mostly formed by carbon) and dispose it in a waste container.

PHOTOS

Fig. 1 The sugar mixture with sulfuric acid will change its color first.

Fig. 2 After a few minutes a black column of carbon emerges from the beaker.

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REACTION OF SUCROSE WITH POTASSIUM NITRATE OR HOW TO MAKE PROPELLANTS FOR MODEL ROCKETS

OVERVIEW

Although this experimental procedure might look different comparing with experiment called *Sucrose oxidation by potassium chlorate* there are many similarities between these two processes. In both cases it's about a very exothermic reaction between sugar and a strong oxidizing agent: in this case the potassium nitrate. In both cases the reaction can be violent so caution is needed both concerning quantities and the place where it is performed.

Due to its exothermicity and abundant gases release the reaction is used to make rocket engines for model rockets which are capable to fly even hundreds of meters in the sky [1].

The reaction can be performed with the class when studying the saccharides, thermochemistry (exothermic reactions) or the properties of the salts (potassium nitrate), or when discussing why a substance acts as an explosive.

MATERIALS AND METHODS

Weigh 7.5 g of KNO_3 and 3.5 g of sugar (sucrose). Grind both solids in order to reach a smaller granulation. Pour both solids on a piece of paper and mix them carefully by moving the paper to and fro, then place the mixture in a form of a cone in a well ventilated place on a fire resistant support. Ignite the mixture carefully using a match. Observe the intense heat generated during the reaction and feel in the air the burning sugar odor.

For an even enhanced effect of the blaze add to the mixture 1.0 g of $NaNO_3$ for a yellow flame, 1.0 g of $Ba(NO_3)_2$ for a green one and a Li salt for red. [2]

HAZARDS

The reaction is strongly exothermic due to the strong oxidizing capacity of the potassium nitrate. Mix carefully the two reagents and avoid friction.

KNO_3

CHEMICAL REACTION EQUATION

$$5C_{12}H_{22}O_{11(s)} + 48KNO_{3(s)} \text{--->} 24K_2CO_{3(s)} + 36CO_{2(g)} + 55H_2O_{(g)} + 24N_{2(g)}$$

CONCLUSIONS

As we can see in the equation of the chemical reaction a large amount of gases is generated which explains the use of this process in the field of model rocketing. The reaction is strongly exothermic and a fire proof support should be used. Sometimes the ceramic tile on which this reaction is performed is breaking apart due to the heat. The sugar acts as a fuel while the potassium nitrate furnishes oxygen allowing the energetic oxidation process in which the C atoms of the sucrose are transformed in CO_2, the H atoms give vapors of H_2O and the nitrogen of the nitrate add a third gas, the N_2.

WASTE MANAGEMENT

The waste remaining after the reaction is basically formed by potassium carbonate, burned sugar and maybe caramelized sugar. They show no potential danger to the environment and may be discarded in the bin.

PHOTOS

Fig. 1 Exothermic reaction between sugar and potassium nitrate with an abundant evolvement of gases.

REFERENCES

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STARCH DETECTION IN BREAD AND POTATO

OVERVIEW

In this experiment we detect starch in bread and potato. Starch is a carbohydrate stock plant. The starch granules are insoluble in cold water, yet they have the capacity to absorb water and swell. By hot water a dense solution (gelation) is formed containing two forms, amylose and amylopectin. Amylose is a linear polymer composed of many molecules of glucose. Amylose is capable of engaging by forming a helical structure around iodine thus shows a light absorption capacity in the resulting blue colour [1].

MATERIALS AND METHODS

In a beaker put one hundred mL of water and add a few drops of an alcoholic iodine tincture. Then place one slice of purified potato in the beaker. Note that within minutes the potato turns blue. Repeat the same procedure with a piece of bread. Observe that the bread becomes blue. The potato and the bread contain starch which is a polysaccharide. The starch is detected by the impact of iodine solution.

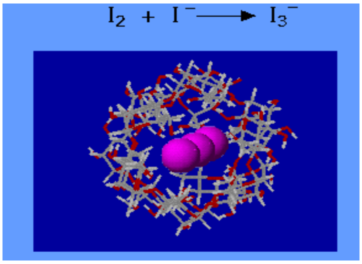
HAZARDS

Attention should be given to the fact that CH_3CH_2OH is flammable and it should not come in contact with an open flame.

CH_3CH_2OH

I_2

CHEMICAL REACTION EQUATION



[2]

CONCLUSIONS

Amylose molecules consist of single, mostly unbranched chains of glucose molecules, shaped like a spring. It is speculated that the iodine gets stuck in the coils of the beta amylose molecules (soluble starch). The starch forces the iodine into a linear arrangement in the middle groove of the amylose coil. There is some transfer of charge between the starch and the iodine. This changes the electron arrangements hence the spacing between energy levels. The new spacing absorbs visible light selectively and gives the complex its intense blue colour. The effect is only seen when both iodine as an element and iodide as an ion are present. Iodine is not very soluble in water and the addition of iodide makes it soluble. Iodine, together with the iodide ion, forms a complex which dissolves in water, unlike iodine on its own. Molecular iodine reacts with iodide which is a negatively charged ion and creates an anion. The anion dissolves easily in water (which is polar). There is still some dispute about the exact mechanism involved in producing the unmistakable colour change, but this charge transfer process is widely accepted as the most likely one. Iodine inside amylose resulting in a blue-black colour [3].

WASTE MANAGEMENT

Dispose the solution in a halogenated organic waste container. Pour the filtrate into a waste container designated for this purpose. Throw bread and potato in a trash.

PHOTOS

Fig. 1 The color of bread and potato turns into blue after the dip in the iodine solution. This is more obvious in bread because it has a higher content of starch than potato.

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SILVER MIRROR USING DIFFERENT SUGARS

OVERVIEW

Silver nitrate is the most common salt of silver. This is produced by dissolving the Ag in HNO_3. It forms colourless crystals, dissolved in water. At 440°C it decomposes in metal, oxygen, nitrogen and nitrogen oxides. It is easily reduced by various reducing agents to metallic silver [1].

Glucose belongs to monosaccharides which are aliphatic polyhydroxy-aldehydes and ketones. The D-glucose is a hexose which is a natural product or a component of many oligosaccharides, polysaccharides and glycosides. Industrially it is produced by hydrolysis of starch. Sucrose is the common sugar and it is very rife in nature. It is found in fruits, sugar cane, beets, honey, grape etc. This is not a reducing sugar because the hemiacetal hydroxyl is not free. By hydrolysis, acidic or enzymatic sugar cane is converted to glucose and fructose. This phenomenon is called inversion. It is used in food industry [2].

In this experiment, we examine the existence of any aldehyde group in different sugars by Tollens’ reaction. Tollens' reagent is a chemical reagent used to determine the presence of aldehyde or alpha-hydroxy ketone functional groups. The reagent consists of a solution of silver nitrate and ammonia. It was named after its discoverer, the German chemist Bernhard Tollens. A positive test with Tollens' reagent is indicated by the precipitation of elemental silver, often producing a characteristic "silver mirror" on the inner surface of the reaction vessel [3].

MATERIALS AND METHODS

In a beaker place 2 mL silver nitrate solution (20%w/v) and two drops of sodium hydroxide (1M). Observe that a dark precipitate is formed which is silver oxide. Then add such an amount of NH_3 (25% weight solution) as necessary to dissolve the silver oxide and form the Tollens reagent. After the precipitate is dissolved, add about 10grams of glucose and dissolve. Then heat the mixture on a hot plate to accelerate the reaction. Observe the silver mirror. Then repeat the same procedure for the sucrose which is the common sugar.

HAZARDS

Attention should be given to the fact that NH_3 is flammable and it should not come into contact with an open flame. Wear protective glasses, gloves and use a respirator filter.

$AgNO_3$

NaOH

NH_3

CHEMICAL REACTION EQUATION

$$2AgNO_{3(aq)} + 2NaOH_{(aq)} \rightarrow Ag_2O_{(s)} + 2NaNO_{3(aq)} + H_2O_{(l)}$$

$$Ag_2O_{(s)} + 4NH_{3(aq)} + 2NaNO_{3(aq)} + H_2O_{(l)} \rightarrow 2[Ag(NH_3)_2]NO_{3(aq)} + 2NaOH_{(aq)}$$

$$C_6H_{12}O_{6(s)} + 2[Ag(NH_3)_2]NO_{3(aq)} + H_2O_{(l)} \rightarrow 2Ag_{(s)} + C_5H_{11}O_5COOH_{(aq)} + 4NH_{3(aq)} + 2HNO_{3(aq)}$$

$$C_{12}H_{22}O_{11(s)} + 2[Ag(NH_3)_2]NO_{3(aq)} + H_2O_{(l)} \rightarrow X$$

CONCLUSIONS

Some minutes after the beginning of heating, the silver mirror is formed due to the oxidation of the aldehyde group which is present in the open -chain form of glucose, because the carbon one is a part of an aldehyde group. Glucose is a reducing sugar and turns into glucuronic acid. In the tube with sucrose we observe that a silver mirror is not formed. The sucrose is a disaccharide of glucose and fructose, which does not contain any aldehyde group thus does not oxidize, as it is a non-reducing sugar.

WASTE MANAGEMENT

Solutions should be disposed in the appropriate waste containers in the laboratory. All glassware must be rinsed at least four times with large volumes of water. If recovery of the silver is desired, acidify the solution with nitric acid before attempting recovery. The flask, or other glassware, can be cleaned with concentrated nitric acid.

PHOTOS

Fig. 1 The silver mirror is formed in the beaker with glucose due to the oxidation of the aldehyde group of $C_6H_{12}O_6$ which is present in the open -chain form of glucose. The silver mirror is not formed in the beaker with sucrose because sucrose is a disaccharide of glucose and fructose, which does not contain any aldehyde group thus does not oxidize, as it is a non-reducing sugar.

Fig. 2 The silver mirror

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3. https://en.wikipedia.org/wiki/Tollens%27_reagent

FEHLING REACTION WITH DIFFERENT SUGARS

OVERVIEW

Glucose belongs to monosaccharides which are aliphatic polyhydroxy-aldehydes and ketones. The D-glucose is a hexose which is a natural product or a component of many oligosaccharides, polysaccharides and glycosides. Industrially, it is produced by hydrolysis of starch. Sucrose is the common sugar, very rife in nature. It is found in fruits, sugar cane, beets, honey, grape etc. This is not a reducing sugar because the hemiacetal hydroxyl is not free. By hydrolysis, acidic or enzymatic sugar cane is converted to glucose and fructose. This phenomenon, which is used in food industry, is called inversion [1].

In this experiment, we examine the existence of aldehyde group in different sugars using the Fehling reaction. Fehling's solution is a chemical reagent used to differentiate between water-soluble carbohydrate and ketone functional groups. It is a test for reducing sugars and non-reducing sugars, supplementary to the Tollens' reagent test. The test was developed by German chemist Hermann von Fehling in 1849. Fehling's can be used to distinguish aldehyde vs ketone functional groups. The tested compound is added to the Fehling's solution and the mixture is heated. Fehling's can be used to screen glucose in urine, thus detecting diabetes [2]. Fehling test includes two different solutions, Fehling A, which is a solution of copper sulphate, Fehling B, which is a solution of sodium hydroxide and potassium tartrate.

MATERIALS AND METHODS

Place in the two different beakers 30 mL of Fehling A (69,2g $CuSO_4 \cdot 5H_2O$ + 1 ml/l of H_2SO_4 (95-97% w/w), and 30ml of Fehling B (352g $C_4H_4O_6KNa \cdot 4H_2O$ + 100 g/l NaOH.). In a beaker add 5 grams of sucrose and in the other beaker 5 grams of glucose and stir until they are dissolved. Place both beakers on a hot plate to accelerate the reaction. After some time we observe the mixtures. In the first beaker a brick-red precipitate is formed.

HAZARDS

Wear protective glasses, gloves and use a respirator filter.

$CuSO_4$		**NaOH**	
Cu_2O		**Na_2SO_4**	

CHEMICAL REACTION EQUATION

$$C_6H_{12}O_{6(s)} + 2CuSO_{4\,(aq)} + 5NaOH_{(aq)} \rightarrow$$
$$C_6H_{11}O_7Na_{(aq)} + Cu_2O_{(s)} + 2Na_2SO_{4(aq)} + 3H_2O_{(l)}$$

$$C_{12}H_{22}O_{11(s)} + 2\,CuSO_{4(aq)} + 5NaOH_{(aq)} \rightarrow \text{X}$$

CONCLUSIONS

Sometime later, after having added the glucose, we observe that in the beaker a brick red precipitate from copper oxide is formed due to oxidation of the aldehyde group present in the open -chain form of glucose, because carbon is a part of an aldehyde group. Glucose is a reducing sugar and turns into a salt of glucuronic acid. On the contrary, the beaker to which sucrose is added, the reaction is not carried out because sucrose is a disaccharide of glucose and fructose, which does not contain aldehyde group, thus does not oxidize, as it is a non-reducing sugar.

WASTE MANAGEMENT

The solutions should be filtered. All solutions are placed in the waste container used for inorganic solvent waste as they contain copper. The red precipitate Cu_2O can be dissolved in concentrated hydrochloric acid.

PHOTOS

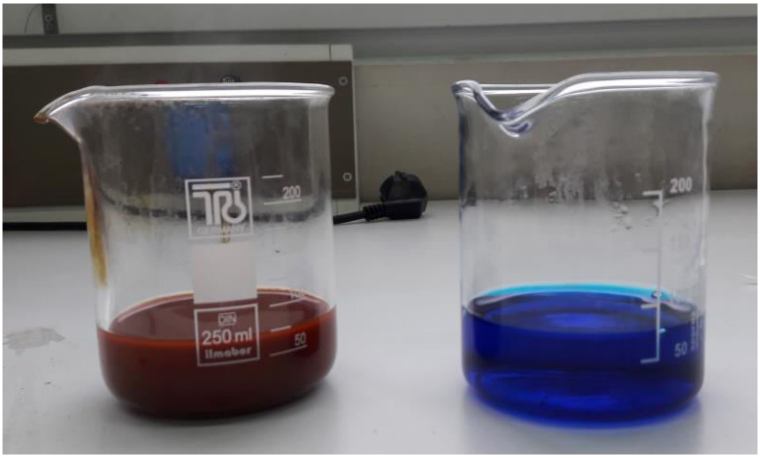

Fig. 1 The brick red precipitated from copper oxide due to oxidation of the aldehyde group in glucose. In the other beaker the reaction was not carried out.

REFERENCES

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DETERMINATION OF GLUCOSE THE FEHLING TEST

OVERVIEW

Glucose, a monosaccharide, is one of the vital carbohydrates for a living. The cells use it as a source of energy and a sub-product in metabolic reactions. Glucose is the basic product of photosynthesis and cellular respiration starts with it. Monosaccharides have a kind of determining reaction, such as, related with the reducing feature of monosaccharides, reaction of colour and forming osazon. In a light alkali medium, during the carboxylic acidity increase of the anomeric carbon of monosaccharides including free half acetyl hydroxyl, the sugar acts as a reducing agent. It reduces ions such as Cu, Ag, Hg, Bi. Such sugars are called as reducing/reductive sugars. Fehling reaction and Benedict reaction are used in the determination of reducing/reductive sugars. The Fehling test, making use of the reducing feature of the monosaccharides, is based on the principle of reducing Cu^{2+} into Cu^{+}.

MATERIALS AND METHODS

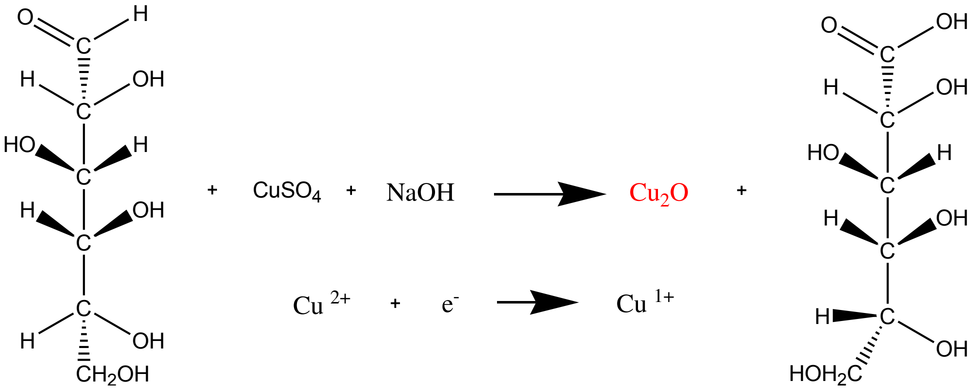
Get 2 test tubes. Put 1 mL Fehlıng I and 1 mL Fehling II solution into the two tubes. Add 4 mL distilled water into one of the tubes and 4 mL of glucose sample. Keep the tubes in hot water. Record the result.

HAZARDS

Fehling I 

Fehling II 

CHEMICAL REACTION EQUATION



CONCLUSIONS

A brick-red coloured precipitate will be observed in the test tubes with positive result.

WASTE MANAGEMENT

The Copper(I) Oxide (C_2O), after thinning, is discharged to the lavatory. Since the Fehling I solution contains Copper II Sulphate ($CuSO_4$), it is harmful to the environment, for this reason it must not be released to the environment. Both solutions (Fehling I, Fehling II) must be stored in leak proof plastic waste containers.

PHOTOS

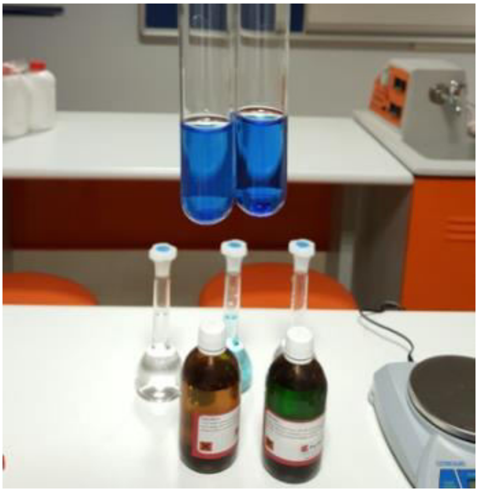


Fig. 1 Put equal volumes of Fehling I and Fehling II solutions in two test tubes.

Fig. 2 Keep one as blank probe and add glucose solution in the second one. The brick red precipitate is a positive test for aldoses.

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SUCROSE OXIDATION BY POTASSIUM CHLORATE OR HOW TO BURN A JELLY BEAR

OVERVIEW

In the experiment "*Reaction of sucrose with potassium nitrate or how to make propellants for model rockets*" we presented the oxidizing properties of the potassium nitrate towards the saccharides (in particular toward sucrose or table sugar). If the reaction is performed with potassium chlorate $KClO_3$ a more intense process occurs. As an adaptation of this reaction we don’t use simple table sugar but a sugary product: a jelly bear.

The reaction between sucrose and $KClO_3$ can be performed in several occasions: during the study of redox reactions, when teaching carbohydrates (saccharides) or when speaking, in different contexts, about salts and their properties. Interesting information can be given to the students about the use of $KClO_3$ in early firearms percussion caps or to signal, when mixed with lactose and rosin (white smoke), the election of the pope. [1]

MATERIALS AND METHODS

Perform the reaction in a wide test tube (ϕ25 mm) in a ventilated hood behind a protective screen. Fix the test tube with a clamp on a laboratory stand on a slight angle oriented toward a neutral zone. Place in the test tube 4-6 g of potassium chlorate $KClO_3$. Melt the salt with a Bunsen or Teclu burner. Add, using a crucible tongs, a jelly bear to the melted salt. You will notice instantly a strange noise, the jelly bear will dance in the test tube as the reaction site becomes intensely hot glowing and abundant smoke is evolved. You can notice the smell of the caramelized sugar due to the sugar burning at the reaction site.[2] Other approaches make use of concentrated sulfuric acid H_2SO_4 to initiate the reaction when the two solids are mixed and placed as a pile on a fire resistant support. [3]

HAZARDS

$KClO_3$ is a very potent oxidizing agent. It reacts with any combustible when hot. It's harmful and toxic to aquatic life with long lasting effects. As a large amount of gases is evolved: carbon dioxide and water vapors, molten potassium chlorate can be thrown out of the test tube.[2]

$KClO_3$

CHEMICAL REACTION EQUATION

$4\ KClO_{3(s)} \text{ ---> } 3\ KClO_{4(s)} + KCl_{(s)}$

$C_{12}H_{22}O_{11(s)} + 8\ KClO_{3(s)} \text{ --->} 12\ CO_{2(g)} + 11\ H_2O_{(g)} + 8\ KCl_{(s)}$

CONCLUSIONS

The potassium chlorate when heated generates oxygen which, at its turn, is capable to oxidize organic substrates. Hot $KClO_3$ disproportionates into potassium chloride KCl and potassium perchlorate $KClO_4$.[1] Subsequently potassium perchlorate $KClO_4$ decomposes releasing oxygen O_2 and generating a new amount of KCl. The generated oxygen oxidizes the sugar generating large amounts of CO_2 and steam H_2O. The dance of the jelly bear and the growling sound is attributable to this gas evolvement. As the oxidizing reaction of sugar is very exothermic the reaction site becomes incandescent.

WASTE MANAGEMENT

The reaction is not performed on a stoichiometric basis. There is unreacted $KClO_3$ at the end of the process. Boil the products with aqueous HCl, neutralize the solution with NaOH and discard the salt solution in the drain.[2]

PHOTOS

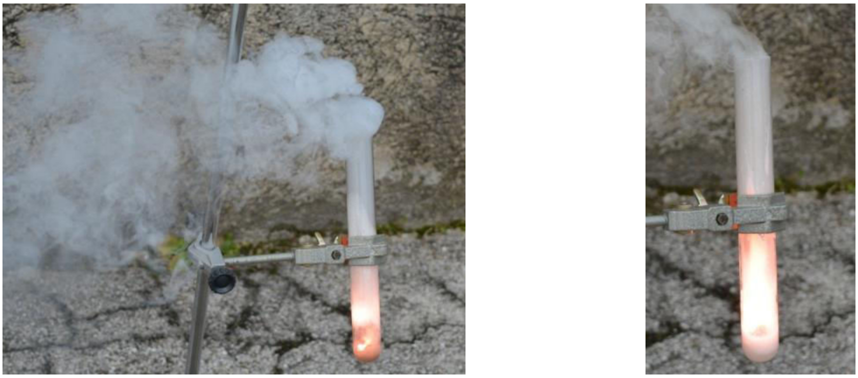

Fig. 1,2 Energetic reaction of molten $KClO_3$ with sugar (in this case a jelly bear) with a zoomed picture of the reaction site.

REFERENCES

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COMBUSTION OF NITROCELLULOSE

OVERVIEW

In this experiment guncotton (nitrocellulose) is produced which, when lighted, burns down producing a glistening flame. The nitrated cotton burns faster because in the exothermic combustion large amounts of gases are formed (nitrogen, nitrous gases, carbon dioxide) which accelerate the reaction. [1]

MATERIALS AND METHODS

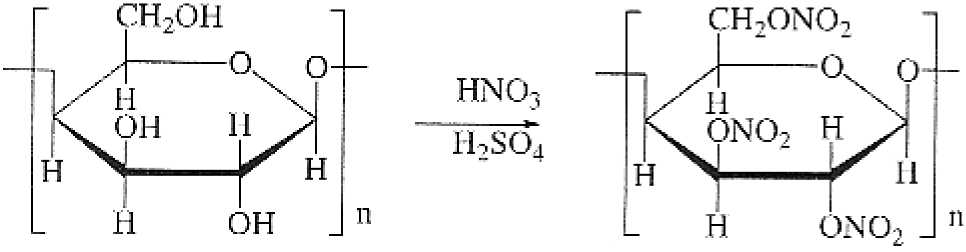
Put a beaker in an ice bath (add some NaCl to the ice water to cool more effective). Put 1,5 g cotton in the beaker and keep the temperature below 5°C (control with thermometer). Mix 20 mL of concentrated sulfuric acid (w = 98%) and 17mL of fuming nitric acid very slowly in the beaker. Knead the cotton with a glass rod for about 10 minutes. Wash with water and wring it out. Pluck the cotton and dry in a desiccator over silica gel. The silica gel has to be replaced until the gun cotton is dried. The prepared guncotton is placed in small portions (no more than about 20 mg) on the open palm of the hand with the help of a second person and lighted. The nitrated cotton burns down in a few seconds.

HAZARDS

When mixing both highly caustic acids it is highly recommended to wear safety gloves and a lab coat that cannot be nitrificated (no cotton). Eye protection is highly requested, because blindness and irritation are caused by both acids if they are in contact with eyes. Skin contact causes acid burn and has to be washed off very quickly. Both acids in combinations form highly explosive reaction products.

H_2SO_4 (98 %) **HNO_3 (65 %)**

CHEMICAL REACTION EQUATION



CONCLUSION

The experiment works best if absorbent (hydrophilic) cotton is used. It is even more likely to succeed when the cotton wool includes much air between the single strings, because it spreads out more easily after nitrating. Nitrated cotton cannot be visibly distinguished from original cotton.

WASTE MANAGEMENT

The mixture of concentrated sulfuric acid and nitric acid is carefully poured into a vessel filled with water. After dilution it may be spilled in to the drain. Guncotton is subject to Explosives Acts and has to be completely destroyed by burning down. Storage is not allowed.

PHOTOS

Fig. 1 Nitrated cotton

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DISSOLUTION OF CELLULOSE IN SCHWEIZER'S REAGENT

OVERVIEW

A very interesting property of tetraammine diaqua copper(II) dihidroxyde $[Cu(NH_3)_4(H_2O)_2]*(OH)_2$, known under the name of Schweizer's Reagent, is its capacity to dissolve cellulose fibers. Moreover, if a solution containing dissolved cellulose is injected in an acidic solution, cellulose is reformed coagulating in longer fibers. This experiment is suitable for the Chemistry classes when studying polysaccharides, other polymers, when discussing some properties of the copper salts and the properties or geometry of different coordinative compounds.

MATERIALS AND METHODS

Synthesis of the Schweizer's Reagent is described in experiment no. 63. Follow the described steps to obtain the reagent but use the following quantities: 5.0 g of $CuSO_4*5H_2O$ dissolved in 90 mL of solution, 0.8 g NaOH dissolved in 10 mL of solution and at the end addition of 50 mL of concentrated ammonia (25%). Pay attention on the filtration step, after precipitating the copper hydroxide, allowing to discard the sulfate ions (SO_4^{2-}) from the reaction mixture.

Dissolution of cellulose in Schweizer's Reagent can be made using medicinal cotton wool which introduced in the Schweizer's Reagent will dissolve almost instantly under constant stirring. Repeat the operation with further amounts of cotton wool until it dissolves in the beaker. Take with a syringe or a Pasteur pipette a few mL of the liquor and gently inject it in a 200 mL beaker containing 100 mL of a 10% solution of sulfuric acid (H_2SO_4). A blue wire is formed which decolorize in a few minutes time.

HAZARDS [1]

$CuSO_4$		**NaOH**	
NH_3		**H_2SO_4**	

CHEMICAL REACTION EQUATION

$$CuSO_{4(aq)} + 2\,NaOH_{(aq)} \text{ ---> } Cu(OH)_{2(s)} + Na_2SO_{4(aq)}$$

$$Cu(OH)_{2(s)} + 4\,NH_{3(aq)} \text{ ---> } [Cu(NH_3)_4(H_2O)_2](OH)_{2(aq)}$$

CONCLUSIONS

The Schweizer's Reagent is well known for its ability to dissolve cellulose [1]. Such a solution, containing dissolved cellulose, when injected in a acid bath, (sulfuric acid 10%) gives a synthetic fiber called rayon (or cuprammonium rayon) which is historically important as it was the first commercially important man made fiber [2]. The coagulation is explained with the regeneration, due to the acidic solution, of the dissolved cellulose. With special machines fine rayon fibers can be obtained allowing their further process until they become a thread.

WASTE MANAGEMENT

The remaining viscous solution is acidified with 1M HCl. The precipitated cellulose is filtered, washed and discarded in the bin. The filtrate, as it contains copper ions and it's harmful for the environment, is disposed in a container for less toxic metallic waste.

PHOTOS

Fig. 1 The Schweizer's Reagent as the only solution capable to dissolve celulose.

Fig. 2 Cotton wool dissolved in the Schweizer's Reagent gives a viscous solution.

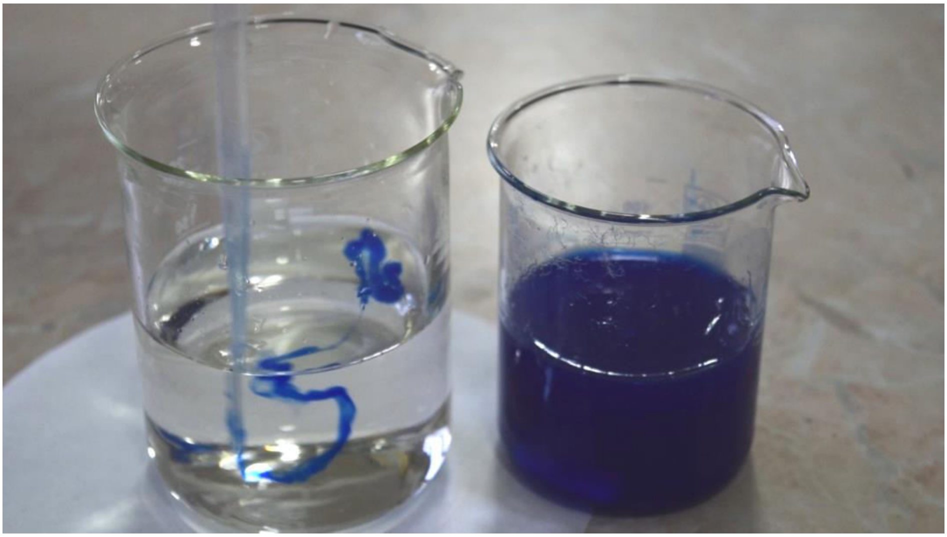

Fig. 3 The cuprammonium solution of cellulose is precipitated using a 10% solution of an acid (in this case H_2SO_4)

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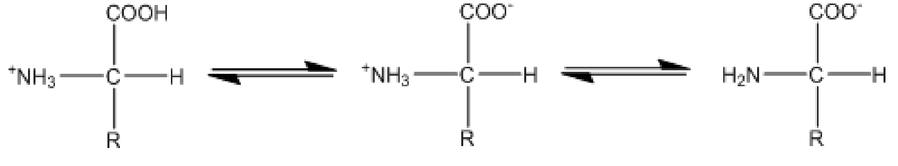
SOME CHEMICAL PROPERTIES OF AMINO ACIDS

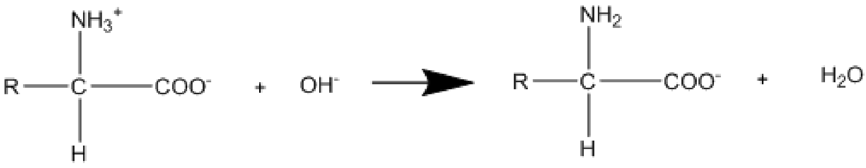
OVERVIEW

Amino acids are essential parts of proteins. They are compounds comprising of both amino groups (-NH_2) and carboxyl groups (-COOH) in their structures. Their general equation is R-CH- NH_2-COOH. Group R is a changeable group. With the changing of group R, 20 kinds of primer or standard amino acid are produced. In their reaction with acids;

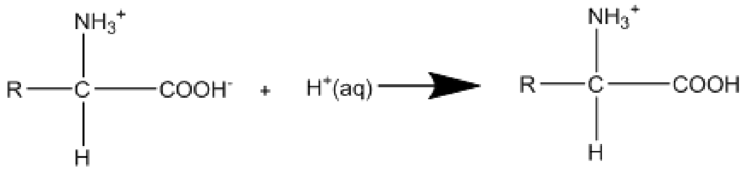


Asidic environment Zwitterion Basic environment

Standard amino acids belonging to a single amino group and single carboxyl group can be crystallized completed in ionized forms known as zwitterion which is a neutral liquid solution. The zero pH value of the net load on an amino acid molecule in the solution is called as iso-electrical point (pI). A standard amino acid does not conduct electricity in a pH medium equal to an iso-electrical point value that is characteristic to it. For this reason, it remains motionless in an electrical field. An amino acid remains in the form of basic anion in pH medium with higher value than an iso-electrical point value. The acid remains in the form of cation in pH medium with lower value than an iso-electrical point. A standard amino acid can act both as a proton donor (acid) and proton acceptor (base).



Notice that this is now a negative ion-no longer a zwitterion

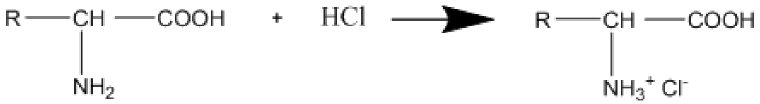


Notice that this is now a positive ion-no longer a zwitterion

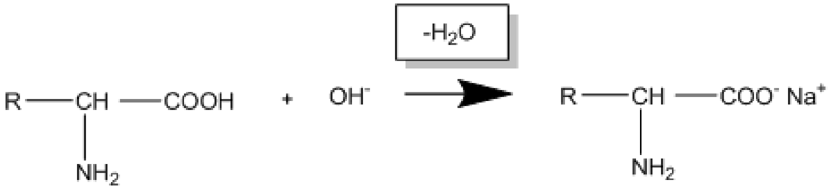
MATERIALS AND METHODS

Put 1 mL of amino acid solution into a test tube. Add 2 drops of HCl (hydrochloric acid) solution and 2 drops of methyl red. Observe the pink colour in the solution. Shake this pink solution with NaOH (sodium hydroxide). In a moment, this pink colour of the solution in the tube will turn yellow.

CHEMICAL REACTION EQUATION



Amino acid chlorine hydrate salt



Sodium salt of amino acids

HAZARDS

HCl 

NaOH 

CONCLUSIONS

On the iso-electrical point, amino acids are found as zwitterion, namely the carboxyl group being negative, and the amino group positive. The molecules at this point will tend to sink down. That is, they are insoluble.

WASTE MANAGEMENT

The yellow coloured solution obtained can be discharged after thinning.

PHOTOS

Fig 1. By adding HCl to the test tube with 1mL of amino acid the acidic amount is increased.

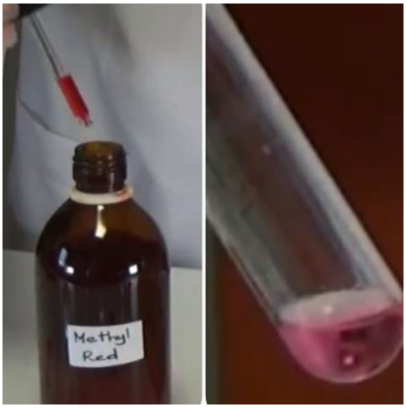


Fig 2. When Methyl red is added to the solution the colour turns to pink, because this solution is pink when in acid.

Fig 3. It is observed that when NaOH solution is being added, the pink colour turns to yellow.

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PROTEIN DENATURATION

OVERVIEW

Protein denaturation … it happens almost every day, in different forms, one of its most common form being the coagulation of egg white due to the heat, but few people are thinking at this process as an irreversible one. The demonstration of protein degradation due to the presence of strong acids, strong bases, even strong electrolytes and of heat can be made when discussing the protein structure and the role of the hydrogen bond or the S-S bonds between cysteines in generating the assembly of the tertiary and quaternary structure of proteins.[1]

IUPAC gave a definition for this process: DENATURATION is a process of partial or total alteration of the native secondary, and/or tertiary, and/or quaternary structures of proteins or nucleic acids resulting in a loss of bioactivity. Denaturation can occur when proteins and nucleic acids are subjected to elevated temperature or to extreme pH, or to nonphysiological concentrations of salt, organic solvents, urea, or other chemical agents [2].

The most common examples of denaturation can be witnessed when a fresh egg or meat are heated, both becoming firm after thermal treatment. If the milk protein is submitted to the action of lactic acid coagulation occurs. Both processes show an irreversible alteration of the protein's structure.

MATERIALS AND METHODS

In order to study the different denaturation factors on proteins a solution of a soluble protein (albumin) is obtained from an egg white. For this an egg white is separated from the yellow core (yolk) in a 150 mL beaker. The egg white is diluted with the same amount of water, is stirred and filtered off. The clear filtrate, the **protein solution**, is used in different demonstrations as follows:

a. Place 5 mL of protein solution in a test tube and heat it.

b. Treat 5 mL of protein solution in a test tube with 5 mL of concentrated sulfuric or nitric acid.

c. Mix in a test tube 5 mL of protein solution with 5 mL of acetone.

d. Mix in a test tube 5 mL of protein solution with a concentrated 7M solution of urea.

HAZARDS

acetone  **HNO_3 (H_2SO_4)**  **(**  **)**

CONCLUSIONS

The protein secondary/tertiary structure is altered due to different external factors. The different interactions between protein chain sequences are altered. The hydrogen bonds which fold the protein in different shapes (alpha helix or beta sheet) are disrupted and the protein folds in a different (mostly random coil) shape. Thus the protein can no longer perform its functions and one of the most evident consequences is that the protein is no longer soluble. This is why a white precipitate appears in the test tube demonstrating the occurred changes. The protein primary structure is not altered as the denaturation does not alter the peptide bond [3]. That means a denaturated protein has the same nutritional value as a raw protein, upon digestion the same amino acids being available for the organism.

All four experiments show the denaturation of soluble albumin due to different influences (thermal, acidic, organic solvents' and concentrated salts). The denaturation mechanism is complex and it's not discussed here. The main conclusion is that proteins are sensitive biomolecules and many external factors can alter their structure respectively their biochemical functions (in most of the cases irreversibly).

WASTE MANAGEMENT

The content of the test tube should be filtered off; the filtrate should be diluted and neutralized with milk of lime than it can be poured down the drain. The filtered solid is to be discarded in the bin. As far as acetone is concerned it should be collected in a recipient for flammable organic liquids but taking in account the small amount used it's safe to discard it in the drain.

PHOTOS

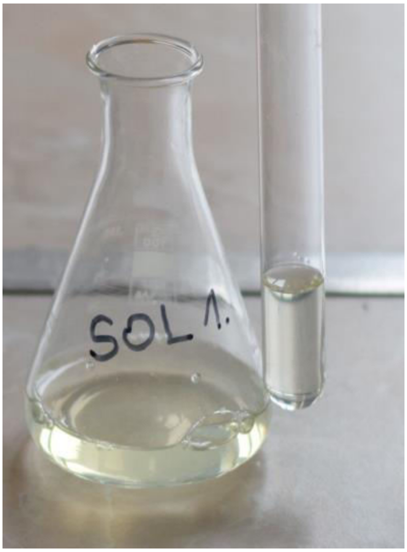

Fig. 1 Clear watery solution of albumin

Fig. 2 Albumin denatured by HNO_3

Fig. 3 Albumin denaturated by heat

Fig. 4 Albumin denaturated by strong acids

Fig. 5 Albumin denaturated by acetone

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THE BIURET TEST FOR PROTEINS

OVERVIEW

This determination of proteins in a basic solution using copper sulfate, known also as *the biuret test*, has analytical value as it determines the presence of the peptide bond hence the presence of soluble protein in solutions. In certain modifications it is not only a qualitative test but it can be used in quantitative measurements, to determine the concentration of the proteins in the range of 5-160 mg/mL [1].

The name ***biuret*** is given because of the resemblance of the peptide bond and its complex with copper to the biuret molecule or carbanylurea ($H_2N\text{-}CO\text{-}NH\text{-}CO\text{-}NH_2$). The beauty of this reaction lies in its simplicity. Everybody are speaking about proteins, they are known as complex biomolecules (which is also true) and few people are aware how simple is to determine their presence in solutions.

MATERIALS AND METHODS

To perform the biuret test aqueous solution of different proteins are needed. In the present experimental work egg, milk, bread and gelatin are used to give protein solutions.

1. Egg white protein solution is obtained after separating the yellow of an egg from its white. Dilute the egg white, in a 150 mL beaker, with the same amount of water, stir it to mix the two phases and then filter it. The filtrate is the SOLUTION1.

2. Take the yolk of an egg in another 150 mL beaker. Add 30 mL of distilled water and stir it well. Filter the mixture the SOLUTION2 being the yellow filtrate.

3. SOLUTION3 is represented by 30 mL of raw milk (which obviously is not a real solution).

4. Take a slice of fresh bread; place a few pieces of core in a beaker containing 50 mL of warm, but not hot, water. Allow to stand for 10 minutes and filter off the mixture. The filtrate is SOLUTION4.

5. SOLUTION5 is obtained by dissolving a spoonful of gelatin in 60 mL of H_2O.

For the test two other solutions are needed: NaOH 10% and $CuSO_4$ 1%.

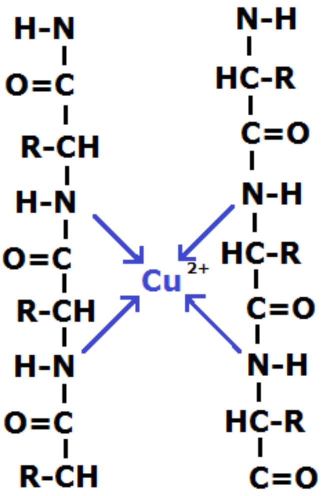
To perform the test take in a test tube 3-4 mL of protein solution (SOLUTION 1 to 5), add the same volume of NaOH solution and in the final add a few drops of CuSO4 1%. Shake the test tube and observe the color. [3] Caution is needed when the copper sulfate solution is prepared as the salt is a pentahydrate. In order to have 1.0 g of $CuSO_4$ we need to weigh 1.562g of $CuSO_4*5H_2O$.

HAZARDS

$CuSO_4$

NaOH

CONCLUSIONS



THE VIOLET BIURET COMPLEX

The test is very sensitive if a soluble protein is present in the solution. Each test is positive if a violet color appears. The color is increasing in intensity as the concentration of the protein is increasing. The violet color is due to the square planar complex of Cu(II) ions formed with the adjacent NH groups from the peptide bonds surrounding the metallic environment. As the color change is drastic it can be observed even in non clear or colored mixtures as the egg yolk and milk which are colored and not transparent.

The literature acknowledge the use of both alkaline bases; either NaOH or KOH [1]. If KOH is used, it has to be freshly prepared and concentrated, otherwise a false negative response will be obtained. With NaOH the reaction is positive even at lower concentrations of the base.

If the original egg white solution or gelatin solution is diluted 2x, 4x, 8x and the reaction is performed again with the same amounts of NaOH and $CuSO_4$ solutions the different shades of violet should demonstrate the applicability of the reaction in the protein’s concentration measurement process.

Try to perform the reaction with any solution containing a soluble protein. We are surrounded with different foods that contain soluble protein. Once the users are acquainted with the test, the determinations are made very easily and the results are quite interesting.

WASTE MANAGEMENT

The content of the test tubes should be discarded in a container for collecting less toxic inorganic compounds as the copper complexes are harmful for the environment.

PHOTOS

Fig. 1 To perform the biuret test take 3-5 mL of protein solution in a test tube, prepare a 10% NaOH solution and a 1% $CuSO_4$ solution.

Fig. 2 Add to the protein solution in the test tube the same volume of 10% NaOH.

Fig. 3 The biuret test is positive if a violet color apperar by adding 1% $CuSO_4$ solution.

Fig. 4 Further addition of $CuSO_4$ gives a more intense violet color.

Fig. 5 Positive biuret test obtained for an egg albumin solution

Fig. 6 Positive biuret reaction for egg yolk solution.

Fig. 7 Positive biuret reaction for milk.

Fig. 8 Positive biuret reaction for a gelatin extract.

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THE XANTHOPROTEIC TEST FOR PROTEINS

OVERVIEW

Another suitable test for identifying proteins is the Xantoproteic Test. It refers to a yellowish color generation after heating a mixture of a protein solution with concentrated nitric acid (60%). If the heating is followed by the addition of a 6M solution of NaOH the color turns orange.

MATERIALS AND METHODS

In order to perform the Xanthoproteic Test a solution of a soluble protein (albumin) is needed. This is obtained in the same way as for the experiment called *"The biuret test"*. For this an egg white is separated from the yellow core in a 150 mL beaker. The egg white is diluted with the same amount of water, is stirred and filtered off. The clear filtrate, the **protein solution**, is used to perform the Xanthoproteic Test. Other protein solutions can be used as well such as: casein, gelatin, etc.

For this 5 mL of protein solution is mixed with 2 mL of concentrated nitric acid HNO_3 60%. Denaturation of proteins occurs giving a white precipitate in the entire volume of the mixture. Heat gently the test tube on an open flame. Observe the changes occurred. Let the content of the test tube to cool down. Add drop by drop a solution of 6M NaOH until the acid is neutralized. Observe the color in this case.

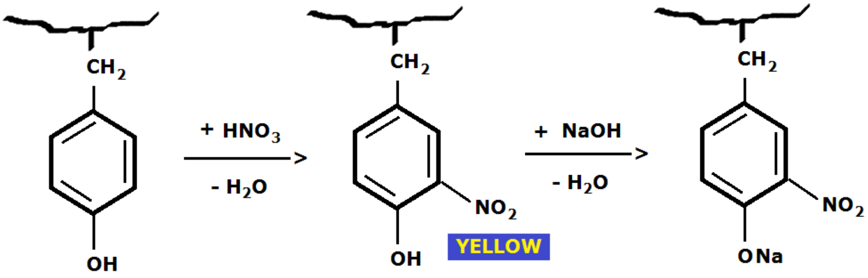
HAZARDS

Caution is needed when handling the concentrated nitric acid as it is very corrosive and releases harmful vapors.

HNO_3

NaOH

CHEMICAL REACTION EQUATION



CONCLUSIONS

The protein secondary/tertiary structure is altered when mixing with strong acids. (For further explanations see the experiment called *Protein denaturation*) This causes the precipitation (coagulation) of the protein giving a white solid mass in the initial solution.

When heated the content of the test tube turns yellow. This is due to the nitration process of the aromatic rings present in the side chains of Tyrosine, Tryptophan and/or Phenylalanine. According to some sources Phenylalanine doesn’t react with hot nitric acid to give xanthoproteic acid. If the mixture is cooled down and concentrated NaOH is added drop by drop the color changes from yellow to orange [1].

If the protein solution is more diluted the degradation process is not observed after the addition of the nitric acid solution, the yellow color still appears but is less intense.

WASTE MANAGEMENT

The content of the test tube should be filtered off; the filtrate should be diluted and neutralized with milk of lime then it can be poured down the drain. The filtered solid is to be discarded in the bin.

PHOTOS

Fig. 1 Clear solution of egg albumin (egg white) used to study the xanthoproteic test.

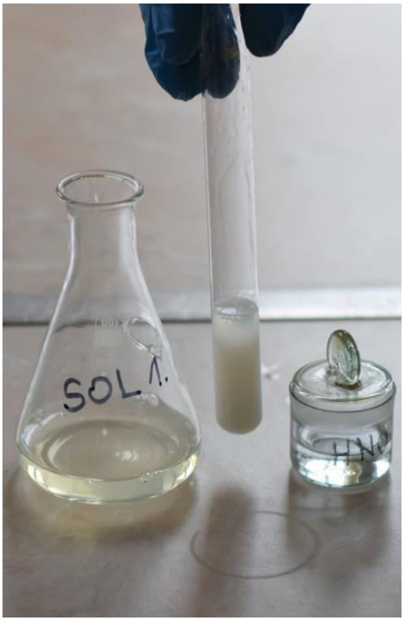


Fig. 2 Egg albumin solution after the addition of 2 mL of concentrated nitric acid (HNO3)

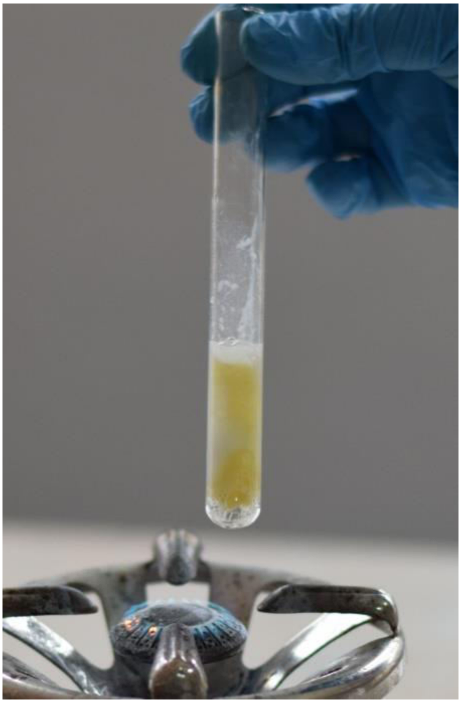


Fig. 3 If the denatured protein is heated, a yellow color is observed due to the nitration of Tyrosine/Tryptophan residues of the protein.

Fig. 4 The addition of the concentrated NaOH (6M) turns the color from yellow to orange.

REFERENCES

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MAKING CHEESE

OVERVIEW

The history of cheese is not clear [1]. We still do not know where the first cheese was produced. What we do know is that recently proof was found that cheese was produced about 7000 years ago in Poland. A theory is that the production of cheese was discovered accidentally; milk was transported in bags made of animal stomachs and probably enzymes still present in the stomachs caused milk to separate in a solid and in a liquid portion. By then, cheese was regarded as a way to conserve milk.

Nowadays a wide variety of cheeses is known. In this experiment we will restrict ourselves to the production of Gouda cheese[2].

The first step is acidification of the milk that causes lactose to convert into lactic acid (reaction 1). The second and most important step in the production of cheese is curdling the milk. Curdling is caused by rennet. Rennet contains a mixture of enzymes (rennin) and can be recovered from the stomach of an unweaned calf. It is a very powerful mixture; one litre of rennet is capable of curdling 10.000 litres of milk at 34°C within 45 minutes! The rennin breaks up proteins by hydrolysis of the bonds between alanine and methionine units (reaction 2). This results in two protein fractions: (i) solid, fat soluble proteins, predominantly casein, and hydrophilic proteins (whey proteins). After separation of water and whey proteins, the remaining solid has to be processed further. These further steps have to be controlled very carefully in order to ensure an excellent product. Therefore, making cheese is just as well an art as a science.

MATERIALS AND METHODS

Requirements:

Milk, lactic acid culture/buttermilk, rennet, pan, heater, thermometer, barrels, cheese press, salt, colander, ladle, beakers, knife.

Procedure:

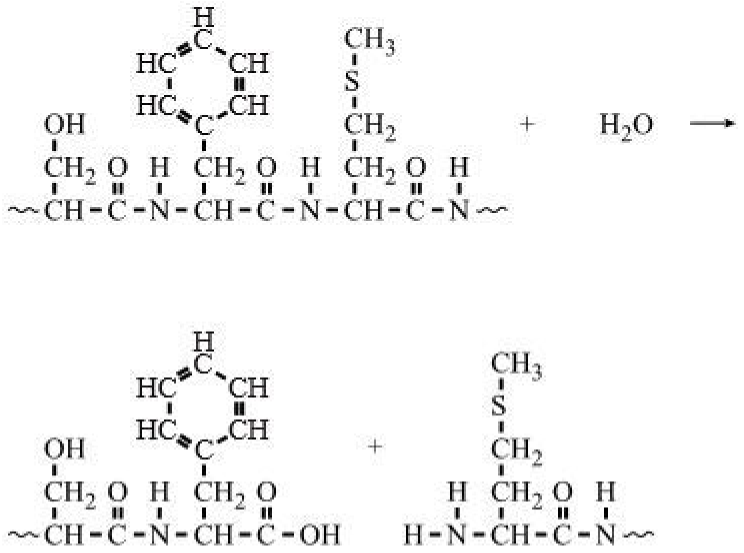
1. put 4.5 L of milk in a pan;
2. add 50 mL of a 1% solution of cheese culture (a lyophilised mixture of bacteria) in milk (or add 100 mL of buttermilk) to acidify;
3. heat the milk at 29oC;
4. add 20 drops of rennet;
5. let the milk curd for 30 minutes;
6. cut the curd and let it deposit for 10 minutes;

7. remove whey (about 1.5 L)
8. heat till 33°C;
9. stir for about 10 minutes;
10. again remove whey (1.5 L)
11. heat till 36°C;
12. stir for about 10 minutes;
13. cover curd and ripen for 30 minutes;
14. Add spices, if desired;
15. Fill barrels without towel;
16. Insert towel and bring the cheese in the barrel upside-down;
17. Leave the barrel upside-down for 10 minutes;
18. Press with 50% of the weight for 30 minutes;
19. Turn the cheese upside -down and press with full weight for 3.5 hours;
20. Ripening without towel for 8 – 10 hours;
21. Salt bath for 6 – 12 hours;
22. Coat the cheese after 24 hours.

From now on it is all a matter of time. The cheese has to be turned upside-down every day. After 4 weeks the cheese is edible; it will have a mild, butter-like taste by then. Who prefers strong taste should have more patience. After a year the cheese is dry and difficult to cut, but with a rich taste.

CHEMICAL REACTION EQUATIONS

$$C_{12}H_{22}O_{11}(aq) + H_2O(l) \rightarrow 4\ C_3H_6O_3(aq) \qquad (1)$$



(2)

CONCLUSIONS

In this experiment we attempted to make Gouda cheese. Evaluation starts with the first taste of our product. And we are not as far yet….

WASTE MANAGEMENT

The only waste produced are Whey proteins (in combination with anabolics a valuable additive for body builders).

PHOTOS

Fig. 1 At the beginning of the experiment

Fig. 2 After separation of the Whey-proteins

Fig. 3 Exerting some pressure at the new-born cheese

Fig. 4 All it takes is time now….

REFERENCES

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FERRIC CHLORIDE TEST FOR PHENOLS

OVERVIEW

The class of phenols, as hydroxyl compounds, is widely studied in the secondary level as there are obvious differences comparing with alcohols.

Useful talks can be conducted in the class concerning the role of phenols (cresols) present in wood smoke used to preserve meat, along with their potential carcinogenic effects. Also the audience is easy to be thrilled when speaking about "*The French paradox*" which refers to the protective role of the polyphenols from red wine to the arteries.

An interesting and easy to make experimental activity, is the ferric test of phenols which give different colors when different phenols are reacted with iron(III) chloride.

MATERIALS AND METHODS

Try to dissolve in 20 mL of water (or 1:1 water+ethanol mixture) around 0.4 g of each of the following phenols: hydrozybenzene (phenol), 1,2,3-trihydroxybenzene (pyrogallol); 1,4-dihydroxy benzene (hydroquinone); 3,4,5-trihydroxybenzoic acid (gallic acid); 2-hydroxybenzoic acid (salicylic acid); sulfosalicylic acid; 2-Naphthol [1]. Any other phenol you might have in the laboratory deserves a try even if not all phenols react with ferric chloride [2]. If a certain phenol is not soluble in water or alcoholic solution then try to dissolve it in a few mL of methylene chloride.

Place each of the resulted solution in a test tube on a row, in a rack and stick a label for each in order to identify the used chemicals. Add a few drops of a 1% solution of $FeCl_3$ to each of the displayed test tubes. Note the colors observed for each phenol.

HAZARDS

Phenol

Ethanol

2-Naphtol

Gallic acid

Hydroquinone

Sulfosalicylic acid

Salicylic acid

Pyrogallol

The phenols are dangerous substances (this is why smoked ham or meat start to be banned from nowadays kitchens as they contain cresols). Phenols are toxic and/or irritants and some of them are harmful for the environment.

CHEMICAL REACTION EQUATION

All authors agree on a complex formation when phenols interact with iron chloride. This is why the most common equation used to describe the reaction between the phenols and Fe(III) [3] should be false [4].

$$\mathbf{3\ Ar\text{-}OH_{(aq)} + FeCl_{3(aq)} \text{--->} (Ar\text{-}O)_3Fe_{(aq)} + 3\ HCl_{(aq)}}$$

a more acceptable representation shows the complex formed around the metallic centre more plausible for the charge transfer which gives the color [4].

$$\mathbf{6\ Ar\text{-}OH_{(aq)} + FeCl_{3(aq)} \text{--->} [Fe(Ar\text{-}O)_6]^{3-}_{(aq)} + 3\ HCl_{(aq)} + 3\ H^{+}_{(aq)}}$$

CONCLUSIONS

The formation of a purple, blue, green or even red color, when diluted $FeCl_3$ is added, is an indication that a phenol is present in the analyzed sample. The ferric solution form a charge transfer complex with the phenols giving a strong spectral sign in the visible domain. The iron ion is hexacoordinated by six phenolato groups. The reaction is very sensitive for all phenols. However for an enhanced color a more concentrated (3%) is advisable for pyrogallol, 2-naphtol and hydroquinone. The observed colors are:

- hydroxybenzene (phenol) - violet
- 1,2,3-trihydroxybenzene (pyrogallol) - red brown
- 1,4-dihydroxybenzene (hydroquinone) - brownish
- 3,4,5-trihydroxybenzoic acid (gallic acid) - black
- 2-hydroxybenzoic acid (salicylic acid) - deep violet
- sulfosalicylic acid -blood red
- 2-Naphthol - green

The dark black color of the gallic acid complex with Fe^{3+} was used in history by people willing to send secret messages as an invisible text, written with a solution of gallic acid, could be made visible at the recipient by spraying a ferric chloride solution on it (for further details see experiment "Secret inks").

WASTE MANAGEMENT

Even if small quantities are used, discard the phenolic solutions in a proper way. Place the waste in a container used to collect organic solutions, as they are harmful, irritant and toxic and some of them harm the environment.

PHOTOS

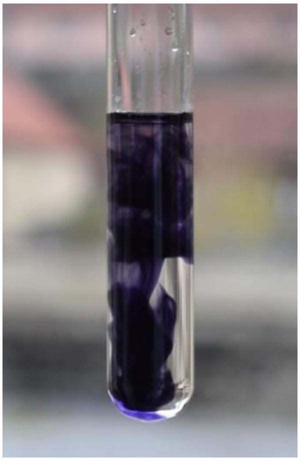

Fig. 1 Phenol

Fig. 2 Hydroquinone

Fig. 3 Gallic acid

Fig. 4 Salicylic acid

Fig. 5 Sulfosalicylic acid

Fig. 6 2-Naphtol

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THE BEILSTEIN TEST AND PLASTICS DETERMINATION

OVERVIEW

The experiment shows a simple way to determine plastics which differ in their density (flotation test) or the presence of a halogen element (Beilstein test) without using any complex analytical methods like IR spectroscopy or melting point determination with a hot stage microscope. [1, 2]

MATERIALS AND METHODS

Plastic materials such as HDPE (high density Polyethylene), LDPE (low density Polyethylene), PS(Polystyrene), PET(Polyethylene terephthalate) and PVC(Polyvinyl chloride) may be used as samples. A copper wire can be obtained from electric insulation material.

Beilstein test

Remove the insulation material from the copper wire fully and heat with a gas burner. Put the plastic sample on the hot copper wire**.** A green flame indicates halogen in the plastic sample (PVC). This experiment must be carried out under a hood because gaseous hydrogen chloride and dioxin compounds may emerge during the combustion of PVC.

Flotation test

Put PET and PS in a beaker filled with saturated NaCl-solution and observe the floating or sinking process. HDPE and LDPE are put in a beaker with water and observe their behaviour. Add methanol to the beakers with HDPE and LDPE and observe again.

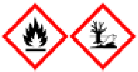
HAZARDS

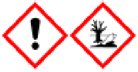
HCl is a poisonous and extremely corrosive gas, but it reacts immediately with CuO to $CuCl_2$. Copper chloride is harmful in contact with skin and if swallowed. Furthermore it causes serious eye damage and is very toxic to aquatic organisms. Copper oxide is also harmful if swallowed and is toxic to aquatic organisms. Methanol is highly flammable and can cause blindness. A fire extinguisher should be available. Protective gloves and protection goggles should be worn.

HCl (g) 

$CuCl_2$ 

Methanol 

Cu 

CuO 

CHEMICAL REACTION EQUATION

$$2\ C_2H_3Cl_{(s)} + 5\ O_{2(g)} \rightarrow 4\ CO_{2(g)} + 2\ H_2O_{(g)} + 2\ HCl_{(g)}$$

$$Cu_{(s)} + \frac{1}{2}O_{2(g)} \rightarrow CuO_{(s)}$$

$$CuO_{(s)} + 2\ HCl_{(g)} \rightarrow CuCl_{2(s)} + H_2O_{(g)}$$

$$2\ C_2H_3Cl_{(s)} + \frac{11}{2}O_{2(g)} + Cu_{(s)} \rightarrow 4\ CO_{2(g)} + 3\ H_2O_{(g)} + CuCl_{2(s)}$$

CONCLUSION

Only PVC shows a green flame in the Beilstein test. It is therefore a reliable test for halogen containing materials. In the flotation test different observations can be made: In sodium chloride solution PS floats and PET sinks. In pure water both HDPE and LDPE float. After adding methanol to the beakers HDPE sinks and LDPE still floats.

Mind that these experiments of plastic determination are not specific for all materials. There are hundreds of types of plastics and most of them have a density in the range of 0.8 to 2 g/cm^3.

WASTE MANAGEMENT

The plastic samples and copper wire can be used for further experiments ($CuCl_2$ reacts to CuO). The liquid waste can be poured in the sink.

PHOTOS

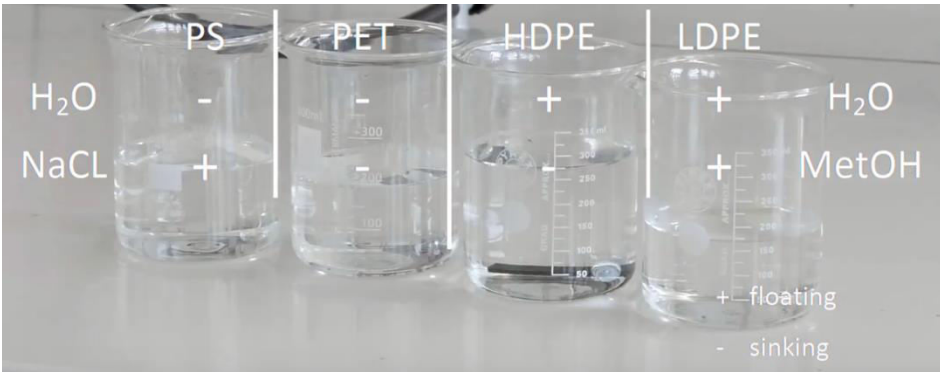


Fig. 1 Result of the flotation test

Fig. 2 Beilstein test: PVC (left), PE (right)

REFERENCES

1. http://www.seilnacht.com/Lexikon/beilst.html(11.02.2016)
2. http://www.sinus-transfer.de/fileadmin/MaterialienBT/Leipzig/Kunststoffe_trennen.pdf (11.02.2016)

SYNTHESIS OF PHENOLPHTHALEIN

OVERVIEW

Phenolphthalein is widely employed as an acid-base indicator. Multiple biological properties of phenolphthalein (contributed to cancerogenic effects observed in various studies) of it are carcinogenic. Exposure of toxic phenol to student's skin should be minimized. The objective of this synthesis is the synthesis product of phenol and phthalic acid anhydride-the clean phenolphthalein.[1,2,3,4]

MATERIALS AND METHODS

Materials

Stand, laboratory claps, laboratory sleeve, magnetic stirrer with integrated heating function, magnetic rod, reflux cooler, inner tubes, security clamps, flat bottomed round butt, Pasteur pipette, Erlenmeyer flask with hose connection, Büchner Funnel, water jet pump

Chemicals

Phenol (s), Phthalic acid-anhydride (s), Deionized water (l), Ethanol p.a. (l), Sulfuric Acid (w/w=96%) [H_2SO_4] (l)

Methods

Set up the synthesis apparatus. Add educts in the flat bottomed round flask (two mol equivalent Phenol (4.71 g) and one mol equivalent phthalic acid anhydride (3.70 g) and 1.5 mL sulfuric acid (96 %). Add a magnetic rod. Start heating and mixing the mixture (at 150-160°C) until it turns deeply red (this might take 15-25 minutes). Add a little bit of water to transfer the whole sample in a liquid phase. After everything is liquidated, add further 150 mL of water (a red-purple solid substance should precipitation). The red-purple solid is phenolphthalein. Separate this solid by tapping machine and wash with 20 mL of boiling ethanol. After this, a completely white powder can be separated from the ethanol and used for e.g. the pH control of an acid/basis reaction.

HAZARDS

Phenol: Dangerous when swallowed and with too long contact with student's skin and inhalation. Further it may lead to genetical malfunctions.

Sulfuric Acid: Causes serious eye damage. Causes severe skin burns and eye damage.

Phenolphthalein: Can cause eye irritations, cancer and cause genetic malfunctions. Suspected of causing cancer.

Phtalic Acid Anhydride: Causes serious eye and skin irritation when in contact and can cause respiratory problems when inhaled.

Ethanol: Highly flammable liquid and vapor. May cause respiratory irritation and drowsiness or dizziness.

Phenol 

Sulfuric Acid (96%) 

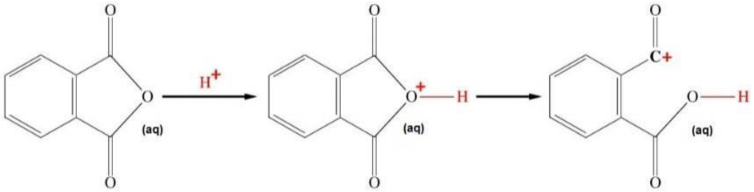
Phenolphthalein 

Ethanol 

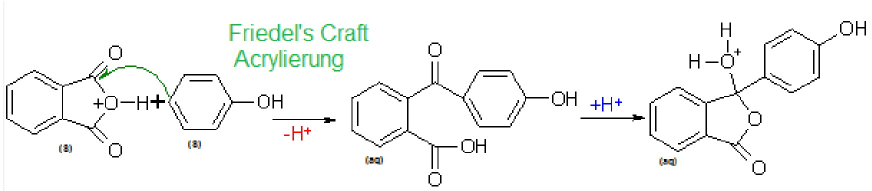
Phtalic acid anhydride 

CHEMICAL REACTION

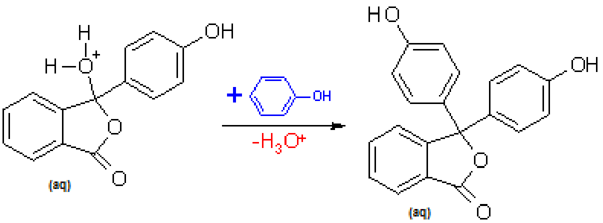
1. Protonation of Phthalic Acid Anhydride

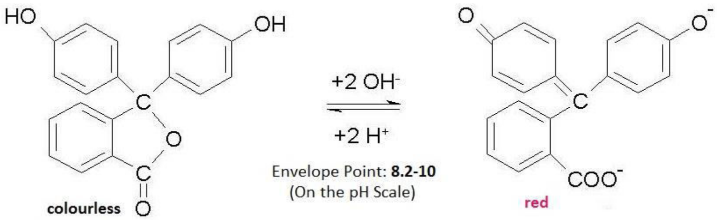


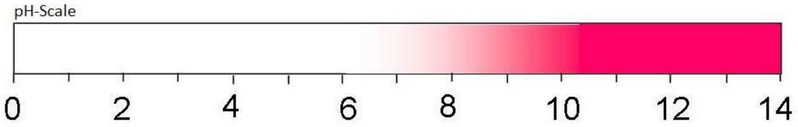
2. Friedel’s Craft Reaction of protonated Phthalic Acid Anhydride with Phenol



3. Addition of second Phenol molecule







CONCLUSIONS

In the first step, the Phthalic Acid Anhydride is protonated with sulfuric acid. In the second step the phenolic compound is added. In the last step one more phenol molecule is being added, which leads to a more stable form of the molecule, because there isn’t an overloaded point in the system This explains the 2 mole phenol: 1 mole phthalic acid anhydride in the method.

WASTE MANAGEMENT

In this synthesis no waste should be produced, except some filtrate from the ethanol precipitation which is collected in a can for water soluble organic solvents.

PHOTOS

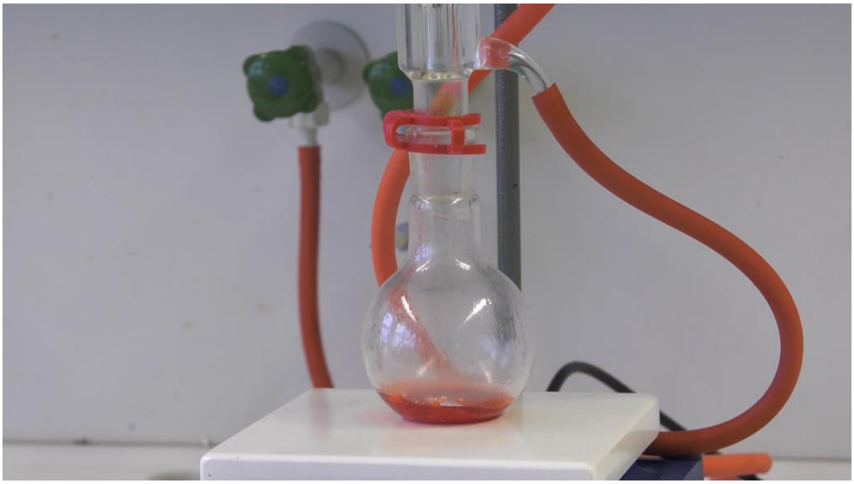


Fig. 1 Precipitation of phenolphthalein (last step of the synthesis)

Fig. 2 Test for alkaline reaction (NaOH and phenolphthaleine)

REFERENCES

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2. http://www.seilnacht.com/Chemie/ch_pheth.html
3. https://www.carlroth.com/de/de
4. http://www.msds-europe.com/id-486-h_p_statements_ghs_clp.html
(all from 31.1.17)

SYNTHESIS OF ASPIRIN

OVERVIEW

In this experiment we will make aspirin out of salicylic acid and acetic anhydride.

MATERIALS AND METHODS

- Salicylic acid ($C_7H_6O_3$)
- Acetic anhydride ($C_4H_6O_3$)
- Phosphoric acid (85%) (H_3PO_4)
- Water
- Ice water
- Beaker
- Pipette
- Pumice
- Scales
- Syringe
- Heating block with sand bath
- Test tube
- Büchner funnel
- Erlenmeyer
- Tripod
- Spatula

Weigh 140 gram of salicylic acid, and place it into a test tube. Add a bit of pumice. Add two drops of phosphoric acid (85%) with a pipette, then add 0,3 mL of acetic anhydride with a syringe. Heat the test tube in a heat block with sand bath, but don’t let it boil. Slowly shake the test tube until all the powder is dissolved. After 5 minutes of heating, add 0,2 mL of water. Then after 2 more minutes of heating add another 0,3 mL of water. Let the test tube cool, until it hits room temperature. Then put the test tube in ice water. Isolate the crystals by filtrating with a small Büchner funnel.

HAZARDS

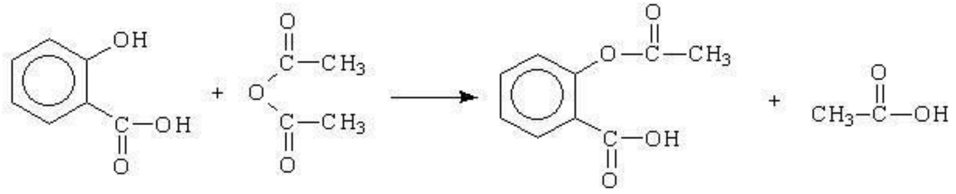
$C_4H_6O_3$

H_3PO_4

$C_7H_6O_3$

CHEMICAL REACTION EQUATION

$$\mathbf{C_7H_6O_3\,_{(s)} + C_4H_6O_3\,_{(l)} \longrightarrow C_9H_8O_4\,_{(s)} + C_2H_4O_2\,_{(l)}}$$



WASTE MANAGEMENT

$C_2H_4O_2$: Throw it into a waste container for other organic substances (OOS). $C_9H_8O_4$: can been thrown into the bin.

PHOTO

Fig. 1 All materials used during the experiment.

REFERENCES

http://www.chem.latech.edu/~deddy/chem104/104Aspirin.htm
http://www.odinity.com/characterization-of-aspirin/

BIODIESEL FROM SUNFLOWER OIL

OVERVIEW

Biodiesel molecules as well as the traditional diesel molecules are basically hydrocarbon chains of about 16 carbon atoms. The structure of biodiesel molecules differs from traditional diesel molecules only at one end by a group known as an "ester." This ester is the dehydrogenation product of a long chain carboxylic acid with methanol. This is why biodiesel is known chemically as "fatty acid methyl ester" or FAME. The biodiesel molecule, freed from the glycerine backbone, is less viscous and burns very well in a diesel engine. Unlike diesel, the ester end makes it much less toxic and much more biodegradable.

To make biodiesel (FAME) we must separate the fatty acid tail from the triglyceride and then add the ester end. This actually occurs in essentially one step, known as "trans-esterification." A catalyst, typically sodium hydroxide (NaOH) or potassium hydroxide (KOH), is dissolved in methanol. This mixture can remove the fatty acid tail from a triglyceride molecule and at the same time add an ester end to it. The methanol is consumed, but the catalyst is not. In general the reaction proceeds at elevated temperatures in order to obtain acceptable reaction rates.

MATERIALS AND METHODS

Materials needed:

Gloves, goggles, aprons, a weight scale, a heater plate, a thermometer, a 250 mL Erlenmeyer flask, a 250 mL separation funnel, a 250 ml beaker, a 100 mL graduated cylinder.

Reagents needed:

Sunflower oil; methanol; KOH flakes.

Biodiesel synthesis:

1. Pour 30 mL of methanol in a 250 ml Erlenmeyer flask;
2. Weigh 1.05 g KOH;
3. Add the KOH to the methanol in the Erlenmeyer flask. Close with a rubber stop and shake until the flakes of KOH are completely dissolved. (The temperature will rise because the reaction is exothermic!);
4. Measure 150 mL of sunflower oil in a 250 mL beaker;
5. Heat the beaker with sunflower oil to 55 - 60 °C;
6. Pour the oil in a 250 mL separation funnel and add mixture of the KOH and methanol;

7. Replace the rubber stop on the separation funnel (make sure it's on tight!!);
8. Shake the separation funnel vigorously for a total of 10 minutes.

Allow the solution to settle. After a minute or two demixing occurs; the glycerine will settle below the diesel.

YOU'VE MADE BIODIESEL!

Biodiesel conversion test (FFA will not fully dissolve in methanol):

When making biodiesel, it's surely important to test the created product to see if you have achieved a full conversion. The test detects di- & triglycerides in the produced biodiesel.

Required for the test:

A sampling vial, two measuring cylinders, a thermometer, methanol, biodiesel.

Performing the test:

1. Measure the temperature of the biodiesel and methanol and make sure that they are both between 19° and 21° C;

IMPORTANT: Temperature is extremely important in this test. If either the biodiesel or methanol isn't at the correct temperature, heat or cool them until they are.

2. Pour 45 mL of methanol in the test vial;
3. Add 5 mL of biodiesel to the test vial;
4. Close the vial & shake it lightly;
5. Leave it for 5-10 minutes;
6. Twist the vial at a 45-degree angle for 10-15 seconds;
7. Look for any demixing at the bottom of the vial.

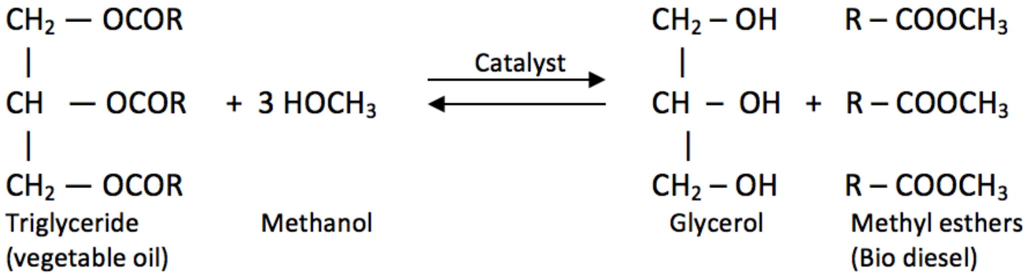
Detection of demixing indicates that some of the oil didn't fully transform into biodiesel. The reaction was successful if there is no noticeable demixing.

HAZARDS

CH_3OH

KOH

CHEMICAL REACTION EQUATION



$$\begin{array}{l} \mathrm{CH_2-OCOR} \\ | \\ \mathrm{CH-OCOR} \\ | \\ \mathrm{CH_2-OCOR} \end{array} + 3\,\mathrm{HOCH_3} \xrightleftharpoons{\text{Catalyst}} \begin{array}{l} \mathrm{CH_2-OH} \\ | \\ \mathrm{CH-OH} \\ | \\ \mathrm{CH_2-OH} \end{array} + \begin{array}{l} \mathrm{R-COOCH_3} \\ \\ \mathrm{R-COOCH_3} \\ \\ \mathrm{R-COOCH_3} \end{array}$$

WASTE MANAGEMENT

Dispose of all liquid residues by pouring them into a can or drum designated for waste collection.

PHOTOS

Fig 1. Requirements

Fig 2. Separation

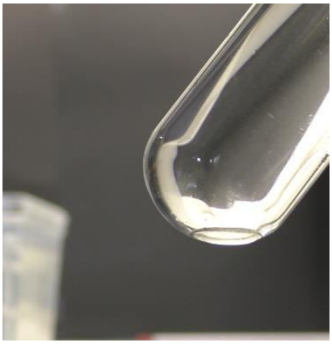


Fig 3. Biodiesel conversion test

REFERENCES

1. http://journeytoforever.org/biodiesel_make.html

SOAP INACTIVATION BY HARD WATER

OVERVIEW

Everyone comes in contact with tap water. However, the calcium ion concentration in the water is different in every region. Water with a relatively high concentration of calcium ions goes by the name “hard water”. Whenever you do the dish washing, soap will eventually be mixed with water, creating a layer of foam. What influence does the water’s calcium ion concentration have on the formation of the foam layer?

MATERIALS AND METHODS

- 4 Erlenmeyer flasks with cork
- Distilled water
- 200 mg calcium ions/L (by using $CaCl_2$)
- Green soap solution
- Measuring cylinder
- Pipette
- Ruler

Firstly, pour 10 mL of calcium ion solution into the first Erlenmeyer flask.
Secondly, add 80 mL of distilled water into this Erlenmeyer flask. Then you add 10 mL of green soup solution into this Erlenmeyer flask. Shake the flask carefully with the cork on it. Measure the formed layer of foam precisely in centimeters. Repeat these steps for 3 other Erlenmeyer flasks, but with the following concentrations instead: For flask 2: 30 mL of calcium ion solution, 60 mL of distilled water; for flask 3: 50 mL of calcium ion solution, 40 mL of distilled water and for flask 4: 70 mL of calcium ion solution, 20 mL of distilled water. Add 10 mL of green soap solution in all flasks.

HAZARDS

$CaCl_2$

CHEMICAL REACTION EQUATION

$$2\ C_{17}H_{35}COONa_{(s)} + Ca^{2+}_{(aq)} \rightarrow (C_{17}H_{35}COO)_2Ca_{(s)} + 2\ Na^{+}_{(aq)}$$

CONCLUSIONS

The amount of formed foam is dependent on the amount of calcium ions in distilled water. A more concentrated calcium ion solution in distilled water gives a smaller amount of foam.

WASTE MANAGEMENT

After finishing the experiment, pour the solutions into the sink.

PHOTOS

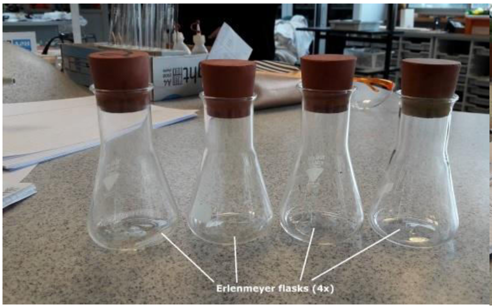


Fig. 1 Four Erlenmeyer flasks, needed for the experiment: 114 – SOAP INACTIVATION BY HARD WATER

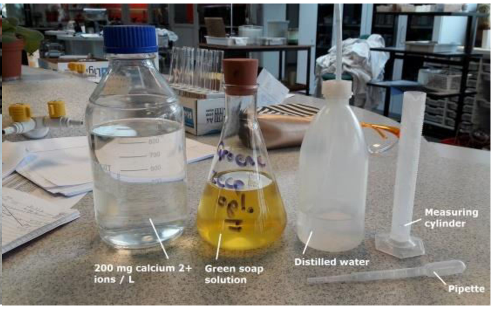


Fig. 2 200 mg calcium^{2+} ions/L; green soap solution; distilled water; measuring cylinder and a pipette, needed for the experiment: 114 – SOAP INACTIVATION BY HARD WATER

REFERENCES

1. http://www.waterhardheid.nl
2. Chemiekaarten (book), 15th edition 2000, pages 790 & 1195
3. https://en.wikipedia.org/wiki/Soap_scum

NYLON SYNTHESIS

OVERVIEW

Nylon can be produced by adding sebacoyl chloride dissolved in benzene to a 1,6 hexanediamine solution.

MATERIALS AND METHODS

- Beakers, 2 of 150 mL and 2 of 50 ml
- 1,6 hexanediamine solution
- Sebacoyl chloride dissolved in benzene
- 2 pincers
- A glass rod

Pour 40 mL of 1,6 hexanediamine solution in a beaker of 150 ml. Gently pour sebacoyl chloride dissolved in benzene on top of the 1,6 hexanediamine solution. A film of nylon forms at the interface, gently pull the nylon string out of the beaker.

HAZARDS

1,6 hexanediamine

Sebacoil chloride

C_6H_6

CHEMICAL REACTION EQUATION

$$C_6H_{16}N_2\ _{(l)} + C_{10}H_{16}Cl_2O_2\ _{(l)} \rightarrow C_{12}H_{22}N_2O_2\ _{(s)} + HCl\ _{(aq)}$$

CONCLUSIONS

After the two chemicals, 1,6 hexanediamine and sebacoyl chloride dissolved in benzene, are put together they, indeed, form nylon.

WASTE MANAGEMENT

Pour all waste from synthesis and solubility tests into container labelled Polymer waste located in the hood.

PHOTOS

Fig. 1 The strands of nylon after the chemical reaction.

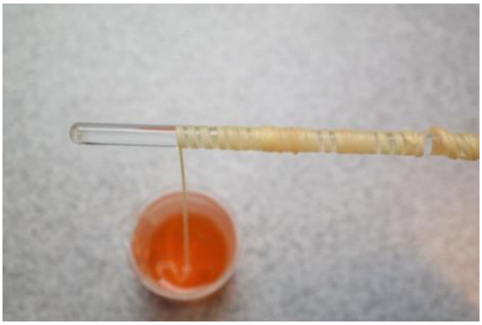


Fig. 2 The nylon attached to a glass rod.

REFERENCES

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HALOFORM REACTION

OVERVIEW

In this experiment we realize a haloform reaction. Substrates that successfully undergo the haloform reaction are methyl ketones and secondary alcohols oxidizable to methyl ketones, such as isopropanol. The only primary alcohol and aldehyde to undergo this reaction are ethanol and ethanal [1]. The haloform reaction is the basis for determining a methyl ketone, a secondary alcohol, ethanol or ethanal through iodoform test in qualitative organic analysis. The sample of an unknown structure is added to a solution containing NaOH and I_2. The formation of the yellow precipitate of iodoform indicates the presence of one of the above compounds [2].

MATERIALS AND METHODS

In a glass test tube add a small quantity of iodine tincture containing I_2 and CH_3CH_2OH, and add a few drops of NaOH 1M. Observe the immediate formation of iodoform.

HAZARDS

Attention should be given to the fact that CH_3CH_2OH is flammable and it should not come in contact with an open flame. In addition, NaOH is very corrosive and one should avoid skin contact. Wear protective glasses and gloves.

CH_3CH_2OH		**I_2**	
NaOH		**CHI_3**	
NaI			

CHEMICAL REACTION EQUATION

$$CH_3CH_2OH_{(aq)} + 4I_{2(aq)} + 6NaOH_{(aq)} \longrightarrow CHI_{3(s)} + HCOONa_{(aq)} + 5H_2O_{(l)} + 5NaI_{(aq)}$$

CONCLUSIONS

Using iodine and sodium hydroxide solution is the most obvious method to chemically detect the presence of ethanol. A positive result is the appearance of a very pale yellow precipitate of iodoform - CHI_3. Apart from its colour, this can be recognized by its faintly "medical" smell. For example, it is used as an antiseptic on the sort of sticky plasters you put on minor cuts [3].

WASTE MANAGEMENT

Combine the contents of the test tube in a beaker, add 1ml of CH_3COCH_3 to destroy any iodide in the test solution and filter the solution to remove CHI_3. Dispose the CHI_3 in a halogenated organic waste container. Pour the filtrate into a waste container designated for the purpose.

PHOTOS

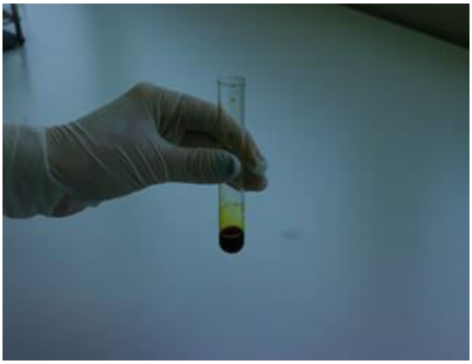

Fig. 1 The beginning of the reaction.

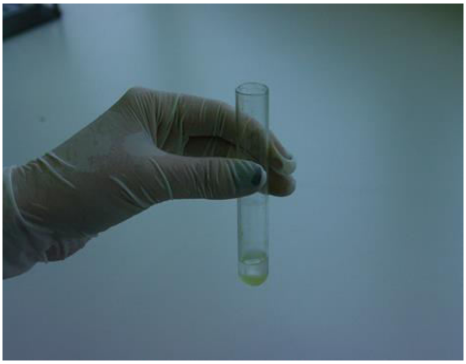

Fig. 2 The yellow precipitate of iodoform.

REFERENCES

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2. Organic Chemistry II, J. Mc Murry, Brooks/Cole Publishing Company, 1996
3. http://www.chemguide.co.uk/organicprops/carbonyls/iodoform.html

pH MEASUREMENT OF ACIDIC AND BASIC SOLUTIONS WITH COLOR-FIXED INDICATOR STICKS

OVERVIEW

In this experiment we measure the pH of various acidic and alkaline solutions using colour-fixed indicator sticks. We can see if a solution is acidic, alkaline or neutral by hydronium ion concentration. Thus we can describe quantitatively the acidity, allowing the concentration of hydronium ions. But since the level of these concentrations are usually very small it serves better to give the acidity function of pH, which is defined as the negative logarithm of the molar concentration of hydronium. The pH is the meter of the acidity of a solution and it is measured in a scale of 1 to 14 [1]. Solutions with a pH less than 7 are acidic and solutions with a pH greater than 7 are alkaline. Pure water is neutral, being neither an acid nor a base. Contrary to popular belief, the pH value can be less than 0 or greater than 14 for very strong acids and bases respectively. The pH of a solution is calculated as follows: $pH=-\log[H_3O^+]$.

One way of measuring the pH is the color-fixed indicator sticks. These sticks are impregnated with indicators and depending on the colour they get, the pH is determined, but not with so much precision as with an electronic pH meter.

MATERIALS AND METHODS

Take five beakers. In the first beaker, place 50 mL of dilute hydrochloric acid (1 M), in the second, one 50 mL of dilute sodium hydroxide (1 M), in the third one 50 mL of water, in the fourth, one 50 mL of lemon juice and the 50, one 50 mL of an aqueous solution of sodium bicarbonate (30%w/v), common baking soda. In each beaker immerse an indicator stick and compare the colour that acquires the indicator stick with the colour of the packaging.

HAZARDS

Wear protective eye shields and gloves. Use a respirator filter.

HCl

NaOH

CONCLUSIONS

Thus the pH of the hydrochloric acid is 0. The pH of the sodium hydroxide is 14. As it is alkaline, the pH of water is 7 because water is neutral, the pH of lemon juice is 3 because it contains citric acid which is a weak organic acid and the pH of the aqueous solution of sodium bicarbonate is 8 because it is alkaline.

WASTE MANAGEMENT

By diluting hydrochloric acid to 5 percent volume to volume or less, and neutralizing it with $NaHCO_3$ to a neutral pH, the solution is no longer corrosive and can, in some jurisdictions, be disposed down the drain. You must partially neutralize the solution of NaOH before be disposed down the drain. All the other solutions can be poured down the drain.

PHOTOS

Fig. 1 The pH of HCl.

Fig. 2 The pH of NaOH.

Fig. 3 The pH of H_2O.

Fig. 4 The pH of lemon juice.

Fig. 5 The pH of $NaHCO_3$.

REFERENCES

1. General Chemistry, E. Gammon, Houghton Mifflin Company, Boston, 2002.

pH DETERMINATION OF ACIDIC AND BASIC SOLUTIONS WITH INDICATORS

OVERVIEW

In this experiment we study the behavior of indicators in alkaline and acidic solutions. A pH indicator is a halochromic chemical compound added in small amounts to a solution so the pH of the solution can be determined visually. Hence, a pH indicator is a chemical detector for hydronium ions or hydrogen ions in the Arrhenius model. Normally, the indicator causes color change of the solution depending on the pH. The indicator is a weak acid or a weak alkaline which in the non-ionized form has a color different from that in the ionized form. For example, if the pointer ionization represented by the equation: $Hln + H_2O \rightleftharpoons In^- + H_3O^+$

then the non-ionized form (Hln) has another color, that of the ionized form (In^-). With gradual acid addition, the balance is shifted to the left where the corresponding color prevails in the solution. In contrast, with addition of base, the equilibrium is shifted to the right so the corresponding color prevails.

To determine the color change range of the index, it is sufficient to write the equation of Henderson-Hasselbach in this case:

$pH = pka + \log [In^-] / [Hln]$ where ka is the ionization constant of Hln indicator.

Experimentally shown that when 90% or more of the indicator is as Hln (i.e. where $[In^-] / [HIn] \approx 0{,}1$) then the HD color prevails in the solution. If 90% or more of the indicator is In- (i.e. when $[In^-] / [HIn] \approx 10$) then the In^- color prevails in the solution. All things considered, the equation of Henderson-Hasselbach is the pH range where the indicator changes color:

$pH = pka + \log [In^-] / [HIn] = pka + \log (0{,}1) = pka - 1$

and

$pH = pka + \log [In^-] / [HIn] = pka + \log (10) = pka + 1$

The above relations show that the indicator changes color between two pH units (when the pH is between pka values + 1 and pka - 1).

So generally:

If $pH < pka - 1$, then there is the Hln color (non-ionized form)

If $pH > pka + 1$, then there is the color of ln^- (ionized form)

If pH = pka then the color is a "mixture" of the color of HIn and In^-.

The amber tea color becomes lighter, when added to lemon juice (citric acid). Juice from red cabbage changes to green and then to yellow, when an alkaline is added. The green and yellow colors change again to red by the addition of acid. The litmus is a common laboratory indicator for alkaline-acids. The pigment which is produced by various species of lichens gets red in acidic solution and blue in alkaline solution. Phenolphthalein which is another laboratory indicator is colorless in acid solution and pink in basic solution. Methyl orange which is another one indicator changes the color to red in acidic solutions and to yellow in alkaline solutions. Finally blue bromothymol changes the color to yellow in acidic solutions and to blue in alkaline solutions [1].

MATERIALS AND METHODS

In five beakers place 20 mL of HCl (1M). In the first beaker add two drops of methyl orange and the color turns red because of the acidity, in the second beaker add two drops of phenolphthalein and the liquid remains colorless because of the acidity, in the third beaker add a few grams of blue bromothymol and the color becomes orange because of the acidity, in the fourth beaker a blue litmus paper is immersed, which turns red because of the acidic conditions, in the fifth beaker add two drops of litmus solution and the color turns light red because of the acidity.

In five beakers place 20 mL of NaOH (1M). In the first beaker add two drops of methyl orange and the color turns yellow because of the basicity, in the second beaker add two drops of phenolphthalein and the color becomes fuchsia because of the basicity, in the third beaker add a few grams of blue bromothymol and the color turns blue because of the basicity, in the fourth test tube a red litmus paper is immersed, which turns to blue because of the alkaline conditions, in the fifth beaker add two drops of litmus solution and the color turns light blue because of the basicity. Thus, we observed the color change of alkaline and acidic solutions using indicators.

HAZARDS

Wear protective eye shields and gloves. Use a respirator filter.

HCl		**NaOH**	
phenolphthalein		**Bromothymol blue**	
methyl orange			

CHEMICAL REACTION EQUATION

$$HIn + H_2O \rightleftharpoons In^- + H_3O^+$$

WASTE MANAGEMENT

By diluting hydrochloric acid to 5 percent volume to volume or less, and neutralizing it with $NaHCO_3$ to a neutral pH, the solution is no longer corrosive and can, in some jurisdictions, be disposed down the drain. You must partially neutralize the solution of NaOH before be disposed down the drain.

PHOTOS

Fig. 1 The color of the solutions using methyl orange.

Fig. 2 The color of the solutions using phenolphthalein.

Fig. 3 The color of the solutions using blue bromothymol.

Fig. 4 The color of the litmus paper.

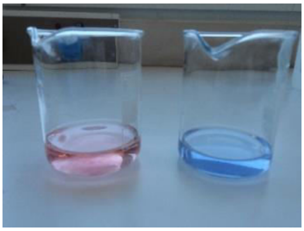

Fig. 5 The color of the solutions using litmus solution.

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ACETIC ACID TITRATION BY SODIUM HYDROXIDE

OVERVIEW

An important method of determining the amount of a particular substance is based on the measurement of the volume of the reagent solution. Titration is the process for determining the amount of a substance A, when you add a carefully measured volume of a known concentration solution of substance B, until the reaction of A and B has been completed [1]. In an acid-base titration, base solution is added to the acidic solution until the acid has been exactly neutralized. This is known as the equivalence point of the reaction. If we are using indicators in an acid-base titration, we are making the assumption that the end point of the indicator (the point at which it suddenly changes color) is the same as the equivalence point. In order to look at this assumption further, we look at the way the pH of a reaction mixture changes during an acid-base titration.

Sodium hydroxide (NaOH), also known as lye and caustic soda, is an inorganic compound. It is a white solid and highly caustic metallic base and alkali of sodium which is available in pellets, flakes, granules, and as prepared solutions at different concentrations. NaOH forms an approximately 50% (by mass) saturated solution with water. It is soluble in water, ethanol, and methanol. This alkali is deliquescent and readily absorbs moisture and carbon dioxide in air. NaOH is used in many industries, mostly as a strong chemical base in the manufacture of pulp and paper, textiles, drinking water, soaps and detergents and as a drain cleaner [2].

Acetic acid is an organic acid. Pure CH_3COOH is often referred as glacial acetic acid because it sets at 17°C taking the form of ice. In contrast, dilutions of the acid are curable at lower temperatures than water. Pure CH_3COOH has been prepared first around 1700A.D. by distillation of vinegar. Since CH_3COOH became a common chemical substance in each laboratory. The older chemists were completely familiar with the solid acid; because of the poor heating in the laboratory pure CH_3COOH was frozen in the flask. While dilute solutions of acid (such as vinegar) are harmless the pure acid is corrosive and causes painful burns. Therefore its handling must be done with caution. Most industries produce acetic acid by CO_2, H_2 and the presence of specific catalysts. The CH_3COOH is an important industrial product mainly used for the preparation of acetates [1].

MATERIALS AND METHODS

From a commercial vinegar 50 mL are obtained and placed in a beaker and diluted with water up to hundred ml. Two drops of phenolphthalein are placed in the solution, which is a suitable indicator, because at pH below eight point three it remains colourless whereas above it, it turns fuchsia. In a burette twenty five mL of NaOH (1M) are added. Add small volumes of the titrant to the solution until the indicator changes colour in reaction to the titrant saturation threshold, reflecting arrival at the endpoint of the titration.

HAZARDS

Wear protective glasses and gloves.

NaOH

CH_3COOH

CHEMICAL REACTION EQUATION

$$\mathbf{CH_3COOH_{(aq)} + NaOH_{(aq)} \rightarrow CH_3COONa_{(aq)} + H_2O_{(l)}}$$

CONCLUSIONS

Depending on the endpoint desired, single drops or less than a single drop of the titrant can make the difference between a permanent and temporary change in the indicator. The endpoint of the reaction is reached, since the colour was fuchsia and the volume of the reactant consumed is measured and used to calculate the concentration of analyte by appropriate formulas.

23ml of NaoH were consumed. In the endpoint of the titration:
$n_{CH3COOH}=n_{NaOH}$

$n_{NaOH}=c_{NaOH}\cdot V_{NaOH}=1\cdot 0.023=0.023$ mol.

Consequently, $n_{CH3COOH}=0.023$ mol.

$c_{CH3COOH}=n_{CH3COOH}/V_{CH3COOH}=0.023/0.1=0.23$ M

So the vinegar is a solution of 0.23 M acetic acid.

WASTE MANAGEMENT

The aqueous solution of CH_3COONa can be poured down the drain.

PHOTOS

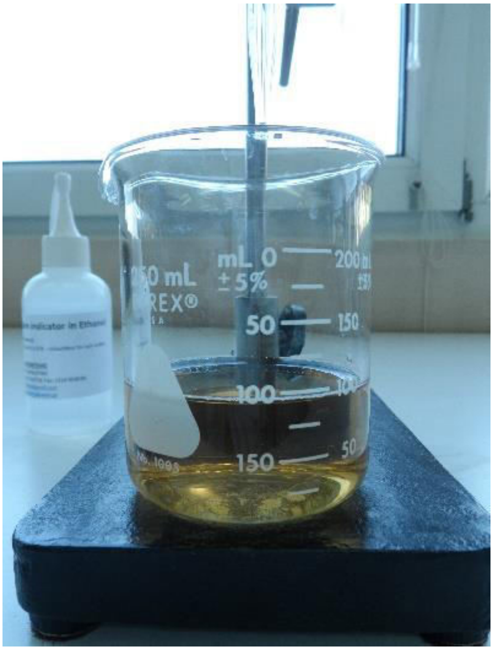

Fig. 1 The solution of vinegar with phenolphthalein.

Fig. 1 The end point of the titration. The color is not pink as expected but darker because of interference from the vinegar. This color appears abruptly during the titration without having as intermediary a pink nuance.

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SEPARATION OF A MIXTURE OF SAND AND SALT

OVERVIEW

In this experiment we study the separation of a mixture of two solid substances, sand and salt. Sodium chloride is a solid and crystalline salt which has a very high melting point. It is found in nature whereas is extracted by crystallization. Most salts are colorless. The color display occurs due to the anion or in the imperfect crystal lattice so-called holes. These holes result from the absence of a cation through the arrangement of the metal ions in the crystal lattice. The gap created is complemented by the electrons of neighboring atoms. This movement of electron cloud absorbs radiation in the visible region of the spectrum. So NaCl is displayed having a blue color before processing. NaCl as an ionic compound can be dissolved in water [1].

Sand is a naturally occurring granular material composed of finely divided rock and mineral particles. It is defined by size, being finer than gravel and coarser than silt. Sand can also refer to a textural class of soil or soil type. The composition of sand varies, depending on the local rock sources and conditions, but the most common constituent of sand in inland continental settings and non-tropical coastal settings is silica (silicon dioxide, or SiO_2), usually in the form of quartz[2]. Sand is not soluble in the water.

MATERIALS AND METHODS

The salt is soluble in water but not the sand. Initially, mix equal amounts of salt and sand. Then, add the mixture to a beaker with one hundred mL water. Then, we use a polyethylene or glass funnel with stem and filter paper. The paper is folded in half, then folded in quarters and the tip of the one corner is torn off to allow for a snug fit in the funnel cone. The funnel is supported with a ring and a clean beaker is placed beneath the funnel. Then, the filter paper is wetted completely with the solvent. This step adheres the filter paper to the funnel walls preventing solid from escaping. Then, the supernatant liquid is poured through the filter first. This allows the initial part of the filtration to proceed faster and may prevent clogging of the filter by the solid. The sand must be scraped onto the filter with a spatula. The paper filter is still the sand, while in the beaker passes the filtrate, the mixture that is water and salt. The filter paper has pores which allows the particles of the fluid to penetrate but not the particles of solid material because they are larger.

The filter paper with the sand is left to dry on a watch glass. Then, the filtrate is subjected to evaporation. Place the beaker on a hot plate (possibly electric stove or plate) for a few minutes until the water vapours.

HAZARDS

No hazards posed.

CONCLUSIONS

We note that in a few minutes the mixture boils and the water vapour begins escaping from the beaker. Upon completion of boiling, the entire amount of water has escaped in the form of water vapour, while the salt remains at the bottom of the beaker. So, the mixture of water-salt is separated through evaporation. The combination of the two methods helped our separation mixture to separate the mixture of sand-salt.

WASTE MANAGEMENT

The waste is environmentally friendly. The salt, the filter paper with the sand can be poured down the drain.

PHOTOS

Fig. 1 The sand that remained on the filter paper.

Fig. 2 The salt that remained on the bottom of the beaker.

Fig. 3 The filtrate.

REFERENCES

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BENZOIC ACID PURIFICATION BY SUBLIMATION

OVERVIEW

Separation methods are generally; the purification of a substance of its dirtiness or obtaining each one of the substances in a mixture. Sublimation is a separation method based on the formation of a new phase through heating or cooling. It is the transition of a substance into gas before being liquid out of solid state. The process seen below is called sublimating and the separated substance is called sublimate.

HEAT

SOLID → GAS → SOLID

This method is used in the analysis of organic and inorganic substances and is applied to substances whose vapour pressure of solid state is extremely high. When the vapour of these substances is cooled down, they again become solid, before being in a liquid state. Thus, a substance that has this property in a mixture, using the sublimation process can easily separate them from other substances which do not have this property.

When the mixture is heated and a cold surface is held over this mixture, the sublimated substance will again go into solid state here, thus separating from the mixture in pure state. Important points about the procedure of this method are given below.

- During the application of this method, the substance may be melted then boiled. However, if cooled directly it will change into solid state from gas state.
- Due to the conditions of solubility, if it is not possible to purify by crystallizing of the solid substances using a solvent, then sublimation must be preferred.
- This method is used in the purification of camphor, benzoic acid, naphthalene, and andesite.

MATERIALS AND METHODS

Mix the carbon dust and solid benzoic acid in a porcelain capsule. Cover it with filter paper. Put the funnel whose neck is closed within the glass cotton on the filter paper. Heat it under melting point (122^{o}C). Record the results. The steam rising through the holes of the filter paper will crystalize on the cold surface of the funnel.

HAZARDS

Benzoic Acid ($C_7H_6O_2$)

CONCLUSIONS

When thermal process is applied to the mixture of carbon powder and benzoic acid mixture, not exceeding 122 ℃, benzoic acid passes into gas phase from solid phase and crystallization in the inner walls of the glass funnel covered over the mixture is observed. While thermal heat is applied, make sure to not inhale it.

WASTE MANAGEMENT

Benzoic acid can easily decompose in nature. It can decompose slower in cold water, but easily in hot water. The separated benzoic acid crystals dissolving in water can be discharged to soil.

PHOTOS

Fig 1. Benzoic acid and carbon dust are mixed in a porcelain capsule, and sealed with filter paper.

Fig. 2 A glass funnel is placed on filter paper and benzoic acid is heated but, not until melting point.

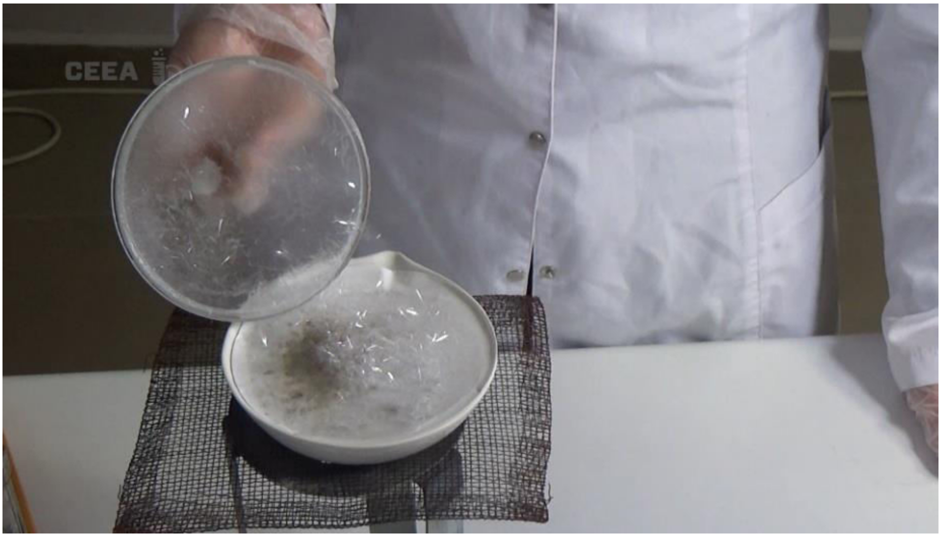

Fig 3. Benzoic acid which is in steam phase hits the surface of the cold funnel and crystallizes.

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DETERMINATION OF THE OXYGEN RATE IN AIR, USING A BURNING CANDLE

OVERVIEW

In a burning candle, stearic acid reacts with oxygen from the environmental air. In case of complete combustion the reaction products are water vapour and carbon dioxide only. However, the yellow flame already indicates that combustion is not complete. Unreacted carbon (soot) and carbon monoxide are produced as well. Assuming that (i) the amount of CO produced is small compared to the amount of oxygen used, and (ii) all CO_2 produced dissolves in water a reliable estimate of the oxygen rate in air can be obtained by allowing a candle to burn in a closed volume until all the available oxygen is exhausted. Measuring the volume decrease should prove that approximately $1/5^{th}$ of the air, that surrounds the candle, is oxygen.

Our measurement shows an oxygen rate of 20%, which is a little less than the theoretical value of 21%.

MATERIALS AND METHODS

Requirements: Glass cylinder, felt-tip pen, Petri dish, coloured water, thin candle, coins.

Procedure: Apply, with the felt-tip pen, a volume scale on the glass cylinder from 0 to 100%. Fasten the candle, using some candle wax, in the middle of the Petri dish. Put four coins around the base of the candle as a support for the glass cylinder, that will be placed, later, upside down over the burning candle. Pour the coloured water in the Petri dish.

Light the candle, and place the cylinder with a quick movement upside down over the burning candle. The candle is burning now in a closed room with a limited amount of oxygen. During the combustion of the stearic acid, oxygen is used and water vapour and carbon dioxide are produced. The air volume of the closed room decreases, while the oxygen is used, probably most of the water vapour condenses and most of the carbon dioxide dissolves in the water. The volume decrease leads to the rising of the water level within the cylinder, until there is no oxygen left. De candle stops burning then. After a few minutes, when the air in the cylinder is back to room temperature, the water level shows a remaining air volume of 80% in the cylinder which means that the oxygen rate was 20%.

NO HAZARDS

CHEMICAL REACTION EQUATIONS

Complete combustion:

$C_{17}H_{35}COOH(s) + 26O_2(g) \rightarrow 18CO_2(g) + 18H_2O(g)$

Incomplete combustion:

$C_{17}H_{35}COOH(s) + ?\ O_2(g) \rightarrow ?\ C(s) + ?\ CO(g) + ?\ CO_2(g) + ?H_2O(g)$

CONCLUSION

Assuming that most water vapour will condense, the amount of CO is negligible and most carbon dioxide will dissolve in the water, the decrease of the air volume will be equal to the volume of the burnt oxygen. In order to obtain a reliable result, the steps have to be executed with care:

1. Immediately, after the cylinder is placed over the candle, the air in the closed room expands. Be aware that the expanding air is not bubbling out of the cylinder, otherwise the measured volume decrease is too high.
2. If placing the cylinder over the candle takes place too slowly, some oxygen in the cylinder will already be consumed before the water closes the combustion room. In that case a lower volume decrease is measured.
3. The oxygen percentage can be calculated using the formula:

$$Rate\ of\ oxygen = \left(\frac{V_{cylinder\ start} - V_{cylinder\ end}}{V_{cylinder\ start} - V_{candle}}\right) \times 100\%$$

   A thin candle makes it therefore easier to derive the oxygen rate from the final water level.

Taking into account these inaccuracies, our result of 20% is acceptable.

WASTE MANAGEMENT

Remaining stearic acid can be reused for new candles or can be treated as normal household waste.

PHOTOS

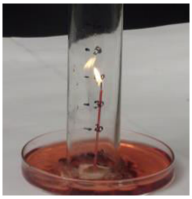

Fig. 1 At the start.

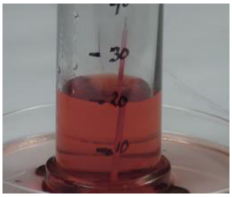

Fig. 2 20% volume decrease.

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OXYGEN CONTENT DETERMINATION, USING GAS SYRINGES AND COPPER WIRE

OVERVIEW

A simple method to determine the rate of oxygen in the air is to use two 100 cm^3 glass gas syringes connected to both sides of a heat resistant silica tubing, containing excess copper wire.

One syringe is empty and the other filled with exactly 100 cm^3 of air. When the copper wire is heated and the air is repeatedly moved from one syringe to the other along the copper wire, the oxygen in the air reacts with the copper to form copper(II) oxide. This results in a decrease of the air volume.

MATERIALS AND METHODS

Requirements:

Two 100 cm^3 glass gas syringes connected to both sides of a heat resistant silica tubing (ordinary glass tubing melts!), copper wire, stands and clamps to hold the gas syringes, Bunsen burner.

Procedure:

The silica tube with copper is strongly heated with a hot flame Bunsen burner and the gas syringes 1 and 2 are moved (approximately 50 times) to and fro to pass the air over the hot copper so ALL the air comes into contact with the hot copper.

Copper oxide is a black solid of little volume (so little error) and none of the other gases in the air reacts with copper, so reduction of the gas volume is solely due to removal of oxygen. Neglecting the volume of the silica tube would lead to an unacceptable systematic error so it has to be taken into account. Another possible source of error is reading the final gas volume at elevated temperature. Make sure the initial and final readings are taken at the same temperature. The gas volume reading reaches a minimum value when all the oxygen in the air has reacted with the copper and when thermal equilibrium with the environment is re-established.

HAZARDS

CuO

CHEMICAL REACTION EQUATION

$2\ Cu_{(s)} + O_{2(g)} \rightarrow 2\ CuO_{(s)}$

CONCLUSIONS

In our set up, the silica tube had a volume of 15 cm^3. The final volume reading appeared to be 76 cm^3. Thus, the measured oxygen rate is:

$$Oxygen\ rate = \frac{((100+15)-(76+15))}{(100+15)} \times 100\ \% = 21\ \%$$

WASTE MANAGEMENT

CuO can be treated as household waste.

PHOTOS

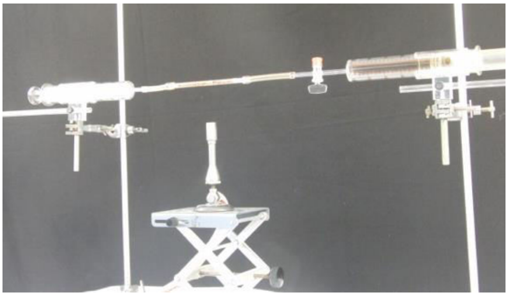

Fig. 1 The set up.

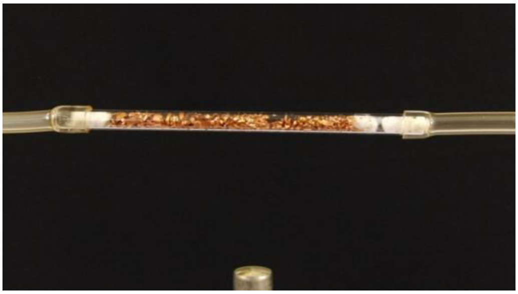

Fig. 2 Silica tube with copper wire

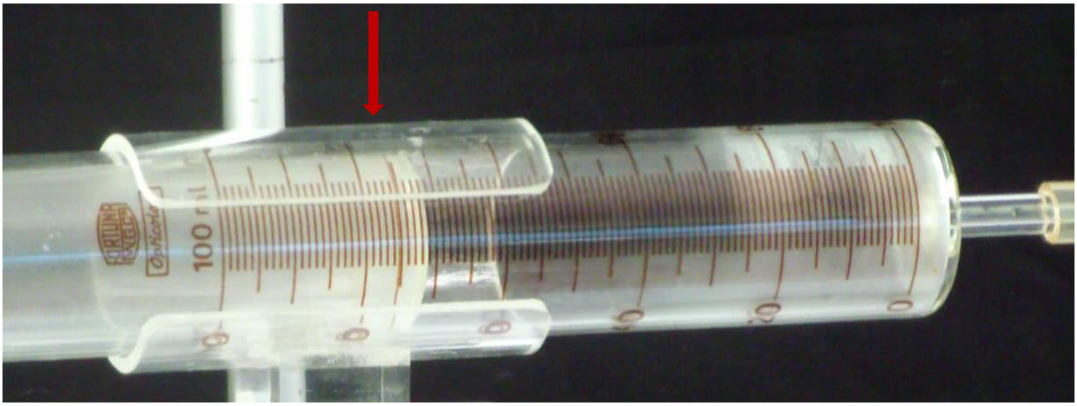

Fig. 3 Final volume reading

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PRODUCTION OF PHOSPHORIC ACID FROM RED PHOSPHORUS

OVERVIEW

This demonstration is suitable for use in lessons on periodicity. Lower secondary school students must be made aware of the fact that the oxides of non-metals form acids, and that these play a role in environmental acid deposition. Older students are often required to know the properties of the period 3 elements and the reactions of their oxides with water.

Many think that the experiment can only be performed with white phosphorus, which is hard to acquire and difficult to handle safely. But this isn't the case – exactly the same astonishing effect can be produced with red phosphorus.

MATERIALS AND METHODS

Equipment:

A clean deflagrating spoon; 1 L borosilicate round bottomed flask stopped with a rubber or cork bung. The deflagrating spoon needs to be inserted through the bung. The diameter of the the spoon handling rod is too small for a standard 2-hole bung, so a special bung needs to be bored; Means to secure the flask (clamp and retort stand or cork ring); Delivery tube attached to the oxygen cylinder; Safety screen if in open lab; Glass rod; Bunsen burner.

Reagents:

Red phosphorus (about 0.4 g); An oxygen cylinder; Methyl orange indicator solution.

Procedure:

Pour 50-100 ml water and several drops of methyl orange indicator solution in the round bottomed flask. Fill the round bottomed flask with oxygen and close it with the rubber stopper. Clamp the flask into position.

Place the phosphorus in the bowl of the deflagrating spoon (about 0.4 g) and lower it into the round bottomed flask. Hold the end of a glass rod in a roaring Bunsen flame for about 30 seconds and then bring the hot end of the rod in contact with the phosphorus, which will lit.

Hypnotic white tendrils of smoke fill the flask, which glows like a miniature sun. Once the reaction is complete, ensure the flask has sufficiently cooled down before swirling the water in de flask to help dissolve the P_4O_{10} mist. The solution will turn red as phosphoric acid is formed.

HAZARDS

P_4

P_4O_{10}

H_3PO_4

Wear eye protection.

Do not attempt to dry red phosphorus in an oven if it looks damp. You may be able to pat it dry enough with a paper towel.

CHEMICAL REACTION EQUATIONS

The reaction we see is:

$$P_4(s) + 5O_2(g) \rightarrow P_4O_{10}(s)$$

Followed by reaction with water to produce phosphoric acid:

$$P_4O_{10}(s) + 6H_2O(l) \rightarrow 4H_3PO_4(l)$$

WASTE MANAGEMENT

The acidic solution produced is dilute enough to be poured down the sink. Any excess red phosphorous on the spoon will need to be burned off in a fume cupboard.

PHOTOS

Fig. 1 At the start.

Fig. 2 Here comes the sun…

REFERENCES

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WATER HARDNESS IDENTIFICATION BY EVAPORATION AND SOAP PRECIPITATION

OVERVIEW

If water contains important amounts of ions which will give precipitation, it is called hard water. Hard water contains calcium, magnesium and heavy metal ions. They form precipitation with soap. Hardness today is defined as the concentration of calcium carbonate equivalent to total concentration of multi-loaded cations in the sample.

Temporary water hardness: Bicarbonate ion contains $HCO_{3.}$ If $HCO_{3\ (aq)}$ is heated, it easily dissolves giving out CO_3^{2-}, CO_2 and water. CO_3, reacting with the cations in water, forms a mixture of precipitation of $CaCO_3$ – $MgCO_3$ and a sediment called lime sludge. The formation of lime sludge may cause serious problems in industrial boilers producing steam and in power plants which operate with steam.

Temporary hard water can be softened by adding burnt limestone [Ca $(OH)_2$] into the water in the refinery and then filtering the metal carbonate deposit. Base, reacting with bicarbonate ion forms water and carbonate ion. Carbonate ion, as Ca^{2+} reacting with Mg^{2+} ions deposits as metal carbonates.

$$HCO_3^- + OH^- \longrightarrow H_2O + CO_3^{2-}$$

$$CO_3^{2-} + Mg^{2+} \longrightarrow MgCO_3$$

Permanent water hardness: HCO_3^- just like SO_4^{2-} contains other cations too. In order to soften the permanent hard water add Na_2CO_3 (sodium carbonate). Cations, such as Ca^{2+} and Mg^{2+} are deposited as carbonates. The remaining water is soft water containing Na^+ ion that contains Ca^{2+}+ Mg^{2+} ions form precipitation with soap and hinders foaming. The deposit seen on the bathtub is a mixture of calcium and magnesium soaps. The formation of this precipitation makes the foaming of soaps and shampoos difficult. The hardness of water caused by the bicarbonates can be removed through ventilation. Because of the released Co_2 there forms carbonates that do not dissolve.

The softening of hard water means the separation of minerals which cause hardness in water. It is important, because

- Less water is consumed when doing washing
- The minerals in hard water accumulate in pipes, causing their plugging and narrowing.

- In some industrial fields soft water is necessary, especially in canned foods, textile, industry, paper production, leather work, ice-making and starch production.

MATERIALS AND METHODS

In this experiment the aim is to determine the pH of water that is, the ability of water to consume soap. The materials required are; a hydrometer bottle hydrometer burette, fixed soap solution (70 mL), tap water, glass funnel, 100g sodium soap and 56% of alcohol.

- Prepare a fixed soap solution by melting the 100g sodium soap in 56% alcohol and fill the hydrometer burette until full.
- Put 40mL tap water into the hydrometer bottle for analysis.
- Slowly, drip soap solution from the hydrotimeter buretta into the sample.
- Close the lid of the bottle tightly and shake the bottle. Continue, until the foam on the liquid remains at least 5 minutes without disappearing.
- The French pH value is to be found by multiplying the used fixed soap solution with factor.

Note: The foam on the liquid must be 1 cm in thickness.

HAZARDS

Ethyl Alcohol (C_2H_6O)

CONCLUSIONS

When ions such as Ca, Mg, Na and K compounds react with soap, as a result of the changing of Na and K ions with Ca and Mg, the ions causing the pH in water are precipitated and thus form foam.

WASTE MANAGEMENT

The substances can be discharged down a sink.

PHOTOS

Fig. 1 Slowly dripping soap solution from the hydrometer burette into the sample.

Fig. 2 Closing the lid of the bottle tightly and shaking the bottle.

Fig. 3 Shake until the foam remains on top of the surface

REFERENCES

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WATER HARDNESS DETERMINATION VIA TITRATION REACTIONS

OVERVIEW

Water hardness is an important factor to take in account when consuming water or when using the water for washing purposes. The **water hardness** refers primarily to the presence of calcium and magnesium salts dissolved in water. There are two types of hardness:

Temporay hardness is caused by the presence of calcium and magnesium bicarbonates {$Ca(HCO_3)_2$ & $Mg(HCO_3)_2$} which are transformed by boiling in insoluble carbonates precipitating out of the solution and leaving the water softer upon cooling.

Permanent hardness is given by the content of calcium and magnesium chlorides and sulfates which do not precipitate out as the temperature increases. [1]

Water hardness is expressed in terms of calcium carbonate ($CaCO_3$) or calcium oxide (CaO) dissolved in a unit of pure water. Various units of hardness are accepted but all refer to the content of different Ca and Mg compounds expressed as:

- - *American degrees* $CaCO_3$ mg/L (or ppm);
- - *German degrees* (1 °dH = 10 mg/L CaO or 17.848 ppm $CaCO_3$);
- - *English degrees* (°e or °Clark = 64.8 mg $CaCO_3$/imperial gallon (4.55 L) or 14.254 ppm)
- - *French degrees* (°fH or °f = 10 mg $CaCO_3$/L; 10 ppm) [1]

Water hardness is important to be known for every household as detergent dosage for washing is adapted to the hardness of water used. A harder water needs more detergent and often detergent producing companies mark the recommended dosage on a chart printed on the package.

MATERIALS AND METHODS

Water hardness is measured following a titration procedure based on the color exchange of a complexometric indicator: the Calmagite.[2] For this a total hardness test kit is used [3] based on a procedure of complexing the Ca and Mg ions with Calmagite at a pH=10 used in conjunction with EDTA. There are several, commercially available, chemical test kits. Most of them are based on the same principle and the used solutions are calibrated to have specific concentrations in such a way that the user can easily determine the water hardness based on the volume of titrating solution used (ml or drop count).

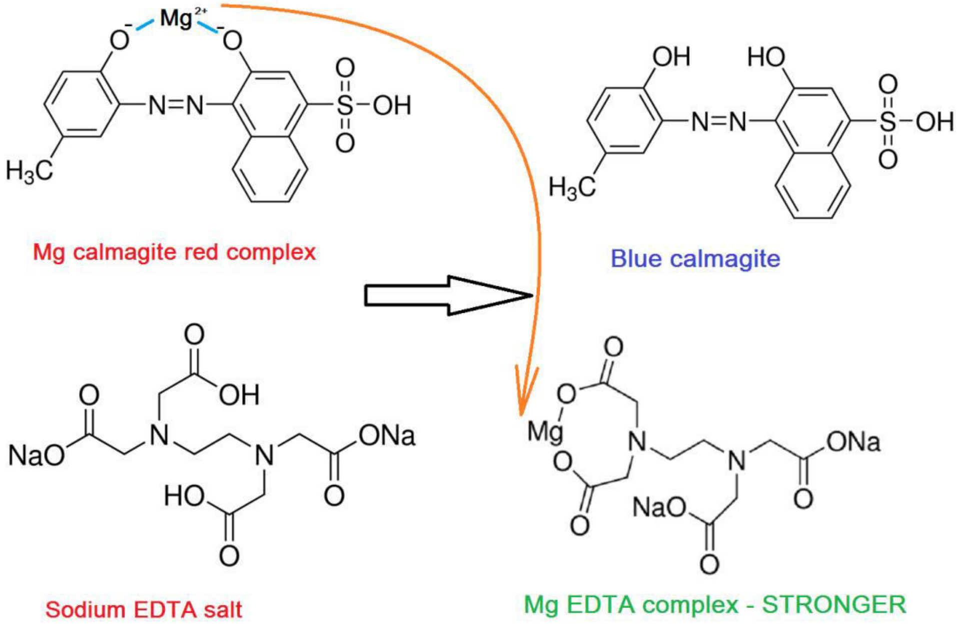
The example described below makes use of one of these commercially available sets. A water sample of 5 mL is treated with 5 drops of Hardness Buffer solution to reach the necessary pH (pH=10). One drop of calmagite indicator solution is added. Observe the red color appearing in the solution. Taking a one mL syringe add drop wise a titrating solution of EDTA until the solution becomes purple. This is the end point of the titration. To measure the water hardness in ppm (mg $CaCO_3$/L) multiply the volume of the added titrating solution with 300.

For a low range result (<30 mg/L) repeat the operation using a larger volume of water to be analyzed. The result will be calculated accordingly (for instance if 25 mL of water is analyzed the volume of titrating (in mL) is multiplied by 60 and if 50 mL of water is measured the multiplication factor will be 30). As the smallest titrating solution increment is a drop of solution which often doesn't allow to stop the titration at the equivalence point a larger volume of water sample is recommended: 25 or 50 mL.

HAZARDS

The reagents used do not represent a potential hazard for the users nor for the environment.

CHEMICAL REACTION EQUATION



CONCLUSIONS

Initially the solution is red due to a Mg (and Ca) calmagite complex. Titration with EDTA (which is a stronger binding agent than calmagite) will coordinate all free Ca^{2+} and Mg^{2+} ions from the solution eventually forcing out the Ca^{2+} and Mg^{2+} ions coordinated in the red calmagite complex. If the titration is done at a pH=10 the end point and the equivalence point are nearly the same allowing the user to stop the titration right when all the Mg^{2+} and Ca^{2+} ions are complexed by the titrant (EDTA) solution. Once the calmagite is set free from the Mg^{2+} and Ca^{2+} ions it changes color from red to blue. A correct equivalence point is obtained if the titration is stopped when the solution turns purple from red. Measuring the volume of the titrating solution and multiplying with the necessary factor (according to the volume of the water sample) we obtain the result directly in ppm (mg $CaCO_3$/L). We can use different hardness charts to transform the result from ppm to the most common German or French hardness degrees.

WASTE MANAGEMENT

The final solutions can be discarded into the drain as there is no harmful effect on the environment for the used chemicals.

PHOTOS

Fig. 1 At the beginning of the experiment the indicator Calmagite is red due to its complexation with Ca^{2+} and Mg^{2+} ions.

Fig. 2 At the end of the titration the indicator Calmagite is blue and non complexed. All Ca^{2+} and Mg^{2+} ions are bind to the EDTA from the titration solution.

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THE INCREASING OF SOAP FOAM FROM OXYGEN

OVERVIEW

The aim of this experiment is to observe the soap foaming properly of oxygen which is released with the breaking-up of hydrogen peroxide.

Oxygen gas is released with the breaking up of peroxides that play an important role in global warming.

Hydrogen peroxide is a dense, odourless colourless and chemically not stationary fluid. It easily decomposes to release water and oxygen.

The areas that hydrogen peroxide is used:

1. Its 3% solution is used as germicide and it is called in public as water with oxygen.

2. Its 3-6% solution is used to turn hair into yellow.

3. Its 30% solution is called perhydrol and is used in chemistry labs.

4. Its 30-50% concentration is used in textile industry to whiten the cotton fabric. Without destroying the quality of the fibre, it whitens the colour defect in a fabric. The whiteness obtained is quite permanent. The same solution is used to whiten paper bulk, wooden surfaces, cotton, silk, oil and some foods.

5. Its 70% solution is used to oxidize excessively the organic matters and inorganic ions. Concentrated hydrogen peroxide is harmful for the skin.

6. Hydrogen peroxide has a reducing and oxidizing property depending on the medium. It destroys and kills the microbes because it is oxidizing. As a result, it can be found in several cleaning materials.

Hydrogen peroxides are commonly known as bleaching agents. Bleaching agents are divided into two:

1. Bleaching agents with chlorine (hypochlorite’s)

a) Calcium Hypochlorite

b) Sodium Hypochlorite

c) Potassium Hypochlorite

2. Bleaching agents with oxygen

a) Hydrogen Peroxide

b) Sodium Perborite

Most of the above bleaching chemical substances, create a greenhouse effect due to to their spreading gas to the air, and thus cause global warming. With the breaking up of hydrogen peroxide one of such bleaching agents, oxygen gas is formed. This contributes to the formation of ozone layer. The layer called ozone layer is formed with the combination of oxygen gas and oxygen atom. As a result of photosynthesis, the oxygen gas accumulates in the stratosphere layer of atmosphere. The ultraviolet rays divide the oxygen molecules of this layer into two oxygen atoms as (oto). These oxygen atoms form the ozone by combining oxygen molecules.

Ozone is found in low ration in the atmosphere compared to 78% nitrogen 21% oxygen and carbon dioxide, etc. Ozone plays an important role in the preservation of living creatures in the world and it affects the climate. Ozone especially together with oxygen absorb most of the ultraviolet rays in the stratosphere layer coming from the sun, thus preventing these rays reaching to the earth and destroys their burning power.

Hydrogen peroxides release oxygen gas when broken up in laboratory conditions. In the experiment of "The increasing of soap foam from oxygen" oxygen gas is obtained with the help of potassium iodine. The oxygen gas which we have obtained here leads to the formation of the ozone layer.

MATERIALS AND METHODS

The materials used are; beaker, spatula, dropper, 30% liquid hand soap

1. Put some liquid soap into a beaker.

2. Later add some solid potassium iodine using a spatula.

3. Shake the beaker slowly.

4. Using a dropper, add hydrogen peroxide of 30% concentration into this mixture.

5. The oxygen gas released with the addition of hydrogen peroxide forms the soap foams.

HAZARDS

Hydrogen Peroxide (H_2O_2)

CHEMICAL REACTION EQUATION

$I^- + H_2O_2 \rightarrow IO^- + H_2O$

$IO^- + H_2O_2 \rightarrow I^- + H_2O + O_2$

CONCLUSIONS

In this experiment oxygen gas has been obtained out of hydrogen peroxide with the help of potassium iodine.

WASTE MANAGEMENT

Hydrogen Peroxide cannot be discarded down a drain it must be emptied in a sealed tight waste bottle.

PHOTOS

Fig. 1 Adding solid potassium iodine using a spatula into the soap

Fig. 2 Using a dropper, add hydrogen peroxide of 30% concentration into this mixture.

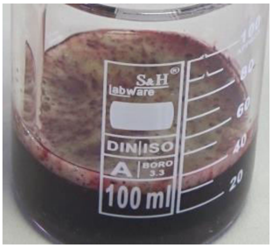

Fig. 3 The oxygen gas released with the addition of hydrogen peroxide forms the soap foams

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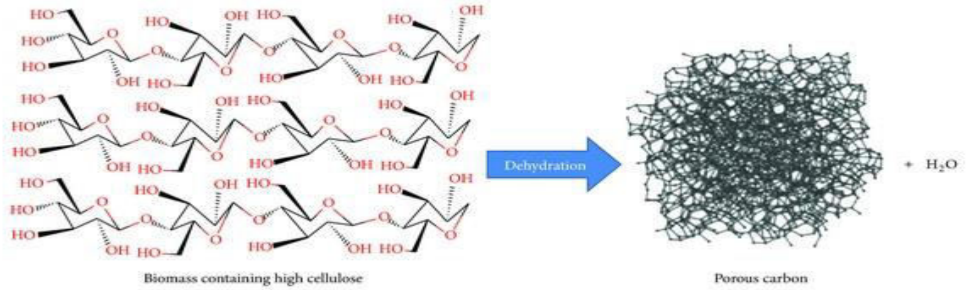
DETERMINATION OF THE MASS PERCENTAGE OF CARBON IN SUGAR

OVERVIEW

At secondary educational level determination of the mass percentage of carbon as for example in sugar is generally demonstrated with the simple method of carbonization or pyrolysis. The outcome should be regarded as a rough estimate because an unknown quantity of carbon escapes in the form of gasses like CO and CH_4. By using the dehydration method instead, more accurate results are likely to be obtained. We were able to achieve surprisingly high accuracies in the carbon content determination.

The dehydration of cellulose into porous carbon is illustrated in the graph below [1]:



Theoretically, 5 moles of water per glucose unit can be extracted by concentrated sulphuric acid to form porous carbon. Besides a more accurate determination of the carbon content, this process provides a green route to synthesis of high content of carbon from glucose derived biomass without any greenhouse gas emissions which only produces water as a by-product.

MATERIALS AND METHODS

Requirements:

250 mL Erlenmeyer flasks, a 50 mL and a 10 mL measurement cylinder, a water bath, a thermometer, a funnel, red litmus paper, filtration paper, a fume hood, a stove, goggles, gloves, aprons.

Reagents:

97% concentrated sulphuric acid, sugar, ethanol, demineralised water, 25% ammonia solution, a soluble aluminium salt.

PROCEDURE

Measure 10 mL of concentrated sulphuric acid with a measure cylinder and pour the content into a 250 mL Erlenmeyer flask. Drop 1.00 gram of sugar into this flask. Heat the Erlenmeyer in a water bath. Make sure, by using a thermometer, that the temperature of the sulphuric acid solution does not exceed 50 ^{o}C. Temperatures above 50 ^{o}C will lead to a competitive oxidation reaction of the sugar, which can be observed by CO_2 gas formation, and will lead to erroneous measurement results.

After about 30 minutes of reaction time the content of the Erlenmeyer will change from first a yellow colour (caramelization) into a deep black colour.

Now, slowly and carefully pour ca. 3 mL of water into the Erlenmeyer. The sulphuric acid will heat up (it is a very exothermic reaction). Cool if necessary. Add 25 mL of demineralised water to which 0.05-0.1 gram of a soluble aluminium salt was added. Neutralize the reaction mixture by carefully adding ca. 25-30 mL 25% ammonia solution (exothermic reaction, cool if necessary) until the pH is alkaline (check with red litmus paper). In the resulting alkaline solution, the aluminium salt will start to act as a flocculent. Leave the Erlenmeyer at rest and observe how the flocculent will coagulate the carbon material (a suspension of predominantly graphene!) into a filterable substance.

Take a filter paper with a diameter of ca. 7 cm. Filter, carefully, the reaction solution over the filter paper of which the mass has been determined with a scale. Wash the black carbon residue on top of the filter paper with 2 portions of ca 10 mL water and make a final wash with also 2 portions of 10 mL pure ethanol. Dry the filter in a stove at 80^{o}C until a constant mass is reached.

Calculate the carbon content as follows: $\frac{\text{residue mass}}{\text{1.00 gram sugar}} \times 100\%$

HAZARDS

$C_{12}H_{22}O_{11}$	**none**	**NH_4OH**	
H_2SO_4		**$AlCl_3$**	
C_2H_5OH			

CHEMICAL REACTION EQUATION

$$C_{6n}(H_2O)_{5n}(s) \rightarrow 6n\ C(s) + 5n\ H_2O(l)$$

WASTE DISPOSAL

Dispose of all liquid residues into a sink and wash away with copious amounts of tap water.

PHOTOS

Fig. 1 Sugar and H_2SO_4

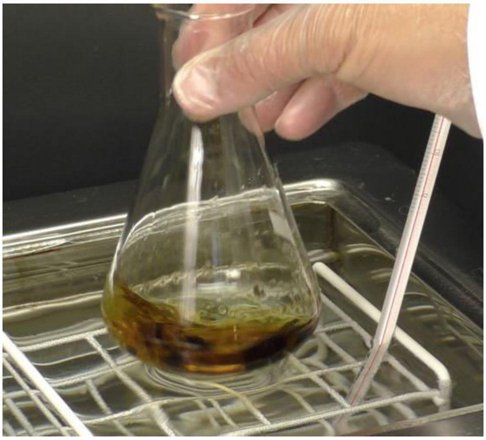


Fig. 2 The colour changes from orange to black

Fig. 3 Filtration, using the water jet pump

Fig. 4 Result after drying: carbon

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DETERMINATION OF THE MASS-PERCENTAGE OF CARBON IN SUGAR

OVERVIEW

Sugar is a carbohydrate; in this test we attempt to obtain an impression of the carbon content in sugar by pyrolysis. Sugar is heated in a test tube over a Bunsen burner flame and thermal agitation causes the molecules to decompose. Ideally, decomposition into carbon and water should occur (reaction 1). Then, the amount of carbon can be determined simply by weighing the remaining solid. In practice however the reaction is more complicated. The product gas appears to be flammable, so light hydrocarbons and / or carbon monoxide are likely to be formed as well (reaction 2). Because the quantity of carbon escaping as gasses is unpredictable the outcome of the experiment has to be treated with care.

MATERIALS AND METHODS

Materials and reagents:

Bunsen burner; test tube; test tube holder; mass scale; sugar.

Procedure:

1. Determine the mass of the empty test tube and note this on paper.
2. Pour about 1 gram of sugar into the test tube and determine the mass of the sugar plus the test tube. Subtract the mass of the test tube to get the mass of sugar and note on paper.
3. Attach test tube holder to the top of the test tube.
4. Heat the test tube over the hottest part of the flame. Point it away from you or anyone else.
5. Lead the product gases through the flame of a second burner in order to investigate flammability;
6. Remove the test tube from the flame before the substance can rise as high as the top of the tube.
7. Allow the test tube to cool in a rack as soon as the formation of gas bubbles has stopped.
8. Find the mass of the test tube with the newly formed carbon substance and calculate the result as the mass percentage of carbon.

CHEMICAL REACTION EQUATIONS

$$\mathbf{C_{12}H_{22}O_{11}(s) \rightarrow 12\ C(s) + 11\ H_2O(g)} \qquad \mathbf{(1)}$$

$$\mathbf{C_{12}H_{22}O_{11}(s) \rightarrow C(s) + \text{mixture of gasses}} \qquad \mathbf{(2)}$$

HAZARDS

$C_{12}H_{22}O_{11}$ **no hazards**

PHOTOS

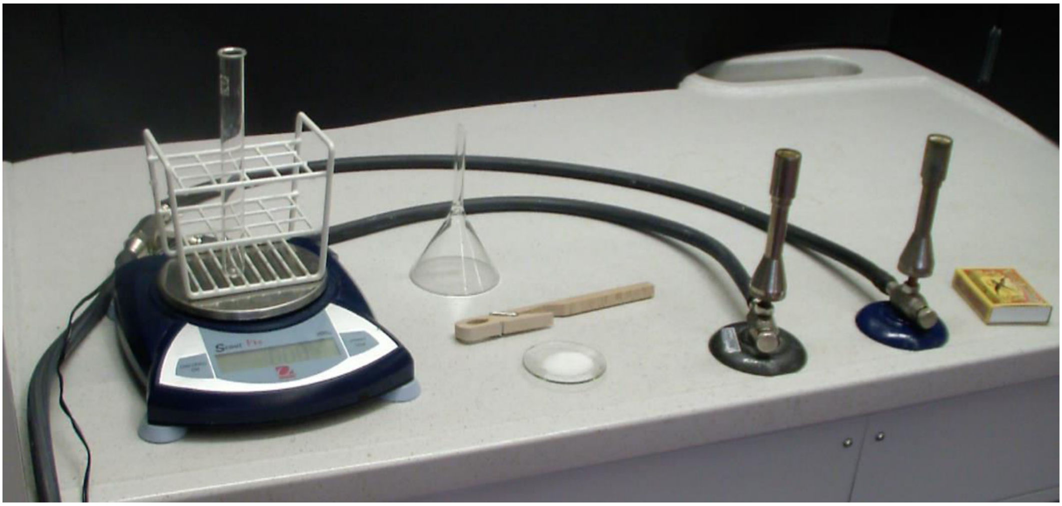

Fig. 1 Requirements.

Fig. 2,3,4 From white to black.

REFERENCES

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THE EXPERIMENT OF BURNING SUGAR

OVERVIEW

Let's look at the process of burning and exactly what it is. Burning is the oxidation reaction of a substance by oxygen. Oxygen takes away electrons from a substance, breaks the bonds between atoms in the substance and inserts itself into the product molecules, which results in a mixture of oxides. The difference between burning and other oxidation reactions (e.g. iron corrosion) is that large amounts of heat are released during burning. There are two conditions needed for a substance to burn:

9. During oxidation the amount of heat released must be sufficient to sustain burning
10. The oxidation reaction must be quite fast

MATERIALS AND METHODS

For this experiment the aim is to burn sugar and produce graphite. In this experiment besides sugar (C12H22O11), potassium chlorate (KClO3) and sulphuric acid (H2SO4) are used. This experiment can be made without using (KClO3), but by speeding up the reaction in order to obtain a much more successful view potassium chlorate (KClO3) is used.

11. First we put the sugar ($C_{12}H_{22}O_{11}$) and potassium chlorate ($KClO_3$) into a cup and break them up into small pieces. The smaller the pieces are, the quicker the reaction will be, so they should be crushed as much as possible.
12. Add sulphuric acid to the mixture obtained and observe that sulphuric acid, without using any flame, will cause the sugar to burn and become coal-like, thus obtaining what is called graphite which is used in the tips of pencils.

HAZARDS

$HClO_3$ (potassium chlorate)

H_2SO_4 (sulphuric acid)

CHEMICAL REACTION EQUATION

This is an exothermic reaction.

$$C_{12}H_{22}O_{11(s)} \longrightarrow 12\,C_{(graphite)} + 11\,H_2O$$

PHOTOS

Fig. 1 The sugar ($C_{12}H_{22}O_{11}$) is added.

Fig. 2 Then the potassium chlorate ($KClO_3$) is added.

Fig. 3 Finally the sulphuric acid (H_2SO_4) is added.

Fig. 4 In just seconds sugar burns, and as a result graphite is observed.

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SILVER IDENTIFICATION WITH POTASSIUM CHROMATE

OVERVIEW

The Silver (Ag) is known from ancient times due to its ability to be present in nature in native state (as a metal) being used mostly in jewelry and mirror production. Later on silver compounds have found different uses both in pharmacy and in the laboratory. Silver ions Ag^+, belonging to the 1st group of cations, have a wide usage in the chemistry laboratory either for the recognition of halogenure ions (F^-, Cl^-, Br^-, I^-) or as coordination compounds (Tollens' reagent) capable to perform different redox reactions.

A specific characteristic of silver ions Ag^+ is the recognition reaction with chromate ions leading to the formation of a specific, bright red precipitate Ag_2CrO_4. [1] A very useful application of the precipitation capability of silver salts, related to the previous description, is the titration of chloride Cl^- ions containing solutions in the presence of chromate CrO_4^{2-} with standard solutions of Ag^+. Useful discussions can be performed in the classroom around this subject, the theme being applicable either when studying precipitation reactions and the solubility product concept or when introducing cation identification methods.

MATERIALS AND METHODS

To present the ability of soluble silver salts ($AgNO_3$) to react with potassium chromate (K_2CrO_4) a 0.1M solution of $AgNO_3$ and respectively a 0.1M solution of K_2CrO_4 are needed:

Solution 1. For this weigh 0.945 g of solid $AgNO_3$, transfer it in a 50 mL volumetric flask followed by solubilization in 35-40 mL of distilled water and finally add the rest of the water to the mark. CAUTION is needed to use distilled or bidistilled water, for the tap water usage will lead to a cloudy solution due to the precipitation of the silver cation with the chloride anions present in the water.

Solution 2. Follow the same procedure to prepare 50 mL of 0.1M solution of K_2CrO_4 by weighing 0.970 g of solid salt and dissolving it in another 50 mL volumetric flask.

The experiment is to be performed in a test tube. Take 2-4 mL of 0.1M solution of K_2CrO_4 and add drop wise, using a Pasteur pipette, a few drops of 0.1M solution of $AgNO_3$. Observe the precipitate formed, its color and structure. Reverse the order of addition of both substances. Observe the similarities and differences.

HAZARDS

Attention is to be paid when handling silver salts as they are very harmful for the environment. Chromate ions containing solutions are also dangerous for the environment and as any Cr(VI) containing salts they are highly toxic, irritants and carcinogenic.[2]

$AgNO_3$

K_2CrO_4

CHEMICAL REACTION EQUATION

$$2\ AgNO_{3(aq)} + K_2CrO_{4(aq)} \text{ ---> } Ag_2CrO_{4(s)} + 2\ KNO_{3(aq)}$$

CONCLUSIONS

When mixing the silver nitrate $AgNO_3$ with potassium chromate K_2CrO_4 a double replacement reaction takes place (see the chemical equation). The silver chromate formed, Ag_2CrO_4, is bright red and insoluble in water whereas the potassium nitrate is colorless and water soluble. Ag_2CrO_4 has a solubility product of $K_{SP}=1.2x10^{-12}$ so it's normal to appear in a solution as a precipitate.

Because the chromate ion doesn't form such red precipitates with other cations, this color being specific for silver, this method is used to identify Ag^+ ions in a solution.

WASTE MANAGEMENT

As the solubility of AgCl is lower than of Ag_2CrO_4 a precipitation reaction is performed with diluted HCl giving solid AgCl which is filtered off and disposed in a silver residues containing bin. The remaining solution has to be placed in a container for toxic inorganic waste.

PHOTOS

Fig. 1 A 0.05M solution of K_2CrO_4 ready to be reacted with silver nitrate solution

Fig. 2 … and the result when treated with $AgNO_3$ 0.1M.

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HEXATHIOCYANATE IRON COMPLEX

OVERVIEW

Iron (Fe) is closely linked to the cosmic world and to humanity. This metal was present in huge amounts during the Solar System formation. It travels the Cosmos associated with nickel (which decays to Fe). It's the most abundant element on Earth [1]. The names "*iron age*" and "*industrial revolution*" are referring to the progress of human civilization.

As it has a very important role in constructions, technology and biochemistry iron related studies are important even at secondary level. The interaction of Fe (III) ions with potassium thiocyanate KSCN is important from analytical point of view (Fe^{3+} recognition), interesting discussions can be conducted when studying Le Châtelier's principle or when conducting practical demonstrations for precipitation reactions, not to mention an interesting application, known in history as *invisible ink* [2] which can ignite student's imagination and allow them to use this reaction for "practical" purposes.

MATERIALS AND METHODS

Prepare two different aqueous solutions: **SOLUTION 1**, one of ferric chloride ($FeCl_3$) 3% and respectively **SOLUTION 2** one of potassium thiocyanate (KSCN) 1%. Fill in a test tube with 5mL of SOLUTION 2. Add with a Pasteur pipette a few drops of SOLUTION 1. Observe the color changes. Repeat the experiment with solution 1 and solution 2 having different degrees of dilutions (for instance try with $FeCl_3$ 0.1% and KSCN 0.5%). Record the observations concerning the color intensity. If exact concentration variations are used these can be applied chemical equilibrium studies (effect of concentration on chemical equilibrium).

HAZARDS

Ferric chloride is irritant for skin and harmful if swallowed. Always use protecting clothing/eye protection. Potassium thiocyanate is harmful if swallowed or inhaled, if in contact with skin. It can release very toxic gases in contact with acids (e.g. hydrogen cyanide and sulfur dioxide) [3].

$FeCl_3$

KSCN

CHEMICAL REACTION EQUATION

$$\mathrm{FeCl_{3\,(aq)}} + \mathrm{KSCN_{(aq)}} \longrightarrow \mathrm{Fe(SCN)Cl_{2\,(aq)}} + \mathrm{KCl_{(aq)}}$$

$$\mathrm{FeCl_{3\,(aq)}} + 3\,\mathrm{KSCN_{(aq)}} \longrightarrow \mathrm{Fe(SCN)_{3\,(aq)}} + 3\,\mathrm{KCl_{(aq)}}$$

$$\mathrm{FeCl_{3\,(aq)}} + 6\,\mathrm{KSCN_{(aq)}} \longrightarrow \mathrm{K_3[Fe(SCN)_6]_{(s)}} + 3\,\mathrm{KCl_{(aq)}}$$

CONCLUSIONS

Fe(III) ions are hydrated in aqueous solutions forming the yellowish hexahydrate ion $Fe(H_2O)_6^{3+}$. In the presence of thiocyanate ions this complex forms the red iron thiocyanate ion $[Fe(SCN)(H_2O)_5]^{2+}$. As this complex has a very specific color for iron (**RED**) and is very easily formed, even in great dilutions, this reaction has an analytical value to identify the iron (III) ions. With well chosen concentration for SOLUTION 2 even rusty water, coming from iron made installations, can be colored highlighting the presence of Fe^{3+} ions. If more concentrated solutions are used an abundant dark red precipitate is formed due to further reactions of the iron ions, forming iron tris-thiocyanate $Fe(SCN)_3$ respectively the potassium hexathiocianato iron (III) complex $K_3[Fe(SCN)_6]$.

Similarities of this complex were found with the oxyhemoglobine as both are octahedral and red colored [4].

Due to its sensitivity to low iron concentrations this experiment can be used to demonstrate "how secret inks works". Also this reaction was performed several times in movies to simulate wounds and blood spills.

WASTE MANAGEMENT

The solution can be poured down the drain after considerable dilution or it can be collected in containers for less toxic inorganic waste.

PHOTOS

Fig. 1 Color recording when dropping FeCl3 0.1% in 0.5% solution of KSCN.

Fig. 2 More intense color if dropping FeCl3 3 % in a 0.5% solution of KSCN.

Fig. 3 If $FeCl_3$ is 3% when added a 1% solution of KSCN it becomes very dark.

Fig. 4 Potassium hexathiocianato iron (III) complex $K_3[Fe(SCN)_6]$

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THE FIRST SYNTHETIC MODERN PIGMENT - THE "PRUSSIAN BLUE"

OVERVIEW

Iron (Fe) is a very important metal for the living cells. It is present in many biochemical processes showing a wide range of properties and this is because Fe^{3+} can form various biochemically important complexes.

The present experiment studies the ability of the ferric and ferrous ions to interact with ligands generating various complexes.

Iron (III) chloride $FeCl_3$ reacting with potassium hexacyanoferrate (II) $K_4[Fe(CN)_6] * 3H_2O$ gives a historically famous complex (the Prussian Blue or Berlin Blue known by the chemists as iron (III) hexacyanoferrate (II)).[1] This substance is the first synthetic modern pigment. Despite its cyanide content the complex itself is considered safe for humans and is used as an antidote for Cs^+, Tl^+ and other heavy metal poisoning or even anticaking agent for both road and table salt [2]. Being quite safe for humans who can eat up to 10g/day without any harm, and due to its low solubility in water, this complex is a byproduct of the iron removal process from vines when treated with potassium hexacyanoferrate (II). As it is very sensitive to small amounts of iron Fe(III) in solution, it has analytical use as test for the ferric ion. Also the Lassaigne Reaction which is a qualitative elemental analysis of elements in organic compounds (N, C, X, S) makes use of this complex when identifying carbon and nitrogen in a form of CN^- ion.

MATERIALS AND METHODS

Prepare two different aqueous solutions:

SOLUTION 1; 3% potassium hexacyanoferrate (II) $K_4[Fe(CN)_6] * 3H_2O$. The color of this solution is straw yellow.

SOLUTION 2, 1% aqueous solution of ferric chloride ($FeCl_3$). This solution is pale yellowish, (in more concentrated solution it's brown and if very diluted (less than 0.1% is almost colorless).

Place in a test tube 4-5 mL of solution1 and a few drops of the solution2. Observe the color formed in the test tube. Repeat the experiment with a double or triple concentrated ferric chloride solution and notice the differences.

HAZARDS

Ferric chloride is irritant for skin and harmful if swallowed. Always use protecting clothing/eye protection.

Surprisingly some legislations are very strict concerning the potassium hexacyanoferrate (II) and potassium hexacyanoferrate (III) but, according to the European legislation and rules, no GHS symbols are allocated- only GHS statements: harmful for aquatic environment and avoid releasing in the environment [3].

$FeCl_3$		**$K_4[Fe(CN)_6] * 3H_2O$**	Harmful for aquatic life

CHEMICAL REACTION EQUATION [2]

$$4\,FeCl_{3\,(aq)} + 3K_4[Fe(CN)_6]_{(aq)} \longrightarrow Fe_4[Fe(CN)_6]_{3\,(s)} + 12\,KCl_{(aq)}$$

$$FeCl_{3\,(aq)} + K_4[Fe(CN)_6]_{(aq)} \longrightarrow KFe[Fe(CN)_6]_{(s)} + 3\,KCl_{(aq)} \quad \text{soluble form}$$

$$4\,FeCl_{3\,(aq)} + 3\,K_4[Fe(CN)_6]_{(aq)} \longrightarrow Fe_4[Fe(CN)_6]_{3\,(s)} + 12\,KCl_{(aq)} \quad \text{insoluble form}$$

CONCLUSIONS

When potassium hexacyanoferrate(II) reacts with a diluted ferric solution a new complex is formed $KFe^{III}[Fe^{II}(CN)_6]$. This is an intense blue colored compound which being colloidal gives the impression of being soluble. If ferric solution is added, or a higher concentration is used, the clear blue solution turns dark blue and a precipitate is observed $Fe^{III}[Fe^{III}Fe^{II}(CN)_6]_3$ written sometimes as $Fe_4^{III}[Fe^{II}(CN)_6]_3$.

There is still a debate whether Turnbull's Blue, called ferrous ferric cyanide identified as a $Fe_3^{II}[Fe^{III}(CN)_6]_2$ is different from the ferric ferrous cyanide. Some authors [1,4] consider the first complex the same with Prussian Blue.

WASTE MANAGEMENT

The solutions can be poured down the drain after considerable dilution and the iron chloride solution can be collected in containers for less toxic metal waste. The potassium ferrocyanide solutions have to be first reacted with Fe(II) or Fe(III) solutions and then the blue complex can be safely discarded in the drain.

PHOTOS

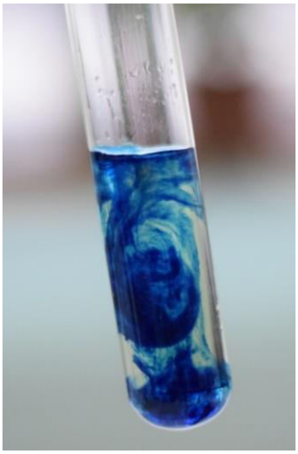

Fig. 1 Dropping $K_4[Fe(CN)_6]$ 0.3% over a 0.1% $FeCl_3$ solution.

Fig. 2 The color of the $KFe^{III}[Fe^{II}(CN)_6]$ complex in solution.

Fig. 3 Dropping $K_4[Fe(CN)_6]$ 3% over a 0.25% $FeCl_3$ solution gives a precipitate.

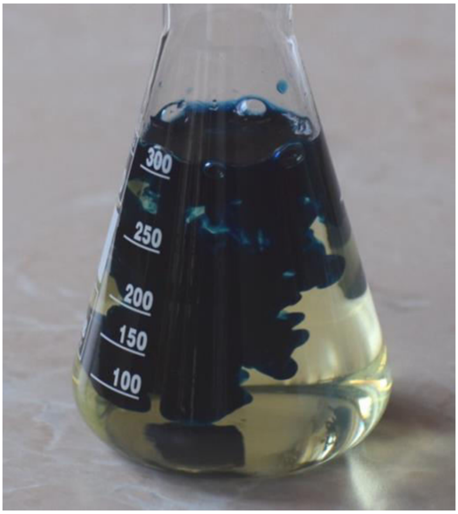

Fig. 4 Diluted potassium complex (0.08%) with diluted $FeCl_3$ (0.5%)

Fig. 5 More concentrated potassium complex (0.25%) with concentrated $FeCl_3$ (5.0%)

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NICKEL IDENTIFICATION WITH DIMETHYLGLYOXIME

OVERVIEW

Nickel (Ni) is the 24th most abundant element in the Earth's crust, comprising about 3% of the composition of the earth. It is the 5th most abundant element by weight after iron, oxygen, magnesium and silicon. It is a member of the transition series and belongs to group VIII B of the periodic table. Nickel metal and its alloys are used widely in the metallurgical, chemical and food processing industries, especially as catalysts and pigments.[1] We can mention here the role of the well known Raney Nickel as catalyst in hydrogenation reactions and our students often can answer to the question: What is a nickel? (Answer A **nickel**, in American usage, is a five-cents coin issued since 1866, being an alloy made of 25% nickel and 75% of copper [2] while its Canadian counterpart issued since 1922 was made of 99.9% nickel). The nickel salts of greatest commercial importance are nickel chloride, sulfate, nitrate, carbonate, hydroxide, acetate and oxide many of them widely found in the chemistry laboratory due to the various properties of the nickel salts. Of a great pedagogical importance are the experiments in which various nickel salts are reacted with solutions of dimethylglyoxime (a chelating agent used in the gravimetric analysis of nickel ores).[3]

This experiment is suitable for the secondary chemistry classes when presenting the students aspects related to the coordination compounds, precipitation reactions, the properties of transitional metals and/or qualitative and quantitative analysis.

MATERIALS AND METHODS

In order to identify the presence of nickel in different solutions a chelate is prepared by using an alcoholic solution of 1% dimethylglyoxime. This oxime is very sensitive to the presence of nickel salts giving an insoluble bright red precipitate when reacting with soluble Ni^{2+} salts (rose if very diluted solutions are used).

To prepare a 1% (m/v) solution of dimethylglyoxime, weigh 0.5g of solid dimethylglyoxime and dissolve it in a previously measured volume of 50 mL of ethanol.

Prepare different aqueous solutions of Ni^{2+} salts having a concentration of 0.05M. For this you have to take into account the structure of the crystalohydrate when weighing them for all the common nickel salts are hydrated (ex. $NiCl_2*6H_2O$, $NiSO_4*7H_2O$, $Ni(NO_3)_2*6H_2O$). For instance a

0.05M solution of $NiCl_2$ is obtained when weighing 0.594 g of $NiCl_2*6H_2O$, transferring it into a 50 mL volumetric flask, dissolving it in approx. 30 mL of distilled water then filling in the flask to the mark with the needed amount of distilled water. For 50 mL of 0.05M solution of $NiSO_4*7H_2O$ the amount of salt needed is 0.702g of crystalohydrate.

A. Take, in a test tube, 5mL of Ni^{2+} solution and add drop wise, from a Pasteur pipette, the 1% dimethylglyoxime solution. Observe the color appeared after mixing the two solutions. B. Do the same inverting the two solutions and adding Ni^{2+} solution to a 1% dimethylglyoxime solution.

Dilute the Ni^{2+} solutions 10 times. Observe the behavior of the diluted solutions when mixed with the dimethylglyoxime solution.

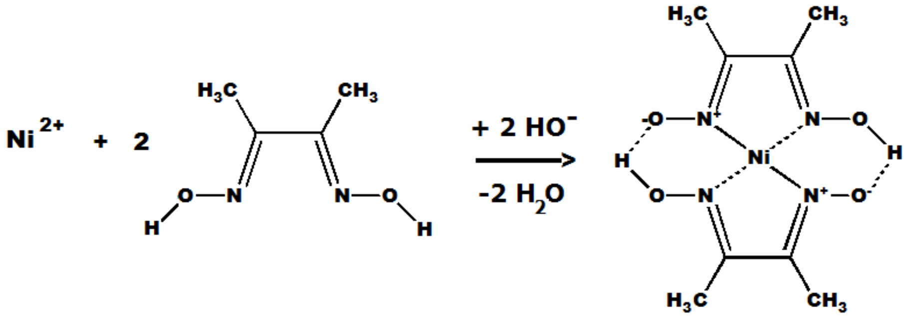
HAZARDS

$NiCl_2$

$NiSO_4$

$Ni(NO_3)_2$

CHEMICAL REACTION



bright red chelate

CONCLUSIONS

When Ni^{2+} salts are in the presence of dimethylglyoxime a chelate is formed [3] in a form of a red precipitate. Two different approaches can be applied.

A) If the 1% dimethylglyoxime solution is dropped over the Ni^{2+} solution, due to the chelating agent solution's lower density, the red insoluble complex is formed on the top of the reaction mixture (see Figure 1).

B) If reversed and the higher density Ni2+ solution is dropped over the alcoholic 1.0% dimethylglyoxime solution we observe the spread of the denser metallic solution in the volume of alcoholic one and the formation of the complex in the mean time, occupying and coloring the entire space. The reaction is specific due to the color formed. And it's also very sensitive for it reacts clearly with a very diluted, for instance 0.005M nickel solution.

WASTE MANAGEMENT

Ni^{2+} salts are dangerous to handle as they are toxic, carcinogenic, causing allergies and irritations, and they are very toxic to the aquatic life with long lasting effects. Caution is necessary when disposing Ni waste. One discarding method is to treat the Ni containing solutions with milk of lime, to filter off the precipitate and collect it in a container for less toxic inorganic salts. As for the filtered supernatant it can be poured down the drain [4]. The same disposal way can be applied for the filtered Ni(bis dimethylglyoximato) red complex.

PHOTOS

Fig. 1 A 1% dimethyl glyoxime solution dropped over a 0.015M $NiSO_4$ solution. The chelate is formed in the upper part due to a lower density of the oxime solution.

Fig. 2 A denser, 0.15M $NiCl_2$ solution, added to the oxime. The heavier, green, Ni solution is visible when added to the alcoholic solution while spreading in the whole space.

Fig. 3 Ni complex with dimethyl glyoxime, bright red complex obtained when reacting 0.05M $NiCl_2$ solution with the chelating agent.

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TITRATION OF CHLORIDE IONS WITH SILVER NITRATE

OVERVIEW

A continuation of experiment *Silver Identification With Potassium Chromate* this experimental work brings in attention the usefulness of this reaction in analytical approaches. Silver chromate formation can be used as indicator of the equivalence point in a titration process of Cl^- containing solutions with silver nitrate.

The titration of chloride (Cl^-) ions containing solutions in the presence of chromate CrO_4^{2-} with standard solutions of Ag^+ it's called Mohr's Titration Method [1] and is based on the solubility difference of the two precipitates: AgCl and Ag_2CrO_4 and on their different color. Useful discussions can be performed in the classroom around this subject, the theme being applicable when studying precipitation reactions, when discussing the solubility product concept or when introducing cation identification methods.

MATERIALS AND METHODS

To perform the titration process we need 0.1M solutions of both reagents $AgNO_3$ and K_2CrO_4 and a solution of NaCl (or other chloride) having a concentration close to 0.1M, but not exact.

Setup a titration equipment: magnetic stirrer and magnetic rod, 20 mL burette fixed on a stand with clamps, 100 mL beaker.

Pour in the beaker around 10 mL of Cl^- solution (with a presumable concentration of 0.1 iong/L), add a few drops of K_2CrO_4 0.1M and titrate it with $AgNO_3$ 0.1M until the content of the beaker starts to be rose. Stop the titration when the solution's color changes from milky white to pale rose, read the added volume and perform the calculations.

HAZARDS

Attention is to be paid when handling silver salts as they are very harmful for the environment. Chromate ions containing solutions are also dangerous for the environment and, as any Cr(VI) containing salts, they are highly toxic, irritants and carcinogenic.[3]

$AgNO_3$

K_2CrO_4

CHEMICAL REACTION EQUATION

$\mathbf{AgNO_{3(aq)} + NaCl_{(aq)} \text{ ---> } AgCl_{(s)} + NaNO_{3(aq)}}$

$\mathbf{2\ AgNO_{3(aq)} + K_2CrO_{4(aq)} \text{ ---> } Ag_2CrO_{4(s)} + 2\ KNO_{3(aq)}}$

CONCLUSIONS

The AgCl precipitate is white and agglomerated whereas the Ag_2CrO_4 is bright red. Interestingly if both anions are present in a solution a competition will start between the chloride and chromate silver precipitate when Ag^+ enters the solution. As AgCl has $K_{SP}=1.77x10^{-10}$ and Ag_2CrO_4 has $K_{SP}=1.2x10^{-12}$ [2] this gives the solubility of AgCl ($1.33x10^{-5}$ mol/L) which is lower than of Ag_2CrO_4 ($6.694x10^{-3}$ mol/L). When the Ag^+ ions enter the solution, the less soluble precipitate is formed (AgCl) until the entire Cl^- is consumed generating a white precipitate and a cloudiness of the reaction site. If no more Cl^- ions are in the solution, further addition of Ag^+ ions will generate the Ag_2CrO_4 precipitate which changes the color of the reaction mixture in red. If this experiment is done with the purpose of determining the content of Cl^- in a solution it is enough to stop the precipitation process when the reaction site becomes reddish, which represents the equivalence point for the Cl^- determination. Knowing the concentration of the silver nitrate solution and the volume needed to reach the equivalence point upon a simple calculation we can obtain the number of moles of Cl^- in the probe, respectively the concentration of the titrated probe.

This reaction is widely used in laboratory to determine the content of NaCl of different foods. From theoretical point of view it's a fine example of applying knowledge about precipitates and their K_{SP} in practical applications.

WASTE MANAGEMENT

As the solubility of AgCl is lower than of Ag_2CrO_4 a precipitation reaction is performed with diluted HCl giving solid AgCl which is filtered off and disposed in a silver residues containing bin. The remaining solution as it contains non reacted K_2CrO_4 which is dangerous for the environment; it has to be collected in a container for less toxic inorganic salts.

PHOTOS

Fig. 1 Cl^- containing solution treated with K_2CrO_4 ready for titration with silver nitrate.

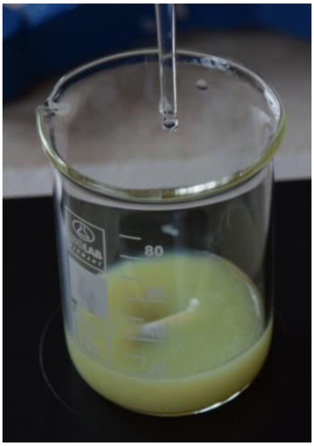

Fig. 2 The added silver nitrate forms the insoluble AgCl causing visible turbidity.

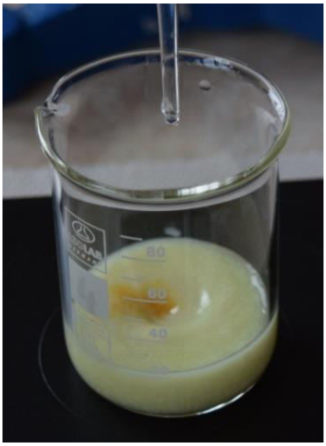

Fig. 3 The very moment when the silver solution droplet is entering the titration mixture.

Fig. 4 The white precipitate is abundant at the end but the content is still white.

Fig. 5 At the end, when all the Cl^- is consumed a red color of silver chromate is appearing.

END OF TITRATION

Fig. 6 An excess of titrant solution generates color enhancement

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DETECTION OF IONS WITH PRECIPITATION REACTIONS

OVERVIEW

In the fourth grade our secondary students get familiar with salts as a group of chemicals consisting of one (or more) positive metal ions and a negatively charged non-metal ion or ion group. Obviously, the ionic bond, predominantly exerted by Coulomb interactions, is characteristic. Because of the charges present, one would expect unlimited solubility in water at first sight, but in lots of cases the interactions between the ions are so strong that remaining a solid is more favourable from an energetic point of view. (in fact we are dealing with a competition between lattice energy and entropy gain upon dissolution).

The differences in solubility behaviour of salts can be used for analytical purposes. The observation (precipitation or not) upon addition of a known salt solution to an unknown one, excludes the presence of a number of ions. Repeating the procedure with different known salt solutions excludes more ions. Consequently, after each step there are less possibilities for ions to be present in the sample.

MATERIALS AND METHODS

Requirements:

Test tubes, copper(II)carbonate, copper(II)chloride, copper(II)nitrate, $AgNO_3$ 0,01M

Procedure:

Suppose we have three blue crystalline materials: one of them is copper(II)carbonate, another is copper(II)chloride and the last one is copper(II)nitrate. Our mission is to identify all three of them.

The procedure is as follows: put a sample of each of the materials in separate test tubes. First, add some water in order to dissolve the samples. One of the samples will not dissolve; this must be copper(II)carbonate. Now, we have to find out which one is the nitrate and which one the chloride. Nitrates are all soluble, so whatever we add to the nitrate, a precipitation with nitrate will never occur. So we have to distinguish between nitrate and chloride by a precipitation reaction with chloride. This can be done with silver ions because silver chloride is sparingly soluble. Addition of silver nitrate solution to one of the two remaining test tubes will finally give the answer.

HAZARDS [1-4]

$Cu(NO_3)_2$

$CuCl_2$

$CuCO_3$

$AgNO_3$

CHEMICAL REACTION EQUATIONS

$$Ag^+_{(aq)} + Cl^-_{(aq)} \rightarrow AgCl_{(s)}$$

CONCLUSIONS

Solubility behaviour of salts can be used for qualitative analytical purposes. The answer to the problem encountered in this experiment was found by simple reactions and observations.

WASTE MANAGEMENT

Solutions containing (heavy) metals should be collected and made available for further processing.

PHOTOS

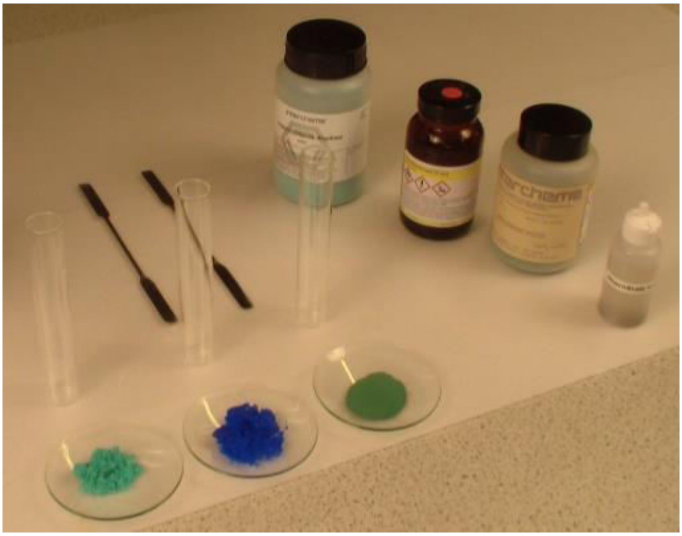

Fig. 1 At the start of the experiment

REFERENCES

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4. http://en.wikiorg/wiki/silver_nitrate

REMOVAL OF UNDESIRABLE PHOSPHATE IONS

OVERVIEW

In this experiment, undesirable phosphate ions are removed from water. The phosphate ions are removed by a technique known as precipitation. By adding a solution of an appropriate salt to the water, the phosphate will precipitate. The obtained phosphate ions are later used for supplements or anti-corrosion materials.

MATERIALS AND METHODS

- Water with phosphate ions
- Zinc chloride solution
- Funnel
- Test tubes
- Filtration paper

Pour the water with the phosphate ions in a test tube. Next, pour some of the zinc chloride solution in the same test tube. You will see the precipitation takes place. Then, fold the filtration paper and put it in the funnel and put the funnel on top of a clean test tube. Pour the suspension in the funnel with the filtration paper and wait for a couple of minutes to complete the filtration. When all the liquid is in the clean test tube the filtration has been completed.

HAZARDS

$ZnCl_2$

$Zn_3(PO_4)_2$

CHEMICAL REACTION EQUATION

After pouring the zinc chloride solution in the test tube, the phosphate will precipitate. The following reaction explains what happens during the precipitation:

$$3\,Zn^{2+}_{(aq)} + 2\,PO_4^{3-}{}_{(aq)} \longrightarrow Zn_3(PO_4)_2{}_{(s)}$$

CONCLUSIONS

The phosphate does precipitate when using a zinc chloride solution. The precipitate is zinc phosphate. It's a white solid substance at the bottom of the test tube. Next filtration was used to get the solid zinc phosphate out of the liquid. On the filtration paper, the zinc phosphate is the residue.

WASTE MANAGEMENT

The filtration paper should be put in the solid waste.

All the other solutions can be poured down the drain.

PHOTOS

Fig. 1 On the left you see the suspension after the two solutions were added. On the right the Erlenmeyer with a funnel and filtration paper

Fig. 2 The precipitate after the filtration.

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BREATHALYZER

OVERVIEW

Alcohol and traffic is known to be a dangerous combination. Therefore, in most countries alcohol concentrations in the blood while driving a vehicle are allowed to 0.5 promille maximum; for starting drivers and/or professional drivers even less.

Police officers are provided with so-called "breathalyzers" (Fig. 1) [1]; a small tube containing potassium dichromate. With alcohol, a redox reaction proceeds turning the orange colour into green. The flow pattern can be regarded as plug flow so a front separating orange and green proceeds through the breathalyzer. If a too large part of the breathalyzer turns green upon contact with your exhaled air you are not allowed to continue your trip and you are taken to the police station for more accurate analysis.

In this experiment, we will simulate the redox reaction with laboratory equipment.

MATERIALS AND METHODS

Chemicals:

Potassium dichromate, sulphuric acid 5 M, silica gel.

Equipment:

Glass tube, length 15 cm, diameter 0.5 cm, cotton wool, measuring cylinder, rubber hose, syringe, wooden skewer.

Procedure:

The first step is to make a mixture of 0.2 grams potassium dichromate and 10 grams of silica gel. Then a few drops of sulphuric acid are to be added. This has to be done carefully in a measuring cylinder, because at first contact the mixture is likely to splash. Mixing should proceed until the silica gel has absorbed the liquid entirely.

Now our laboratory breathalyzer has to be constructed. At one end of the glass tube some cotton wool is pushed in with the skewer. The cotton wool should not be packed too tight because air should be able to flow through it without too much resistance. Then the tube is filled with the silica gel mixture for about 10 centimetres. The other end of the bed is blocked with cotton wool as well. Now we drench a piece of cotton wool in alcohol and bring it into the syringe. The syringe is connected to the glass tube with a rubber hose. Then the alcohol vapour is pressed through the silica bed and the colour changes from orange into green.

HAZARDS *[2,3]*

$K_2Cr_2O_7$

H_2SO_4

C_2H_5OH

$Cr_2(SO_4)_3$

CH_3COOH

CHEMICAL REACTION EQUATIONS

$$Cr_2O_7{}^{2-}(aq) + 14\ H^+(aq) + 6\ e \rightarrow 2\ Cr^{3+}(aq) + 7\ H_2O(l) \qquad (1)$$

$$C_2H_5OH(aq) + H_2O(l) \rightarrow CH_3COOH(aq) + 4\ H^+(aq) + 4\ e \qquad (2)$$

Overall:

$$2\ Cr_2O_7{}^{2-}(aq) + 16\ H^+(aq) + 3\ C_2H_5OH(aq) \rightarrow$$

$$4\ Cr^{3+}(aq) + 3\ CH_3COOH(aq) + 11\ H_2O(l) \qquad (3)$$

CONCLUSIONS

With some simple laboratory equipment, we simulated the breathalyzer. Indeed, the reaction of potassium dichromate in acid medium upon contact with ethanol proceeds immediately as can be seen by change of colour from orange into green.

WASTE MANAGEMENT

The chrome-containing species should be collected in vessels designed for collection.

PHOTOS

Fig. 1 Industrial breathalyzer

Fig. 2 Requirements.

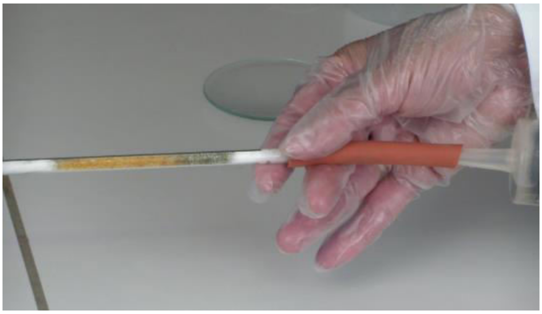

Fig. 3 The colour turns from orange to green upon contact with alcohol vapour.

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THE USE OF INDICATORS

OVERVIEW

Indicators, as used in this context, show that in a chemical reaction a certain situation is reached, usually by a drastic change in colour. One of the most well-known examples is the detection of iodine: addition of a few drops of a starch solution results in a deep blue colour due to the formation of a complex between iodine and starch. When iodine is consumed by some chemical reaction a sudden change from intensely blue to colourless indicates that there is no more iodine present.

Also for acid-base reactions indicators are widely used for visual determination of the equivalence point. The indicators themselves are weak acids or bases, the conjugated acids and bases having distinct different colours. The intensity of the colours allows accurate observation already upon addition of a few drops.

In this experiment we will make our own acid-base indicator. The colour of *red cabbage* is caused by a pigment *cyanidine,* the colour of which being strongly dependent the value of the pH. In this experiment we will investigate the applicability of red cabbage juice as a pH-indicator.

MATERIALS AND METHODS

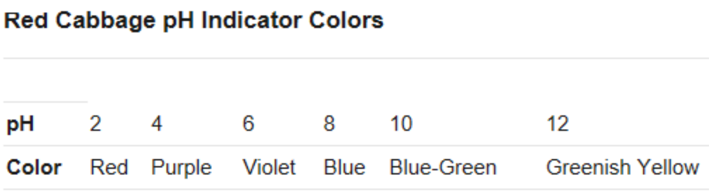
Chemicals:

1.0 M HCl; 1.0 M NaOH; red cabbage.

Procedure:

Take the outer leaves of a red cabbage and cut them into slices. Make sure not to lose too much of the pigment. Boil the slices in the beaker in 500 ml water for about 30 minutes. Store the extract in the flask.

Fill a test tube with 10 ml 1.0 M HCl solution (pH = 0), and add a few drops of the extract. Take 1 ml of this mixture, dilute 10 times (pH = 1) and again add a few drops of the extract. Repeat dilution until a pH value of 6 is reached. Then we start with 1.0 M NaOH (pH = 14), downwards to pH = 8. Finally we add a few drops of the extract to demineralized water. Now, we can compare the colours at unity pH intervals. The following results are to be expected [1]:



Red Cabbage pH Indicator Colors

pH	2	4	6	8	10	12
Color	Red	Purple	Violet	Blue	Blue-Green	Greenish Yellow

The results are visualized in figures 1 and 2.

We used our home-made indicator to estimate the pH-value of some commonly used products (see figure 3). Soft drinks turn out to be extremely acidulous (pH 2-3) but the acid taste is covered by addition of huge amounts of sugar. The third bottle contains calcium carbonate remover (pH 3-4); then mineral water (pH =7), detergent (9-10) and hand soap (10-11). All measured values seem reasonable.

HAZARDS [2,3]

NaOH and HCl:

CONCLUSIONS

Red cabbage juice shows several colour changes depending on the pH-value. Our results match well with literature data.

We used our indicator to estimate pH-values of commonly used products. Although the colour changes are not sharp, it can be used to obtain a first impression over a wide range of pH values (0-14).

WASTE MANAGEMENT

All chemicals involved can be treated as waste water.

PHOTOS

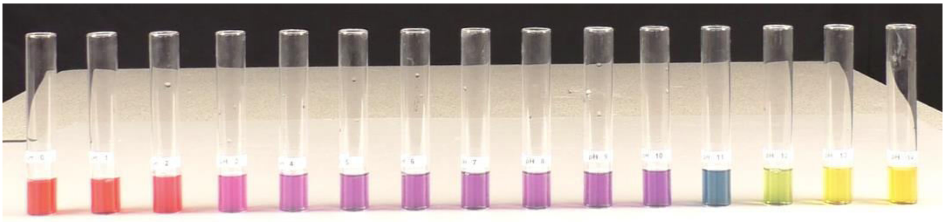


Fig. 1 Colour of red cabbage juice at 15 different pH-values

Fig. 2 Colour of red cabbage juice at 15 different pH-values.

Fig. 3 pH measurement of commonly used products with red cabbage juice.

REFERENCES

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MEASURING pH OF DIFFERENT SOLUTIONS WITH INDICATORS AND ELECTRODE

OVERVIEW

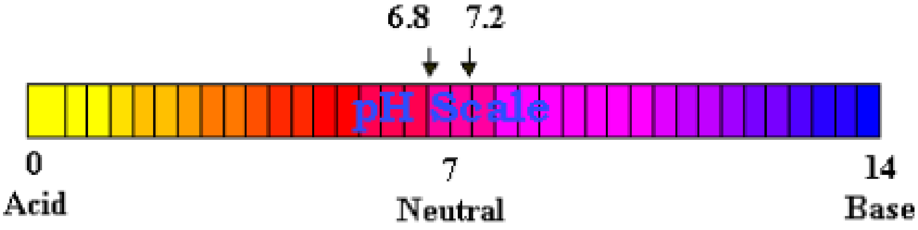
pH is a numeric scale used to specify the acidity or basicity of an aqueous solution. In English it is the abbreviated form of "potential of hydrogen" or "power of hydrogen". In addition to this in pH, p is the mathematical symbol of p minus logarithm. It is therefore the chemical symbol of hydrogen element. Namely, more precisely it is the negative of the base 10 logarithm of the molar concentration measured in units of moles per litre of hydrogen ions. More precisely, it is the negative of the logarithm to base 10 of the activity of the hydrogen ion.

$$\mathbf{pH = -\log [H^+] \;\; or \;\; pH = -\log [H3O^+]}$$

As a matter of fact, H+ ion is not found freely in water, on the contrary it is in the state of a mixture with water molecules, that is it is in the form of hydronium ion (H_3O^+).

The pH value of water is influenced by the acidic and basic elements which are dissolved in water. With the pH value of water the total influence of these elements has been measured. The more acidic element is found in water the lower its pH value.

The pH value of water is measured using a scale of 0-14. The hydrogen ions (H+) of pure water concentration equals the hydroxide (OH-) ions concentration as a result its pH value is neutral, that is 7.



On the other hand, with the release of negative logarithm of the molar concentration of OH-, the pOH equation can be obtained. Even if it is less pure water is in the dissociation state and the dissociation constant of water is 1,0x10-14.

MATERIALS AND METHODS

Calibrate the pH meter before pH measurement. Prepare in different beakers the solutions pH measurement will be made measure the pH values as potentiometric. Observe the acidity and basic properties of the same solutions using a pH paper. Compare the results taken with the results obtained through pH meter.

CONCLUSIONS

The pH value in all acidic solutions is expected to be between 0-7. If the solutions have a strong acidic property, its pH value can be between 0-3. In the solutions with basic property pH is expected to be between 7-14.
If the solution has a strong basic property, its pH value can be between 11-14.
The values obtained using indicators are expected to coincide with the results taken through pH meter. But faults in pH measurement with glass electrode can cause deviations in the results.

WASTE MANAGEMENT

All the solutions that pH values have been measured are discharged from the tap. The pH papers used are collected in the appropriate solid waste container.

PHOTOS

Fig. 1 The pH meter to be used in the pH measurement is first calibrated with buffer solutions.

Fig. 2 The measurements taken with pH meter are compared with the indicator measurements and the acidic and basic values of the samples are determined.

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TITRATION OF ACID SOLUTION WITH NaOH

OVERVIEW

In the fifth grade, our students have to study quantitative analyses. Titrimetry is a suitable subject to make students familiar with a wide scope of laboratory techniques to ensure a high degree of accuracy. It requires development of experimental skills and dedicated way of working. Theoretical knowledge from acid – base reactions is required for data reduction. After about 15 x 50 minutes of experimental exercises our students are supposed to be ready for the final test; an assignment they have to work out all by themselves rewarded with a mark that counts for their final exam.

Introduction:

Titration is a classic method of quantitative analyses. The principle is as follows: in a sample, the quantity of a certain compound has to be determined. This is done by dropwise addition of a reagent until the reaction is completed: the so-called "equivalence point". The equivalence manifests itself by a drastic change of a detectable property like colour, conductivity, potential or acidity. From the amount of added reagent the quantity of the compound to be determined can be calculated.

In this experiment we will restrict ourselves to acid-base titrimetry. The equivalence point is marked by a steep change of the pH-value that can be made visible by the use of a suitable indicator.

Theoretical background:

While titrating a strong acid with NaOH-solution, the reaction equation reads:

$$\mathbf{H_3O^+(aq) + OH^-(aq) \rightarrow 2\ H_2O(l)} \qquad \mathbf{(1)}$$

and for a weak acid:

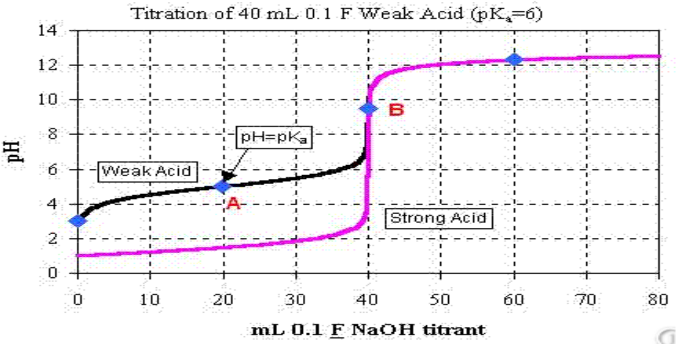
$$\mathbf{HZ(aq) + OH^-(aq) \rightarrow Z^-(aq) + H_2O(l)} \qquad \mathbf{(2)}$$

In the first case the equivalence point is reached when pH = 7. In the second case, the equivalence point is reached at a higher pH. The value can be estimated from:

$$[OH^-] = \sqrt{K_B . C} \qquad \mathbf{(3)}$$

with K_B the base constant of the conjugated base and C the (approximated) initial concentration. Note that at the equivalence point, the mixture can be regarded as a solution of the weak base Z^- with concentration C.

A schematic presentation of titration curves for both cases is given in the graph on the right.



MATERIALS AND METHODS

Requirements:

NaOH solution, hydrochloric acid, acetic acid (in vinegar), citric acid (in lemons), phenolphthalein, volumetric pipette, burette, beakers, volumetric flasks.

Experimental procedure

Before titration can be conducted, a sample has to be taken and diluted such that the concentration amounts to the order of magnitude of 0,1 M. Because we are dealing with quantitative analysis, all steps have to be carried out with maximum care and accuracy.

An accurate volume of the sample can be taken with a volumetric pipette. Before filling the pipette, preparative work is required to ensure that the concentration remains unchanged.

Filling of the volumetric pipette:

- Flush the pipette twice with demineralized water.
- Pour some of the sample in a beaker; fill the pipette with it and turn the pipette in such a way that the sample makes contact with the entire inner surface.
- Pour the liquid in a beaker in such a way that while pouring the inner side of the glass is wetted; flush one more time, then empty the beaker.
- Pour some new sample in the beaker. Repeat the procedure.
- Again pour some sample in the beaker. Fill the pipette, using a balloon, to 3 cm above the mark. Then remove the balloon and close the pipette with your index finger.

- Dry the tip of the pipette with a filter paper and adjust the liquid level to the mark by gently reducing the pressure exerted by your finger. Ensure the direction of observation is perpendicular to the marked spot.
- Then empty the pipette in a volumetric flask; keep the pipette in vertical position and the flask at an angle of 45^0.
- Wait for about 10 seconds to ensure the pipette is really empty.

Filling of the volumetric flask:

- Flush the flask one time with tap water and two times with demineralized water;
- Pour the sample to be diluted in the flask, using a funnel and a paperclip;
- Add demineralized water, just to slightly under the mark;
- Dry the part above the mark with filter paper;
- Complete filling by dropwise adding demi-water with a Pasteur pipette. Ensure that the direction of observation is perpendicular to the mark;
- Close the flask with a dry and clean stopper and shake at least ten times in order to homogenize the solution.

Filling of the burette:

- Flush twice with demineralized water;
- Flush twice with titrant;
- Fill the burette to above zero level using a glass funnel and make sure there are no air bubbles present;
- Remove the funnel and dry the inner side above the liquid level with filter paper;
- Let titrant escape dropwise to just underneath zero level;
- Read the initial volume in two decimals;

After intensive preparation we are ready for the titration itself. Take a portion of the diluted sample from the volumetric flash with a volumetric pipette, following the flushing and filling procedure exactly as described before. Pour the sample in an Erlenmeyer flask, add two or three drops of phenolphthalein solution and titrate.

Titration:

- Titrate dropwise, continuously shaking, until the expected pink colour is visible. The colouring has to be stable for at least 30 seconds.
- Now read and write down the final volume in two decimals.
- Repeat titration at least twice; the differences in added volume may not be bigger than 0.1 mL.

Chemicals:

At our premises we conducted titrations with:

- NaOH solution;

- hydrochloric acid;
- acetic acid (in vinegar);
- citric acid (in lemons).

HAZARDS:

NaOH

HCl

CH_3COOH

Citric acid

Direct contact with the human body should be avoided.

CHEMICAL REACTION EQUATIONS

$$H_3O^+(aq) + OH^-(aq) \rightarrow 2\ H_2O(l)$$

$$HZ(aq) + OH^-(aq) \rightarrow Z^-(aq) + H_2O(l)$$

CONCLUSIONS

As a final exercise, our students have to determine the amount of acetic acid in table vinegar. Most of them find an answer of about 4 m% which is reasonable.

The exam assignment is determination of the amount of citric acid in a lemon. Most students find values of 50 – 60 grams per lemon.

WASTE MANAGEMENT

Dilute solutions can be treated as waste water.

PHOTOS

Fig. 1 At the start….

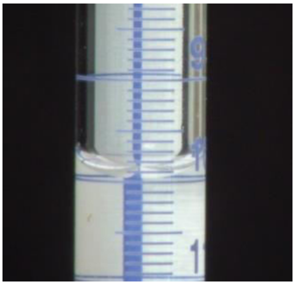


Fig. 2 Close-up of the burette

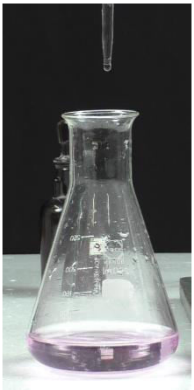


Fig. 3 Reaching the equivalence point.

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TITRATION OF AN IRON (II) SOLUTION WITH POTASSIUM PERMANGANATE

OVERVIEW

The reactions depending on the electron flowing among atoms, ions or molecules are called redox reactions (oxydoreduction reactions). Oxidation is the state of an atom, ion or a molecule to lose one or more electrons. These electrons are taken in by another atom, ion or molecule. The atom, ion or the molecule taking these electrons are reduced. When an element matter is oxidized, an equivalent amount of matter, at the same time, is reduced. Using this element principle a titrimetric determination of several matters having a redox reaction can be made. The point of a complete consumption of the matter to be titrated is called **equivalence point**.

All the titrations where $KMnO_4$ solution used is called **manginametri**. The number of the electrons received during this titration changes due to the medium being acidic or basic.

Potassium permanganate is one of the elements mostly used in redox titrations. The basic reasons for this are:

1. Permanganate ion is a highly strong oxidizing agent, and it's being able to oxidize any kind of weak concentrations.
2. So as to determine the turning point, an indicator is not needed.
3. It being cheap and easily obtainable however, apart from all these superior sides;
   a) It's not durable
   b) Its oxidizing sulphur is the two negative aspects of permanganate ions, when sulphuric acid is added.

MATERIALS AND METHODS

Prepare 0,02M potassium permanganate and put it into a burette. 0,75M 50 mL sulphuric acid solution is added into the 30 mL iron (II) sulphate solution in a beaker. Shake the obtained solution with the potassium permanganate solution in the burette. Continue shaking until a pink colour becomes permanent.

HAZARDS

$FeSO_4$

$KMnO_4$

H_2SO_4

CHEMICAL REACTION EQUATION

$5\ Fe^{2+} + MnO_4^- + 8\ H^+ \rightarrow 5\ Fe^{3+} + Mn^{2+} + 4\ H_2O$

CONCLUSIONS

In this titration, the Mn^{7+}in the MnO_4^- (permanganate) ion take $5e^-$ from the titrated element and thus reduces to Mn^{2+} ion, so the characteristic purple colour of MnO_4^- ion fades during titration. However upon the equivalent point the MnO_4^- will not be reduced, and makes the solution colourful and determines the end of the reaction.

WASTE MANAGEMENT

At the end of the experiment the thinned solution is discharged into a sink.

PHOTOS

Fig. 1 Iron(II) sulphate solution (a yellow colored solution) is obtained.

Fig. 2 A pink color is seen, when the equivalent point in titration is excessed.

Fig. 3 The shaking continues until the yellow solution turns into a pink color.

Fig 4. The shaking is ended when the solution turns into a stable pink color.

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TITRATION OF METANOIC ACID WITH SODIUM HYDROXIDE

OVERVIEW

Methanoic acid or formic acid is commonly found in nature. Remember when an ant bite you?!? Than a small amount of metanoic acid was injected under the skin provoquing pain. In fact formic acid comes from *formis* (lat. = ant). [1] And when by accident you touched a common nettle or stinging nettle (*Urtica dioica)*. You remember the pain?!?! The plant bears some tiny needles which inject several chemicals when touched, chemicals which can be very painful, among them being the formic acid.

Formic acid is a carboxilic acid. It's acidity is expected to be lower that those of strong mineral acids as HCl, HNO_3, H_2SO_4, $HClO_4$ but higher than those of other carboxilic acids. The present experimetal approach will deal with the titration of a ~ 0.1M formic acid solution with a strong alkaline base (NaOH 0.1M) in order to make a comparison with other acids' neutralisation behaviour.

MATERIALS AND METHODS

Setup a titration apparatus consisting of a magnetic stirrer, an electronic pH meter, a 25 mL burette with the necessary burette support and clamps, and a 100 mL beaker. Prepare a 0.1M solution of NaOH by weighing 0.4 g of solid NaOH and dissolving it in a 100 mL volumetric flask.

Place the beaker on the magnetic stirrer, add a magnetic bar followed by 10 mL of HCOOH ~0.1M. Fill in the burette with the NaOH 0.1M solution. Place the pH electrode in the solution on the edge of the beaker. Add distilled water (~10 more mL of water) until the solution covers the active part of the pH electrode. Add also a few drops of a mixture of an acid-base indicator: methyl orange and bromthymol blue.

Record the initial pH of the solution, notice its color and start the titration with the NaOH solution. Record the pH changes after small steps of titrating solution added (for instance every 0.2 mL). Record the pH when the first color change occurs. Observe at what amount of added base the pH is changing dramatically. Observe again the second color change. Add the base until the pH raises above 11. Fit the titration curve for the process reprezenting the pH against the added volume of base. Find the position of the equivalence point on the graph. Represent the pH interval where the methyl orange changes color and the interval when the bromthymol blue changes its color.

HAZARDS [2]

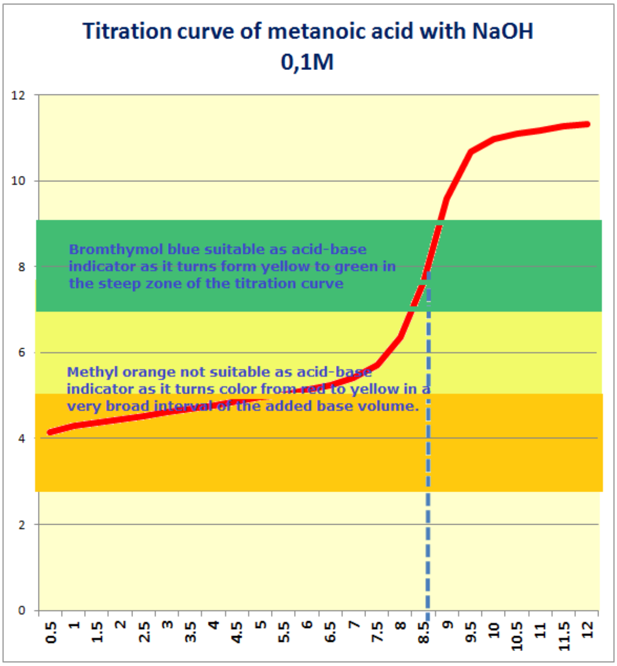
HCOOH  **NaOH** 

CHEMICAL REACTION EQUATION

$$HCOOH_{(aq)} + NaOH_{(aq)} \longrightarrow HCOONa_{(aq)} + H_2O_{(l)}$$

CONCLUSIONS

As the base is added to the formic acid solution we observe a color change at arround 5.5 mL of NaOH when the metyl orange turns from red to yellow. However this color change is not neat but very discreet over a 1.0 mL of base added in the beaker. As the pH is increasing slowly after this point, and the indicator already changed its color it's obvious this compound is not suitable to determine the equivalence point of the titration as it's not situated in the vertical area of the titrating curve (see figure). Continuing the addition there is neat color change from yellow to deep green (the resulted color between the blue of the bromthymol blue and the yellow methyl orange) at arround 8.3 mL of base added. In this moment a sudden pH raise is also noticed suggesting that the titration reached the equivalence point. This pH change is steep but not vertical as in the case of the titration of strong acids with strong base. From the titration's curve allure it's obvious that the formic acid is not a strong acid.



WASTE MANAGEMENT

Sodium formate is a food additive known under the code E237. The resulted solutions can be discarded in the drain as there is no chemical hazard for them.[3]

PHOTOS

Fig. 1 Start of the titration of formic acid ~0.1M with NaOH 0.1M

Fig. 2 End of the titration of formic acid with NaOH 0.1M.

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TITRATION OF PHOSPHORIC ACID WITH ALKALINE BASES

OVERVIEW

Phosphoric acid H_3PO_4 is an important acid, present in every living cell, being intricately connected with acid/base equilibrium, energy storage and release, cell multiplication, DNA etc. As being a weak tribasic acid it regularize, in two steps, the pH level of different biological environments. This is the reason why, at upper level secondary classes we can have a discussion with the students about its role in the living cells, the pH regulation mechanisms through different acid/base equilibrium. In order to see how phosphoric acid is reacting with bases a titration process is proposed with the recording of the pH at every step of added base. This will allow the teacher to address other subjects such as neutralization reaction, acid-base titration, pH measurements and graphic representation of different measurements.

The titration is recorded on a graph called **titration curve.** This will contain the volume of the titrant as independent variable and the pH as dependent variable [1]. The aim of this experimental work is to show that a poliprotic acid loses its protons in several steps according to the acidity constants. Using the graph, the equivalence points will be highlighted and the different acid dissociation constants (pKa) could be estimated.

MATERIALS AND METHODS

This experiment deals with an experimental work carried out in order to pinpoint the equivalence points of a titration when a triprotic acid (phosphoric acid H_3PO_4) is neutralized (titrated) with a strong monoacid base. Two different approaches are proposed in the present experimental work.[2]

1. The titration of 10 mL of H_3PO_4 aprox. 0.1M with a 0.1M solution of alkaline base measuring, with a benchtop pH-meter, the pH level of the solution at every 0.5 mL of added alkaline solution to the acid. In order to make a comparison between the efficiency of different bases during the neutralization reaction NaOH and KOH are used alternatively.

2. The second approach could be envisaged by measuring with an accurate benchtop meter the pH variation during the titration, associated with the color turn of two acid/base indicators used in the process (methyl orange and bromothymol blue). This will develop specific titration skills for the one who operates the experimental

assembly as it will allow him/her to observe what color is a certain indicator when an equivalence point is reached.

The first approach is presented as follows: 20 mL of H_3PO_4 0.1M is poured in a 150 mL (preferably tall) beaker and placed on a magnetic stirrer. The Berzelius beaker is preferred to the Erlenmeyer flask in order to allow the sensor to fit in, to have a higher level of the solution to be titrated to cover the sensor and to keep space for the titration process as well. Fill in a burette with a freshly prepared 0.1M KOH solution and place it over the beaker. Insert a small stirring rod in the acid solution and turn on the magnetic stirrer. Adjust the revolution speed of the magnetic rod in order to ensure a calm but efficient stirring without spilling out the solution. Setup a pH bench meter with a proper pH sensor, calibrate the pH-meter and insert the sensor in the beaker in such a way that you avoid collision between the stirring rod and the sensor. Avoid mechanical shocks as it is made of a very thin glass. If the level of the solution is not high enough distilled water can be added to the system. This will not alter the equivalence points to be measured but will allow the sensor to be fully immersed in the solution. Different approaches can be applied concerning the burette.

a) It can be a 50 mL burette which will allow the user to observe two equivalence points (after 20 and 40 mL of alkaline solution has been titrated) followed by a refilling of the burette and continuation of titration.

b) A larger burette will allow titrating with 3x more solution than the volume of the initial acid.

If the operator wants to keep accurate volume reading and in the mean time it wants to avoid the refilling of the burette a 0.2M solution of base might be envisaged. In this case the equivalence points will be measured at 1/2 and 1/1 of the added volume of base compared to the initial volume of the acid.

In order to avoid reading errors during the titrations the titrating solution has to be added drop wise directly into the acid solution avoid spilling or dropping on the edge of the flask. Once titration has been started, read and write down the measured pH on the benchtop's screen at each 0.5 mL of titrated KOH solution into the acid solution. Redo the experiment with a 0.1M (or 0.2M) NaOH solution. Draw a graph plotting the pH against the volume of the added base. Compare both graphs and try to interpret the results.

HAZARDS

Alkaline bases are corrosive as it is the phosphoric acid. Handle the solution with care! Pay attention when fixing or filling in with solutions the burette.

H_3PO_4

KOH, NaOH

CHEMICAL REACTION EQUATION

$$H_3PO_{4(aq)} + NaOH_{(aq)} \longrightarrow NaH_2PO_{4(aq)} + H_2O_{(l)}$$

$$NaH_2PO_{4(aq)} + NaOH_{(aq)} \longrightarrow Na_2HPO_{4(aq)} + H_2O_{(l)}$$

CONCLUSIONS

The graph is created on a millimeter Engineering Graph Paper in order to be easy to read any data once the titration curve is drawn. Mark on the paper the sets of coordinates you get when measuring the pH during the titration. The added volume of the base solution is the x value and the measured pH is the y value. Draw the graph and look on the curve you get. This curve has two inflexion points in the middle of two pH jumps corresponding to the neutralization steps carried out during the titration. The first and most visible jump is the $H_3PO_4/H_2PO_4^-$ neutralization step and the second is corresponding to the $H_2PO_4^-/HPO_4^{2-}$ neutralization step. The third one HPO_4^{2-}/PO_4^{3-} is not visible due to the slight acidity of the third hydrogen of the phosphoric acid. To calculate the pKa corresponding to both ionization steps we have to measure on the graph the pH of the point halfway from the beginning of the curve and the first equivalence point respectively the pH of the point situated at the halfway between the two equivalence points. In these points the pH=pKa so pKa_1 is corresponding for the $H_3PO_4/H_2PO_4^-$ equilibrium and the pKa_2 is for the $H_2PO_4^-/HPO_4^{2-}$ neutralization step.[3]

WASTE MANAGEMENT

Phosphates levels in water generally are an indicator of human activity. Large levels of these become detrimental when they over fertilize aquatic plants creating stepped up eutrophication. If too much phosphate is present in the water the algae and weeds will grow rapidly. [4] This is why we have to avoid large quantities of phosphates to being released in the water. Waste water treatment plants are applying procedures to reduce phosphate levels in the water before being released in the nature as phosphates are accumulating from human and animal waste, laundry, cleaning, industrial processes and fertilizers runoff.

The resulted solutions, as they are very diluted and in very small quantities to harm the environment, may be poured down the drain.

Attention has to be paid not to pour very acidic nor very basic solution down the drain.

PHOTOS

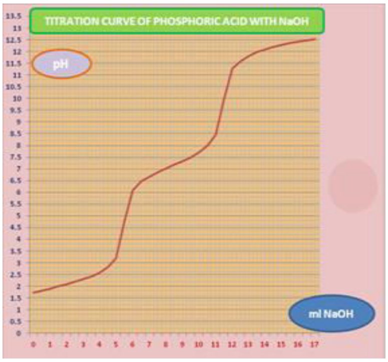

Fig. 1 The shape of the titration curve when 10 mL of phosphoric acid H_3PO_4 0.1M is titrated with NaOH 0.2M. We observe only two inflexion points on the curve when NaH_2PO_4 and respectively Na_2HPO_4 are completely formed. No signs of the Na_3PO_4 formation are seen when recording the pH during the titration.

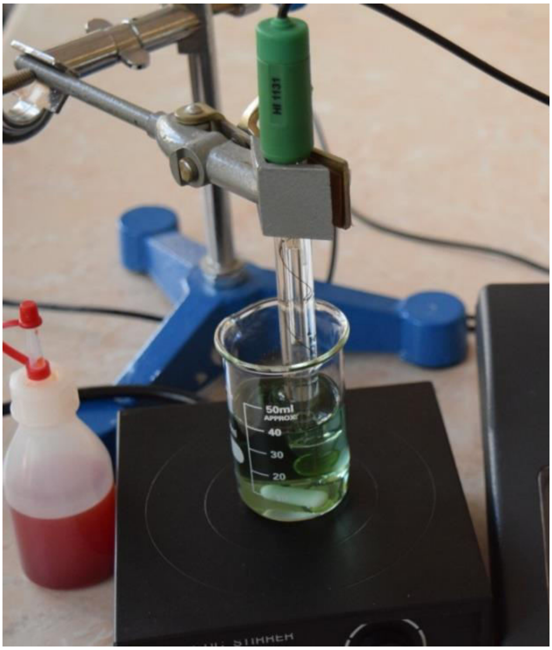

Fig. 2 Experimental setup to record pH change during the titration of H_3PO_4 with NaOH

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PREPARATIONS OF DIFFERENT SOLUTIONS HAVING AN ACCURATE CONCENTRATION

OVERVIEW

Lab experiments and types of research often require preparation of chemical solutions in their procedure. We look at preparation of these chemical solutions by weight (w/v) and by volume (v/v). The glossary below cites definitions to know when your work calls for making these and the most accurate molar solutions. To this we add information designed for understanding how to use the pH scale when measuring acidity or alkalinity of a solution.

<u>Glossary basic terms for better understanding;</u>

<u>Solute</u> - The substance which dissolves in a solution

<u>Solvent</u> - The substance which dissolves another to form a solution. For example, in a sugar and water solution, water is the solvent; sugar is the solute.

<u>Solution</u> - A mixture of two or more pure substances. In a solution one pure substance is dissolved in another pure substance homogenously. For example, in a sugar and water solution, the solution has the same concentration throughout, ie. It is homogenous.

<u>Mole</u> - A fundamental unit of mass (like a "dozen" to a baker) used by chemists. This term refers to a large number of elementary particles (atoms, molecules, ions, electrons, etc.) of any substance. 1 mole is 6.02 x 1023 molecules of that substance. (Avogadro's number).M

In this experiment Potassium permanganate is used. Potassium permanganate is an oxidising agent with disinfectant, deodorising and astringent properties. Its chemical formula is KMnO4 and it is sometimes called by its common name Condy's crystals. In its raw state potassium permanganate is an odourless dark purple or almost black crystal or granular powder. The main form of use is a potassium permanganate solution that is made by dissolving crystals or powder in water. There is also a more convenient 400mg tablet form of potassium permanganate available to prepare topical solutions. Crystals and tablets are available in pharmacies and garden shops.

Potassium permanganate is difficult to obtain in a very pure state and cannot therefore be used as a primary standard. Potassium permanganate decomposes slowly in solution, especially in sunlight, and readily oxidises organic matter. Potassium permanganate decomposes slowly in solution, especially in sunlight, and readily oxidises organic matter. Because of this the solution should be used as soon as possible after standardisation and should be stored in a dark bottle in the absence of light.

Potassium permanganate is useful in the following dermatological conditions:

a) Infected eczema and blistering skin conditions

b) Wound cleansing, especially weeping ulcers or abscesses

c) Fungal infections such as athlete's foot

MATERIALS AND METHODS

- 1. 0.4g potassium permanganate is weighted on scales.
- 2. After being weighted the potassium permanganate is put into a beaker and dissolved with distilled water.
- 3. This is then added into a 250mL flask.
- 4. Finally it is completed by adding distilled water until the line, in which produces 0, 01 molar potassium permanganate solution.

HAZARDS

Potassium permanganate (KMnO4)

WASTE MANAGEMENT

Potassium permanganate crystals and concentrated solutions are caustic and can burn the skin. Even fairly dilute solutions can irritate skin and repeated use may cause burns. If redness or irritation continues, notify your doctor. When preparing solutions make sure that the crystals or tablets are fully dissolved in water before usage. Potassium permanganate soaks are not suitable for dry skin conditions. Note that potassium permanganate may leave a brown stain on skin and nails as well as the bath or vessel holding the solution. Store locked up. Keep container tightly closed and in a well-ventilated place, where it is in a cool dry place and away from incompatible materials.

PHOTOS

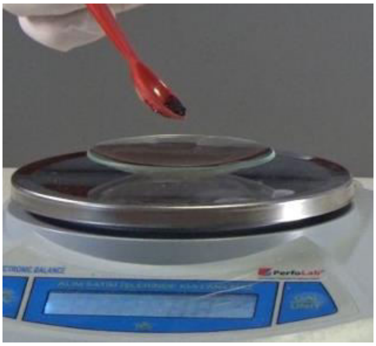

Fig. 1 0.4 g of Potassium permanganate KMnO4 is weighted on scales.

Fig. 2 After being weighted the potassium permanganate is put into a beaker and dissolved with distilled water. This is then added into a 250 mL flask.

Fig. 3 Finally, it is completed by adding distilled water until the line, in which produces 0.01 molar potassium permanganate solution.

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BUFFERING PROPERTIES OF AMINO ACIDS

OVERVIEW

Amino acids, the wonder molecules staying at the basement of every living organism on Earth! Amino acids, the building blocks of the proteins, are peculiar molecules, generating special classes of compounds and behaving in a very particular way. Due to them, the pH of the living cells and body fluids can be kept in a safe range allowing the organisms to develop and to fulfill their functions.

The following experimental procedure demonstrates the ability of different amino acids to counterbalance the influence of strong mineral acids or strong alkaline bases due to their special properties. Amino acids posses a carboxyl group so they can serve as acids and an amino group thus they can serve as bases. [1] This means amino acids are amphiprotic but due to the fact that the acidic group can transfer its proton to the weak base, represented by the amino group, the molecule itself is neutral containing a positively charged end and a negatively charged end. Such a molecule is called zwitterion.[2] An amino acid contains a weak acid group and a weak basic group therefore it is a buffer, which are solutions of weak acids and their salts respectively solutions of weak bases and their salts. Buffers resist changes in pH because they contain both bases and acids capable to neutralize added H^+ or HO^- ions [1].

Imagine the following exercise: to ~99 mL water we add 1 mL of 1M solution of HCl. What is the expected pH of the final solution if its volume is 100 mL? Taking in account the formula of the molar concentration $C_M=m/V_S$ the final concentration of the HCl after dilution is $m=C_M*V_S$, $m=1.0*0.001=10^{-3}$ moles HCl in 100 mL of solution -> $C_{Mf}=0.001/0.1=0.01M$. As HCl is a strong monobasic acid $[H_3O^+]=10^{-2}$ ion g/L -> pH=2. This means that an addition of a 1 mL of a 1M strong acid solution to 100 mL of water will cause a **drop of the pH of 5 units**: from 7 to 2 (the concentration of the $[H_3O^+]$ ion is raising 100.000 times). If the same question is asked when 1 mL of 1M NaOH solution is added to 99 mL of pure water the result will be pH=12, **a rise of 5 units**.

Taking in account these theoretical calculations an experiment is setup in which the pH variation is measured when 1M strong acid and 1M strong base solutions are added to a 0.1M solution of an amino acid. Comparing the pH variations we obtain we have a qualitative and quantitative estimation on how amino acids are acting as buffers. In the mean time information about properties of amino acids is refreshed in the classroom making this experiment suitable for at least two different lessons.

MATERIALS AND METHODS

In order to setup the experimental method 1M of the following solutions are prepared: HCl 1M; HNO_3 1M; NaOH 1M; KOH 1M; glycine 0.1M, alanine 0.1M, aspartic acid 0.1M and proline 0.1M. For details on how to prepare a solution of exact concentration please refer also to the experiment "Preparing solutions of exact concentrations".

For 100 mL 1M HCl solution you have to measure 8.88 mL HCl 35% (d=1.173g/ mL at 20°C), pour it into a 100 mL volumetric flask and add water until the mark is reached. The same procedure is applied for 100 mL 1M NHO_3 solution when you need 7.68 mL NHO_3 60% to be diluted to a volume of 100 mL (d=1.3667g/ mL at 20°C). Proceed accordingly if you have solutions of other concentration. A 1M NaOH solution is obtained if you measure 4.00g NaOH and dissolve it into a 100 mL volumetric flask. For 100 mL KOH you need 5.610g solid KOH.

The amino acid solutions are prepared in a larger amount as they will be used in 100 mL portions at least 4 times. Therefore these solutions are to be prepared in 500 mL volumetric flask and the needs are: 3.7533 g Gly; 4.454 g Ala; 6.655 g Asp; 7.756 g Pro.

Take a 150 mL beaker, add 100 mL of 0.1M solution of Glycine. Place it over a magnetic stirrer and stir it at low speed. Insert in the solution the pH electrode of a previously calibrated pH meter. Measure the initial pH of the amino acid solution. Add slowly (using a 1 mL syringe) 1 mL of HCl 1M. Measure the pH after the addition. Repeat two times the addition of 1 mL of HCl 1M and measure the pH and respectively the pH drop/raise compared with the initial pH. Do the same experiment with HNO_3 1M; NaOH 1M; KOH 1M for all 4 amino acids. Mark your experimental observations in the table:

Table 1.

Reagent	**Glycine**		**Alanine**		**Aspartic acid**		**Proline**	
initial pH	7.23		7.26		2.48		7.50	
	pH	**ΔpH**	**pH**	**ΔpH**	**pH**	**ΔpH**	**pH**	**ΔpH**
HCl 1M 1 mL	3.80	**-3.43**	3.69	**-3.57**	2.18	**-0.30**	3.36	**-4.14**
HCl 1M 2 mL	3.38	**-3.85**	3.29	**-3.97**	1.92	**-0.56**	3.00	**-4.50**
HCl 1M 3 mL	3.16	**-4.07**	3.07	**-4.19**	1.70	**-0.78**	2.79	**-4.71**
HNO_3 1M 1 mL	3.52	**-3.71**	3.55	**-3.71**	2.06	**-0.42**	3.20	**-4.30**
HNO_3 1M 2 mL	3.16	**-4.07**	3.15	**-4.11**	1.73	**-0.75**	2.82	**-4.68**
HNO_3 1M 3 mL	2.94	**-4.29**	2.93	**-4.33**	1.48	**-1.00**	2.60	**-4.90**
NaOH 1M 1 mL	9.07	**1.84**	9.05	**1.79**	2.91	**0.43**	9.84	**2.34**
NaOH 1M 2 mL	9.42	**2.19**	9.37	**2.11**	3.25	**0.77**	10.16	**2.66**
NaOH 1M 3 mL	9.63	**2.40**	9.60	**2.34**	3.58	**1.10**	10.39	**2.89**

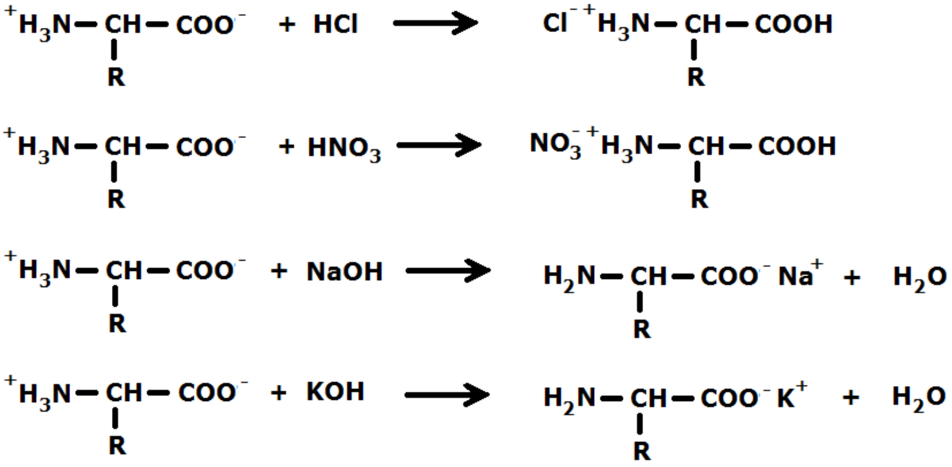
Reagent	Glycine		Alanine		Aspartic acid		Proline	
KOH 1M 1 mL	8.97	**1.74**	8.93	**1.67**	2.88	**0.40**	9.79	**2.29**
KOH 1M 2 mL	9.22	**1.99**	9.27	**2.01**	3.21	**0.73**	10.14	**2.64**
KOH 1M 3 mL	9.47	**2.24**	9.49	**2.23**	3.53	**1.05**	10.36	**2.86**

HAZARDS

Alkaline bases and mineral acids are corrosive.[3] Handle with care!

HCl, NaOH  **HNO_3**  **KOH** 

CHEMICAL REACTION EQUATION



$${}^{+}H_3N-\underset{R}{CH}-COO^{-} + HCl \longrightarrow Cl^{-}\,{}^{+}H_3N-\underset{R}{CH}-COOH$$

$${}^{+}H_3N-\underset{R}{CH}-COO^{-} + HNO_3 \longrightarrow NO_3^{-}\,{}^{+}H_3N-\underset{R}{CH}-COOH$$

$${}^{+}H_3N-\underset{R}{CH}-COO^{-} + NaOH \longrightarrow H_2N-\underset{R}{CH}-COO^{-}Na^{+} + H_2O$$

$${}^{+}H_3N-\underset{R}{CH}-COO^{-} + KOH \longrightarrow H_2N-\underset{R}{CH}-COO^{-}K^{+} + H_2O$$

CONCLUSIONS

When a strong acid enters an aqueous solution containing an amino acid, due to its zwitterionic structure it reacts with the carboxilato group protonating it and leaving protonated the amino group. A weak organic acid, accompanied by a salt end will be present in the solution instead of a strong mineral acid. This explains why the amino acid solutions resist pH changes. The same explanation is given for the addition of base to a solution of an amino acid. The base, entering the amino acid solution will neutralize the protonated amino end generating a weak base accompanied by a carboxilato group. The weak base formed will raise much less the pH comparing how a strong base would do normally. Looking at the data in Table 1. we see that different amino acids resist in a different ways to pH change. Most of them are buffering more efficiently the addition of bases than the strong acids.

WASTE MANAGEMENT

The resulted solutions, are diluted ones and in very small quantities. As they are usual salts of amino acids they don't harm the environment, so we may pour them down the drain. Attention has to be paid to properly neutralize all the remaining solutions before discarding them in the sewage system.

PHOTOS

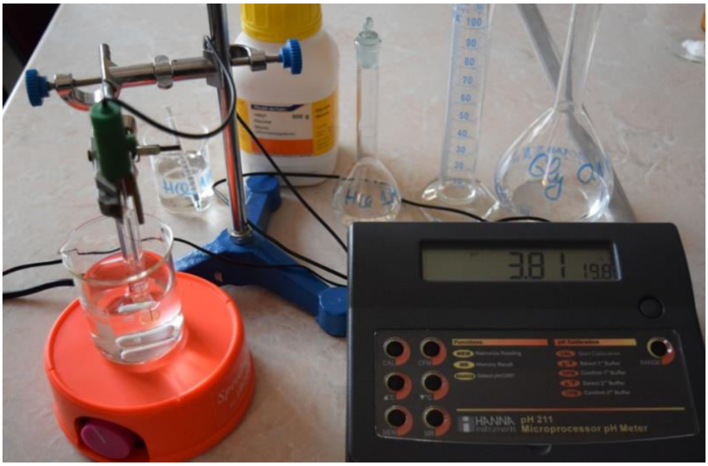


Fig. 1 Experimental setup for measuring the buffer properties of amino acid solutions (0.1M) when strong acid is added. Gly 0.1M reacted with 1 mL of HCl 1M (∆pH=3.43)

Fig. 2 Experimental setup for measuring the buffer properties of amino acid solutions (0.1M) when strong alkali base is added. Ala 0.1M treated with 1 mL of NaOH 1M (∆pH=1.79).

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COPPER OXIDE REACTION WITH PRIMARY AND SECONDARY ALCOHOLS

OVERVIEW

A simple experiment is demonstrating the capacity of copper (II) oxide CuO to remove hydrogen from primary and secondary alcohols, thus oxidizing them to aldehydes or ketones [1].

The copper (II) oxide (CuO) is well known as oxidant agent, especially with hydrogen or carbon monoxide when it undergoes reduction giving copper dust and releases water respectively carbon dioxide. A less common experiment is the demonstration of the oxidative properties of the CuO toward primary and secondary alcohols giving, as in the previous examples metallic copper.

MATERIALS AND METHODS

The experiment is very simple but very effective. Put in three test tubes 10-15 mL of different primary and secondary alcohols. Take a copper wire; coil it over a pencil in order to get a spiral. Heat the spiral over an oxidative flame of a gas burner. (The flame has to be blue as we learn from other analytical chemistry experiments). Pay attention not to melt the copper wire as its melting point is 1085°C. Keep the wire in the flame until it becomes black than immerse it at once in one of the chosen alcohols. Observe the changes occurred on the surface of the wire. Repeat the experiment with other alcohols, write down your observations and give an explanation to the observed phenomenon. As an experiment with negative results you may try to immerse a hot, black covered, copper wire in a tertiary alcohol.

HAZARDS [2]

The copper oxide is very harmful for the environment but the amount we get on a burned copper wire is insignificant. The used alcohols are flammable, they are irritants and methanol is toxic. Pay attention when using these reagents as open flame is used to burn the copper wire.

Ethanol, Isopropanol

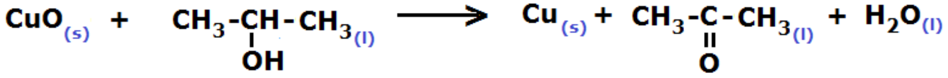
CuO

Methanol

CHEMICAL REACTION EQUATION

$$CuO_{(s)} + CH_3OH_{(l)} \longrightarrow Cu_{(s)} + CH_2O_{(l)} + H_2O_{(l)}$$

$$CuO_{(s)} + CH_3\text{-}CH_2\text{-}OH_{(l)} \longrightarrow Cu_{(s)} + CH_3\text{-}CHO_{(l)} + H_2O_{(l)}$$



$$CuO_{(s)} + CH_3\text{-}\underset{\displaystyle OH}{\underset{|}{CH}}\text{-}CH_{3(l)} \longrightarrow Cu_{(s)} + CH_3\text{-}\underset{\displaystyle O}{\underset{\|}{C}}\text{-}CH_{3(l)} + H_2O_{(l)}$$

CONCLUSIONS

The copper (II) oxide has demonstrated oxidative capacity toward many substances (H2, carbon, carbon monoxide, other metals) and it can even be used to safely dispose hazardous materials such as cyanide, hydrocarbons and halogenated hydrocarbons.[3] Its oxidative capacity toward alcohols is shown by generating aldehydes (from primary alcohols) or ketones (from secondary alcohols) itself reducing it to metallic copper.

When immersed, the hot black covered, copper wire becomes bright reddish due to the fresh layer of copper deposed on the surface replacing the superficial layer of copper (II) oxide. This process goes on only if the temperature is over 300°C. The experiment can be repeated, always ensuring enough copper (II) oxide on the surface of the wire by burning it. This experiment has analytic value as is used, in some cases, to identify primary and/or secondary alcohols in a mixture. Tertiary alcohols are not supposed to undergo oxidation. Therefore these alcohols does not reduce hot copper oxide.

WASTE MANAGEMENT

Discard the alcohols in a container for flammable organic solvents.

PHOTOS

Fig. 1 Oxidized copper wire after heating in oxidizing flame before immersing in ethanol.

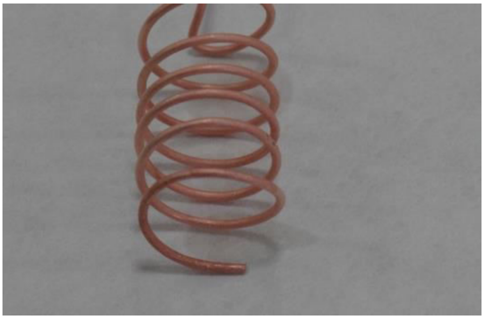


Fig. 2 Copper wire after heating and subsequent immersion in ethanol.

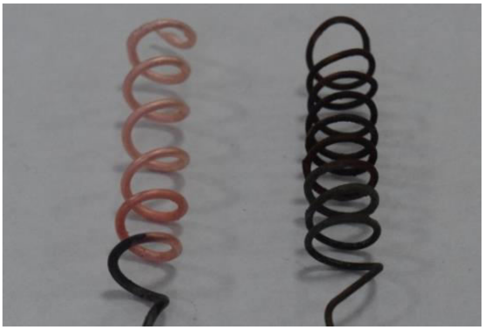


Fig. 3 Oxidized copper wire after heating in oxidizing flame and before immersing in methanol. Comparison with an already ethanol reducted wire (left).

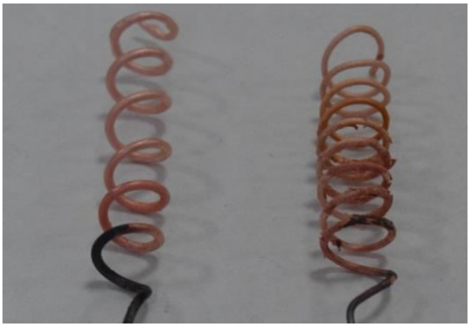


Fig. 4 Copper wire after heating and subsequent immersion in methanol (right). Comparison with an ethanol reducted copper wire (left)

Fig. 5 Hot oxidized copper wire immersed in methanol

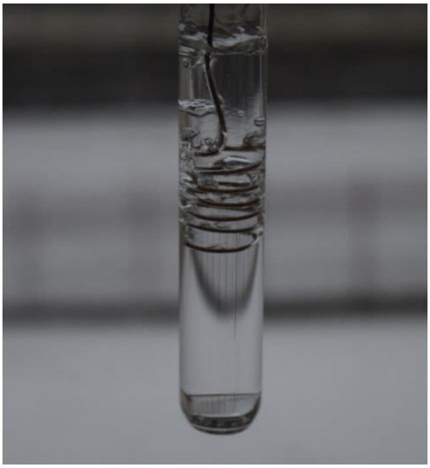


Fig. 6 Hot oxidized copper wire immersed in ethanol

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SPECTROPHOTOMETRIC DETERMINATION OF NITRITES IN WATER

OVERVIEW

Nitrite ions NO_2^- are found in water as intermediary states for ammonia oxidation process or for the nitrates reduction pathways. Their presence in water is associated with a high risk of cancers not to mention their contribution to cause methemoglobinemia. Therefore they must not be present in drinking water. [1]

In order to be aware of the risks certain water sources (family wells) can pose to the health of their users, an experimental procedure is demonstrated for the upper secondary students, at chemistry lessons, in order to measure the levels of nitrites in drinking waters.

In contrast to the experiment "*Water quality monitoring using nitrate and nitrite test strips*" the present approach will make use of a photometer to measure the color intensity of a water sample treated with the specific reagents in order to determine the nitrites level of a water sample.

MATERIALS AND METHODS

The used photometer is a commercially available analytic device, offering accuracy for nitrite concentrations in the range of 0.000 - 0.600 ppm (mg NO_2^-/L). The measurement principle is as follows:

- Switch on the photometer, check the battery level.

- Fill the glass cuvette, provided with the photometer, with 10 mL of water sample to the mark. Replace the cap and place the cuvette in the photometer. Press the "ZERO" button in order to set the zero point for the measuring device.

- Remove the cuvette, add the precalibrated reagent, available in a sealed packet, swirl gently the mixture for 20 seconds and place it back in the cuvette locking system of the photometer.

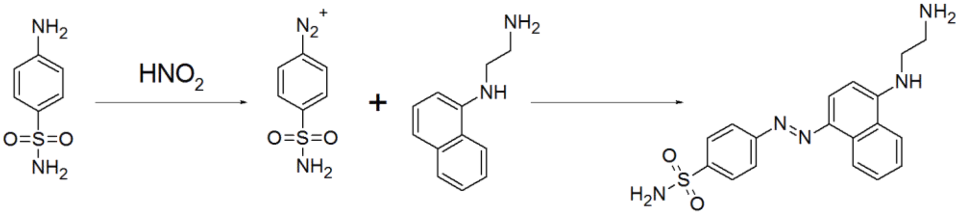
- Pres the "READ" button more than 3 seconds and the built in timer is activated. The photometer will measure the nitrites level after 15 minutes past. The result is displayed in ppm.

HAZARDS

No hazards are associated with the chemicals used to measure nitrites levels. Attention is needed when operating the reagents from the package. Do not breathe the dust which can be formed when taking out the reagent from its original package.

CHEMICAL REACTION EQUATION [2]

The procedure set up to measure the levels of nitrites in water is an adaptation of the EPA approved Diazotization method 354.1 to measure nitrite concentrations. [3] The reaction EPA Diazotization method 354.1 is represented in the following image:



CONCLUSIONS

When the reagent is added to samples containing nitrite, the sample will turn a pink tint; the greater the concentration, the deeper the color. The associated color change is then colorimetrically analyzed according to the Beer-Lambert Law. This principle states that light is absorbed by a complimentary color, and the emitted radiation is dependent upon concentration. For low levels of nitrite, a narrow band interference filter at 525 nm (green) allows only green light to be detected by the silicon photo detector and omits all other visible light emitted from the tungsten lamp. As the change in color of the sample increases, absorbance of the specific wavelength of light also increases, while transmittance decreases. [3,4]

Most of the times the analyzed water, coming from the tap and from several wells, show a negative response for the nitrites measurements. In some cases there is a positive response for the nitrite levels, the analyzed sample showing a very light pink coloration.

The image shown below represents the result of the nitrite level measurement for a solution in which there was a single tiny crystal dropped in a 1 liter flask of water. The color change for such a small concentration is dramatic, demonstrating how sensible the method is.

WASTE MANAGEMENT

The reagents used are not hazardous; the solutions can be safely discarded in the drain.

PHOTOS

Fig. 1 Experimental setup for nitrites determination. The result shown here is for a solution with more than 0.600 ppm (0.700 mg/l) NO_2^-. Observe the intense color for a tiny concentration.

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SPECTROPHOTOMETRIC DETERMINATION OF AMMONIA IN WATER

OVERVIEW

Ammonia when dissolved in water forms the ammonium cation in an equilibrium reaction with $K_b=1,8*10^{-5}$. Natural groundwater contain a very low level of ammonia (bellow 0.2 mg/L) In some cases (forests, strata rich in humic substances) the ammonia level can raise up to 3 mg/L. A higher level than 3 mg/L is an important indicator for fecal pollution. This is why the level of ammonia in drinking water has to be permanently monitored. [1]

It's important to say that the threshold odor concentration in water is aprox. 1,5 mg/L whereas the taste threshold of 35 mg/L is proposed for the ammonium cation while in contrast up to 4000 mg of ammonia is produced endogenously in the human intestine by the enzymatic breakdown of food. Ammonia is also formed due to the deamination of amino acids in liver, as a metabolite in nerve excitation and muscular activity. About 99% of metabolically produced ammonia is absorbed from the gastrointestinal tract and transported to the liver, where it is incorporated into urea as part of the urea cycle. Due to the high levels of ammonia found in the human body this compound is not of direct importance for health in the concentrations to be expected in drinking-water. [1] Despite this ammonia levels in water are monitored as a quality indicator of the water source used for drinking.

In the chemistry laboratory ammonia can be qualitatively determined using the Nessler reagent (see the experiment dealing with the *Synthesis Of The Nessler's Reagent And Ammonium Ion Determination*). Quantitative measurements can be performed using specially designed photometers which measure ammonia concentrations of less than 3.00 mg/L (ppm) ammonia-nitrogen. The procedure is based on the measurement of the yellow color of a sample treated with Nessler's Reagent. The greater the ammonia concentration, the deeper the color is. The associated color change is then colorimetrically analyzed according to the Beer-Lambert Law.[2]

MATERIALS AND METHODS

The following experimental procedure is dealing with water sample analysis for determining the ammonia concentration using a photometer. The used photometer is a commercially available analytic device, offering accuracy for ammonia concentrations in the range of

0.00 - 3.00 ppm (mg NH_3-N/L). The measurement principle is as follows:[2]

- Switch on the photometer, check the battery level.

- Fill the glass cuvette, provided with the photometer, with 10 mL of water sample to the mark. Replace the cap and place the cuvette in the cuvette holder and ensure that the notch on the cap is positioned securely into the groove. Press the "ZERO" button in order to set the zero point for the measuring device. The meter is zeroed and ready for measurement.

- Remove the cuvette, add 4 drops of the HI93700A reagent.

- Replace the cap and swirl the solution.

- Add 4 drops of HI93700AB reagent, replace again the cap and swirl again. Observe the yellow tint appearing if the water sample contains ammonia.

- Replace the cuvette into the holder and ensure that the notch on the cap is positioned securely into the groove

- Pres the "READ" button more than 3 seconds and the built in timer is activated counting 3.5 min. The instrument directly displays concentration in mg/L of ammonia nitrogen (NH3-N) on the LCD. To convert the reading to mg/L of ammonia (NH3), multiply the display by a factor of 1.214.

HAZARDS [3]

Mercury compounds present in HI93700AB are highly toxic, harmful and all are cumulative poisons. Sodium hydroxide is a highly corrosive reagent.

Nessler reagent

HgI_2

NaOH

CHEMICAL REACTION EQUATION

As for the use of the Nessler reagent as an analytical tool to determine the presence of the ammonium ion, the following reaction takes place.

$$2K_2[HgI_4]_{(aq)} + 4KOH_{(aq)} + NH_4^+{}_{(aq)} \xrightarrow[-K^+]{} HgO{*}Hg(NH_2)I_{(s)} + 3H_2O + 7KI_{(aq)}$$

CONCLUSIONS

The photometer measures the transmittance of the light passing through the colored sample. The transmittance is related to the absorbance which is directly related to the concentration of the ions to be measured [4]. As a general observation: the ammonia level of water coming from different wells is low, demonstrating the sources are not contaminated. The method used is very sensitive; a small addition of ammonium chloride (a few crystals) to a 250 mL water sample determines the apparition of the specific yellow color when treated with the specific reagents.

WASTE MANAGEMENT

As mercury (II) compounds are very toxic and harmful for the environment all the solutions or suspensions containing mercury (II) ions have to be neutralized with special reagents or collected in containers for toxic inorganic waste.

PHOTOS

Fig. 1 Very sensitive measurement -positive reaction when a tiny crystal of NH4Cl is dropped in a glass of 250 mL of water which is then analyzed.

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SPECTROPHOTOMETRIC DETERMINATION OF SULFATES IN WATER

OVERVIEW

Sulfates occur naturally in numerous minerals, including barite (BaSO4), epsomite ($MgSO_4 \cdot 7H_2O$) and gypsum ($CaSO_4 \cdot 2H_2O$). These dissolved minerals contribute to the mineral content of many drinking-waters. [1] In fact permanent hardness of water is due also to the content in calcium and magnesium sulfates dissolved in it, along with the chlorides of these two metals.

Reported taste threshold concentrations in drinking-water are 250–500 mg/L (median 350 mg/L) for sodium sulfate, 250–1000 mg/L (median 525 mg/L) for calcium sulfate and 400–600 mg/L (median 525 mg/L) for magnesium sulfate. [1]

As any water contains sulfates, their concentration is an important factor in determining the quality of a drinking water.

The sulfate ion (SO_4^{2-}) is very sensitive to the presence of Ba^{2+} ions as they form a fine crystalline precipitate of barium sulfate $BaSO_4$. Some water analysis methods rely on this equation to measure the level of sulfates.

In the last year of high school, students are taught precipitation reactions, the method presented bellow being suitable for this chapter, not to mention the optional course on water and food quality assessment methods where they can extensively use this reaction in their analytics studies.

MATERIALS AND METHODS

The following experimental procedure is dealing with water sample analysis for determining the sulfate concentration using a photometer. The used photometer is a commercially available analytic device, for sulfate concentrations in the range 0-150 mg/L (ppm) with accuracy of 1 ppm. The measurement principle is as follows:[2]

- Switch on the photometer, verify the battery level.

- Fill the provided glass cuvette with 10 mL of water sample to the mark. Replace the cap and place the cuvette in the cuvette holder and ensure that the notch on the cap is positioned securely into the groove. Press the "ZERO" button in order to set the zero point for the measuring device. The meter is zeroed and ready for measurement.

- Remove the cuvette and add the content of one packet of sulfate reagent. The reagent pack contains barium chloride dihydrate $BaCl_2*2H_2O$ and citric acid. A white turbidity appears in the cuvette due to the formation of barium sulfate. Shake the cuvette for 1 minute.

- Replace the cuvette into the holder and ensure that the notch on the cap is positioned securely into the groove

- Pres the "READ" button more than 3 seconds and the built in timer is activated. After 5 minutes the device is measuring the sulfate content. If on the display, the number 150 ppm is flashing this means the level of sulfates is higher than 150 ppm and the probe has to be diluted with distilled water and the procedure repeated the measurement being than interpreted according to the dilution of the probe.

The cuvette is a very sensitive piece. Do not stain it, avoid scratching it and avoid actions which can damage the cuvette's surface and interfere with the measurements.

HAZARDS [3]

According to the manufacturer, the content of the reagent pack is toxic and corrosive as it contains $BaCl_2$ and respectively citric acid. It can cause serious eye damage.

Barium salts

CHEMICAL REACTION EQUATION

$$BaCl_{2\,(aq)} + MgSO_{4\,(aq)} \longrightarrow BaSO_{4\,(s)} + MgCl_{2\,(aq)}$$

CONCLUSIONS

Water get enriched in minerals as it passes through different minerals in its way to the groundwater. The sulfates are quite common in water as some of them are water soluble contributing to the mineral content of the drinking waters. The analysis of water with a spectrophotometer relies on the sulfate capacity to create turbidity when reacted with barium salts. The reagent packet contains barium chloride which is soluble in water releasing Ba^{2+} ions which precipitate with the present sulfate ions giving barium sulfate $BaSO_4$. The white turbidity occurred in the reaction sample, as the barium sulfate formed is insoluble, is measured by a photometer specially calibrated for this type of measurements. The photocell of the device detects the amount of light transmitted through the sample. All photometers contain electronics that convert transmittance values to absorbance values, as absorbance is directly proportional to the concentration of ions in the sample.[4]

The higher the content of sulfates in water, the more turbid is the solution to be measured and according to Lambert-Beer's Law the higher is the result of the measurement.

WASTE MANAGEMENT

As barium (II) compounds are very toxic and harmful for the environment all the solutions or suspensions containing barium ions have to be collected in containers for toxic inorganic waste.

PHOTOS

Fig. 1 Experimental setup for sulfates analysis. Observe the turbidity of the analyzed water sample (at a measured level of 56 mg/L)

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CHLORIDE CONTENT DETERMINATION BY TITRATION WITH MERCURY NITRATE

OVERVIEW

Chloride is maybe the most common anion in water so its content is an important analytical indicator. Magnesium and calcium chloride are part of the permanent hardness for water whereas sodium chloride in water is responsible for the salinity of the sea water. Therefore measuring the chloride levels in water may be important.

The chloride content of a solution can be measured in different ways: either using silver nitrate with potassium chromate as indicator or with mercury nitrate using diphenylcarbazone as indicator. A third alternative for chloride ion determination is the mercuric thiocyanate method in the presence of ferric ions.

The present procedure refers to the second method, using mercury nitrate solution making use of commercially available chemical test kits.

MATERIALS AND METHODS

There are several, commercially available, chemical test kits. These are based on one of the described principles and the used solutions are calibrated to have specific concentrations in such a way that the user can easily determine the chloride content based on the volume of titrating solution used (mL or drop count).

The example described below makes use of one of these commercially available sets. A water sample of 5 mL is treated with 2 drops of Diphenycarbazone indicator and swirled to get a homogenized solution. The solution will become a reddish-violet color. From the second bottle add drop wise a solution of nitric acid until the solution turns yellow. The titration is performed with a standardized solution of mercuric nitrate added with a 1 mL syringe, the equivalence point being reached when the solution turns from yellow to violet. The content of chloride is determined by multiplying the volume of the titrant solution with 1000 the result being expressed in mg cloride/1L (ppm).[1]

If the results are lower than 100 ppm the precision of the method can be improved by repeating the procedure using a 50 mL probe and calculating the content of chloride accordingly (by multiplying the result with 100).

HAZARDS

Mercuric solutions are toxic and hazardous when handled. Attention is needed when handling the titrant solutions. Nitric acid is corrosive.

$Hg(NO_3)_2$ **HNO_3**

CHEMICAL REACTION EQUATION

$$Hg(NO_3)_{2\,(aq)} + 2Cl^- \longrightarrow HgCl_{2\,(s)} + 2NO_3^-{}_{(aq)}$$

CONCLUSIONS

Mercuric ions in the titrant solutions react with chloride in the sample to form mercuric chloride. After all of the chloride is in the form of mercuric chloride, the mercuric ions react with diphenylcarbazone indicator to form a pink-purple complex, which shows the endpoint of the titration. [2]

WASTE MANAGEMENT

As the reagents used for titration contain mercury ions, collect the reacted samples for safe disposal. Dispose the reacted solutions in a container for toxic inorganic waste.

PHOTOS

Fig. 1 Experimental setup for chlorine determination after adjusting the pH with nitric acid solution.

Fig. 2 The end point of the titration process is when the solution turns from yellow to violet. A single drop of titrating solution can cause at the end a dramatic change in color.

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DETERMINATION OF SULFITES IN WATER

OVERVIEW

Water can contain sulfites and their content is important to be measured as concentrations over 20 ppm (mg/L) can cause damage in boilers. High levels of sulfites are consistent with low pH, promoting thus corrosion. Sulfite ions are toxic to aquatic life forms; the chemical demand that sulfide produces on oxygen in water can destroy the delicate ecological balance of lakes, rivers and ponds. [1]

An experimental procedure is setup in order to measure the content of sulfites in waters.

MATERIALS AND METHODS

The method is using commercially available chemical test kits. The kit used in this experimental procedure is based on a iodometric titration method relying on the iodine generation process between iodate and iodide in the presence of sulfuric acid and subsequent reduction due to sulfite ions. The excess of iodine, after all sulfite was consumed by titration, will be visualized by reacting with starch giving a blue color which determines the end point of the titration. [2] The used chemical kit contains a standard solution of sulfamic acid, a standard solution of EDTA, a solution of sulfuric acid, a starch solution used as indicator and a titrating solution containing a mixture of potassium iodate and iodide which is added using a 1 mL syringe.

Take first a 5 mL sample of water in the provided graded beaker. Add 4 drops of the sulfamic acid solution followed by 4 drops of the EDTA solution. Acidify the sample with 2 drops of sulfuric acid and add a few drops of starch indicator. Proceed to the titration of the probe with the provided iodide/iodate solution until the colorless solution turns blue. Observe that adding a drop of titrating solution causes the colorization of the sample, which disappears in a second if the solution is swirled gently. Stop the titration when the solution is constantly blue in color. Read the volume of the titrating agent added (in mililiters) and multiply by 200 to obtain the result in ppm sulfites in the sample.

If the result is low, or the blue color appears right after the first drop of titrating solution, you can increase sensitivity using a 10 times bigger water sample. Add accordingly the necessary reagents: 40 drops of the sulfamic acid solution, 40 drops of the EDTA solution. Acidify the sample with 20 drops of sulfuric acid and add a few drops of starch indicator.

Repeat the titration this time multiplying the result of the titration (in mililiters) with 20.

HAZARDS

Sulfamic acid is irritant. [3]

H_2SO_4

Sulfamic acid H_3NO_3S

CHEMICAL REACTION

$$KIO_3 + 5KI + 3H_2SO_4 \longrightarrow 3I_2 + 3K_2SO_4 + 3H_2O$$

$$SO_3^{2-} + I_2 + H_2O \longrightarrow SO_4^{2-} + 2HI$$

CONCLUSIONS

Once the sample prepared with the necessary reagents the titration starts. The iodide/iodate solution is added in small samples using a 1 mL syringe. The iodide ions react with the iodate in the presence of the sulfuric acid to generate iodine (see reaction1). The iodine is subsequently reduced by the sulfites (if present in water) back to iodide ions. The titration goes on until the generated iodine is not reduced any more to iodide, so its presence is marked by the blue color of the complex made with the starch, present in the sample. The color change determines the end point of the titration. The solutions used for the titrations are of determined concentrations in order to make the procedure simple and effective. The titrating solution has a concentration that allows the direct determination of the sulfite content only by multiplying the volume of the added solution with a factor (which is 20 for a 50 mL sample).

If other volumes of sample are used add the necessary solutions and multiply the result accordingly. For instance a 5 mL sample will require 4 drops of sulfamic acid respectively of EDTA solutions; 2 drops of sulfuric acid and the multiplier factor will be x 200.

WASTE MANAGEMENT

The solutions can be poured down the drain as they don’t harm the environment.

PHOTOS

Fig. 1 Titration experimental setup for the measurement of sulfites in water

Fig. 2 End point of the titration process for determining sulfites content in water

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SULFITES DETERMINATION IN WINE

OVERVIEW

The wine is a well known drink, used by large masses of people, many times on a regularly base. Many specialists consider wine almost a living organism due to the many biochemical processes that can be witnessed during the wine lifetime between fermentation to the end of its existence, when it is poured in our glass.

Wine is prone to oxidation. An oxidized wine loses its quality (with some exceptions) and generally it's not suitable any more for consumption. To prevent this, SO_2 is added to the wine, in order to keep its freshness and prevent oxidation. The presence of this additive (E220 and related sulfites E228) [1], widely used in winemaking but in other food processing approaches too, is generally harmless unless you suffer from asthma [2].

Sulfites can occur even during the normal process of fermentation. The amount of sulfites that a wine can contain is regulated around the world. If a wine contains more than 10 ppm (mg/L) it must affix the label "*contains sulfites*". [2,3] A white wine can have up to 210 ppm of sulfites, a red one up to 160 ppm and a sweet wine can contain up to 400 ppm of this additive. Red wines contain typically less sulfites as they can be decolorized by this additive, and having more tannin which stabilize also the wine.

An experimental procedure is setup in order to measure the content of sulfites in wine.

MATERIALS AND METHODS

The method is an adaptation of the sulfite determination procedure in water using commercially available chemical test kits. The kit used in this experimental procedure is based on a iodometric titration method relying on the iodine generation process between iodate and iodide in the presence of sulfuric acid and subsequent reduction due to sulfite ions until all sulfite ions are removed from the sample. The excess of iodine will be visualized by reacting with starch giving a blue color which determines the end point of the titration. [4] The used chemical kit contains a standard solution of sulfamic acid, a standard solution of EDTA, a solution of sulfuric acid, a starch solution used as indicator and a titrating solution containing a mixture of potassium iodate and iodide which is added using a 1 mL syringe.

Take a 25 mL of wine sample in a beaker (preferably use the provided graded beaker). Add 20 drops of the sulfamic acid solution followed by

20 drops of the EDTA solution. Acidify the sample with 10 drops of sulfuric acid and add a few drops of starch indicator. Titrate the probe with the provided iodide/iodate solution until the colorless solution turns blue. Read the volume of the titrating agent added (in mililiters) and multiply by 40 to obtain the result in ppm sulfites in the sample.

HAZARDS

Sulfamic acid is irritant. [5]

H_2SO_4

Sulfamic acid H_3NO_3S

CHEMICAL REACTION

$$KIO_3 + 5KI + 3H_2SO_4 \longrightarrow 3I_2 + 3K_2SO_4 + 3H_2O$$

$$SO_3^{2-} + I_2 + H_2O \longrightarrow SO_4^{2-} + 2HI$$

CONCLUSIONS

Once the sample prepared with the necessary reagents the titration starts. The iodide/iodate solution is added in small samples using a 1 mL syringe. The iodide ions react with the iodate in the presence of the sulfuric acid to generate iodine (see reaction1). The iodine is subsequently reduced by the sulfites present in wine back to iodide ions. The titration goes on until the generated iodine is not reduced any more to iodide, so its presence is marked by the blue color of the complex made with the starch, present in the sample. The color change determines the end point of the titration. The solutions used for the titrations are of determined concentrations in order to make the procedure simple and effective. The titrating solution has a concentration that allows the direct determination of the sulfite content only by multiplying the volume of the added solution with a factor (which is 40 for a 25 mL sample).

If other volumes of sample are used add the necessary solutions and multiply the result accordingly. For instance a 5 mL sample will require 4 drops of sulfamic acid respectively of EDTA solutions; 2 drops of sulfuric acid and the multiplier factor will be x 200.

WASTE MANAGEMENT

The solutions can be poured down the drain as they don’t harm the environment.

PHOTOS

Fig. 1 Wine sample prepared for titration to measure its sulfite content

Fig. 2 End point of titration with excess iodine forming a blue complex with starch

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REFRACTOMETRIC DETERMINATION OF SUCROSE IN BEVERAGES

OVERVIEW

Sucrose, commercially known as sugar, was a rare commodity and traders of sugar became wealthy. In the mod 18th century, in Silesia, a method was discovered to extract it from beetroot so it was not any more necessary to rely on the sucrose from sugar cane which was imported from the Caribbean Islands later being the subject of a British blockade during the Napoleonic Wars. This was a period when sugar consumption had become regarded as a human necessity. Today sugar is still produced in a greater amount from cane as the beet sugar gave in 2009 only 20% of the total amount produced worldwide … but we are speaking about 175 million tons produced every year (2013). [1]

Globally in 2013 more than 45 million children were obese and these figures are expected to raise to 70 million by 2025, not to mention the figures for adults. [2] This obesity is attributed to a growing consumption of processed food, including heavily sweetened beverages. These figures are worrying and different measures are taken in different countries. The Romanian legislation is banning the selling of "soda" type beverages and sweets in schools but this is not enough to stop the wide usage of sweets and of sugary beverages among students.

Maybe a more profound approach in which the student (and not only) is taught how much sugar is in the products he is consuming on every day basis will raise conscience toward the high amounts of sugar (mainly sucrose) entering the body per year or in a lifetime. This is why in an optional Chemistry course given to the senior students in the high school we measure the sugar (sucrose) content of different beverages and make comparisons between different foods and their contents.

MATERIALS AND METHODS

The content of different sugary beverages is measured using a refractometer. There are several options to use such devices; most of them are relying on the refractive index of a sugar containing fluid which can generate the refraction of light either in an optical equipment or in an electronic one. So the sugar concentration can be determined by measuring how light is refracted in a sugar containing solution.

The described procedure make use of an electronic refractometer having a stainless steel probe fountain over a sensor capable to measure the refractive index and to display the result directly calibrated in % sugar or Brix degrees °Bx which represent grams of sucrose dissolved in 100g of solution.[3]

First you have to calibrate the refractometer with distilled water in order to set the "zero" point. Refer to the user's manual to do this as there are several devices commercially available.

Take different probes of sugary beverages (you may measure also "light" drinks to see the difference) and apply a drop on the fountain of the refractometer. Press the "read" button and record the value on the display which is either given in °Bx or in g of sugar/100mL of liquid. Taking in account that a tablespoon of sugar contains 12 g of pure sucrose count the amount of pure sugar you consume when drinking a glass of the analyzed beverage.

HAZARDS

No hazards are related to the measurements done with a refractometer. Just keep in mind not to scratch the probe fountain when cleaning it as any scratch will irreversibly alter the measurements. Therefore you shall use only soft tissue paper or soft cotton cloths to clean the fountain.

CHEMICAL REACTION EQUATION

This is a physical method of determining the content of a substance in a product. There are no chemical reactions occurring during the test.

CONCLUSIONS

A few examples of beverages and their sugar (sucrose) content is relevant to understand the amount of this disaccharide introduced in our body during the consumption of different soft drinks.

The great majority of soda beverages contain around 10% sucrose (9.9%, 10.2%, 10.3%), that means a 500 mL flask will contain ~50 g sucrose (more than 4 tablespoons of sugar in it).

The fleshy orange drink has 12% sucrose which corresponds to 5 table spoons of sugar in a 500 mL flask. A "cola light" drink contains 0.2% sugar so the sweet taste is attributable to the artificial sweeteners therein. The small amount of sugar present there comes from the caramel used in this type of beverages. An energizer contains more sugar: 15% which means that a 330 mL can contain the same amount of sugar as the 500 mL "soda" flask. The tea drinks have a 4,6% sugar content which suggest the use of artificial sweeteners to reach the same sweet taste of the product.

Measuring two homemade soft drinks (a black currant syrup and an elderberry one) contain 19.5% respectively 19.2%. This means that a

330 mL flask will represent for the consumer a sugar intake of 64.35 grams of sucrose (5.36 table spoons of sugar). Interesting discussions can be made after measuring the sugar content of the most common beverages. The most important is not to be aware of the concentration of sugar in our beverages, drunk on a daily basis but to be aware of the amount of sugar (sucrose) they represent and what means this for our body.

WASTE MANAGEMENT

Discard the remaining beverages in the sink with plenty of water.

PHOTOS

Fig. 1, 2 Some common beverages and their sugar content measured with a refractometer.

Fig. 3, 4 On the left a common soda drink and on the right elderberry syrup used often in Spring as a natural soft drink.

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REFRACTOMETRIC DETERMINATION OF SUGARS IN FRUITS

OVERVIEW

Sugar content of fruits is an important nutritional aspect and it has to be taken in account when attention is required for your diet. Some fruits have a high sugar content (grapes, plums, bananas) whereas others have low sugar (mostly the berry fruits: blackberries, raspberries, strawberries and the citrus fruits: grapefruit, orange, nectarines, mandarins).[1] There are examples for each end of the fruit-sugar-chart. The other fruits are considered fruits with a medium content of sugar. Of course the sugar content of a fruit can greatly vary depending of its sort, the place it was cultivated, the ripeness, etc. So we keep our analysis on thinking globally and taking in account all these aspects. On the other hand the sugar content of the fruits can be different concerning the three main sugars we can found there: glucose, fructose, sucrose. Whereas the apples and pears have a high content of fructose, the apricots, nectarines, peaches and mangoes have most of their fruit content as sucrose. Other fruits contain glucose and fructose in almost equal amounts (cherries, sour cherries, star fruit, blueberry, strawberry, blackberry, grapes).[2,3]

MATERIALS AND METHODS

The content in sugars of different fruits is measured using a refractometer. The principle is the same as for experiment "*Refractometric determination of sucrose in beverages*", using a different device which is capable to measure inverted sugar as most of the fruits are containing both fructose and glucose.

The described procedure make use of an electronic refractometer having a stainless steel probe fountain over a sensor capable to measure the refractive index and to display the result directly calibrated in % sugar. First you have to calibrate the refractometer with distilled water in order to set the "zero" point. Refer to the user's manual to do this as there are several devices commercially available. Take different fruits and place a few drops of fruit juice over the sensor in the stainless steel well of the refractometer. Press the "READ" button and record the measured values.

HAZARDS

No hazards are related to the measurements done with a refractometer. Use only soft tissue paper or soft cotton cloths to absorb the probe and to clean the well. Use distilled water to wash the remaining fruit juices to avoid measurement errors. Keep in mind not to scratch the probe fountain when cleaning it as any scratch will irreversibly alter further measurements.

CONCLUSIONS

Measurements performed over several types of grapes revealed sugar concentrations spanning between 15% and 22%. Surprisingly the sugar content of a lemon was 9.1%, an orange gave 7.3% fruits whereas a clementine was measured with 8.3%. Even for apples there were concentrations of sugar up to 15% the average being situated around 10%. Discussions can be made with the students, after performing the measurements, concerning their diet and the type of fruits they should eat on a daily basis.

WASTE MANAGEMENT

No special attention is needed for the waste disposal. Just keep clean and tidy the working place and use the bin for the fruits whose content was measured.

PHOTOS

Fig. 1-4 Some common fruits and their sugar content measured with a refractometer. (from left to right grapes, clementine (up), apple and pear (down))

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REFRACTOMETRIC DETERMINATION OF SUGARS IN WINES

OVERVIEW

The sweetness of a wine is determined by several factors, including not only the sugar content but also the alcoholic degree, the content in different acids, the presence of tannins and even if the wine is sparkling or not. It is well-known that alcohol is produced by the fermentation of the so called "fermentescible" sugars present in the grape must. Some of those sugars may not undergo the alcoholic fermentation process (due to different techniques used) and remain in the wine becoming "residual sugar". due to the fact that glucose is fermented in a much higher rate than fructose, arresting the fermentation will give a wine with mostly fructose as residual sugar. Also, there are some "non fermentescible" sugars in the grape must, mainly pentoses, which are found also in the wine as residual sugar. [1]

The sugar content of wines is important as they are labeled accordingly: DRY up to 4 g/L (some sources consider <2 g/L [2]); MEDIUM DRY between 4 and 12 g/L; MEDIUM between 12 and 45 g/L and SWEET over 45 g/L. Measuring/knowing the sugar content of a wine you may have a quality index of it. This is important as residual sugar can hide some of defects in the wine. Therefore it's said that you can best judge a wine if it's dry.

MATERIALS AND METHODS

The residual sugar content of different wines is measured using a refractometer using a refractive principle as described in the previous two experiments. For this analysis we use a device specially designed for measuring the sugars in wines, as this device takes in account also the alcohol content of the liquid and its influence on the refractive index. The described procedure makes use of an electronic refractometer having a stainless steel probe fountain over a sensor capable to measure the refractive index and to display the result directly calibrated in % sugar. [3]

Refer to the user's manual how to calibrate the refractometer and set the "ZERO" point, as there are several devices commercially available. Take different wines and use a Pasteur pipette to place a few drops of liquid over the sensor of the refractometer. Press the "READ" button and record the measured values. Than clean the well for the next probe.

HAZARDS

No hazards are related to the measurements done with a refractometer. Use only soft tissue paper or soft cotton cloths to absorb the probe and to clean the well. Use distilled water to wash the remaining wine to avoid measurement errors, especially when sweet wine is measured.

CONCLUSIONS

Measuring the sugar content of the wines you can make a classification of them in DRY, MEDIUM DRY, MEDIUM (SEMI-SWEET) or SWEET. You may ask the students what other parameters, measured for wines, are giving details about the quality of the product. As this measurement is based on the refraction of an incident beam any other substance which can alter the refractive index (acids, alcohol content different from the average) will shift the measured values causing either positive or negative errors.

WASTE MANAGEMENT

No special attention is needed for the waste disposal. You can discard the wine probes in the drain and keep clean your working place.

PHOTOS

Fig. 1-4 Sugar content measurement of four different types of wines from different countries.

REFERENCES

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DETERMINING THE COMMERCIAL GLUCOSE IN HONEY

OVERVIEW

The aim of this experiment is the determination of whether a honey sample contains commercial glucose through experiments.

Honey is the natural product produced by industrious bees made by using the nectar of flowering plants, the secretion of the living parts of the flowers or the secretion of the flower-sucking bugs living on the lively parts of the flowers. The bees pickup all these secreted materials and, change them by mixing them with their own special materials, reduce their water content and store all in the beehive until they become mature enough to be eaten. Fake honey is made directly using commercial glucose, molasses syrup, corn syrup, starch, flour, gelatine and jam just by adding honey perfume to give the odour and taste.

It is extremely difficult to differentiate it from the natural gas in respect to odour and taste. As a result, to be able to understand whether the sample we have is fake honey or not, we will use an experiment method.

MATERIALS AND METHODS

The materials used are; honey sample, diethyl ether, resorcin, concentrated HCl solution, (or Potassium iodine (KI) and iodine (I_2), a small beaker, test tube, pestle. The aim of the experiment is the determination of commercial glucose in honey is based on two principles qualitatively.

a) The monosaccarittes are partially broken due to the high temperature in the process of the production of glucose syrup added into honey. As a result of their extraction with ether, their forming colourful complex compound with resorcin is observed.

b) Due to the polysaccharides in the raw materials or the commercial glucose syrups, various colours changing from red to purple are observed with the iodine solution.

<u>1. Method:</u> Put about 3g honey into a pestle. Add 5mL ether and mix them. Wait 1-2 minutes. Take out the ether on the mixture into a dry beaker of 100mL. Repeat this process three times. Collect the excess ether into the same beaker each time. Evaporate the collected extract of ether in a water of 30-35^{o}C until about 5mL of extract is left.

Take the solution at the bottom of the beaker into the test tube. Add 2mL of resorcin and shake the tube strongly. One minute after the addition if resorcinol, there happens to be a cherry-like redness in the extract it shows the addition of commercial glucose or inverted sugar in the honey. If yellow or light pink colours mean no indication of fake honey.

Resorcinol solution 1ml: Dissolve 1g of freshly sublimated resorcinol in 100mL concentrated hydrochloric acid (37%). Mix the solution and store it in an air-free glass bottle. Prepare it newly before used. Shake diethyl (saturated with water) and 50-100 mL ether in a separate funnel with 20-30 mL water. After waiting for 20-30 minutes, take out the water at the bottom layer. This obtained ether saturated with water must be used maximum in one week.

2. Method: In the iodine solution (1g iodine and 1.4g KI in 50mL water) method mix 1mL honey with 1mL water in a clean tube. Add 4-5 drops of iodine solution, and shake the tube. Red and violet colours mean the honey is fake and that there is commercial inverted sugar in the honey.

HAZARDS

$(C_2H_5)_2O$ (Ether)

WASTE MANAGEMENT

The ether must be evaporated in a beaker put in hot water and the honey can be disposed in a rubbish bin.

PHOTOS

Fig. 1 4 g of honey is put into a pestle**.**

Fig. 2 Diethyl ether is dropped on the honey.

Fig. 3 The ether taken from the honey is heated.

Fig. 4 The resorcin is added.

Fig. 5 After the resorcin is added and 10 minutes have passed the result is observed. In this case it is not fake.

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THE INFLUENCE OF TEMPERATURE OF THE REACTION SPEED OF Mg WITH HCl

OVERVIEW

The goal of this experiment is establishing the influence of temperature on the reaction speed of Mg with HCl. We expect the speed of the reactions to increase when the temperature rises.

MATERIALS AND METHODS

- 2 Petri dishes
- HCl 1 M
- Thermometer
- Stand for test tubes
- Magnesium ribbon
- Scissors
- 2 measuring cylinders
- Ruler
- 2 test tubes
- Water bath of 80 degrees Celsius

Take 2 test tubes. Place 5 mL HCl in each test tube. Measure the 5 mL with a cylinder. Cut two 2 cm pieces of magnesium ribbon. Start with test tube 1. The temperature of the HCl in this test tube is 20°C. Put the 5 mL HCl of test tube 1 in a petri dish. Put test tube 2 with 5 mL HCl in the water bath of 80°C and wait a few minutes for the solution to heat up. Then put the HCl of test tube 2 in the other petri dish. Put a 2 cm pieces of magnesium ribbon in each petri dish at exactly the same time and measure the time it requires for the magnesium to fully react.

HAZARDS

HCl

Mg

CHEMICAL REACTION EQUATION

$\mathbf{2\,HCl_{(aq)} + Mg_{(s)} \longrightarrow MgCl_{2(aq)} + H_{2\,(g)}}$

CONCLUSIONS

When the temperature rises the speed of the reaction increases. At 20°C the speed of the reaction is much lower than the speed of the reaction at 80°C. You can also see that when the temperature is higher the magnesium ribbon is bubbling much more.

WASTE MANAGEMENT

Dilute the solution at the end of the reaction with water and put it in the waste bin of aqueous waste.

PHOTOS

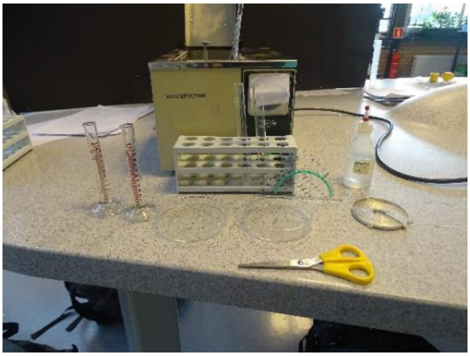

Fig. 1 All materials needed for the experiment.

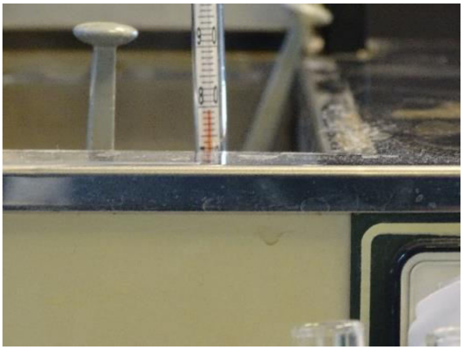

Fig. 2 The water bath of 80 degree Celsius.

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THE INFLUENCE OF CONCENTRATION ON THE REACTION SPEED OF HCl WITH Mg

OVERVIEW

The goal of the experiment is establish the influence of the concentration of HCl on its reaction speed with Mg.

MATERIALS AND METHODS

- 2 petri dishes
- Test tubes
- A magnesium ribbon
- 3 bottles of HCl in the following concentrations: 0.01 M, 1 M and 1.5 M.
- A pair of scissors
- A measuring cylinder

Take the petri dishes apart and put them in 2 pairs. Cut the magnesium ribbon into 4 parts of about 1.5 cm long. Pour 10 mL hydrochloric acid into two petri dishes, 0.01 M in one, 1 M in the other. Put one part of the magnesium ribbon in both dishes and compare the time it takes for both of the reactions to completely finish. Do this again, but this time compare 1 M with 1.5 M.

HAZARDS

HCl

H_2

Mg

CHEMICAL REACTION EQUATION

$$\mathbf{Mg_{(s)} + 2HCl_{(aq)} \longrightarrow H_{2\,(g)} + MgCl_{2\,(aq)}}$$

CONCLUSIONS

In the reaction, the hydrochloric acid has reacted with the magnesium to form hydrogen gas and magnesium chloride. A higher concentration of reactants, in this case HCl, results in a higher reaction speed.

WASTE MANAGEMENT

Add water to the Petri dishes and flush it down the sink.

PHOTO

Fig. 1 The materials needed for the experiment.

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THE INFLUENCE OF THE REAGENT ON THE REACTION SPEED

OVERVIEW

The purpose of the experiment is to show that using different materials like zinc and magnesium to react with hydrochloric acid will result in a different reaction speed.

MATERIALS AND METHODS

- Magnesium powder (0.3 g)
- Zinc powder (0.8 g)
- Hydrochloric acid (1 M)
- 2 petri dishes
- 2 crucibles
- 1 set of scales
- 2 test tubes

The molar amount should be equal for the zinc and magnesium powder, which is why you should take 0.3 grams of magnesium powder and 0.8 grams of zinc powder. Weigh the materials carefully and place them in the petri dishes. Pour 15 mL hydrochloric acid in 2 test tubes. The amount of hydrochloric acid in the test tubes has to be exactly the same. Put the magnesium in test tube 1 and the zinc in test tube 2. Let the hydrochloric acid react with the two other materials and watch the two petri dishes carefully to observe a difference in reaction speed. Measure the time it takes for both metals to completely react with the hydrochloric acid.

HAZARDS

Mg		**Zn**	
HCl		**$MgCl_2$**	
$ZnCl_2$		**H_2**	

CHEMICAL REACTION EQUATION

$$Mg_{(s)} + 2\,HCl_{(aq)} \rightarrow MgCl_{2\,(aq)} + H_{2\,(g)}$$

$$Zn_{(s)} + 2\,HCl_{(aq)} \rightarrow ZnCl_{2\,(aq)} + H_{2\,(g)}$$

CONCLUSIONS

After the two reactions took place you have seen that the hydrochloric acid had a much higher reaction speed with magnesium than with zinc. So different reagents have a different influence on the reaction speed.

WASTE MANAGEMENT

The reaction products are to be thrown away in an organic waste bin.

PHOTOS

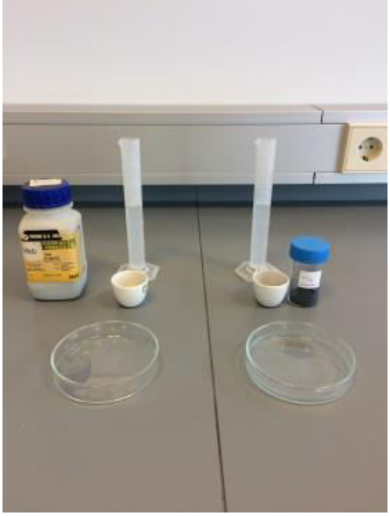

Fig. 1 In this picture you can see the start of the experiment.

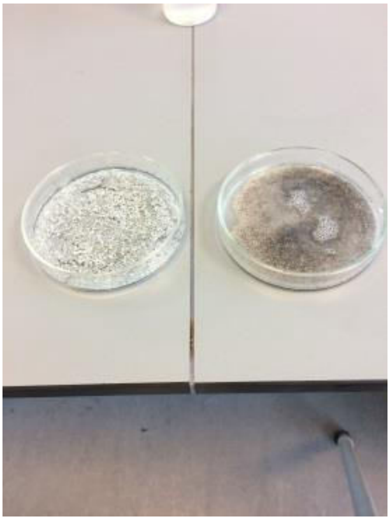

Fig. 2 In this picture you can you see the reaction products between hydrochloric and magnesium (right) and hydrochloric acid and zinc (left).

REFERENCES

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4. https://en.wikipedia.org/wiki/Zinc_chloride
5. https://en.wikipedia.org/wiki/Magnesium_chloride

THE INFLUENCE OF CATALYST ON THE REACTION SPEED H_2O_2 DECOMPOSITION

OVERVIEW

The goal of the experiment is to understand the influence on the reaction speed by using a catalyst (H_2O_2 with and without MnO_2).

MATERIALS AND METHODS

- Pipette
- Spatula
- 2 test tubes
- Stand for test tubes
- Measuring cylinder
- Petri dishes
- Hydrogen peroxide (H_2O_2, 3%)
- Manganese (IV) oxide (MnO_2)

Take a measuring cylinder of 10 mL and fill 2 test tubes with 8 mL H_2O_2 by using a pipette. Take 2 petri dishes and fill one of these with a spatula tip of MnO_2. Divide the Mno2 evenly over the petri dish. Take both test tubes and pour the H_2O_2 evenly over the two petri dishes at the same time. Observe the reaction.

HAZARDS

H_2O_2 ⚠

MnO_2 ⚠

CHEMICAL REACTION EQUATION

$$2\ H_2O_{2(aq)} \longrightarrow 2\ H_2O_{(aq)} + O_{2(g)}$$

CONCLUSIONS

When the MnO2 is used as a catalyst with H2O2, the decomposition reaction increases because the activation energy is lowered. When MnO2 is not used the decomposition reaction does not start. Therefor the use of a catalyst increases the reaction speed.

WASTE MANAGEMENT

Pour the reaction products in an inorganic barrel.

PHOTOS

Fig.1 All materials needed for the experiment.

Fig. 2 Result of the experiment.

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INFLUENCE OF THE GRANULATION ON THE REACTION RATE

OVERVIEW

Chemical kinetics is an important chapter in each chemistry book for it shows or demonstrates the speed of different processes/reactions happening around us. In a natural way, the role of different factors affecting the reaction rate is studied (see the previous experiments for a more detailed analysis of the factors affecting the reaction rate).

The present experiment, having its roots in the collision theory, shows how the granulation of the reagent is affecting the reaction rate. As a practical example marble chips, $CaCO_3$, having different granulations, are reacted with a 2M hydrochloric acid solution. As this reaction is producing a gas escaping from the system we can accurately measure the amount of CO_2 released during the experiment and plotting the mass of evolved CO_2 against time on a graph can give quantitative results concerning the influence of the granulation on the reaction rate. [1]

MATERIALS AND METHODS

Prepare 500 mL of a 2M solution of HCl and make available marble chips of different sizes. Measure 2.000 g of marble chips, using plastic weighing boats. Leave them on the plate of the analytical balance and place next to each other, on the same plate, a conical flask containing 40 mL of the acid solution. Press the ***tare*** button so your display will record 0.000 g. Have a chronometer available for the next step of the experiment.

With quick, but precise, movements pour the marble chips in the conical flask avoiding them to spread around the flask. Put the weighing boat on the plate and in the same time start the chronometer. Record at each 10 seconds the loss of the CO_2, recorded by the balance with negative mass. Measure at least 6-7 intervals of 10 seconds and draw a graph in which the mass loss is plotted against time. [1]

Repeat the experiment with the same quantities of acid and marble but using marble chips of different granulations. Do the experiment also with powdered $CaCO_3$. Record in each case the mass loss and plot it against time in the same graph.

HAZARDS

Hydrochloric acid is a strong, corrosive acid.

HCl

CHEMICAL REACTION EQUATION

$CaCO_{3(s)} + 2\ HCl_{(aq)} \dashrightarrow CaCl_{2(aq)} + CO_{2(g)} + H_2O_{(l)}$

CONCLUSIONS

The calcium carbonate when reacting with hydrochloric acid releases carbon dioxide. This is a gaseous compound and escapes from the system lightening the mixture, process observed on the scale of the analytical balance where the negative mass is giving a measure of the amount of gas evolved.

As the reaction involves a solid ($CaCO_3$) the reaction takes place at a faster rate when the solid is in a form of smaller granules or even powder rather than as lumps. [2] By measuring the mass of CO_2 evolved with time we can monitor the rate of the reaction. Plotting all masses on the same graph, for several experiments (see fig. 2), we can observe the influence of the surface area on the reaction rate. The gradient provides an indication on the reaction rate. The gradient of the graph is much steeper with powdered calcium carbonate because the surface area available for the reaction with hydrochloric acid is much greater than with lumps, allowing more collisions per second. [3] More collisions per second means more chances for the reagents to have the correct collision geometry and to surmount the energy barrier (the activation energy) leading to more reaction products.

You can plot also the volume of the CO_2 against time, the graph will be different but will have the same allure with the same rules: the steeper the curve the faster the reaction.

When performing the reaction it is important to use the same volume of the hydrochloric acid solution with the same concentration and the same amount of calcium carbonate with different granulations.

WASTE MANAGEMENT

The used chemicals are not dangerous for the environment although the acid is used 400% in excess. Thus the remaining solution will be acidic and it has to be neutralized before discarding it in the drain. Neutralize the solution either with milk of lime or with a solution of sodium carbonate/bicarbonate before discarding it in order to avoid corrosion of the sewage system.

PHOTOS

Fig. 1 Experimental setup for analyzing the influence of the surface area of solids on the reaction rate for the reaction between calcium carbonate chips $CaCO_3$ and hydrochloric acid, HCl.

Fig. 2 Mass loss after 5 seconds from the reaction starts when $CaCO_3$ was powdered. 0.759g CO_2 evolved

Fig. 3 Mass loss after 5 seconds from the reaction starts when $CaCO_3$ was 5 mm marble chips. 0.015g CO_2 evolved

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3. Harwood R.; Lodge I.; *Cambridge IGCSE Chemistry Coursebook, 4th edition;* Cambridge University Press, Cambridge; 2015, p. 183

AUTOCATALYTIC REACTION

OVERVIEW

A catalyst is a chemical compound capable to speed up a chemical reaction dramatically lowering its activation energy but is itself unchanged in the process. We witness every day chemical processes whose reaction rate increases as catalysts are added. One of most common example in people's mind when speaking about catalyst is "*catalytic filter*" or *"catalytic converter"* for petrol or diesel powered engines used to reduce the air pollution. Furthermore the great majority of biochemical processes undergoing in the living cell are catalyzed by different organic substrates such as *enzymes* acting as catalysts.[1]

The autocatalytic reaction is a great example showing how a substance can increase the reaction rate of a reaction. Moreover, due to the changes occurred in the catalyst itself by the reagents (and this is a special case of a catalytic reaction), is a perfect example to demonstrate how different substances have different catalytic actions on the same chemical process.

MATERIALS AND METHODS

Prepare a 10% solution of sodium potassium tartrate (Seignette salt) by weighing 13.428g of $NaKC_4H_4O_6*4H_2O$ and adding water until the solution reaches the 100g mass.

The second solution you need to perform the reaction is a 15% solution of oxygenated water or hydrogen peroxide (H_2O_2). Some authors advise the use of a 5% solution of hydrogen peroxide (H_2O_2) [2].

Put in a 250 mL beaker 10 mL of the Seignette solution and add 10 mL of oxygenated water. Observe the mixture and make notes of what you observe.

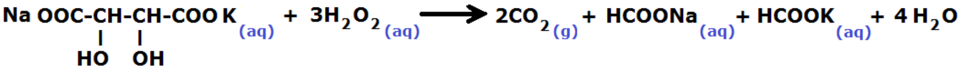
Repeat the experiment by heating first the two reagents to 70°C and mixing the two solutions, in the already specified amounts or you may increase the volumes to 45 mL each. Make notes on what you observe. Add in the beaker a few drops of a 5% $CoCl_2$ solution (obtained by weighing 1.832 g of $CoCl_2*6H_2O$ and adding water until the solution weights 20 g). Observe the changes; make associations of the solution's color and the intensity of the CO_2 evolvement.

HAZARDS

$NaKC_4H_4O_6$ No GHS symbols **$CoCl_2$**

H_2O_2

CHEMICAL REACTION EQUATION



$$Na\,OOC\text{-}\underset{HO}{\underset{|}{CH}}\text{-}\underset{OH}{\underset{|}{CH}}\text{-}COO\,K_{(aq)} + 3H_2O_{2\,(aq)} \longrightarrow 2CO_{2\,(g)} + HCOONa_{(aq)} + HCOOK_{(aq)} + 4\,H_2O$$

CONCLUSIONS

Oxygenated water oxidizes the potassium sodium tartrate (known both under the name of *Seignette salt* or *Rochelle salt* as Pierre Seignette was first to prepare it in 1675 in La Rochelle, France) [3]. The process cannot be seen when the two reagents are mixed at room temperature. Heated at 70°C the hydrogen peroxide doesn't acts either on the organic substrate and no evolvement of CO_2 is noticed.

The addition of the pink $CoCl_2$ in the beaker speeds up the reaction, a slight CO_2 evolvement being witnessed. This demonstrates the $CoCl_2$ (more specifically the Co^{2+} ions) act as catalyst for this process. The cobalt catalyst provides an alternative route for the reaction to occur. This alternative route has lower activation energy and the reaction proceeds much more quickly.

Due to the presence of the oxygenated water, the pink Co^{2+} ions are oxidized to green Co^{3+} ions whose catalytic action is higher. This is why, when the solution's color turns green, the CO_2 evolvement is the most abundant demonstrating the maximal capacity of Co^{3+} ions to speed up the decomposing reaction of the Seignette salt (Na and K tartrate). Eventually the green Co^{3+} is reduced to Co^{2+} the reaction site becoming pink again and the CO_2 generation cease. Further addition of the H_2O_2 solution will generate the same succession of steps if sodium potassium tartrate has remained in the beaker. [4]

The use of a large beaker is advisable (250 mL for around 30 mL of reagent solutions) because, during the reaction, abundant foam is generated which can cause spills on the reaction site.

WASTE MANAGEMENT

As the final solutions are containing Co^{2+} salts which are very toxic for the aquatic life and are harmful and irritants for humans it is advisable to discard the solutions in a container used for collecting toxic inorganic waste.

PHOTOS

Fig. 1 A 10% Seignette salt solution is treated with an equal volume of 15% H_2O_2. No reaction is observed even when hot.

Fig. 2 10 drops of $CoCl_2$ 5% are added turning the solution pink. If the solution is cold no reaction is observed.

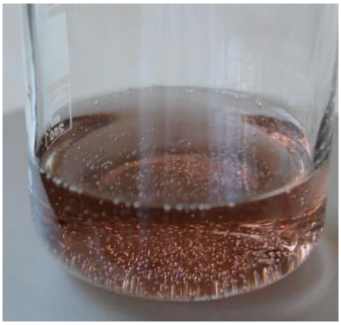


Fig. 3 Once the mixture is heated the oxidation reaction begins, catalyzed by Co^{2+}, and CO_2 evolvement is noticed.

Fig. 4 The most abundant CO_2 evolvement (highest reaction speed) is noticed when Co^{2+} is oxidized to Co^{3+}.

Fig. 5 After H_2O_2 is consumed the catalyst (Co^{3+}) returns to its initial state (Co^{2+}).

Fig. 6 Further addition of H_2O_2 repeat the oxidation process.

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ELECTRIC CONDUCTIVITY OF DIFFERENT MATERIALS

OVERVIEW

Different classes of materials (metals, salts or molecular substances) show different behaviour as far as electric conductivity is concerned. For a material to conduct electricity, two conditions have to be fulfilled: (i) charges have to be present, and (ii) the charged particles have to be mobile. The differences in behaviour can be explained by the way these classes of materials are composed.

Metals can be regarded as a lattice consisting of positive ions; the valence electrons moving freely through the metal. Solid salts consist of positive metal ions and negative non-metal ions or ion groups, arranged in a crystal lattice. In molecular substances partial separation of charges occurs in the case of polar bonds. The structure can be regarded as molecules being the individual building blocks. Transfer of charges between the molecules does not occur.

MATERIALS AND METHODS

Equipment

A beaker, two electrodes, an AC power source, an Ammeter.

It is recommended to take alternating current with a frequency of at least 1000 Hz in order to prevent electrochemical reactions to proceed at the electrodes.

Procedure

To start with, a piece of lead is connected to two electrodes to be integrated in an electrical circuit. Upon slowly increasing the voltage the ammeter reveals high conductivity. Then, the lead is melted in a beaker and now the molten lead is used to close the circuit. Again, appreciable conductivity is observed.

Secondly, the experiment is repeated with potassium nitrate. When the two electrodes are put into a layer of crystals no conductivity is measured. When melting however, the liquid appears to conduct electricity.

Finally, we conduct the experiment with water. Ice cubes do not conduct, nor does liquid water. The results are summarized in table 1.

HAZARDS [1,2]

KNO_3

Pb

CONCLUSIONS

Based on our assumptions we would expect metals to conduct electricity in the solid as well as in the liquid phase. In solid salts, the ions are fixed to their position in the crystal lattice so conductivity is not to be expected. In the liquid phase however they can move freely. Molecular substances are expected not to conduct electricity in both phases. The outcome of our experiment is in complete agreement with our assumptions.

WASTE MANAGEMENT

All materials can be used repeatedly so there is no waste.

PHOTOS

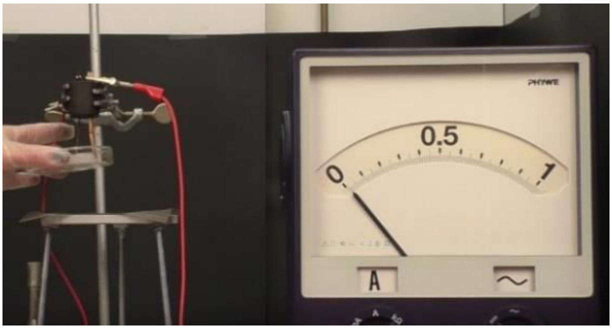

Fig 1. Ice, showing no conductivity.

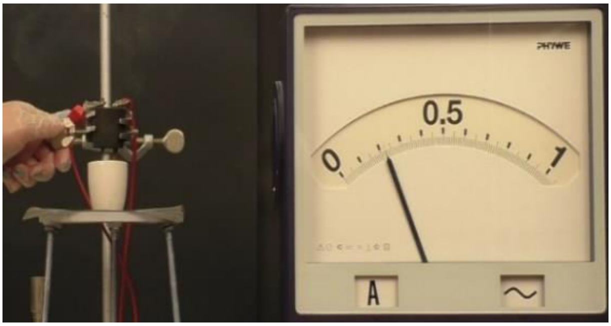

Fig 2. Liquid lead, allowing electrical current to pass.

Fig 3. Melting potassium nitrate.

Table 1.

Conductivity	liquid	solid
molecular	-	-
salt	+	-
metal	+	+

REFERENCES

1 https://en.wikiorg/wiki/Lead
2 https://en.wikiorg/wiki/potassium_nitrate

ELECTRIC CONDUCTIVITY OF DIFFERENT SALT SOLUTIONS

OVERVIEW

Solutions of salts are known to be conductors of electrical current because (i) upon dissolution the salt breaks up into charged particles (ions), and (ii) these ions can move freely in the solution. The conductivity of a salt solution will depend on (i) its concentration, (ii) the charge of the particles, and (iii) their mobility. In general, mobility of ions is comparable except for H^+ and OH^- ions.

In this experiment we will compare the conductivities of 0.1 M solutions of (i) kitchen salt, NaCl, (ii) soda, Na_2CO_3, (iii) KNO_3 and (iv) K_2SO_4. We expect the conductivities of solutions (i) and (iii) to be comparable because they both consist of singly charged ions. Solutions (ii) and (iv)contain doubly charged negative ions and twice as many positive ions as solutions (i) and (iii). Therefore, the conductivities of solution (ii) and (iv) are expected to be significantly higher.

MATERIALS AND METHODS

Requirements:

Four beakers, two electrodes, an AC power source, an ammeter, 0.1 M solutions of NaCl, Na_2CO_3, KNO_3 and K_2SO_4, a LED.

It is recommended to take alternating current with a frequency of at least 1000 Hz in order to prevent electrochemical reactions to proceed at the electrodes.

Procedure:

The two electrodes are placed in the beaker. Then a serial circuit is constructed with the power supply, the beaker, the ammeter (fig 2) and the LED. The salts are dissolved in water to yield 0.1 M solutions. The beaker is filled with the first solution to investigate and the voltage is adjusted such that a significant electric current is measured. The procedure is repeated for the three other salts, the voltage being the same in all four experiments. The resulting electric current is a direct measure of the conductivity. The LED emits more intensively as the conductivity is higher.

HAZARDS [1-4]

Na_2CO_3	⚠	**KNO_3**	⚠ 🔥
K_2SO_4	**none**	**NaCl**	**none**

CHEMICAL REACTION EQUATIONS

$NaCl_{(s)} \rightarrow Na^+_{(aq)} + Cl^-_{(aq)}$ **(1)**

$KNO_{3\,(s)} \rightarrow K^+_{(aq)} + NO_3^-{}_{(aq)}$ **(2)**

$Na_2CO_{3\,(s)} \rightarrow 2\ Na^+_{(aq)} + CO_3^{2-}{}_{(aq)}$ **(3)**

$K_2SO_{4\,(s)} \rightarrow 2\ K^+_{(aq)} + SO_4^{2-}{}_{(aq)}$ **(4)**

CONCLUSIONS

The electrical currents through solutions (i) and (iii) did not differ significantly. The current through solutions (ii) and (iv) appeared to be appreciably higher. This observation is in agreement with our expectations.

WASTE MANAGEMENT

The salt solutions do not contain any harmful components so they can be treated as waste water.

PHOTOS

Fig. 1 The solutions to be investigated.

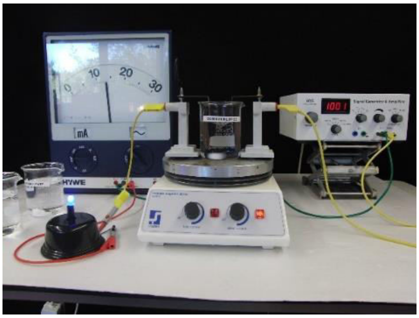

Fig 2. The equipment in action.

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3. http://en.wikipedia.org/wiki/potassium_nitrate;
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AUTOIONISATION OF WATER

OVERVIEW

Conductivity is the digital definition of a liquid solution to conduct electricity. In other words, conductivity is the electricity flow between to electrodes in a solution. The flow of electricity in a solution is obtained with the ion activity in the solution. Conductivity is a property to determine the purity of water. The less the conductivity of water is, the less the ions in it is. The factors providing all the electrical properties in water are the carbondioxide and mineral salt ions dissolved in water, their activity and heat. Most inorganic acids, bases and salt solutions are better conductors. However, the molecules of organic compounds are weak conductors due to their non-ionization of water solutions. The unit of conductivity is Siemens/cm ($S=Siemens=Ohm^{-1}$). The conductivity of pure water is almost o,055 µS/cm.

As a result of the experiments carried out, water is observed to conduct electricity if it is so slight. Water becomes ionized taking the form of on hydroxide anion and hydronium cation of two molecules of water. Through this, it conducts electricity.

The ionization of water simply can be shown as;

$$H_2O \rightleftarrows H^+ + OH^-$$

The ionization equation of water is;

$$H_2O + H_2O \rightleftarrows H_3O^+ + OH^-$$

The balance relation of the ionization of water;

$$\boldsymbol{K_{eq} = \frac{[H^+][OH^-]}{[H_2O]}}$$

K_{eq} is the balance whose digital value change upon only temperature H_2O is the molar concentration of water. The molar concentration of water is 55,6mol/L in 4^0C and this value is a constant value for 4^0C. In this context, $K_{eq} * [H_2O] = K_{eq} * 55,6$ products are also constant. This constant value is defined as the ionization of water, and shown as $\mathbf{K_w}$ symbol. As a result of this, the autoionization balance relation of water can be defined as;

$$\boldsymbol{K_w = [H^+].[OH^-]}$$

K_w value in 25^oC is $1.0X10^{-14}$

Temperature (°C)	K_w Values
0	$0.114 \cdot 10^{-14}$
10	$0.295 \cdot 10^{-14}$
20	$0.676 \cdot 10^{-14}$
25	$1.00 \cdot 10^{-14}$
60	$9.55 \cdot 10^{-14}$

The $[H^+]$ and $[OH^-]$ concentration in

$K_w = [H^+]\,[OH^-] = 1.0*10^{-14}$

$X^2 = 1.0*10^{-14}$
$X = 1.0*10^{-7}$

$[H^+]$ in distilled water at 21°C = $[OH^-] = 1.0*10^{-7}$ mol/L

MATERIALS AND METHODS

One of the beaker is filled with distilled water and the other is filled with tap water. Electrodes are dipped into the beaker with the tap waterand then these electrodes are connected to the power source with connecting cables. Between the two cables, one is connected to the power source while the other one is connected to one of the electrodes, a bulb is connected and it will light brightly. The same experiment is carried out using the beaker with with distilled water. This time, we will observe a dim, weak light in the bulb.

CHEMICAL REACTION EQUATION

$$H_2O_{(aq)} + H_2O_{(aq)} \rightleftarrows H_3O^+ + OH^-$$

CONCLUSIONS

Since there are mineral salt ions and carbondioxide solutions in tap water, severe conductivity of electricity is observed in the autoionization. Carbondioxide is found in water in three forms; dissolved gas (CO_2), ions (HCO_3^-, CO_3^{2-}) and carbonic acid (H_2CO_3). Conductivity has a direct proportion with the amount of ions in water. Depending on this, the ionic forms of carbondioxide in water and the ions of the mineral salts found dissolved in water will increase conductivty. Since pure water, including some important minerals in water such as sodium, calsium, iron is obtained by separating all the elements in it, its electricity conductivity is less than that of the tap water.

PHOTOS

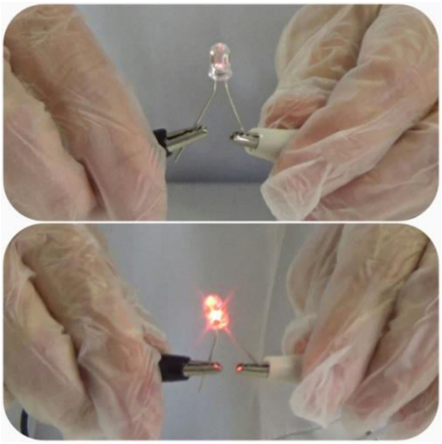

Fig. 1 Since there is no ion which will intensify the conductivity of electricity in water, the light is low. Since tap water contains all mineral salts, its electricity conductivity is high and the light is extremely bright.

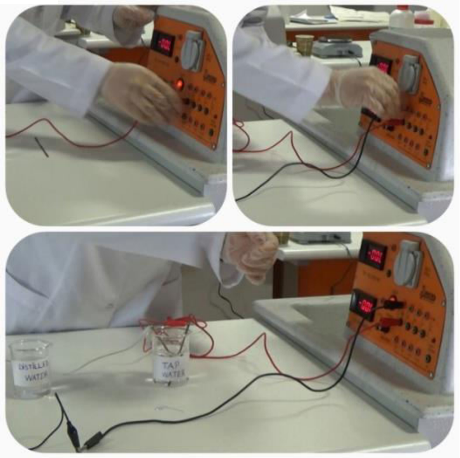

Fig. 2 A black crocodile cable is connected to the positive pole (+) of the power source and a red crocodile cable connected to the negative pole.

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ELECTROLYSIS OF WATER AND REACTION OXYGEN AND HYDROGEN

OVERVIEW

We can describe electrolysis in several ways.

- The realization of an electrical energy with the help of chemical reaction not forming on its own,that is, the transformation of electrical energy to chemical energy.
- Electrolyte is a distilled liquid or melted solution including cation and anion conducting electrical current. Electrolysis is the process by which ionic substances are decomposed into simpler substances when an electric current is passed through them.
- The decomposition process by which some substances are broken down through transforming the electrical energy to chemical energy.

Redox Reaction: The reactions in which gaining and losing go side by side are called redox reaction.

Electrode: The place where the chemical process takes place

Anode: The electrode which is attached to the (+) end of the electrical device

Cathode: The electrode which is attached to the (-) end of the electrical device.

Pure substance: Is a sample of matter with both definite and constant composition with distinct properties.

Electrolysis of water:
"Electrolysis of water" is the oxidation-reduction reaction of water that occurs when a voltage is applied to water. In this process, a reduction reaction occurs at the cathode (-) to generate hydrogen, and an oxidation reaction occurs at the anode (+) to generate oxygen.

In an electrolytic solution*, water (H2O) has separated into hydrogen ions (H+) and hydroxide ions (OH-).

$$2H_2O \rightarrow 2H^+ + 2OH^-$$

When a voltage is applied to an electrode in this electrolytic solution, electrons are supplied to the hydrogen ions at the cathode (-), causing reduction and generating hydrogen.

$$2H^+ + 2e^- \rightarrow H_2\uparrow$$

At the anode (+), the hydroxide ions are deprived of electrons, and oxidation occurs to generate oxygen and water.

$$2OH^- \rightarrow H_2O + 1/2O_2\uparrow + 2e^-$$

In the course of electrolysis the amount of material to change in anode and cathode is in direct proportion to the electrical current through the circuit.

Faraday Laws

- The rules that study the relation between electrical current and chemical change are known as Faraway Laws.
- Since an electron charge is $1.6x10^{-7}x6$, 02.1023 = 96500 coulomb that is 96500 coulomb charges (1 Faraday) is equivalent to 1 mole electron charge.
- The amount of matter (material) released in anode and cathode in an electrolysis circuit is in direct proportion with the electrical charge passing through the circuit.
- When a 1 Faraday charge passes through the circuit, one equivalent gram matter is released in anode and cathode.
- The gram amount of the matters released through electrolysis made in different electrolysis plates with equal amounts of electrical charge is equal.

The gram amount of the matters accumulated in the anodes or cathodes of the electrolysis plates are tied in series.

MATERIALS AND METHODS

Pour 50 mL water into a beaker. Prepare a solution, putting a little amount of sulphuric acid or 30 g of Na_2CO_3 (Sodium carbonate). Fill the two test tubes with this solution, by closing with a finger so as not to have air penetration and turn it upside down. Plunge it into the solution in the beaker. Fix the tubes with Bunsen pincers and place the tips of the electrodes into the tubes. Tie the crocodiles with the electrodes and the other tips to the direct current of the power supply. Upon the completion of gas accumulation inside the tubes, take them out without air penetration. Then a match is lit next to the blaze in which the burning gas is hydrogen and the gas making the blaze brighter being oxygen gas.

HAZARDS

Sulfuric acid **(H_2SO_4)**

CHEMICAL REACTION EQUATION

Cathode: $2\ H_2O + 2\ e \rightarrow H_{2(g)} + 2\ OH^-_{(aq)}$
Anode: $H_2O \rightarrow \frac{1}{2}\ O_{2(g)} + 2\ e + 2\ H^+_{(aq)}$

$3\ H_2O \rightarrow H_{2(g)} + \frac{1}{2}\ O_{2(g)} + 2\ OH^- + 2H^+_{(aq)}$

CONCLUSIONS

In the course of electrolysis process, water is oxidized in the anode (giving off electron) and it forms 0_2 gases, water is reduced in cathode (taking electron) and it forms H_2 gas. A few seconds later the hydrogen gas (H_2) accumulated in the tubes doubles the oxygen gas (O_2).

WASTE MANAGEMENT

At the end of the experiment, the thinned sulphuric acid solution is discharged into a sink.

PHOTOS

Fig. 1 The electrode fixed to the positive pole of the power device is anode whereas the negative pole is cathode.

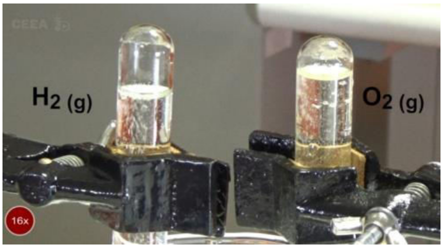


Fig. 2 While electrolysis hydrogen gas (H_2) accumulates in the tube fixed to the cathode (-), oxygen gas (O_2) accumulates in the anode (+).

Fig. 3 The mass of the collected hydrogen gas (H_2) in the test tube is double in mass of the oxygen gas (O_2).

REFERENCES

1. http://www.dersimiz.com/deneyler/Suyun-Elektrolizi-21696.html
2.http://yunus.hacettepe.edu.tr/~msahin06/fen_ve_teknoloji/mutlu_sahin_suyun_elektrik_enerjisi_ile_ayrismasi.pdf
3. http://www.panasonic.co.jp/ap/FC/en_construction_02.html

SOLUBILITY OF SALTS IN WATER - SOLUBILITY CHART

OVERVIEW

A solubility chart refers to a matrix where positive ions are listed vertically and negative ions horizontally. The matrix elements denote whether the corresponding combination of ions is soluble, sparingly soluble or insoluble. The chart to be completed in this experiment shows the solubility of various salts in water at atmospheric pressure and room temperature (293 K). Any element that reads "soluble" means complete dissolution, while "sparingly soluble" and "insoluble" denotes that solid salt remained.

MATERIALS AND METHODS

Requirements:

Spatula, test tubes, salts, demineralized water

Procedure:

We proceed with some simple experiments to fill the solubility chart with data. We will attempt to dissolve the following salts in water:

Copper carbonate	$CuCO_3$
Calcium hydroxide	$Ca(OH)_2$
Sodium sulphate	Na_2SO_4
Copper chloride	$CuCl_2$
Zinc nitrate	$Zn(NO_3)_2$
Calcium sulphate	$CaSO_4$
Ammonium acetate	NH_4CH_3COO
Iron(III) chloride	$FeCl_3$

We take a bit of the salt to be investigated on a spatula, put it in a test tube and add 2 mL of demineralized water. Start wagging and watch the content of the test tube to observe whether the salt dissolves. The results are qualified as follows:

Complete dissolution:	S	from soluble
No dissolution:	I	from insoluble
Slightly dissolved:	sS	from sparingly soluble

The results can be put into the solubility chart (see table 1).

HAZARDS [1-8]

$CuCO_3$	⚠	**$Zn(NO_3)_2$**	⚠ 🔥
$Ca(OH)_2$	🧪	**$CaSO_4$**	**none**
Na_2SO_4	**none**	**NH_4CH_3COO**	⚠
$CuCl_2$	⚠ 🌳	**$FeCl_3$**	⚠ 🧪

WASTE MANAGEMENT

The salts are harmless and used in such minimal quantities that they can be treated as waste water.

PHOTOS

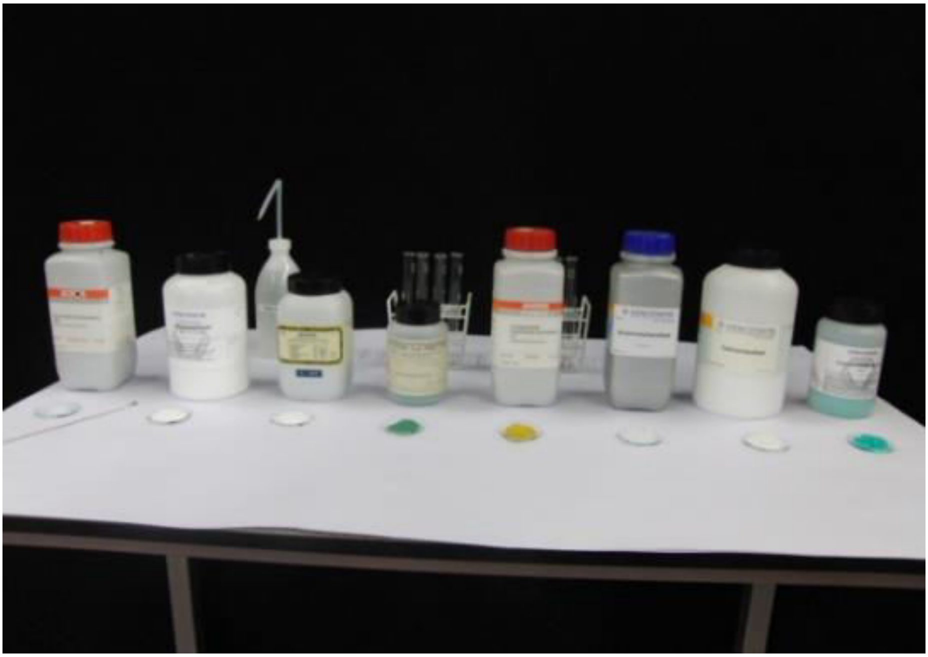


Fig. 1 The salts to investigate.

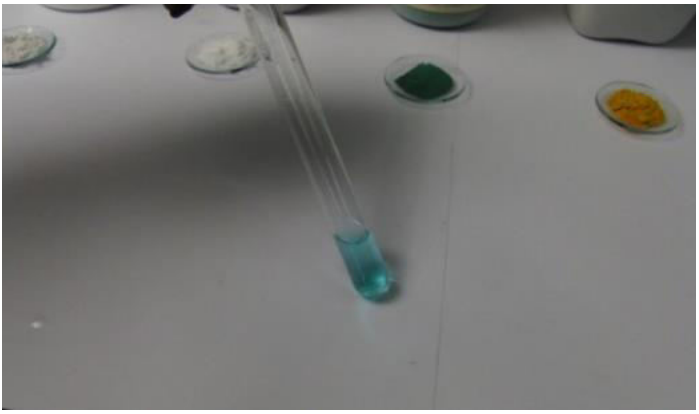


Fig. 2 Simply in a test tube…

Table 1.

	NO_3^-	CH_3COO^-	Cl^-	SO_4^{2-}	OH^-	PO_4^{3-}	CO_3^{2-}
Na^+				S			
K^+							
NH_4^+		S					
Cu^{2+}			S				I
Zn^{2+}	S						
Mg^{2+}							
Ca^{2+}				I	sS		
Ba^{2+}							
Pb^{2+}							
Fe^{3+}			S				
Ag^+							

REFERENCES

1. http://en.wikiorg/wiki/copper(II)_carbonate;
2. http://en.wikiorg/wiki/calcium_hydroxyde;
3. http://en.wikiorg/wiki/sodium_sulphate;
4. http://en.wikiorg/wiki/copper(II)_chloride;
5. http://en.wikiorg/wiki/zinc_nitrate;
6. http://en.wikiorg/wiki/calcium_sulphate;
7. http://en.wikiorg/wiki/ammonium_acetate;
8. http://en.wikiorg/wiki/iron(III)_chloride.

HOMOGENEOUS BALANCE OF NO_2 AND N_2O_4

OVERVIEW

Our students are made familiar with the phenomenon of chemical equilibrium in an early stage because knowledge of this topic is required to understand the behaviour of weak acids and bases. Furthermore, equilibrium reactions are important in chemical industry, e.g. the manufacture of synthesis gas and ammonia, and in biochemical processes.

In principle, every reaction is an equilibrium reaction. The value of the equilibrium constant *K* is determined by:

$$R.T\ ln\ K = -\ \Delta G_0 \qquad (1)$$

with R the universal gas constant, T the absolute temperature and ΔG_0 is the Gibbs free energy of reaction at standard conditions. The equilibrium constant depends on temperature only and is expressed in terms of activities which in the case of ideal mixtures reduce to molar fractions. For strongly negative values of ΔG_0 the equilibrium constant K is very high and the reaction is said to be complete.

For a reaction a A + b B $\leftarrow\rightarrow$ c C + d D we express the equilibrium constant *K* in terms of concentrations (moles/liter) according to:

$$K(T) = \frac{[C]^c[D]^d}{[A]^a[B]^b} \qquad (2)$$

Note that when c+d does not equal a+b, K is not dimensionless.

When a system in chemical equilibrium is disturbed, the system will counteract the disturbance as much as possible. So when a system in equilibrium is heated, the endothermic reaction is promoted in order to convert a part of the sensible heat added into chemical energy.

In this experiment we study the system NO_2/N_2O_4. NO_2 can be easily recognized because of its dark brown colour, while N_2O_4 is colourless. The dimerization reaction is exothermic with ΔH = -57 kJ / mole N_2O_4 [1]. So, we expect the NO_2 concentration to increase with increasing temperature.

MATERIALS AND METHODS

Requirements:

Sealed tubes with NO_2 and N_2O_4, ice cubes, kettle.

Procedure:

We have three sealed tubes with NO_2 and N_2O_4 to our disposal. All tubes are at room temperature and the colour is yellowish. One of them we put in melting ice, the second one in water at room temperature and the third one in hot water of about 80°C.

The colour of the tube in melting ice turns to colourless, indicating the formation of more N_2O_4. The tube in water at room temperature stays yellowish and the tube in hot water is dark brown. In the latter case, nearly all the gas molecules are present as NO_2. When the position of the three tubes is interchanged, their colours change immediately, indicating that the equilibrium adapts itself to the "new" conditions.

HAZARDS

NO_2, N_2O_4

CHEMICAL REACTION EQUATION

$2\ NO_2\ (g) \leftrightarrow N_2O_4\ (g)$

CONCLUSIONS

The observed changes in colour clearly indicate that the predicted shifts in equilibrium indeed occur.

WASTE MANAGEMENT

Does not apply.

PHOTO

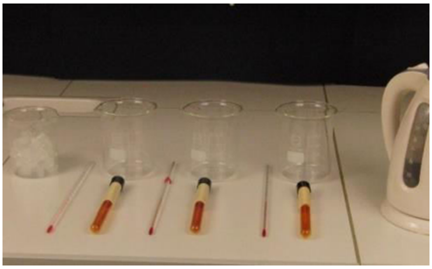


Fig. 1 At the start: Three identical tubes with NO_2 / N_2O_4.

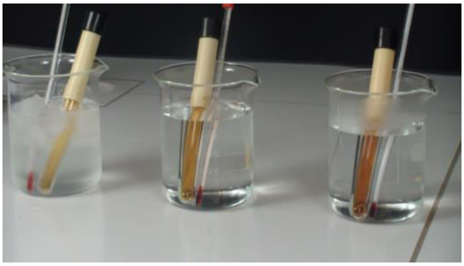


Fig. 2 Colours at different temperatures

REFERENCES

1. https://en.wikipedia.org/wiki/nitrogen_dioxide

ELECTRIC CONDUCTIVITY OF PURE AND DILUTED ACETIC ACID

OVERVIEW

In this experiment we will compare the electrical conductivity of pure, liquid acetic acid with an aqueous solution of acetic acid.

In pure acetic acid, protolysis occurs due to hydrogen bond formation according to:

$$2\ CH_3COOH \leftrightarrow CH_3COO^- + CH_3COOH_2^+ \qquad (1)$$

The equilibrium constant for reaction 1 is much lower than for the protolysis of water, so liquid acetic acid is expected not to conduct electricity at all.

As acetic acid is a weak acid in aqueous solution, dissociation will occur upon dissolution to some extent according to:

$$H_2O + CH_3COOH \leftrightarrow CH_3COO^- + H_3O^+ \qquad (2)$$

So the addition of water to the acetic acid will result in the formation of freely moving ions and conduction of electrical current is to be expected.

MATERIALS AND METHODS

Equipment:

A beaker, two electrodes, an AC power source, an ammeter.

It is recommended to take alternating current with a frequency of at least 1000 Hz in order to prevent electrochemical reactions to proceed at the electrodes.

Procedure:

The two electrodes are placed in the beaker. Then a serial circuit is constructed with the power supply, the beaker, the ammeter and the LED. The beaker is filled with pure acetic acid and submitted to electrical voltage. Then, some water is added to dilute the acetic acid and the procedure is repeated. The resulting electric current is a direct measure of the conductivity. The LED emits more intensively as the conductivity is higher.

HAZARDS

CH_3COOH

CHEMICAL REACTION EQUATIONS

$$2\ CH_3COOH \leftrightarrow CH_3COO^- + CH_3COOH_2^+$$

$$CH_3COOH + H_2O \leftrightarrow CH_3COO^- + H_3O^+$$

CONCLUSIONS

Pure acetic acid appears to be an insulator. Addition of water, being an insulator as well, yields a conducting liquid. This phenomenon was to be expected and can be explained by partial dissociation of acetic acid molecules into ions.

WASTE MANAGEMENT

Acetic acid is harmless and can be treated as waste water.

PHOTOS

Fig. 1 Pure acetic acid showing no conduction.

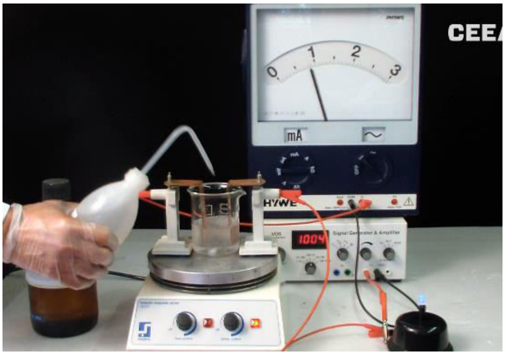


Fig. 2 Addition of water to pure acetic acid yields a conducting liquid.

REFERENCES

1. http://en.wikiorg/wiki/acetic_acid

ELECTRICAL CONDUCTIVITY OF HYDROCHLORIC vs. ACETIC ACID

OVERVIEW

The aim of this experiment is to experimentally reveal the dissociation behaviour of strong and weak acids. In this experiment we take hydrochloric acid and acetic acid as examples.

A strong acid will dissociate completely upon solution in water, whereas dissociation of weak acids is only fractional. The weaker the acid, the more molecules will be present in molecular form and hence the smaller the fraction of ionized particles will be.

The electrical conductivity of a substance depends on (i) presence of electrical charges; and (ii) their ability to migrate. For aqueous solutions of electrolytes both conditions are undoubtedly fulfilled.

For equimolar solutions of electrolytes, as in our case, the number of charge particles per unit volume and hence conductivity will depend on the degree of dissociation. As hydrochloric acid is known to be a strong acid and acetic acid to be weak, we expect the conductivity of the solutions to be substantially different.

MATERIALS AND METHODS

Requirements:

A power supply, two electrodes, a LED, a 500 mL beaker, a conductivity meter, demineralized water, tap water, mineral water, solutions of 0.1 M acetic acid, 1.0 M acetic acid, 0.1 M HCl and 1.0 M HCl.

Procedure:

After filling the beaker with the first solution to investigate, it has to be stirred in order to ascertain a homogeneous distribution of the concentration. The conductivity can be directly measured with the conductivity meter. The conductivity can be visualised with an electrical circuit with a LED. First we measured the conductivities of demineralized water, tap water and mineral water as references. Then we measured the conductivities of 0.1 M Acetic acid, 1.0 M acetic acid, 0.1 M hydrochloric acid and 1.0 M hydrochloric acid resp.

HAZARDS [1,2]

HCl

CH_3COOH

CONCLUSIONS

The outcome of our experiment was, in order of increasing conductivity:

demineralized water < tap- & mineral water < 0.1 M acetic acid < 1.0 M acetic acid < 0.1 M HCl < 1.0 M HCl

as was to be expected. The conductivity of 0.1 M acetic acid turned out to be only slightly higher than the conductivity of tap water and mineral water.

Comparison of equimolar solutions of hydrochloric acid and acetic acid indicates that the conductivity of acetic acid is much less than that of hydrochloric acid. This result could be expected because hydrochloric acid is a strong acid, which means that upon dissolution it breaks up completely into ions. Acetic acid however is a weak acid and remains molecules for the major part.

WASTE MANAGEMENT

All chemicals involved can be treated as waste water.

PHOTOS

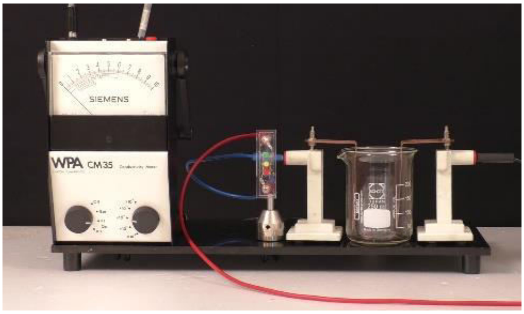


Fig. 1 At the start of the experiment.

Fig. 2 Conductivity meter indicating electron flow.

REFERENCES

1. http://en.wikiorg/wiki/hydrochloric_acid;
2. http://en.wikiorg/wiki/acetic_acid

DANIELL CELL

OVERVIEW

Redox reactions are characterized by transfer of one or more electrons. As long as both half reactions proceed in the same solution, the electron flow cannot be used to convert its energy into a different form of energy. When the half reactions proceed in separate reactors however the electrons can be forced to flow through a wire.

Daniell was the first to imply electrochemical phenomena into a working electrochemical cell. It was known that a rod of zinc placed in a solution of zinc sulphate becomes negatively charged because zinc atoms tend to dissolve, leaving their two valence electrons behind in the rod. Reversely, copper(II) ions in a solution of copper(II)sulphate tend to precipitate on a copper rod, resulting in a positively charged copper rod. When the two rods are connected with a wire and the circuit is completed with a salt bridge we have a cell that is capable of delivering electrical energy until one of the reactants is exhausted.

The Daniell cell can be regarded as the precursor of our batteries.

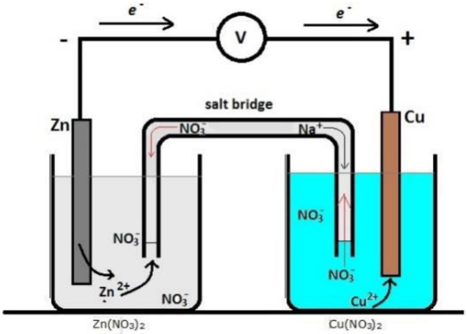
MATERIALS AND METHODS

Requirements:

Two beakers, copper and zinc electrodes, salt bridge (Na^+/NO_3^- in agar [3]), wires, voltmeter, copper(II)- and zinc solutions.

Procedure:

A schematic experimental set-up is shown in the graph on the right. In our half cells 0.1 M Cu(II)- and Zn-nitrate solutions were used. The standard potentials predict the voltage to be 1.1 V.



HAZARDS [1,2]

$Zn(NO_3)_2$

$Cu(NO_3)_2$

CHEMICAL REACTION EQUATIONS

$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ $E_0 = +0{,}34$ V (1)

$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$ $E_0 = -0{,}76$ V (2)

CONCLUSIONS

Our version of the Daniell element has shown to be capable of delivering electrical energy; the voltage however being less than is to be expected from equilibrium data. The standard potentials predict the voltage to be 1.1 V, our cell delivered 0.7 V. The difference can be explained by the fact that the reported E_0-values are only valid at equilibrium conditions and at concentrations of 1.0 M. Now we are dealing with a dynamic system and lower concentrations; reactions are irreversible so energy losses are likely to occur.

WASTE MANAGEMENT

Solutions of copper(II)- and zinc nitrate should be collected and further processed by a recognized waste processing company.

PHOTOS

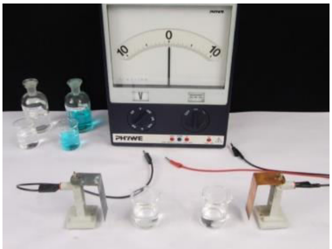


Fig. 1 At the start of the experiment.

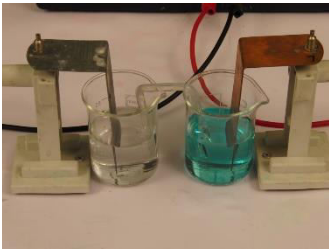


Fig. 2 Daniell cell.

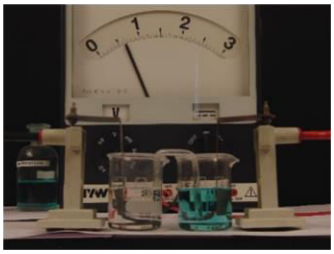


Fig. 3 Daniell cell in action.

REFERENCES

1. https://en.wikipedia.org/wiki/Copper(II)_nitrate
2. https://en.wikipedia.org/wiki/Zinc_nitrate
3. https://en.wikipedia.org/wiki/Salt bridge

LEMON BATTERY

OVERVIEW

Lemons are known to be able to generate electrical energy. When two pieces of metal are inserted in a lemon a potential difference arises between the metals. Arranging a number of lemon batteries in series produces a high enough voltage to supply a led with sufficient energy.

Zinc and copper are widely used as electrode materials. Zinc atoms are oxidized; the zinc ions dissolve in the electrolyte. The electrons flow through the wire to the anode where dissolved oxygen is reduced. The anode itself does not react.

Throughout this experiment we used different combinations of electrode materials and different fruit as well.

MATERIALS AND METHODS

Equipment

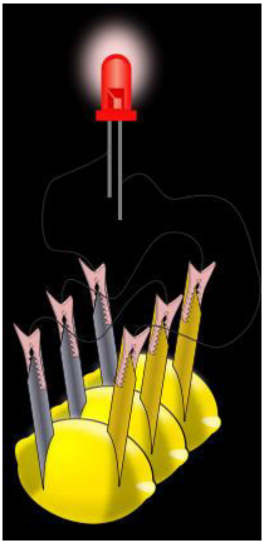
- wires;
- Voltmeter;

Materials

- pieces of cathode materials;
- pieces of anode materials;
- different sorts of fruits.

Procedure

Insert the pieces of cathode and anode material at a certain distance into the fruit. Then connect them with a Volt-meter; cathode on the negative side and anode on the positive side.



HAZARDS

Citric acid 

CHEMICAL REACTION EQUATIONS

Cathode: $Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2\ e^-$

Anode: $O_{2(aq)} + 4\ H^+_{(aq)} + 4\ e^- \rightarrow 2\ H_2O_{(l)}$

CONCLUSIONS

A wide variety of fruits and electrode materials can be used to gain electrical energy from fruit. A lemon supplied with a zinc and a copper electrode however appeared to be the most powerful source.

WASTE MANAGEMENT

There is no waste to handle.

PHOTOS

Fig.1 Our materials at the start.

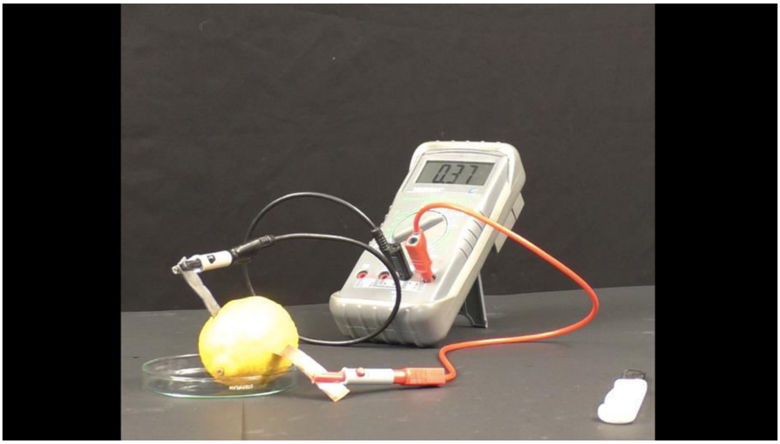


Fig. 2 A single lemon, Cu / Pb

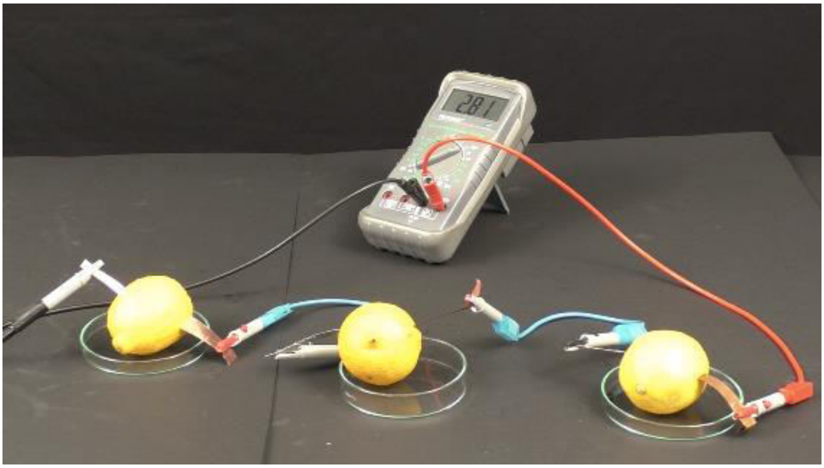


Fig. 3 Three lemons in series, Cu / Zn

Table 1.

Materials	Potential difference (V)
Cu / Zn	0.92
Cu / Pb	0.37
Zn / Pb	0.49
Brass / Zn	0.09
Fe / Zn	0.07
Cu / Fe	0.93
Fruits, 3 in series	**Potential difference (V)**
Lemons	2.81
Potatoes	2.71
Dragon fruits	2.70
Bananas	2.66
Red potatoes	2.68
Oranges	2.80
Dragon/orange/banana	2.85

Potential differences for (i) combinations of electrode materials, measured with one lemon, and (ii) for different kinds of fruits, 3 in series, all with Cu/Zn.

REFERENCES

1. https://en.wikipedia.org/wiki/Lemon_battery

THE APPLE BATTERY

OVERVIEW

The aim of this experiment is to make electrical current with the use of an apple. This experiment also displays how electricity is generated by using iron, copper and an acid.

MATERIALS AND METHODS

- One apple
- Flat piece of copper
- Flat piece of zinc
- Multimeter
- Two wires
- Knife
- Two alligator clips

Start by cutting the apple in half. After that, stick the flat piece of zinc in one half of the apple and stick the flat piece of copper on the other side of the same piece of apple. Attach the clips on each end of the two wires. Connect the wires to the multimeter correctly to measure in volts (DC). Attach the wires with the clips to the piece of iron and the piece of copper. Lastly you can switch on the multimeter and measure the amount of volt created.

HAZARDS

None

CHEMICAL REACTION EQUATION

$\mathbf{Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-}$

$\mathbf{2H^+_{(aq)} + 2e^- \rightarrow H_{2\,(aq)}}$

CONCLUSIONS

It is possible to create electrical current with the use of an apple, a piece of flat zinc and a piece of flat copper. Zinc reacts as a reducing agent, giving away electrons to flow to the copper electrode, where they react with free hydrogen ions in the acidic solution in the apple.

WASTE MANAGEMENT

There is no special waste management. You can throw the apple away in the trash can and recycle the zinc and copper plates.

PHOTOS

Fig. 1 In this picture you see the materials needed for this experiment**.**

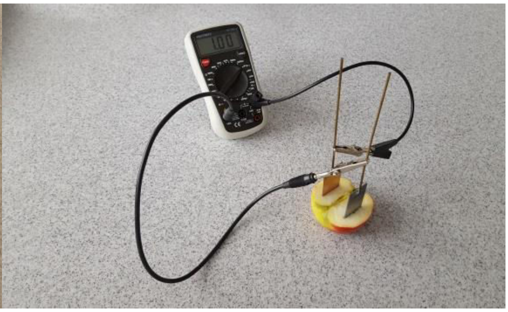


Fig. 2 In this picture you see the set-up for the experiment. You see how the apple is correctly connected to the multimeter.

REFERENCES

1.http://www.wetenschapsite.be/
2.http://how-things-work-science-projects.com/
3.https://en.wikipedia.org/wiki/Lemon_battery

GALVANIC CELL WITH DIFFERENT VEGETABLES

OVERVIEW

A cell a battery is a device ased to store and the chemical energy in science and technology transform, into electrical form. Batteries(cells) are sources of direct current. The essence of all batteries (cells) is the same. That is, they consist of two different electrodes sonted into on electrolyte. Due to the reduction and oxidation happening in the battery there exists a potential difference between the (+) and (-) poles. There exists a lower potential in the excessind, electron loaded negative pole and a higher potential in the scarcely electron loaded positive pole. When these two poles are attached to each other withen conductive wire, then the electrons move from then negative pole to the positive one.

Out of fruits containing acid it is possible to make simple batteries cells. Due to their transforming the chemical energy to electrical energy, they are called voltaic batteries (cells). Cells are made by placing two different metals into an acidic solution. The juice of the fruit used consists of the necessory acid and thus ensured the production of electricity. The basic principle is that in the structure of these fruit and vegetables there are electrons and acid that produce electrical current. The electrical current obtained is so weak. İt gives a very slight light in the LED. What makes electricity possible is the mobility of electrones. A cell has a positive and a negative pole. The electrons move from the negative pole to the positive pole.

In fact, the facturs that make energy possible is the acid, namely the H^+ ion in the fruit. Acid is a conductive fluid. The zinc atoms dissolve in the solution and become a (+) positively loaded ions. İn this sense the anote pole (the zinc plate), gain (+) positive load and the copper plate (the cathode pole) gain (-) negative load. The positive loads in the zinc electrode, with the help of a conductor, upon reaching the copper electrode, the hydrogen ions coming from the solution become neutralized and accumulate on the elecrtode as a gas.

The process taking place in the cathode copper electrode is;

$$2H^+ + 2e^- \rightarrow H_2$$

İn time, when the hydrogen gas around copper electrode multiplies, then the voltaic cell doesn’t produce any current.

MATERIALS AND METHODS

Using an emery clean the tips of the copper wires that will contact the copper and zinc plates. Fix a copper plate and a zinc plate an each of the oranges reciprocally. Connect the copper an zinc plates fized on the oranges as **Cu-Zn-Cu-Zn-Cu-Zn-Cu-Zn** using a copper wire. Connect the open ends of the copper wires to the LED lamp as anode on Zinc plate, as cathode on copper plate observe the light in the lamp. By using this same method we can observe the same situation by using tomatoes.

CONCLUSIONS

When a zinc plate is placed to one side of on orange and a copper plate to the other side the hydrogen ions(H+in the orange) help the flow of electricity, thus functioning as a cell. Then the LED lamp, attached to the circuit, gives light.

PHOTOS

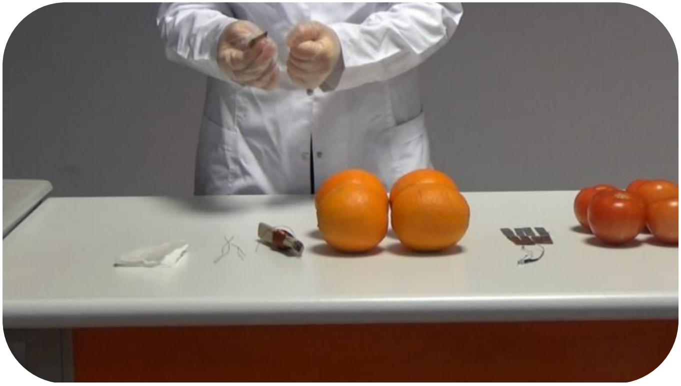


Fig. 1 If the copper wire is not sand papered, it will not transmit the electricity.

Fig. 2 The plates must be put in the oranges in Cu-Zn-Cu-Zn order and must be connected with copper wire to each other.

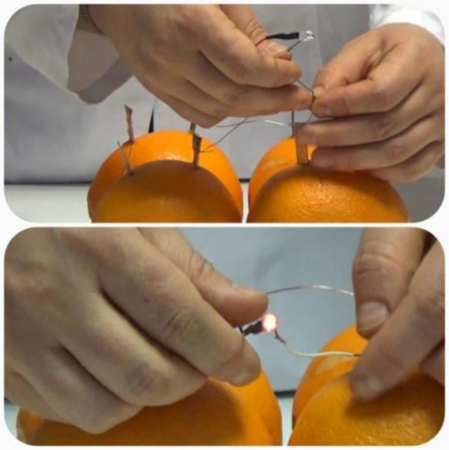


Fig. 3 Connect the open ends of the copper wires to the LED lamp as anode on Zinc plate, as cathode on copper plate observe the light in the lamp.

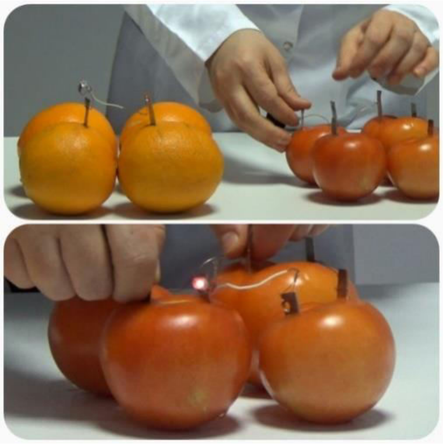


Fig. 4 When the copper and zinc plates are wiped they are put in the tomatoes the same way and it can be observed that the LED bulb lights.

REFERENCES

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2.http://sorumakinasi.blogcu.com/uretecler-elektrik-ureten-kaynaklar/5371597
3.http://yemek.com/yiyeceklerden-elektrik-uretmek/#.WH05iFWLTDe

THE HEAT OF COMBUSTION OF PEANUTS

OVERVIEW

From the beginning on, our students are being taught that chemical reactions exhibit heat effects. When energy is released we refer to the reaction as exothermic. When energy is consumed, the reaction is endothermic.

Combustion reactions in general are exothermic and without their energy release life as we know it would not exist. We burn natural gas for cooking purposes, and we use several kinds of fossil fuels to convert chemical energy into kinetic energy. These processes all have in common that they proceed at elevated temperatures and exhibit visible effects like flames or explosions.

Even in our own bodies energy is supplied by combustion of food. Fats, hydrocarbons and proteins all can react with oxygen to yield combustion products and energy. Of course these processes are much different from the reactions mentioned in the previous section: needless to say there are no flames in our cells and reaction temperatures are restricted to 36 – 37 ^{0}C. The process consists of a complicated series of redox reactions being catalysed by enzymes. However, the energy contents involved (to be more correct: enthalpies) are state functions which means that the energy effect is determined by the final and initial states only and not on the path between them.

In this experiment we will demonstrate the enormous heat content of peanuts. They consist approximately of 30 m% fat, 50 m% proteins and 18% carbohydrates. With a heat of combustions of 6 kcal / gram [1] peanuts are real fatteners. According to the definition of the calorie, 1 gram of peanuts releases enough energy to let 6 liters of water rise one degree in temperature.

MATERIALS AND METHOD

Equipment

- Erlenmeyer flask 200 mL;
- tweezers;
- tea light;
- thermometer;
- balance.

Procedure

We fill the Erlenmeyer flask with 20 mL of water and put a thermometer in it. Weigh the peanut, hold it with the tweezers and ignite it in the flame of a tea light. Now hold the burning peanut as close as possible under the bottom of the Erlenmeyer flask. When the peanut is burnt completely, a substantial rise of the water temperature is observed. From the weight of the peanut and the mentioned value of the combustion energy it can be estimated what fraction of the released heat is lost to the environment.

HAZARDS

CHEMICAL REACTION EQUATIONS

Fats:

$$\mathbf{CnHmCOOH_{(s)} + (n + \tfrac{1}{4}(m+1))\ O_{2(g)} \rightarrow (n+1)\ CO_{2(g)} + \tfrac{1}{2}\ (m+1)\ H_2O_{(l)}}$$

Carbohydrates:

$$\mathbf{(C_6H_{10}O_5)_{n(s)} + 3\ \tfrac{1}{2}\ n\ O_{2(g)} \rightarrow 6n\ CO_{2(g)} + 5n\ H_2O_{(l)}}$$

Proteïns:

$$\mathbf{C_nH_mN_pO_qS_{r(s)} + (n + \tfrac{1}{4}(m+1) + p.x + r - q)O_{2(g)} \rightarrow n\ CO_{2(g)} + \tfrac{1}{2}(m+1)\ H_2O_{(l)} + p\ NO_{x(g)} + r\ SO_{2(g)}}$$

CONCLUSIONS

Although no quantitative conclusions can be drawn from this simple experiment, the substantial temperature rise of the water indicates that a huge amount of energy is released upon combustion of a peanut.

WASTE MANAGEMENT

The solid remaining after combustion does not contain hazardous components so it can be treated as normal waste.

PHOTOS

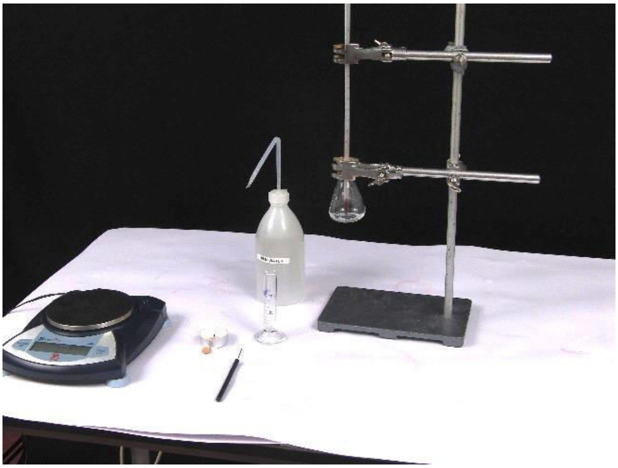

Fig. 1 Ready for take-off…

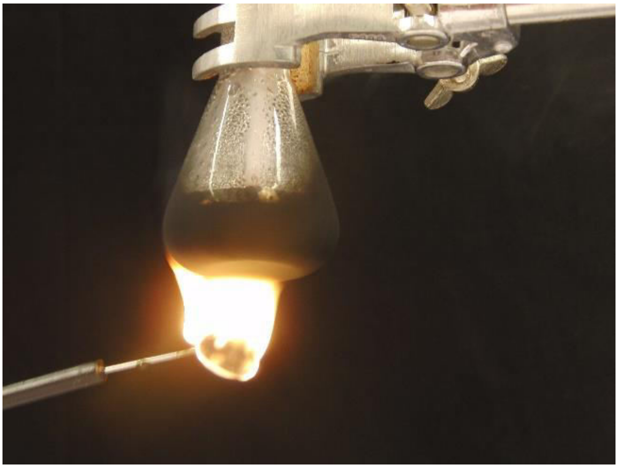

Fig. 2 Notice the enormous flame rising from one peanut

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EXTRACTING OILS WITH ACETONE

OVERVIEW

The aim of this experiment is to extract oil from peanuts by using acetone. During this experiment acetone will be used as a dissolvent to extract the oil. After the extraction, the acetone will be evaporated to obtain pure and natural oil. This is a simplified way of extracting pure oil from peanuts.

MATERIALS AND METHODS

- 5g Peanuts
- Mortar
- Petri dish
- Acetone 50 ml
- Funnel
- Filter disc
- Scales
- Erlenmeyer flask
- Spatula
- 2 Beakers

Measure 5 grams of peanuts using scales and mash these peanuts using the mortar. Make sure that the peanuts are mashed finely. Transfer the mashed peanuts into a beaker. Add 50 mL acetone to the beaker with the mashed peanuts. Leave the peanuts to soak in the acetone for an hour. Fold a filter disc in a cone shape. Filter the peanut-acetone mixture by using the filter-cone, funnel and the Erlenmeyer flask. Pour the filtrate in a petri dish. Put the petri dish on the heating, so the acetone can evaporate. Leave the petri dish until all the acetone has evaporated. The residue in the dish is peanut oil.

HAZARDS

Acetone

CHEMICAL REACTION EQUATION

No chemical reaction occurred.

CONCLUSIONS

Acetone works as a dissolvent for peanut oil, and therefore extracts oil from peanuts. By doing so, after filtration you can evaporate the acetone from the extracted mixture to form purified peanut oil.

WASTE MANAGEMENT

The acetone should be disposed in the inorganic waste bin and the peanuts can be disposed in a regular bin.

PHOTOS

Fig. 1 Peanuts after mashing.

Fig. 2 Mixture after being soaked for an hour.

REFERENCES

https://en.wikipedia.org/wiki/Peanut_oil

MEASURING THE RATE OF RUSTING

OVERVIEW

Rusting is the process in which iron turns into iron oxide. It happens when iron comes into contact with water and oxygen. The process is a type of corrosion that occurs easily under natural conditions. The goal of this experiment is to cause rusting artificially and to measure the rate of rusting when exposed to different amounts of oxygen. With the help of this experiment we will find out if the amount of oxygen iron is exposed to can affect the reaction speed.

MATERIALS AND METHODS

- Erlenmeyer flask
- Test-tube
- Burner
- Match sticks
- Stopwatch
- 2 Porcelain dishes
- Crucible tongs
- Nipper
- Iron(II)oxalate ($FeC_2O_4 \bullet 2H_2O$)
- Spatula
- Standard procedure materials (e.g. lab coat, latex-gloves, glasses etc.)
- A scale

Prepare the experiment by gathering all the necessary items. Light the burner and attach it horizontally to the stand. Pour 2.0 grams of iron(II)oxalate into both the test-tube and the Erlenmeyer flask. Change the colour of the flame to a roaring blue flame. Be careful not to get too close to the flame. Place the test-tube above the flame with a crucible tong and immediately start timing the amount of time it took for the iron(II)oxalate to turn red due to the heat. After the colour changes to red, write down the time on the timer, place the reaction product into the petri dish and spread it out on the surface. Now do the same procedure but use the Erlenmeyer flask instead of the test-tube.

HAZARDS

$FeC_2O_4 \bullet 2H_2O$

CHEMICAL REACTION EQUATION

$$FeC_2O_{4(s)} \rightarrow Fe_{(s)} + 2\ CO_{2(g)}$$

$$4\ Fe_{(s)} + 3\ O_{2(g)} \rightarrow 2\ Fe_2O_{3(s)}$$

CONCLUSION

When pure iron is exposed to oxygen from the air it starts to undergo a reaction in which iron oxide is formed. The amount of oxygen the iron is exposed to has an effect on the reaction speed. It is concluded after the experiment that if iron is exposed to more oxygen the reaction speed goes up, which means that the rust forming is faster.

WASTE MANAGEMENT

Put the ironoxide in the solid waste bin.

PHOTOS

Fig. 1 These are all the materials you need for this experiment.

Fig. 2 Formed ironoxide

REFERENCES

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DNA ISOLATION OF WHEAT GERM

OVERVIEW

DNA was first observed by a German biochemist named Frederich Minescher in 1869. But for many years, researchers did not realize the importance of this molecule. It was not until 1953 that James Watson, Francis Crick, Maurice Wilkins and Rosalind Franklin figured out the structure of DNA- a double helix- which they realized could carry biological information. Watson, Crick and Wilkins were awarded the Nobel Prize in Medicine in 1962.

DNA is made up of molecules called nucleotides. Each nucleotide contains a phosphate group, a sugar group and a nitrogen base. The four types of nitrogen bases are adenine (A), thymine (T), guanine (G) and cytosine (C). The order of these bases is what determines DNA's instructions, or genetic code. Similar to the way the order of letters in the alphabet can be used to form a word, the order of nitrogen bases in a DNA sequence forms genes, which in the language of the cell, tells cells how to make proteins. Another type of nucleic acid, ribonucleic acid, or RNA, translates genetic information from DNA into proteins.

MATERIALS AND METHODS

13. The experiment is as follows;
14. 1 g of wheat is put in 50mL test tube.
15. Add 20mL hot water and stir for 3 minutes.
16. By adding 2mL of dish washer detergent shake it for 5 minutes every minute. Try not to make bubbles.
17. By tilting the tube put about 15mL of ethyl alcohol. The alcohol will stay on the surface.
18. After alcohol is added wait for 15 minutes.
19. The DNA that is between the alcohol and the compound is taken by the help of a pipette and put into another tube.

HAZARDS

Detergent		**C_2H_6O (Ethyl Alcohol)**	

PHOTOS

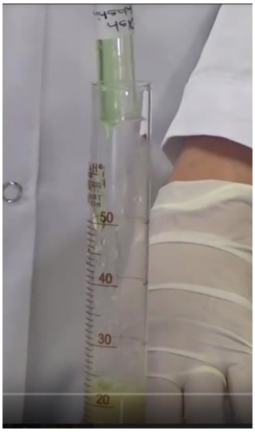

Fig. 1 The dish washing detergent is added.

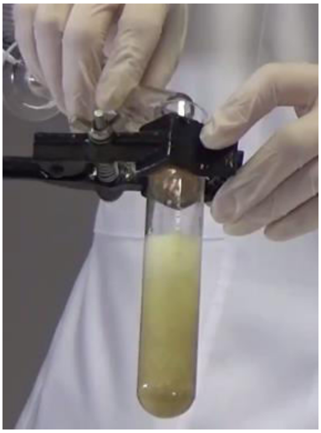

Fig. 2 Ethyl alcohol is added.

Fig. 3 The DNA of the wheat germ in the solution is observed.

Fig. 4 The DNA out of the test tube.

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DNA ISOLATION FROM ONION

OVERVIEW

DNA was first observed by a German biochemist named Frederich Minescher in 1869. But for many years, researchers did not realize the importance of this molecule. It was not until 1953 that James Watson, Francis Crick, Maurice Wilkins and Rosalind Franklin figured out the structure of DNA- a double helix- which they realized could carry biological information. Watson, Crick and Wilkins were awarded the Nobel Prize in Medicine in 1962.

DNA is made up of molecules called nucleotides. Each nucleotide contains a phosphate group, a sugar group and a nitrogen base. The four types of nitrogen bases are adenine (A), thymine (T), guanine (G) and cytosine (C). The order of these bases is what determines DNA’s instructions, or genetic code. Similar to the way the order of letters in the alphabet can be used to form a word, the order of nitrogen bases in a DNA sequence forms genes, which in the language of the cell, tells cells how to make proteins. Another type of nucleic acid, ribonucleic acid, or RNA, translates genetic information from DNA into proteins.

MATERIALS AND METHODS

<u>Preparation of solution A:</u>

20g salt (iodine free, plain salt), 20mL liquid detergent, +180mL distilled water, so the final volume being 200mL. The other materials are; salt, liquid detergent, distilled water, absolute ethyl alcohol (99.5%), filter paper, diced onion, glass tubes of 10mL each and a blender.

20. Take out 4 cubes a side of which is 4cm, out of the inner part of an onion.
21. Add solution A that was prepared before, and put the onion pieces into a blender, add 100mL of solution A, prepared before, into it.
22. Break the pieces in the blender for 1 minute.
23. Filter the mixture, using a coffee filter machine.
24. Add 30mL pineapple juice into it. Mix and stir it slowly.
25. Take out 6mL of the mixture and put it into the test tube. Add 6mL of cold absolute (99.5%) ethyl alcohol. Observe two phases.
26. Fix the tube and wait 4-5 minutes (until the release of bubbles ends).
27. Observe the onion DNA on the upper phase.

HAZARDS

Ethyl alcohol (C_2H_6O) 🔥

PHOTOS

Fig. 1 When preparing solution A, we add 20g of iodine free salt into distilled water.

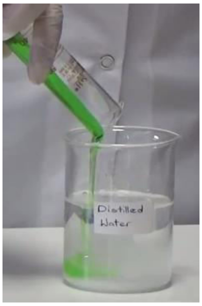

Fig. 2 20mL liquid detergent is added to the solution.

Fig. 3 Solution A is formed.

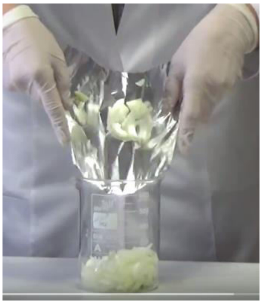

Fig. 4 The diced onions are put in a beaker.

Fig. 5 Solution A is added to the onions in the beaker.

Fig. 6 The mixture is blended.

Fig. 7 After the mixture is blended the mixture is filtered with filter paper through a funnel.

Fig. 8 Ethyl alcohol is added to the filtered mixture.

Fig. 9 The DNA of the onion can be seen after 10 minutes.

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HOW IS TINCTURE OF IODINE MADE?

OVERVIEW

The goal of this experiment is the explanation of solubility and the factors influencing solubility.

Solution and its kinds

Solution is the homogenous mixture of two or more substances and it comprises of at least two components. One of the components is solvent, the other is dissolved. The less the components forming the solution are called "dissolved", the more is called "solvent". The solvent, in general, determines the physical state of the solution. So, solutions can be in a state of gas, liquid or solid. For example, the air we breathe is a "gaseous solution" of various gases such as oxygen, argon dissolved in nitrogen. Sea water is a "liquid solution" of salt and several matters found in water. Amalgam which is used in tooth filling is a "solid solution" formed by the dissolving of silver and mercury.

Solvent and dissolved naming is completely arbitrary. A few numbers of mixtures are accepted as solvents. For instance, though the amount of water is very little, in a solution of sulphuric acid and water, sulphuric acid is "dissolved" and water is "solvent".

Most of the chemical reactions can not only be made among pure, solid, liquid and gases but also can be realized between ions and molecules dissolved in water or other solvents. The most common solutions are liquid solutions. In liquid solutions, if the solvent is water, then the solution is called "diluted". Water dissolves a lot of compounds of ionic and molecular structure. Most of the dissolved compounds decompose into their ions and these ions take place in chemical reactions. In general chemistry labs, most of the reactions are realized in diluted (watery) solutions.

Formation of Dissolution And Solution

Dissolution occurs with the homogeneous mixture of solvent and dissolved. When a substance is dissolved in another substance the particles of the dissolved are distributed homogeneously in the solvent and thus the particles of the dissolved take place among the molecules of solvent. Three important reciprocal interactions in this state (dissolved- dissolved, solvent-solvent, dissolved-solvent) are explained below.

1. Leaving the dissolved particles each other.

2. Leaving of the solvent particles each other to open a space for placing the dissolved.

3. Attracting the dissolved particles by the solvent particles.

If the attracting power between the molecules of solvent and dissolved is almost the same, molecules mix randomly and thus form a homogeneous mixture that is a new solution.

If the attractive power between the solvent- dissolved molecules is more than the attractive power of the solvent-solvent, dissolved-dissolved, again the dissolution occurs and a solution is formed.

With the increase of irregularity in solution, in the states of the attractive power of solvent –dissolved molecules are less than that of the solvent-solvent and dissolved-dissolved a solution may occur. In order to explain this situation two factors causing this state must be considered. As known, all the systems want to be stable by decreasing their energy and also they want to be in the most irregular state. As the state of solution forms much more irregular state than the dissolved state, the dissolved- solvent interaction even if not enough to meet the total interaction between dissolved-dissolved and solvent-solvent in terms of energy, due to the increase in irregularity, dissolution occurs.

Solubility: Solubility is the name given to an amount of matter that can dissolve, at any temperature, in a certain amount of solvent. At a certain temperature, prepare a saturated solution. When temperature is adjusted to a point where solubility is lower (generally temperature is reduced) generally the excess of the dissolved settles. However, sometimes no settlement is seen. The amount of dissolved matter in the solution becomes more than necessary that can dissolve at that temperature. Such solutions are called excessively saturated solutions. An excessively saturated solution is unstable, by adding some dissolved crystals (these extra crystals play the role of vaccination), the excessively dissolved part (or the excessively dissolved amount) settles and changes into "saturated solution".

MATERIALS AND METHODS

The materials used are; iodine, sodium chlorine or potassium iodine, ethyl alcohol, mixer, beaker, mortar.

The experiment is carried out by:

28. 7g of I_2 is pounded to make powder.
29. 3g Kl, in powder is mixed with I_2.
30. Put 90g ethyl alcohol into the mixture and then the tincture of iodine is ready.
31. In a 100g tincture of iodine, there exist 7g I_2, 3g Kl, 90g C_2H_5OH.

HAZARDS

C_2H_6O (Ethyl alcohol)

CHEMICAL REACTION EQUATION

$KI + I_2 \rightarrow KI_3$

$I^- + I_2 \rightarrow I_3^-$

CONCLUSIONS

While explaining solubility, the making of tincture of iodine can be given as an example. This will be an experiment to draw the attention of the students, just by explaining both the topic of solubility and the influence of shared ion upon solubility.

WASTE MANAGEMENT

The iodine left from the experiment must be put in a sealed waste container.

PHOTOS

Fig. 1 The iodine is crushed with a mortar and pestle.

Fig. 2 Potassium iodine is added to the crushed iodine.

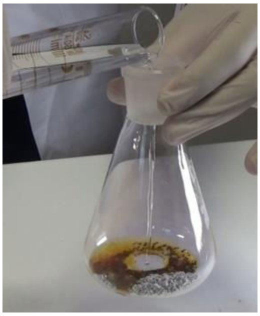

Fig. 3 After the Potassium iodine, and crushed iodine is put in the flask finally, ethyl alcohol is added to the mixture.

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THE DIFFUSION RATE OF IONS IN WATER

OVERVIEW

The aim of the experiment is to show that different ions have a different diffusion rate in water. This is done by adding two different salts on both sides of a Petri dish and determine where the precipitation of two ions occurs.

MATERIALS AND METHODS

- Potassium thiocyanate (KSCN)
- Lead(II) nitrate (Pb(NO3)2)
- Potassium iodide(KI)
- Iron (III) chloride hexahydrate (FeCl3.6H2O)
- Distilled water
- 2 Petri dishes
- Spoon

Gather all the materials. Pour 1 cm distilled water in both Petri dishes. In Petri dish 1 add 1 spatula point of lead nitrate and potassium iodide. Make sure one of the chemicals is on the right side and the other on the left. Do the same in the other Petri dish using iron(III) chloride and potassium thiocyanate. Wait several minutes to let the ions diffuse and precipitate.

HAZARDS

$Pb(NO_3)_2$		**KI**	
$FeCl_3.6H_2O$		**KSCN**	

CHEMICAL REACTION EQUATION

Dissolving

$KSCN_{(s)} \rightarrow K^+_{(aq)} + SCN^-_{(aq)}$

$Pb(NO_3)_{2(s)} \rightarrow Pb^{2+}_{(aq)} + 2NO_3^-{}_{(aq)}$

$KI_{(s)} \rightarrow K^+_{(aq)} + I^-_{(aq)}$

$FeCl_3.\,6H_2O_{(s)} \rightarrow Fe^{3+}_{(aq)} + 3Cl^-_{(aq)} + 6H_2O_{(aq)}$

Precipitation

$\mathbf{Fe^{3+}_{(aq)} + 3SCN^{-}_{(aq)} \rightarrow FeSCN_{3(s)}}$

$\mathbf{Pb^{2+}_{(aq)} + 2\ I^{-}_{(aq)} \rightarrow PbI_{2(s)}}$

CONCLUSIONS

It is clear that the iron ions have travelled less distance than the thiocyanate ions, since the precipitation takes place closer to where the iron chloride was added than where the potassium thiocyanate was added. Therefor the diffusion rate of thiocyanate is higher than the diffusion rate of iron ions.

In a similar way it can be said that the iodine ions have a lower diffusion rate than the lead ions.

WASTE MANAGEMENT

All the reaction products should be thrown away in the heavy metals waste bin.

PHOTOS

Fig. 1 The reaction between potassium thiocyanate and Iron (III) chloride.

Fig. 2 The reaction between potassium iodide and lead(**II**) nitrate.

REFERENCES

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SURFACE TENSION SPEEDBOAT

OVERVIEW

Besides having the ability to emulsify fats, soap can reduce the surface tension of water. In this experiment we will test this using a creative approach.

MATERIALS AND METHODS

- Scissors
- Aluminum foil
- A plastic box of approximately 25 by 45 centimetres
- Water
- A plastic syringe
- A test tube containing soap

First, gather all the required materials. Fill the plastic box with cold water to approximately 1 cm high and make sure the surface is perfectly horizontal so that all the water is evenly distributed. Use your scissors to cut out a shape of aluminium foil which looks like the aluminium pieces as seen in figure 2. Make sure it is 8 centimetres in length. Put the cut out shape on the surface of the water at one end of the box, pointing at the other end. Use your syringe to obtain soap from the test tube and drop a drop of soap in the hole of the aluminium ‘boat’.

HAZARDS

Al

CHEMICAL REACTION EQUATION AND EXPLANATION

No chemical reaction occurs in this experiment, yet there is an explanation. Soap is a surfactant, which means it has a big influence on the surface tension. To be more precise it reduces the surface tension of water. This reaction pushes water away from the spot where soap is dropped on water. The specific shape makes it possible to use this tension reduction in order to be propelled forward. Because there is only one way for water to be pushed away, the aluminium piece is propelled forward.

CONCLUSIONS

Soap indeed reduces the surface tension of water concluded from the observation that the aluminium boat was propelled forward after the soap was dropped in the cut out hole.

WASTE MANAGEMENT

Water and soap can be flushed down the drain. Aluminium is a metal, but because this is aluminium foil it can be thrown in the trash.

PHOTOS

Fig. 1 Here you can see the required materials for this experiment included the specific aluminium piece.

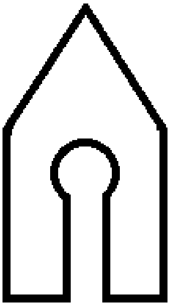

Fig. 2 Shape of the aluminium 'boat'.

REFERENCES

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HOTPACK AND SUPERSATURATED SOLUTIONS

OVERVIEW

Once upon a time there was a little girl who got a very interesting present in a freezing Christmas Day. She found in her present two pink hearts, as big as her hands were, and the hearts were filled in with a liquid, making them soft and flexible. If she pressed a metallic blade which was inside the hearts suddenly the content become firm and warm. It was a pocket hand warmer for freezing days. Once the blades pressed she put the hearts in her pockets and it was warm and pleasant for her hands as she went to school or she returned home. The warmth lasted a few minutes, long enough to have several times the opportunity to find shelter for the freezing fingers. And surprise, if she boiled the hearts in hot water they become again soft and flexible, ready for another tour outside.[1]

Later on, she started asking questions, first in her mind and later on to her Science teacher. And she found an answer … see "Conclusions".

Supersaturated solutions have practical applications in medicine for the hot packs widely used in emergency cases. All hot packs rely on the same principle: a supersaturated solution is seeded, it crystallizes releasing heat which help paramedics to keep warm parts of the body of an injured.

MATERIALS AND METHODS

Take preferably a new 200 mL beaker. Wash it thoroughly but avoid to scratch the inner surface, and weigh 125 g of sodium acetate trihydrate (CH_3-$COONa^{\cdot}3H_2O$). Pour over the solid 30 mL of distilled water.[2] Gently heat the mixture until a clear solution is obtained. As you will notice a layer of crystals as a ring on the inner wall of the beaker, right above the solution's surface, pour the hot solution in another clean and new beaker, cover it with a watch glass and allow to cool down. As it is cooled at room temperature you will notice that the solution remain clear and colorless. Measure the temperature of the solution than place a tiny crystal in the middle of the surface. Notice how from the spot where the initial crystal was dropped a hedgehog of acicular crystals are growing occupying eventually the entire volume of the solution. Measure again the temperature, notice the hot evolved and you will have the explanation of the hot packs used in medicine.

Another example used to demonstrate the principle of a hot pack due to supersaturated solutions crystallization is the sodium thiosulfate pentahydrate example $Na_2S_2O_3 \cdot 5H_2O$. Take another clean (preferably new) 200 mL beaker and fill it with solid sodium thiosulfate pentahydrate ($Na_2S_2O_3*5H_2O$). Gently heat the solid and notice how it melts in its own crystal water giving a clear homogenous liquid. [3] Pour the solution in a second clean 200 mL beaker and allow the solution to cool to room temperature. To initiate crystallization drop a single crystal of $Na_2S_2O_3 \cdot 5H_2O$ into the cooled liquid. Crystallization begins and the beaker becomes warm (42°C).

Both beakers can be reheated, the clear liquids poured in another beaker and the demonstration can be repeated after cooling, another time with another group of students. The supersaturated solutions of both chemicals, sodium acetate and sodium thiosulfate, are stable for days/weeks if the cleanliness principle is respected.

HAZARDS

There are no hazards related to the used chemicals. Just obey the general rules corresponding to the manipulation of chemicals and to the heating and hot solutions manipulation.

CHEMICAL REACTION EQUATION

$$Na^+_{(aq)} + CH_3\text{-}COO^-_{(aq)} + 3\ H_2O_{(l)} \text{--->} CH_3\text{-}COONa*\ 3H_2O_{(s)}$$

(-19,7kJ/mol)

$$2\ Na^+_{(aq)} + S_2O_3^{2-}{}_{(aq)} + 5\ H_2O_{(l)} \text{--->} Na_2S_2O_3*5\ H_2O_{(s)}$$

(-48 kJ/mol)

CONCLUSIONS

The palm hearts of the little girl were filled in with a mixture of water and sodium acetate, acetate which dissolves in the provided water only when hot. A saturated solution was thus created, which upon cooling at room temperature remained as a solution - in this case a supersaturated one. When the metallic blade from inside was turned tiny iron crystals were seeding the solution which started to crystallize in the entire inner volume of the heart making it firm and releasing heat .. not too much ... only up to 36-37 °C. When a hot pack is used the inner bag containing a supersaturated solution of sodium thiosulfate is squeezed. The solution reaches a space where a few crystals of solid sodium thiosulfate are placed which, in contact with the solution, induce the crystallization process releasing the necessary heat. The negative enthalpy explains why these supersaturated solutions are releasing heat when crystallizing. The crystallization heat for the sodium acetate trihydrate is -19.7 kJ/mole and the one of the sodium thiosulfate is higher (-48 kJ/mole). This explains why the sodium acetate solutions are warming up to 36-38°C whereas the sodium thiosulfate can warm up to 48°C upon crystallization.

The supersaturated sodium thiosulfate solutions can crystallize in a very peculiar way giving birth to huge mono crystals.

WASTE MANAGEMENT

The crystallization process can be repeated several times so there is no need to discard the used reagents.

PHOTOS

Fig. 1/2 Heat evolvement upon crystallization of supersaturated sodium acetate CH_3-COONa.

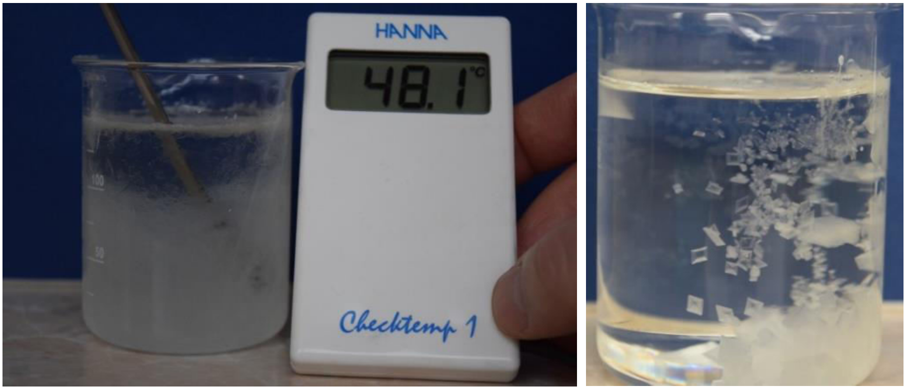

Fig. 3/4 Crystallization and heat evolvement of supersaturated sodium thiosulfate solution.

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THE PROCESS BEHIND AN INSTANT COOLPACK

OVERVIEW

The aim of the experiment is finding out how an instant coolpack works. An endothermic reaction will take place which extracts heat from the environment. The solution will cool down slowly and regain room temperature very rapidly.

MATERIALS AND METHOD

- 5 millilitres of water
- Measuring cylinder (10 ml)
- ammonium nitrate, 11 grams
- 2 spoons
- 1 thermometer
- Beaker
- Scale

Gather your materials. Measure 5 mL of water and pour it in the beaker. Measure the temperature of the water. Leave the thermometer in the beaker. Measure 11 grams of ammonium nitrate. Add the ammonium nitrate to the water and stir until dissolved. Measure the temperature of the mixture with the thermometer.

HAZARDS

NH_4NO_3

CHEMICAL REACTION EQUATION

$NH_4NO_{3(s)}$ ---> $NH_4^+{}_{(aq)}$ + $NO_3^-{}_{(aq)}$

CONCLUSION

When dissolving ammonium nitrate in water, heat is extracted from the environment which makes the mixture cool down to about 2°C within one minute. This dissolving of ammonium nitrate is therefor an endothermic process.

WASTE MANAGEMENT

The mixture can be flushed down the drain.

PHOTOS

Fig. 1 These are all the materials you will need during the experiment.

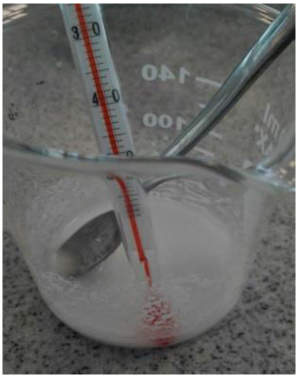

Fig. 2 The solution after you have added the ammonium nitrate.

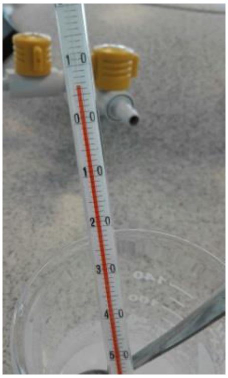

Fig. 3 The temperature when the solution is at its coldest point.

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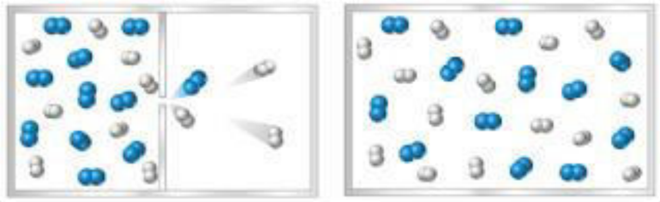
THE MEASUREMENT OF THE DIFFUSION SPEED OF GASES

OVERVIEW

Two gases or more, in terms of the temperature (E_K) and molecular mass, can mix homogenously in short time in a container. This is called **diffusion**. If a hole is opened in a closed container where there are gas molecules they escape this hole due to the pressure, temperature(E_K) and molecular mass between the container and outside, this is called **effusion** (expansion outside). When any type of perfume or natural gas leakage is realized in a room, the reason for this is the expansion or spreading of the gas molecules. All gases expand in each other and form a homogeneous mixture, whether the same type of gas or in different conditions



The mixture of gases with each other can best be explained by their striking to the walls of the container and to each other while moving permanently and randomly. During this striking, a flow of kinetic energy occurs among molecules. However, the total kinetic energy of the system does not change, even if there are different gases in the system, the average kinetic energy of all the gases is the same at an even temperature which can be shown as below;

$$\frac{1}{2}m_1V_1^2 = \frac{1}{2}m_2V_2^2 \longrightarrow m_1V_1^2 = m_2V_2^2$$

$$\sqrt{\frac{V_1^2}{V_2^2}} = \sqrt{\frac{m_2}{m_1}} \qquad \longrightarrow \qquad \frac{V_1}{V_2} = \sqrt{\frac{m_2}{m_1}}$$

$$\boldsymbol{\frac{V_1}{V_2} = \sqrt{\frac{m_{A2}}{m_{A1}}} = \sqrt{\frac{d_2}{d_1}}}$$

Due to, the kinetic energy of the gases in the same temperature will be equal their expansion speed will be affected from the molecular mass.

MATERIALS AND METHODS

A 50 cm glass tube is fixed parallel to the floor using two supporting rods. Then, cotton is fixed to the inside of the rubber corks using the pins of HCl and NH_3 solutions. 7-8 drops are dripped into the cotton at the same time, and the corks are fitted to both ends of the glass tube. Finally, observe the tube, and a white ring appears from the HCl and NH_3 gases and then the length of the ring to both ends of the tube is measured.

HAZARDS

NH_3

HCl

CONCLUSIONS

The diffusion speed of a gas is inversely proportional with the square of its molecular weight. The molecular of ammonium is 17,3 g/M and the molecular mass of hydrochloric acid is 36,5 g/M. Ammonium gas whose molecular mass moves faster than the hydrochloric acid is observed in the glass tube.

WASTE MANAGEMENT

After the observation of an ammonium chlorine cloud in the glass tube, remove the cotton from the tip of the mechanism. After this, the sweeping of the vapor is done through a fume hood, and the glass tube and cotton are washed.

PHOTOS

Fig. 1 Due to HCl and NH_3 being toxic and spreading very quickly this experiment must be conducted in a fume hood.

Fig. 2 The cotton which HCl and NH_3 were dropped are screwed on either end of a glass tube and the spreading of the gas is observed.

Fig. 3 Due to the gas that has spread through the tube a white cloud is observed. This cloud is ammonium chloride [NH_4Cl].

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PAPER CHROMATOGRAPHY OF TEXTILE DYES

OVERVIEW

Chromatography is the general term of an analytical method in which the separation, especially of closely related compounds, by allowing a solution or mixture to seep through an absorbent so each compound becomes absorbed into a separate, often colored layer, quantitatively and qualitatively. The study in such methods is based on two compounds, constant phase and moving (mobile) phase or are the names given to these compounds. The constant phase can be solid and liquid; the moving (mobile) phase can be liquid and gas. The compounds in the moving (mobile) phase, due to their interaction with the filling materials of constant phase, will be held for a while. This holding will be in different amounts for different compounds shown in the example. Thus, the compounds since they advance in different speeds towards the end of the constant phase will leave the constant phase in different times separately. As a result, the concentration of the compounds coming from the constant phase can be measured appropriately and the graphics, named "**chromatogram**" are obtained in respect to time and the volume of the moving phase used. In our day, chromatography is used to

32. Find out how many compounds the obtained mixture is composed of.
33. Determine what each of these compounds are and how much their ratio is.
34. Find out if any element is found in any mixture.
35. Understand whether a continuing reaction has finished.
36. Obtain each of the compounds forming the mixture purely.

The moving of each material with the mobile phase is different. Thus, it is possible to give characteristic values for an element defining the conditions applied. This is called "drifting degree" or "holding factor" and shown as R_f.

$$R_f = {}^{\text{The distance the element advenced(cm)}}\Big/_{\text{the distance the solvent advanced(cm)}}$$

$$R_f = {}^{a}/_{b}$$

R_f is constant for the same constant phase and the same solvent for any matter. Using the R_f values, it is easy to determine the element in an unknown solution.

MATERIALS AND METHODS

One drop of gren nutrient dye is dropped onto a dry filter paper. Then, just to the center of the dye, a comple of water drops are dropped. After waiting for a moment, if not enough, water is dropped again. The dyes in the green nutrient dye will be separated as blue and yellow. Water, after carrying the to colors in the dye to different lengths, dissolves from the ends of the filter paper. After the water reaches to the ends 3/2 of the paper, it is left to dry. The same experiment can be made by cutting a strip from the filter paper. A small quantity of dye is dropped in the center of the filter paper. The end of the filter paper is dipped into the beaker, full of water. We will observe that to the upper end of the filter paper, different colours will move.

CONCLUSIONS

Solvent selection is made considering the interaction of constant phase and active phase the interaction. There, the element required to separate into it components with the active phase is more than its interaction with the constant phase. The element holds the active phase firmly and is drift. Chromotographic separation is made using this principle. As a consequence, due to the interaction between the components of constant and active phases, the components are separated.

WASTE MANAGEMENT

Since the filter paper used is not contaminated with hazardous chemicals is accumulated in a solid waste container. As the mobile phase used in this test is liquid, it can be discharged into the tap.

PHOTOS

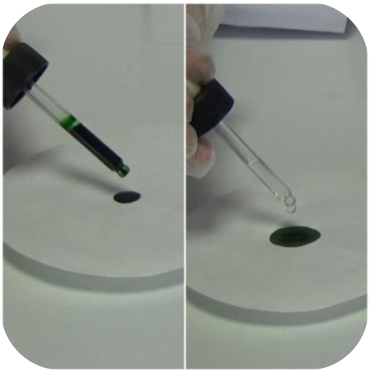

Fig. 1 A drop of dye right in the middle of a filter paper is dropped, and then a drop of the solvent is added.

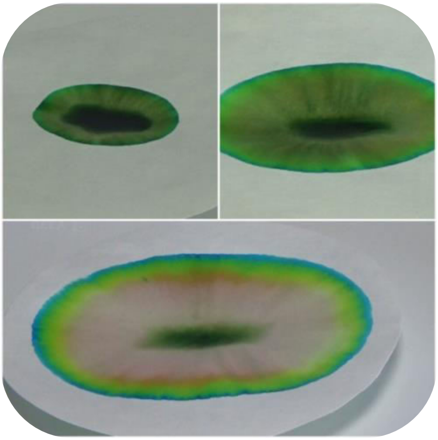

Fig. 2 It is observed that the solvent moves the dye and that the main colors in the dye separate within time.

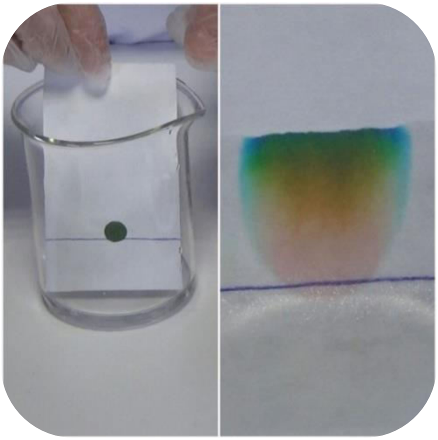

Fig. 3 The same process with the filter paper is conducted when dipped in the solvent.

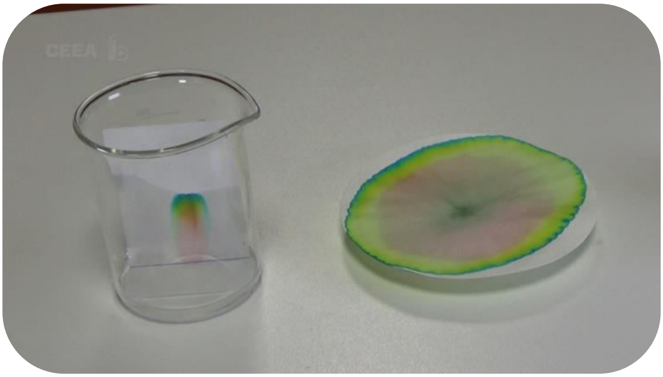

Fig. 4 In the end the components of the dye are separated due to the excessive and stable reaction difference.

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