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EXPERIMENTAL LABORATORY MANUAL
by discipline GENERAL CHEMISTRY
for students of technical specialties

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CONTENTS

INTRODUCTION

The study programs for scientific and technical degrees include classroom theoretical teaching, analysis and resolution of real exercises, and some more practical teaching in the laboratory. The laboratory teaching focuses on the application of the theoretical concepts and principles to practical cases, in order to complement the theoretical classes and to reinforce the understandability of the main theoretical concepts.

Commonly, the professor provides the students with a detailed recipe about what to do and how to do in the lab. The recipe contains the theoretical basis and the detailed description of the work in the lab in order to reach the desired result. However, there is different approach to the laboratory teaching. The professor can provide the student with the theoretical concepts and the practical tools to get the desired results without a detailed recipe. It makes the laboratory teaching more dynamic and creative. However, it implies more work for the professor and the student. The professor needs to supervise the work of the student more closely; and the student has to organize his/her work before to go to the lab.

Chemistry is the science of matter, its properties, and changes. In your classroom work in chemistry, you will learn a great deal of the information that has been gathered by scientists about matter. But, chemistry is not just information. It is also a process for finding out more about matter and its changes. Laboratory activities are the primary means that chemists use to learn more about matter. The activities in the Laboratory Manual require that you form and test hypotheses, measure and record data and observations, analyze those data, and draw conclusions based on those data and your knowledge of chemistry. These processes are the same as those used by professional chemists and all other scientists.

Organization of Activities:

- Introduction: following the title and number of each activity, an introduction provides a background discussion about the problem you will study in the activity.
- Problem: the problem to be studied in this activity is clearly stated.
- Objectives: the objectives are statements of what you should accomplish by doing the investigation. Recheck this list when you have finished the activity.
- Materials: the materials list shows the apparatus you need to have on hand for the activity.
- Safety: precautions Safety symbols and statements warn you of potential hazards in the laboratory. Before beginning any activity, refer to page 23 to see what these symbols mean.
- Pre-Lab: the questions in this section check your knowledge of important concepts needed to complete the activity successfully.

- Procedure. The numbered steps of the procedure tell you how to carry out the activity and sometimes offer hints to help you be successful in the laboratory. Some activities have CAUTION statements in the procedure to alert you to hazardous substances or techniques.
- Hypothesis: this section provides an opportunity for you to write down a hypothesis for this activity.
- Data and Observations: this section presents a suggested table or form for collecting your laboratory data. Always record data and observations in an organized way as you do the activity.
- Analyze and Conclude. The Analyze and Conclude section shows you how to perform the calculations necessary for you to analyze your data and reach conclusions. It provides questions to aid you in interpreting data and observations in order to reach an experimental result. You are asked to form a scientific conclusion based on what you actually observed, not what "should have happened." An opportunity to analyze possible errors in the activity is also given.
- Real-World Chemistry. The questions in this section ask you to apply what you have learned in the activity to other real-life situations. You may be asked to make additional conclusions or research a question related to the activity.

When scientists perform experiments, they make observations, collect and analyze data, and formulate generalizations about the data. When you work in the laboratory, you should record all your data in a laboratory report. An analysis of data is easier if all data are recorded in an organized, logical manner. Tables and graphs are often used for this purpose.

Title: The title should clearly describe the topic of the report.

Hypothesis: Write a statement to express your expectations of the results and as an answer to the problem statement.

Materials: List of all laboratory equipment and other materials needed to perform the experiment.

Procedure: Describe each step of the procedure so that someone else could perform the experiment following your directions.

Results: Include in your report all data, tables, graphs, and sketches used to arrive at your conclusions.

Conclusions: Record your conclusions in a paragraph at the end of your report. Your conclusions should be an analysis of your collected data.

HOW TO PREPARE FOR CHEMISTRY LAB?

Chemistry lab is a required component of most chemistry courses. Learning about lab procedures and performing experiments helps you to learn techniques and reinforces textbook concepts. Make the most of your time in the lab by coming to lab prepared. Review these pre-lab tips before starting an experiment.

Complete any pre-lab assignments or homework. The information and calculations are intended to make the lab exercise quicker and easier.

Know the location of the lab safety equipment and understand how to use it. In particular, know the location of the emergency exit, fire extinguisher, eye wash station and safety shower.

Read through the experiment before going to lab. Make sure you understand the steps of the experiment. Write down any questions you have so that you can ask them before starting lab.

Start filling out your lab book with information about the experiment. It's a good idea to draw out your data table in advance so all you need to do in lab is fill it in with numbers.

Review the Material Safety Data Sheets (MSDSs) of the chemicals you will be using during lab.

Make certain you have all of the glassware, materials and chemicals needed to complete the lab before starting any part of the procedure.

Understand disposal procedures for the chemicals and other items used in your experiment. If you are unclear about what to do with your experiment after it has been completed, ask your instructor about it. Don't throw items in the trash or dump liquids down the drain or in waste disposal containers until you are certain it is acceptable to do so.

Be prepared to take data in the lab. Bring your notebook, a pen and a calculator.

Have personal safety gear, such as a lab coat and goggles, clean and ready to use before lab.

SAFETY IN THE LABORATORY

The chemistry laboratory is a place to experiment and learn. You must assume responsibility for your own personal safety and that of people working near you. Accidents are usually caused by carelessness, but you can help prevent them by closely following the instructions printed in this manual and those given to you by your teacher. The following are some safety rules to help guide you in protecting yourself and others from injury in a laboratory.

1. The chemistry laboratory is a place for serious work. Do not perform activities without your teacher’s permission. Never work alone in the laboratory. Work only when your teacher is present.

2. Study your lab activity before you come to the lab. If you are in doubt about any procedures, ask your teacher for help.

3. Safety goggles and a laboratory apron must be worn whenever you work in the lab. Gloves should be worn whenever you use chemicals that cause irritations or can be absorbed through the skin.

4. Contact lenses should not be worn in the lab, even if goggles are worn. Lenses can absorb vapors and are difficult to remove in an emergency.

5. Long hair should be tied back to reduce the possibility of it catching fire.

6. Avoid wearing dangling jewelry or loose, draping clothing. The loose clothing may catch fire and either the clothing or jewelry could catch on chemical apparatus.

7. Wear shoes that cover the feet at all times. Bare feet or sandals are not permitted in the lab.

8. Know the location of the fire extinguisher, safety shower, eyewash, fire blanket, and first-aid kit. Know how to use the safety equipment provided for you.

9. Report any accident, injury, incorrect procedure, or damaged equipment immediately to your teacher.

10. Handle chemicals carefully. *Check the labels of all bottles before removing the contents.* Read the labels three times: before you pick up the container, when the container is in your hand, and when you put the bottle back.

11. Do not return unused chemicals to reagent bottles.

12. Do not take reagent bottles to your work area unless specifically instructed to do so. Use test tubes, paper, or beakers to obtain your chemicals. Take only small amounts. It is easier to get more than to dispose of excess.

13. Do not insert droppers into reagent bottles. Pour a small amount of the chemical into a beaker.

14. Never taste any chemical substance. Never draw any chemicals into a pipette with your mouth. Eating, drinking, chewing gum, and smoking are prohibited in the laboratory.

15. If chemicals come into contact with your eyes or skin, flush the area immediately with large quantities of water. Immediately inform your teacher of the nature of the spill.

16. Keep combustible materials away from open flames. (Alcohol and acetone are combustible.)

17. Handle toxic and combustible gases only under the direction of your teacher. Use the fume hood when such materials are present.

18. When heating a substance in a test tube, be careful not to point the mouth of the tube at another person or yourself. Never look down the mouth of a test tube.

19. Use caution and the proper equipment when handling hot apparatus or glassware. Hot glass looks the same as cool glass.

20. Dispose of broken glass, unused chemicals, and products of reactions only as directed by your teacher.

21. Know the correct procedure for preparing acid solutions. *Always add the acid slowly to the water*.

22. Keep the balance area clean. Never weigh chemicals directly on the pan of the balance.

23. Do not heat graduated cylinders, burettes, or pipettes with a laboratory burner.

24. After completing an activity, clean and put away your equipment. Clean your work area. Make sure the gas and water are turned off. Wash your hands with soap and water before you leave the lab.

25. The below given table is helpfull in an accident situation. Read and remember them:

Situation	Safe Response
Burns	Flush with ethanol solution or diluted solution of $KMnO_4$.
Cuts and Bruises	Treat as directed by instructions included with first aid kit
Fainting or collapse	Provide person with fresh air, have him recline in a position so that his head is lower than their body
Foreign Matter in Eyes	Flush about 15 min with plenty of water, then go to the doctor
Severe bleeding	Apply pressure or a compress directly to the wound and get medical attention immediately
1. Spills, general 2. Acid burns 3. Base burns	1. Wash area with plenty of water, use safety shower if needed 2. Use sodium hydrogen carbonate (baking soda) 3. Use 3 % of boric acid or acetic acid

HOW TO WRITE A LAB REPORT?

Lab reports are an essential part of all laboratory courses and usually a significant part of your grade. If your instructor gives you an outline for how to write a lab report, use that. Some instructors require the lab report be included in a lab notebook, while others will request a separate report. Here's a format for a lab report you can use if you aren't sure what to write or need an explanation of what to include in the different parts of the report.

A lab report is how you explain what you did in experiment, what you learned, and what the results meant. Here is a standard format. If you prefer, you can print and fill in the science lab report template or download the pdf version of the template.

Title Page. Not all lab reports have title pages, but if your instructor wants one, it would be a single page that states:

- The title of the experiment.
- Your name and the names of any lab partners.
- Your instructor's name.
- The date the lab was performed or the date the report was submitted.

Title. The title says what you did. It should be brief (aim for ten words or less) and describe the main point of the experiment or investigation. An example of a title would be: "Effects of Ultraviolet Light on Borax Crystal Growth Rate". If you can, begin your title using a keyword rather than an article like 'The' or 'A'.

Introduction / Purpose. Usually the Introduction is one paragraph that explains the objectives or purpose of the lab. In one sentence, state the hypothesis. Sometimes an introduction may contain background information, briefly summarize how the experiment was performed, state the findings of the experiment, and list the conclusions of the investigation. Even if you don't write a whole introduction, you need to state the purpose of the experiment, or why you did it. This would be where you state your hypothesis.

Materials. List everything needed to complete your experiment.

Methods. Describe the steps you completed during your investigation. This is your procedure. Be sufficiently detailed that anyone could read this section and duplicate your experiment. Write it as if you were giving direction for someone else to do the lab. It may be helpful to provide a Figure to diagram your experimental setup.

Data. Numerical data obtained from your procedure usually is presented as a table. Data encompasses what you recorded when you conducted the experiment. It's just the facts, not any interpretation of what they mean.

Results. Describe in words what the data means. Sometimes the Results section is combined with the Discussion (Results & Discussion).

Discussion or Analysis. The Data section contains numbers. The Analysis section contains any calculations you made based on those numbers. This is where you interpret the data and determine whether or not a hypothesis was accepted. This is also where you would discuss any mistakes you might have made while conducting the investigation. You may wish to describe ways the study might have been improved.

Conclusions. Most of the time the conclusion is a single paragraph that sums up what happened in the experiment, whether your hypothesis was accepted or rejected, and what this means.

Figures & Graphs. Graphs and figures must both be labeled with a descriptive title. Label the axes on a graph, being sure to include units of measurement. The independent variable is on the X-axis. The dependent variable (the one you are measuring) is on the Y-axis. Be sure to refer to figures and graphs in the text of your report.

References. If your research was based on someone else's work or if you cited facts that require documentation, then you should list these references.

LABORATORY EQUIPMENT AND INSTRUMENTS

Glass is the preferred material for most laboratory purposes. It is resistant to nearly all chemicals except hydrofluoric acid. However, prolonged exposure to strong bases should also be avoided. Only the fluorocarbons are more chemically resistant. Most laboratory glassware is composed of borosilicate glass such as pyrex or kimax. This material has excellent chemical resistance, and may be heated to high temperatures without risk of breaking or significant deformation. Unfortunately, many reagent bottles are made of high sodium glass. These vessels may shatter under extremes of temperature.

Many types of glassware are equipped with ground glass joints. These allow one to assemble all-glass apparatus without the need for corks or rubber stoppers. Bottles made of high sodium glass are often fitted with low-quality ground glass stoppers. These may not seal well, and they cannot be used with the standard interchangeable joints. Borosilicate glassware will have more carefully ground surfaces. The most common joints are the conical type (cone and socket), although spherical joints may also be used. Most conical joints have the "standard taper" of 1 mm change in diameter for every 10 mm in length. Standard taper joints are designated by a two number system. For example, the common "24/40" joint has a maximum diameter of 24 mm and a ground surface length of 40 mm. These are made to be interchangeable, so that any 24/40 cone should fit any 24/40 socket.

Non-volumetric Glassware. The least expensive types of glassware are the non-volumetric pieces. For this reason, operations in the laboratory not requiring precise measurement of volume or those occurring after or prior to volume measurement are best conducted with non-volumetric glassware. Beakers and flasks may be stored either in drawers or cabinets. They are best left upside down so that their interiors are not contaminated by settling dust.

Beakers. Beakers are most often used for mixing solutions, and dissolving solids in water. Solutions in beakers may be protected from atmospheric dust by placing an appropriately sized watch glass over the top.

Erlenmeyer Flasks. Erlenmeyer Flasks are frequently used as the receiving vessel during titrations. Some are made with ground glass stoppered tops. These may be used with reflux condensers (COD test) and other specialized glassware for a variety of purposes.

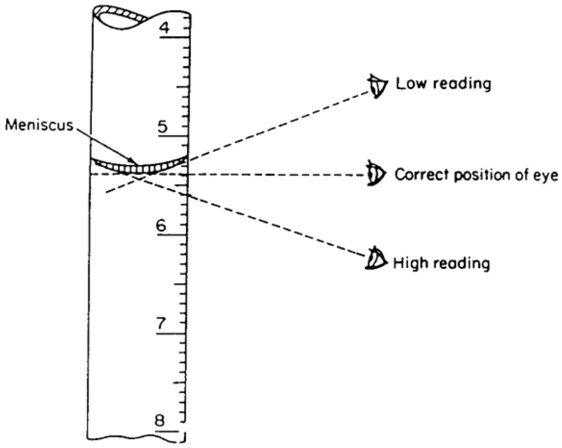
Reagent Bottles. Reagent bottles may have either flat-topped or round-topped stoppers. The flat-topped stoppers may be placed upside-down on the lab bench while the reagent is being dispensed. The round-topped stoppers must be either held between the fingers or set down on a clean watch glass, or in a clean beaker during use.

Volumetric Glassware. Volumetric glassware always contains one or more circular etched lines. When the bottom of the liquid meniscus just touches that line, the volumetric glassware is ready to either deliver the rated volume (pipets); actually contains the rated volume (flasks); or can be read to determine the precise volume delivered (burets). Volumetric glassware should be read with your eyes at the level of the meniscus, in order to avoid parallax errors. This is especially important when making buret readings.

Volumetric glassware is rated at 20°C and should always be used at or near that temperature. This is necessary, because glass expands and contracts much less than water, so that changes in temperature change the mass of water contained or delivered by volumetric glassware. Burets and pipets should be stored in drawers. Frequently used pipets may be kept upside down on the bench top in a specially designed pipet rack. Volumetric flasks should be stoppered and stored upright in wall cabinets. Volumetric glassware should be dried at room temperature. However, it need not be perfectly dry before re-use. It is a good practice to pre-rinse volumetric glassware in the liquid to be transferred, especially when the glassware is not perfectly dry. There is no need to fill the glassware with liquid, just add a small amount, swirl to wet all surfaces, and discard the rinse solution.

Technique for reading a meniscus

Measure the meniscus at eye level from the center of the meniscus (figure 1). In the case of water and most liquids, the meniscus is concave. Mercury produces a convex meniscus.



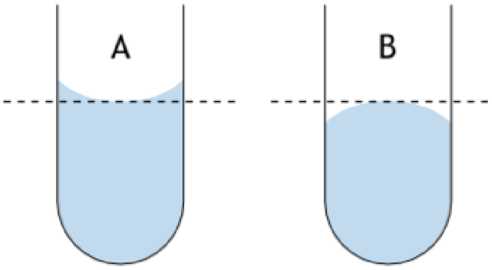


Figure 1 - Technique for reading a meniscus

How to read meniscus. The meniscus is the curve seen at the top of a liquid in response to its container. The meniscus can be either concave or convex. A concave meniscus (e.g., water in glass) occurs when the molecules of the liquid are more strongly attracted to the container than to each other. A convex meniscus (e.g., mercury in glass) is produced when the molecules of the liquid are more strongly attracted to each other than to the container.

In some cases, the meniscus appears flat (e.g., water in some plastics). When you read a scale on the side of a container with a meniscus, such as a graduated cylinder or volumetric flask, it's important that the measurement accounts for the meniscus. Measure so that the line you are reading is even with the center of the meniscus. For water and most liquids, this is the bottom of the meniscus. For mercury, take the measurement from the top of the meniscus. In either case, you are measuring based on the center of the meniscus.

Volumetric Pipets. Pipets are used to accurately measure out volumes from 1 mL to 100 mL. Larger volumes can be handled with volumetric flasks; smaller volumes with micropipets and syringes. Two types are available; the highly accurate fixed-volume transfer pipets, and the more flexible measuring pipet. They are filled by creating a suction with a pipet bulb. Never pipet by mouth! We often use Fisher 3-way pipet bulbs. Be careful not to draw liquid into the pipet bulb. If this happens, the $18 bulb may have to be discarded.

Pipets of 1 mL or greater are generally designated "to deliver" (i.e., TD). This means that they will deliver the stated volume when filled to the mark and allowed to drain by gravity while the tip is just submerged or contacting the side of a receiving vessel. You should wait about 20 seconds after all of the liquid has discharged to be sure that the meniscus has reached the correct, equilibrium level. Thus, the last small amount that remains in the tip due to capillary forces must not be forced out. Gently, wipe the outside of the pipet tip with a kimwipe. Note that you should also contact the pipet tip to the vessel while drawing up liquid. Always rinse pipets immediately after use. It is very difficult to remove dryed deposits from the insides of pipets.

Volumetric Flasks. Volumetric flasks are most commonly used for the preparation of standard solutions. Volumetric flasks are normally designated "to contain" (i.e., TC). This means that in order to deliver the stated volume, the liquid must be quantitatively transferred. Volumetric flasks often have round-topped (pennyhead) stoppers which must be either held between the fingers or set down on a clean watch glass, or in a clean beaker while filling or emptying the flask. Some volumetric flasks have polyethylene snap caps in place of ground glass stoppers.

Burets. A buret is a very precisely formed glass tube with calibrated gradations throughout most of its length. Near the lower extremity it contains a stopcock and drain tip. Note from Table 11.1 that the smaller the buret, the smaller the error. Thus you should always use the smallest buret that will completely deliver the total anticipated titrant volume. Proper use of burets will be discussed in the lab #7 handout (Water hardness, acid/base titration).

Graduated Cylinders. Graduated cylinders are commonly used when highest accuracy is not required, especially when odd volumes must be measured. Most also come with polyethylene safe-gard bumpers to prevent breakage in case of tipping.

Buret. Burets are available both with glass and teflon stopcocks. The traditional glass stopcocks must be lubricated with silicone grease. It is usually held in place with a metal tension clip. The teflon stopcocks do not require grease. They are held in place with a teflon washer and nut. The two types of stopcocks are not interchangeable, as they have different tapers. Burets equiped with glass stopcocks will often become contaminated with silicone grease. The beading of water on the inner surface, especially below the stopcock, is indicative of this type of contamination. Be sure that when the buret is full, the tip below the stopcock does not contain any air bubbles. If an air bubble exists here at the start of a titration, it may slowly fill in and introduce significant error. Burets are often used with a loose fitting cap for the purpose of keeping dust out and vapors in.

Be sure that your eyes are at the same height as the meniscus when you read a buret. In this way you minimize error due to parallax. Parallax causes one to overestimate the liquid height when looking at the meniscus from above and vice versa. Be sure that there is no liquid clinging to the inside walls of the buret above the meniscus.

Filtration Apparatus. Filtration, or the separation of solids from liquid, is a very important step in gravimetric analysis. The most common labware used for filtration is the "Millipore-type" apparatus. This consists of a glass or plastic base which supports the filter disc, and a cylindrical reservoir that sits on top of the base. The entire assembly fits on a filter flask by means of a rubber stopper. Typically, the filter flask is connected in series to another filter flask which is ultimately connected to a vacuum. The second filter flask serves as a safety barrier to minimize the chance that liquid will drawn into the vacuum source.

Most environmental filtrations will use either membrane filters or glass fiber filters. The former have better defined pores, and therefore, they result in sharper and more accurate separations. The membrane filters are most commonly used for microbiological determinations. The glass fiber filters, however, are cleaner and more inert than the membrane filters. They are also quite heat resistant. The glass fiber filters are generally used for thermo-gravimetric analyses and for trace organic analysis.

Glassware Cleaning. For the successful analysis of environmental samples, it is imperative that glassware should be scrupulously clean throughout all "crucial" steps of the analysis. Although an experienced analyst will know which steps are crucial, and where a less rigorous cleaning regimen is appropriate, the beginner will not have this advantage. Thus you are advised to always use one of the following procedures:

- General Use. Soak for 10-15 minutes in warm detergent solution. Avoid soaking for very long periods of time, as this may tend to roughen the glass surface (A variety of commercially available laboratory detergents may be dissolved in tap water for this purpose; e.g., 1% Alconox, 2% RBS-35). Rinse 3 times with tap water, followed by 3 rinses with distilled water. In the UMass Environmental Engineering laboratories, we often use a 10% hydrochloric acid rinse just prior to the final 3 rinses with distilled water. Place clean non-volumetric glassware upside down in a 110oC drying oven. Volumetric glassware may deform and loose their precision upon extended heating, therefore, they must be dryed at room temperature. This is the procedure we will be using for most of the experiments in CE 572. The removal of silicone grease presents a special challenge. It should be mechanically removed with kimwipes and pipe cleaners to the extent possible. Chemical removal is best achieved by a series of acid/base soaking cycles. Sometimes and ultrasound bath can be used to remove stubborn particles that may cling to glass surfaces. These devices should be filled with water when in use.

- Trace Organic Analysis. Rinse first with distilled water. Then soak for 20 min in chromic acid bath. Chromic acid may be prepared from commercially available chromerge or according to directions on the bottle (add slowly to a 9 lb bottle of conc. sulfuric acid), or by simply dissolving 25 g of sodium dichromate in 15 mL of water and slowly adding 500 mL of concentrated sulfuric acid. The eventual appearance of green Cr(III) indicates that the solution is exhausted. This cleaning procedure is commonly used for pesticide analysis by gas chromatography. Chromic acid should never be heated! It should always be used in a hood or in a covered bath to prevent atmospheric contamination. Hot chromic acid, or chromic acid contaminated with chloride may give off toxic, and even carcinogenic, fumes. A less toxic substitute called “Nochromix” is commercially available. Although it may be a less powerful oxidant, this product is preferred over chromic acid for reasons of safety.
- Trace Inorganic Analysis. Rinse first with distilled water. Then soak in re-distilled nitric acid. This procedure is commonly used for heavy metal analysis by atomic absorption spectrophotometry. Some soak glassware in a hot dilute (0.004M, pH 12) solution of EDTA for 20 minutes to remove metal ions.
- Oxidant Analysis. Rinse first with distilled water. Then soak overnight with the appropriate oxidant solution. Use a solution that is slightly more concentrated than the unknowns. In some cases it may be most convenient to leave the glassware in a permanent oxidant soak (e.g., chlorine demand free glassware). Other oxidants may be too transient, and glassware must then be prepared immediately before use (e.g., ozone demand free glassware). This procedure is commonly used for the analysis of residual chlorine.

Centrifuge

Analytical balance

Balance (electronic)

Bunsen burner

Crucible tongs

Pipets

Forceps

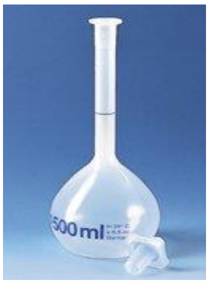
Funnels

Burner Crucible

Graduated Cylinders

Buret & Stand

Pipets and Bulbs



Volumetric flask

Watch dishes

Beakers

Vials

Petri dish

Buret

Flasks

Desiccator

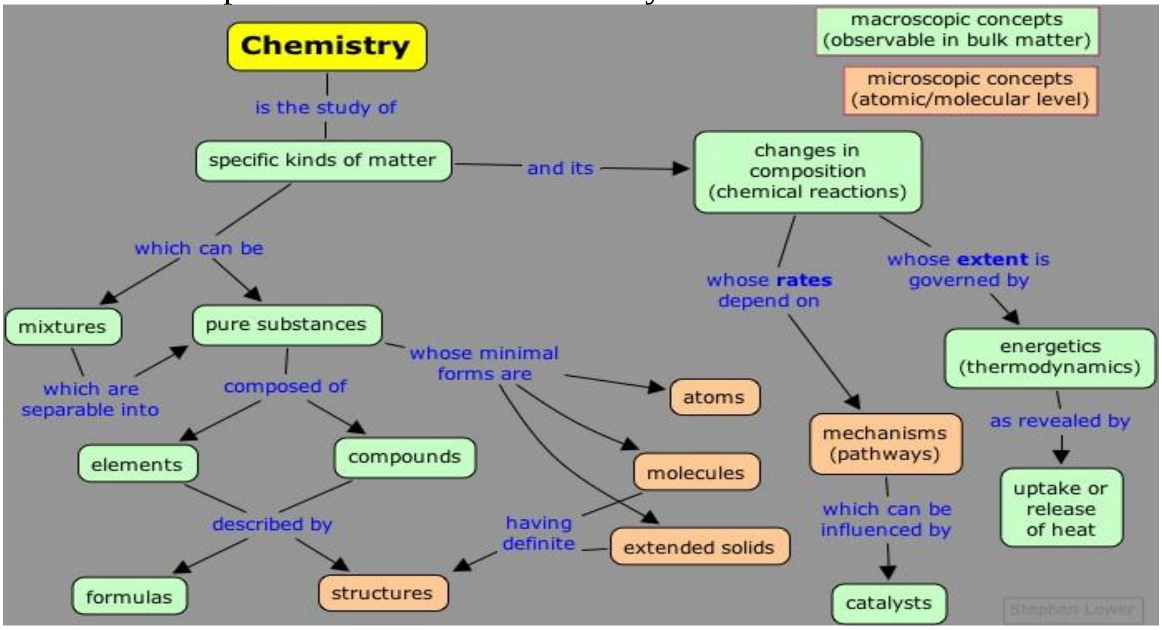
Drying cabinet

Muffle furnace

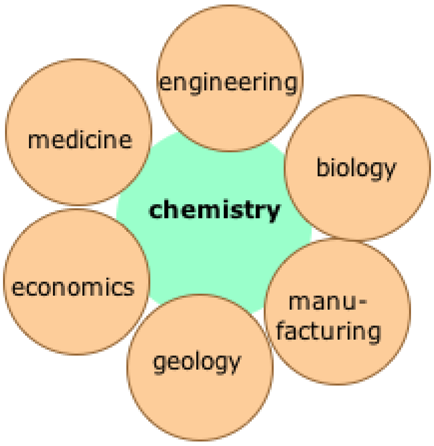
Water bath

WHAT IS CHEMISTRY?

Chemistry is a branch of physical science that studies the composition, structure, properties and change of matter. Chemistry is chiefly concerned with atoms and molecules and their interactions and transformations, for example, the properties of the chemical bonds formed between atoms to create chemical compounds. As such, chemistry studies the involvement of electrons and various forms of energy in photochemical reactions, oxidation-reduction reactions, changes in phases of matter, and separation of mixtures. Preparation and properties of complex substances, such as alloys, polymers, biological molecules, and pharmaceutical agents are considered in specialized fields of chemistry.



Chemistry is the study of matter and the changes that take place with that matter. Matter is everything that you can touch, see, feel, or smell. Matter is defined as anything that has rest mass and volume (space) and is made up of particles. Everything on Earth, everything in our solar system, everything in our galaxy is made up of matter. Because understanding chemistry helps in understand the world around you. The importance of chemistry is that it explains the world around you. If you take classes in chemistry, you'll apply math and logic, which can make studying chemistry a challenge.



Civilizations used technologies that would eventually form the basis to the various branches of chemistry. Examples include fermenting beer and wine, extracting metals from ores, extracting chemicals from plants, making alloys like bronze, making glass, making pottery and glazes and for medicines and perfumes. Chemists later began looking for substitutes of natural resources which were essential but rare to obtain and thus began the search for perfect or precise form of experiments. Chemistry is often called the *central science* because of its role in connecting the physical sciences, biological, medical and environmental sciences.

At the most fundamental level, chemistry can be organized along the lines shown here:

- *Dynamics* refers to the details of that rearrangements of atoms that occur during chemical change, and that strongly affect the rate at which change occurs.
- *Energetics* refers to the thermodynamics of chemical change, relating to the uptake or release of heat. More importantly, this aspect of chemistry controls the direction in which change occurs, and the mixture of substances that results.
- *Composition and structure* define the substances that are results of chemical change. Structure refers specifically to the relative arrangements of the atoms in space. The extent to which a given structure can persist itself is determined by energetics and dynamics.
- *Synthesis*, strictly speaking, refers to formation of new (and usually more complex) substances from simpler ones, but in the present context we use it in the more general sense to denote the operations required to bring about chemical change and to isolate the desired products.

Mastering General Chemistry

1) Success in general chemistry requires a blend of ingredients. It requires a clearly presented body of information which requires lucid instructions.

2) The central principle of modern chemistry revolves around the general ideas of chemistry involving the chemical changes precisely arrangement of molecules.

3) The perfect combinations of atoms into molecules, complicated structures which accounts to the different characteristics of materials.

4) Modern chemistry also emerged very strongly in the 18th century, when researchers began using the systematic balance as a tool in their research.

5) During this time chemists learn to balance essentials in measuring mass, which is nothing but the quantity of matter in the material.

Principles of General Chemistry. Chemists understood the central principle and the true nature of chemicals, atoms, molecules and thus began the journey of shaping molecules to order. The basic principle of Chemistry involves molding of chemical order beginning with chemical reactions, types, the general characteristics of chemicals and the basic stoichiometry involved in turning these simple elements into complex, essential large molecules.

General Chemistry Help. The main intention of introduction to General Chemistry is to provide an introduction to the most important principles and applications of chemistry. Its roots lie in descriptive chemistry, which focuses on the physical and chemical characteristics of the elements and their compounds and stresses a practical knowledge of reactivity and applications. Over the years General Chemistry has become more complex with a strong emphasis on the mathematical and theoretical aspects of science.

Chemistry is the study of nature and its behavior with matter. The atom is the basic building block of matter, representing the smallest unit of a chemical element. An atom in turn is composed of subatomic particles called protons, neutrons and electrons. An atom is classified according to the number of protons and neutrons in its nucleus.

General Chemistry Formulas. A chemical equation is the symbolic representation of a chemical reaction. The coefficients next to the symbols and formulae of entities are the absolute values of the stoichiometric numbers. A chemical equation describes what happens in a chemical reaction. Balancing a chemical equation refers to establishing the mathematical relationship between the quantity of reactants and products. Each side of the chemical equation must represent the same quantity of any particular element.

A chemical formula is a way of conveying data about the proportions of atoms that constitute a particular chemical compound, using a single line of chemical element symbols, numbers, and other symbols, such as dashes, brackets, parentheses, plus (+) and minus (-) signs. From the law of conservation of mass, the mass of the reactants in a reaction must be equal to the mass of the products. More specifically chemical equations must be balanced so that there is the same number of atoms of each element in the products as there is the same number of atoms of each element in the products as there is in the reactants.

General Chemistry Topics. More than thousands of years, human beings have fashioned natural materials into useful products. The topics included in General Chemistry are:

- Elements, Compounds, and Mixtures
- Redox Reactions
- Liquids
- Gases

- Equations of State
- Units & Conversions
- Matter
- Chemical Equations
- Chemical Reactions
- Solutions and mixture

General Chemistry Practice Problems. Below you could see problems. Solved Examples.

Question 1: What is the percentage of lead in galena, PbS? Calculate to 0.1%

Solution: The formula weight of PbS is obtained by adding the atomic weights of lead and sulfur, which we obtain from the standard table.

Weight of one lead atom (1 Pb) = 207.2
Weight of one sulfur atom (1 S) = 32.1
Weight of 1 molecule PbS (galena) = 239.3

Hence 239.3 of PbS contains 207.2 of lead. We see that 100.0g of PbS would contain:

$$\frac{207{,}2\ of\ lead}{239{,}3\ of\ galena}\cdot 100g\ of\ galena = 86{,}6g\ of\ lead$$

Hence the percentage of lead in galena is 86.6%.

Question 2: A propellant for rocket can be made by mixing powdered potassium perchlorate $KClO_4$, and powdered carbon (carbon black), C with a little adhesive to bind the powdered materials together. What weight of carbon should be mixed with 1000g of potassium perchlorate in order that the products of the reaction be KCl and CO_2

Solution: Taking the equation for the reaction as

$$KClO_4 + 2C \rightarrow KCl + 2CO_2$$

we first calculate the formula weight of $KClO_4$:

weight of 1K = 39.1
weight of 1Cl = 35.5
weight of 4O = 4 * 16.0 = 64.0
Weight of $KClO_4$ = 39.1+35.5+64=138.6

The atomic weight of carbon is 12.0; the weight 2C is 24.0. Hence the weight of carbon required is 24.0/138.6 times the weight of potassium perchlorate:

$$\frac{24{,}0\ C}{138{,}6\ KClO_4}\cdot 1000g\ (KClO_4) = 173g\ (C)$$

Hence about 173g of carbon is required for 1000g of potassium perchlorate.

SI UNIT

The first measurements were probably based on the human body (the length of the foot, for example). In time, fixed standards developed, but these varied from place to place. Each country or government (and often each trade) adopted its own units. As science became more quantitative in the seventeenth and eighteenth centuries, scientists found that the lack of standard units was a problem. < They began to seek a simple, international system of measurement. In 1791 a study committee of the French Academy of Sciences devised such a system. Called the metric system, it became the official system of measurement for France and was soon used by scientists throughout the world. Most nations have since adopted the metric system or, at least, have set a schedule for changing to it.

SI Base Units and SI Prefixes. In 1960 the General Conference of Weights and Measures adopted the International System of units (or SI, after the French le Systyme International d'Unites), which is a particular choice of metric units. This system has seven SI base units. In this chapter, we will discuss four base quantities: length, mass, time, and temperature.

One advantage of any metric system is that it is a decimal system. In SI, a larger or smaller unit for a physical quantity is indicated by an SI prefix, which is a prefix used in the International System to indicate a power of 10. For example, the base unit of length in SI is the meter (somewhat longer than a yard), and 10^{-2} meter is called a centimeter. Thus 2.54 centimeters equals $2.54*10^{-2}$ meters. The SI prefixes used in our manual are presented in Table 1.

Quantity	Name of unit	Abbreviation
Mass	Kilogram	Kg
Lenght	Meter	m
Temperature	kelvin	K
Amount of substance	Mole	Mol
Time	Second	S
Electric current	Ampere	A

Some quantities are expressed as a function of more than one fundamental units known as derived units. For example velocity, acceleration, work, energy etc. The quantities used in this manual are presented in Table 2.

Quantity with Symbol	Unit (S.I.)	Symbol
Velocity (v)	Metre per sec	ms^{-1}
Area (A)	Square metre	m^2
Volume (V)	Cubic metre	m^3
Density (r)	Kilogram m^{-3}	Kg m^{-3}
Energy (E)	Joule (J)	Kg m^2s^{-2}
Frequency (n)	Hertz	Cycle per sec
Pressure (P)	Pascal (Pa)	Nm^{-2}
Electrical charge	Coulomb (C)	A-s (ampere – second)

Length, Mass, and Time

The **meter (m)** is the SI base unit of length. By combining it with one of the SI prefixes, you can get a unit of appropriate size for any length measurement. For the very small lengths used in chemistry, the nanometer (nm; 1 nanometer =10^{-9} m) or the picometer (pm; 1 picometer = 10^{-12} m) is an acceptable SI unit. A non-SI unit of length traditionally used by chemists is the **angstrom (E),** which equals 10^{-10} m. (An oxygen atom, one of the minute particles of which the substance oxygen is composed, has a diameter of about 1.3 A. If you could place oxygen atoms adjacent to one another, you could line up over 75 million of them in 1 cm.)

The **kilogram (kg)** is the SI base unit of mass, equal to about 2.2 pounds. This is an unusual base unit in that it contains a prefix. In forming other SI mass units, prefixes are added to the word gram (g) to give units such as the milligram (mg; 1 mg = 10^{-3} g).

The **second (s)** is the SI base unit of time. Combining this unit with prefixes such as milli-, micro-, nano-, and pico-, you create units appropriate for measuring very rapid events. The time required for the fastest chemical processes is about a picosecond, which is on the order of how fast supercomputers can perform a single calculation.

When you measure times much longer than a few hundred seconds, you revert to minutes and hours, an obvious exception to the prefix–base format of the International System.

Temperature is difficult to define precisely, but we all have an intuitive idea of what we mean by it. It is a measure of “hotness.” A hot object placed next to a cold one becomes cooler, while the cold object becomes hotter. Heat energy passes from a hot object to a cold one, and the quantity of heat passed between the objects depends on the difference in temperature between the two. Therefore, temperature and heat are different, but related, concepts.

A thermometer is a device for measuring temperature. The common type consists of a glass capillary containing a column of liquid whose length varies with temperature. A scale alongside the capillary gives a measure of the temperature. The **Celsius scale** (formerly the centigrade scale) is *the temperature scale in general scientific use.* On this scale, the freezing point of water is 0°C and the boiling point of water at normal barometric pressure is 100°C. However, *the SI base unit of temperature* is the **kelvin (K),** a unit on an *absolute temperature* scale. (See the first margin note on the next page.) On any absolute scale, the lowest temperature that can be attained theoretically is zero. The Celsius and the Kelvin scales have equal-size units (that is, a change of 1°C is equivalent to a change of 1 K), where 0°C is a temperature equivalent to 273.15 K. Thus, it is easy to convert from one scale to the other, using the formula: $T_K = (t_C \times \frac{1K}{1^\circ C}) + 273.15K$

where T_K is the temperature in kelvin and t_C is the temperature in degrees Celsius. A temperature of 20°C (about room temperature) equals 293 K.

Volume. Volume is defined as length cubed and has the SI unit of cubic meter (m^3). This unit is too large a unit for normal laboratory work, so we use either cubic decimeters (dm^3) or cubic centimeters (cm^3, also written cc). Traditionally, chemists have used the liter (L), which is a unit of volume equal to a cubic decimeter (approximately one quart).

In fact, most laboratory glassware is calibrated in liters or milliliters (1000 mL = 1 L). Because 1 dm equals 10 cm, a cubic decimeter, or one liter, equals $(10\ cm)^3 = 1000\ cm^3$. Therefore, a milliliter equals a cubic centimeter. In summary

$$1\ L = 1\ dm^3 \text{ and } 1\ mL = 1\ cm^3$$

Density. The density of an object is its mass per unit volume. You can express this as $\rho = \frac{m}{V}$

where ρ is the density, m is the mass, and V is the volume. Suppose an object has a mass of 15.0 g and a volume of 10.0 cm^3. Substituting, you find that

$$\rho = \frac{m}{V} = \frac{15g}{10cm^3} = 1.50 g / cm^3$$

The density of the object is 1.50 g/cm^3 (or 1.50 $g*cm^{-3}$).

Density is an important characteristic property of a material. Water, for example, has a density of 1.000 g/cm^3 at 4°C and a density of 0.998 g/cm^3 at 20°C. Lead has a density of 11.3 g/cm^3 at 20°C. Oxygen gas has a density of $1.33*10^{-3}$ g/cm^3 at normal pressure and 20°C. (Like other gases under normal conditions, oxygen has a density that is about 1000 times smaller than those of liquids and solids.) Because the density is characteristic of a substance, it can be helpful in identifying it.

Density can also be useful in determining whether a substance is pure. Consider a gold bar whose purity is questioned. The metals likely to be mixed with gold, such as silver or copper, have lower densities than gold. Therefore, an adulterated (impure) gold bar can be expected to be far less dense than pure gold.

Question 2: A colorless liquid, used as a solvent (a liquid that dissolves other substances), is believed to be one of the following:

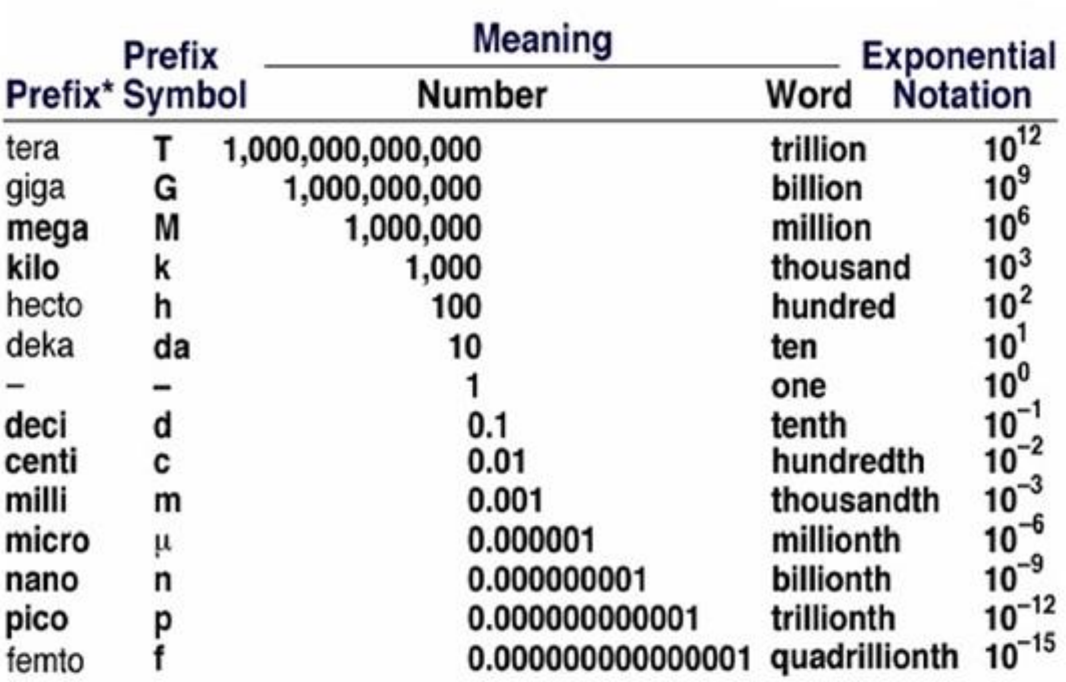
Substance	***Density (in g/mL)***
n-butyl alcohol	0.810
ethylene glycol	1.114
isopropyl alcohol	0.785
toluene	0.866

To identify the substance, a chemist determined its density. By pouring a sample of the liquid into a graduated cylinder, she found that the volume was 35.1 mL. She also found that the sample weighed 30.5 g. What was the density of the liquid? What was the substance?

Solution: You substitute 30.5 g for the mass and 35.1 mL for the volume into the equation: $\rho = \frac{m}{V} = \frac{30.5g}{35.1\ mL} = 0.869 g / mL$

The density of the liquid equals that of **toluene.**

Table 3 - Common demical prefixes used with SI Units.



Prefix*	Prefix Symbol	Meaning		Exponential Notation
		Number	Word	
tera	T	1,000,000,000,000	trillion	10^{12}
giga	G	1,000,000,000	billion	10^{9}
mega	M	1,000,000	million	10^{6}
kilo	k	1,000	thousand	10^{3}
hecto	h	100	hundred	10^{2}
deka	da	10	ten	10^{1}
–	–	1	one	10^{0}
deci	d	0.1	tenth	10^{-1}
centi	c	0.01	hundredth	10^{-2}
milli	m	0.001	thousandth	10^{-3}
micro	µ	0.000001	millionth	10^{-6}
nano	n	0.000000001	billionth	10^{-9}
pico	p	0.000000000001	trillionth	10^{-12}
femto	f	0.000000000000001	quadrillionth	10^{-15}

LABORATORY WORK
CHEMICAL THERMODYNAMICS

Objective: to determine the thermal effect of neutralization reaction of strong acid by strong base by calorimetric method
Materials Required: Calorimeter, Thermometer, Stirrer
Reagents: acids – HCl, HNO_3, H_2SO_4, base – NaOH

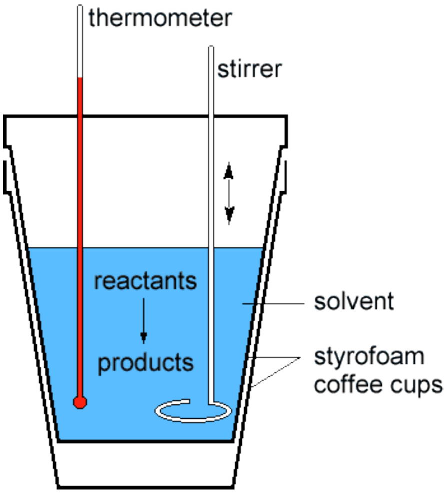
Thermodynamics is the branch of chemistry which deals with all changes in energy or transfers of energy which accompany chemical or physical changes.

Chemical thermodynamics is the branch of physical chemistry which deals with transport of heat either as a result of chemical change or physical change.

The thermal effect of chemical reaction is the amount of heat which absorbed or released as a result of this reaction.

The heat of neutralization of an acid is defined as the amount of heat evolved when one equivalent of an acid and one equivalent of a base undergo a neutralization reaction to form water and a salt. Similarly the heat of neutralization of a base is the amount of heat evolved when 1 g equivalent of the base is completely neutralised by a strong acid in a dilute solution:

$HA + BOH = AB + H_2O + 57{,}3$ kJ



For a constant pressure (P=const), **calorimeter** measures the heat effects of variety of reactions such as neutralisation reactions, heat of solution and heat of dilutions. A coffee cup calorimeter is basically constructed from a polystyrene cup with a lid, in which, the cup is filled with a known amount of water and a thermometer inserted measures the heat changes associated with the reaction.

Heat Q is given the system and consumed to increase the internal energy ΔU and to perform work against external forces A:

$$\mathbf{Q = \Delta U + A}$$

For isobar-isothermal process: p, T = const $A_{gas} = p \cdot \Delta V$

$\Delta U = U^2_{final} — U^{1i}{}_{nitial}$

$\Delta V = V^2_{final} – V^1{}_{initial}$

$Q = \Delta U + A$

$Q = (U^2 — U^2) + (pV^2 — pV^1)$

$Q = (U^2 + pV^2) — (U^1 + pV^1)$

U + pV = H – Enthalpy function

Hess's law (the law of constant heat summation). This law was presented by Hess in 1840. According to this law "*If a chemical reaction can* *be made to take place in a number of ways in one or in several steps, the total enthalpy change (total heat change) is always the same, i.e. the total enthalpy change is independent of intermediate steps involved in the change*". The enthalpy change of a chemical reaction depends on the initial and final stages only.

$$Q = -\Delta H$$

The heat of reaction (or enthalpy of reaction) is actually the difference between the enthalpies of the products and the reactants when the quantities of the reactants indicated by the chemical equation have completely reacted. Mathematically, Enthalpy of reaction (heat of reaction):

$$\Delta H_{chem\ reaction} = \Sigma \Delta H^{\circ}_{products} - \Sigma \Delta H^{\circ}_{reac\tan ts}$$

$$aA + bB = cC + dD$$

$$\Delta H^{\circ}_{A} \quad \Delta H^{\circ}_{B} \quad \Delta H^{\circ}_{C} \quad \Delta H^{\circ}_{D}$$

$$\Delta H_{chem\ react} = (\Delta H^{\circ}_{D} \cdot d + \Delta H^{\circ}_{C} \cdot c) - (\Delta H^{\circ}_{B} \cdot b + \Delta H^{\circ}_{A} \cdot a)$$

Applications of Hess's law:

a) For the determination of enthalpies of formation of those compounds which cannot be prepared directly from the elements easily using enthalpies of combustion of compounds.

b) For the determination of enthalpies of extremely slow reactions.

c) For the determination of enthalpies of transformation of one allotropic form into another.

d) For the determination of bond energies:

$\Delta H_{reaction}$ = Σ Bond energies of reactants – Σ Bond energies of products.

e) For the determination of resonance energy.

f) For the determination of lattice energy.

If the reaction is generated heat ($Q>0$), enthalpy of the system is lowered ($\Delta H < 0$), in this case the reaction is called *exothermic*.

Examples: (1) $C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)} + 393.5$ kJ (at constant temperature and pressure)

or $C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$, $\Delta H = -393.5$ kJ

(2) $H_{2(g)} + 1/2\ O_{2(g)} \rightarrow H_2O_{(l)}$, $\Delta H = -285.8$ kJ

(3) Fermentation is also an example of exothermic reaction.

If the reaction is absorbed heat ($Q<0$), the system increases the enthalpy ($\Delta H > 0$), and it called the *endothermic reaction*.

Example: (1) $N_{2(g)} + O_{2(g)} \rightarrow 2NO_{(g)}$; H = +180.5 kJ

(2) $C_{(s)} + 2S_{(s)} \rightarrow CS_{2(l)}$; H = +92.0 kJ

(3) Preparation of ozone by passing silent electric discharged through oxygen is the example of endothermic reaction.

(4) Evaporation of water is also the example of endothermic reaction.

Factors which influence the heat of reaction: There are a number of factors which affect the magnitude of heat of reaction.

1) Physical state of reactants and products: Heat energy is involved for changing the physical state of a chemical substance. For example in the conversion of water into steam, heat is absorbed and heat is evolved when steam is condensed. Considering the following two reactions:

$$H_{2(g)}+\frac{1}{2}O_{2(g)}=H_2O_{(g)};\ \Delta H=-57.8\ kcal$$

$$H_{2(g)}+\frac{1}{2}O_{2(g)}=H_2O_{(l)};\ \Delta H=-68.32\ kcal$$

It is observed that there is difference in the value of H if water is obtained in gaseous or liquid state. ΔH value in second case is higher because heat is evolved when steam condenses. Hence, physical sate always affects the heat of reaction.

2) Allotropic forms of the element: Heat energy is also involved when one allotropic form of an element is converted into another. Thus, the value of H depends on the allotropic form used in the reaction. For example, the value of H is different when carbon in the form of diamond or in amorphous form is used.

$$C_{\text{(diamond)}} + O_{2(g)} \text{–>} CO_{2(g)},\ \Delta H = –94.3\ \text{kcal}$$

$$C_{\text{(amorphous)}} + O_{2(g)} \text{–>} CO_{2(g)},\ \Delta H = –97.6\ \text{kcal}$$

The difference between the two values is equal to the heat absorbed when 12g of diamond is converted into 12g of amorphous carbon. This is termed as heat of transition.

3) Temperature: Heat of reaction has been found to depend upon the temperature at which reaction is occurring. The variation of the heat of reaction with temperature can be ascertained by using Kirchhoff's equation:

$$\Delta HT_2 – HT_1/T_2 – T_1 = C_P$$

Kirchhoff's equation at constant volume may be given as,

$$ET_2 – ET_1/T_2 – T_1 = C_v$$

Orientation process, or the possibility of chemical reaction is determined by the Gibbs energy. Gibbs energy or the isobaric - isothermal potential – is a manifestation of the cumulative effect of the enthalpy H and entropy S factors:

- At any temperature: $\Delta G_T=\Delta H-T\cdot\Delta S$
- At standard condition (298K): $\Delta G = \Sigma\ \Delta\ G_{\text{products}} - \Sigma\ \Delta\ G_{\text{reactants}}$

If $\Delta G <0$, then the reaction is possible.
If $\Delta G> 0$, then the reaction is impossible (is the reverse reaction).
If $\Delta G = 0$, then the system is in equilibrium.

Basic laws of Thermodynamics:

- Zeroth law: If any two systems are in thermal equilibrium with the third system, then they are also in thermal equilibrium with each other.
- First law: First law of thermodynamic states that energy can neither be created nor be destroyed but it can only be converted from one form to another.

- Second law: This law states that "all processes in nature tend to occur with an increase in entropy and the direction of change always lead to the increase in entropy."
- Third law: This law states that "The entropy of a perfect crystal of each element and a compound is zero at absolute zero."

PROCEDURE:

The following section gives an example of a laboratory determination of the enthalpy change for chemical reaction in which reactants are mixed in water. After the reaction has taken place the final mixture consists of products, unused reactants and water. The principles of such measurement are as follow.

The reaction is carried out in an insulated container which ideally prevents the reaction mixture from losing heat to (or gaining heat from) the surroundings.

The chemical reaction should be rapid, so that the energy change complete in a short period of time. This is achieved by rapidly mixing the reactants.

An exothermic chemical reaction increases the amount of heat energy contained in the reaction mixture. This raises the temperature of the final mixture by ΔT°C, defined as $T_{final} – T_{initial}$. An endothermic reaction decreases the amount of heat energy contained in the reaction mixture, so lowering the temperature of the final mixture.

The change in the amount of heat energy due to chemical reaction, Q joules, is calculated using the equation: **Q = m*C* ΔT**
(note the negative sign) where **m** is mass and **C** (for pure water $4{,}184*10^3$ J/kg*°C = 1 kal/g*°C) the average specific heat capacity of the final mixture (mainly water) whose temperature is being measured. The value of Q is negative if the chemical reaction is exothermic (when ΔT is positive); the value of Q is positive if the chemical reaction is endothermic (when ΔT is negative).

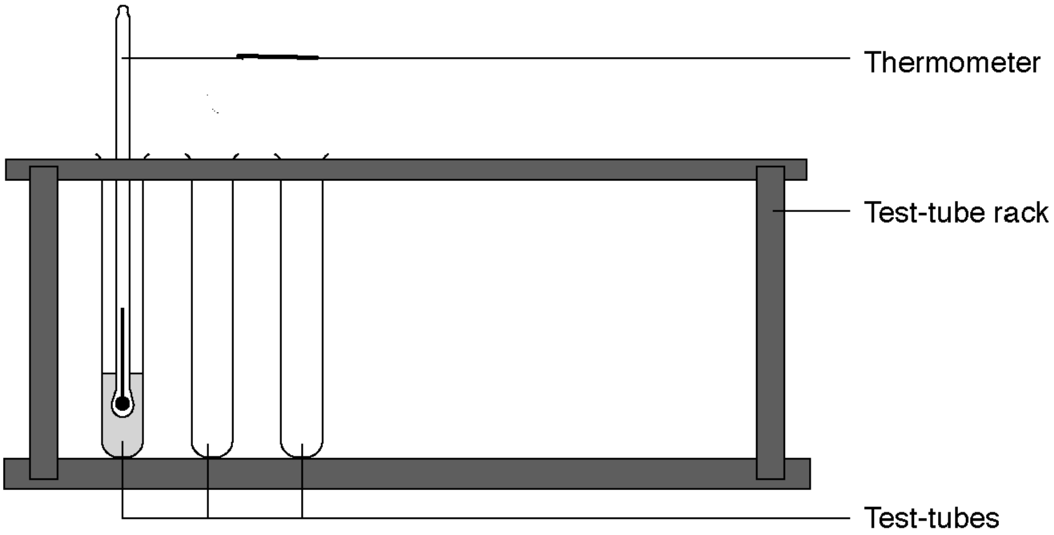
At constant pressure, the heat change undergone produced by the chemical reaction is equal to the enthalpy change of the reaction, i.e. Q = –ΔH. Note that Q and ΔH have the different sign.

***Task to you*:** determine the thermal effect of the neutralization reaction, calculate the enthalpy change, please fill out the table and write thermochemical reaction equations:

$$aA + bB = cC + dD \quad \Delta H>0 \text{ (endothermic reaction)}$$

or

$$aA + bB = cC + dD \quad \Delta H<0 \text{ (exothermic reaction)}$$



1. Put 5 cm^3 of each acid in a test-tube.
2. Record the initial temperature of the each acid in table.
3. Add and mix 5 cm^3 of alkali NaOH.
4. Watch what happens? How the temperature of acid is changed: increases or decreases?
5. Carefully stir, using the thermometer, and record the final temperature again.
6. Calculate the thermal effect of neutralization reaction by formula: **Q = m*C* ΔT**
7. All measurement results tabulated:

Volume of acid (5 ml)	Volume of base (5 ml)	Temperature		Δt, °C	Volume of mixture, ml	The heat effects of reaction, kJ/mol
		T_{final}	$T_{initial}$			
HCl	NaOH					
HNO_3	NaOH					
H_2SO_4	NaOH					

EXERCISE: 1) Use the ΔH°, S°, ΔG° data to calculate next quantities of the following reaction:

- heat effect and type of reaction (exothermic or endothermic)
- entropy change under standard conditions (298K)
- determine the possibility of spontaneous reaction under standard conditions (Gibbs free energy)
- determine the possibility of spontaneous reaction at a temperature of 1500 K
- the temperature at which system is being in the equilibrium

$$2HCl + Mg(OH)_2 = MgCl_2 + 2H_2O$$

ΔH°, kJ/mol	- 92,30	-924,7	- 644,80	- 285,8
S°, J/mol*K	186,80	63,2	89,5	69,9
ΔG°, kJ/mol	- 95,30	- 833, 8	- 595, 3	- 273,2

2) Perform a test:

№	Question	Variants
1	Combustion reactions of fuels is an example of	A) Endothermic reaction B) Exothermic reaction C) Neither Endothermic nor Exothermic D) Both A and B
2	Exothermic reaction is a chemical reaction that	A) releases heat B) absorbs C) can release heat or absorb heat D) None of the above
3	Which of the following option states an endothermic reaction	A) Conversion of Graphite to diamond B) Burning of a substance C) Evaporation of water D) Both A and C
4	Endothermic reaction is a chemical reaction that	A) releases heat B) absorbs C) can release heat or absorb heat D) None of the above
5	Silver chloride(AgCl) turns grey in sunlight. $2AgCl(s) = 2Ag(s) + Cl_2$	A) is an exothermic reaction B) is a decomposition reaction C) is an endothermic reaction D) Both B and C
6	What form of energy is causing the decomposition reactions	A) heat B) light C) electricity D) All of the above

LABORATORY WORK
CHEMICAL KINETICS

Objective: to investigate the effect of changing the concentration and temperature on the reaction rate
Materials Required: tubes and graduated tubes, beakers, thermometer, stopwatch, water bath, glass
Reagents: solutions of 0.5% sulfuric acid H_2SO_4 and 0.5% disodium thiosulfate $Na_2S_2O_3$, distilled water

Chemical kinetics deals with the study of reaction rate. Every chemical reaction occurs at a definite rate under a given set of conditions. Some reactions are very fast and some other reactions are comparatively slow. ***Reaction rate can be defined as,*** "***The change in the concentration of a***

reactant or product per unit time and per unit volume (for homogeneous reactions) or per unit area (for heterogeneous reactions)":

$$\upsilon = \frac{C^1 - C^2}{\tau_1 - \tau_2} = \pm\frac{\Delta C}{\Delta \tau}$$

Thus, chemical reaction speed is the reverse quantity of the reaction time. The rate of reactions is measured in terms of change in concentrations of reactants or products mol/*l**s. For the reaction to be useful, either in the laboratory or in nature, it must occur at a reasonable rate.

Factors affecting reaction rate:

1) Nature of reactants:

a) Physical state of reactants: This has considerable effect over rate of reaction: Gaseous state > Liquid state > Solid state —> Decreasing rate of reaction

b) Physical size of the reactants: Among the solids, rate increases with decrease in particle size of the solid.

c) Chemical nature of the reactants:

- Reactions involving polar and ionic substances including the proton transfer reactions are usually very fast. On the other hand, the reaction in which bonds is rearranged, or electrons transferred are slow.
- Oxidation-reduction reactions, which involve transfer of electrons, are also slow as compared to the ionic substance.
- Substitution reactions are relatively much slower.

2) Effect of temperature: The rate of chemical reaction generally increases on increasing the temperature. The rate of a reaction becomes almost double or tripled for every rise in temperature.

Temperature coefficient (*j*) of a reaction is defined as the ratio of rate constants at two temperatures differing by 10°C:

$$\frac{V_{t_2}}{V_{t_1}} = j^{\frac{t_1 - t_2}{10^\circ}} \qquad \text{or} \qquad \frac{K_{T_2}}{K_{T_1}} = j^{\frac{T_2 - T_1}{10^\circ}}$$

Arrhenius proposed a quantitative relationship between rate constant and temperature as, $k = A \cdot e^{-\frac{E_a}{RT}}$

The equation is called Arrhenius equation. In which constant A is known as frequency factor. This factor is related to number of binary molecular collision per second per litre; E_a is the activation energy; T is the absolute temperature and R is the gas constant. Both A and E_a are collectively known as Arrhenius parameters.

Taking logarithm equation may be written as, **log k = log A – E_a/2.303 RT**

The value of activation energy (E_a) increases, the value of k decreases and therefore, the reaction rate decreases.

The physical meaning of the activation energy (E_a) is that it is the minimum relative kinetic energy which the reactant molecules must possess for changing into the products molecules during their collision. This means that the fraction of successful collision is equal to $e^{-E_a/RT}$ called Boltzmann factor.

3) Concentration of reactants (Law of Mass action): The rate at which a substance reacts is directly proportional to its active mass and the rate at which a reaction proceeds is proportional to the product of the active masses of the reacting substances:

$$aA + bB = cC + dD$$

$$V_1 = k_1 \cdot [A]^a \cdot [B]^b$$

$$V_2 = k_2 \cdot [C]^c \cdot [D]^d$$

• For a reaction, aA + bB –> product

Rate = (dx/dt) $\propto [A]^a[B]^b$; (dx/dt) = $k[A]^a[B]^b$

where k is called **rate constant or velocity constant.**

When [A] = [B] = 1 mol/litre, then dx/dt = k

Thus, rate constant k is also called **specific reaction rate.**

• The value of rate constant depends on, nature of reactant, temperature and catalyst. It is independent of concentration of the reactants.

• Unit of rate constant = $[\text{litre/mol}]^{n-1} \times \text{sec}^{-1} = [\text{mol/litre}]^{1-n} \times \text{sec}^{-1}$

Where n order of reaction.

4) Presence of catalyst: The function of a catalyst is to lower down the activation energy. The greater the decrease in the activation energy caused by the catalyst, higher will be the reaction rate.

5) Effect of sunlight: There are many chemical reactions whose rate are influenced by radiations particularly by ultraviolet and visible light. Such reactions are called photochemical reactions. For example, Photosynthesis, Photography, Blue printing, Photochemical synthesis of compounds etc.

The radiant energy initiates the chemical reaction by supplying the necessary activation energy required for the reaction.

Types of chemical reactions. On the basis of reaction rates, the chemical reactions have been classified into the following three types,

1) Very fast or instantaneous reactions: These reactions occur at a very fast rate generally these reactions involve ionic species and known as ionic reactions. It is almost impossible to determine the rates of these reactions.

Examples:

a) Precipitation reaction:

$$AgNO_{3(aq)} + NaCl_{(aq)} = AgCl\downarrow_{(solid)} + NaNO_{3(aq)}$$

b) Neutralization reaction: $HCl_{(aq)} + NaOH_{(aq)} = NaCl_{(aq)} + H_2O_{(l)}$

2) Moderate reaction: These reactions proceed with a measurable rates at normal temperature and these reactions are studied in chemical kinetics. Mostly these reactions are molecular in nature.

Examples:

a) Decomposition of: H_2O_2 : $2H_2O_2 \to 2H_2O + O_2$

b) Decomposition of: N_2O_5 : $2N_2O_5 \to 2N_2O_4 + O_3$

3) Very slow reactions: These reactions are extremely slow and take months together to show any measurable change.

Examples:

a) Rusting of iron: $Fe_2O_3 + xH_2O \to Fe_2O_3 \cdot xH_2O_{(rust)}$

b) $2H_2 + O_2 \xrightarrow{Room\ \ temperature} 2H_2O$

Chemical Equilibrium. State the irreversible reactions in which the starting materials are fully converted into reaction products, i.e. reaction goes to completion. Signs of the irreversibility:

a) precipitation: $Na_2SO_4 + BaCl_2 \to 2NaCl + BaSO_4\downarrow$;

b) the allocation of gas: $Na_2CO_3 + H_2SO_4 \to Na_2SO_4 + H_2O + CO_2\uparrow$;

c) formation of a weak electrolyte: $2NaOH + H_2SO_4 \to Na_2SO_4 + H_2O$.

State the reversible reactions in which the final products interact to form the starting materials.

Such reactions are not going to end, to a state of equilibrium.

Chemical equilibrium – a state of the system in which speed forward and reverse reactions are equal. Equilibrium is called concentration, which are installed on the equilibrium state (for the initial equilibrium concentration of the substance is the amount of material which remained at the time of equilibrium for the reaction products - is the amount of matter, which was formed at the time of equilibrium).

Chemical equilibrium is characterized by the equilibrium constant K_p, which is the ratio of product concentrations of the reaction products to the product of the concentrations of substances in the initial degrees are stoichiometric coefficients. In accordance with the law of mass action for the reversible reaction:

$$aA + bB \to cC + dD$$

expression of the K_p can be written as follows:

$$K_p = \frac{k_1}{k_2} = \frac{V_2}{V_1} = \frac{[C]^c \cdot [D]^d}{[A]^a \cdot [B]^b}$$

Thermodynamic equilibrium condition: $\Delta G = 0$, $\Delta F = 0$.

K_p shows how many times the rate of direct reaction greater than the rate of reverse reaction:

- If the $K_p > 1$, the faster the direct reaction; $\Delta G < 0$.
- If the $K_p < 1$, then quickly goes back reaction; $\Delta G > 0$.
- If $K_p = 1$, then $\Delta G = 0$ (equilibrium state).

Кр depends on the nature of the reactants and temperature, and does not depend on the concentration of the catalyst.

Displacement of chemical equilibrium - is the transition system from one equilibrium state to another.

Le Chatelier's principle: *if the system is in equilibrium, to produce effects (change the concentration, pressure, temperature), the equilibrium is shifted in the direction of the reaction, which weakens this effect.*

Experiment based on the following reaction between sodium thiosulfate and sulfuric acid. As result of reaction after a certain period of time in solution we observe turbidity of the solution, and then the yellow precipitate sulfur:

$$Na_2S_2O_{3\ (aq)} + H_2SO_{4\ (aq)} = Na_2SO_{4\ (aq)} + H_2O_{(l)} + SO_{2\ (g)}\uparrow + S_{(s)}\downarrow$$

1st Experiment: Investigate the dependence of the rate reaction on the concentration change of reactant.

PROCEDURE:

1) Take four test tubes and sign their as 1st, 2nd, 3rd and 4th, one volumetric tube signed as “**Acid**”

2) Measure by volumetric tube 2 ml $Na_2S_2O_3$ solution and pour into “1st“ test tube, then diluted with 6 ml distilled water

3) Measure by volumetric tube signed as “**Acid**” 2 ml of sulfuric acid H_2SO_4 and mix into a diluted solution of sodium thiosulfate in “1st“ test tube. At the same time, start the stopwatch for fixing a moment time when appear turbidity of the solution, and then the yellow precipitate sulfur in solution

4) Record the measurements in the table and calculate the reaction rate

5) Repeat experiments with subsequent test tubes

	Solution of thiosulfate and water, *ml*		Sulfuric acid volume, *ml*	The relative concentration of the solution: $C_{Na_2S_2O_3} = \frac{V_{Na_2S_2O_3}}{V_{Na_2S_2O_3} + V_{H_2O}}$	Time of reaction (turbidity appears) $\tau;\ s$	The relative rate of reaction $\upsilon = \frac{1}{\tau};\ s^{-1}$
	$V_{Na_2S_2O_3},\ ml$	$V_{H_2O},\ ml$	H_2SO_4			
1	2	6	2			
2	4	4	2			
3	6	2	2			
4	8	0	2			

Task to you: please, show on the graph the dependence of the rate reaction on the relative concentration of disodium thiosulfate in solution in next coordinates $-C_{Na_2S_2O_3}$.

2nd Experiment: Investigate the dependence of the rate reaction on the temperature change of reactant solution.

PROCEDURE:

1) Take four test tubes and sign them as 1st, 2nd, and 3rd, one volumetric tube signed as “**Acid**”.

2) Measure by volumetric tube 8 ml $Na_2S_2O_3$ solution and pour into all test tubes signed as 1st, 2nd and 3rd .

3) Take 1st test tube with thiosulfate solution, insert the thermometer into the test tube, fix the temperature of the solution, and then heat the test tube on a water bath to a temperature of 30°C.

4) At that time, measure by volumetric tube signed as “**Acid**” 2 ml of sulfuric acid H_2SO_4 and mix into a heated solution of sodium thiosulfate in “1st“ test tube. At the same time, start the stopwatch for fixing a moment time when appear turbidity of the solution, and then the yellow precipitate sulfur in solution

5) Record the measurements in the table and calculate the reaction rate

6) Repeat experiments with subsequent test tubes.

№	$Na_2S_2O_3$ thiosulfate solution volume, *ml*	H_2SO_4 sulfuric acid volume, *ml*	Heating temperature of the solution, $t℃$	Time of reaction (turbidity appears) $\tau;\ s$	The relative rate of reaction $\upsilon=\frac{1}{\tau};\ s^{-1}$
1	8	2	30°C		
2	8	2	40°C		
3	8	2	50°C		

Task to you: please, show on the graph the dependence of the rate reaction on the temperature of solution in next coordinates $\upsilon-\frac{1}{t^\circ}$.

3rd Experiment. Altering an Equilibrium System: Le Chatelier’s Principle

1 KSCN and $Fe(NO_3)_3$ form ions in water:

$$KSCN \leftrightarrow K^+ + SCN^-$$

$$Fe(NO_3)_3 \leftrightarrow Fe^{3+} + 3NO_3^-$$

2 Na_2HPO_4 binds strongly to Fe^{3+} ions, essentially removing them from solution.

3 KSCN is toxic in high concentrations. Avoid contact with skin.

Purpose:

- To examine the equilibrium: $\underset{\text{(colourless)}}{Fe^{3+} + SCN^-} \leftrightarrow \underset{\text{(brown)}}{FeSCN^{2+}}$
- To use Le Chatelier’s principle to explain changes in the equilibrium

PROCEDURE:

1 Obtain 4 clean test tubes, a clean 50 mL beaker, and a glass stirring rod.

2 Into the beaker place approximately 40 mL of 0.001 M KSCN.

3 Add four drops of $Fe(NO_3)_3$ solution to the beaker containing the KSCN. Stir. Rinse off stirring rod.

Colour of KSCN __________, Colour of Fe(NO$_3$)$_3$ ______________, Colour when the two are added __________. What product/ion provides the colour? ___________

4 Divide the coloured solution into the 4 tubes (the levels only need to be roughly equal).

5 The first test tube will act as a control (i.e. you will compare the colour of the other tubes to tube #1).

6 To tube #2 add some KSCN(s) (about ¼ - ½ a small spoonful). Stir. Record colour change in chart.

7 To tube #3 add three drops of Fe(NO3)3. Stir with clean glass rod. Record any colour change.

8 To tube #4 add 5-10 small crystals of Na2HPO4. Stir with clean glass rod. Record any colour change.

9 Using any of the chemicals in the lab, see if you can return one of the tubes to the colour of tube 1. See if you can change the colour of one of the tubes from brown to colourless and then back to brown.

10 Dump the contents of test tubes and beakers down the sink. Rinse and return all equipment.

Tube	Addition of	Colour change	Explanation according to Le Chatelier’s principle
2	KSCN		
3	Fe(NO$_3$)$_3$		
4	NaH$_2$PO$_4$		

11. Which of the tubes has a different Kc value than tube #1? ____________
Explain: ____

12. For each of the tubes indicate how the level of each ion was affected (use ↑ for increased, ↓ for decreased, – for no change). Also, predict which chemicals (KSCN, Fe(NO$_3$)$_3$, and/or Na$_2$HPO$_4$), if any, could be used to return the tubes to exactly their original equilibrium concentrations for <u>all</u> three ions (Fe^{3+}, SCN$^{–}$ and FeSCN^{2+}).

Tube	Stress imposed in equilibrium	Change from original concentration (↑, ↓, or –)			Chemical that could exactly restore [ ]s
		Fe^{3+}	SCN$^{–}$	FeSCN^{2+}	
2	↑ SCN$^-$				
3					
4					

EXERCISE: write down mathematical expressions of the speeds of straight and reverse reaction:

$$Al(OH)_3\downarrow_{(s)} + 3HCl_{(aq)} = AlCl_{3\ (aq)} + 3H_2O_{(l)} \qquad \Delta H < 0$$

1) Determine the state of the system (homogeneous or heterogeneous)
2) Predict how many times will increase the speeds of straight reaction if you increase the concentration of hydrochloric acid 3 times
3) Write down a mathematical expression for the equilibrium constant for this reaction
4) Predict how to change the direction of the equilibrium with the following changes in the conditions of the reaction: a) if temperature of system increases; b) if pressure decreases; c) if you decrease the concentration of hydrochloric acid; d) if you decrease the concentration of aluminum chloride.
5) Perform test

№	Question	Variant
1	Chemical kinetics is the branch of chemistry which deals with the study of:	a) Speed or rate of chemical reaction. b) The factors affecting the rates of the reaction. c) The mechanism by which the reactions proceed. d) All of these.
2	The unit of the rate of reaction is:	a) mol L^{-1} min $^{-1}$ b) mol $^{-1}$ L min^{-1} c) Both a and b d) none of these
3	The rate of reaction depends upon the molar concentration of reactants which:	a) Keep on increasing with passage of time. b) Keep on decreasing with passage of time. c) Remains same with passage of time. d) Does not depend upon the time.
4	The factors on which the rate of reaction depends is:	a) Temperature. b) Presence of catalyst. c) Presence of light. d) All of these
5	Activation energy of a chemical reaction can be determined by:	a) Evaluating velocities of reaction at two different temperature. b) Changing concentrations of reactants. c) Evaluating rate constants at two different temperatures. d) Evaluating rate constant at standard temperature.
6	In the Arrhenius equation K = A exp (-E/RT), A may be termed as	a) High temperature. b) Infinite temperature.

	the rate constant at:	c) Low temperature. d) Moderate temperature.
7	The rate of a reaction doubles when the temperature changes from 27°C to 37°C. Calculate the energy of activation.	a) 54.6 kJ mol^{-1} b) 53.6 kJ mol^{-1} c) 52.6 kJ mol^{-1} d) 54.1 kJ mol^{-1}
8	The role of catalyst in a chemical reaction is to change:	a) Equilibrium constant. b) Activation energy. c) Heat of reaction. d) Products of reaction.
9	The ionic reactions are very fast because:	a) These reactions are highly exothermic. b) The energy of interaction between charged ions is greater than between neutral molecules. c) It does not involve bond breaking. d) The number of collisions between ions per unit volume per second is very large.
10	The catalyst used for heterogeneous catalysis is:	a) Solid b) Gas c) Liquid d) None of these
11	Ionic reactions are:	a) Very slow b) Fast c) Slow d) Very fast

LABORATORY WORK
SOLUTIONS

Objective: to prepare a solution and determine its concentration
Materials Required: graduated cylinder and beakers, balance, volumetric flasks of 100 ml, 250 ml, glass rod, filter papers, watch glass, densimeters set (areometer)
Reagents: sodium chloride, sodium carbonate, barium chloride, distiller water

A solution is a homogeneous mixture of two (or more) substances, the composition of which may vary between certain limits.

This concept is valid if we take the example of alcohol and water i.e. both substances are in the same phase. But what about sugar syrup, which contains 60% sugar and 40% water? Will the sugar be the solvent? Answer is No, Still the sugar is solute. Then what is the right definition for this case.

Solvent is that component of the binary mixture which is in same physical state as the solution will be. A solution consisting of two components is called ***binary solution.*** The component which is present in large quantity is called ***solvent*** and the component which is small in quantity is called ***solute***. If both components are in same physical state.

Type of Solutions. All the three states of matter (gas, liquid or solid) may behave either as solvent or solute. Depending on the state of solute or solvent, mainly there may be the following seven types of binary solutions:

Solute	**Solvent**	**Example**
Gas	Gas	Air
Gas	Liquid	Aerated water ($CO_2 + H_2O$)
Gas	Solid	Hydrogen in palladium
Liquid	Liquid	Alcohol in water, benzene in toluene
Liquid	Solid	Mercury in zinc amalgam
Liquid	Gas	CO_2 dissolved in water
Solid	Liquid	Sugar in water, common salt in water
Solid	Gas	Smoke
Solid	Solid	Various alloys

For a given solution, the amount of the solute dissolved in a unit volume of solution (or a amount of solvent) is termed as the **concentration of the solute**. Solutions containing relatively high concentration of solute are called **concentrated solutions** while those of relatively low concentrations of solute are termed as **dilute solutions**.

Solubility of a substance is its maximum amount that can be dissolved in a specified amount of solvent at a specified temperature. It depends upon the nature of solute and solvent as well as temperature and pressure.

Solutions of Solids in Liquids

Solutions of this type are most common. In solutions of solids in liquids, the liquid is invariably referred to as a solvent and the solid dissolved in it as the solute. If a solute is added in small amounts at a time to a given amount of a solvent at a constant temperature, with vigorous stirring of the solvent after each addition, a stage is reached when the added solute no more disappears, i.e. goes into solution but remains undissolved. The solution is then said to be saturated. A solution which remains in contact with undissolved solute is termed as saturated. It can also be defined as one which is in equilibrium with the excess of solid at a particular

temperature.

The amount of solute dissolved in 100g of a solvent to form a saturated solution at a given temperature is termed the solubility of the solute in the given solvent at that temperature. Each substance has a characteristic solubility in a given solvent at a definite temperature.

When a solid is added to the solvent, the particles from the solid diffuse into it. The solute and solvent molecules move constantly in the solution phase. Some of the particles of the solute return to the solid state due to collisions. Thus, two opposite processes operate simultaneously.

Dissolution: Particles of solute leaving the solid and dissolving in the solvent.

Recrystallisation: Solute particles returning to the solid form.

When these two processes move with the same speed, an equilibrium stage is reached.

$$\textbf{Solute (solid)} \rightleftharpoons \textbf{Solute (dissolved)}$$

Thus, a dynamic equilibrium exists in a saturated solution. When a saturated solution prepared at a higher temperature is cooled, it gives a solution which contains usually more of solute than required for the saturated solution at that temperature. Such a solution is referred to as a supersaturated solution. It is usually unstable and changes to saturated solution when excess of solute comes out in solid state.

The following factors influence the solubility of a solid in a liquid:

- **Nature of solute :** The solutes (solids) can be classified as ionic and non-ionic solids. The ionic solids consist of positively and negatively charged ions. It is the force of attraction between the ions, i.e., lattic energy which opposes the tendency of a solute to dissolve. This force of attraction is different in different ionic solids depending on the charges present on the ions and distance between ions (ionic radii). The ionic solutes having high less lattice energy have more solubility. The ions are solvated by the solvent molecules and in this process energy (known as hydration energy) is released. When the hydration energy is high, the ionic solid is more soluble.

Many non-ionic substances dissolve in polar solvents due to hydrogen bonding. Generally, if the solute and solvent have similar characteristics, i.e. both polar or both non-polar, the solubility is high and if both are dissimilar, the solubility is found low.

- **Nature of solvent:** Ionic solids dissolve to a larger extent in a solvent having a high dielectric constant as compared to solvents of low dielectric constants. Dielectric constant of water is 80 while that of methyl alcohol is 33.5 An ionic solid, therefore, dissolves more readily in water than in methyl alcohol. Benzene has a very low dielectric constant of 2.3 and, hence, ionic solids do not dissolve in benzene.

For non-ionic solids, the guiding principle is ‘like dissolves like, i.e., if the solvent is polar, it will dissolve the polar solutes and if it is non-molar, it will dissolve the non-polar solutes in it.

- **Temperature:** The solubility of a solute in a given solvent varies appreciably with temperature.

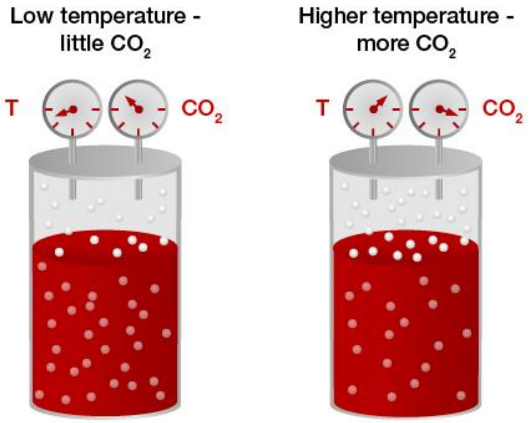
It is observed that the solubility of NaCl increases very slightly with an increase in temperature whereas those of KNO_3, $NaNO_3$, $AgNO_3$, KI, etc., increase greatly. A sharp break in a solubility curve indicates the formation of a compound whose solubility is different from that of the substance from which it has been formed as in the case of $Na_2SO_4 10H_2O$. It losses its water of crystallization at 32.3°C and is converted into anhydrous form. There are few substances like calcium acetate, cerium sulphate, calcium chromate etc., which show a decrease in solubility with rise in temperature.

Generally, solubility depends on heat of solution. If a substance dissolves with absorption of heat, the solubility increases with rise of temperature. On the other hand, if a substance dissolves with evolution of heat, the solubility decreases with rise of temperature.

Solutions of Gases in Liquids

All gases are soluble in water as well as in other liquids to a greater or lesser extent. The solubility of a gas in liquids depends upon the following factors:

- Nature of the gas
- Nature of the solvent
- Temperature
- Pressure



Generally, the gases which can be easily liquefied are more soluble in common solvents. For example, CO_2 is more soluble than hydrogen or oxygen in water. the gases which are capable of forming ions to aqueous solutions are much more soluble in water than in other solvents. Gases like hydrogen chloride (HCl) and ammonia (NH_3) are highly soluble in water but not in organic solvents in which they do not ionize.

The solubility of most gases in liquids decreases with increase of temperature. When a solution of a gas is heated, the gas is usually expelled. However, some gases are more soluble at higher temperature than at lower.

The most important factor which influences the solubility of a gas in liquid is the pressure. the quantitative connection between the solubility and pressure is given by Henry’s law. According to this law, *“The mass of a gas* *dissolved by a given volume of a liquid, at constant temperature, is proportional to the pressure of the gas".* Or *"The solubility of a gas in a liquid is directly proportional to the partial pressure of the gas present above the surface of liquid or solution".* Or *"Mole fraction of gas in the solution is proportional to the partial pressure of the gas over the solution".*

The most commonly used form of ***Henry's law*** *states that "The partial pressure of the gas in vapour phase (P) is proportional to the mole fraction of the gas (X) in the solution"***.** Mathematically,

$$p = K_H X$$

here K_H is the Henry's law constant.

It has been observed that most gases obey Henry's law provided,

- The pressure is not too high.
- The temperature is not too low.

Applications of Henry's law:

- Scuba divers must cope with high concentrations of dissolved gases while breathing air at high pressure underwater. Increased pressure increases the solubility of atmospheric gases in blood. When the divers come towards surface, the pressure gradually decreases. This releases the dissolved gases and leads to the formation of bubbles of nitrogen in the blood. This blocks capillaries and creates a medical condition known as bends, which are painful and dangerous to life. To avoid bends, as well as, the toxic effects of high concentrations of nitrogen in the blood, the tanks used by scuba divers are filled with air diluted with helium (11.7% helium, 56.2% nitrogen and 32.1% oxygen).
- To increase the solubility of CO_2 in soft drinks and soda water, the bottle is sealed under high pressure.
- At high altitudes the partial pressure of oxygen is less than that at the ground level. This leads to low concentrations of oxygen in the blood and tissues of people living at high altitudes or climbers. Low blood oxygen causes climbers to become weak and unable to think clearly, symptoms of a condition known as anoxia.

Solutions of Liquids in Liquids

When one liquid dissolves in another, the molecules of the solvent are caused to move apart so as to accommodate the solute molecules. Similarly, the solute molecules must also be separated so that they can take their places in the mixture. In both these processes energy is required. Finally, as the solute and solvent molecules are brought together, energy is released because of the attractive forces between them. When solute and solvent molecules are strongly attracted to each other, more energy is released in the final step. Three cases may arise under these circumstances. The overall dissolution process results either in evolution of heat or absorption of heat, or energy released in the final step is the same as the absorbed in the first two, i.e., net change is zero.

Examples:

- Benzene and carbon tetrachloride : No evolution or absorption of Heat.
- Acetone and water: Evolution of heat.
- Ethyl alcohol and water: Absorption of heat.

A liquid may or may not be soluble in another liquid. Depending upon the relative solubility of a liquid in another, the following three cases are possible.

- Liquids that are completely miscible. Examples: Benzene and toluene; Ethyl alcohol and water; carbon tetrachloride and benzene.
- Liquids that are partially miscible. Examples: Ether and water; Phenol and water; Nicotine and water.
- Liquids that are practically immiscible. Examples: Benzene and water; carbon tetrachloride and water; Benzene and alcohol.

Colloids

A **colloid** is a dispersion of particles of one substance (the dispersed phase) throughout another substance or solution (the continuous phase). Fog is an example of a colloid: it consists of very small water droplets (dispersed phase) in air (continuous phase). A colloid differs from a true solution in that the dispersed particles are larger than normal molecules, though they are too small to be seen with a microscope. The particles range from about $1*10^3$ pm to about $2*10^5$ pm in size.

Although a colloid appears to be homogeneous because the dispersed particles are quite small, it can be distinguished from a true solution by its ability to scatter light. The scattering of light by colloidal-size particles is known as the **Tyndall effect.** For example, the atmosphere appears to be a clear gas, but a ray of sunshine against a dark background shows up many fine dust particles by light scattering. Similarly, when a beam of light is directed through clear gelatin (a colloid, not a true solution), the beam becomes visible by the scattering of light from colloidal gelatin particles. The beam appears as a ray passing through the solution (Figure 2). When the same experiment is performed with a true solution, such as an aqueous solution of sodium chloride, the beam of light is not visible.

Types of Colloids. Colloids are characterized according to the state (solid, liquid, or gas) of the dispersed phase and of the continuous phase. Table 4 lists various types of colloids and some examples of each. Fog and smoke are **aerosols,** which are liquid droplets or solid particles dispersed throughout a gas. An **emulsion** consists of liquid droplets dispersed throughout another liquid (as particles of butterfat are dispersed through homogenized milk). A **suspension** consists of solid particles dispersed in a liquid.

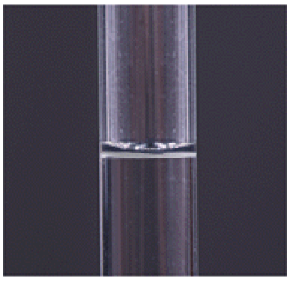
Table 4 - Types of Colloids

Continuous phase	**Dispersed phase**	**Name**	**Example**
Gas	Liquid	Aerosol	Fog, mist
Gas	Solid	Aerosol	Smoke
Liquid	Gas	Foam	Whipped cream
Liquid	Liquid	Emulsion	Mayonnaise (oil dispersed in water)
Liquid	Solid	Suspension	AgCl(*s*) dispersed in H_2O
Solid	Gas	Solid foam	Pumice, plastic foams
Solid	Liquid	Gel	Jelly, opal (mineral with liquid inclusions)
Solid	Solid	Solid sol	Ruby glass (glass with dispersed metal)

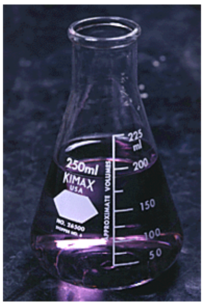
FOR PREPARING THE SOLUTION WE NEED:

A volumetric flask is used to make up a solution of fixed volume very accurately. This volumetric flask measures 500 mL ± 0.2 mL. This is a relative uncertainty of 4 x 10^{-4} or 400 parts per million.

To make up a solution, first dissolve the solid material completely, in less water than required to fill the flask to the mark.

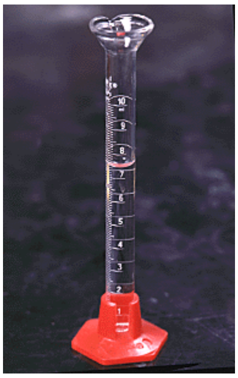


After the solid is completely dissolved, very carefully fill the flask to the 500 mL mark. Move your eye to the level of the mark on the neck of the flask and line it up so that the circle around the neck looks like a line, not an ellipse. Then add distilled water a drop at a time until the bottom of the meniscus lines up exactly with the mark on the neck of the flask. Take care that no drops of liquid are in the neck of the flask above the mark.



After the final dilution, remember to mix your solution thoroughly, by inverting the flask and shaking.

Erlenmeyer flasks and beakers are used for mixing, transporting, and reacting, but not for accurate measurements. The volumes stamped on the sides are approximate and accurate to within about 5%.



Graduated cylinders are useful for measuring liquid volumes to within about 1%. They are for general purpose use, but not for quantitative analysis. If greater accuracy is needed, use a pipet or volumetric flask.

Using an Analytical Balance

Turn the balance on by pressing the control bar. The display lights up for several seconds, then resets to 0.0000.

Place creased, small weighing paper on the balance pan.

Close the sliding glass doors. Wait for the green dot on the left to go out. This is the stability indicator light, indicating that the weight is stable.

Press the control bar to cancel out the weight of the container or paper. The display will again read 0.0000.

Carefully add the substance to be weighed up to the desired mass. Do not attempt to reach a particular mass *exactly*.

Before recording the mass, close the glass doors and wait until the stability detector lamp goes out. Record mass of solid.

Solution Definition: A solution is a homogeneous mixture of two or more substances which composed of only one phase. But a solution may exist in any phase.

Aqueous Solution Definition: An aqueous solution is any solution in which water (H_2O) is the solvent.

Solvent Definition: The component of a solution that is present in the greatest amount. It is the substance in which the solute is dissolved.

Solute Definition: The substance that is dissolved in a solution. For solutions of fluids, the solvent is present in greater amount than the solute.

$$m_{s\ln} = m_{solute} + m_{solvent}$$

$$m_{solvent} > m_{solute}$$

$$V_{s\ln} = \frac{m_{s\ln}}{\rho} \quad \Rightarrow \quad m_{s\ln} = V_{s\ln} \cdot \rho$$

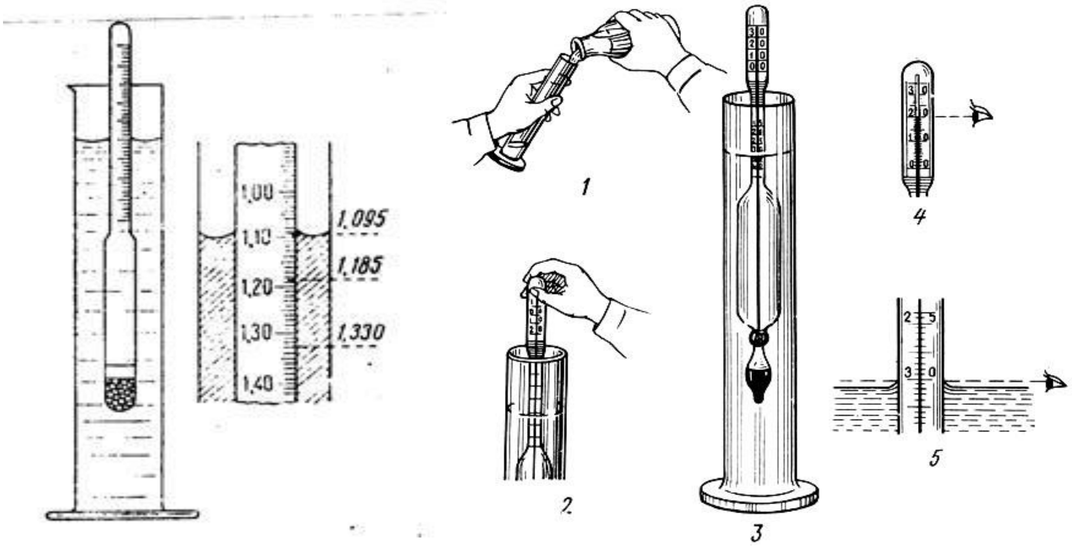


Figure 2 - How to use a densimeter (The density hydrometer measures the density of light or heavy liquid)

PROCEDURE:

1) Weighed in the balance the just about ***m*** grams of chemically pure substance ($\boldsymbol{m}_{\text{solute}}$).

2) Fill the volumetric flask of 250 ml about halfway with distilled water or deionized water ($\boldsymbol{V}_{\text{sln}}$ = 250 ml).

3) Transfer the solid salt to the volumetric flask.

4) Rinse the weighing dish with the water to make certain all of the solute is transferred into the flask.

5) Stir the solution until the solute is dissolved. You may need to add more water (solvent) or apply heat to dissolve the solid.

6) Fill the volumetric flask to the mark (meniscus) with distilled or deionized water.

7) Pour the obtained solution into the cylinder to the mark (meniscus) and measure its density by areometer. Write the measurement result (ρ_{sln}, *g/ml*)

8) Using the results, calculate the percentage, molarity, normality, molality concentrations and the titer of the solution.

9) Record the measurements into the table:

№	***Solute***	m_{solute}, g	V_{sln}, *ml*	ρ_{sln}, g/ml	m_{sln}, g	m_{svt}, g	C%	C_M, *mol/l*	C_N, *mol/l*	C_m, mol/kg	T, *g/ml*
1	NaCl	22,67	250								
2	Na_2CO_3	10,40	250								
3	$BaCl_2$	15,80	250								

Scientists use units of concentration to describe the amount of a chemical substance (solute) dissolved in a given volume or mass of liquid. There are many different units for this purpose, including percent by weight or volume, molarity, molality, normality and titer.

When two solutions of the same chemical but having different concentrations are combined, the concentration of the resulting mixture will be different from either of the two starting solutions. You can calculate the concentration of the final mixture using a mathematical formula involving the volumes of the two combined solutions, as well as the initial concentrations of the two solutions.

1) **Percent composition by mass** is the mass of the solute divided by the mass of the solution, multiplied by 100 (%):

$$C_{\%} = \frac{m_{solute}}{m_{s\ln}} \cdot 100\% = \frac{m_{solute}}{m_{solute} + m_{solvent}} \cdot 100\% = \frac{m_{solute}}{V_{s\ln} \cdot \rho} \cdot 100\%$$

2) **Molarity** is the number of moles of solute per liter of solution (*mol/l*):

$$C_M = \frac{\nu}{V} = \frac{m_{solute}}{M \cdot V_{s\ln}}$$

3) **Normality** is equal to the gram equivalent weight of a solute per 1 liter of solution (*mol*eq/l*):

$$C_N = \frac{m_{solute}}{Eq \cdot V_{s\ln}}$$

4) **Molality** is the number of moles of solute per 1 kilogram of solvent (*mol/kg*):

$$C_m = \frac{\nu \cdot 1000}{m_{solvent}} = \frac{m_{solute} \cdot 1000}{M_{solute} \cdot m_{solvent}}$$

5) **Titer** is equal to the gram of a solute per 1 milliliter of solution (*g/ml*):

$$T = \frac{m_{solute}}{V_{sln}\ (ml)} = \frac{C_N \cdot Eq}{1000}$$

EXERCISES:

1) Calculate the molarity, molality, normality concentrations and titer of 25% aluminum sulfate solution with a density of 1,025g/ml

2) How many grams of salt and water necessary for preparing 650 ml of a 10% solution with a density of 1,005g / ml?

3) How much water should be added to a solution of 40% nitric acid with volume 100 ml and density 1,307g / ml to obtaining a 15% solution of this acid?

4) To neutralize 25 ml of sulfuric acid solution required 0.02N 45.5 ml of sodium hydroxide solution. Calculate the normality and the titer of sulfuric acid solution.

5) An aqueous solution is 8.50% ammonium chloride, NH_4Cl, by mass. The density of the solution is 1.024 g/mL. What are the molality, mole fraction, and molarity of NH_4Cl in the solution?

6) An aqueous solution is 27.0% lithium chloride, LiCl, by mass. The density of the solution is 1.127 g/mL. What are the molality, mole fraction, and molarity of LiCl in the solution?

7) A sample of potassium aluminum sulfate 12-hydrate, $KAl(SO_4)_2*12H_2O$, containing 118.6 mg is dissolved in 1.000 L of solution. Calculate the following for the solution:

a. The molarity of $KAl(SO_4)_2$.

b. The molarity of SO_4^{2-}.

c. The molality of $KAl(SO_4)_2$, assuming that the density of the solution is 1.00 g/mL.

8) A sample of aluminum sulfate 18-hydrate, $Al_2(SO_4)_3*18H_2O$, containing 159.3 mg is dissolved in 1.000 L of solution. Calculate the following for the solution:

a. The molarity of $Al_2(SO_4)_3$.

b. The molarity of SO_4^{2-}.

c. The molality of $Al_2(SO_4)_3$, assuming that the density of the solution is 1.00 g/mL.

9) A solution is made up by dissolving 15.0 g $MgSO_4*7H_2O$ in 100.0 g of water. What is the molality of $MgSO_4$ in this solution?

10) A solution is made up by dissolving 15.0 g $Na_2CO_3*10H_2O$ in 100.0 g of water. What is the molality of Na_2CO_3 in this solution?

11) An aqueous solution is 15.0% by mass of copper (II) sulfate pentahydrate, $CuSO_4*5H_2O$. What is the molarity of $CuSO_4$ in this solution at 20°C? The density of this solution at 20°C is 1.167 g/mL.

12) An aqueous solution is 20.0% by mass of sodium thiosulfate pentahydrate, $Na_2S_2O_3$*$5H_2O$. What is the molarity of $Na_2S_2O_3$ in this solution at 20_C? The density of this solution at 20°C is 1.174 g/mL.

13) Perform test

№	Question	Variant
1	Which of the following factors does not affact the solubility of solids in liquid to large extent?	a) Nature of solvent b) Temperature c) Nature of solute d) Pressure
2	A solution which remains in contact with undissolved solute is termed as	a) ideal solution b) non-ideal solution c) saturated solution d) unsaturated solution
3	Which of the following liquid pairs is completely miscible in each other?	a) Benzene and water b) Carbon tetrachloride and water c) Benzene and alcohol d) Ethyl alcohol and water
4	To increase the solubility of CO_2 in soft drinks and soda water, the bottle is sealed under …	a) high pressure and low temperature b) high pressure and high temperature c) low pressure and high temperature d) low pressure and low temperature
5	Molarity, normality and molality of a solution is expressed as:	a) the number of moles of a solute present in one litre of the solution. b) the number of moles of a solute present in 1000 gm of the solvent. c) the number of gram equivalent of solute present in one litre of solution. d) the ratio of the number of moles of solute to the total number of moles of solute.
6	Which of the following is not a volatile substance?	a) Camphor b) Petrol c) Acetone d) Acetanilide
7	If 50 ml of 0.50 M NaCl solution is diluted with water to a volume of 500 ml the new concentration of solution is:	a) 0.16 M b) 0.05 M c) 0.08 M d) 0.04 M
8	The value of 0.03 M $Ca(OH)_2$ required to neutralise 20 ml of 0.025 M H_3PO_4 is	a) 25 ml. b) 50 ml. c) 40 ml. d) 55 ml.

9	The molarity of 4.6 N H_2SO_4 solution is:	a) 0.46 M. b) 0.23 M. c) 4.6 M. d) 2.3 M.
10	Which of the following pair of liquids are immiscible?	a) Acetone + water. b) Benzene + water. c) Ethanol + water. d) Acetic acid + water.
11	Find the molarity of 0.585g NaCl present in 500 ml of solution.	a) 0.2 M. b) 0.01 M. c) 0.1 M. d) 0.02 M.
12	Unit of molarity is:	a) Kg / litre. b) mol / litre. c) gm / litre. d) none of these.

LABORATORY WORK
COLLIGATIVE PROPERTIES OF SOLUTIONS

Objective: *to investigate the colligative properties of non-electrolytes and electrolytes solutions*

When a non-volatile (substances that do not readily form vapors), non-electrolytic (substances that do not form ions and do not conduct electricity when placed in water) solute such as sugar is dissolved in a given volume of solvent to form a sugar solution, it changes the set of properties of the pure solvent entirely. In this regard, the set of properties such as freezing point, boiling point, vapor pressure, and osmotic pressure of a solvent are affected by the presence of the solute particles in the solution. This set of properties will depend only on the number of dissolved particles in the solution and not on their identity. These properties are collectively known as *colligative properties of solution.*

This module will focus on four colligative properties of electrolyte and non-electrolyte solutions namely:

- *vapor pressure lowering,*
- *boiling point elevation,*
- *freezing point depression,*
- *osmotic pressure.*

The Colligative Properties of Solution	**Mathematic equation for**	
	Non-electrolytes Solution	**Electrolytes Solution**
Vapor pressure lowering (1st Raoult's Law)	$\frac{P^\circ - P}{P^\circ} = N_{X_{s-te}} = \frac{n_{s-te}}{n_{s-te}+n_{svt}}$	$\frac{P^\circ - P}{P^\circ} = i \cdot N_{X_{s-te}} = i \cdot \frac{n_{s-te}}{n_{s-te}+n_{svt}}$

Boiling point elevation (2nd Raoult's Law)	$t_{boil\,s\ln} = t_{boil\,svt} + \Delta t_{boil}$ $\Delta t_{boil} = C_m \cdot E_{svt}$	$t_{boil\,s\ln} = t_{boil\,svt} + \Delta t_{boil}$ $\Delta t_{boil} = i \cdot C_m \cdot E_{svt}$
Freezing point depression (2nd Raoult's Law)	$t_{freez\,s\ln} = t_{freez\,svt} - \Delta t_{freez}$ $\Delta t_{freez} = C_m \cdot K_{svt}$	$t_{freez\,s\ln} = t_{freez\,svt} - \Delta t_{freez}$ $\Delta t_{freez} = i \cdot C_m \cdot K_{svt}$
*Osmotic pressure (*Van't Hoff law)	$P_{osm} = C_M \cdot R \cdot T$	$P_{osm} = i \cdot C_M \cdot R \cdot T$

There are:

- ***E***$_{svt}$ *is boiling point elevation constant, equivalent to 0.52℃/m for aqueous solutions*. This means that, for example, 1 mole of sugar (non-electrolyte) in 1 kilogram of water will increase the boiling point from 100°C to 100,52°C. And C_m is the molal concentration of solute.
- ***K***$_{svt}$ *is the freezing point depression constant equivalent to -1,86°C/m for aqueous solutions.* Again, for example, 1 mole of sugar (non-electrolyte) in 1 kilogram of water will decrease the freezing point from 0°C to -1,86°C.
- *The* ***i*** *Van't Hoff factor gives the number of particles per formula unit of the electrolyte solute.* The degree of dissociation is associated with an isotonic factor by next ratio:

$$\alpha = \frac{i-1}{n-1} \quad \Rightarrow \quad i = \alpha \cdot (n-1) + 1$$

$$Al_2(SO_4)_3 \Leftrightarrow 2Al^{3+} + 3SO_4^{2-}$$

$$n = 2\ kation + 3\ anion = 5$$

1st Experiment: *Investigate and classify colligative and non-colligative properties of solution*

Materials and Reagents: 1% (w/v) brown sugar solution, 0.5 % (w/v) salt solution (NaCl), thermometer

Procedure:

Prepare the brown sugar and salt solutions of concentration 1% and 0.5% respectively. Observe the following properties in the two solutions and identify whether the property observed is a colligative or non-colligative property based on the definition.

№	**Properties of Solution**	**0.5 % (w/v) salt solution**	**1% (w/v) brown sugar solution**
1	Solubility (ability of the solute to be dissolved in a solvent)		
2	Viscosity of solution ("viscosity" is a property of solution to resist flow)		
3	Color of the solution		
4	Taste of solution		

5	Temperature upon boiling		
6	Temperature upon freezing		

Analysis:

1. Which of the following properties of solution can be considered colligative? Why?

2. Which of the following properties of solution can be considered non-colligative?

Justify your answer.

2nd Experiment: *Investigate the vapor-pressure lowering of non-electrolyte sugar solution*

Materials and Reagents: water, aqueous sugar solution, two glasses, one sealed enclosed container

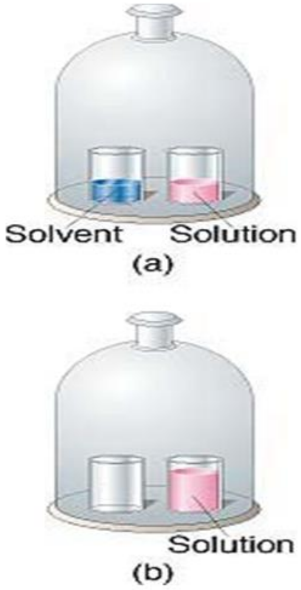


Figure 1.1 Experiment set-up for vapor-pressure lowering

Procedure:

1. Get two glasses placed side by side in a sealed enclosed container. One glass contains pure water, the other an equal volume of an aqueous solution of sugar. Take note of the volume of water and sugar solution. (See figure 3.1a)

2. Leave the set-up until the next day. Gradually measure the volume of the sugar solution and that of the pure water. Is there any change in their original volume? (See figure 3.1b).

After finishing the activity, you should now see why the vapor pressure of the solvent decreases upon addition of a non-volatile, non-electrolyte solute (sugar).

Analysis:

1. What can you infer about the change in volume of the sugar solution and that of pure water?

2. What has caused this change in volume of the sugar solution and that of pure water?

3rd Experiment: *Investigate the Boiling Point Elevation*

Materials and Reagents: three eggs, 1 tea spoon of salt, 1 tea spoon of sugar, boiling water.

PROCEDURE:

1. Put the first egg in two cups of water and take note of the time until the water boils.

2. Again, using two cups of water, put the second egg in the water and add 1 tbsp of salt. Record the time it will take the water to boil.

3. Repeat step 2, but add 1 tbsp of sugar instead of salt. Again, take note of the time it will take the water to boil.

4. Compare the time it will take for water to boil and cook the hard-boiled egg in step 1 to step 3. Record your observations.

Analysis:

1. Which set-up took less time to cook hard-boiled eggs? Why?
2. Did the water take more time to boil when electrolyte (NaCl) solute was added? Why?
3. Did the water take more time to boil upon addition of non-electrolyte (sugar) solute? Justify your answer.

4th Experiment: *Investigate the Freezing Point Depression*

Materials and Reagents: one tsp. salt, one glass of water, one glass of crushed Ice, stirrers
Spoon, thermometer

PROCEDURE:

1. Using one glass of water and one glass of crushed ice, stir the mixture; then using a thermometer, observe and record this temperature.
2. Add one tsp. of salt to the water/ice mixture, then observe and record the temperature. You should repeat this procedure until the temperature reaches 10^oC. More ice should be added if it is necessary.

Analysis:

1. What happens to water and ice when salt is added to this mixture?
2. What happens to the temperature when salt is added to the mixture?
3. What variables would cause these differences?

EXERCISES:

1) What is the osmotic pressure of the solution containing 0.2 M sugar ($C_{12}H_{22}O_{11}$) solution at 25^oC?

2) At what temperature will the sugar solution boil and freeze if 20 g sucrose ($C_{12}H_{22}O_{11}$) is added to 1.5 kg of water?

3) What is the boiling and freezing points of a 0.20 m NaCl solution in water? For water, E is 0.512^oC*kg/mol; K is 1.86^oC*kg/mol

4) Choose the letter of the best answer. Write the chosen letter on a separate sheet of paper:

№	Question	Answer
1	A substance whose water solution conducts an electric current is called a(n)	a. electrolyte b. nonelectrolyte c. polar substance d. nonpolar substance
2	A substance that vaporizes easily is known as	a. volatile b. nonvolatile c. electrolyte d. nonelectrolyte
3	Adding a solute such as NaCl to water increases its	a. boiling point b. melting point c. osmotic pressure

		d. vapor pressure
4	Which of the following is NOT a colligative property of solution?	a. solubility b. boiling point c. melting point d. osmotic pressure
5	Which of the following statements is TRUE about vapor pressure?	a. The vapor pressure of the solvent is less than the solution. b. The vapor pressure of the solution is higher than the pure solvent. c. The vapor pressure of the pure solvent is higher than the solution. d. The vapor pressure of the pure solvent and the solution are the same.
6	The i factor gives the number of particles per formula unit of the solute. What is the *i* factor for NaCl?	a. 0 b. 1 c. 2 d. 3
7	What happens during osmosis?	a. Pure solvent diffuses through a membrane but solutes do not. b. Pure solutes diffuse through a membrane but solvent does not. c. Gases diffuse through a membrane into a solution and build up pressure. d. Pure solvent and a solution both diffuse at the same time through a membrane.
8	What is the temperature at which the vapor pressure equals the atmospheric pressure?	a. 100^oC b. freezing point c. boiling point d. melting point
9	The freezing point of a solution is always ___________ the pure solvent.	a. less than b. same as c. greater than d. equal to the freezing point of solvent
10	What is the van't hoff factor for $CaCl_2$?	a. 0 b. 1 c. 2 d. 3
11	According to 1st Raoult's law, which statement is FALSE?	a. The vapor pressure of a solvent over a solution decreases as its mole fraction increases. b. The solubility of a gas increases as the temperature decreases. c. The vapor pressure of a solvent over a

		solution is less than that of pure solvent. d. The greater the pressure of a gas over a solution the greater its solubility.
12	Which of the following statement is TRUE about the boiling point of the solution?	a. The boiling point of the solvent is greater than that of the solution. b. The boiling point of the solution is less than that of the pure solvent. c. The boiling point of the solution is the same as that of the pure solvent. d. The boiling point of the solution is greater than that of the pure solvent.
13	The boiling point of an impure compound is generally ___________ that of pure solid.	a. same as b. less than c. greater than d. equal to the boiling point of pure solid
14	The melting point of an impure compound is generally ___________ that of the pure solid.	a. less than b. the same as c. greater than d. equal to the boiling point of pure solid
15	Adding sodium chloride to water will cause the:	a. boiling point to rise and the freezing point to lower. b. boiling point to lower and the freezing point to rise. c. both boiling point and freezing point to rise. d. both boiling point and freezing point to lower.
16	Consider a solution made from a nonvolatile solute and a volatile solvent. Which statement is TRUE?	a. The osmotic pressure is the same as vapor pressure of the solution. b. The vapor pressure of the solution is always greater than the vapor pressure of the pure solvent. c. The boiling point of the solution is always greater than the boiling point of the pure solvent. d. The freezing point of the solution is always greater than the freezing point of the pure solvent.
17	Dissolving a solute such as NaCl in a solvent such as water results in:	a. an increase in the melting point of the liquid b. a decrease in the boiling point of the liquid c. a decrease in the vapor pressure of the liquid

		d. no change in the boiling point of the liquid

5) TRUE or FALSE. Write **TRUE** if the statement is correct. Otherwise, write **FALSE**.

__________ 1. Colligative properties will only depend on the number of dissolved particles in the solution and not on their identity.

__________ 2. The addition of a solute will increase the boiling point of the solution.

__________ 3. The van't Hoff factor is only applicable for non-electrolyte solutions.

__________ 4. The boiling point of the solution is always greater than that of the pure solvent.

__________ 5. All biological membranes are regulated via osmosis.

6) Match the correct equation in Column B with the principles in Column A.

Column A	**Column B**	**Your answer**
1. Raoult's law	a. $\Delta T_b = E_b * C_m$	
2. Dalton's law of partial pressures	b. $P = C_M RT$	
3. Boiling point elevation	c. $P_1 = P^o{}_1 * X_1$	
4. Freezing point depression	d. $P_{total} = P_A + P_B$	
5. Osmotic pressure	e. $\Delta T_f = K_f * C_m$	
	f. $\Delta P = P_2 - P_1$	

LABORATORY WORK
IONIC REACTION

Objective: *investigate the reactions of ions in solution*

- Students understand terms such as double displacement reactions & different types of such reactions like neutralization, precipitation.
- Students classify the compounds that give double displacement reactions.
- Students acquire the skill to perform a double displacement reaction using barium chloride and sodium sulphate.
- Students will be able to distinguish a double displacement reaction from a given set of chemical reactions in future.

The purpose of these experiments is to:

- write balanced equations
- study ionic reactions

- write net ionic reactions

***Materials and Reagents*:** test tubes, indicators, electrolytes solutions

Ions are atoms (or groups of atoms) that carry electrical charge. Positively charged ions are called cations, whereas negatively charged ions are called anions.

Ionic compounds have certain properties common to all of them:

1) *High melting and boiling points* (they are non-volatile)

2) *Soluble in water*. Water is polar solvent and tends to dissolve ionic compounds. Soluble ionic compounds **dissociate** into ions in water (This process is called ionic dissociation):

$$Mg(NO_3)_{2(solid)} \xrightarrow{+H_2O} Mg^{2+}{}_{(aq)} + 2NO^{-}_{3(aq)}$$

3) *Solutions of ionic compounds conduct electricity*. Conduct electricity when molten or in aqueous solution: solid ionic compounds do not conduct electricity, when the substance is melted or dissolved in water, however, the ions move freely and carry an electric current.

4) *Ionic crystals shatter easily*.

COMPOUNDS		
Electrolytes *are substances that dissociate into ions when dissolved in water and conduct electricity*		**Nonelectrolytes** *are substances that do not form ions and do not conduct electricity when placed in water*
Strong electrolytes:	**Weak electrolytes:**	
• completely dissociate into ions, • solutions strongly conduct electricity, • typical compounds: soluble ionic compounds (salts), strong acids and alkalies	• incompletely dissociate into ions, • solutions weakly conduct electricity, • typical compounds: weak acids and bases, slightly soluble and insoluble salts	• no dissociation, • solutions don't conduct electricity, • typical compounds: molecular organic compounds

Acids are substances that increase the concentration of H^+ when dissolved in water (by Arrhenius theory) or proton donors (by Bronsted–Lowry theory). There are only seven strong acids:

- Hydrochloric (HCl)
- Hydrobromic (HBr)
- Hydroiodic (HI)
- Nitric (HNO_3)
- Sulfuric (H_2SO_4)
- Chloric ($HClO_3$)
- Perchloric ($HClO_4$)

Bases are a substances that increase the concentration of OH^- when dissolved in water or proton acceptors. Strong bases are group 1 hydroxides (soluble bases are called alkali: NaOH, KOH) and soluble group 2 hydroxides ($Ca(OH)_2$, $Ba(OH)_2$, $Sr(OH)_2$)

Neutralization Reactions occurs when a solution of an acid and a base are mixed and as a result forms a salt and water:

$$HCl_{(aq)} + NaOH_{(aq)} \rightarrow H_2O_{(l)} + NaCl_{(aq)}$$

The term **neutralization** is often used to describe a reaction in which equal amounts of acid and base react with each other. Acid and bases can react with some compounds to change their color, these compounds are dyes that change color as the pH changes are referred to as indicators. This is one method used to determine the point at which exact amounts of acid and base have been reacted in neutralization reactions.

Strength of the electrolyte is described by the following values: the degree (α) and the dissociation constant (K_{diss}) of the electrolyte. Degree of ionization 'α' may be defined as a fraction of total number of molecules of an electrolyte which dissociate into ions. The arbitrary border between weak and strong electrolytes is 30 % dissociation in solution.

$$\alpha = \frac{n}{N} \cdot 100\%$$

$$\alpha = \frac{Number\ of\ molecuiles\ dissociated\ into\ ions}{Total\ number\ of\ molecules}$$

where n – Number of molecules dissociated as ions, N – Total number of molecules of electrolyte dissolved. For strong electrolytes $\alpha>0,3$ (30%), and for weak electrolytes $\alpha<0,3$ (0 – 30%).

Ions present in solution constantly re-unite to form neutral molecules and, thus, there is a state of dynamic equilibrium between the ionized the ionized and non-ionised molecules, i.e.,

$$\mathbf{A_xB_y \rightleftharpoons x\,A^+ + y\,B^-}$$

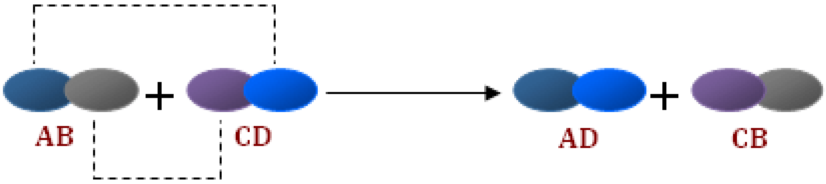
Applying the law of mass action to above equilibrium:

$$K_{diss} = \frac{[A^+]^x \cdot [B^-]^y}{[A_xB_y]}$$

where K is known as dissociation (ionization) constant. The electrolytes having high value of K_{diss} are termed strong electrolytes and those having low constant value of K_{diss} as weak electrolytes.

A chemical equation which shows dissociation of electrolyte and written as dissociated ions of electrolyte is known as **ionic equation**.

Double displacement reactions may be defined as the chemical reactions in which one component each of both the reacting molecules is exchanged to form the products. During this reaction, the cations and anions of two different compounds switch places, forming two entirely different compounds. The general equation which represents a double displacement reaction can be written as:



- Metathesis reactions will lead to a change in solution if one of three things occurs:
  - an insoluble solid is formed (When a chemical reaction forms such a solute, the insoluble solute comes out of solution and is called a ***precipitate***),
  - weak or nonelectrolytes are formed, or
  - an insoluble gas is formed (H_2S, CO_2, SO_2, NH_3).

Let's write one balanced ionic equation for one of the precipitate reaction between next electrolytes in their aqueous solution.

1) *First write the balance chemical molecular equation:*

$$(NH_4)_2S + Cd(NO_3)_2 \rightarrow 2NH_4NO_3 + CdS$$

2) Then write the total ionic equation by determining the dissociation of reactants and products in solution. Remember, a strong electrolyte is written as ions while a weak electrolyte is written as molecule because it cannot be dissociate completely in solution. Total ionic equation includes only those compounds and ions that undergo a chemical change in a reaction in an aqueous solution. (Not spectator ions):

$$Cd^{2+} + 2NO_3^- + 2NH_4^+ + S^{2-} \rightarrow CdS\downarrow + 2NO_3^- + 2NH_4^+$$

Spectator ions are ions that didn't take part in a chemical reactions. They are found on both sides of the reaction. Spectator Ions NO_3^- and NH_4^+, therefore must be excluded in net ionic equation

$$Cd^{2+} + \cancel{2NO_3^-} + \cancel{2NH_4^+} + S^{2-} \rightarrow CdS\downarrow + \cancel{2NO_3^-} + \cancel{2NH_4^+}$$

3) Hence net ionic equation should be written as follow:

$$Cd^{2+} + S^{2-} \rightarrow CdS\downarrow$$

1st Experiment: *Investigate Ionic Reaction of Precipitate*

PROCEDURE:

1. Place 2 ml of 1st solution into a test tube.
2. Place 2 ml of 2nd solution into another test tube.
3. Record observations of both solutions.
4. Pour the contents of one test tube into the other test tube and record observations:
   - On mixing aqueous solutions of silver nitrate and sodium chloride, a white curdy precipitate of silver chloride is formed.
   - On mixing an aqueous solution of barium chloride with that of copper sulphate, a white precipitate of barium sulphate is formed.

- On mixing an aqueous solution of lead nitrate with sodium sulphate, a white precipitate of lead sulphate is formed.
- On passing hydrogen sulphide gas through copper sulphate solution, a black precipitate of copper sulphide is formed.
- On adding a solution of lead nitrate to sodium iodide solution, a yellow precipitate of lead iodide is formed.
- Cobalt (II) chloride reacts with sodium carbonate to form pink/red coloured precipitate of cobalt(II) carbonate and sodium chloride.
- On adding aluminium sulphate solution to calcium chloride solution, a precipitate of calcium sulphate is formed.

5. Write the molecular, complete ionic and net ionic equations for each of the following reactions.

2nd Experiment: *Investigate Ionic Reaction of forming a Weak Electrolyte*

PROCEDURE:

The goal of this experiment is to identify seven aqueous solutions using qualitative analysis.

a. Qualitative analysis is a systematic method of recording precipitation reactions, color changes, and other visible changes to determine chemical composition.

b. Acids and bases will be identified using pH paper and phenolphthalein indicator.

Acids	*Color of universal paper indicator in acid*	*Base*	*Color of Phenolphthalein indicator in base*
HCl		NaOH	
HNO_3		NaOH	
H_2SO_4		NaOH	

c. After a pH analysis, all solutions will be mixed together to determine if a reaction has taken place. Write the molecular, complete ionic and net ionic equations for each of the following reactions.

d. Write the molecular, complete ionic and net ionic equations for each of the following reactions.

3rd Experiment: *Investigate Ionic Reaction of forming an Insoluble Gas.*

PROCEDURE:

Another way a double-replacement reaction occurs if the reaction produces an unstable compound that decomposes into a **gas** and water.

The three most common unstable compounds which give **gases** are:

- Carbonic acid: $H_2CO_3 \rightarrow CO_{2(g)} + H_2O_{(l)}$
- Sulfurous acid: $H_2SO_3 \rightarrow SO_{2(g)} + H_2O_{(l)}$
- Ammonium hydroxide: $NH_4OH \rightarrow NH_{3(g)} + H_2O_{(l)}$

- Hydrogen sulphide: $2H^+ + S^{2-} = H_2S$

1) Marble chips ($CaCO_3$) react with dilute hydrochloric acid to form calcium chloride and carbonic acid. Carbonic acid is unstable and readily decomposes to form carbon dioxide and water. Since H_2CO_3 decomposes into CO_2 and H_2O, you can see bubbles forming. That's a sign that the double-replacement reaction is occurring.

- Locate the 0.1M HCl solution and the test tube with calcium bicarbonate ($CaCO_3$). Transfer a small amount of the bicarbonate to an empty test tube. See image for amount to use.
- This time don't add any water to the sodium bicarbonate that you transferred. There is already water in the 0.1 M HCl.
- Add 0.1 M HCl to the calcium bicarbonate powder.

2) Many sulfide salts will react with acids to form gaseous hydrogen sulphide. Sodium sulfide (Na_2S) react with dilute HCl acid to form sodium chloride and hydrogen sulphide, as a result you can feel the smell of rotten eggs.

By paying attention to colors and smells of gases released by a reaction one can often infer the chemical identity of the original compound. Be careful, however, many of the smells are strong and some are toxic. To smell a gas, don’t hold a solution under your nose and inhale, rather hold the solution away from you and waft the gases toward your nose with your hand.

3) Mix the following solutions in tube: ammonium chloride and sodium hydroxide. Then heat the tube. Attach to the wall test tube blue litmus paper. In the formation of white gas, look at the color of litmus paper, as a result it will be dark blue. Because ammonium hydroxide is formed, which decomposes into gaseous ammonia and water. Smell cautiously gas: ammonia has a pungent peculiar smell.

Write the molecular, complete ionic and net ionic equations for each of the following reactions.

4th Experiment: *Investigate Ionic Reaction of dissolving the precipitate of electrolyte*

PROCEDURE:

1. Place 5 ml of aluminium chloride ($AlCl_3$) solution into a test tube.
2. Add by pipette 2 ml of 0.1N sodium hydroxide (NaOH) solution into this test tube until you will see a white precipitate of aluminium hydroxide.
3. Divide the precipitate into two test tubes.
4. Add an excess of hydrochloric acid to the first test tube, until the precipitate dissolved. Then add an excess of sodium hydroxide to the second test tube, until the precipitate dissolved too.
5. Aluminum hydroxide is amphoteric base, so it′s precipitate react and dissolve in strong acids and alkalis. Write the molecular, complete ionic and net ionic equations for each of the following reactions.

EXERCISES:

1) Write the molecular, complete ionic and net ionic equations for each of the following reactions:

a) Aqueous solutions of sodium sulfide and calcium nitrate are mixed.

b) Aqueous solutions of barium chloride and potassium sulfate are mixed.

c) Aqueous solutions of silver nitrate and ammonium chloride are mixed.

2) Write the equation for the dissociation of the following electrolytes, call them, indicate their strength, for the weak electrolytes write the equation of the dissociation constant: a) HNO_3 b) $CuCl_2$ c) $Al(OH)_3$

3) Write 2 example of the molecular equations for next net ionic reaction:

$$\mathbf{2\,I^{-}\,(aq) + Pb^{2+}\,(aq) \rightarrow PbI_{2\,(s)}}$$

4) Perform test:

№	Question	Variant
1	The products obtained when HCl reacts with NaOH are ……………………. & ………………… respectively. Such reactions are called …	a. Cl_2, H_2O, Neutralisation b. H_2 , O_2, Combustion c. H_2O, NaCl, Double displacement d. Cl_2, H_2O, decomposition
2	What is the precipitate formed on mixing the solution of barium chloride and sodium sulphate?	a. Barium sulphate b. Barium sulphite c. Sodium chloride d. Sodium sulphite
3	Identify the product A in the reaction: $Al_2O_3(s) + HClO_4(aq) \text{-->} A\ (aq) + H_2O\ (l)$	a. $Al(ClO_4)_3$ b. $Al_2(ClO_4)_3$ c. Al_3ClO_4 d. $AlClO_4$
4	Four students were asked to study the reaction between barium chloride and sodium sulphate. They reported their experiment as follows. Which one is a correct report ?	a. On mixing the powder of barium chloride and sodium sulphate, the colour of the mixture changes to brown. b. On adding powdered sodium sulphate to barium chloride solution; solution becomes white. c. On adding the powder of barium chloride to sodium sulphate solution; solution turns white. d. On mixing solution of barium chloride and sodium sulphate, white solid substance is formed.
5	Which of the following reactions represents the reaction between barium chloride and sodium sulphate correctly?	a. $NaSO_4 + BaCl_2 \rightarrow BaSO_4 + NaCl$ b. $Na_2SO_4 + BaCl \rightarrow BaSO_4 + NaCl$ c. $Na_2SO_4 + BaCl_2 \rightarrow BaSO_4 +$

		2NaCl d. $NaSO_4$ + BaCl → $BaSO_4$ + NaCl
6	Identify the product(s) of the following double displacement reaction $CoCl_2$ (aq) + Na_2CO_3 (aq) -->	a. NaCl(s) and $CoCO_3$(s) b. NaCl(aq) and $CoCO_3$(aq) c. NaCl(aq) and $CoCO_3$(s) d. NaCl(s) and $CoCO_3$(aq)
7	Which of the following is not a precipitation reaction?	a. Reaction of $NaNO_3$ and KI b. Reaction of $AgNO_3$ and $MgCl_2$ c. Reaction of $Pb(NO_3)_2$ and NaI d. Reaction of $FeBr_2$ and $Sr(OH)_2$
8	Consider the reaction Aluminium sulphate + Calcium hydroxide--> Aluminium hydroxide +Calcium sulphate. The stoichiometric ratio for the balanced equation are respectively:	a. 1:3:2:3 b. 2:1:2:1 c. 1:2:3:1 d. 1:2:1:3
9	What is the acid formed by the reaction between lead acetate and hydrochloric acid?	a. Propanoic acid b. Formic acid c. Carbonic acid d. Acetic acid
10	Which one of the following is a neutralization reaction?	a. Reaction between ammonium chloride and KOH b. Reaction between calcium carbonate and HCl c. Reaction between NaOH and HCl d. Reaction between copper sulphate and hydrogen sulfide
11	Identify the precipitate obtained by the reaction between lead nitrate and sodium sulphate.	a. $PbSO_4$ b. $NaNO_3$ c. $Pb(SO_4)_2$ d. PbS
12	Which one of the following is true about a double displacement reaction?	a. Reaction between MgO and water is an example for this reaction. b. Only a single product is formed. c. Two compounds exchange ions to form new compounds. d. Electrolysis of water is an example of this reaction.
13	Which one of the following represents a double displacement reaction?	a. A + B --> C b. C --> A +B c. A + BC --> AC + B

		d. AB + CD --> AD + CB
14	Geethu added a small quantity of dil. HCl into a precipitate of barium sulphate in a test tube obtained by adding barium chloride solution to sodium sulphate solution. Which one of the following is the correct observation?	a. The precipitate is soluble in HCl b. The colour of the precipitate turns to yellow c. Formation of brisk effervescence d. Precipitate remains insoluble

LABORATORY WORK
SALTS HYDROLYSIS

Objective: *investigate indicators and reaction of salt′s hydrolysis.*

***Skills to develop*:**

- *Predict the acidity of a salt solution.*
- *Calculate the pH of a salt solution.*
- *Calculate the concentrations of various ions in a salt solution.*
- *Explain hydrolysis reactions.*

Water molecules are highly polar and are in continuous motion. Occasionally, the collisions between water molecules are energetic enough to transfer a hydrogen ion from one water molecule to another.

Self ionization of water – the reaction in which water molecules produce ions. A water molecule that loses a hydrogen ion becomes a negatively charged hydroxide ion. A water molecule that gains a hydrogen ion becomes a positively charged hydronium ion:

$$HOH_{(l)} \Leftrightarrow H^+_{(aq)} + OH^-_{(aq)}$$

In pure water at 25˚C, the equilibrium concentration of hydrogen ions and hydroxide ions are each only 1×10^{-7}.

Any aqueous solution in which H^+ and OH^- are equal is a **neutral solution**:

$[H^+] = [OH^-] = 10^{-7} mol/l$

Not all solutions are neutral. When some substances dissolve in water, they release hydrogen $[H^+]$ ions or hydroxide $[OH^-]$ ions.

If $[H^+] > [OH^-]$ take place in a solution this solution is called **acidic** $[H^+] > 10^{-7} mol/l$.

If $[H^+] < [OH^-]$ take place in a solution this solution is called **basic**. Basic solutions are also known as **alkaline solutions** $[H^+] < 10^{-7} mol/l$.

For aqueous solutions, the product of the hydrogen ion concentration and the hydroxide ion concentration equals: $[H^+] \cdot [OH^-] = 10^{-14} mol/l$. This equation is true for all dilute aqueous solutions at 25˚C.

<u>Ion-Product Constant for Water</u> (K_w) – the product of the concentrations of the hydrogen ions and hydroxide ions in water: $K_W = [H^+]\cdot[OH^-] = 10^{-14}$

The quantitative designation of the acidity or alkalinity of a solution may be still further simplified by using the **hydrogen ion concentration index** (pH). The relation defines this index:

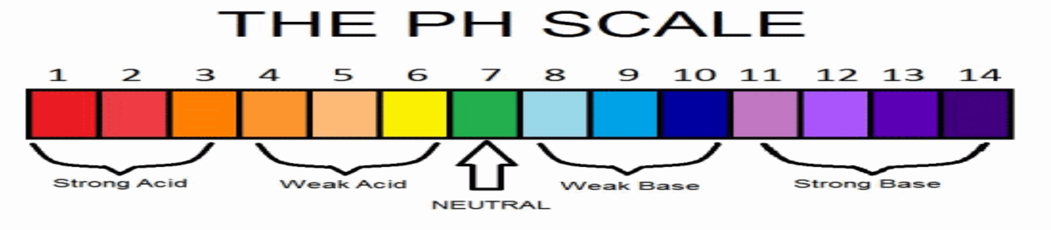
$pH = -\lg[H^+]$ and $pOH = -\lg[OH^-]$

The range of the pH scale is from 0 to 14: $pH + pOH = 14$

$[H^+] = 10^{-pH}$ and $[OH^-] = 10^{-pOH}$

Thus Solutions can be classified based on their pH values:

- $[H^+] > 10^{-7}\,mol/l$ and pH < 7.0, the solution is acidic;
- $[H^+] = [OH^-] = 10^{-7}\,mol/l$ and pH = 7.0, the solution is neutral;
- $[H^+] < 10^{-7}\,mol/l$ and pH > 7.0, the solution is basic.



In practice, the acidity or alkalinity of a solution is conveniently determined by means of **indicators – substances that change color depending on the relative concentrations of H^+ and OH^- ions**. *This method is called a colorimetric. Thus,* ***indicator is*** *a halochromic chemical compound* (weak organic acids or bases that react with ions in solution) *that is added in small amounts to a solution so that the pH (acidity or basicity) of the solution can be determined visually.* The color change of different indicators occurs at different hydrogen ion concentrations, which is important for chemical analysis.

For more precious measuring of pH the special tools are widely used– **pH-meters**, which provides assurance of measuring within the limits of ± 0,01.

1st Experiment. *Determine the pH of water and various electrolyte solutions using acid-base indicators*

Materials and Reagents: test tubes, acid-base indicators, electrolytes solutions

PROCEDURE:

1. Data Table 1 lists the formulas of the 3 electrolytes solutions you will be using.

2. Set up a test tube rack with three test tubes in a row.

3. Use the marking pen to number the test tubes to correspond to the numbers of the first indicator in Table 1.

4. Add 10 mL of each electrolyte solution to each of the test tubes.

5. Use the thin stem pipet to add 3 drops of first indicator to each test tube. Swirl to mix the drops into the solution.

6. Note the color of the indicator in each tube and enter the data into the Table 1.

№	**Indicator type**	**ΔpH range which indicator changes color**	**The indicator color in …. solution**		
			Acidic (HCl_{aq})	**Neutral (H_2O_{aq})**	**Basic ($NaOH_{aq}$)**
1	Methyl orange (sln)	**3,2 – 4,4**			
2	Blue litmus paper	**5,0 – 8,0**			
3	Phenolphthalein (sln)	**8,2 – 10,0**			
4	Universal Ind (sln)	**0 – 14**			
5	Universal Ind paper	**0 – 12**			

Analysis:

1. Which indicator is used to determine the acidic solution?
2. Which indicator is used to determine the basic solution?

A **salt** is formed between the reaction of an acid and a base. Usually, a neutral salt is formed when a strong acid and a strong base is neutralized in the reaction: $\mathbf{H^+ + OH^- = H_2O}$

There are four possible ways of forming salts:

1) If the salt is formed from a *strong* base and *strong* acid, then the salt solution is neutral, indicating that the bonds in the salt solution will not break apart (indicating no hydrolysis occurred) and is **neutral (pH=7)**.
2) If the salt is formed from a *strong* acid and *weak* base, the bonds in the salt solution will break apart and becomes **acidic (pH<7)** and hydrolyzes.
3) If the salt is formed from a *strong* base and *weak* acid, the salt solution is **basic (pH>7)** and hydrolyzes.

4) If the salt is formed from a *weak* base and *weak* acid, it will hydrolyze, but the acidity or basicity depends on the equilibrium constants of K_a and K_b. If the K_a value is greater than the K_b value, the resulting solution will be acidic and vice versa:
   - If K_a(cation) > K_b(anion) the solution of the salt is acidic.
   - If K_a(cation) = K_b(anion) the solution of the salt is neutral.
   - If K_a(cation) < K_b(anion) the solution of the salt is basic.

Table 5 - Four possible ways of forming salts and their hydrolysis

№	**Type of salt is involved in hydrolysis**	**Mechanism of hydrolysis**	**Cations types**	**Anions types**
1	*A salt formed between a strong acid and a strong base is an* ***neutral salt***	**No hydrolysis (pH=7)**	***Strong base cations***	***Strong acid anions***
			Na^+, K^+, Rb^+, Cs^+, Ca^{2+}, Sr^{2+}, Ba^{2+}	Cl^-, Br^-, I^-, NO_3^-, SO_4^{2-}, ClO_4^-
2	*A salt formed between a strong acid and a weak base is an* ***acid salt***	**Cationic hydrolysis (pH<7)**	***Weak base cations***	***Strong acid anions***
			Insoluble in water base **cations:** NH_4^+, Ag+, Cu^{2+}, Zn^{2+}, Al^{3+}, Cr^{3+} and others	Cl^-, Br^-, I^-, NO_3^-, SO_4^{2-}, ClO_4^-
3	*A salt formed between a weak acid and a strong base is a* ***basic salt***	**Anionic hydrolysis (pH>7)**	***Strong base cations***	***Weak acid anions***
			Na^+, K^+, Rb^+, Cs^+, Ca^{2+}, Sr^{2+}, Ba^{2+}	F^-, CH_3COO^-, CN^-, NO_2^-, S^{2-}, CO_3^{2-}, SiO_3^{2-}, SO_3^{2-}, PO_4^{3-}
4	*A salt formed between a weak acid and a weak base* ***can be neutral, acidic, or basic*** *depending on the relative strengths of the acid and base.*	**Cationic-anionic hydrolysis (pH≈7)**	***Weak base cations***	***Weak acid anions***
			Insoluble in water base cations: NH_4^+, Ag+, Cu^{2+}, Zn^{2+}, Al^{3+}, Cr^{3+} and others	F^-, CH_3COO^-, CN^-, NO_2^-, S^{2-}, CO_3^{2-}, SiO_3^{2-}, SO_3^{2-}, PO_4^{3-}

The reaction of salt takes place in the solution. In reality, and looking at a wider variety of 'salts', the picture is much more complicated and a 'salt' solution may be acid, neutral or alkaline depending on the nature of the interaction of the salt ions with water. **The reaction of the salt with water, whereby the salt is dissociated and decomposed to form a weak electrolyte (weak acid or weak base) called <u>hydrolysis</u>** ("chemical decomposition by water," 1880, formed in English from <u>hydro-</u> + Greek *lysis* "a loosening, a dissolution," from *lyein* "to loosen, dissolve"). Hydrolysis is the reverse of neutralization.

To determine acidity or basic of a salt:

1. Write the neutralization reaction in reverse (salt and water makes base and acid).
2. Break up the salt into its ions. Water is weak electrolyte and does not dissociate.
3. To write the base, add an OH^- to the cation (the positive ion) for each positive (+) charge.
4. To write the acid, add H^+ to the anion (the negative ion) for each negative (–) charge.

If you for a strong base, the salt is basic (NaOH, KOH, LiOH, $Ca(OH)_2$ ...etc.)

If you form a strong acid it is acidic. (HCl, HNO_3, $HClO_4$, HBr...)

If you form both a strong acid and a strong base it is neutral.

EXAMPLES

1) SODIUM NITRITE ***$NaNO_2$***

$NaNO_2 + H_2O \Leftrightarrow NaOH + HNO_2$

strong ***weak***
base ***acid***

$Na^+ + NO_2^- + H_2O \Leftrightarrow Na^+ + \mathbf{OH^-} + HNO_2$

$\mathbf{NO_2^-} + H_2O \Leftrightarrow \mathbf{OH^-} + HNO_2$ **(basic of a salt, pH>7)**

anion
hydrolyzed

Since NaOH is a strong base it breaks up and yields OH^-, the salt is basic. HNO_2 is a weak acid and will form (does not break up in water).

2) AMMONIUM BROMIDE ***NH_4Br***

$NH_4Br + H_2O \Leftrightarrow NH_4OH + HBr$

weak ***strong***
base ***acid***

$NH_4^+ + Br^- + H_2O \Leftrightarrow NH_3 + H_2O + \mathbf{H^+} + Br^-$

$\mathbf{NH_4^+} + H_2O \Leftrightarrow NH_3 + H_2O + \mathbf{H^+}$ **(acidity of a salt, pH<7)**

cation
hydrolyzed

Since HBr is a strong acid it breaks up and yields H^+, the salt is acidic. NH_4OH is a weak base. They generally stay together, however this is actually breaks down into ammonia and water.

2nd **Experiment**. *Investigate the reaction of salt hydrolysis by colorimetric method.*

Materials and Reagents: test tubes, universal indicator solution and paper, salts solutions

PROCEDURE:

Identify the relative strengths (weak or strong) of the acidic ions and basic ions in solution. Predict whether the resulting solution will be acidic, basic or neutral. Use a pH meter (or pH paper) to determine the pH of each solution. Write a balanced equation for each hydrolysis reaction. Which salt is not hydrolyzed?

1. Data Table 5 lists the formulas of the four salts you will be using. Using your knowledge of hydrolysis, form a prediction about which of the salt solutions will be acidic, which will be basic, and which, if any, will be neutral. Record your prediction in the space provided.

2. Set up a test tube rack with four test tubes in a row.

3. Use the marking pen to number the test tubes to correspond to the numbers of the first four salts in the following table.

№	**Formula of salt**	**Structure of salt**		**Color of indicator in salt solution**	**Approximate pH value on scale indicator**	**The reaction medium (acidity, basic or neutral)**
		Base (weak or strong)	**Acid** (weak or strong)			
1						
2						
3						
4						

4. Use the spatula to place a small amount of the first six salts in the corresponding test tubes 1 – 4. Put one salt in each tube. Clean the spatula with a paper towel between each use.

5. Add 10 mL of distilled water to each of the test tubes.

6. Close each test tube with a rubber stopper, and shake each tube until the solid salt is dissolved completely.

7. Remove the rubber stoppers. Use the thin stem pipet to add 3 drops of universal indicator to each test tube. Swirl to mix the drops into the solution. (Alternatively, use pH paper or pH probes to test the pH of the solutions.)

8. Use the chart of pH and indicator colors to determine the pH of each solution. Record the value in Data Table 1. Identify whether the pH is acidic, basic, or neutral in the column provided in Data Table 5.

9. All solutions except $Pb(NO_3)_2$ can be disposed of down the sink, flushing with plenty of water. Collect the lead (II) nitrate in a container provided by your teacher to be disposed of properly.

10. Repeat steps 2-9 with the last three solutions.

11. Some of the 4 salts you tested will cause water to undergo hydrolysis to form weak acids or weak bases. Other salts will not cause hydrolysis. Complete the last two columns in Data Table 5 by writing the formula for the species produced in the correct column. Refer to the background information example if you need help with this.

Analysis:

1) Write a net ionic equation to show the species formed for each salt that undergoes a hydrolysis reaction with water. Use the equation in the background information as an example. Write “NO REACTION” if the salt does not cause hydrolysis.

2) Complete the data table showing the weak acid or weak base formed in the products. Use your net ionic equations to determine this.

3) Be prepared to discuss in class:

- Which chemical salts formed acidic solutions?
- Which chemical salts formed basic solutions? How do you identify ions that come from a weak acid?
- How do you identify ions that come from a weak base?
- How do you write an equation for the hydrolysis of water?

EXERCISES:

1) What is the value of pH and pOH in solutions in which the concentration of hydrogen ions is: a) $1\cdot10^{-7}$; b) $1\cdot10^{-5}$; and the concentration of hydroxide ions is: c) $1\cdot10^{-10}$; d) $1\cdot10^{-12}$ mol/l?

2) What is pH value of 0,01M sodium hydroxide solution if its dissociation is complete?

3) What is pH value of 0,05N sulfur acid solution if its dissociation is complete?

4) What salt aqueous solution will have acidic medium? Write the corresponding equations:

a) Aluminium nitrate, Potassium sulfide, Potassium chloride;
b) Potassium carbonate, Iron (II) sulfate, Lithium chloride;
c) Copper (II) chloride, Sodium chloride, Lithium carbonate.

5) What salt aqueous solution will have alkaline medium? Write the corresponding equations:

a) Calcium chloride, Aluminium chloride, Sodium silicate;
b) Lithium nitrate, Potassium carbonate, Zinc nitrate;
c) Potassium chloride, Aluminium chloride, Sodium sulfide.

6) Perform test

№	Question	Variant
1	Choose substances that form water solutions with alkaline medium:	a) Sodium chloride b) Hydrogen chloride c) Sodium carbonate d) Ammonium chloride

		e) Sodium nitrite
2	The medium of water solution of sodium hydrogen phosphate is:	a) Slightly acidic b) Strongly acidic c) Slightly alkaline d) Strongly alkaline e) Neutral
3	Note the solution that has most alkaline condition:	a) $[H^+] = 10^{-7}$ mol/l b) $[OH^-] = 5\times10^{-8}$ mol/l c) $[OH^-] = 10^{-4}$ mol/l d) $[OH^-] = 5\times10^{-10}$ mol/l e) pH = 5,0
4	The pH of a solution is equal to 5,0. Find the hydrogen ion concentration:	a) $1,0\times10^{-14}$ b) $1,0\times10^{-9}$ c) $1,0\times10^{-5}$ d) 5,0 e) 9,0
5	Calculate the pH of a acetic acid solution, if its dissociation degree is equal to 1%:	a) 0,76 b) 2,76 c) 3,76 d) 4,76 e) 5,76
6	What medium will be in a solution of aluminium nitrate:	a) Alkaline b) Neutral c) Acidic d) none of these
7	Find compounds that reduce a degree of hydrolysis of a sodium cyanide solution:	a) Nitrate of potassium b) Water c) Sodium hydroxide d) Barium chloride e) Potassium hydroxide
8	Hydrolysis of the salt causes acidic reaction of medium in water solution:	a) BaF_2 b) $ZnCl_2$ c) Na_2S d) CH_3COONH_4 e) Na_2CO_3
9	The hydrogen ion concentration in electrolyte solution with pOH=9 is equal to:	a) $1,0\times10^{-14}$ mol/dm^3 b) $1,0\times10^{-5}$ mol/dm^3 c) 5,0 mol/dm^3 d) 9,0 mol/dm^3 e) $1,0\times10^{-9}$ mol/dm^3
10	Where is the hydrolysis not possible:	a) NaCl b) $FeCl_3$ c) $Al_2(SO_4)_3$ d) K_2SO_4

		e) NH_4NO_3
11	Note the ions that can not be in considerable amounts in a neutral solution:	a) Ba^{2+} b) CO_3^{2-} c) NO_3^- d) Li^+ e) SO_4^{2-}
12	Where does the acid base indicator phenolphthalein gain of crimson colour:	a) HCl b) KOH c) Na_2CO_3 d) NaCl e) H_2CO_3
13	The alkaline solution can be received at dissolution in water of:	a) Aluminium sulphate b) Sodium hydrogen phosphate c) Sodium dihydrogen phosphate d) Iron (III) chloride e) Ammonium chloride
14	The water solution of the following compounds have acidic medium:	a) Sodium chloride b) Copper sulphate c) Sodium carbonate d) Ammonium chloride e) Sodium sulphide
15	Find basis and acid that form salt XA, if such ionic - molecular equation $A^- + H_2O = HA + OH^-$ correspond to its hydrolysis:	a) Strong acid and weak base b) Weak acid and strong base c) Weak acid and weak base d) Strong acid and strong base

LABORATORY WORK
WATER HARDNESS

Objective: *to determine in laboratory a) by titration with 0,1N HCl solution temporary water hardness and b) by titration with 0,05N EDTY solution general water hardness*

Skills to develop:

- to be able to titrate
- to prepare solutions
- to measure the hardness of the water

Hard water is water that has high mineral content (in contrast with "soft water"). Hard water is formed when water percolates through deposits of calcium and magnesium-containing minerals such as limestone, chalk and dolomite.

Hard drinking water is generally not harmful to one's health, but can pose serious problems in industrial settings, where water hardness is monitored to avoid costly breakdowns in boilers, cooling towers, and other equipment that handles water. In domestic settings, hard water is often indicated by a lack of suds formation when soap is agitated in water, and by the formation of lime scale in kettles and water heaters:

$$CaSO_4 + 2RCOONa = (RCOO)_2Ca\downarrow + Na_2SO_4$$

$$MgCl_2 + 2RCOONa = (RCOO)_2Mg\downarrow + 2NaCl$$

Water's hardness is determined by the concentration (mmol-eq/L) of multivalent cations (Ca^{2+} and Mg^{2+}) in 1 liter of the water. The total water hardness is the sum of the molar concentrations of Ca^{2+} and Mg^{2+}, in mmol/L units. One unit of hardness (1 mmol/L) corresponds to the content of calcium ions equal to 20,04 mg / L of magnesium ions or equal to 12,16 mg / liter:

$$GH = \frac{m_{Ca^{2+}}}{20{,}04} + \frac{m_{Mg^{2+}}}{12{,}16}$$

Theoretically any water hardness (H) can be calculated by the formula:

$$H = \frac{m_{salt}}{Eq_{(salt)} \cdot V_{H_2O(L)}}$$

On scale hardness distinguishes soft, hard and very hard natural water:

Value of hardness (mmol/L)	**Water hardness**	**Example**
H > 1,5	Very soft	Most soft water is rain, snow (0.1 mmol/L)
H = 1,5 – 4,0	Soft	Rain water
H = 4,0 – 8,0	Moderately hard	Drinking water
H = 8,0 – 12	Hard	
H > 12,0	Very hard	The most hard - water of oceans (up to 130 mmol/L)

General hardness is the sum of temporary and permanent hardness: **GH = TH + RH**

Temporary (carbonate) hardness is a type of water hardness caused by the presence of dissolved bicarbonate minerals: $Ca(HCO_3)_2$ and $Mg(HCO_3)_2$. Temporary hardness can be reduced either by boiling the water, or by the addition of lime (calcium hydroxide) through the softening process of lime softening. Boiling promotes the formation of carbonate from the bicarbonate and precipitates calcium carbonate out of solution, leaving water that is softer upon cooling:

$$Ca(HCO_3)_2 \xrightarrow{t^\circ} CaCO_3 \downarrow + CO_2 \uparrow + H_2O$$

$$Mg(HCO_3)_2 \xrightarrow{t^\circ} MgCO_3 \downarrow + CO_2 \uparrow + H_2O$$

Unit of measure carbonate hardness: in German degrees of hardness. 1AdH corresponds to 18 mg/L $CaCO_3$:

- below 10 AdH – soft
- from 20A dH – hard
- above 30A dH – very hard

Permanent hardness is hardness (mineral content) that cannot be removed by boiling. When this is the case, *it is usually caused by the presence of calcium sulfate and chloride ($CaSO_4$, $CaCl_2$), magnesium sulfates and chloride ($MgSO_4$, $MgCl_2$) in the water, which do not precipitate out as the temperature increases.* Ions causing permanent hardness of water can be removed by using a water softener, or ion exchange column. It can be removed by soda (Na_2CO_3), surfactants (Na_3PO_4) or lime ($Ca(OH)_2$):

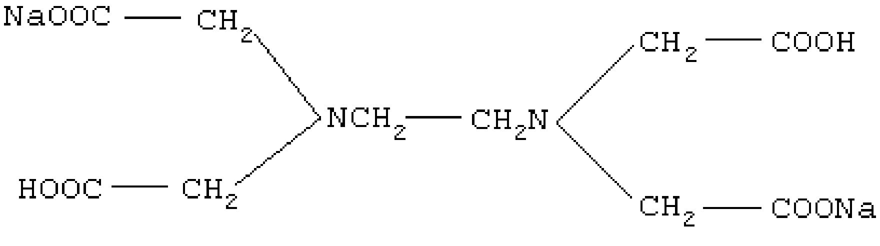
$$CaSO_4 + Na_2CO_3 \rightarrow Na_2SO_4 + CaCO_3 \downarrow$$

$$MgSO_4 + Ca(OH)_2 \rightarrow Mg(OH)_2 \downarrow + CaSO_4$$

$$3CaCl_2 + 2Na_3PO_4 \rightarrow Ca_3(PO_4)_2 \downarrow + 6NaCl$$

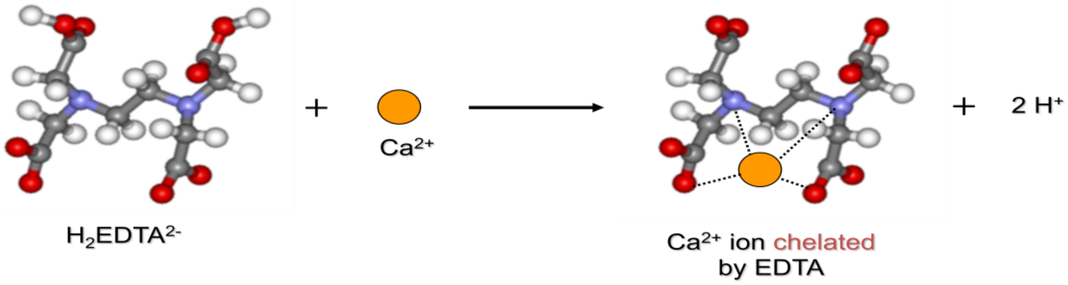
Water hardness can be determined by the following 2 methods:

- By titration with 0,1N HCl solution in the presence of a titration acid-base indicator methyl orange. In this method temporary hardness can be measured.
- By titration with disodium edetate 0,05N Na_2H_2Y solution (Trilon B) in the presence of a titration indicator Eriochrome Black-T. It is a salt of EDTA (ethylenediaminetetraacetic acid). In this method general hardness can be measured.



Molecular weight of disodium salt of EDTA: $(CH_2COOH)_2N_2(CH_2)_2(CH_2COONa)_2*2H_2O$ = (12+1*2+12+16*2+1)×2 +14*2+ (12+2)*2+ (12+1*2+12+16*2+23)×2 + 2*18 = 118+ 28+28+162+36 = 372 g/mol

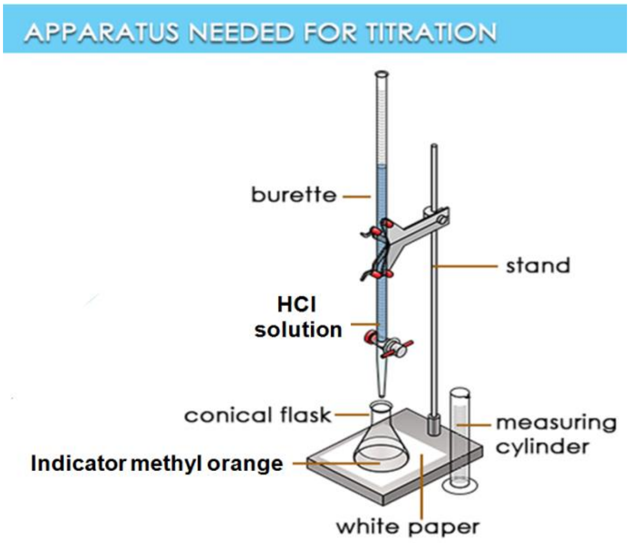
EDTA has a greater affinity for Ca^{2+} and Mg^{2+} when it is in the form of the dihydrogen anion H_2EDTA^{2-}. This is the ionic form of EDTA at pH 10. H_2EDTA^{2-} binds to a Ca^{2+} ion by forming four special covalent bonds are called coordinate covalent bonds:



In today’s experiment, you will determine the total concentration of calcium and magnesium ions in a hard water sample using EDTA in solution ammonia buffered (NH_4OH – NH_4Cl) to a pH of 10.

Preparation of ammonia buffer solution: 145 ml of liquor ammonia (NH_4OH) of specific gravity 0.88 + 15 gm NH_4Cl + distilled water to make 250 ml solution to give a pH of 10.

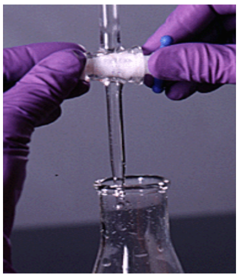
Erio - T indicator or Eriochrome Black-T indicator is used in this titration. When it is chealted or acidifies, it produces a pink red solution. When it is not chelated and under basic conditions it is blue.



A titration is a method or process of determining the concentration of a dissolved substance in terms of the smallest amount of reagent of known concentration required to bring about a given effect in reaction with a known volume of the test solution in the titration flask. A **buret** is used to deliver the second reactant to the flask and an indicator or pH-meter is used to detect the endpoint of the reaction.

This method relies on observing a color change in the solution. Indicators are weak organic acids or bases that are different colors in their dissociated and undissociated states. Because they are used in low concentrations, indicators do not appreciably alter the equivalence point of a titration. The point at which the indicator changes color is called *the end point*. The volume added to achieve the end point may be calculated using this formula (equivalent law):

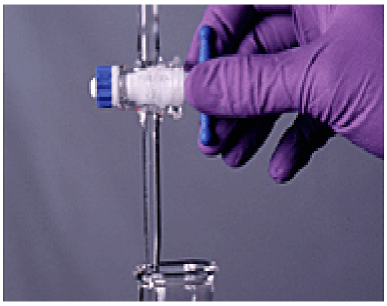
$$V_{titrant} \cdot N_{titrant} = V_{analyte} \cdot N_{analyte} \quad \Rightarrow \quad N_{analyte} = \frac{V_{titrant} \cdot N_{titrant}}{V_{analyte}}$$



Begin by preparing your buret, as described on the buret page. Your buret should be conditioned and filled with titrant solution. You should check for air bubbles and leaks, before proceding with the titration.

Prepare the solution to be analyzed by placing it in a clean Erlenmeyer flask or beaker. If your sample is a solid, make sure it is completely dissoloved. Put a magnetic stirrer in the flask and add indicator.

Use the buret to deliver a stream of titrant to within a couple of mL of your expected endpoint. You will see the indicator change color when the titrant hits the solution in the flask, but the color change disappears upon stirring.



Approach the endpoint more slowly and watch the color of your flask carefully. Use a wash bottle to rinse the sides of the flask and the tip of the buret, to be sure all titrant is mixed in the flask.

As you approach the endpoint, you may need to add a partial drop of titrant. You can do this with a rapid spin of a teflon stopcock or by partially opening the stopcock and rinsing the partial drop into the flask with a wash bottle. Ask your TA to demonstrate these techniques for you, in the lab.

Make sure you know what the endpoint should look like. For phenolphthalein, the endpoint is the first permanent pale pink (purpule). The pale pink fades in 10 to 20 minutes.

If you think you might have reached the endpoint, you can record the volume reading and add another partial drop. Sometimes it is easier to tell when you have gone past the endpoint.

If the flask looks like this, you have gone too far!

When you have reached the endpoint, read the final volume in the buret and record it in your notebook.

Subtract the initial volume to determine the amount of titrant delivered. Use this, the concentration of the titrant, and the stoichiometry of the titration reaction to calculate the number of moles of reactant in your analyte solution.

1st Experiment. ***Determine the Temporary Hardness of Water by Acid-base titration***

Materials and Reagents: 0.1N hydrochloric acid (HCl) solution, acid-base indicator methyl orange, Erlenmeyer flasks (250 ml), Graduated cylinder (100 ml), buret (25 ml).

PROCEDURE:

For determining temporary hardness, we measured by cylinder 100 ml hard water and taken into a 250 ml conical flask. Then few 3-4 drops of methyl orange is added in it as an indicator. Now titration is carried out by adding 0.1N cold HCl until the yellow color of methyl orange turns red. Titration was repeated 3 times

During titration in solution occurs the following reaction:

$$Ca(HCO_3)_2 + 2HCl = CaCl_2 + H_2O + CO_2\uparrow$$

$$Mg(HCO_3)_2 + 2HCl = MgCl_2 + H_2O + CO_2\uparrow$$

Protocol of analysis:

1. The volume of water taken for titration $V_{H_2O} = 100\ ml$
2. Volume of acid which went on titration for the first time $V^1_{HCl} =$
3. Volume of acid which went on titration for the second time $V^2_{HCl} =$
4. Volume of acid which went on titration for the third time $V^3_{HCl} =$
5. The average amount of acid which went on titration $V^*_{HCl} = \dfrac{V^1 + V^2 + V^3}{3}$
6. Normality of acid solution $N_{HCl} = 0{,}1N$

Calculate TH (mmol/L) by following formula: $TH = \frac{V^*_{HCl} \cdot N_{HCl}}{V_{H_2O}} \cdot 1000$

2nd Experiment. ***To determine the General Hardness of Water by Complexometric titration***

Materials and Reagents: 0.05N disodium salt of EDTA solution, complexometric indicator Eriochrome Black-T, ammonia buffer solution, Erlenmeyer flasks (250 ml), Graduated cylinder (100 ml), buret (25 ml).

Procedure:

- Use a graduated cylinder to dispense 50.00 mL of analyses water into a 250 mL flask.
- Add 2 mL of pH 10 ammonia buffer solution, and then 7 - 10 drops of Eriochrome Black T indicator. The resulting solution will turn the dark red color.
- Titrate the solution with 0,05N EDTY from your buret. As you near the endpoint, the solution will turn purple. Continue to slowly add EDTY until the solution turns blue, with no trace of red.

Protocol of analysis:

1. The volume of water taken for titration $V_{H_2O} = 50\ ml$
2. Volume of EDTA which went on titration for the first time $V^1_{EDTY} =$
3. Volume of EDTA which went on titration for the second time $V^2_{EDTY} =$
4. Volume of EDTA which went on titration for the third time $V^3_{EDTY} =$
5. The average amount of EDTA which went on titration
$V^*_{EDTY} = \frac{V^1 + V^2 + V^3}{3}$
6. Normality of acid solution N_{EDTY} = 0,05N

Calculate GH (mmol/L) by following formula: $GH = \frac{V^*_{EDTY} \cdot N_{EDTD}}{V_{H_2O}} \cdot 1000$

EXERCISES:

1) The total hardness of water is equal to 5.25 mmol / L. Calculate how many grams of sodium phosphate should be added to 500 liters of water to remove its hardness.

2) Calculate the water hardness when in 1 liter of water was dissolved 40.04 mg of calcium ions and 24.32 mg magnesium ions.

3) Determine what is the temporary hardness of the water, if the titration of 20 ml of water was spent on average 5.25 ml of 0.1 N hydrochloric acid solution?

4) Why pure water is very weak electrolyte?

a) It ionizes to a large extent.

b) It ionizes to very small extent.

c) It does not ionizes at all.
d) It ionizes completely.
5) Water molecule is a highly polar solvent due to:
a) Presence of two lone pair of electrons.
b) Presence of three lone pair of electrons.
c) Absence of lone pair of electrons.
d) None of these.

LABORATORY WORK
REDOX REACTIONS

Objective: *to investigate the redox reactions and chemical properties of main oxidants.*

Reactions involving oxidation and reduction processes are very important in our everyday world. They make batteries work and cause metals to corrode (or help to prevent their corrosion). They enable us to obtain heat by burning fuels in factories and in our bodies.

Oxidation-Reduction Reactions *are all reactions that involve the change of an oxidation number, and transfer of electrons among the reacting substances.*

Oxidation is the loss of electrons. When a species loses electrons it is said to be oxidized:

$$Fe^{3+} - 1e \rightarrow Fe^{2+}$$

Reduction is the gain of electrons. When a species gains electrons it is said to be reduced:

$$MnO_4^- + 8H^+ + 5e \rightarrow Mn^{2+} + 4H_2O$$

Overall redox equations can be created by combining the half-equations for the oxidation process and reduction processes, after multiplying all the coefficients of the species in one of the half-equations by a factor which ensures that the number of electrons gained is equal to the number of electrons lost.

$Fe^{3+} - 1e \rightarrow Fe^{2+}$	1	5	oxidation
$MnO_4^- + 8H^+ + 5e \rightarrow Mn^{2+} + 4H_2O$	5	1	reduction

Multiplying all coefficients in the oxidation reaction by 5:

$5Fe^{3+} - 5e \rightarrow 5Fe^{2+}$

means that 5 electrons are gained and five are lost
and overall equation:

$MnO_4^- + 8H^+ + 5Fe^{2+} \rightarrow Mn^{2+} + 4H_2O + 5Fe^{3+}$

A species which can accept (+ne) electrons from another species is an **oxidising agent** (Cl_2, $KMnO_4$, $K_2Cr_2O_7$, HNO_3 concentrated, H_2O_2).

Oxidising agents are reduced during redox reactions. MnO_4^- *is the oxidizing agent* in the above reaction.

A species which can donate (–ne) electrons to another species is a **reducing agent** (metals, SO_2, H_2S gas, Na_2SO_3, $SnCl_2$ solution). Reducing agents are oxidised during redox reactions. Fe^{2+} *is the reducing agent* in the above reaction.

The main property of oxidising agent and reducing agent are their equivalent weight (mass) that *react with one mole of electrons in a redox reaction:*

$$Eq_{Ox/\mathrm{Red}} = \frac{M_{Ox/\mathrm{Red}}}{\pm n\bar{e}}$$

Types of Redox reactions

1) **Single-Replacement Redox Reactions: A + BC → AC + B**

- Conventional (Molecular) Equation

$Zn_{(s)} + Cu(NO_3)_{2\,(aq)} \rightarrow Zn(NO_3)_{2(aq)} + Cu_{(s)}$

- Ionic Equation

$Zn_{(s)} + Cu^{+2}_{(aq)} + NO_3^-{}_{(aq)} \rightarrow Zn^{+2}_{(aq)} + NO_3^-{}_{(aq)} + Cu_{(s)}$

- Net Ionic Equation

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

2) **Combination reaction** is a reaction in which two or more substances are combined to form a single product: **A +B + C → ABC**

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

3) **Decomposition reaction** is a reaction in which a single compound reacts to give two or more substances, usually requiring a raise in temperature:

$$\mathbf{ABC \rightarrow A + B + C}$$

$$2KClO_{3\,(s)} \rightarrow 2KCl_{(s)} + 3O_{2\,(g)}$$

4) **Combustion reaction** is a reaction of a substance with oxygen, usually the rapid release of heat produces a flame:

$$CH_{4\,(g)} + 2O_{2\,(g)} \rightarrow CO_{2\,(g)} + 2H_2O_{(g)}$$

5) **Disproportionation reaction** is a intramolecular reaction in which the atoms of one element is reduced and simultaneously increase the degree of oxidation state.

$$\overset{+1}{H_2}\overset{-1}{\mathbf{O_2}} \rightarrow \overset{+1}{H_2}\overset{-2}{\mathbf{O}} + \overset{0}{\mathbf{O_2}}$$

To predict the products of a redox reaction, look at the reagents given to see if there is both an oxidizing agent and a reducing agent. When a problem mentions an acidic or basic solution, it is probably redox.

Common oxidizing agents	**Products formed**
MnO_4^- in acidic solution	Mn^{2+}
MnO_2 in acidic solution	Mn^{2+}

MnO_4^- in neutral or basic solution	MnO_2 (s)
$Cr_2O_7^-$ in acidic solution	Cr^{3+}
HNO_3, concentrated	NO_2
HNO_3, dilute	NO
H_2SO_4, hot, concentrated	SO_2
metalic ions	metalous ions
free halogens	halide ions
Na_2O_2	NaOH
$HC1O_4$	Cl^-
H_2O_2	H_2O

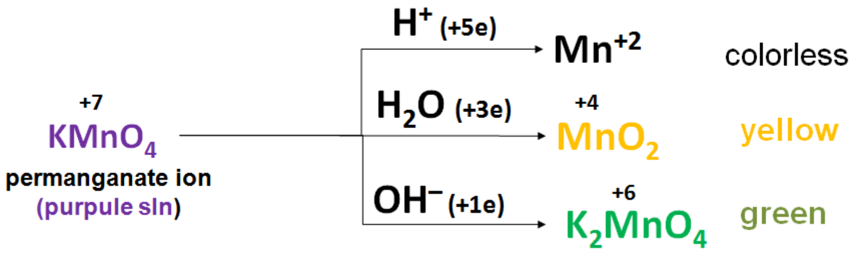
Common reducing agents	**Products formed**
halide ions	free halogen
free metals	metal ions
sulfite ions (SO_3^{2-}) or SO_2	sulfate ions (SO_4^{2-})
nitrite ions (NO_2^-)	nitrate ions (NO_3^-)
free halogens, dilute basic solution	hypohalite ions
free halogens, conc. basic solution	halate ions
metalous ions	metalic ions
H_2O_2	O_2
$C_2O_4^{2-}$	CO_2

1st Experiment. *Investigate oxidizing properties of Potassium Permanganate ($KMnO_4$) depending from pH*

Materials and Reagents: water, solutions of 2N sulfuric acid (H_2SO_4), 0,1N sodium hydroxide (NaOH), 2N potassium permanganate ($KMnO_4$) and sodium sulfite (Na_2SO_3), test tubes and disposable plastic beral pipets

Potassium permanganate consists of dark purple or bronze-like crystals and it is a strong oxidizing agent. It is used extensively in laboratory work and also in dyeing wood, bleaching, photography, and tanning. Dilute solutions are mildly irritating to the skin and high concentrations are caustic. Potassium permanganate stains skin and clothing like silver nitrate.

Oxidizing properties of Potassium permanganate is depended from pH:



PROCEDURE:

1) Take three test tubes and sign them as 1st, 2nd, 3rd

2) Measure by volumetric tube 2 ml KMnO4 solution and pour into all test tubes, then diluted 1st tube with 5 – 7 drops of sulfuric acid, 2nd tube with

5 – 7 drops of distilled water and 3rd tube with 5 – 7 drops of sodium hydroxide.

3) Measure by volumetric tube 2 ml of sodium sulfite (Na_2SO_3) and mix into each tubes with permanganate. At the same time, solutions in tubes begin to change their colors. In acidic solution ($pH<7$) permanganate ion is colorless, in neutral medium ($pH=7$) solution it is yellow and in basic medium ($pH>7$) solution it is green.

4) Write a balanced redox reaction between $KMnO_4$ and Na_2SO_3 in acidic, neutral and basic solutions. Determine the oxidation number for each atom. Identify which substance is oxidized, which substance is reduced, the oxidizing agent, and the reducing agent, and their calculate equivalent masses:

a) $KMnO_4 + H_2SO_4 + Na_2SO_3 \rightarrow$
b) $KMnO_4 + H_2O + Na_2SO_3 \rightarrow$
c) $KMnO_4 + NaOH + Na_2SO_3 \rightarrow$

2nd Experiment. ***Investigate oxidizing and reducing properties of iron (III, II)compounds***

Materials and Reagents: solutions of potassium iodide (KI), iron (III) chloride ($FeCl_3$), 2N sulfuric acid (H_2SO_4), 2N potassium permanganate ($KMnO_4$), iron (II) sulfate ($FeSO_4$), 1% starch solution, test tubes and disposable plastic pipets.

A) Property of compound iron (Fe^{3+})

1) Take two test tubes.

2) In the 1st test tube measure by volumetric tube 2 ml solution of iron (III) chloride ($FeCl_3$).

3) In the 2nd test tube measure by volumetric tube 2 ml solution of potassium iodide (KI).

4) In both test tubes add 2 drops of 1% starch solution (solutions in tubes does not change).

5) Then mix both solutions in tubes: $FeCl_3$ + KI. Solution colouring in dark blue. Why? What is a product of reaction and colour of it?

6) Write a balanced redox reaction between KI and $FeCl_3$ [among the products are I_2 and Fe^{2+}]. Determine the oxidation number for each atom. Identify which substance is oxidized, which substance is reduced, the oxidizing agent, and the reducing agent, and their calculate equivalent masses:

$$FeCl_3 + KI \rightarrow$$

B) Property of compound iron (Fe^{2+})

1) In test tube measure by volumetric tube 2 ml $KMnO_4$ solution, then add 5-6 drops of H_2SO_4 and solution of $FeSO_4$ until solution is colourless.

2) Write a balanced redox reaction between $KMnO_4$ and $FeSO_4$ in acidic solution [among the products are Mn^{2+} and Fe^{3+}]. Determine the oxidation number for each atom. Identify which substance is oxidized, which

substance is reduced, the oxidizing agent, and the reducing agent, and their calculate equivalent masses:

$$KMnO_4 + H_2SO_4 + FeSO_4 \rightarrow$$

3^{rd} Experiment. *Investigate Disproportionation reaction*

Materials and Reagents: solutions of 2N sulfuric acid (H_2SO_4), 0,1N sodium hydroxide (NaOH), iodine water (I_2), test tubes and disposable plastic beral pipets

PROCEDURE:

1) In test tube measure by volumetric tube 2 ml light brown iodine water (I_2), then add 5-6 drops of sodium hydroxide (NaOH) and mixed solution until it is colourless. What happens in solution?

2) Write a balanced redox reaction between I_2 and NaOH. Determine the oxidation number for each atom. Identify which substance is oxidized, which substance is reduced, the oxidizing agent, and the reducing agent, and their calculate equivalent masses:

$$I_2 + NaOH \rightarrow$$

3) Then in that tube add some drops of sulfuric acid (H_2SO_4) until solution colour in light brown again. Why? What forms in solution and colour of it?

4) Write a balanced redox reaction. Determine the oxidation number for each atom. Identify which substance is oxidized, which substance is reduced, the oxidizing agent, and the reducing agent, and their calculate equivalent masses.

EXERCISES:

1. Determine the oxidation number for each atom in the formulas: $Ca_3(PO_4)_2$, $Al_2(SO_3)_3$, $(NH_4)_2S$, P, Cl_2O_7, $H_3[PtCl_6]$, $(Cr_2O_7)^{2-}$, $(CO_3)^{2-}$

2. Write balanced equations for the complete combustion of a) $C_8H_{18(l)}$, b) $CH_3OH_{(l)}$, and c) $C_3H_7SH_{(l)}$.

3. Classify each of these reactions with respect to the following categories: combination reaction, decomposition reaction, combustion reaction, and single-displacement reaction.

a) $2Fe_2O_{3\,(s)} + 3C_{\,(s)} \rightarrow 4Fe_{\,(l)} + 3CO_{2(g)}$

b) $Cl_{2\,(g)} + F_{2\,(g)} \rightarrow 2ClF_{(g)}$

c) $2Pb(NO_3)_{2\,(s)} + 4NO_{2(g)} + 2PbO_{\,(s)} + O_{2\,(g)}$

d) $C_5H_{11}SH_{\,(l)} + 9O_{2\,(g)} \rightarrow 5CO_{2\,(g)} + 6H_2O_{\,(l)} + SO_{2\,(g)}$

4. Electrons are _____________ found unattached to atoms. Thus, for one element or compound to lose electrons and be _____________, another element or compound must be there to gain the electrons and be _____________. In other words, _____________ (loss of electrons) must be accompanied by _____________ (gain of electrons).

5. Reactions in which electrons are _____________, resulting in oxidation and reduction, are called oxidation-reduction reactions.

6. The separate oxidation and reduction equations are called _____________.

7. A(n) _____________ is a substance that loses electrons, making it possible for another substance to gain electrons and be reduced.

8. A(n) _____________ is a substance that gains electrons, making it possible for another substance to lose electrons and be oxidized.

9. Oxidation is defined as the complete or _____________ loss of electrons, reduction as the complete or partial gain of electrons.

10. Just think of oxidation numbers as tools for keeping track of the _____________ in redox reactions.

11. If any element undergoes a(n) _____________ of oxidation number in the course of a reaction, the reaction is a redox reaction. If an element's oxidation number _____________ in a reaction, that element is oxidized. If an element's oxidation number _____________ in a reaction, that element is reduced. The reactant containing the element that is oxidized is the _____________ agent. The reactant containing the element that is reduced is the _____________ agent.

12. In combination reactions, _____________ elements or compounds combine to form one compound.

13. In decomposition reactions, _____________ compound is converted into two or more simpler substances.

14. In a combustion reaction, oxidation is very rapid and is accompanied by _____________ and usually _____________.

15. When any substance that contains carbon is combusted (or burned) completely, the carbon forms _____________.

16. When a substance that contains hydrogen is burned completely, the hydrogen forms _____________.

17. When any substance that contains sulfur burns completely, the sulfur forms _____________.

18. In single-displacement reactions, atoms of one element in a compound are displaced (or replaced) by atoms from a(n) _____________.

19. Determine the oxidation number for each atom, and decide whether the reaction is a redox reaction or not. If it is redox, identify which substance is oxidized, which substance is reduced, the oxidizing agent, and the reducing agent, balanced equation and calculate equivalent masses of oxidant and reductant:

$$KCl + HNO_3 + O_2 \rightarrow KNO_3 + Cl_2 + H_2O$$

20. Perform test

№	*Question*	*Variant*
	Which of the following characterizes the hydrogen.	a. Element losing electrons is losing oxygen, element gaining electrons is gaining oxygen. b. Element gaining electrons is oxidation–reduction relationship? oxidized, element losing electrons

		is reduced. c. Element gaining electrons is losing hydrogen, element losing electrons is gaining d. Element losing electrons is oxidized, element gaining electrons is reduced.
2	What kind of process is represented by this reaction? $Na \rightarrow Na^+ + e^-$	a. oxidation b. reduction c. redox d. neutralization
3	The chemical equation above shows a redox reaction: $Mg + O_2 = 2MgO$ Which of these best represents what has occurred?	a. Magnesium is the oxidizing agent and was reduced. b. Oxygen is the reducing agent and was oxidized. c. Magnesium is the reducing agent and was oxidized. d. Oxygen is the oxidizing agent and was oxidized.
4	$Si + 2Cl_2 = SiCl_4$ The redox reaction above can be used to determine the change in oxidation numbers for chlorine and silicon. According to this equation, what changes occurred?	a. Silicon went from +1 to +4, and chlorine went from +2 to -4. b. Silicon went from 0 to -4, and chlorine went from 0 to +1. c. Silicon went from +1 to -4, and chlorine went from +4 to -4. d. Silicon went from 0 to +4, and chlorine went from 0 to -1.

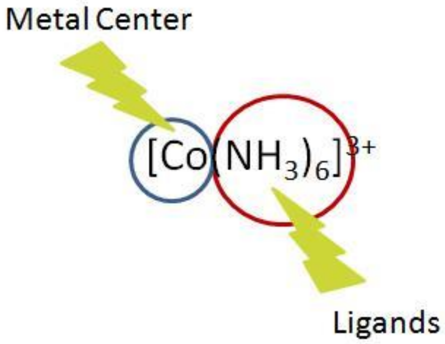
LABORATORY WORK
COORDINATION COMPOUNDS

Objectives: *To prepare a special type of compound called a coordination compound in the laboratory.*

One of the important aspects of any chemical research is the preparation of new compounds. This process is called synthesis. In this experiment you will synthesize a special kind of compound called a coordination compound.

Complexes or **coordination compounds** are molecules that possess a metal center that is bound to ligands (atoms, ions, or molecules that donate electrons to the metal).

Coordination complexes (a complex ion) is comprised of two important parts: the central atom and its surrounding ligands. The central atom can be any



metallic ion (usually a transition metal). The overall charge can be positive, negative, or neutral.

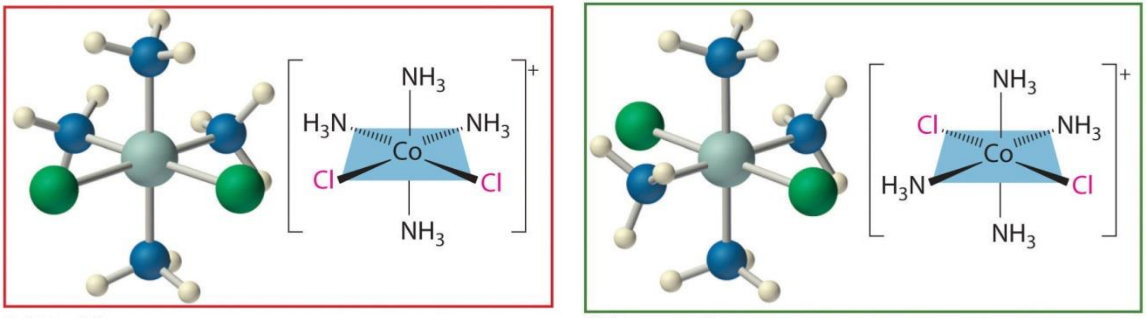
Coordination compounds are complex or contain complex ions, for example:

- Complex Cation: $[CO(NH_3)_6]^{3+}$
- Complex Anion: $[CoCl_4(NH_3)_2]^{-}$
- Neutral Complex: $[CoCl_3(NH_3)_3]^{\circ}$
- Coordination Compound: $K_4[Fe(CN)_6]$

A ligand can be an anion or a neutral molecule that *donates an electron pair* to the complex. Ex: NH_3, H_2O, Cl^-. The number of ligands that attach to a metal depends on whether the ligand is monodentate, bidentate, or polydentate.

A coordination compound consists of a central metal ion which is chemically bonded to one or more atoms or groups of atoms by coordinate covalent bonds. The metal ion contains one or more empty orbitals which can receive pair(s) of electrons and the atom or group of atoms bonded to the metal ion (ligands) contain one or more pairs of electrons which can be donated to the metal ion. When a covalent bond (a bond formed by sharing of one or more pairs of electrons) contains a pair of electrons which comes from only one atom in the bond it is called a coordinate covalent bond.

Coordination compounds are usually brightly colored compounds and some of them play important roles in biological processes; e.g., heme which is found in hemoglobin, is a coordination compound of iron and chlorophyll is a coordination complex of magnesium.



(a) Red form

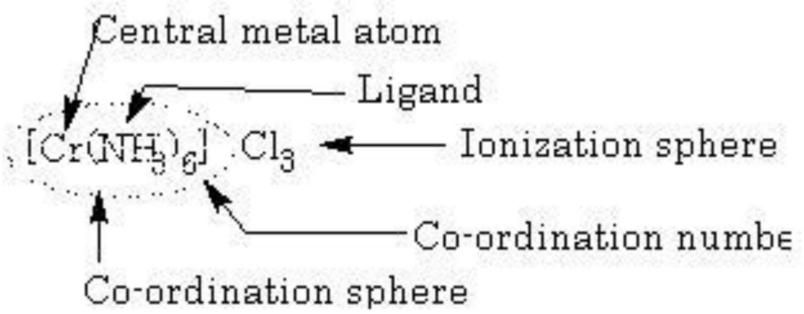
(b) Green form

The modern theory of coordination chemistry is based largely on the work of Alfred Werner (1866–1919; Nobel Prize in Chemistry in 1913).

Central Metal Ion. It is an acceptor atom containing vacant orbitals to which a fixed number of ligands are attached via co-ordinate bonds in definite geometrical arrangement.

Coordination Sphere: Central atom and ligands comprise the *inner coordination sphere*; complex ion enclosed in square bracket, it behaves as a single unit.

Ionization Sphere: Part of compound presents outside coordination sphere, *an outer coordination sphere* constitutes a positively or negatively charged ions that are on more distance from the central ion or associated there with:



Coordination Number: It is the total number of ligands attached to the central metal atom through coordinate bonds or the number of atoms of a ligand attached to the same central atom, e.g. hexadentate ligand should be counted as forming six co-ordinate bonds.

Oxidation number: It is the charge which the central atom appears to have if all the ligands are removed along with the electron pairs that are shared with the central atom.

$[Cr(H_2O)_4Cl_2]NO_3 \Rightarrow [Cr(H_2O)_4Cl_2]^+ + NO_3^-$

$[Cr(H_2O)_4Cl_2]^+$

$x + (4 \times 0) + (-1 \times 2) = +1$ [because the ligand H_2O is neutral and $2Cl^-$ carries - 2 charge]

$x + 0 - 2 = +1$

$x = 3\ (Cr^{3+})$

Ligands. It is an ion or molecule capable of donating a pair of electrons to the central atom via a donor atom.

Types of ligands:

- ***Unidentate ligands:*** Ligands with only one donor atom, e.g. NH_3, Cl^-, F^- etc.
- ***Bidentate ligands:*** Ligands with two donor atoms, e.g. ethylenediamine, $C_2O_4^{2-}$(oxalate ion) etc.
- ***Tridentate ligands:*** Ligands which have three donor atoms per ligand, e.g. (dien) diethyl triamine.
- ***Hexadentate ligands:*** Ligands which have six donor atoms per ligand, e.g. EDTA.
- ***Chelating Ligands:*** Multidentate ligand simultaneously co-ordinating to a metal ion through more than one site is called chelating ligand. These ligands produce a ring like structure called chelate. Chelation increases the stability of complex. This effect is called chelation effect.

Example:

$H_2C — \ddot{N}H_2$ $\quad\quad$ $O = \overline{C} — O:$

$|$ $\quad\quad\quad\quad\quad\quad\quad$ $|$

$H_2C — \ddot{N}H_2$ $\quad\quad$ $O = \overline{C} — O:$

The composition and structure of complex compounds:

Component parts	**Cationic complex $[Cu(NH_3)_4]SO_3$**	**Anionic complex $Na_2[Pt(Br)_6]$**	**Neutral complex $[PtCl_4(NH_3)_2]$**
Complexing agent	Cu^{2+}	Pt^{4+}	Pt^{4+}
Ligand	NH_3	Br^-	Cl^- and NH_3
The inner coordination sphere	$[Cu(NH_3)_4]$	$[Pt(Br)_6]$	$[PtCl_4(NH_3)_2]$
An outer coordination sphere	SO_3^{2-}	Na^+	-
Coordination number	4	6	6
Charge of the complex ion	$[Cu(NH_3)_4]^{2+}$	$[Pt(Br)_6]^{2-}$	$[PtCl_4(NH_3)_2]^{\circ}$

Preparation of complex compounds

Complex compounds prepared by such reactions:

1) Connection reactions: $HgI_2 + 2KI \rightarrow K_2[HgI_4]$
2) Substitution reactions: $[Cu(H_2O)_4]\ SO_4 + 4NH_3 \rightarrow [Cu(NH_3)_4]\ SO_4 + 4H_2O$
3) Exchange reactions: $2\ ZnCl_2 + K_4\ [Fe(CN)_6] \rightarrow Zn_2[Fe(CN)_6] + 4KCl$
4) Redox reactions: $2Al + 6KOH + 6H_2O \rightarrow K_3[Al(OH)_6] + 3H_2\uparrow$

Classification of complex compounds

1) By the charge of the complex ion:

Cationic	**Anionic**	**Neutral**
$[Ag(NH_3)_2]^+Cl$	$Na[Al(OH)_4]^-$	$[Fe(CO)_5]^{\circ}$
$[Cu(NH_3)_4]^{2+}\ SO_4$	$H_2[Pt(Cl)_6]^{2-}$	$[Pt(NH_3)_2Cl_2]^{\circ}$
$[Co(NH_3)_4]^{3+}Cl_3$	$K_3[Fe(CN)_6]^{3-}$	$[Cr(H_2O)_3F_3]^{\circ}$

2) By the nature of the ligand:

Complex name	**Ligand, its name**	**Example of connection**
Ammines	NH_3 — ammine	$[Zn(\mathbf{NH_3})_4]SO_4$
Aquacomplexes	H_2O — aqua	$[Al(\mathbf{H_2O})_6]Cl_3$
Hydroxocomplexes	OH^- – hydroxo	$Na_2[Zn(\mathbf{OH})_4]$
Carbonilocomplex	CO – carbonil	$[Fe(\mathbf{CO})_5]^{\circ}$
Acidocomplexes	Cl^- — chlorine CN^- — cyanide NO_2^- — nitrite CO_3^{2-} — carbonate SCN^- —	$K_3[Al\mathbf{Cl}_6]$ $K_4[Fe(\mathbf{CN})_6]$ $Na_3[Co(\mathbf{NO_2})_6]$ $[Co(NH_3)_4\mathbf{CO_3}]$ $(NH_4)_2[Hg(\mathbf{SCN})_4]$

	thiocyanide SO_4^{2-} — sulphate	$K_2[Be(\mathbf{SO_4})_2]$

There are also *intracomplex* compounds – chelates polynuclear complexes, clathrates, fullerenes. Complex compounds denticity. One characteristic is their denticity ligands. *Denticity* - is the number of seats occupied by ligands in the inner coordination sphere of the complex.

Complex compounds denticity

Denticity	Ligands
monodentate	NH_3, H_2O, OH^-, Cl^-, Br^-, F^-, NO_2^-, SCN^-, CN^-
bidentate	CO_3^{2-}, SO_4^{2-}, $S_2O_3^{2-}$, $C_2O_4^{2-}$
polydentate	Aminopolycarboxylic acid (complexones), proteins

Naming Of Coordination Compounds

Rule 1: To name a co-ordination compound, no matter whether the complex ion is the cation or the anion, always name the cation before the anion.

Rule 2: Part (a) Name the ligand first, in alphabetical order, then the metal atom or ion.

Part (b) For anionic ligands end in "- o -". For anions that end in "- ide -" (e.g. chloride), " – ate -" (e.g. sulfate, nitrate) and " – ite -" (e.g. nitrite), change the ending as follows:-ide ——→ -ido, -ate ——→ -ato, -ite ——→ -ito

Part (c) For neutral ligands, the common name of the molecule is used e.g. $H_2NCH_2CH_2NH_2$ (ethylenediamine). Important exceptions: For H_2O ——→ aqua, NH_3 ——→ ammine, CO ——→ carbonyl and N_2 and O_2 are called dinitrogen and dioxygen.

Anionic ligands	Neutral ligands
Br^- - Bromo	NH_3 – ammine
F^- - Fluoro	H_2O – aqua
O^{2-} - Oxo	NO – nitronyl
OH^- - Hydroxo	CO – Carbonyl
CN^- - Cyano	O_2 – Dioxygen
$C_2O_4^{2-}$ - Oxalato	N_2 – Dinitrogen
CO_3^{2-} - Carbonato	H_2NH_2C | H_2NH_2C — ethylene diamine
CH_3COO^- - aceto	

Name of Metals in Anionic Complexes:

Iron (Fe^{2+})	Ferrate	Silver (Ag^+)	Agrenate
Copper (Cu^{2+})	Cuprate	Gold (Au^{3+})	Aurate
Lead (Pb^{2+})	Plambate	Tin (Sn^{2+})	Stannat

Rule 3: Greek prefixes are used to designate the number of each type of ligand in the complex ion, di-, tri- and tetra-. If the ligand already contains a

Greek prefix (e.g. ethylendiamine) if it is polydenate ligands the prefixes, bis-, tris-, tetrakis-, pentakis- are used instead.

Rule 4: After naming the ligand, name the central metal. If the complex ion is a cation, metal is named same as the element. But if the complex is an anion, the name of metal ends with the suffix –ate- for Latin name.

Rule 5: The oxidation state of the metal in the complex is given as a roman numeral in parentheses.

For example: Name the following coordination complexes

(a) $[Cr(NH_3)_5(H_2O)][(NO_3)_3]$ (b) $[Cr(NH_3)_4Cl_2]Cl$

Solution:

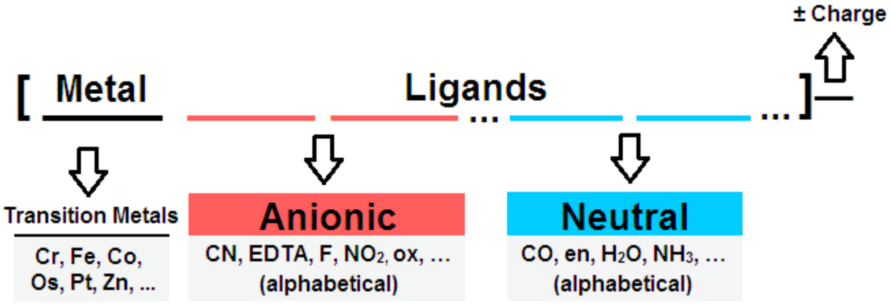
(a) The complex contains the $[Cr(NH_3)_5(H_2O)]^{3-}$ ion and three nitrate ions to balance the charge on the complex ion. It is therefore known as pentaammineaquochromium (III) nitrate.

(b) Two of the three chloride ions are ligands that are coordinated to the Cr^{3+} ion, and the other is a Cl^- ion that balances the charge on the $[Cr(NH_3)_4Cl_2]^+$ ion. The compound is therefore known as dichlorotetraamminechromium (III) chloride.

For historic reasons, some coordination compounds are called by their common names. For example, $[Fe(CN)_6]^{3-}$ and $[Fe(CN)_6]^{4-}$ are named ferricyanide and ferrocyanide respectively, and $[Fe(CO)_5]$ is called iron carbonyl.

Writing Formulas of Coordination Complexes

The formula of a coordination complex is written in a different order than its name. The chemical symbol of the metal center is written first. The ligands are written next, with anion ligands coming before neutral ligands. If there is more than one anion or neutral ligand, they are written in alphabetical order according to the first letter in their chemical formula.



In a coordination compound's name, when one of the ions is just an element, the number of atoms is not indicated with a prefix. Since it still has to be written in the formula, it is determined by balancing the overall charge of the compound.

For example: Determine the formulas of the following compounds.

(a) potassium hexacyanoferrate (II)

(b) tris(ethylenediamine)chromium (III) chloride

Solution:

(a) The compound contains enough K^+ ions to balance the charge on the negatively charged ferrate ion. Because the ferrate ion is formed from an Fe^{2+} ion and six CN^- ligands, it carries a net charge of -4. The formula for the compound is therefore $K_4[Fe(CN)_6]$.

(b) The compound contains a Cr^{3+} ion coordinated to three bidentate ethylenediamine or “en” ligands. It therefore must contain three Cl^- ions to balance the charge on the complex ion. Thus, the formula for the compound is $[Cr(en)_3]Cl_3$.

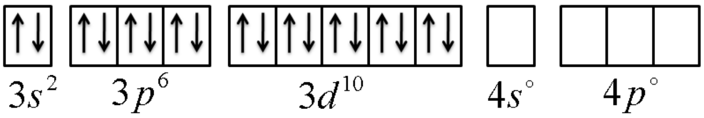
Nature of chemical bonding in complex compounds

Structure physico-chemical and biological properties of complex compounds depend on the nature of chemical bonds in them. Currently, the nature of chemical bonds in complex compounds such theories explain:

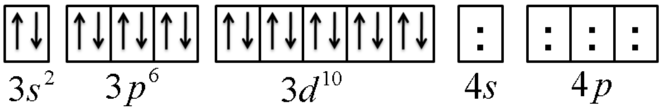
1) the method of valence bonds ;
2) the crystal field theory ;
3) the method of molecular orbitals .

Let us consider the simpler *method of valence bonds.* Under this method, the connection between the central atom and the ligand is formed by the donor- acceptor mechanism. The donor is a ligand that donates a lone electron pair acceptor - Central heating atom, which has free orbitals.

For example, during the formation of the ion $[Zn(NH_3)_4]^{2+}$ is a donor molecule of ammonia, in which the nitrogen atom has an unshared electron pair. Acceptor is a zinc atom, which has free d - orbitals. Electron- graphic formula of the third and fourth energy levels of zinc ions Zn^{+2} has the form:



Lone pairs of the nitrogen atoms of four ammonia molecules located on the ligand-free 4s-and 4p- orbitals of the zinc ion:



Properties of complex compounds

1) *Dissociation* of complex compounds.

Complex compounds are strong electrolytes. In aqueous solution, they readily dissociate to a complex ion and the outer sphere. This is called the primary dissociation. Examples of primary dissociation of complexes are given in ***Table***

Dissociation of the complex compounds

Types of complex connections	**Dissociation equation**
Cationic complex	$[Ag(NH_3)_2]Cl \leftrightarrow [Ag(NH_3)_2]^+ + Cl^-$

	$[Cr(H_2O)_6]Cl_3 \leftrightarrow [Cr(H_2O)_6]^{3+} + 3Cl^-$
Anionic complex	$K_4[Fe(CN)_6] \leftrightarrow 4K^+ + [Fe(CN)_6]^{4-}$ $Na_2[Zn(OH)_4] \leftrightarrow 2Na^+ + [Zn(OH)_4]^{2+}$

The formed complex ion may also dissociate, albeit weaker. This *secondary* dissociation, which is stepwise:

I step $[Ag(NH_3)_2]^+ \leftrightarrow Ag(NH_3)^+ + NH_3$

II step $Ag(NH_3)^+ \leftrightarrow Ag^+ + NH_3$

We can write the equation in a secondary dissociation step of:

$$[Ag(NH_3)_2]^+ \leftrightarrow Ag^+ + 2NH_3$$

2) *Stability* of complex compounds. In general, chemical equilibrium is reached when the forward reaction rate is equal to the reverse reaction rate and can be described using an equilibrium constant, *K*. Complex ion equilibria are no exception to this and have their own unique equilibrium constant. This *formation constant* (K_f), describes the formation of a complex ion from its central ion and attached ligands. This constant may be called a **stability constant** or **association constant**; the units depend on the specific reaction it is describing.

At its most basic level, K_f can be explained as the following, where *M* is a metal ion, *L* is a ligand, and *x* and *y* are coefficients:

$$K_f = \frac{[M_x L_y]}{x[M^{+n}] \cdot y[L]}$$

K_f values tend to be very large in magnitude because complex ion formation reactions heavily favor the product.

For example, observe the formation reaction of the dicyanoargentate (I) ion and the resulting formation constant expression:

$$K[Ag(CN)_2] = K^+ + [Ag(CN)_2]^-$$

$$[Ag(CN)_2]^- = Ag^+ + 2CN^-$$

$$K_f = \frac{[Ag(CN)_2]^-}{[Ag^+] \cdot [CN^-]^2} = 5{,}6 \cdot 10^{18}$$

The larger the K_f value of a complex ion, the more stable it is. Due to how large formation constants often are it is not uncommon to see them listed as logarithms in the form $\log K_f$. You may also see them in the form of a *dissociation constant*, K_d (which should not be confused with pK_a). This is merely the inverse value of the formation constant and such describes the instability of a complex ion. It may help to think of K_f and K_d as stability and instability constants, respectively.

Using the above example, the dissociation constant expression would be:

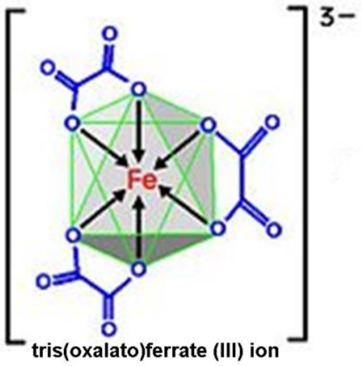
$$K_f = \frac{1}{K_d}$$

$$K_f = \frac{[Ag(CN)_2]^-}{[Ag^+]\cdot[CN^-]^2} = 5{,}6\cdot10^{18}$$

$$K_d = \frac{[Ag^+]\cdot[CN^-]^2}{[Ag(CN)_2]^-} = 1{,}8\cdot10^{-19}$$

In logarithm form K_f and K_d would be: $\log K_f = \log(5.6\times1018) = 18.7$
$\log K_d = \log(1.8\times10-19) = -18.7$

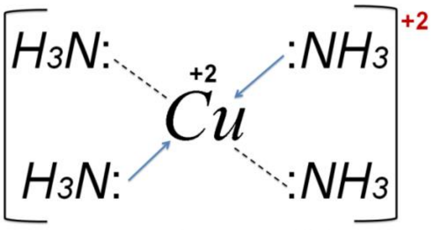
Today we will synthesize two of these compounds. One of them is a type of coordination compound called a chelate. In a chelate, the ligand has more than one point of attachment to the metal atom or ion. The ligand is called polydentate ("many toothed"). Three oxalate ions attach to an iron ion to give:



tris(oxalato)ferrate (III) ion

The oxalate ion, $C_2O_4^{2-}$, has two points of attachment to the iron ion, so it is bidentate. Three oxalate ions arrange themselves about the iron ion in an octahedral geometry. Note that the iron ion must have a 3+ charge in this substance in order to account for the 3– charge on the ion. The iron chelate will be crystallized out of solution as the potassium salt.

The other coordination compound you will prepare has a unidentate ligand, ammonia, attached to a copper ion in four positions:



tetramminecopper (II) ion

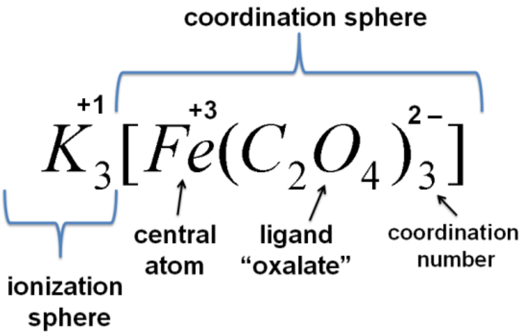
The nitrogen in ammonia covalently bonds to the copper using its unbonded pair of electrons. Four ammonia molecules arrange themselves about the copper ion in a square planar geometry. The copper has a 2+ charge in this substance. The copper complex will be crystallized out of solution as the sulfate salt.

1st Experiment. Synthesis of Potassium tris(oxalato)ferrate (III)

The reaction is:

$$FeCl_3 + 3K_2C_2O_4 \rightarrow K_3[Fe(C_2O_4)_3] + 3KCl$$

All of the substances are in aqueous solution. You will allow the green complex iron salt to precipitate by forming a mixed solvent. In this case, the mixed solvent is created by adding ethanol to a water solution of the iron complex. When the complex ion salt precipitates, it carries three molecules of water with it, so that the solid crystals have the formula: $K_3[Fe(C_2O_4)_3]\cdot 3H_2O$.



Procedure:

While you are making the iron compound, you will also prepare a solution of copper sulfate that will be used in the 2nd experiment.

Set up a hot water bath.

Add water to the beaker to about the 150 ml mark. Heat the water in the beaker to boiling. Once it boils, reduce the flame to low heat so that the water simmers.

For Part 2 of the experiment, weigh out 1.0 gram of $CuSO_4*5H_2O$ on a tared piece of paper. First fold the paper to put a crease in it, and then unfold it. The crease will help to pour the solid into a test tube. Add the solid to a test tube. Add 2.5 ml of deionized water to the test tube and place it in the hot water bath.

For Part 1, use the transfer pipets on the reagent bottles to add 1 ml of 1 M $FeCl_3$ solution and 2 ml of a saturated $K_2C_2O_4$ solution to the other test tube and place the test tube in the hot water bath. Leave it in the bath for 5 minutes.

Stir the contents of the copper sulfate test tube for Part 2 until all of the solid is dissolved, then remove the test tube from the bath and set it in the test tube rack.

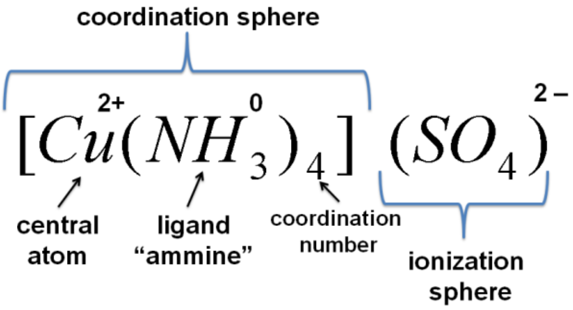
Turn the bunsen burner off. Use the transfer pipet on the reagent bottle to add 1.5 ml of ethanol to the 50 ml glass beaker. Do not bring ethanol to your work station until the burner is off. Pour the hot, green contents of the test tube still in the hot water bath into the beaker and stir. Put the watch glass on the beaker and set the beaker out of the way. The watch glass goes on curved side down. Crystals should form while you work on Part 2. If no crystals appear within 15 minutes, add an inch of ice to a 250 ml beaker and set the 50 ml beaker on top of the ice. Rub the stirrer across the inside bottom of the beaker a few times. This can seed the solution with nanometer sized chunks of glass that often initiate crystal growth. The crystals should grow for at least 30 minutes.

While you are waiting, go on to Part 2 of the experiment.

Then decant the liquid from the crystals into the waste beaker. Add 2 ml of ethanol to rinse the crystals. Swirl, and decant the liquid from the rinse into the waste beaker. Use the rubber policeman to transfer the crystals onto a piece of filter paper setting on a double folded piece of paper towel. Spread the crystals out and allow to dry. Weigh the crystals on a piece of weighing paper. Put them in the container provided and keep them. Label the container with the formula and weight of the substance, as well as your name and the date. Use ink on white tape for the label. Show the sample to the instructor when you turn in your lab. Discard liquid from the waste beaker into the container in the fume hood marked “Iron Waste”.

The iron compound is light sensitive. Over time, if it is kept in the light, it will turn a rusty brown.

2nd Experiment. Synthesis of Tetramminecopper (II) sulfate



The compound which will be prepared in this experiment is tetraamminecopper (II) sulfate monohydrate, a coordination compound which is also a hydrate. A hydrate is a compound which contains loosely bound water (called water of hydration) in its structure. You can recognize the formula of a hydrate by the dot with a certain number of water molecules after which follows the formula for the anhydrous compound (the compound without the water. The equation for the overall reaction is: $CuSO_4 + 4NH_3 + H_2O \text{ --> } [Cu(NH_3)_4]SO_4 \cdot H_2O$

The synthesis will be carried out in three steps:

1) The anhydrous copper (II) sulfate ($CuSO_4$) will be dissolved in water:

$$CuSO_{4(s)} \text{ ---> } Cu^{2+}_{(aq)} + SO_4{}^{2-}{}_{(aq)}$$

2) The dissolved copper (II) ion (in step one) will be reacted with ammonia to form the complex ion $[Cu(NH_3)_4]^{2+}$:

$$Cu^{2+}_{(aq)} + 4NH_{3(aq)} + SO_4{}^{2-}{}_{(aq)} \text{ ---> } [Cu(NH_3)_4]^{2+}_{(aq)} + SO_4{}^{2-}{}_{(aq)}$$

3) The final product will be precipitated with ethanol in which it is insoluble:

$$[Cu(NH_3)_4]^{2+}_{(aq)} + SO_4{}^{2-}{}_{(aq)} + H_2O \;\overset{\text{ethanol}}{\text{--->}}\; [Cu(NH_3)_4]SO_4 \cdot H_2O_{(s)}$$

The reaction equation for step (2) indicates that one mole of Cu2+, which comes from one mole of CuSO4, reacts with four moles of NH3 (ammonia) to form one mole of the tetraamminecopper (II) ion. One mole of coordination compound is produced in step (3) from one mole of the tetraamminecopper (II) ion; therefore, in the overall reaction, one mole of coordination compound is obtained for every one mole of copper (II) sulfate reacted. The amount of ammonia required to react with any given amount of copper (II) sulfate can be calculated from the equation in step (2).

Safety Disposal Methods: Addition of ammonia should be performed under the hood. Disposal of the filtrate should be done in the drains under the hood with lots of water running. The precipitate should be disposed of in the labeled container under the hood.

PROCEDURE:

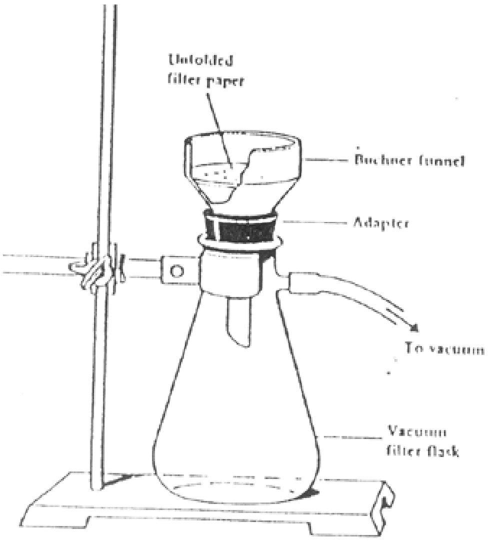
Part One. Preparation of the compound.

1. After zeroing the top-loader balance, place a clean, dry 100 mL beaker on the pan and press the bar to set the balance at 0.00 g again (this is called taring the balance). Using the spatula provided, add small amounts of anhydrous copper (II) sulfate to the beaker until you have exactly 2.00 g (to the nearest 0.01 g). Record the weight of the copper (II) sulfate on your report sheet.

2. Add 10 mL of distilled water to the beaker and stir to dissolve the copper (II) sulfate. Leave the stirring rod in the solution. Note the color of the solution, which is due to the copper (II) ion, on your report sheet.

3. Under the hood, add the required volume (as designated by the instructor) of concentrated ammonia to the beaker and stir until any precipitate is completely dissolved. The first addition of ammonia may cause a light blue precipitate of copper (II) hydroxide to form, but this should dissolve to form the tetraammine-copper (II) ion upon further addition of ammonia. Note any color changes on your report sheet which occur during the addition of ammonia. The color of the solution is due to the presence of tetraamminecopper (II) ion.

4. While stirring the solution in the beaker, add 20 mL of ethanol. Fill a 1000 mL beaker half full of crushed ice, position the beaker containing your compound in the crushed ice, and allow its contents to cool for about 10 minutes. Record your observations during the cooling on your report sheet.



Part Two. Isolation of the Compound by Vacuum Filtration.

5. Set up a vacuum filtration flask as shown in the figure to the right. Place a piece of filter paper into the Buchner funnel so that all the holes are covered. Moisten the paper with water. Connect the vacuum tubing from the filter flask to the vacuum outlet at your desk. Turn the vacuum handle until it is in the completely open position. Press down on

the Buchner funnel to set the adapter in tightly and ensure a good vacuum.

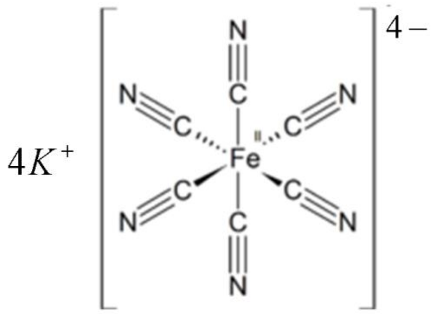
6. Pour the cold solution in the beaker and the precipitate into the Buchner funnel. When all the liquid has passed through, turn off the vacuum.

7. Pour 5 mL of ethanol into the beaker and rinse any remaining precipitate into the Buchner funnel. Stir CAREFULLY to avoid tearing the filter paper. Turn on the vacuum again and pull the liquid through. Turn off the vacuum and wash the precipitate by adding 5 mL of acetone to the funnel, stirring carefully. Turn on the vacuum and pull the wash liquid through. Repeat the washing one more time with a second 5 mL portion of acetone, turning off the vacuum between washings. After the second washing, allow the suction to continue for 2-3 minutes to speed the during of the product.

8. Remove the product from the Buchner funnel with your scoopula and place it on a paper towel. Spread the product out on the towel and break up any lumps with a stirring rod. Allow the product to air-dry for 15 minutes. Dispose of the liquid in the filter flask (called the filtrate) in the drain under the hood.

9. Tare the top-loader balance with a clean, dry 100 mL beaker and then scrape the dried product into the beaker. Record the weight of the product on your report sheet.

3rd Experiment. Properties of Potassium hexacianoferrate (II)

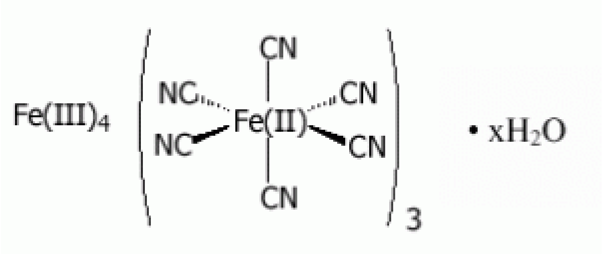


Potassium hexacyanoferrate (II)

Potassium ferrocyanide (hexacianoferrate) is the inorganic compound with formula $K_4[Fe(CN)_6] \cdot 3H_2O$. It is the potassium salt of the coordination complex $[Fe(CN)_6]^{4-}$. This salt forms lemon-yellow monoclinic crystals.

It is a famous reaction which involves treatment with ferric salts to give Prussian blue. Prussian blue is a dark blue pigment with the idealized formula $Fe_7(CN)_{18}$. To better understand the binding situation in this complex compound the formula can also be written as $Fe_4[Fe(CN)_6]_3 \cdot$ x H_2O. Another name for the color is Berlin blue or, in painting, Parisian or Paris blue.

It was one of the first synthetic pigments used by humans. It is employed as a very fine colloidal dispersion, as the compound itself is not soluble in water. In medicine, Prussian blue is used as an antidote for certain kinds of heavy metal poisoning, e.g., by cesium and thallium. In particular it was used to absorb $137Cs^+$ from those poisoned in the Goiania accident. Prussian Blue is orally administered. The therapy exploits Prussian Blue’s ion exchange properties and high affinity for certain “soft” metal cations. Here is the molecular structure of a colloidal dispersion of Prussian Blue:



Prussian blue is formed as a colored precipitate when two solutions of ionic, inorganic compounds are mixed together. There are several different inorganic compounds, which can be used to produce this pigment. In this experiment, the reaction of ferric chloride ($FeCl_3$) and potassium ferrocyanide ($(K_4(Fe(CN)_6)$) will produce Prussian blue:

$$4FeCl_3 + 3K_4[Fe(CN)_6] \rightarrow \mathbf{Fe_4[Fe(CN)_6]_3} + 12KCl$$

intense blue

Reagents: Iron (III) Chloride ($FeCl_3$) and Potassium Ferrocyanide ($K_4[Fe(CN)_6]$)

Equipment: Balance, Buchner funnel; Filter flask; Tubing; Filter-vac rubber rings (for suction flasks); Scissors and Filter paper

Safety Notes: $FeCl_3$ is an irritant (used to deodorize sewage!) and is hygroscopic. Please keep the bottle tightly capped.

PROCEDURE:

1. Prepare a saturated solution of iron (III) chloride by placing 3.7 g $FeCl_3$ in a small beaker with 5 mL distilled water. Use a graduated cylinder to measure the volume of water you used. Stir to dissolve.

2. Separately, prepare a saturated solution of potassium ferrocyanide by placing 1.39 g $K_4[Fe(CN)_6]$ in another beaker with 5 mL distilled water. Describe the appearance of these solutions in your LNJ.

3. Make the Prussian Blue by pouring the potassium ferrocyanide solution into the beaker with the ferric chloride solution. Stir with a glass rod. Describe in your notebook EXACTLY what you see happening when the solutions are mixed.

4. Obtain filter paper that fits the Buchner funnel (so it lays flat in the bottom). Set up the aspirator by connecting a piece of tubing from the little side-arm on the filter flask to the similar arm sticking out from the side of the faucet. Put the neck of the Buchner funnel through a Filter-vac rubber ring, then seat it into the top of the filter flask. If your filter flask is small, you may want to clamp it into place.

5. Turn on the faucet to create suction. Then pour your reaction mixture into the funnel. Scrape the entire blue product into the funnel: use a little distilled water to rinse the beaker.

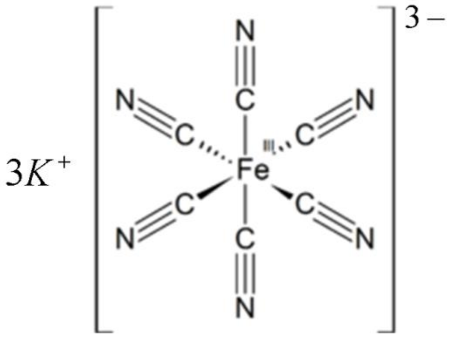
6. Once the liquid has all drained into the flask, gently remove the hose from the filter flask to break the suction. Turn off the water. Put your filter paper on a few paper towels inside your lab drawer, and let your Prussian

Blue dry until the next lab period. Describe the appearance of the pigment, both while it is wet and after it dries.

Reflections:

1. Write the reaction in ionic form.
2. Decribe structure of complex Prussian Blue and name it.
3. Find several other purposes for this substance, besides as a colorant in paints.

4th Experiment. Properties of Potassium hexacyanoferrate (III)



Potassium hexacyanoferrate (III)

Potassium ferricyanide is the chemical compound with the formula $K_3[Fe(CN)_6]$. This bright red salt contains the octahedrally coordinated $[Fe(CN)_6]^{3-}$ ion. It is soluble in water and its solution shows some green-yellow fluorescence.

Ferricyanide $[Fe(CN)_6]^{3-}$ ion is used to make Turnbull's blue.

In former times, the addition of Fe(II) salts to a solution of ferricyanide was thought to afford a material different from Prussian blue. The product was traditionally named Turnbull's Blue (TB). It has been shown, however, by means of X-ray diffraction and electron diffraction methods, that the structures of PB (Prussian Blue) and TB are identical. The differences in the colors for TB and PB reflect subtle differences in the method of precipitation, which strongly affects particle size and impurity content:

$$3FeSO_4 + 2K_3[Fe(CN)_6] \rightarrow \underset{\textbf{deep blue}}{\mathbf{Fe_3[Fe(CN)_6]_2}} + 3K_2SO_4$$

Reagents: Iron (II) sulphate ($FeSO_4$) and Potassium Ferricyanide ($K_3[Fe(CN)_6]$)

Equipment: Balance, Buchner funnel; Filter flask; Tubing; Filter-vac rubber rings (for suction flasks); Scissors and Filter paper

Safety Notes: Potassium ferricyanide; Low toxicity as long as it is not heated, it will release cyanide gas.

PROCEDURE:

1. Weigh accurately about 1.5 g of $FeSO_4$ in a small beaker labeled Turnbull blue and mix with 5 mL of distilled water. Mix the solution until it is all dissolved.

2. Weigh accurately 0.6 g of potassium ferricyanide, $K_3[Fe(CN)_6]$ into a small beaker and mix with 5 mL of distilled water . Stir until dissolved.

3. Slowly pour the solution of potassium ferrocyanide into the beaker labelled Turnbull Blue and stir the mixture. Let the reaction continue and record the results you see.

4. Weigh accurately a small filter paper and place in the filtration apparatus. Filter the precipitate of TB that forms using the house vacuum.

5. Allow the pigment to dry uncovered on the filter paper until the next lab period in your lab drawer on top of a paper towel.

6. Weigh the dried pigment and filter and subtract the initial weight of the filter paper to obtain the weight of the pigment.

Reflections:

1. Write the reaction in ionic form.
2. Decribe structure of complex Prussian Blue and name it.
3. Find several other purposes for this substance, besides as a colorant in paints.

EXERCISES:

1. Give the systematic names for the following coordination compounds:

a) $[Cr(NH_3)_3(H_2O)_3]Cl_3$
b) $[Pt(NH_3)_5Cl]Br_3$
c) $[Pt(H_2NCH_2CH_2NH_2)_2Cl_2]Cl_2$
d) $[Co(H_2NCH_2CH_2NH_2)_3]_2(SO_4)_3$
e) $K_4[Fe(CN)_6]$
f) $Na_2[NiCl_4]$
g) $[Pt(NH_3)_2Cl_4]$
h) $[Pt(NH_3)_2Cl_2]Cl_2$
i) $Fe(CO)_5]$
g) $(NH_4)_2[Ni(C_2O_4)_2(H_2O)_2]$
k) $[Ag(NH_3)_2][Ag(CN)_2]$
l) $[CoBr(NH_3)_5]SO_4$
m) $[Fe(NH_3)_6][Cr(CN)_6]$
n) $[Co(SO_4)(NH_3)_5]^+$
o) $[Fe(OH)(H_2O)_5]^{2+}$

2. Can you give the molecular formulas of the following coordination compounds?

a) hexaammineiron (III) nitrate
b) ammonium tetrachlorocuprate (II)
c) sodium monochloropentacyanoferrate (III)
d) potassium hexafluorocobaltate (III)

3. Can you give the name of the following coordination compounds?

4. Which of the following is the correct oxidation state of the gold atom in $K[Au(CN)_2]$

a) +2 b) +1 c) -1 d) -2

5. Provide the proper name for the coordination complex $[Co(NH_3)_5Cl]SO_4$

a) pentaamminechloridocobalt (III) sulfate
b) pentammoniachloridocobalt (III) sulfate
c) sulfoxidepentaamminechloridocobalt (III)
d) pentaamminechloridocobalt (II) sulfate

6. All of the following properties may be affected by cis/trans isomerization of a coordination complex EXCEPT:
a) color of the complex in solution
b) oxidation state of the metal
c) boiling point
d) melting point

7. Which of the following is an example of structural isomerization of a coordination complex?
a) cis to trans isomerization
b) linkage isomerization
c) Λ to Δ isomerization
d) mer to fac isomerization

8. Which of the following is an example of a bidentate ligand?
a) EDTA b) chloride c) ethylenediamine d) carbon monoxide

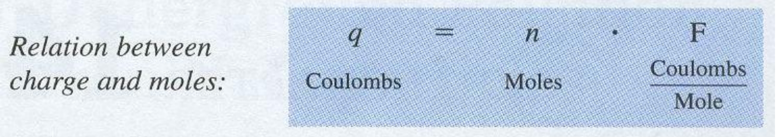
9. Calculate the coordination number in the complex $[Cd(en)_2(CN)_2]$. Note that en = ethylenediamint
a) eight b) six c) four d) two

LABORATORY WORK
ELECTROCHEMICAL CELL TYPES

Learning Objectives:

- Properties and nature of cell will be understood. The differences between electrolytic and electrochemical cells will be learnt. Interconversion of electrical and chemical energy will be emphasized.
- Metal – metal electrode and standard hydrogen electrode construction, the electrode potential and Nernst equation will be learnt.
- Construction of cell with specific electrodes and the EMF calculations will be studied. Daniell cell and it′s representation. EMF of cell from two half cell potentials will also be calculated.
- Complete electrochemical cell representation and writing oxidation and reduction half cell reactions will be studied with suitable example.

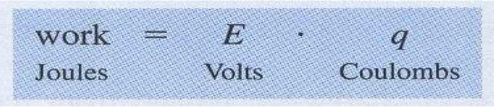
Electrochemistry: the study of the interchange of chemical and electrical energy. Electric charge (**q**) is measured in coulombs (C). The magnitude of the charge of a single electron (or proton) is 1.602×10^{-19} C. A mole of electrons therefore has a charge of $(1.602\times10^{-19}\ C)*(6.022\times10^{23}/mol) = 9.649\times10^{4}$ C/mol, which is called the Faraday constant, F.



Relation between charge and moles:

$$\underset{\text{Coulombs}}{q} = \underset{\text{Moles}}{n} \cdot \underset{\frac{\text{Coulombs}}{\text{Mole}}}{F}$$

The difference in electric potential between two points measures the work that is needed (or can be done) when electrons move from one point to another:



$$\underset{\text{Joules}}{\text{work}} = \underset{\text{Volts}}{E} \cdot \underset{\text{Coulombs}}{q}$$

An electrochemical cell is a system consisting of electrodes that dip into an electrolyte and in which a chemical reaction either uses or generates an electric current, these reactions are called electrochemical, they happen on electrode-solution surface.

Electrode is the material: a metallic rod / bar / strip which conducts electrons into and out of a solution.

An electrode in an electrochemical cell is referred to as either an *anode* or a *cathode*:

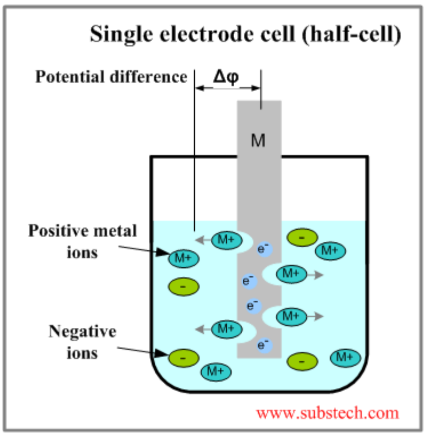
- **Anode:** is the electrode where oxidation takes place, it is the negative (-) electrode (example, active metals are soluble in water).
- **Cathode:** is the electrode where reduction takes place, it is the positive (+) electrode (example, passive metals are insoluble in water).

A half cell is a structure that contains a conductive electrode and a surrounding conductive electrolyte separated by a naturally-occurring Helmholtz double layer. Chemical reactions within this layer momentarily pump electric charges between the electrode and the electrolyte, resulting in a potential difference between the electrode and the electrolyte. The typical reaction involves a metal atom in the electrode being dissolved and transported as a positive ion across the double layer, causing the electrolyte to acquire a net positive charge while the electrode acquires a net negative charge. The growing potential difference creates an intense electric field within the double layer, and the potential rises in value until the field halts the net charge-pumping reactions.

A standard half cell, used in electrochemistry, consists of a metal electrode in a 1 molar (1 mol/L) aqueous solution of the metal's salt, at 298 kelvin (25 °C). The electrochemical series, which consists of standard electrode potentials and is closely related to the reactivity series, was generated by measuring the between the metal half cell in a circuit with a standard hydrogen half cell, connected by a salt bridge.

1st Experiment. Construction of a half cell (metallic electrode)

If a piece of a metal M is dipped into a solution capable to dissolve it the metal begins to oxidize giving the electrons e^- to the metal specimen and forming positive ions M^{+n} transferring to the electrolyte solution. As a result a potential difference (Δφ) between the metal piece and electrolyte forms, it is called ***electrode potential*** ([E] = V).



An electrode potential is affected to the nature of electrode, M^{+n} ions concentration in electrolyte solution and temperature. This relationship may be described by ***Nernst equation*** (at room temperature 25°C):

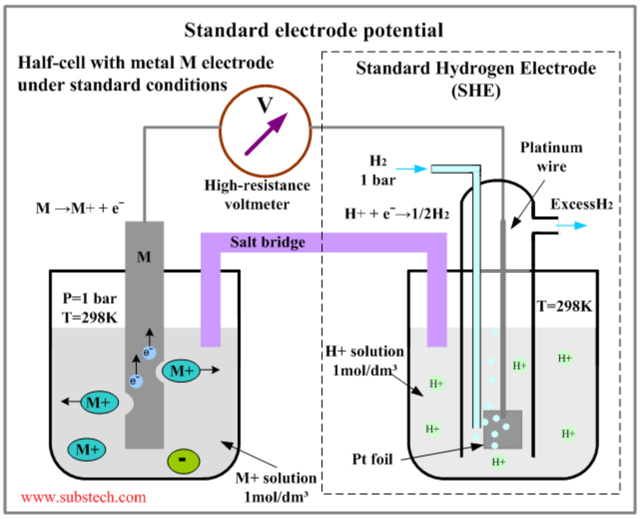
$$E_{Me} = E^{\circ}_{Me^{+n}} + \frac{0,059}{n}\cdot \lg C_{M(Me^{+n})}$$

where **E°** is a ***Standard Electrode Potential.*** SEP is a measure of the tendency of the metallic electrode to loose or gain electrons when it (metal electrode) is dipped in its own salt solution of unit concentration (1M), at 25°C and atmospheric pressure (1 atm = 101,325kPa).

The absolute value of the potential difference cannot be measured since the measurement would mean inserting another electrode into the electrolyte and formation another potential difference between them. For relative measurements of potentials of various metals in various solutions galvanic cell is used (Figure 2).

Using the standard hydrogen electrode

The standard hydrogen electrode is attached to the electrode system you are investigating - for example, a piece of magnesium in a solution containing magnesium ions.



The whole of this set-up is described as a cell. It is a simple system which generates a voltage. Each of the two beakers and their contents are described as half cells.

The salt bridge is included to complete the electrical circuit but without introducing any more bits of metal into the system. It is just a glass tube filled with an electrolyte like potassium nitrate solution. The ends are "stoppered" by bits of cotton wool. This stops too much mixing of the contents of the salt bridge with the contents of the two beakers.

The electrolyte in the salt bridge is chosen so that it doesn't react with the contents of either beaker.

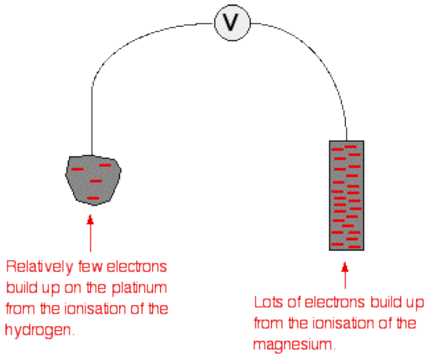
What happens? These two equilibria are set up on the two electrodes (the magnesium and the porous platinum):

$$Mg^{\circ} - 2\bar{e} \to Mg^{2+}$$

$$2H^{+} + 2\bar{e} \to H_2^{\circ}$$

Magnesium has a much greater tendency to form its ions than hydrogen does. The position of the magnesium equilibrium will be well to the left of that of the hydrogen equilibrium.

That means that there will be a much greater build-up of electrons on the piece of magnesium than on the platinum. Stripping all the rest of the diagram out, apart from the essential bits:



There is a major difference between the charge on the two electrodes - a potential difference which can be measured with a voltmeter. The voltage measured would be ΔE=2.37 volts and the voltmeter would show the magnesium as the negative electrode (– anode) and the hydrogen electrode as being positive (+ cathode).

This sometimes confuses people! Obviously, the platinum in the hydrogen electrode isn't positive in real terms - there is a slight excess of electrons built up on it. But voltmeters don't deal in absolute terms - they simply measure a difference:

$$\Delta E = E_{cathode\ (+)} - E_{anode\ (-)} = E^{\circ}_{H_2} - E^{\circ}_{Mg}$$

$$2{,}37 = 0 - E^{\circ}_{Mg}$$

$$E^{\circ}_{Mg} = 0 - 2{,}37 = -2{,}37\ V$$

The magnesium has the greater amount of negativeness - the voltmeter records that as negative. The platinum of the hydrogen electrode isn't as negative - it is relatively more positive. The voltmeter records it as positive.

E° values give you a way of comparing the positions of equilibrium when these elements lose electrons to form ions in solution:

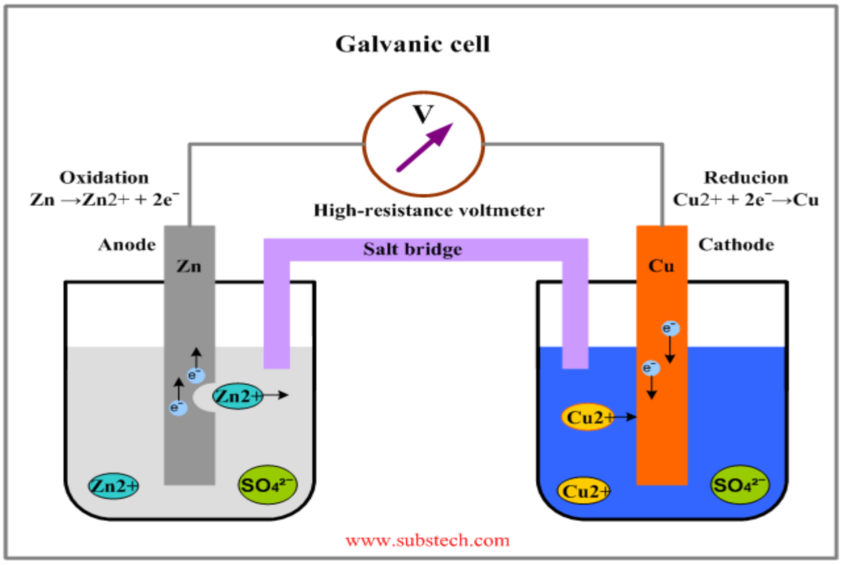
- The more negative the E° value, the further the equilibrium lies to the left - the more readily the element loses electrons and forms ions (metal atoms are soluble in own solution and oxidation occurs).
- The more positive (or less negative) the E° value, the further the equilibrium lies to the right - the less readily the element loses electrons and forms ions (metal ions are reduced in solution).

The main type of an electrochemical cell are:

- *A voltaic, or galvanic, cell* is an electrochemical cell in which a spontaneous reaction generates an electric current.
- *An electrolytic cell* is an electrochemical cell in which an electric current drives an otherwise non-spontaneous reaction.

2nd Experiment. Construction of Galvanic cell

Galvanic cells convert different forms of energy (chemical fuel, sunlight, mechanical pressure, etc.) into electrical energy and heat.



A galvanic Daniell cell consists of two half-cells. Each half-cell has: (1) an electrode, which in the figure are the plates of Zn (zinc) and Cu (copper); and (2) an electrolyte, which in the figure are aqueous solutions of $ZnSO_4$ and $CuSO_4$. The metal of a metallic electrode tends to go into solution, thereby releasing positively charged metal ions into the electrolyte, and retaining negatively charged electrons on the electrode. Thus each half-cell has its own half-reaction. For the Daniell cell, depicted in the figure, the Zn atoms have a greater tendency to go into solution than do the Cu atoms. More precisely, the electrons on the Zn electrode have a higher energy than the electrons on the Cu electrode. Because the electrons have negative charge, to give electrons on it a higher energy the Zn electrode must have a more negative electrical potential than the Cu electrode. However, in the absence of an external connection between the electrodes, no current can flow.

When the electrodes are connected externally (as in the figure, with wire and a light bulb), the electrons tend to flow from the more negative electrode (Zn) to the more positive electrode (Cu). Because the electrons have negative charge, this produces an electric current that is opposite the electron flow. At the same time, an equal ionic current flows through the electrolyte. For every two electrons that flow from the Zn electrode through the external connection to the Cu electrode, on the electrolyte side a Zn atom must go into solution as a Zn^{2+} ion, at the same time replacing the two electrons that have left the Zn electrode by the external connection.

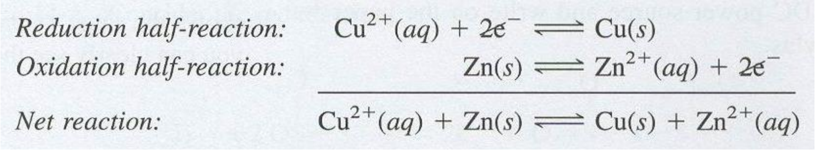
By definition, the anode is the electrode where oxidation (removal of electrons) takes place, so in this galvanic cell the Zn electrode is the anode. It oxidizes and gives up the electrons of the atoms dissolving in the electrolyte in form of positive ions:

Because the Cu has gained two electrons from the external connection, it must release two electrons at the electrolyte side, where a Cu^{2+} ion plates onto the Cu electrode. By definition, the cathode is the electrode where reduction (gain of electrons) takes place, so the Cu electrode is the cathode:

$$Cu^{2+} + 2\bar{e} \rightarrow Cu^{\circ}$$

$$E^{\circ}_{Cu^{2+}} = +0{,}34\ V$$

Thus the reaction that is going on is really:



Reduction half-reaction:	$Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$
Oxidation half-reaction:	$Zn(s) \rightleftharpoons Zn^{2+}(aq) + 2e^-$
Net reaction:	$Cu^{2+}(aq) + Zn(s) \rightleftharpoons Cu(s) + Zn^{2+}(aq)$

If the value of the electric current is negligible (high resistance voltmeter is used) the measured **potential difference between the electrodes is equal to the electromotive force (EMF)** of the galvanic cell. The measurement of electromotive force is usually made under standard conditions:

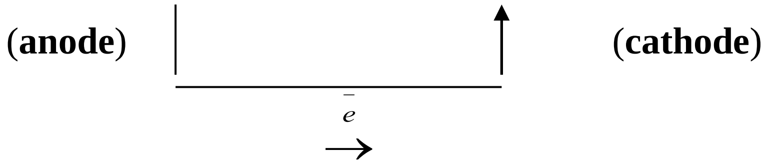
- Temperature: 77ºF (25ºC, 298K);
- Pressure of gases: 1 atm;
- Electrolyte solution activity (concentration): 1 M (1 mol/dm^3).
- Electromotive force measured under standard conditions has a notation E°.

EMF of a galvanic cell is a resulting sum of potential differences of the anode and cathode (at standard condition):

$$\Delta E = E_{cathode\ (+)} - E_{anode\ (-)} = E^{\circ}_{Cu} - E^{\circ}_{Zn} = 0{,}34 - (-0{,}76) = +1{,}10\ V$$

The galvanic cells, as the one shown in the figure, are conventionally described using the following notation:

$$(-)\ Zn / ZnSO_{4\ (aq)}\ 1M \,//\, CuSO_{4\ (aq)}\ 1M / Cu\ (+)$$

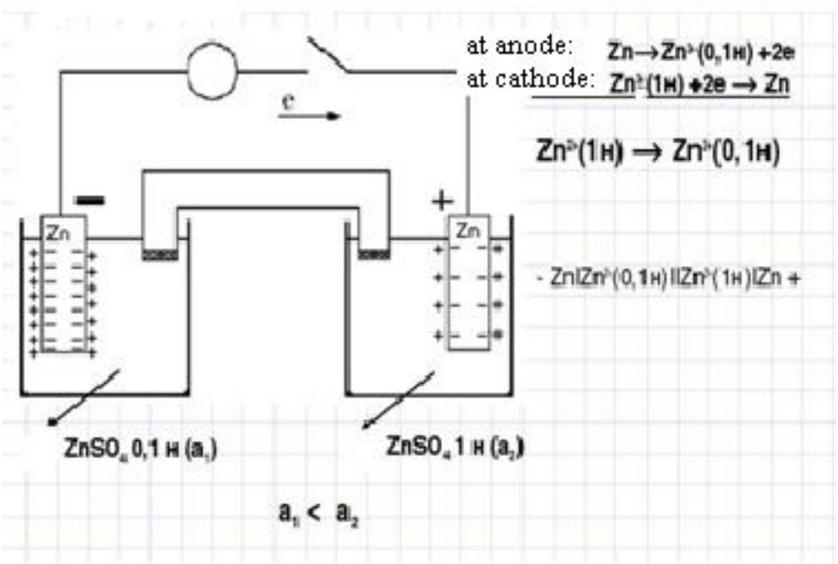


where: (s) denotes solid; (aq) means aqueous solution; the vertical bar, /, denotes a phase boundary; and the double vertical bar, //, denotes a liquid junction.

If the cell is operated under non-standard conditions, the potentials must be adapted using the Nernst equation.

3rd Experiment. Construction of concentration cell

Construction of concentration cell



If two plates of the same metal are dipped separately into two solutions of the same electrolyte and are connected with a salt bridge, the whole arrangement is found to act as a galvanic cell.

The cells in which EMF arises due to transfer of matter from one half cells to the other because of a difference in the concentration of species involved are known as ***concentration cell.***

- In these cells, electrodes are identical but these are immersed in solutions of the same electrolyte of different concentrations.
- The source of electrical energy in the cell is the tendency of the electrolyte to diffuse from a solution of higher concentration to that of lower concentration.
- With the expiry of time, the two concentrations tend to become equal. Thus, at the start the EMF of the cell is maximum and it gradually falls to zero. Such a cell is represented in the following manner: (C_1 is greater than C_2).
- **$- M / M^{n+} (C_2) // M^{n+} (C_1) / M +$**
- Example: $(Zn|Zn^{2+} (C_2))$/Anode || $(Zn^{2+} (C_1)|Zn)$/Cathode

$(Zn / Zn^{2+} (C_2))$/Anode // $(Zn^{2+} (C_1) /Zn)$ / Cathode	
Anode (-)	**Cathode (+)**
$Zn°_{(s)} - 2e \rightarrow Zn^{2+}(C_2)$	$Zn^{2+}(C_1) + 2e^- \rightarrow Zn°_{(s)}$

- The emf of the cell is given by the following expression at 25° C:

$$EMF = \Delta E = \frac{0,058}{n} \cdot \lg \frac{C_1}{C_2} \quad (C_{1\ cathode} > C_{2\ anode})$$

- The concentration cells are used to determine the solubility of sparingly soluble salts, valency of the cation of the electrolyte and transition point of the two allotropic forms of a metal used as electrodes, etc.
- For every concentration cell E° = 0.

EXERCISES:

1) What is the potential of a half cell consisting of aluminum electrode in 0,003N $AlCl_3$ solution 25°C (E°=1,66 V)

2) Determine the standard EMF of the cell: Zn/Zn^{2+} // Ni^{2+}/Ni

3) Calculate the EMF of the cell: Fe/$FeSO_4$ 0,2N // $CuCl_2$ 0,001M / Cu

4) **Question 1:** Which of the following equations represents calomel electrode?

a. $Hg_2Cl_2 + e^- \rightleftharpoons Hg + Cl^-$ b. $AgCl\ (s) + e^- \rightarrow Ag + Cl^-$

c. $Fe^{3+} + e^- \rightarrow Fe^{2+}$ d. $H_2 \rightleftharpoons 2H^+ + 2e-$

Question 2: Magnitude of electrode potential does not depend on.

a. Pressure b. Temperature c. Concentration d. Nature of electrode

Question 3: In which of the following elements is not a component of SHE?

a. Platinum b. Hydrogen c. Mercury d. Gold

Question 4: SHE can be used as..

a. only anode b. only cathode d. both anode and cathode c. neithet anode nor cathode

Question 5: E° value for Cu^{2+}/Cu volt is

a. 0.34 V b. 0.58 V c. 1.02 V d. 1.51 V

Question 6: Anode of Daniell cell is made up of..

a. Zn b. Cu c. Mg d. Al

Question 7: At anode of Denial cell

a. oxidation of zinc ions take place

b. oxidation of copper ions take place

c. oxidation of zinc take place

d. oxidation of copper take place

Question 8: In Denial cell direction of electron flow is

a. toward cathode b. toward anode c. toward zinc electrode

d. from copper to zinc electrode

Question 9: Daniell cell has the emf value.

a. 1.09 V b.0.19 V c.9.01 V d. 10.9 V

Question 10: Consider the reaction, $2Ag^+ + Cd \rightarrow 2Ag + Cd^{2+}$

The standard electrode potentials for $Ag^+ \rightarrow Ag$ and $Cd^{2+} \rightarrow Cd$ couples are 0.80 volt and -0.40 volt, respectively. What is the standard potential ΔE^o for this reaction?

a. 1.20 V b. 12.0 V c. 120 V d. 0.12 V

Question 11: What will be the emf of the following cell $Zn|Zn^{2+}(aq)(1.0M)||Cu^{2+}(aq)(1.0M)|Cu$

The standard electrode potentials are: $Cu^{2+} + 2e^- \rightarrow Cu(aq)$ $E^o = 0.350$ volt

$Zn^{2+} + 2e^- \rightarrow Zn(aq)$ $E^o = -0.763$ volt

a. 1.113 V b. 12.43 V c. 1.92 V d. 5.324 V

Question 12: What will be the electrode potential at a copper electrode dipped in a 0.1 M solution of copper sulphate at 25°C? The standard electrode potential of Cu^{2+}/Cu system is 0.34 volt at 298 K.

a. 10.113 V b. 0.31045 V c. 11.92 V d. 9.324 V

Question 14: For a concentration cell E° =

a. 0 V b. 1 V c. 2 V d. 3 V

Question 15: In concentration cells EMF arises due to transfer of:

a. electrons b. charge c. matter d. protons

Question 16: Which among the following is not a usage of standard hydrogen electrode?

a. Used for pH measurement of a given solution.
b. Used in acid-base titrations.
c. Used as a primary reference electrode.
d. None of these.

Question 17: In calomel electrode the bottom of tube contains:

a. Hg
b. KCl solution
c. HCl solution
d. Both b and c

Question 18: In high energy density batteries Lithium is used as an electrode because:

a) it is reactive.
b) it does not corrode easily.
c) it has high negative reduction potential.
d) it is the lightest metal.

Question 19: A standard hydrogen electrode has zero electrode potential because:

a) hydrogen is the lightest element.
b) hydrogen is easiest to oxidize.
c) this electrode potential is assumed to be zero.
d) hydrogen has only one electron.

Question 20: The constitute of a Danlell cell is:

a) Cu - Ag cell.
b) Zn – Ag cell
c) Zn – Cu cell
d) None of these.

Question 20: Which is a secondary cell?

a) Mercury cell.
b) Dry cell.
c) H_2 – O_2 cell.
d) Ni – Cd cell.

Question 21: If the electrode potentials of both the electrodes become equal in magnitude but opposite in sign then:

a) an electrochemical cell stops working.
b) an electrochemical cell starts working.
c) an electrochemical cell is reversed.
d) none of these.

Question 22: …………. cell directly converts the free energy of a chemical reaction into electricity.
a) Lead storage battery.
b) Fuel cell.
c) Leclanche cell
d) Concentration cell.

Question 23: In an electrochemical cell:
a) potential energy changes into electrical energy.
b) chemical energy changes into electrical energy.
c) kinetic energy decreases.
d) potential energy decreases.

LABORATORY WORK
ELECTROLYSIS

Objective: *to dentify the products of electrolysis. This experiment allows students to investigate the products formed by electrolysing various solutions.*

This experiment enables students to carry out the electrolysis of various solutions and to investigate the identity of the products formed at the electrodes. They should be able to link their practical experiences with theory and learn how to construct simple ionic equations.

Laboratory conditions: It should be carried out in a well-ventilated laboratory as significant amounts of toxic chlorine, bromine and iodine can be produced in some cases, as well as highly flammable hydrogen.

Reagents: Access to the following aqueous solutions (all approx. 0.5 M concentration) potassium iodide KI, sodium sulfate Na_2SO_4, tin chloride $SnCl_2$, copper sulfate $CuSO_4$, 1% starch solution

Materials:

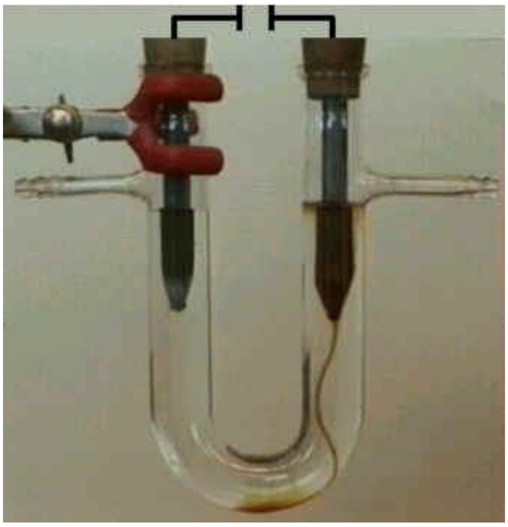
- Electrolysis apparatus (see diagram)
- Graphite electrodes (about 5 mm diameter),
- Large rubber bung to fit electrolysis cell, with holes to carry the graphite electrodes
- Small test-tubes (to fit over the electrodes),
- DC power supply (6 V)
- Small light bulb in holder (6 V, 5 W)
- Leads and crocodile clips
- Wooden spills
- Small pieces of emery paper

- Strips of Universal indicator paper
- Diposable plastic gloves
- Clamp and stand

Recommendations on the implementation of the laboratory experience and the use of electrolysis

***Notes to the student*:**

1. The electrolysis apparatus shown below can be purchased ready-made. Alternatively, it can be made from thick glass tubing of 8 - 10 cm diameter, professionally cut into lengths of about 12 cm (fig. 1). A suitable glass bottle, with a wide-necked top and its base cut off, could be used instead (as shown in the diagram). The graphite rods should be well sealed into the holes, 2 - 3 cm apart, of the rubber bung, otherwise the electrolyte may leak onto the external wiring, causing it to corrode. Once made, this apparatus should last for several sessions, but the graphite rods tend to erode away quite quickly, particularly if students use larger than recommended voltages. The rods do eventually become thin and snap fairly easily, but they are cheap enough to replace.



2. Once copper (II) sulfate has been electrolysed (preferably last), a deposit of copper will have formed on the cathode. This has to be removed before the cells can be used again. Immersing the plated part of the electrode in a small quantity of 50% concentrated nitric acid (CORROSIVE) in a small beaker can be used to do this. Gloves and eye protection should be worn and the cleaning done in a fume cupboard by a suitably qualified person.

3. Depending on the volume of the electrolysis apparatus, each group of students needs enough solution to cover the electrodes plus about 2 cm to enable the full test-tubes of liquid to be inverted over the electrodes.

a) Clamp the electrolysis cell and pour in enough of the first electrolyte so that the tops of the electrodes are covered with about 1–2 cm of liquid. Fill the two test-tubes with the same electrolyte. Wearing gloves, close the end of each test-tube in turn with a finger and invert it over an electrode, so that no air is allowed to enter (see diagram). During electrolysis it may be necessary to lift the test-tubes slightly to ensure that the electrodes are not completely enclosed, preventing the flow of current.

b) Connect the circuit, and mark the polarity of each electrode on the bung. The circuit should be checked *before* being switched on.

c) Switch on the circuit, then:

- observe whether or not the lamp lights up;

- look for the substances produced at each electrode – ie gaseous, solid or in solution;
- write down results *after each observation*, not when all the experiments are finished.

d) Only carry out the electrolysis for long enough to make the necessary observations. Prolonging the electrolyses unnecessarily causes toxic gases such as chlorine and bromine to be produced in unacceptably hazardous quantities.

e) After each electrolysis switch off the current and remove the test-tubes from the cell to test any gases present by lifting them slowly in turn to let any remaining solution drain out before closing the end with a finger. Carry out the tests on the gases as instructed.

f) (Optional) After removing the test-tubes from the cell, quickly pour the liquid down the sink with plenty of water. Wipe a piece of Universal indicator paper over each electrode and note any colour changes.

g) Wash the cell with plenty of water and dry the outside with a paper towel before fixing it back into position and re-connecting the power supply. It is important to connect the leads according to the polarities marked on the bung.

h) Repeat the experiment with each of the other four solutions, trying to keep to the order given in the table. Zinc chloride and copper nitrate should be the last electrolytes tested. This is because they deposit solids on the cathode. If zinc chloride is electrolysed first, the solid deposit on the cathode can be easily removed with a piece of emery paper or dipping the end of the electrode in some dilute hydrochloric acid in a beaker.

Notes to the teacher: The electrolysis of aqueous solutions, rather than molten salts, is easier and safer for students to do for themselves, Unfortunately the theory is more complicated, because the presence of water complicates what students may decide are the products formed at the electrodes.

Ensure that students do not attempt to smell directly any of the halogen fumes produced. It is important that you are aware of any students who are asthmatic or who might have an allergic reaction to these toxic gases. In this context do not allow the electrolyses of the halide solutions to proceed any longer than is absolutely necessary.

When testing for hydrogen or oxygen, the mouth of the test-tube can be closed with a gloved finger, and the test-tube transported to a central area, where a single naked flame has been set up, well away from the experiments. A supply of spills can also be kept in this area for the tests.

For the hydrogen test, students may well ask why there is little or no ‘pop’ or ‘squeak’. Explain that pure hydrogen – rather than a mixture of hydrogen and air – is being tested if the test-tube was full of gas before it was removed.

For the oxygen test, care should be taken that the dampness at the mouth of the test-tube does not extinguish the ‘glow’, causing the test to fail.

Once the electrolysis of zinc chloride or copper nitrate has been done, a deposit of metal will have formed on the cathode. This will have to be cleaned before the cell can be used again. These metal deposits can be removed using emery paper. Alternatively, small quantity of 50% concentrated nitric acid (CORROSIVE) in a small beaker can be used to remove the copper, providing gloves are worn and the operation is done in a fume cupboard by a suitably qualified person. Similarly dilute hydrochloric acid will remove the zinc.

4. ***Results and conclusions:***

Experimental results must be tabulated:

Lamp lights?		**KI**	**Na_2SO_4**	**$SnCl_2$**	**$CuSO_4$**
Observations	**anode (–)**				
	cathode (+)				
Test used for products	**anode (–)**				
	cathode (+)				
Identity of product formed	**anode (–)**				
	cathode (+)				

Write oxidation and reduction reactions occurring at the electrodes during electrolysis of these solutions by using ***rules of electrolysis*** in solution:

Anode (–): **the anion which is stronger reducing agent (low value of standard potential) is liberated first at the anode**	**Cathode (+):** **the ion which is stronger oxidizing agent (high value of standard potential) is discharged first at the cathode**
1. If anions of halogens (Cl^- , Br^- and I^-) are present in the solution, halide ion is oxidized at the anode and the halogen is produced: $\mathbf{\Gamma^{-n} - ne = \Gamma_2}$	1. If the cations of active metals are present in the solution (*they stay above to hydrogen in the electrochemical series*), at the cathode hydrogen ions are recovered: $\mathbf{2H^+ + 2e = H_2}$
2. If the solution contains no halide ions, such as NO_3^-, SO_4^{2-},CO_3^{2-}, hydroxide OH^- ion is oxidized at the anode and oxygen is produced: $\mathbf{OH^- - e = OH^\circ}$ $\mathbf{4OH^\circ = 2H_2O + O_2}$	2. If in solution present cations of passive metals (they stay after hydrogen in the electrochemical series), at the cathode these cations will be recovered: $\mathbf{Me^{+n} + ne = Me^\circ}$

For examples: a) electrolysis of silver nitrate solution

$$AgNO_3 \Rightarrow Ag^+ + NO_3^-$$
$$HOH \Rightarrow H^+ + OH^-$$

At Cathode (–):	**At Anode (+)**
Silver ions, Ag^+ ions are selectively discharged by receiving electrons to form Ag atom. This is because Ag^+ ions has a lower position and high E° than H^+ ions in the electrochemical series.	Hydroxide ions, OH are selectively discharged by donating electrons. This is because OH-ions has a lower position than SO_4^{2-} ions in the electrochemical series.
$Ag^{+1} + 1e = Ag°$	NO_3^-
H^+	**$OH^- - e = OH°$**
	$4OH° = 2H_2O + O_2$
• A silvery grey solid is deposited at cathode. • Silver, Ag metal is formed at the cathode.	• Colourless gas bubbles are released at anode. (The gas relights a glowing wooden splinter.) • Oxygen gas, O_2 and water, H_2O are produced at the anode.

$$4AgNO_3 + 2H_2O = 4Ag + O_2 + 4HNO_3$$

b) electrolysis of magnesium bromide solution:

$$MgBr_2 \Rightarrow Mg^{2+} + 2Br^-$$
$$HOH \Rightarrow H^+ + OH^-$$

Cathode (–):	**Anode (+)**
Hydrogen ions, H^+ ions are selectively discharged by receiving electrons to form H_2. This is because H^+ ions has a lower position and high E° than Mg^{2+} ions in the electrochemical series.	Bromide ions, Br- are selectively discharged by donating electrons. This is because in the first order halide ions are oxidized in solution
Mg^{2+}	**$2Br^- - 2e = Br_2$**
$2H^+ + 2e = H_2$	OH^-
• Colourless gas bubbles are released at cathode. • Hydrogen gas, H_2 is produced at the cathode.	• Brown vapor is released at anode. The gas turns the blue litmus paper to red. • Bromine gas, Br_2 is produced at the anode.

$$MgBr_2 + 2H_2O = H_2 + Br_2 + Mg(OH)_2$$

1st Experiment. ***The electrolysis of salt potassium iodide solution performed in a U-shaped test tube with universal indicator.***

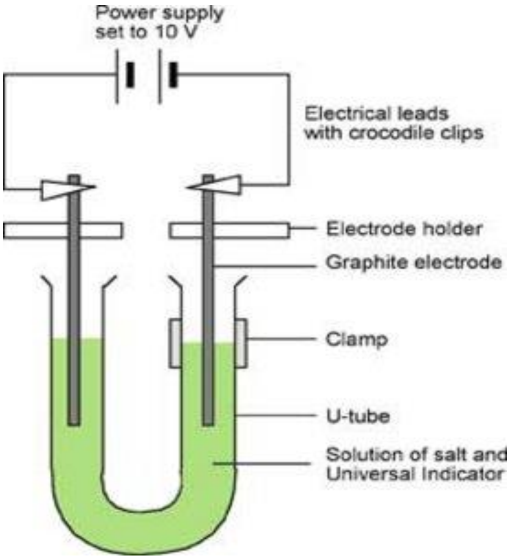
1) The products of the electrolysis of the salt solution are all more hazardous than the starting materials. Hydrogen is EXTREMELY FLAMMABLE, iodine is TOXIC and DANGEROUS FOR THE ENVIRONMENT, and sodium hydroxide is CORROSIVE. Ensure that the current is turned off a soon as a trace of chlorine is detected.

If the directions in the procedure notes are followed then very little iodine is produced. Sodium hydroxide is CORROSIVE. Ensure that students wear eye protection, especially when they are clearing up the experiment.

2) If distilled water is a problem, then tap water could be used. But it may affect the colours produced, especially in areas with hard water.

3) If electrode holders are not available, another suitable means of securing the electrodes could be used. Do not use bungs (заглушки, пробки) because the products are gases.

Procedure:



a) Put about 75 cm^3 distilled water into the beaker. Add about 2 heaped spatulas of potassium iodide.

b) Stir until the salt dissolves. Then add several drops of Universal Indicator solution. Stir to mix thoroughly. You need enough indicator to give the water a reasonable depth of green colour.

c) Pour coloured salt solution into the U-shaped test tube and clamp it as shown in the diagram.

d) Wash the carbon electrodes carefully in distilled water and then fix them so that there is about 3 cm of electrode in each side of the U-tube – see diagram. This is most easily done using electrode holders.

e) Attach leads and connect to a power pack set to 10 V (d.c.).

f) Turn on the power pack and observe closely what happens. A piece of white paper held behind the U-tube can help. Make sure the U-tube is kept very still during the experiment.

g) ***Turn off the power as soon as you notice any change*** at the positive electrode, or when you smell a ‘bleachy, swimming pool’ smell. This will probably take less than 5 minutes.

This experiment is an interesting introduction to the electrolysis of brine. It is probably best not used as the first electrolysis that students encounter. They would really struggle to explain for themselves what is going on. It could be followed by the electrolysis of salt solution in industry.

Students should be able to notice bubbles of gas at each electrode. At the positive electrode, the indicator turns red initially. This indicates the presence of iodine. At the negative electrode the indicator turns purple. The remainder of the solution stays green.

The product at the negative electrode is hydrogen. This can be difficult for students to understand.

Some of the water will ionise, that is, turn to hydrogen (H^+) and hydroxide (OH^-) ions.

When the KI is dissolved in water, the ions forming the ionic solid separate out. This means that there are actually 4 ions present in the solution: H^+, OH^-, K^+ and I^-.

The negative ions are attracted to the positive electrode. The iodide ions are discharged (giving iodine) in preference to the hydroxide ions. These are left behind in solution.

At the negative electrode, the hydrogen ions are discharged (producing hydrogen gas) in preference to the sodium ions. These are also left behind in solution. Thus, sodium hydroxide solution remains. This is the cause of the purple colour of the indicator at the negative electrode.

In time, the green colour of the indicator in the middle would change too, as the ions diffuse through the resulting solution.

Equations: $KI => K^+ + I^-$

$H_2O = > H^+ + OH^-$

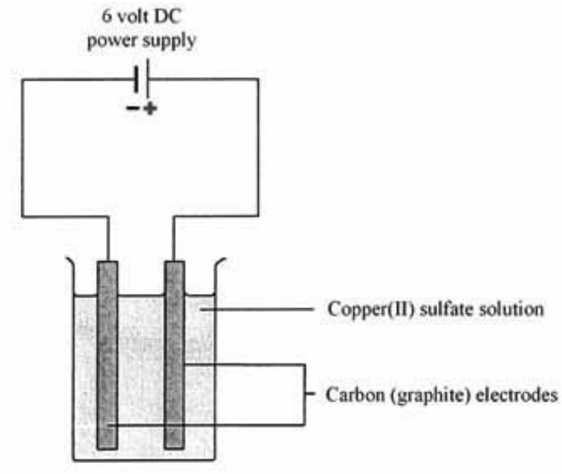
At anode (+): $2I^- \rightarrow I_2 + 2e^-$

At cathode (–): $2H^+ + 2e^- \rightarrow H_2$

In solution: $K^+ + OH^- = KOH$

Overall equation: $2KI + 2H_2O = H_2 + I_2 + 2KOH$

2nd Experiment. The electrolysis of copper sulfate solution



This experiment enables students to carry out the electrolysis of copper (II) sulfate solution and to link their findings with the industrial electrolytic refining of copper.

Copper (II) sulfate solution, $CuSO_4(aq)$ – see CLEAPSS Hazcadr and Cleapss Recipe Book. At the suggested concentrations, the copper (II) sulfate solution is LOW HAZARD If the concentrations are increased, the solutions must be labelled with the correct hazard warnings. Copper (II) sulphate solution is HARMFUL if concentration is equal to or greater than 1M.

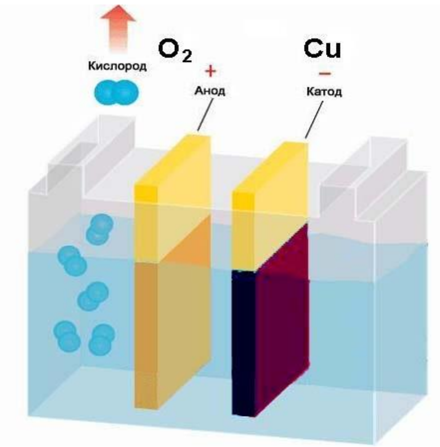
1) There are several ways of securing the graphite electrodes. Using a retort stand and clamp is probably the most convenient. They can also be fixed using Blutac on to a small strip of wood resting on the top of the beaker.

2) A bulb can be included in the circuit to indicate that there is a flow of current.

3) As an extension to the basic experiment, strips of copper can be used in place of the graphite rods.

Procedure:

A) Experiment with inert carbon electrodes.



Ask the students to set up the cell as shown. They should watch for any activity on each of the electrodes, and write down their observations.

The cathodes can be cleaned using emery paper.

Students should see a deposit of copper forming on the cathode. This will often be powdery and uneven. You should explain that, if the current used is much lower, then the solid coating is shiny, impermeable and very difficult to rub off; this process forms the basis of electroplating.

Bubbles of gas (oxygen) are formed at the anode.

Equation: $CuSO_4 => Cu^{2+} + SO_4^{2-}$

$H_2O => H^+ + OH^-$

At anode : $OH^- – e = OH°$

$4OH° \rightarrow 2H_2O + O_2(g)$

At cathode: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

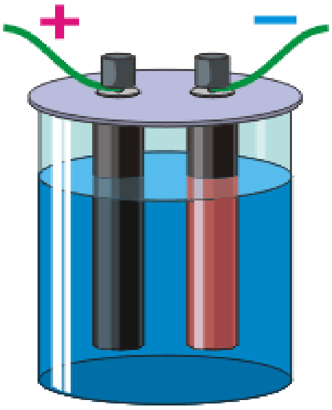
In solution: $2H^+ + SO_4^{2-} = H_2SO_4$

Overall equation: $2CuSO_4 + 2H_2O = 2Cu + O_2 + 2H_2SO_4$

Notes: With inert carbon (graphite) electrodes, the oxygen usually reacts with the anode to form CO_2. If copper is used for the electrodes, the copper anode dissolves. The reaction is the reverse of the cathode reaction.

The results of this experiment can lead to a discussion about electroplating and the electrolytic refining of copper.

It can be instructive to allow students to copper-plate metal objects supplied by the school and previously tested for their suitability. Personal items should not be used. In many cases, an alternative redox reaction often takes place before any current is actually passed. This happens for instance in items made of metals above copper in the reactivity series. It is wise not to complicate electrolytic deposition with chemical displacement - valued articles can be effectively ruined.



B) Experiment with active (copper) electrodes for copper refining.

1) After doing the electrolysis as described above, the electrodes can be interchanged. Students can then see the copper disappearing from the surface of the copper-coated anode:

$$Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$$

This leads to a discussion as to why, during electrolytic refining:

- the anode consists of an unrefined sample of the metal

- the cathode is made of pure copper or a support metal such as stainless steel.

2) The electrolysis can be done using two weighed copper strips. This is to confirm that the mass gained at the cathode is equal to the mass loss at the anode.

3) Quantitative electrolysis of aqueous copper (II) sulfate. This demonstration experiment shows how to find the value of the Faraday constant using electrolysis.

This demonstration is designed to find the value of the **Faraday constant** – the amount of electric charge carried by one mole of electrons - from the electrolysis of aqueous copper (II) sulfate solution, using weighed copper electrodes.

4) This is a relatively straightforward demonstration to set up and carry out, but there is not much to see while it is taking place. This provides an ideal opportunity to ask students to prepare a suitable results sheet and to explain how the masses of the electrodes, the current passed and the time elapsed can be used to calculate a value for the Faraday constant. The experiment will take around 40 minutes.

5) Students can be told that the calculations are relevant for industrial applications of electrolysis, for example in the extraction of elements such as sodium and chlorine, and in the electroplating industry.

Procedure:

These should be 2 - 3 cm wide and long enough to reach from the bottom of a 250 cm^3 beaker to the rim, where they can be folded over and gripped by the crocodile clips on the leads.

a) Clean the copper electrodes with emery paper, rinse under the tap and dry thoroughly using paper towels and a hot-air blower.

b) Mark the electrodes as **+** or **–** at one end, and weigh them separately. Record the masses.

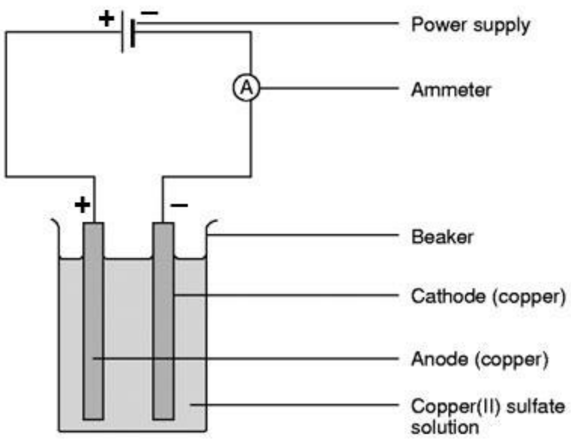
c) Set up the circuit as shown in the diagram, clamping the electrodes to the rim of the beaker using the crocodile clips. Make sure the electrodes do not touch.

d) Fill the beaker with copper (II) sulfate solution to just below the crocodile clips.

e) Start the stop clock and switch on the current, setting the rheostat so that a current of 0.50 A passes through the solution.

f) The current alters slightly throughout the electrolysis, so it is important to make continual adjustments to the rheostat to keep the current constant.

g) After about 30 min switch off the current and stop the clock.



h) Remove the electrodes from the electrolyte, wash them carefully under running water, then rinse them in a beaker of propanone in a fume cupboard. Finally dry them by allowing the propanone to evaporate in a well-ventilated laboratory, and away from any naked flames.

i) Re-weigh the dry electrodes.

The results should be entered in a table, for example:

Mass of cathode (-) before /g	
Mass of cathode (-) after /g	
Change in cathode's mass/g	
Mass of anode (+) before /g	
Mass of anode (+) after /g	
Change in anode's mass /g	
Current used /A	
Time for electrolysis /s	

Notes: It is difficult to avoid fluctuations in current throughout the electrolysis, and thus, obtain an accurate value to use in the calculations.

Be careful to avoid any movement of the electrodes during the electrolysis.

Thorough washing and drying of the electrodes is vital. A tiny amount of moisture left on the electrodes causes a considerable error in the mass calculations. The electrodes should NOT be heated in a flame to dry them as the copper will oxidise on the surface to form black copper (II) oxide. After initial drying by evaporation of the propanone, a hot-air blower could be used.

The student questions work through the process of calculating the Faraday constant. Students could also be asked to use the official value of 96500 Coulombs per mole to calculate the percentage error.

The calculated value for the Faraday constant should be correct to within about 10%. The major error comes from the value for the current, which is hard to keep constant to two significant figures.

The reaction at the cathode is: $Cu^{2+}(aq) + 2e^{-} \rightarrow Cu(s)$

The concentration of Cu^{2+} ions in solution remains constant. In a given time as many of these ions are produced at the anode as are being removed at the cathode.

The coating of solid copper on the cathode tends to be 'spongy' and some can be lost when electrode is washed or dried. Some may also have fallen off the cathode during electrolysis. Thus, the mass gain of the cathode is an unreliable measure of the quantity of copper actually deposited there.

3rd Experiment. The electrolysis of tin (II) chloride solution

Tin chloride offers a safer alternative to lead bromide for demonstrating the electrolysis of molten salts. Lead bromide decomposes to its elements just by heating without the need for electricity. The electrolysis of lead bromide must be carried out in a fume cupboard.

The electrolysis of tin chloride should be carried out in a fume cupboard. The chlorine produced at the positive electrode is toxic and dangerous for the environment.

Tin (II) chloride, $sncl_2(s)$ is corrosive, dangerous for the environment.

Chlorine, $Cl_2(g)$, is toxic, dangerous for the environment. Chlorine is a product of the electrolysis.

Procedure:

Set up apparatus as shown in diagram as in 1st experiment.

Bubbles of gas (oxygen) are formed at the anode, bubbles of gas (hydrogen) are formed at the cathode and precipitate of tin hydroxide is formed in solution.

Equation: $SnCl_2 \Rightarrow Sn^{2+} + 2Cl^{-}$

$H_2O = > H^{+} + OH^{-}$

At anode (+): $2Cl^{-} \rightarrow Cl_2 + 2e^{-}$

At cathode (–): $2H^{+} + 2e^{-} \rightarrow H_2$

In solution: $Sn^{2+} + 2OH^{-} = Sn(OH)_2$

Overall equation: $SnCl_2 + 2H_2O = H_2 + Cl_2 + Sn(OH)_2$

4th Experiment. The electrolysis of sodium sulfate solution

The electrolysis of an aqueous solution of sodium sulphate using inert electrodes produces hydrogen at the cathode and oxygen at the anode and a neutral solution of sodium sulphate remains unaltered by the electrolysis.

Equation: $Na_2SO_4 \Rightarrow 2Na^{+} + SO_4^{2-}$

$H_2O \Rightarrow H^{+} + OH^{-}$

At anode : $OH^{-} - e = OH^{\circ}$

$4OH^{\circ} \rightarrow 2H_2O + O_2(g)$

At cathode: $2H^{+} + 2e^{-} \rightarrow H_2$

In solution: $2Na^+ + SO_4^{2-} = Na_2SO_4$
The overall cell reaction is: $\cancel{Na_2SO_4} + 2H_2O = 2H_2 + O_2 \cancel{+ Na_2SO_4}$

If the reaction is carried out in a Hofmann Voltammeter, with some universal indicator in the solution, it will be noticed that around the cathode the solution becomes alkaline and around that anode the solution becomes acidic. This is explained as follows:

At the cathode: Hydrogen ions are being removed from solution, thereby leaving an excess of hydroxyl ions which makes the solution alkaline, and

At the anode: Hydroxyl ions are being removed, so leaving an excess of hydrogen ions which makes the solution acidic.

Procedure:

1) Prepare a very concentrated (it need not be saturated) solution of sodium sulfate in water. Add a generous portion of bromothymol blue solution to the aqueous sodium sulfate solution. If it is necessary add either some sulfuric acid solution or some sodium hydroxide solution in order to keep the electrolyte light green in color. Sodium sulfate solution should color the indicator green (pH = 7). The addition of acid or base may be necessary to counteract local water and/or air conditions.

2) Any DC power supply will work for this experiment.

3) The electrodes for this experiment must be platinum. Many of the commercially available platinum electrodes look black. The black coating is a layer of finely divided platinum (platinum black) deposited on the electrodes in order to increase the surface area of the platinum and thus increase the current that can flow through the system.

4) For this experiment the glassware must be very clean. I keep the Hoffman apparatus filled with distilled water until just before use. After the electrolysis has continued for sufficient time that measurable quantities of gas are produced, stop the reaction. Be careful when discussing the observations that you do not use terms that have not yet been defined in class.

5) Interpret the observations and on the blackboard write the balanced half-reactions for the processes that took place at each of the electrodes.

6) On the basis of the observations and the equations that have been written by the students, ask them to predict the color that will be produced when the cell is inverted into the beaker that was originally used for filling the cell. The students will look at the two equations, and because they have not taken account of the different volumes of hydrogen and oxygen that are produced, they will predict that the resultant solution will be acidic and the indicator will turn yellow. Based on what has been written on the board this is the only prediction that can be made.

7) Pour the contents of the Hoffman apparatus back into the beaker. You must be very careful to get every drop of electrolyte back into the beaker. The reaction is very sensitive and the loss of only a drop of the electrolyte may have a very negative effect on the reaction results. The students are amazed that the resultant solution is green. Use the results of this reaction to define terms such as anode, cathode, oxidation and reduction. The idea that the number of electrons lost at the cathode and gained at the anode are equal will "fall" out of the experiment.

8) Graphite (carbon) electrodes cannot be used in this experiment because the oxygen produced is adsorbed on the surface of the electrode (anode) and the net results will not show the quantitative 2:1 ratio of hydrogen to oxygen that is expected.

9) The choice of electrolytes and acid-base indicators is very limited. The pH of the solution must be 7.0, and the indicator must change through three colors, the middle one at pH 7.0.

5th Experiment. Preferential discharge of cations during electrolysis

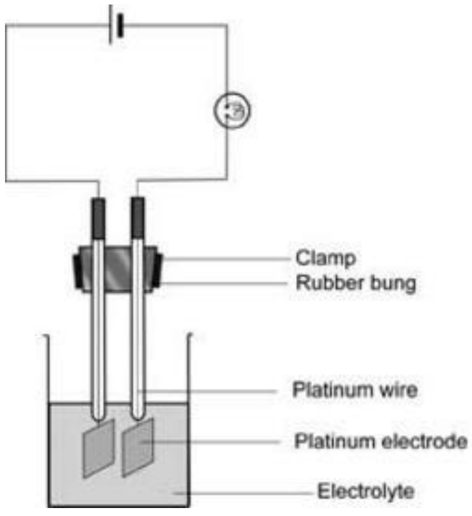
This experiment shows the order in which cations are discharged during electrolysis.

This experiment is designed to show that metal cations are preferentially discharged, in relation to the position of the metal in the reactivity series.

There is probably insufficient time for each group to do all six electrolyses. It is probably best to assign two metal ions and one mixture to each group, and pool the results from all the groups.

Reagents: Copper (II) sulfate, about 0.5 M, 200 cm^3; Iron (II) sulfate, about 0.5 M, 200 cm^3; Zinc sulfate, about 0.5 M, 200 cm^3 (IRRITANT, DANGEROUS FOR THE ENVIRONMENT); Nitric acid, about 4 M (CORROSIVE), 20 cm^3; Aqueous ammonia, about 4 M (IRRITANT), 10 cm^3; *Optional:* Solution of mercury (II) chloride (VERY TOXIC) mixed with ammonium thiocyanate (HARMFUL)

Materials: Beaker (250 cm^3); Beaker (100 cm^3), 2 Boiling tube, Test-tubes, 2 or 3, Teat pipette, Bunsen burner, Heat resistant mat, Tripod and gauze, Platinum electrodes, 1 cm square, with platinum leads sealed through glass tubes, both supported in a rubber bung or cork so that the electrodes are about 2 cm apart, DC power pack for supplying about 3–4 V; Light bulb (5 V) and holder; Several lengths of connecting wire, including two fitted with crocodile clips; Emery paper.



Procedure:

1) Platinum electrodes work best – alternatives include graphite.

2) If it is required, to test for zinc in the presence of copper (see teaching notes), the solution containing mercury (II) chloride and ammonium thiocyanate can be prepared by dissolving 2.7 g of mercury (II) chloride and 3 g of ammonium thiocyanate in 100

cm^3 of water. The mixed solution at this concentration should be labelled TOXIC.

a) Half-fill the 250 cm^3 beaker with water, and heat with a Bunsen burner until boiling. Stop heating and carefully transfer the beaker onto a heat resistant mat.

b) Half-fill the 100 cm^3 beaker with the copper (II) sulfate solution.

c) Clamp the electrodes so that they dip into the copper (II) sulfate solution and connect up the rest of the circuit as shown in the diagram, with the power supply switched off.

d) Switch on the current. The bulb should light up. Set the voltage to about 3–4 V, and electrolyse the solution for about 2 min.

e) In the meantime, use the teat pipette to transfer no more than 3–4 cm^3 of nitric acid into the boiling tube, and then place this in the beaker of hot water.

f) Switch off the current and remove the electrodes.

g) Place the electrodes under running water, ensuring that all parts that have been in contact with the solution have been washed thoroughly.

h) Dip the electrodes one at a time into the hot nitric acid, leaving them in the acid for about 10 seconds. If the electrodes do not reach into the acid or are difficult to remove from the bung, pour the hot acid into a small beaker first.

i) Test the solution of nitric acid for the presence of copper ions, using the test described at the end of the procedure.

j) Wash the electrodes, the contents of the boiling tube and the 100 cm^3 beakers using tap water.

k) Repeat steps **b–j** using the iron (II) sulfate solution, and then the zinc sulfate solution, using the tests for iron (II) ions and zinc ions in each case.

l) Finally use mixtures of two electrolytes and do the tests for the appropriate metal ions to find out which ion is discharged in preference to the other. Suitable combinations are:

- copper (II) sulfate and iron (II) sulfate
- iron (II) sulfate and zinc sulfate
- copper (II) sulfate and zinc sulfate

Tests for metal ions

Copper ions (Cu^{2+}): Using the teat pipette, transfer five drops of the solution into a test-tube. Wash out the teat pipette and then use it to add the aqueous ammonia drop by drop. Shake the test-tube thoroughly throughout the addition. Once all the nitric acid has been neutralised, a pale blue precipitate of copper (II) hydroxide forms initially, but with excess ammonia, this precipitate dissolves and is replaced by a dark blue solution:

$$CuSO_4 + 2NH_4OH = \underset{\text{blue precipitate}}{Cu(OH)_2\downarrow} + (NH_4)_2SO_4$$

$$Cu(OH)_2 + 4NH_4OH = \underset{\text{dark blue complex}}{[Cu(NH_3)_4](OH)_2} + 4H_2O$$

Iron (II) ions (Fe^{2+}): Using the teat pipette, transfer five drops of the solution into a test-tube. Wash out the teat pipette and then use it to add the aqueous ammonia drop by drop. Shake the test-tube thoroughly throughout the addition. Once all the nitric acid is neutralised, a dirty, dark green precipitate of iron (II) hydroxide eventually forms.

$$FeSO_4 + 2NH_4OH = \underset{\text{dark green precipitate}}{Fe(OH)_2\downarrow} + (NH_4)_2SO_4$$

Zinc ions (Zn^{2+}): Using the teat pipette, transfer five drops of the solution into a test-tube. Wash out the teat pipette and then use it to add the aqueous ammonia drop by drop. Shake the test-tube thoroughly throughout the addition. Once all the nitric acid is neutralised a white precipitate of zinc hydroxide forms, but this then disappears and is replaced by a colourless solution. If the ammonia is added too quickly the white precipitate will not be seen.

$$ZnSO_4 + 2NH_4OH = \underset{\text{white precipitate}}{Zn(OH)_2\downarrow} + (NH_4)_2SO_4$$

$$Zn(OH)_2 + 2NH_4OH = (NH_4)_2ZnO_2 + 4H_2O$$

Notes: However tempting might be, electrolysing solutions of other metal ions give poor results in practice. This is because the metals concerned tend not to adhere to the platinum cathode very effectively. However, copper, iron and zinc usually work very well.

It is essential that the washing process described in **g** is done thoroughly, otherwise the solution to be tested will contain the original ions, in addition to those derived from the electrolysis.

There are grounds for criticising these tests, as there is the possibility of two metals being deposited on the cathode. For example, during the electrolysis of a mixture of copper ions with zinc ions, if there is any zinc deposited on the cathode in addition to copper, the test described for copper will mask that for zinc. More complex test reactions for zinc ions do exist, producing definitive coloured products, but they would not be understood by the students.

In theory, and indeed in practice, no such mixture of ions should be present, because the experiment is designed to show preferential electrolytic discharge.

However, if you want to know how to test for zinc in the presence of copper, you can use the following technique. Transfer a few drops of the solution to a test-tube using a teat pipette. Add a few drops of the mercury(II) chloride-ammonium thiocyanate reagent. A yellow precipitate is formed by copper alone, but if zinc is also present the colour is darker.

Obviously you need to try this out beforehand so that the test can be described and demonstrated with a prior knowledge of the colour changes involved. During the class experiments it may be helpful to have both colours available on display to enable students to compare their results with yours.

Questions to the student:

1) For each of the individual solutions being electrolysed, on which electrode (positive or negative) is a metal deposited?

2) Write an equation for the electrical discharge of each metal ion: Cu^{2+}, Fe^{2+}, Zn^{2+}.

3) When mixtures of electrolytes are used, is a mixture of metals deposited at an electrode, or is it a single metal? If so, which?

4) What conclusion can you make about the priority of metal ions discharged during electrolysis?

EXERCISES:

1) Calculate the number of moles of hydrogen released when 5 amps of current passes for 3000 seconds through a solution of sulphuric acid.
2) Calculate the number of coulombs needed to deposit 6.35g of copper at the cathode in an electrolysis of $CuCl_2$ molten. Calculate the time that a current of 4 amps must pass to deposit this mass of copper.
3) A current of 0,5 ampere was passed through a solution of $AgNO_3$ for one hour. Calculate the masses of substances deposited on the electrodes.
4) When one Faraday or 96500 Coulombs of electricity is passed through silver nitrate solution, 108 gram of silver are deposited. Calculate electrochemical equivalent of silver.
5) How many grams of oxygen is liberated by the electrolysis of water after pass in of 0,0565 ampere for 185 sec.
6) Perform test:

№	Question	Variant
1	A device in which electric current is produced at the expense of spontaneous chemical reaction is:	a) voltaic cell b) voltameter c) galvanic cell d) both a and c
2	Which among the following is not true for electrochemical cell?	a) It consists of a battery. b) It consists of generally two electrolytes. c) Anode acquires negative charge. d) It needs a porous partition.
3	Among the following which is not a true statement for Faraday’s laws of electrolysis?	a) The weight of substance deposited is directly proportional to the quantity of electricity passed. b) The weight of substance deposited at the respective electrodes are directly proportional to their

		equivalent weights. c) Both a and b d) None of these
4	The quantity of electricity required to liberate or deposit 1 gram equivalent of a substance from its solution during electrolysis is known as:	a) Faraday b) Ampere c) Coulomb d) None of these
5	Salt bridge is an inverted U-glass tube containing saturated solution of:	a) KCl b) NH_4NO_3 c) K_2SO_4 d) Any one of these
6	In a redox reaction if the emf is positive then the reaction is:	a) Non-spontaneous b) Spontaneous c) Both a and b d) None of these.
7	For the following reduction reaction how many coulombs are required? 1 mol of Cu^{2+} to Cu	a) 193000 C b) 96500 C c) 482500 C d) 386000 C
8	Find the valency of the metal whose atomic weight is 96 and 0.3605g of a metal is deposited on the electrode by passing 1.2 ampere current for 15 minutes:	a) 1 b) 2 c) 3 d) None
9	For a given concentration of a solution with increase in temperature:	a) electrolytic conduction increases, metallic conduction decreases. b) electrolytic conduction decreases, metallic conduction increases. c) both electrolytic and metallic conduction increases. d) Both electrolytic and metallic conduction decreases.
10	When electrolysis of aqueous NaCl solution is carried out, the products are:	a) Na at cathode and Cl_2 at anode. b) H_2 at cathode and O_2 at anode. c) Na at cathode and O_2 at anode. d) H_2 at cathode and Cl_2 at

		anode.
11	…………….. is the amount of electricity required to deposit 1 mol of aluminium from a solution of $AlCl_3$.	a) Faraday b) 1 ampere c) 3 Faradays d) 3 Ampere
12	An aqueous solution of Na_2SO_4 is electrolysed using platinum electrodes. The products at the anode and the cathode are:	a) O_2, Na b) SO_{4-}, H_2 c) SO_{4-}, Na d) O_2, H_2
13	Due to the presence of …………… NaCl conducts electricity.	a) free ions b) free atoms c) free molecules d) free electrons
14	Which one of the following is a false statement?	a) Salt bridge maintains electrical neutrality b) When salt bridge is removed then the potential of the cell drops to zero. c) Salt bridge increases the emf of the cell. d) Salt bridge connects two half cells.
15	A current of x A flowing for 10 min deposits 3g of the metal which is monovalent. The atomic mass of the metal is 50. Find the value of x?	a) 10.5 b) 9.65 c) 8.75 d) 6.72
16	The chemical equivalent of the metal is 19.3. When 3A of electric current is passed for 50 min. what will be the amount of metal deposited from an electrolyte?	a) 1.8 g b) 0.9 g c) 3.6 g d) 4.2 g
17	NaCl in solid form does not conduct electricity because:	a) It does not ionize b) It ionizes c) It is present in fused form d) None of these.
18	When 10A current is passed for 96.5 sec for electrolysis of Al $(NO_3)_3$. Find the number of aluminium atoms deposited at cathode?	a) 2×10^{23} b) 2×10^{24} c) 2×10^{22} d) 2×10^{21}
19	X ampere current is passed for y seconds through copper and silver voltameters. The metal that is deposited more is:	a) Ag b) Na c) Cu d) Al
20	A current of 2 ampere was passed	a) 3.16g

	through solutions of $CuSO_4$ and $AgNO_3$. The weight of copper deposited is 0.636 g. What is the weight of silver deposited?	b) 4.32g c) 2.61g d) 2.16g
21	A certain quantity of electricity was passed through $CuSO_4$ solution for 6 minutes 20 sec. The amount of Cu deposited is 0.317g. (At.wt.of Cuis 63.5 and Faraday = 96500 coulombs). Find the amount of electricity consumed in ampere?	a) 25.0 b) 27.3 c) 2.54 d) 2.73
22	1 C contains ……………… number of electrons.	a) 6.24×10^{18} b) 6.03×10^{16} c) 6.05×10^{21} d) 6.24×10^{17}
23	What is the correct efficiency, if an electricity of 10800C decomposes 14.25 gm of $AgNO_3$?	a) 100% b) 50% c) 25% d) 75%
24	Calculate the time required for a current of 2.5 A to obtain 1% coating of Ag on a tea-pot weighing 0.27 Kg.	a) 965 sec b) 40 sec c) 1930 sec d) 486 sec.

LABORATORY WORK
METALS

Objectives: 1) Students understand the terms like metals, metal displacement reaction, reactivity series, etc.

2) Students acquire skills to perform & visualize the reactions of Al, Zn, Fe and Cu with the following salt solutions:

- Aluminium sulphate
- Zinc sulphate
- Ferrous sulphate
- Copper sulphate

3) Students can analyze the meaning of reactivity series of metals based on the inferences from the experiment.

4) Students acquire skills to perform an experiment to determine the reactivity of metals in salt solutions.

Materials and reagents: strip of metals – Zn, Cu, Al, Fe; solutions – $ZnSO_4$, $CuSO_4$, $Al_2(SO_4)_3$, $FeSO_4$, four 100 mL beakers, forceps

Precautions:

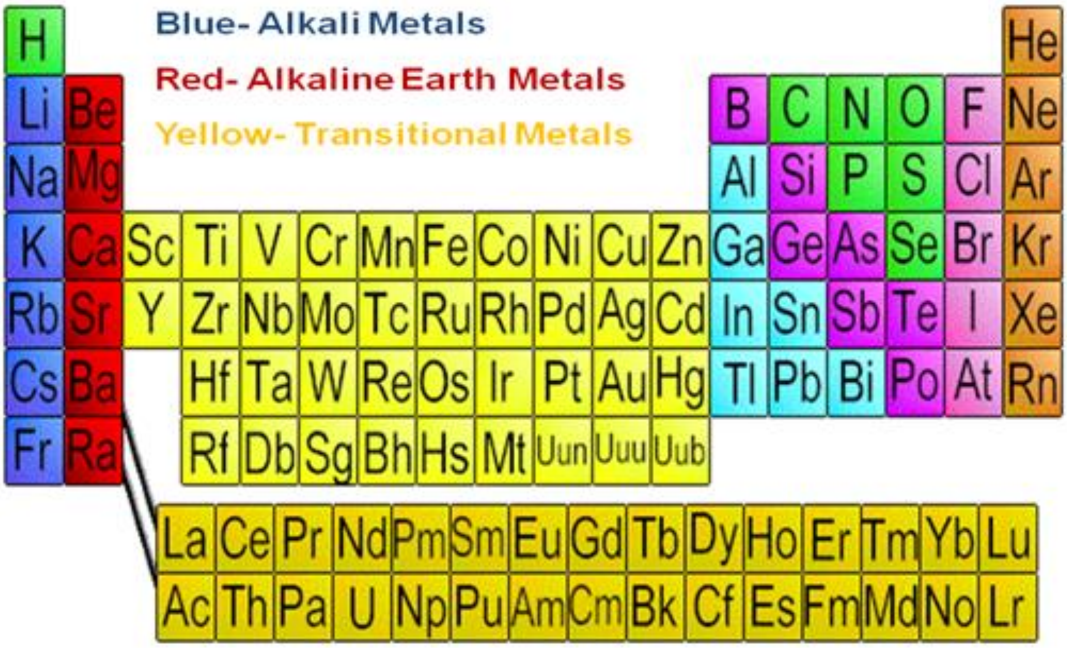
- Do not touch any chemical.

- Label the beakers properly and put them in sequences. It is better to label the beakers both with the chemical names and also by chemical symbols.
- Till the reaction goes on, keep the beakers at a safe place. Cover them so that insects or dust may not pollute the solutions.
- It is very important that the apparatus you use must be very clean and dry.
- Always clean the metallic strips before use with a sand paper. Use the same strips of metals of same size, weight and number.

Metals are elements and are good conductors of heat and electricity. Most metals are electropositive in nature and the metal atoms lose electrons in chemical reactions to form cations. The more reactive a metal, the greater tendency it has to form a positive ion in a chemical reaction. For example:

$$Na \longrightarrow Na^{+} + e^{-}$$

How are metals arranged in a periodic table? Metals occupy the bulk of the periodic table. The alkali metals are a series of chemical elements in the periodic table.



- Alkali metals comprise group 1 in the periodic table along with hydrogen. Alkali metals are lithium (Li), sodium (Na), potassium (K), rubidium (Rb), cesium (Cs) and francium (Fr).
- The group 2 elements in the periodic table are called alkaline earth metals. Alkaline earth metals are beryllium (Be), magnesium (Mg) calcium (Ca), strontium (Sr), barium (Ba) and radium (Ra). Group1 and group 2 elements are together called s-block elements.
- Elements of group 3 to 12 are transition elements. These are also metallic in nature and are called transition metals. They are also called d-block elements.

- Non-metallic elements occupy the right side of the periodic table. A diagonal line from boron to polonium separates metals from non-metals.

Physical properties of metals:

- Good conductors of heat and electricity: In metals, the positive ions are surrounded by a sea of electrons and these are responsible for their conductivity.
- High melting point and boiling point.
- They are malleable and ductile, so they can be bent and stretched without breaking.
- In most metals, the atoms are highly close packed, so they have high density.
- They have a shiny appearance.
- Metals exist in solid state at room temperature except for mercury, which is in a liquid state at room temperature.

Chemical properties of metals:

A more reactive metal readily reacts with other elements.

- The most reactive metals will react with even water, while the least reactive metals will not react even with acid. Most metals, on reacting with water produce hydroxide.

Example: If we put a small piece of sodium metal in water, sodium reacts exothermically with water producing hydrogen and metal hydroxide.

$$2Na(s) + 2H_2O(l) \longrightarrow 2\,NaOH(aq) + H_2(g)$$

Magnesium reacts mildly with water but vigorously with steam. Zinc and iron react mildly with steam. Copper, gold and silver do not react with water at all.

- Metals react with oxygen in the air to form oxides.

Examples:

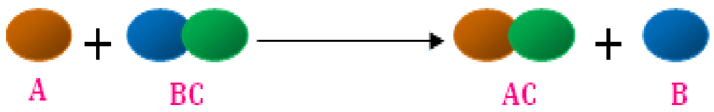
$$4Na + O_2 \longrightarrow 2Na_2O$$

$$4Al + 3O_2 \longrightarrow 2Al_2O_3$$

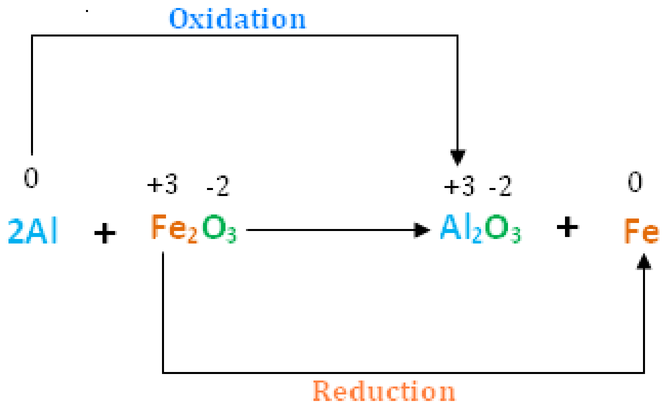
- Some metals react with acid and replace hydrogen from the acid.

$$Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2$$

- Most metals corrode when they are exposed to atmosphere. For example, the iron gets rusty after sometime if it is not painted. Titanium is highly resistant to corrosion.
- *Single displacement reactions*: A single displacement reaction is an important type of chemical reaction. It is also called substitution reaction. In these reactions, a free element displaces another element from its compound, producing a new compound. The reaction is usually written as:



Single displacement reactions are all oxidation-reduction reactions. For example,



Definition of Displacement reaction: *The chemical reaction by which one element takes the position or place of another element in a compound.*

THE ACTIVITY SERIES OF METALS

Element	Symbol	Group No.
Potassium	K	1
Sodium	Na	1
Lithium	Li	1
Calcium	Ca	2
Magnesium	Mg	2
Aluminium	Al	3
Zinc	Zn	Transition metal
Iron	Fe	Transition metal
Tin	Sn	4
Lead	Pb	4
[Hydrogen]	H	Non-metal
Copper	Cu	Transition metal
Silver	Ag	Transition metal
Gold	Au	Transition metal
Platinum	Pt	Transition metal

Most reactive

Decreasing chemical reactivity

Least reactive

Displacement reactions are very common in metals. They can be used to find out the relative reactivities of metals. In a displacement reaction, a more reactive metal can displace a less reactive metal from its salt solution. The reaction is often known as metal displacement reaction.

Some of the commonly used metals have been arranged in the decreasing order of reactivity. This is known as the **reactivity series or activity series**. The activity series of metals is an important concept in chemistry. The activity series of metals is an important tool for predicting the products of displacement reactions and the reactivity of metals in other reactions. Potassium is the most reactive metal, while platinum is the least reactive.

The higher of metal in the series, the more reactive it is and the more vigorously it reacts with water, oxygen and acid. A metal in the activity series can displace any metal below it in the series from its compound. The elements potassium, sodium, lithium and calcium are very reactive and they react with cold water to produce hydroxides and hydrogen gas. The elements magnesium, aluminium and iron are also considered as active metals and they react with steam to produce oxides and hydrogen gas. The metals above hydrogen are more reactive than hydrogen. These metals can displace hydrogen from acids or water and liberate hydrogen gas. The metals like copper, silver, gold and platinum are less reactive than hydrogen and they do not replace hydrogen from water or acid.

Examples for metal-displacement reactions:

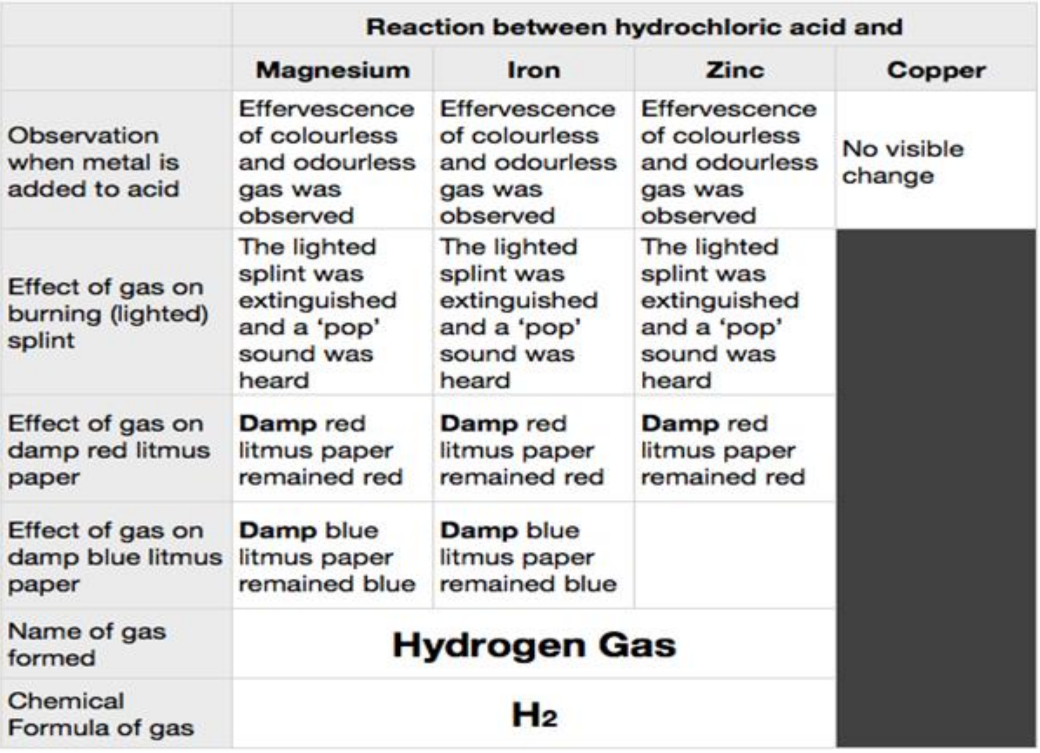
- Reaction with acids:

a) Potassium, sodium, lithium and calcium react violently with dilute H_2SO_4 and dilute HCl, forming the metal salt (either sulphate or chloride) and hydrogen gas:

$$2Na + 2HCl = 2NaCl + H_2 \uparrow$$

$$Mg + H_2SO_4 = MgSO_4 + H_2 \uparrow$$

$$Zn + H_2SO_4 = ZnSO_4 + H_2 \uparrow$$



	Reaction between hydrochloric acid and			
	Magnesium	**Iron**	**Zinc**	**Copper**
Observation when metal is added to acid	Effervescence of colourless and odourless gas was observed	Effervescence of colourless and odourless gas was observed	Effervescence of colourless and odourless gas was observed	No visible change
Effect of gas on burning (lighted) splint	The lighted splint was extinguished and a 'pop' sound was heard	The lighted splint was extinguished and a 'pop' sound was heard	The lighted splint was extinguished and a 'pop' sound was heard	
Effect of gas on damp red litmus paper	**Damp** red litmus paper remained red	**Damp** red litmus paper remained red	**Damp** red litmus paper remained red	
Effect of gas on damp blue litmus paper	**Damp** blue litmus paper remained blue	**Damp** blue litmus paper remained blue		
Name of gas formed	**Hydrogen Gas**			
Chemical Formula of gas	**H_2**			

b) Zinc with dilute sulphuric acid is often used for the laboratory preparation of hydrogen. The reaction is slow at room temperature, but its rate can be increased by the addition of a little copper (II) sulphate. Zinc displaces copper metal, which acts as a catalyst.

c) Metals below hydrogen (copper, silver, gold and platinum), will not react with dilute acids. They cannot displace hydrogen from the non-metal anion.

- **Reaction with concentrated acids as HNO_3 and H_2SO_4:**

Hydrogen gas is not evolved when metals react with nitric acid (HNO_3) because it is a strong oxidising agent and it oxidises the H_2 produced to water and is itself reduced to nitrogen dioxide.

1) With active metals:

$$Mg + HNO_{3(dilut)} = Mg(NO_3)_2 + H_2O + NH_3\,(NH_4NO_3)$$

$$Mg + 6HNO_{3(conc)} = Mg(NO_3)_2 + 4H_2O + 2N_2O$$

2) With passive metals:

$$Cu + HNO_{3(dilut)} = Cu(NO_3)_2 + H_2O + NO$$

$$3Cu + 8HNO_{3(conc)} = 3Cu(NO_3)_2 + 4H_2O + 2NO_2$$

3) Reaction with concentrated sulfuric acid:

$$Me + H_2SO_4\,(conc) = MeSO_4 + H_2O + (H_2S, S, SO_2)$$

4) Fe and Al will not react with conc H_2SO_4 acid, they are passivated.

5) One part of HNO_3 and three parts of HCl are mixed together, the mixture wich obtained is called *Aqua Regia*. Noble metals such as gold, platinum are not dissolved either of the acid alone but in Aqua Regia they easily dissolve due to liberation of atomic Cl which is highly reactive:

$$HNO_3 + 3HCl \longrightarrow NOCl + 2H_2O + 2Cl^-$$

- Zinc can displace copper from copper sulphate solution and iron from ferrous sulphate solution. So zinc is more reactive than iron and copper.

$$Zn(s) + CuSO_4(aq) \longrightarrow ZnSO_4(aq) + Cu(s)$$

$$Zn(s) + FeSO_4(aq) \longrightarrow ZnSO_4(aq) + Fe(s)$$

- If a piece of zinc metal is added to aluminium sulphate solution, no reaction takes place. So aluminium is more reactive than zinc.
- If small pieces of copper are added into solutions of zinc sulphate, aluminium sulphate and ferrous sulphate, no reactions take place. So, copper is less reactive than zinc, aluminium and iron.
- Aluminium can displace Zn from zinc sulphate solution, Cu from copper sulphate solution and Fe from ferrous sulphate solution.

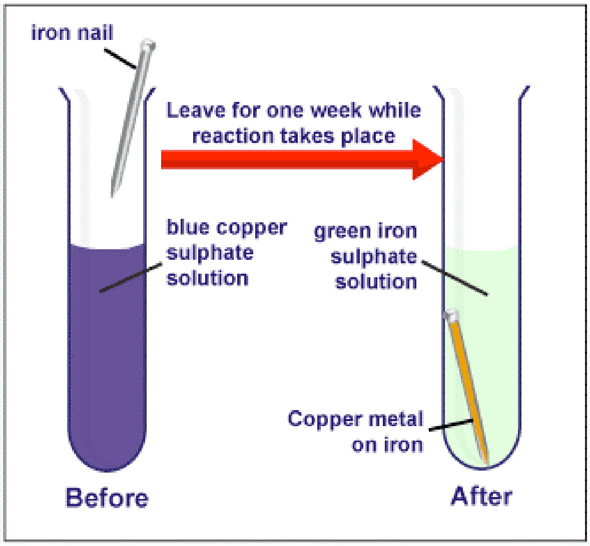
$$2Al + 3ZnSO_4 \longrightarrow 3Zn + Al_2(SO_4)_3$$

$$2Al + 3CuSO_4 \longrightarrow 3Cu + Al_2(SO_4)_3$$

$$2Al + 3FeSO_4 \longrightarrow 3Fe + Al_2(SO_4)_3$$

- Iron displaces Cu from copper sulphate solution, but it does not displace Al and Zn from their salt solutions.

$$Fe(s) + CuSO_4(aq) \longrightarrow Cu(s) + FeSO_4(aq)$$



- Not only metals but also non-metals can take part in displacement reactions.

For example, chlorine gas displaces bromine from potassium bromide solution. The solution of potassium bromide acquires a yellowish and orange color of the liberated bromine gas.

$$Cl_2(g) + 2KBr(aq) \longrightarrow 2KCl(aq) + Br_2(g)$$

Occurrence of Metals. The earth's crust is the major source of metals. Seawater also contains some soluble salts such as sodium chloride, magnesium chloride, etc. The elements or compounds, which occur naturally in the earth's crust, are known as minerals. At some places, minerals contain a very high percentage of a particular metal and the metal can be profitably extracted from it. These minerals are called ores.

Extraction of Metals: Some metals are found in the earth's crust in the free state. Some are found in the form of their compounds.

The metals at the bottom of the activity series are the least reactive. They are often found in a free state. For example, gold, silver, platinum and copper are found in the free state. Copper and silver are also found in the combined state as their sulphide or oxide ores.

The metals at the top of the activity series (K, Na, Ca, Mg and Al) are so reactive that they are never found in nature as free elements.

The metals in the middle of the activity series (Zn, Fe, Pb, etc.) are moderately reactive. They are found in the earth's crust mainly as oxides, sulphides or carbonates. The ores of many metals are oxides. This is because oxygen is a very reactive element and is very abundant on the earth.

Thus on the basis of reactivity, we can group the metals into the following three categories:

1 Metals of low reactivity (after H atom);
2 Metals of medium reactivity (between Al and Pb);
3 Metals of high reactivity (above Mg metal atom).

Different techniques are to be used for obtaining the metals falling in each category. Metallurgy as Per Reactivity:

- Electrolysis of molten is used for highly active metals such as K, Na, Ca, Mg and Al.
- Reduction with carbon is used for medium active metals such as Zn, Fe, Pb and Cu.
- Least reactive metals such as Ag and Au are found in free state in nature.

Refining of Metals: The metals produced by various reduction processes described above are not very pure. They contain impurities, which must be removed to obtain pure metals. The most widely used method for refining impure metals is electrolytic refining.

Electrolytic Refining: Many metals, such as copper, zinc, tin, nickel, silver, gold, etc., are refined electrolytically. In this process, the impure metal is made the anode and a thin strip of pure metal is made the cathode. A solution of the metal salt is used as an electrolyte. On passing the current through the electrolyte, the pure metal from the anode dissolves into the electrolyte. An equivalent amount of pure metal from the electrolyte is deposited on the cathode. The soluble impurities go into the solution, whereas, the insoluble impurities settle down at the bottom of the anode and are known as anode mud.

1st Experiment. Reaction of metals with water

We have the following metals powder in laboratory: Cu, Fe, Zn, Mg, Na, Al, K, Ca, Pb. Please, make up a series of activity of these metals in accordance with their standard potentials from active to passive:

Me									
E°, V									

PROCEDURE:

1) Primarily we carry out experiments with alkali metals: for the alkali metals prepare porcelain cups with water and metal tweezers. Take a small piece of each metal by tweezers and gently lower it into the water in cup. Be careful, the reaction proceeds vigorously. After the metal is completely reacted with the water, drip 1-2 drops phenolphthalein indicator in cup and note the color of the solution. What products are formed? What metal is the most active? Write the equation redox reactions of alkali metals with water:

$K + H_2O =>$
$Na + H_2O =>$
$Ca + H_2O = >$

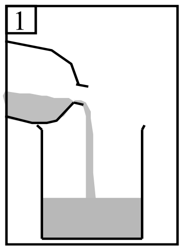
2) For other metals, prepare the six test tubes and label them. Put a few of the metal powder in each tube and add a small amount of water and heat the tube until the reaction starts. Then, add 1 drop of phenolphthalein indicator in each tube and note the change in colors of solutions in test tubes. What products are formed? What is the metal most active? What metals do not react with water? Write the equation redox reactions of metals with water.

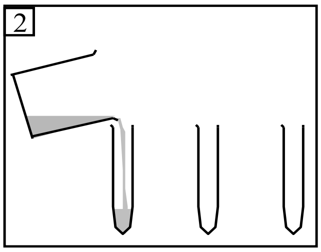
2nd Experiment. Reaction of metals with diluted acids as HCl and H_2SO_4.

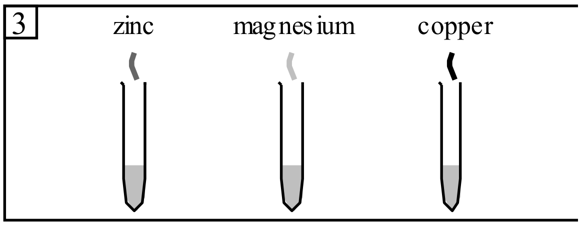
Hazards: *Dilute hydrochloric acid irritates the eyes and magnesium ribbon is highly flammable. When a metal reacts with acid, an acid mist is formed which irritates the eyes and throat. Hydrogen gas is produced in the reaction and it is highly flammable.*

Procedure:

1. Add dilute hydrochloric acid to the beaker until it is half full.
2. Put five test tubes in the test tube rack.
   Pour some of the hydrochloric acid into the first test tube to a depth of about 4 cm.
   Pour the same volume of acid into the other four test tubes.







3. Add a piece of zinc to the first test tube.
   Add a piece of magnesium to the second test tube.
   Add a piece of copper to the third test tube.
   Add a piece of iron to the fourth test tube
   Add a piece of aluminium to the fifth test tube
4. Watch carefully what happens in each test tube.
   - the name of each metal
   - whether bubbles of gas were given off or not
   - the speed at which the bubbles were given off
   - write the equation redox reactions of metals with water.

5 Repeat the experience with diluted 0.01M sulfuric acid H_2SO_4. What products are formed? What metal is the most active? What metals do not react with acid? Write the equation redox reactions of metals with water.

3rd Experiment. Reaction of metals with nitric acid HNO_3.

Nitric acid is a strong oxidizing agent as shown by its large positive E° values:

$$NO_3^-{}_{(aq)} + 2H^+{}_{(aq)} + 1e^- \rightarrow NO_{2\,(g)} + H_2O \qquad E^\circ = 0.79\ V$$

$$NO_3^-{}_{(aq)} + 4H^+{}_{(aq)} + 3e^- \rightarrow NO_{(g)} + 2H_2 \qquad E^\circ = 0.96\ V$$

Being a powerful oxidizing agent, nitric acid reacts violently with many non-metallic compounds and the reactions may be explosive. Depending on the acid concentration, temperature and the reducing agent involved, the end products can be variable. Reaction takes place with all metals except the precious metal series and certain alloys. As a general rule, oxidizing reactions occur primarily with the concentrated acid, favoring the formation of nitrogen dioxide (NO_2).

Nitric acid dissolves most metals including iron, copper, and silver, with generally the liberation of lower oxides of nitrogen rather than hydrogen:

$$Cu + 4HNO_{3\,(conc)} \rightarrow Cu(NO_3)_2 + 2NO_2 + 2H_2O$$

The acidic properties tend to dominate with dilute acid, coupled with the preferential formation of nitrogen oxide (NO):

$$3Cu + 8HNO_{3\,(dil)} \rightarrow 3Cu(NO_3)_2 + 2NO + 4H_2O$$

Since nitric acid is an oxidizing agent, hydrogen (H) is rarely formed. Only magnesium (Mg) and calcium (Ca) react with *cold, dilute* nitric acid to give hydrogen:

$$Mg_{(s)} + 2HNO_{3\,(aq)} \rightarrow Mg(NO_3)_{2\,(aq)} + H_{2\,(g)}$$

$$4Mg + 10HNO_{3\,(conc)} \rightarrow Mg(NO_3)_2 + 3H_2O + NH_4NO_3$$

Procedure:

1) ***Safety:*** *This acid is dangerous because it is highly corrosive and has oxidizing properties and as such it must be handled with great care. When handling nitric acid, one must wear protective gears especially to protect the eyes. The skin must also be protected because when nitric acid comes in contact with the skin, it results in a yellow discolouration. Perform this reaction in the fume hood.*

2) Prepare the seven test tubes and label them.

3) Place 20-25 drops of 16 M nitric acid (concentrated) in all test tubes and support the test tubes in a beakers in the hood.

4) Add a small amount of each metals powder to the test tubes. Allow the metals approximately five minutes to dissolve. The brown gas produced is toxic nitrogen dioxide (NO_2). When all of the metal has dissolved, fill the test tube approximately ¼ full with distilled water.

5) What metal is the most active? Write the equation redox reactions of metals with nitric acid.

6) Repeat this experiment with dilute nitric acid. What products are formed? What metal is the most active? Write the equation redox reactions of metals with acid.

4nd Experiment. Single displacement reaction of metals with salt.

In this experiment you will be ranking some of the metals according to their activities, from most to least active. One way to do this is to observe the relative vigor of their reactions (tendency to react) with nonmetals such as oxygen or chlorine. Another method, which differentiates more clearly, is to note whether one metal can replace another in a chemical compound. The general rule is that the more active metal replaces the less active one. This is easily observed in reactions between metals and metal ions (cations) in solution.

The general reaction for the replacement of a metal ion by another metal may be written as follows (We are arbitrarily showing only one electron per atom being transferred):

$$\underset{\text{Free metal A}}{\mathbf{M_A}} + \underset{\text{Salt of metal B}}{\mathbf{M_B^+X^-}} \rightarrow \underset{\text{Salt of metal A}}{\mathbf{M_A + X^-}} + \underset{\text{Free metal B}}{\mathbf{M_B}}$$

Or, in net ionic form, $M_A + M_B^+ \rightarrow M_A^+ + M_B$

PROCEDURE:

1. Select four 100 ml beakers and label them as zinc sulphate, copper sulphate, aluminium sulphate and ferrous sulphate.
2. Add 20ml of 1M solution of $ZnSO_4$, $CuSO_4$, $Al_2(SO_4)_3$ and $FeSO_4$ in the labelled beaker respectively.
3. Take a strip of cleaned zinc metal and cut it into small pieces of suitable size.
4. Add two pieces into each beaker containing $CuSO_4$, $Al_2(SO_4)_3$ and $FeSO_4$.

5 Keep the beakers undisturbed for about 1 hour.

6 After about an hour, note the change in colour of solutions, appearance of metal surface or any other changes.

7 Write an equation for the reaction between metal and salt.

8 Repeat the similar procedure by adding –

- Copper strips to $ZnSO_4$, $Al_2(SO_4)_3$ and $FeSO_4$ solutions.
- Aluminium strips to $ZnSO_4$, $CuSO_4$, and $FeSO_4$ solutions.
- Iron strips to $ZnSO_4$, $CuSO_4$ and $Al_2(SO_4)_3$ solutions.

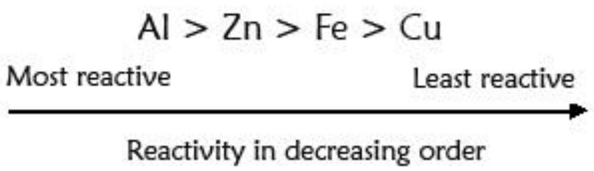
Observations:

Metal	*Salt solution in which added*	*Colour change of solution*	*Appearance of metal surface*	*Inference*
Zinc	$CuSO_4$	Blue color of $CuSO_4$ disappears and red-brown copper particles settle down at the bottom of the beaker.	Changes	Zn can displace Cu from $CuSO_4$ solution. $Zn + CuSO_4 \dashrightarrow ZnSO_4 + Cu$
	$Al_2(SO_4)_3$	Solution remains colorless.	No change	Zn cannot displace Al from $Al_2(SO_4)_3$ solution.
	$FeSO_4$	Green colour of ferrous sulphate disappears. Iron metal is settling down at the bottom of the beaker.	Changes	Zn can displace Fe from $FeSO_4$ solution. $Zn + FeSO_4 \dashrightarrow ZnSO_4 + Fe$
Copper	$ZnSO_4$	Solution remains colorless.	No change	Cu cannot displace Zn from $ZnSO_4$ solution.
	$Al_2(SO_4)_3$	Solution remains colorless.	No change	Cu cannot displace Al from $Al_2(SO_4)_3$ solution.
	FeSO4	Solution remains colorless.	No change	Cu cannot displace Fe from $FeSO_4$ solution.

Aluminium	$ZnSO_4$	Solution remains colorless. Zinc metal is settling down at the bottom of the beaker.	Changes	Al can displace Zn from $ZnSO_4$ solution. $2Al + 3ZnSO_4 \text{---}> 3Zn + Al_2(SO_4)_3$
	$CuSO_4$	Blue colour of $CuSO_4$ disappears. The brown colored copper particles settle down at the bottom of the beaker.	Changes	Al can displace Cu from CuSO4 solution. $2Al + 3CuSO_4 \text{---}> 3Cu + Al_2(SO_4)_3$
	$FeSO_4$	Green colour of $FeSO_4$ disappears. Iron metal is settling down at the bottom of the beaker.	Changes	Al can displace Fe from $FeSO_4$ solution. $2Al + 3FeSO_4 \text{----}> 3Fe + Al_2(SO_4)_3$
Iron	$ZnSO_4$	Solution remains colorless.	No change	Fe cannot displace Zn from ZnSO4 solution
	$CuSO_4$	Blue colour of $CuSO_4$ changes to light green colored $FeSO_4$. Copper metal is formed in the beaker.	Changes	Fe can displace Cu from CuSO4 solution. $Fe + CuSO_4 \text{-}=> Cu + FeSO_4$
	$Al_2(SO_4)_3$	Solution remains colorless.	No change	Fe cannot displace Al from $Al_2(SO_4)_3$ solution.

Conclusions:

- Aluminium is able to displace Zn, Cu and Fe from their salt solutions.
- Zinc is able to displace Cu and Fe from their salt solutions.
- Fe is able to displace Cu from its salt solution.
- Cu cannot displace Zn, Al or Fe from their salt solutions.
- Thus, aluminium is a more reactive metal. The reactivity of given metals in decreasing order is as given below:



Perform test:

№	Question	Variant
1	Na^+ and K^+ elements have:	a) tendency to gain electrons. b) positive reduction potential. c) negative reduction potential. d) negative oxidation potential.
2	Ca and Cu metals when coupled together gives maximum emf for a voltaic cell because:	a) It shows greater difference in standard reduction potential values. b) It shows less difference in standard reduction potential value. c) It shows same standard reduction potential value. d) None of these.
3	If half cell reaction $A + e^- ? A^-$ has large negative reduction potential, it shows that:	a) A is readily reduced b) A^- is readily reduced. c) A^- is readily oxidized. d) None of these.
4	Four alkali metals K, L, M and N are having standard electrode potentials as -4.05, -2.66, -0.50 and 0.70 V respectively. Which one of the following is the most reducing?	a) N b) M c) L d) K
5	The oxidation potential of Mg and Al are +2.37 and +1.66 volt respectively. The Mg in chemical reaction:	a) will not replace Al at all. b) will replace Al. c) will be replaced by Al d) none of the above.
6	Three metal cations A, B and C are having standard reduction potential values as 0.73, -2.17 and -1.15 V respectively. The order of reducing power of corresponding metals is:	a) C > A > B b) A > B > C c) B > C > A d) C > B > A
7	Use the activity series given to you to determine which of the following reactions will NOT take place:	a) $Zn + HNO_3$ --> b) $Ag + AuCl_3$--> c) $Ca + Pb(NO_3)_2$--> d) $Cu + HCl$-->

8	$Zn(s) \Rightarrow Zn^{2+}(aq) + 2e$ The chemical equation above shows a corrosive half-reaction for zinc. Which of these best represents what is occurring?	a) Zinc is reduced and is acting as the oxidizing agent. b) Zinc is oxidized and is losing electrons. c) Zinc is reduced and is losing electrons. d) Zinc is oxidized and is gaining electrons.
9	Corrosion, a type of redox reaction, can cause cars and buildings to rust. Rusting happens when iron is oxidized by oxygen in the air. This process happens when …...	a) oxygen acts as the reducing agent and gains electrons b) iron acts as the reducing agent and loses electrons c) oxygen acts as the oxidizing agent and loses electrons d) iron acts as the oxidizing agent and gains electrons
10	In order to develop building materials that are resistant to corrosion, manufacturers must be aware of ways to promote…	a) anodic inhibition, which prevents the reduction of the metal b) cathodic reactions, which allow reduction of the metal c) anodic inhibition, which prevents the oxidation of the metal d) cathodic reactions, which allows oxidation of the metal
11	Some silverware is not made entirely of silver but is electroplated, or coated, with silver. The process of electroplating a fork would involve the migration of .	a) silver ions oxidized at the anode and depositing on the fork, which acts as the cathode b) ions oxidized from the fork (cathode) and depositing on the silver metal (anode) c) silver ions reduced at the anode and depositing on the fork, which acts as the cathode d) ions oxidized from the fork (anode) and depositing on the silver metal (cathode)
12	Which of the following is true about metals?	a) Irregular order of arrangement of atoms b) Metallic luster c) Bad conductors of electricity d) Low boiling point
13	On the basis of sequence of	a) A, C

	reactions, identify the most and least reactive elements: A + BX → AX + B, C + AY → CY + A	b) C, A c) B, C d) C, B
14	Which one of the following metals does not displace Ag from a solution of $AgNO_3$?	a) Au b) K c) Ca d) Cu
15	Name a least reactive metal.	a) Sn b) Na c) Ca d) Pt
16	If iron filings are put in different test tubes containing $ZnSO_4$, $CuSO_4$, $Al_2(SO_4)_3$, $CaCl_2$, in which one changes will be observed?	a) $ZnSO_4$ b) $CuSO_4$ c) $CaCl_2$ d) $Al_2(SO_4)_3$
17	What is the correct decreasing order of the reactivity of the metals based on the reactions given below? $Fe(s) + Pb(NO_3)_2(aq) \rightarrow Fe(NO_3)_2(aq) + Pb(s)$, $Cu(s) + 2AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2Ag(s)$, $Zn(s) + FeSO_4(aq) \rightarrow ZnSO_4(aq) + Fe(s)$, $Pb(s) + CuCl_2(aq) \rightarrow PbCl_2(aq) + Cu(s)$	a) Zn > Fe > Pb > Cu > Ag b) Pb > Zn > Fe > Cu > Ag c) Fe > Zn > Cu > Ag > Pb d) Ag > Cu > Pb > Fe > Zn
18	Which one of the following is true about displacement reaction?	a) One component of both the reacting molecules gets exchanged to form the product. b) A single compound splits into two or more simple substances. c) Highly reactive metal replaces the least reactive one from its salt solution. d) Only a single product is formed.
19	Some crystals of copper sulphate were dissolved in water. What colour would be the solution obtained?	a) Green b) Blue c) Red d) Brown
20	Identify the set up in which reaction occurs.	a) $FeSO_4 + Cu \Rightarrow$ b) $Ag + HCl \Rightarrow$

		c) Al + $CuSO_4$ = > d) H_2 + CaO =>
21	A student puts one big iron nail in each of four test tubes containing solutions of zinc sulphate, aluminium sulphate, copper sulphate and iron sulphate. A reddish brown coating was observed only on the surface of the iron nail which was put in which solution?	a) Zinc sulphate b) Copper sulphate c) Iron sulphate d) Aluminium sulphate
22	Iron nails were dipped in a solution kept in a test tube. After half an hour, it was observed that the colour of the solution had changed. Which solution was in the test tube?	a) Zinc sulphate b) Iron sulphate c) Copper sulphate d) Aluminium sulphate
23	When an aluminium strip is kept immersed in freshly prepared ferrous sulphate solution taken in a test tube, what change is observed?	a) The green solution slowly turns brown. b) The green solution turns colourless. c) A colourless gas with smell of burning sulphur is observed. d) Light green solution changes to blue.
24	Which of the following statements is not correct for the following reaction ? Cu + $ZnSO_4$ → $CuSO_4$ +Zn?	a) The reaction does not occur. b) Cu is less reactive than Zn. c) The colourless solution is $ZnSO_4$. d) Zn is less reactive than Cu.
25	Four test tubes were taken and marked A, B, C and D respectively. 2 mL of solution of $Al_2(SO_4)_3$ in water was filled in each of the four test tubes. Clean piece of metal Zinc was placed in test tube A, clean iron nail was put in test tube B, clean copper wire was placed in test tube C and a clean aluminium wire was placed in test tube D. It was observed that no change occurred in any of the test tubes. What is the correct inference?	a) Zinc is more active than Aluminium. b) Zinc is more active than Copper. c) Copper is more active than Aluminium. d) Zinc, Iron and copper are less active than Aluminium.

26	In which case will not a chemical reaction take place?	a) Zinc metal and copper sulphate solution b) Copper metal and copper sulphate solution c) Iron metal and copper sulphate solution d) Aluminium metal and copper sulphate solution
27	What is the correct procedure to show that zinc metal is more reactive than copper?	a) Prepare copper sulphate solution and dip zinc strip into it. b) Prepare zinc sulphate solution and dip a copper strip into it. c) Heat zinc and copper strips. d) Add dilute nitric acid on both strips.
28	A strip of copper was placed in zinc sulphate solution in a beaker. What was seen while observing on the next day?	a) Copper strip was unchanged. b) Copper strip became thinner. c) Copper strip became thicker. d) Colour of copper strip changed.
29	In which case does a chemical reaction take place?	a) Zinc metal and zinc sulphate solution b) Aluminium metal and zinc sulphate solution c) Iron metal and zinc sulphate solution d) Copper metal and zinc sulphate solution

LABORATORY WORK
METAL CORROSION

Objectives: *Students will be able to observe and record the corrosive nature of oxidation-reduction reactions and to determine the electro-chemical series of selected metals (relative strengths of oxidizing and reducing agents).*

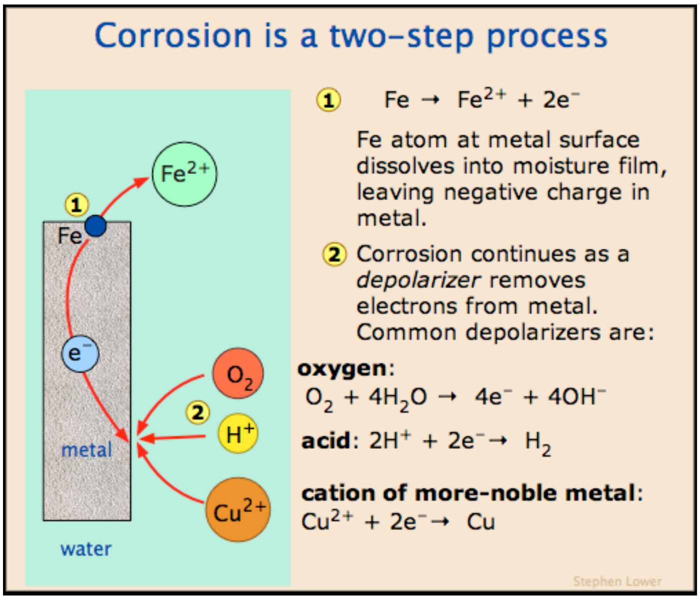
Corrosion can be defined as the degradation of a material when it comes in contact with the environment.

The dissolving of a material by a corrosive liquid is called *chemical corrosion*. The material continues to dissolve until either it is consumed or the liquid is saturated. A simple example is salt dissolving in water.

The removing of metal atoms from a solid material as the result of an electric circuit is called *electrochemical corrosion*. In this form of corrosion, metal atoms lose electrons and thus become ions forming a byproduct.

Electrochemical corrosion occurs most frequently in aqueous mediums, in which ions are present in water or moist air. In this process, an electric circuit is created and the system is called an electrochemical cell. Corrosion of a steel pipe or a steel automobile panel, creating holes in the steel and rust as the byproduct, are examples of this reaction.

Corrosion:



- Silver articles become black after some time when exposed to air. This is because it reacts with sulphur in the air to form a coating of silver sulphide.
- Copper reacts with moist carbon dioxide in the air and slowly loses its shiny brown surface and gains a green coat. This green substance is copper carbonate.
- Iron when exposed to moist air for a long time acquires a coating of a brown flaky substance called rust.

Prevention of Corrosion:

- The rusting of iron can be prevented by painting, oiling, greasing, galvanising, chrome plating, anodising or making alloys.
- Galvanisation is a method of protecting steel and iron from rusting by coating them with a thin layer of zinc. The galvanised article is protected against rusting even if the zinc coating is broken.
- Alloying is a very good method of improving the properties of a metal. An alloy is a homogeneous mixture of two or more metals, or a metal and a nonmetal. It is prepared by first melting the primary metal, and then, dissolving the other elements in it in definite proportions. It is then cooled to room temperature. We can get the desired properties by this method.
- For example, iron is the most widely used metal. But it is never used in its pure state. This is because pure iron is very soft and stretches easily when hot. But, if it is mixed with a small amount of carbon (about 0.05 %), it becomes hard and strong.
- When iron is mixed with nickel and chromium, we get stainless steel, which is hard and does not rust. Thus, if iron is mixed with some other substance, its properties change.

Rusting of metals is a special case of metal oxidation. Iron will oxidize to form rust. Water will cause metals to rust; this reaction can be accelerated by adding salts. In the corrosion process, metals get oxidized. For example in mild steel (which is greater than 99% iron) the metal corrodes according to the following: $Fe \rightarrow Fe^{2+} + 2\bar{e}$ $(the\ removal\ of\ \bar{e})$

These electrons are consumed by reacting with another substance (usually oxygen but it can be H^+ in acids) in reduction as in

$$O_2 + 4\bar{e} = 2H_2O \rightarrow 4OH^- (the\ gaining\ of\ \bar{e})$$

In an acid solution, the reduction is: $2H^+ + 2\bar{e} \rightarrow H_2^\circ$

These equations indicate that in order for metals to corrode (rust), two reactions occur; an oxidation that converts metal to metal ions and electrons and a second reaction which consumes those electrons by converting oxygen and water to hydroxide ions. In order for these reactions to occur, the electrons must be transported from the place where the metal dissolves to the place where the oxygen is consumed and an ionic current must also flow between the sites to complete the circuit. This ionic current flows more easily through water containing electrolytes (i.e., NaCl). This accounts for the rapid rusting of unprotected steel in a salty environment.

The final product of iron oxidation (rust) is usually a ferric oxide (often hematite Fe_2O_3). The initial corrosion product of the anodic reaction is ferrous (Fe^{2+}) ion. This is subsequently oxidized to Fe^{3+} by exposure to oxygen.

Hydrogen ions are being consumed by the process. As the iron corrodes, the pH in the droplet rises. Hydroxide ions (OH^-) appear in water as the hydrogen ion concentration falls. They react with the iron (II) ions to produce insoluble iron (II) hydroxides or green rust:

$$\mathbf{Fe^{2+}_{(aq)} + 2OH^-_{(aq)} \dashrightarrow Fe(OH)_{2(s)}}$$

The iron (II) ions also react with hydrogen ions and oxygen to produce iron (III) ions:

$$\mathbf{4Fe^{2+}_{(aq)} + 4H^+_{(aq)} + O_{2(aq)} \dashrightarrow 4Fe^{3+}_{(aq)} + 2H_2O_{(l)}}$$

The iron (III) ions react with hydroxide ions to produce hydrated iron(III) oxides (also known as iron (III) hydroxides):

$$\mathbf{Fe^{3+}_{(aq)} + 3OH^-_{(aq)} \dashrightarrow Fe(OH)_{3\,(s)}}$$

The loose porous rust or $Fe(OH)_3$ can slowly transform into a crystallized form written as $Fe_2O_3.H_2O$ the familiar red-brown stuff that is called "rust" forming tubercles as shown here. Since these processes involve hydrogen ions or hydroxide ions, they will be affected by changes in pH. With limited O_2, magnetite is formed (Fe_3O_4).

The corrosion process may be slowed by coating the metals with other metals or polymers in order to protect the metal from the corrosive environment. Examples of this can be seen in food cans which have a polymer coating and in galvanized steel where iron is coated with zinc.

When we put two metals in direct contact, one can oxidize (rust) while the other reduces oxygen. This reaction sets up a voltage and is the primary reaction in a battery. By measuring this voltage, it is possible to construct a list ranking the metal's oxidation tendencies. If metals which are far apart in oxidation tendencies are placed in contact with each other and with an electrolyte solution, severe corrosion of one metal can occur.

1st Experiment. Rusting of iron.

Only Iron and steel rust. Other metals corrode. Rusting is an oxidation process. What we normally call rust is a flaky red-brown solid which is largely hydrated iron?

The primary corrosion product of iron is $Fe(OH)_2$ (or more likely $FeO*nH_2O$), but the action of oxygen and water can yield other products having different colors:

- $Fe_2O_3*H_2O$ (hydrous ferrous oxide, sometimes written as $Fe(OH)_3$) is the principal component of red-brown rust. It can form a mineral called hematite.
- $Fe_3O_4*H_2O$ ("hydrated magnetite" or ferrous ferrite, Fe_2O_3*FeO) is most often green but can be deep blue in the presence of organic complexants as shown here.
- Fe_3O_4 ("magnetite") is black.

Materials and reagents: test tubes, water, 12% acetic acid solution (CH_3COOH) or vinegar, sodium hydroxide (NaOH) concentrated solution, nails.

Procedure:

1) For the first experiment put a 5 cm finishing nail brushed with steel wool in a test tube containing 8 mL of tap water. The picture was taken three days later.

2) The second experiment was carried out with an acidic solution containing 12% acetic acid. The second picture was taken an hour after the start of the experiment. You can produce a similar reaction with vinegar which typically contains 3% acetic acid in water. However, you will have to be more patient than we were.

3) In the third experiment a pellet of sodium hydroxide (caustic 'Drano') was dropped in the test tube of the second experiment. The test tube was capped to prevent air to react with the solution. After three hours blue rust precipitates had more or less settled out as you can see in this picture:

1

Red-brown flaky rust

2

Black magnetite

3

Blue/green unstable rust

Formula	Color	Oxidation State	Structure / comments
$Fe_2O_3*H_2O$ or $Fe(OH)_3$	red brown	Fe^{3+}	a-form Hematite, b-form used in recording tapes
Fe_3O_4	black	Fe^{2+} and Fe^{3+}	Magnetite / lodestone
$Fe(OH)_2$	blue/green	Fe^{2+}	Soluble, the color going from yellow to green and blue by changing the pH of the solution from acidic to very basic
FeO	black	Fe^{2+}	Pyrophoric

2nd Experiment. Cathodic Protection

This is a little corrosion experiment where we will show how Cathodic Protection works and how is possible to protect a metal, an iron nail in this case, from corrosion using the principle of Galvanic Corrosion.

The same principle is used to protect Ship and Offshore structure like Oil&Gas Platform from Corrosion.

This experiment shows also how nails rust when are exposed to water. What we will do to expose two identical iron nails to the same electrolyte, in this case fresh water with two spoons of NaCl, salt that you use for cooking. One of the nails will be connected through an electrical wire to a piece of aluminium foil, the same will be used for cooking, while the second nail will be immersed in the same electrolyte.

We will leave the two nails for two days immersed in water and then we will see if Cathodic Protection has protected the first nail from corrosion.

Materials and reagents: two identical iron nails; a piece of aluminium foil; an electrical wire; something to hang the second nail, we've used an elastic a glass; two spoons of salt, NaCl, the same that you use for cooking

Procedure:

Fig 1 - The electrical wire and elastic

Fig 2 – The aluminium foil

In the 1 and 2 pictures below you can see how we've connected the nail to the cable and to the aluminium foil. To fix the aluminum foil just wrap it around the electrical cable and be sure that there is a contact between the aluminum and copper.

Fig 3 - Glass with water and salt.

You can see how we have immersed the nails. This picture has been taken after two days of exposure. The nail that is not connected to aluminium is corroded.

Fig 4 – Nails types

Fig 4 - On the left you can see the nail without cathodic protection, in the middle the nail connected to the aluminium foil, cathodically protected, on the right a new nail.

Fig 5 -Nails after two days of immersion

You can see how the nail without cathodic protection, that is not connected to the aluminium is severely corroded, see the red rust, while the second nail shown only minor sign of corrosion.

Fig 6 - Rusted Nail after two days of exposure

You can see the red rust, iron oxide (fig. 6).

In this experiment we've shown how cathodic protection works. When you connect a metal, in this case aluminium, that in the Galvanic series of metals is below Iron, has a lower potential and is anodic, corrodes and protect Iron from Corrosion. The nail that is not connected to aluminium will corrode severely only after two days of exposure.

The amount of corrosion is a function of several factors, like the temperature of water, the amount of salt dissolved, the chemical composition of the nail. In this case as you can see in the picture, the nails are galvanized, it means that they have a thin superficial layer of zinc deposited on the surface, and this gives greater resistance to corrosion compared to steel. If we use a nail that is not galvanized it corrodes faster.

Perform test:

№	Question	Variant
1	Identify the correct statement	a) Rusting is the oxidation of any metal b) Corrosion is the oxidation of any metal with rusting being about iron only c) Rusting is the oxidation of any metal
2	When iron rusts, which of these ion-electron equations is correct?	a) $Fe_{(s)} \rightarrow Fe_{(aq)}$ b) $Fe^{2+}_{(aq)} + 2e \rightarrow Fe_{(s)}$ c) $Fe(s) \rightarrow Fe^{2+} + 2e-$
3	What is the colour change that Ferroxyl indicator shows when iron is oxidised to ions?	a) Yellow → pink b) Blue → yellow c) Yellow → blue
4	A motoring organisation noticed that cars rust faster when owned by people who live by the sea in comparison to people who live inland. What is the main reason for this?	a) Sea water is a salt solution full of ions b) The wind increases the concentration of oxygen at the seaside c) Seagull droppings are acidic
5	How does salt speed corrosion?	a) Salt contains oxygen b) Salt is an electrolyte and can carry charged particles c) Salt is sodium chloride and sodium is reactive
6	Elise investigated rusting by attaching steel to a car battery. Which is the true statement?	a) Attaching steel to the negative terminal of a battery protects it from rusting b) Attaching steel to the positive terminal of a battery protects it from rusting c) Attaching steel to a battery has no effect on rusting
7	Identify the oxidation equation involved in corrosion.	a) $Fe^{2+}_{(aq)} + 2e^- \rightarrow Fe_{(s)}$ b) $Fe^{2+}_{(aq)} \rightarrow Fe^{3+}_{(aq)} + e^-$ c) $2H_2O_{(l)} + O_{2(g)} + 4e^- \rightarrow 4OH^-_{(aq)}$
8	Identify the reduction equation involved in corrosion.	a) $Fe^{2+}_{(aq)} + 2e^- \rightarrow Fe_{(s)}$ b) $Fe^{2+}_{(aq)} \rightarrow Fe^{3+}_{(aq)} + e^-$ c) $2H_2O_{(l)} + O_{2(g)} + 4e^- \rightarrow 4OH^-_{(aq)}$
9	Magnesium metal can be attached to the steel pipeline	a) Galvanising b) Sacrificial

	to prevent rusting. What name is given to this type of protection provided by magnesium?	c) Anodising
10	Steel cars can be protected from rusting in a number of ways. Dipping steel in molten zinc can also physically protect steel from rust. Why would the steel not rust even if the zinc is scratched?	a) The molten zinc will fill in the scratch b) Zinc is less reactive than steel and will give electrons to steel c) Zinc is more reactive than steel and will give electrons to steel
11	Which of these metal elements has to be stored under oil?	a) mercury b) sodium c) aluminium
12	Which of these metal elements produces hydrogen gas when it is placed in hydrochloric acid?	a) iron b) copper c) silver
13	Which of these metal elements will form an alkali when dropped into cold water?	a) lead b) aluminium c) calcium
14	Which of the following groups of 4 metals is arranged in descending order of their chemical reactivity?	a) potassium, magnesium, lead, iron b) sodium, aluminium, tin, gold c) calcium, zinc, aluminium, silver
15	Which of these sets of elements contain metals all known to mankind in prehistoric times?	a) aluminium, silver and mercury b) mercury, silver and gold c) calcium, copper and tin
16	In which of the following pairs of metal elements are both easily extracted from their ores by heating with carbon?	a) tin and copper b) magnesium and silver c) zinc and potassium
17	When sodium reacts with water the two products of the reaction are:	a) sodium oxide and hydrogen b) sodium hydroxide and oxygen c) sodium hydroxide and hydrogen
18	Which of the following shows two metals which will both react quickly with cold water?	a) aluminium and calcium b) potassium and sodium c) sodium and magnesium

19	Identify the pair of substances which would react together to give a gas as one of their products?	a) iron and dilute sulphuric acid b) copper and water c) magnesium and oxygen
20	Choose the pair of metal elements where both are required to be extracted from their ores by means of electrolysis.	a) copper and zinc b) magnesium and iron c) potassium and aluminium

QUESTIONS FOR THE EXAM

Elements of Chemical Thermodynamics

1. Subject and tasks of chemical thermodynamics. Chemical thermodynamics as the basis of bioenergetics. Isolated, closed and open systems.
2. The first law of thermodynamics. Internal energy, heat and work. Isobaric and isochoric thermal processes. Enthalpy.
3. Hess’s law and its corollaries. Standard formation and combustion heat. Thermochemical calculations and their usage for energetic characteristic of biochemical processes.
4. Interconnection between the processes of metabolism and energy exchange. Caloric value of main constituents of food and some food products. Energy consumption at different modes of moving activity.
5. Thermodynamically reversible and irreversible processes. The second law of thermodynamics. Entropy. Statistic and thermodynamic explanation of entropy. Standard entropy.
6. The Gibbs free energy (isobaric-isothermal potential). Enthalpy and entropy factors. Endothermic and exothermic processes.
7. Thermodynamics of chemical equilibrium. Reversible and irreversible reactions. Concept of chemical equilibrium. Constant of the chemical equilibrium. The interconnection between the constant of chemical equilibrium and the Gibbs free energy. Equations of isotherm and isobaric curve of a chemical reaction.

Elements of Chemical Kinetics

8. Main concepts of chemical kinetics. Simple and complex, homogeneous and heterogeneous reactions. The speed of homogeneous chemical reactions and methods of its measuring.
9. The main postulate of chemical kinetics. The order of reaction and the reaction speed constant. The Law of mass action for the speed of the reaction and its sphere of application.
10. Kinetic equations of the reactions of zero, first and second order. Period of semi-transformation. Molecularity of the reaction.
11. Theory of active collisions. Arrhenius' equation. Energy of activation. Vant-Hoff's rule. Temperature coefficient of the reaction speed for enzymatic processes.
12. Catalysis and catalysts. The theories of catalysis. The mechanism of homogeneous and heterogeneous catalysis. Enzymes as biological catalysts, peculiarities of their action.

Colligative Properties of Solutions

13. Thermodynamics of solution formation.
14. Osmose and osmotic pressure of solutions. Vant-Hoff's law.
15. Osmotic pressure, osmolarity and osmolality of some biological fluids. The concept of isotonic, hypertonic and hypotonic solutions.
16. The role of osmotic phenomena in biological processes.
17. The pressure of saturated vapor of solvent above the solution. Raoul's first law.
18. Boiling and freezing temperatures of solvents. Raoul's second law. Cryoscopy. Ebullioscopy.
19. Colligative properties of electrolyte solutions. Isotonic coefficient.

Electrolyte solutions. Acidity and basicity of aqueous solutions, pH.

20. The theory of weak electrolyte solutions. Main characteristics of a weak electrolyte: pH, K_{ion}, pK_{ion}.
21. The theory of strong electrolyte solutions. Main characteristics of a strong electrolyte: a, fa, I.
22. Protolytic theory of acids and bases.
23. The ion product of water. Hydrogen ion exponent pH.
24. Calculation of solution pH of weak and strong acids and bases.
25. Determination of hydrogen ion exponent.
26. Role of hydrogen ions in biological processes.
27. Buffer systems, their classifications.
28. Calculation of pH of acid and basic buffer solutions.
29. Mechanism of action of buffer systems.
30. Buffer capacity.

Atomic Properties and Periodic Trends

31. The electron shell of the atom. Energy levels. The wave nature of microparticles motion. The uncertainty principle. The electron cloud.

32. The quantum numbers. The principal quantum number n. The angular momentum quantum number ℓ. Magnetic quantum number $m\ell$. Electron spin. The shapes of atomic orbitals.

33. Electron configurations of the elements. The Pauly principle. Hund's rule. Order of orbital energies and assignments. Orbital energies. Electron configurations of the main group elements. Electron configurations of the transition elements.

34. The periodic table. Atomic properties and periodic trends. Atomic size. Ionization energy. Electron affinity.

Chemical Bond

35. Chemical bond formation.Valence electrons. The formation of a covalent bond in H_2 molecule.

36. Properties of covalent bond. The bond order. Single and multiple bonds. Sigma and π-bond. The donor–accepter mechanism of formation of covalent bond. Bond length. Bond energy. 60. Polarity and electronegativity. Oxidation numbers. Ionic bond.

37. Orbital hybridization. Hybrid orbitals. Molecular Shape. Molecular polarity.

38. Orbital hybridization in CH4 , H2O and NH3 molecules.

39. Theories of chemical bonding. The valence bond (VB) theory and the molecular orbital (MO) theory. Molecular orbital theory: principles of molecular orbital theory, bonding and antibonding molecular orbitals, bond order. Electron configurations for H_2 molecule.

Complex (Coodination) Compounds

40. The structure of complex compounds: inner sphere, outer sphere, central atom, ligands, coodination number, dentation of ligands.

41. The nature of the chemical bond in complex compounds.

42. Classification and nomenclature of complex compounds. Cyclic complexes or chelates.

43. Dissociation of complex compounds in solutions. Destruction of complex compounds.

44. Equilibria and the processes in solutions with complex compounds.

Dispersion Systems

45. Dispersion systems, their peculiarities and classification.

46. Molecular-kinetic properties of colloidal systems. Sedimentation.

47. Optical properties of colloidal systems . Opalescence.

48. Structure of colloid particles.

49. Methods of obtaining and purification of colloidal systems. Peptization.

50. Colloid protection and its importance.

GLOSSARY

Acid - a compound that, when dissolved in water, gives a pH of less than 7.0 or a compound that donates a hydrogen ion

anion - negatively charge ions

atom - a chemical element in its smallest form, and is made up of neutrons and protons within the nucleus and electrons circling the nucleus

atomic number - the number representing an element which corresponds with the number of protons within the nucleus

atomic orbital - the region where the electron of the atom may be found

absolute zero - Absolute zero is °K. It is the lowest possible temperature. Theoretically, at absolute zero, atoms stop moving.

accuracy - Accuracy is a measure of how close a measured value is to its true value. For example, if an object is exactly a meter long and you measure it as 1.1 meters long, that is more accurate than if you measured it at 1.5 meters long.

acid - There are several ways to define an acid, but they include any chemical that gives off protons or H^+ in water. Acids have a pH less than 7. They turn the pH indicator phenophthalein colorless and turn litmus paper red.

acid anhydride - An acid anhydride is an oxide that forms an acid when it is reacted with water. For example, when SO_3^- is added to water, it becomes sulfuric acid, H_2SO_4.

actual yield - The actual yield is the amount of product you actually obtain from a chemical reaction, as in the amount you can measure or weigh as opposed to a calculated value.

addition reaction - An addition reaction is a chemical reaction in which atoms add to a carbon-carbon multiple bond.

alkali metal - An alkali metal is a metal in Group I of the periodic table. Examples of alkali metals include lithium, sodium, and potassium.

alkaline earth metal - An alkaline earth metal is an element belonging to Group II of the periodic table. Examples of alkaline earth metals are magnesium and calcium.

allotrope - Allotropes are different forms of a phase of an element. For example, diamond and graphite are allotropes of carbon.

alpha particle - An alpha particle is another name for a helium nucleus, which contains two protons and two neutrons. It's called an alpha particle in reference to radioactive (alpha) decay.

Base - a substance that accepts a proton and has a high pH; a common example is sodium hydroxide (NaOH)

boiling - the phase transition of liquid vaporizing

bond - the attraction and repulsion between atoms and molecules that is a cornerstone of chemistry

burette (also buret) - glassware used to dispense specific amounts of liquid when precision is necessary (e.g. titration and resource dependent reactions)

binary compound - A binary compound is one made up of two elements.

binding energy - Binding energy is the energy that holds protons and neutrons together in the atomic nucleus.

bond energy - Bond energy is the amount of energy required to break one mole of chemical bonds.

bond length - Bond length is the average distance between the nuclei of two atoms that share a bond.

buffer - A liquid that resists change in pH when an acid or base is added. A buffer consists of a weak acid and its conjugate base. An example of a buffer is acetic acid and sodium acetate.

Catalyst - a chemical compound used to change the rate (either to speed up or slow down) of a reaction, but is regenerated at the end of the reaction.

cation - positively charged ion

calorimetry - Calorimetry is the study of heat flow. Calorimetry may be used to find the heat of reaction of two compounds or the heat of combustion of a compound, for example.

cathode - A cathode is the electrode which gains electrons or is reduced. In other words, it is where reduction occurs in an electrochemical cell.

chemical equation - A chemical equation is a description of a chemical reaction, including what reacts, what is produced, and which direction(s) the reaction proceeds.

chemical property - A chemical property is a property that can only be observed when achemical change occurs. Flammability is an example of a chemical property, since you can't measure how flammable a substance is without igniting it (making/breaking chemical bonds).

covalent bond - A covalent bond is a chemical bond formed when two atoms share two electrons.

critical mass - Critical mass is the minimum quantity of radioactive material needed to cause a nuclear chain reaction.

critical point - The critical point is the endpoint of the liquid-vapor line in a phase diagram, past which a supercritical liquid forms. At the critical point, the liquid and vapor phases become indistinguishable from one another.

crystal - A crystal is an ordered, repeating three-dimensional pattern of ions, atoms, or molecules. Most crystals are ionic solids, although other forms of crystals exist.

centrifuge - equipment used to separate substances based on density by rotating the tubes around a centred axis

cell potential - the force in a galvanic cell that pulls electron through reducing agent to oxidizing agent

chemical Law - certain rules that pertain to the laws of nature and chemistry – examples

chemical reaction - the change of one or more substances into another or multiple substances

colloid - mixture of evenly dispersed substances, such as many milks

compound - a substance that is made up of two or more chemically bonded elements

condensation - the phase change from gas to liquid

conductor - material that allows electric flow more freely

covalent bond - chemical bond that involves sharing electrons

crystal - a solid that is packed with ions, molecules or atoms in an orderly fashion

cuvette - glassware used in spectroscopic experiments. It is usually made of plastic, glass or quartz and should be as clean and clear as possible

Dissolution or solvation - the spread of ions in a solvent

double bond - sharing of two pairs of electrons

diffusion - Diffusion is the movement of particles from an area of higher concentration to one of lower concentration.

dilution - Dilution is when solvent is added to a solution, making it less concentrated.

dissociation - Dissociation is when a chemical reaction breaks a compound into two or more parts. For example, NaCl dissociates into Na^+ and Cl^- in water.

double displacement reaction - A double displacement or double replacement reaction is when cations of two compounds switch places.

Earth metal - see alkaline earth metal electrolyte - a solution that conducts a certain amount of current and can be split

electrochemical cell - using a chemical reaction's current, electromotive force is made

electron - a subatomic particle with a net charge that is negative

electron shells - an orbital around the atom's nucleus that has a fixed number

electric charge - a measured property (coulombs) that determine electromagnetic interaction

element - an atom that is defined by its atomic number

energy - A system's ability to do work

enthalpy - Enthalpy is a measure of the amount of energy in a system (usually symbolized as H)

entropy - Entropy is a measure of the disorder or randomness in a system (usually symbolized as S)

electrolysis - Electrolysis is using electricity to break the bonds in a compound to break it apart.

electrolyte - An electrolyte is an ionic compound that dissolves in water to produce ions, which can conduct electricity. Strong electrolytes completely dissociate in water, while weak electrolytes only partially dissociate or break apart in water.

Faraday constant - a unit of electrical charge widely used in electrochemistry and equal to ~ 96,500 coulombs. It represents 1 mol of electrons, or the Avogadro number of electrons: 6.022×10^{23} electrons. F = 96 485.339 9(24) C/mol

Faraday's law of electrolysis - a two part law that Michael Faraday published about electrolysis

freezing - phase transition from liquid to solid

Galvanic cell - battery made up of electrochemical with two different metals connected by salt bridge

gas - particles that fill their container though have no definite shape or volume

Gibbs energy - value that indicates the spontaneity of a reaction (usually symbolized as G)

Halogens - Group 17 on the Periodic Table and are all non-metals

heat - energy transferred from one system to another by thermal interaction

Iodium – Latin name of the halogen element iodine

indicator - a special compound added to solution that changes color depending on the acidity of th e solution; different indicators have different colors and effective pH ranges

inorganic compound - compounds that do not contain carbon, though there are exceptions (see main article)

inorganic chemistry - a part of chemistry concerned with inorganic compounds

ion - a molecule that has gained or lost one or more electrons

ionic bond - electrostatic attraction between oppositely charged ions

Kinetics - A sub-field of chemistry specializing in reaction rates

Kinetic energy - The energy of an object due to its motion.

Lanthanides - Elements 57 through 71

lattice - Unique arrangement of atoms or molecules in a crystalline liquid or solid.

liquid - A state of matter which takes the shape of its container

law of conservation of energy - The law of conservation of energy states the energy of the universe may change form, but its amount remains unchanged.

ligand - A ligand is a molecule or ion stuck to the central atom in a complex. Examples of common ligands include water, carbon monoxide, and ammonia.

Metal - Chemical element that is a good conductor of both electricity and heat and forms cations and ionic bonds with non-metals.

melting - The phase change from a solid to a liquid

metalloid - A substance possessing both the properties of metals and non-metals

microcentrifuge - a small plastic container that is used to store small amounts of liquid

mole - abbreviated mol - a measurement of an amount of substance; a single mole contains approximately 6.022×10^{23} units or entities

molecule - a chemically bonded number of atoms that are electrically neutral

molecular orbital - region where an electron can be found in a molecule (as opposed to an atom)

Neutron - a neutral unit or subatomic particle that has no net charge

nucleus - the centre of an atom made up of neutrons and protons, with a net positive charge

noble gases - group 18 elements, those whose outer electron shell is filled

non-metal - an element which is not metallic

number density – a measure of concentration of countable objects (atoms, molecules, etc.) in space; number per volume

Orbital - may refer to either an atomic orbital or a molecular orbital

organic compound - compounds that contain carbon

organic chemistry - a part of chemistry concerned with organic compounds

oxidation number is the apparent charge on an atom. For example, the oxidation number of an oxygen atom is -2.

Poor metal - Metallic elements in the p-block, characterized by lower melting and boiling points than other metals

proton - a positive unit or subatomic particle that has a positive charge

protonation - the addition of a proton (H+) to an atom, molecule, or ion

pH - the measure of acidity (or basicity) of a solution

radioactivity - Radioactivity occurs when the atomic nucleus is unstable and breaks apart, releasing energy or radiation.

Raoult's Law - Raoult's Law states that the vapor pressure of a solution is directly proportional to the mole fraction of solvent.

rate determining step - The rate determining step is the slowest step in any chemical reaction.

rate law - A rate law is a mathematical expression relating the speed of a chemical reaction as a function of concentration.

redox reaction - A redox reaction is a chemical reaction that involves oxidation and reduction.

Salts - ionic compounds composed of anions and cations.

salt bridge - devices used to connection reduction with oxidation half-cells in an electrochemical cell.

saline solution - general term for NaCl in water.

s-block elements - Group 1 and 2 elements (alkali and alkaline metals), which includes Hydrogen and Helium.

single bond - sharing one pair of electrons.

solid - one of the states of matter, where the molecules are packed close together, there is a resistance of movement/deformation and volume change.

solution - homogeneous mixture made up of multiple substances. It is made up of solutes and solvents.

solvent - the part of the solution that dissolves the solute (H2O in saline water).

spectroscopy - study of radiation and matter, such as X-ray absorption and emission spectroscopy.

speed of light - the speed of anything that has zero rest mass (Energyrest = mc^2 where m is the mass and c is the speed of light).

Standard conditions for temperature and pressure or SATP - a standardisation used in order compare experimental results (25 °C and 100.000 kPa).

state of matter - matter having a homogeneous, macroscopic phase; gas, plasma, liquid, and solid are the most well known (in increasing concentration).

strong acid - A strong acid is an acid that completely dissociates in water. An example of a strong acid is hydrochloric acid, HCl, which dissociates into H^+ and Cl^- in water.

sublimation - Sublimation is when a solid changes directly into a gas. At atmospheric pressure, dry ice or solid carbon dioxide goes directly into carbon dioxide vapor, never becoming liquid carbon dioxide.

synthesis - synthesis is making a larger molecule from two or more atoms or smaller molecules.

system - system includes everything you are evaluating in a situation.

subatomic particles - particles that are smaller than an atom; examples are protons, neutrons and electrons.

substance - material with definite chemical composition.

Temperature - the average energy of microscopic motions of particles.

theory - a model describing the nature of a phenomenon.

thermochemistry - the study of absorption/release of heat within a chemical reaction.

thermodynamics - the study of the effects of changing temperature, volume or pressure (or work, heat, and energy) on a macroscopic scale.

thermodynamic stability - when a system is in its lowest energy state with its environment (equilibrium).

titration - the process of titrating one solution with another, also called volumetric analysis.

Tyndall effect - the effect of light scattering by colloidal (mixture where one substance is dispersed evenly through another) or suspended particles

Unit cell - the smallest repeating unit of a lattice.

universal or ideal gas constant - proportionality constant in the ideal gas law (0.08206 L·atm/(K·mol)).

Valence electron - the outermost electrons of an atom, which are located in electron shells.

Valence bond theory - a theory explaining the chemical bonding within molecules by discussing valencies, the number of chemical bonds formed by an atom.

van der Waals force - one of the forces (attraction/repulsion) between molecules.

Vapor - when a substance is below the critical temperature while in the gas phase.

vapour pressure - pressure of vapour over a liquid at equilibrium.

Water - H_2O - a chemical substance, a major part of cells and Earth, and covalently bonded.

work - the amount of force over distance and is in terms of joules (energy)

X-ray - form of ionizing, electromagnetic radiation, between gamma and UV rays.

X-ray photoelectron spectroscopy - a spectroscopic technique to measure composition of a material.

Zone melting - a way to remove impurities from an element by melting it and slowly travel down an ingot (cast).

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