KING SAUD UNIVERSITY
COLLEGE OF SCIENCES
CHEMISTRY DEPARTMENT

CHEM 201

Laboratory of General Chemistry (2)

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1442 H

List of Experiments

Grading

Description	Marks
Reports	**10**
Quiz	**5**
Final exam	**15**

INTRODUCTION

Our aims in this course are:

- Understand the importance of applying safety and security in chemical laboratories.
- Skill and experience in using tools and devices, and how to deal with different chemicals used in laboratories.
- Learn about chemical changes and scientific phenomena and be able to explain and understand them.

LAB SAFETY

In any laboratory, safety is of paramount importance. The experiments you will perform are designed to minimize hazards, but dangerous materials are involved and accidents can happen. The safety rules given below are meant to prevent accidents and to minimize injuries.

EYE PROTECTION

In labs, there is the potential for spills and splashes of corrosive chemicals and explosions involving broken glass. Getting a chemical in your eyes can be agonizing, even if it would cause no problem anywhere else on your body. (Think about soapy water!) Of course, your eyes are more susceptible to major injury than other parts of your body. For these reasons, eye protection is crucial in the laboratory.
You must always wear safety glasses while in the laboratory. Wearing contact lenses in lab is strongly discouraged.

CLOTHING AND PERSONAL ITEMS

Students are urged to dress with potential lab hazards in mind. Clothing should protect as much of the body as possible. Clothing may have to be immediately removed if grossly contaminated with chemicals or ignited. The following rules apply:

1. Shirt must cover shoulders, frontal area, and extend approximately 6" below the waist
2. Pants must extend to ankles (no tights, leggings or capri pants)
3. Torso must be covered when bending (no bare midriffs)
4. Shoes must cover toes and heels (no flip flops)
5. Long hair must be tied back
6. Do not wear valuable jewelry while working in the lab.
7. All radios and headsets are forbidden in the lab.

HANDLING CHEMICALS

The easiest routes for possibly dangerous chemicals and vapors to enter the body are via inhalation or ingestion. Avoid inhalation of fumes of any kind. No eating or drinking in the laboratory. Never bring food into the lab and never taste any chemicals in the lab. Also, do not place your mouth on any piece of equipment in the laboratory.

MATERIALS SAFETY DATA SHEETS (MSDS)

Full safety data on all chemicals used in the laboratory are included in their Material Safety Data Sheet (MSDS). This includes toxicology, detailed first aid and proper disposal and handling instructions. MSDS sheets are available at http://www.ncsu.edu/ehs/MSDS.htm. They are also available on some chemical distributor web pages or by calling any chemical distributor or manufacturer of the chemical in question.

Experiment (1)
Percent Composition of Zinc and Copper in coins.

Objectives:
1. Determine densities and use this information to find out the percentage of copper and zinc in coins.
2. To become familiar with Excel to tabulate, calculate, analyze, and graph scientific data
3. To evaluate the uncertainty (error) in scientific measurements and understand the causes of the underlying uncertainty.

Theoretical Information:
Some international coins are made of metallic copper plated onto a zinc core. You are going to determine the density of pure copper and zinc metals as well as the density of the coin. From this data you will estimate the percentages of copper (Cu) and zinc (Zn) in the coin.
In an object composed of multiple materials (like Britch pence), the density is a weighted average of the densities of the pure substances that make up the object.
If the density of pure copper is d_{Cu} and the density of pure zinc is d_{Zn}, then the density (d) of a composite of copper and zinc is:

$$d = \frac{Pd_{Cu} \quad + \quad qd_{Zn}}{100}$$

In this equation, p = the % Cu by mass and q = % Zn by mass.
Since the object contains only copper and zinc: p + q = 100% and therefore q = 100-p.
So, equation 1 may be rewritten as:

$$d= \frac{Pd_{Cu} \quad + \quad (100-P)d_{Zn}}{100}$$

Materials and equipment:
- A specimen (metal piece)
- A graduated cylinder.
- Balance.
- Water.

Procedure:
Part (1): The Density of Cu, Zn, and coins:
1. Zero the balance with an empty beaker on its top.
2. Weigh your metal pieces. Record this mass.

Part (2): The volume of your metal pieces:
1. Pour water into the graduated cylinder. Record the exact volume as V_{before}.
2. Place your metal piece which you used in the first part inside the graduated cylinder. Record the new volume as V_{after}.

Results and calculation:

1. You will now create a excel table for of the metals might look like this:

objects	mass (g)	Volume $_{before}$ (mL)	Volume $_{after}$ (mL)	Volume $_{object}$ (mL)	d (g/mL)
Cu					
Zn					
Coin					
				Average	
				% Error	
				% Precision	

2. You will now create a table that correlates the density of a Cu/Zn composite (coins) to the percent Cu present.

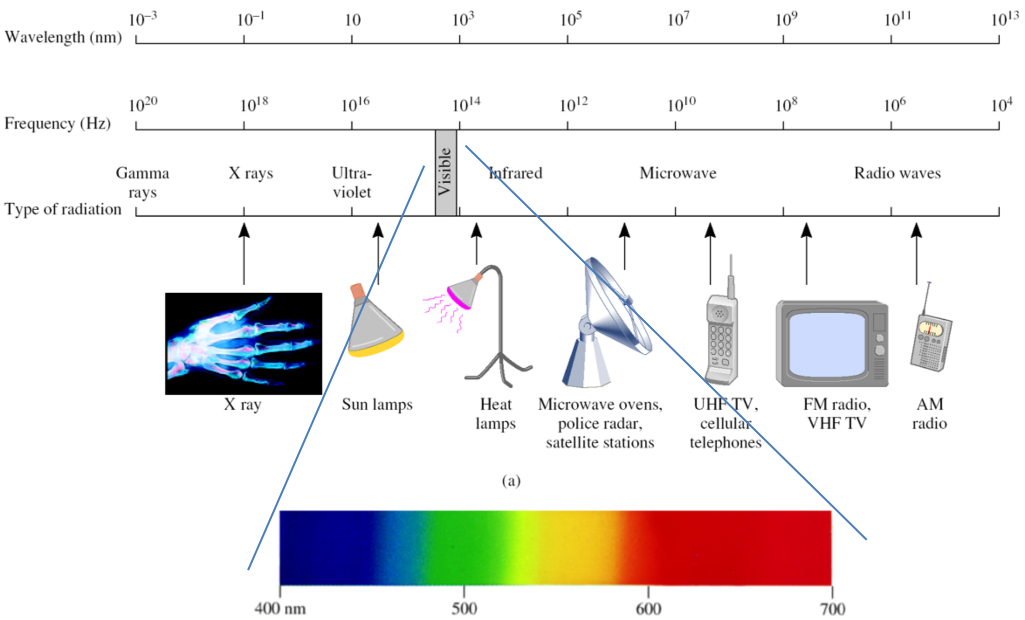
p % Cu	Density of Cu/Zn composite(g/mL)
0	
10	
20	
50	
70	
100	

3. You will now create a graph that correlates the density of a Cu/Zn composite (coins) to the percent Cu present.
   a) Add a linear trend line and place the result on your graph.
   b) Do a LINEST analysis as outlined in your Error Analysis handout.
   c) Using the average coin density that you determined experimentally. and the equation for the line in your graph, determine the % composition of Cu and Zn in the coin.

Experiment (2):
Atomic Emission spectra-Flame Tests Experiment

Objectives:
To determine the spectra of atomic emission and calculate the energy of its wavelength.

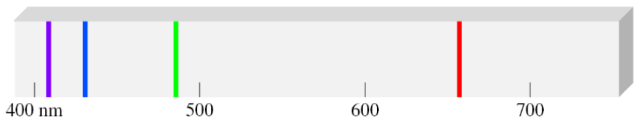
Theoretical Information:
The flame test is a qualitative test used in chemistry to help determine the identity or possible identity of a metal or metalloid ion found in an ionic compound. If the compound is placed in the flame of a gas burner, there may be a characteristic colour given off that is visible to the naked eye (in the visible region).
Visible light is the most familiar example of electromagnetic radiation. Differences in the wavelengths of visible light are manifested as different colours,



Emission
When an atom or ion absorbs energy, its electrons can make transitions from lower energy levels to higher energy levels. The energy absorbed in the form of heat in flame tests, subsequently returns from higher energy levels to lower energy levels, and the energy is released predominantly in the form of *electromagnetic radiation*. In a spectrum of an atom, light of only a certain wavelength is emitted or absorbed, rather than a continuous range of wavelengths.

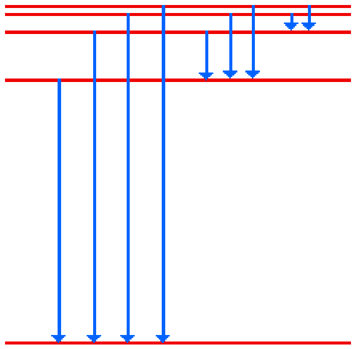
Example:

1. The light emitted by hydrogen atoms is red because, of its four characteristic lines, the most intense line in its spectrum is in the red portion of the visible spectrum, at 656 nm.



2. With sodium, we observe a yellow colour because the most intense lines in its spectrum are in the yellow portion of the spectrum, at about 589 nm.

The light seen from an atom is created from transitions from one energy state to another. The spacing between energy levels in an atom determines the sizes of the transitions that occur, and thus the energy and wavelengths of the collection of photons emitted. Some will consist of Ultra Violet light, so they create no colour, and some will consist of Visible light radiation and show colour.



Colour of light absorbed	Approx. λ ranges / nm	Colour of light transmitted
Red	700-620	Green
Orange	620-580	Blue
Yellow	580-560	Violet
Green	560-490	Red
Blue	490-430	Orange
Violet	430-380	Yellow

Materials and equipment

- Beakers–250 mL, 100 mL.
- wire loop made out of platinum.
- powder or solution to (ionic metal salt)
- flame (blue part).
- Hydrochloric acid (6M)

SAFETY CAUTION:
Always use good safety techniques. Wear chemical splash approved goggles. Wear a chemical apron. Practice the flame test under the supervision of a chemistry teacher.

Procedure

1. Use a clean wire loop made out of platinum.
2. Dip the loop into the powder or solution to be tested (ionic metal salt), and then place it into the hottest portion of a flame (blue part).
3. The resulting colour of the flame is observed and this may be an indication of the presence of a particular ion.
4. To clean the wire: dip the wire into hydrochloric acid. Test the loop by placing it into a gas burner flame. If there is a burst of colour, then you did not clean it sufficiently. If there is no distinct colour, then it is ready for use.

Results and calculation:

- Match the colour to the list below and find the metal.
- Find the wavelength range of the absorbed colour from the table above.
- Calculate the energy of the wavelength.

$E = h\nu.$

$E = \frac{hc}{\lambda}$ h= $6.62607004 \times 10^{-34}$ m^2 kg / s

c= 3×10^8 m / s

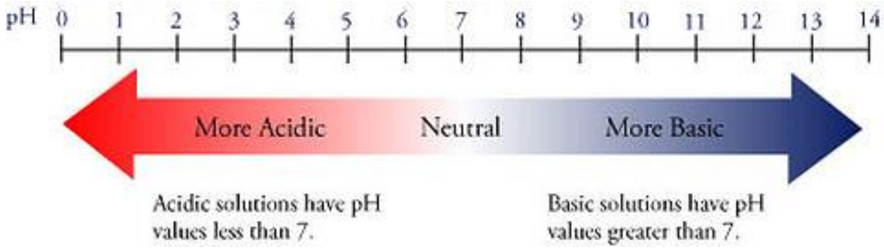
Metal atom	Flame colour
Li	Deep red
Na	Bright yellow
Ca	Brick orange
K	Light violet
Ba	Yellow/green
Sr	Orange/crimson red
Cu	Green/blue

***Note*:** that the flame colour is the emitted (transmitted) colour, not the absorbed one.

Experiment (3):
Measuring the pH of a Solution

Objectives:
1. Determination of the pH values of solution using a pH meter
2. Determination the concertation of strong acid and weak acid.

Theoretical Information:
PH is the quantitative measure of the acidity or basicity of aqueous or other liquid solutions. The pH is determined by measuring the concentration of the hydrogen ion into numbers between 0 and 14 (which ordinarily ranges between about 1 and 10^{-14})



- In pure water,which is neutral,the $[H^+]$ is 10^{-7} ,which corresponds to a pH of 7.
- In a solution with a pH less than 7 is considered acidic.
- In a solution with a pH greater than 7 is considered basic, or alkaline.

pH is calculated as the negative log of the H^+ concentration ($pH = -\log[H^+]$).
pH meter is an electronic instrument used to measure hydrogen-ion activity (acidity or alkalinity) in solution. Fundamentally, a pH meter consists of a pH-responsive electrode that sensitive to hydrogen ions. The pH-responsive electrode is usually glass and this electrode contacted to an electronic meter that measures and displays the pH reading.

Materials and equipment
- CH_3COOH solution
- HCl solution
- NaOH solution
- pH meter with glass electrode.
- Beaker
- Graduated burette
- Pipette

Procedure
1. Put 25 ml of a strong acid solution (HCl solution) in 150ml-beaker.
2. Measure using pH meter by Place the electrode into the solution and record the pH shown on the meter. (Note: Never leave the electrode standing out of a solution)
3. Add some amount of basic solution (NaOH solution) to the beaker and measure pH of the total solution again.
4. Repeat the previous step until you notice that there is no change in pH value.
5. After you finish wash the electrode with distilled water and repeat all the previous steps with a weak acid solution (CH_3COOH solution) instead of strong acid solution.

6. Show your values as curve then make the requirement calculations.

Results and calculation:
<u>Part One: Using HCl Solution.</u>

V_{NAOH}	PH
0	
5	
10	
15	
20	
22	
23	
24	
26	
27	
30	
35	
40	

- Draw a curve between pH values and the volume added from NaOH and then from the curve determine :
  - ✓ The volume of NaOH at equivalent point …………………………......
  - ✓ pH value at equivalent point ………………………………………

- Calculate the Molarity of HCl, Normality, Molecular weight of HCl and Strength of solution.

Part two: Using CH_3COOH Solution.

V_{NAOH}	PH
0	
5	
10	
20	
25	
28	
29	
30	
31	
32	
35	
40	
45	

- Draw a curve between pH values and the volume added from NaOH and then from the curve determine :
  - ✓ The volume of NaOH at equivalent point ………………………….......
  - ✓ pH value at equivalent point …………………………………………

- Calculate the Molarity of CH_3COOH, Normality, Molecular weight of CH_3COOH and Strength of solution. ($ka = 1.76 \times 10^{-5}$)

Experiment No (4)
Molecular Structure

Objective:

- To predict the three-dimensional structure of molecules and molecular ions

Introduction

In 1916, G. N. Lewis proposed **the octet rule** in which atoms form bonds by sharing **valence electrons** until each atom of the molecule has the same number of valence electrons (eight, two for helium) as **the nearest noble gas** in the periodic table. The resulting arrangement of atoms formed the **Lewis structure** of the compound. In most cases we can construct a Lewis structure in the following steps.

1. Calculate the total number of valence electrons from all the atoms (remember to add or subtract the number of electrons necessary to give the total charge on the ion).
2. Divide the total number of electrons by two to give the number of electron pairs.
3. Arrange the atoms to show specific connections. When there is a central atom, it is usually the least electronegative element in the compound, but there are many well-known exceptions like (H_2O and NH_3).
4. Distribute the electrons in pairs so that there is one pair of electrons forming a single bond between each pair of atoms bonded together. Beginning with the terminal atoms to supply the remaining electron pairs to form lone pairs until each atom has an octet. If the central atom has fewer electrons than an octet, use lone pairs from terminal atoms to form multiple (double or triple) bonds to the central atom to achieve an octet.
5. Then assign formal charges.
   -For each atom, count the electrons in lone pairs and half the electrons it shares with other atoms.
   -Subtract that from the number of valence electrons for that atom: the difference is its formal charge.

Note: The Lewis octet rule is not always obeyed. There are some exceptions when the atoms surrounded themselves in molecules by more or less than eight electrons. Resonance between Lewis structures also can be existed to provide great stabilization.

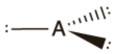
Although, a Lewis structure accounts for the bonding based on the valence electrons, it does not predict the three-dimensional structure for a molecule. The development of the **V**alence **S**hell **E**lectron **P**air **R**epulsion (**VSEPR**) theory provides insight into the three-dimensional structure of the molecule. VSEPR theory proposes that the three-dimensional (3-D) structure of a molecule is determined by the repulsive interaction of electron pairs in the valence shell of its central atom. The three-dimensional orientation is such that the distance between the electron pairs is maximized so that the electron pair interactions are minimized.

Note: We can refer to the electron pairs as electron domains.

In a double or triple bond, all electrons shared between those two atoms are on the same side of the central atom; therefore, they count as one electron domain.

Table 1 represents a summary of the VSEPR theory for predicting the geometrical shape and approximate bond angles of a molecule or ion.

Figure 1 VSEPR and geometric shapes of molecules and molecular ions

Valence Shell Electron Pairs	Bonding Electron Pairs	Nonbonding Electron Pairs	VSEPR Formula	Three-Dimensional Structure	Bond Angle	Geometric Shape
2	2	0	AX_2		180°	Linear
3	3	0	AX_3		120°	Planar triangular
	2	1	AX_2E		<120°	V-shaped
4	4	0	AX_4		109.5°	Tetrahedral
	3	1	AX_3E		<109.5°	Trigonal pyramidal
	2	2	AX_2E_2		<109.5°	Bent
5	5	0	AX_5		90°/120°	Trigonal bipyramidal
	4	1	AX_4E		>90°	Irregular tetrahedral
	3	2	AX_3E_2		<90°	T-shaped
	2	3	AX_2E_3		180°	Linear
6	6	0	AX_6		90°	Octahedral
	5	1	AX_5E		>90°	Square pyramidal
	4	2	AX_4E_2		90°	Square planar

(**A**) refers to the central atom, (**X_m**) refers to m number of bonding pairs of electrons on A, (**E_n**) refers to n number of nonbonding pairs of electrons on A.

Dry Lab Procedure

Several of simple molecules and molecular ions will be assigned, and their three-dimensional structure and approximate bond angles will be determined. The Lewis structure and the VSEPR adaptation of the Lewis structure are used for analysis. A web application (https://molview.org/) is also used to draw a few number of molecules structures and analyzing their 3D structure.

Exercises:

Here are selections of suggested molecules and molecular ions for which their three-dimensional structures (geometric shapes) and approximate bond angles are to be determined. Consider **VSEPR** formulas and their related geometrical shapes as represented in Table 1, while completing the following tables.

A) Complete the following table for many molecules and ions, all of them obey the Lewis octet rule.

Molecular or Ions	Lewis structure	Valence shell e pairs	Bonding electron pairs	Nonbonding electron pairs	VSEPR formula	3-D structure	Approx. bond angle	Geometric shape
CH_4								
NH_3								
NH_4^+								
PF_3								
PF_4^-								
H_2O								
PO_4^{-3}								

B) Complete the following table (as the previous one) for many molecules and ions, none of them obey the Lewis octet rule.

Molecular or Ions	Lewis structure	Valence shell e pairs	Bonding electron pairs	Nonbonding electron pairs	VSEPR formula	3-D structure	Approx. bond angle	Geometric shape
SF_6								
XeF_4								
PCl_2F_3								
BF_3								

C) Draw the following molecules using MOLVIEW website (https://molview.org/) and investigate their shapes and properties like bond angle, bond length and compare the results by which you are expected above.

a. CH_4 b. H_2O c. NH_3 d. SF_6 e. SO_3 f. $SO3^{-2}$

D) Arrange the molecules (a, b, c) in the increasing order of bond angle and discuss the reason behind this order.

Experiment (5):
Chemical equilibrium
(IONIC EQUILIBRIUM IN SOLUTION)

Objectives:

1. Study of shift in equilibrium in the reaction of ferric ions and thiocyanate ions by increasing the concentration of any one of these ions.
2. Study the effect of temperature changes (hot and cold) on the aqueous equilibrium.

Theoretical Information:

The state of equilibrium in any reaction, is recognized by the constancy of an observable property (macroscopic property), like the color intensity of the solution. In this unit we will study the shift in equilibrium in various reactions.

The equilibrium reaction between ferric chloride and potassium thiocyanate is conveniently studied through the change in the intensity of the color of the solution.

$$Fe^{3+}{}_{(aq)} + SCN^{-}{}_{(aq)} \rightarrow [Fe(SCN)]^{2+}{}_{(aq)}$$

(pale yellow) + (colorless)→ (Blood red colour)

The equilibrium constant for the above reaction may be written as:

$$K = \frac{\left[[Fe(SCN)]^{2+}{}_{(aq)}\right]}{\left[Fe^{3+}{}_{(aq)}\right]\left[SCN^{-}{}_{(aq)}\right]}$$

Here K is constant at a constant temperature. Increasing the concentration of either Fe^{3+} ion or thiocyanate ion would result in a corresponding increase in the concentration of $[Fe(SCN)]^{2+}$ ions. In order to keep the value of K constant, there is a shift in equilibrium, in the forward direction and consequently an increase in the intensity of the blood red color which is due to $[Fe(SCN)]^{2+}$., color intensity remains constant at equilibrium.

Materials and equipment

- Ferric chloride (0.1M)
- Potassium thiocyanate (0.1M)
- Potassium chloride (0.1M)
- Beakers
- Boiling tubes
- graduated cylinder
- Test tube stand
- Glass rod
- Ice water
- A water bath

Procedure

Part A: Concentration Changes

1. Mix 10 mL of ferric chloride solution with 10 mL of potassium thiocyanate solution. Blood red color will be obtained.
2. Add 150 mL of water to the beaker by a graduated cylinder.
3. Take four boiling tubes of the same size and mark them as a,b,c and d .
4. Add 10 mL of blood red solution to each of the boiling tubes from the graduated cylinder.
5. Add 5 mL of water to the boiling tube ‘a’ so that the total volume of the solution in the boiling tube ‘a’ is 15 mL. Keep it for reference.
6. Add 5 mL of ferric chloride solution to the boiling tube ‘b’ so that the total volume of the solution in the boiling tube ‘b’ is 15 mL.
7. Add 5 mL of potassium thiocynate solution to the boiling tube ‘c’ so that the total volume of the solution in the boiling tube ‘c’ is 15 mL.
8. Add 5 mL of potassium chloride solution to the boiling tube ‘d’ so that the total volume of the solution in the boiling tube ‘d’ is 15 mL.
9. Compare the color intensity of the solution in each boiling tube with the color intensity of the reference solution in boiling tube ‘a’.
10. Record your results in a tabular form as in Tables 1 .

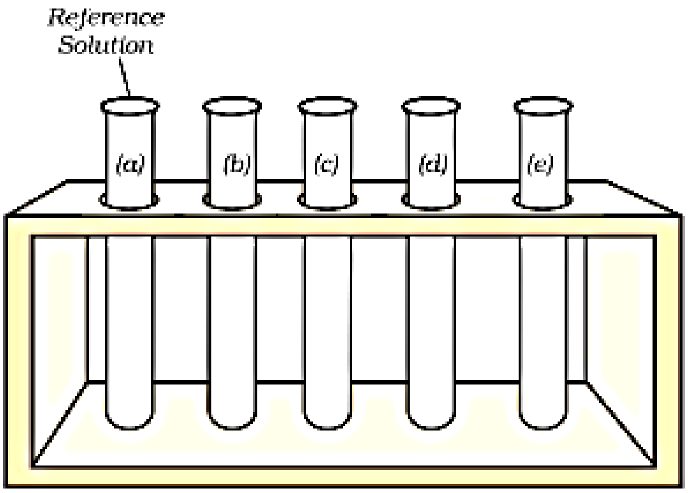


Fig. 1 : Set up of the experiment for observing equilibrium, each boiling tube contains 15 mL solution

Note: The color intensity of the solution will decrease very much on dilution. It will not be deep blood red color.
total volume in each test tube is 15 ml, each test tube has 10 mL equilibrium mixture.

Part B: Changes in Temperature

1. Take three boiling tubes of the same size and mark them as 1,2, and 3
2. Add 10 mL of blood red solution to each of the boiling tubes from the burette.
3. Set one tube aside as a color standard against which to judge color changes in the other tubes.
4. Warm the second tube in a hot water bath on a hot plate. Do not boil the solution. Observe.
5. Cool the third tube in a beaker of ice water. Observe.

Results and calculation:

Part A: Concentration Changes

- **Table (1):** Equilibrium shift on increasing the concentration:

Boiling tube	Substance added at equilibrium	Change in the color intensity as matched with the reference solution in boiling tube "a"	Effect on the concentration of $[Fe(SCN)(H_2O)_5]^{2+}$	Direction of shift in equilibrium
A	5 ml of water			
B	5 ml of 0.1M $FeCl_3$ solution			
C	5 ml of 0.1M KSCN solution			
D	5 ml of 0.1M KCl solution			

Part B: Changes in Temperature

- **Table (2): The effect of temperature changes (hot and cold) on the aqueous equilibrium after placing the test tubes in hot water and ice water**

Boiling tubes	Water temperature	Change in the color intensity as matched with the reference solution in boiling tube "a"	Direction of shift in equilibrium
1			
2			
3			

What is Le Chatelier's principle?

What are the factors affecting the Position of Equilibrium?

Experiment (6):
Chemical equilibrium
(Determination of the equilibrium constant for formation of important ethyl acetate)

Objectives:
To determine the equilibrium constant.

Theoretical Information:

The equilibrium constant of a chemical reaction is the value of its reaction quotient at chemical equilibrium, a state approached by a dynamic chemical system after sufficient time has elapsed at which its composition has no measurable tendency towards further change.

Ethanol reacts with acetic acid to give the ester, ethyl acetate:

$$CH_3CH_2OH_{(aq)} + CH_3COOH_{(aq)} \leftrightarrow CH_3COOCH_2CH_{3\,(aq)} + H_2O_{(1)}$$

The equilibrium constant for the above reaction is defined as,

$$Kc = \frac{[CH_3COOCH_2CH_{3(aq)}]}{[CH_3CH_2OH_{(aq)}][CH_3COOH_{(aq)}]}$$

Balance Equation at Titration:

$$NaOH_{(aq)} + CH_3COOH_{(aq)} \leftrightarrow CH_3COONa_{(aq)} + H_2O_{(1)}$$

Materials and equipment
- Ethanol solution (1M).
- acetic acid solution (1M).
- NaOH solution (0.5 M).
- 250 cm^3 volumetric flask.
- Standard glassware for calibration.

Procedure
1. Set up the 50 cm^3 burette containing 0.5 M NaOH solution.
2. Pipette 10.0 cm^3 of the sample solution into a conical flask.
3. Add two drops of phenolphthalein indicator into the conical flask.
4. Take the initial burette reading. Titrate the sample solution against the sodium hydroxide solution. Swirl continuously during the addition of the titrant.
5. Towards the end point, add the NaOH solution dropwise and swirl. Stop the addition when one drop of titrant causes the indicator to change from colorless to pink. The titration should be completed in the shortest possible time.
6. Take the final burette reading and calculate the titre volume.
7. Repeat titrations until the titre volume are within ± 0.20 cm^3 consistency.
8. Repeat the experiment to check the reliability of results.

Results and calculation:

$V_{initial}$ (ml)	V_{final} (ml)	V (ml)	$V_{Average}$ (ml)

Let the [CH_3COOH] at equilibrium that is determined from the titration be x M.

	$CH_3CH_2OH_{(aq)}$	$CH_3COOH_{(aq)}$	$CH_3COOCH_2CH_{3\,(aq)} + H_2O_{(1)}$
Initial conc. (M)			
At equilibrium(M)			

$$Kc = \frac{[CH_3COOCH_2CH_{3(aq)}]}{[CH_3CH_2OH_{(aq)}][CH_3COOH_{(aq)}]}$$

Why isn't the [H_2O] included in the equilibrium expression, Kc = $\boldsymbol{\frac{[CH_3COOCH_2CH_{3(aq)}]}{[CH_3CH_2OH_{(aq)}][CH_3COOH_{(aq)}]}}$***?***

As the ethanoic acid is titrated with the aqueous NaOH, wouldn't the decreasing concentration of the ethanoic acid causes the position of equilibrium to shift left, hence affect the accuracy of the result?

Experiment No (7)

Kinetic study of Sodium Thiosulfate reaction with Hydrochloric Acid

Objective:

To determine the order of reaction (reaction law) with respect to $S_2O_3^{-2}$ and H^+.

Theoretical Information:

The oxidation-reduction reaction that occurs between hydrochloric acid and sodium thiosulfate, produces **insoluble** sulfur as a product.

$$Na_2S_2O_3 + 2HCl \rightarrow S_{(s)} + SO_{2(g)} + H_2O_{(l)}$$

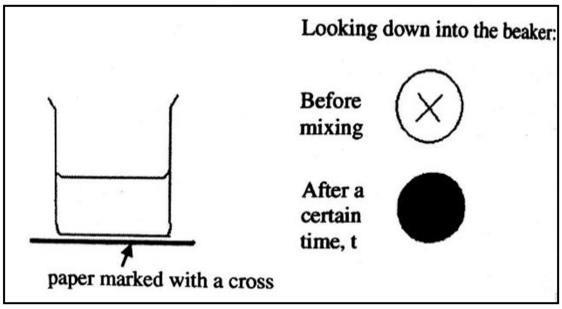
The time required for the **cloudiness** of sulfur to appear is a measure of the reaction rate. Thus, we can measure the rate by measuring the time required for a **fixed** amount of sulfur to cover a piece of paper with printed text. To determine the order of reaction with respect to $S_2O_3^{-2}$ and H^+ we need to vary the concentration of $S_2O_3^{-2}$ while keeping the concentration of H^+ constant. Then we would need to analyze how the rate is affected by changing the concentration of $S_2O_3^{-2}$. Likewise, repeat the whole experiment by changing the concentration of H^+ instead while keeping the concentration of $S_2O_3^{-2}$ constant.

Materials and equipment

- Sodium thiosulfate solution
- HCl solution
- Beaker
- Burette
- Graduated cylinder.
- Stopwatch (electric timer)
- Distilled or deionized water
- paper printed with printed X.

Procedure

1. Use a measuring burette to measure 4 ml of HCl into the beaker.
2. Use another measuring cylinder to measure 5 ml of water into the beaker.
3. Use a different measuring cylinder to measure 20 ml of $Na_2S_2O_3$ into the beaker.
4. Start the stopwatch immediately, swirl the contents and place the beaker over the printed material.
5. Record the time taken t in the table below.
6. Repeat the next set of experiment according to the tables below.
7. Analyze your results graphically (as explained in the next section) and deduce the rate law of the reaction.

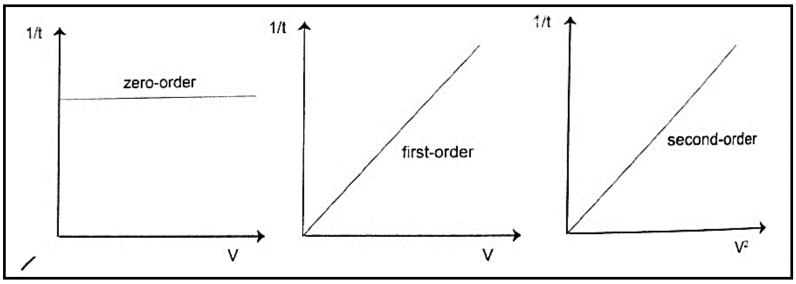


No.of exp	V($Na_2S_2O_3$)	V(H_2O)	V(HCl)	V	V^2	T (sec)	1/t
1	25	0	4				
2	20	5	4				
3	15	10	4				
4	10	15	4				
5	5	20	4				

No.of exp	V($Na_2S_2O_3$)	V(H_2O)	V(HCl)	V	V^2	T (sec)	1/t
1	10	0	5				
2	10	1	4				
3	10	2	3				
4	10	3	2				
5	10	4	1				

Using the graphical approach to treat the results.

Plot a graph of 1/t versus V for both sodium thiosulfate and HCl. Check if the graph resembles zero or first order as their shapes represented in the following Picture and write the reaction rate law.



Note: If the rate equation is, R = K[A]^n then $\frac{1}{t} \propto V^n$ .Thus if n = 0 a graph of 1/t vs V would give a horizontal straight line and if n = 1 a graph of 1/t vs V would give a linear straight line.

Experiment (8):
Oxidation-reduction reactions (Redox reactions).

Objectives:
To determine relative oxidizing and reducing strengths of a series of ions.

Theoretical Information:
The ability of an element to react with another element is called its **activity**. The easier it is for an element to react with another substance, the greater its activity. A more reactive metal displaces a less reactive metal from salt solution. Such reactions are called **displacement reactions.** The displacement reactions involve simultaneous Oxidation and Reduction which is why they are called **redox** reactions.

Oxidation is electron loss - the atom/ion/molecule losing one or more electrons is said to be oxidized.

Reduction is electron gain - the atom/ion/molecule gaining one or more electrons is said to be reduced.

Oxidizing agent is the species that is being reduced (gaining electrons).

Reducing agent is the species that is being oxidized (losing electrons).

For example, magnesium is more reactive than copper. It displaces copper from copper sulfate solution.

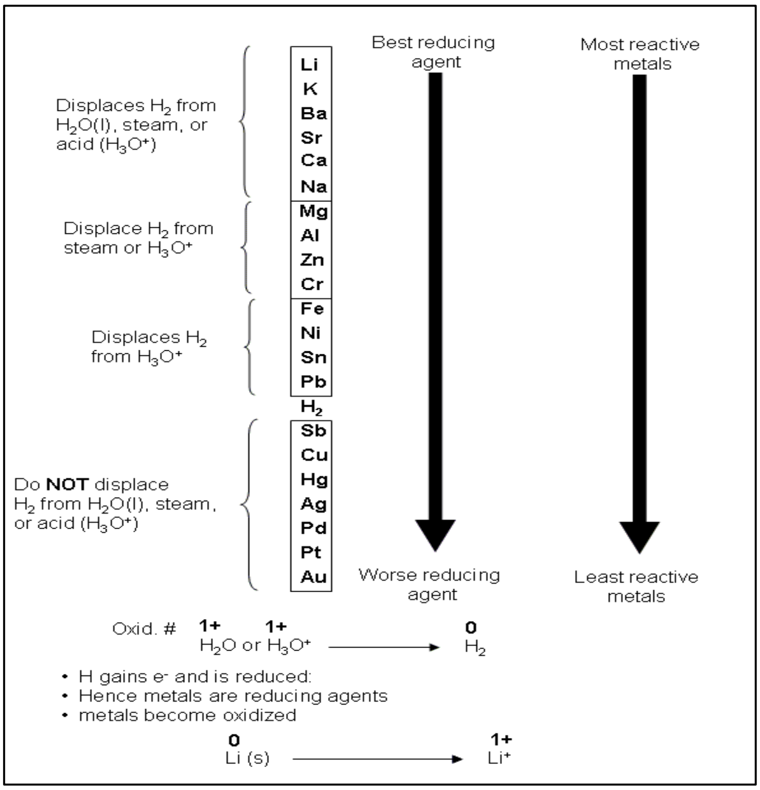
$$Mg_{(s)} + CuSO_{4(aq)} \rightarrow MgSO_{4(aq)} + Cu_{(s)}$$

Magnesium atoms lose electrons - they are oxidized, copper ions gain electrons - they are reduced. Even though the redox reactions happen at the same time this equation can split into two half equations:

$$Mg_{(s)} \rightarrow Mg^{2+}_{(aq)} + 2e^{-} \text{ (oxidation)}$$

$$Cu^{2+}_{(aq)} + 2e^{-} \rightarrow Cu_{(s)} \text{ (reduction)}$$

The **activity series** is a list or table of elements organized by how easily they undergo a reaction Metals higher on the activity series are more likely to reacts relative to those lower on the activity series. The activity series can be used to predict products of a single displacement reaction, and to predict if a reaction will even occur. Activity series of some of the more common metals, listed in decreasing order of reactivity.



Materials and equipment

- Zinc (piece).
- Copper (piece).
- HCl solution. (2M)
- Copper Sulfate solution
- Zinc Sulfate solution
- Silver nitrate solution
- Test tubes and test tube stand.
- Dropper
- Cyclohexane
- Chlorine water
- Iodine water (0.2 M I_2 in dropper bottle)
- Potassium iodide solution (0.5 M)
- Potassium bromide solution (0.5 M)
- Potassium chlorine solution (0.5 M)

Procedure

Part one: Potential Series for metals

A. Activity Series for metals and Hydrogen

1. Place a piece of Cu and Zn metal into two separate, clean test tube, add small amount of hydrochloric acid into each test tube.

B. Silver, Copper, and Zinc activities

1. Place a piece of Cu in a clean test tube, add small amount of zinc Sulfate solution.
2. Place a piece of Zn in a clean test tube, add small amount of copper Sulfate solution.
3. Place a piece of Cu in a clean test tube, add small amount of silver nitrate solution.

Examine each reaction mixture and record your observations on the Report Sheet. If you conclude from your observations that a reaction has occurred, write its net ionic equation. If no reaction occurs, do not write an equation, write N.R.

Part two: Potential Series for Halogens

A. Potassium Iodide and Chlorine

1. Put 2 mL of 0.5 M aqueous potassium iodide solution in a test tube and add 1 mL of cyclohexane. Mix well.
2. To the test tube with aqueous potassium iodide/cyclohexane add 1 ml of Cl_2 water. Mix well and record the color of the cyclohexane layer.

Now you will be able to identify any of these three halogens in the cyclohexane layer.

B. Potassium Bromide and Chlorine

1. Put 2 mL of 0.5 M aqueous potassium bromide solution in a test tube and add 1 mL of cyclohexane. Mix well. Observe the colors of the two layers and record them on your report sheet.
2. To the test tube with aqueous potassium bromide/cyclohexane add 1 mL of Cl_2 water. Mix well and record the color of the cyclohexane layer.

Now you will be able to identify any of these three halogens in the cyclohexane layer If you conclude from your observations that a reaction has occurred.

C. Potassium Bromide and Iodine

1. Put 2 mL of 0.5 M aqueous potassium bromide solution in a test tube and add 1 mL of cyclohexane. Mix well.
2. To the test tube with of aqueous potassium bromide/cyclohexane add 1 mL of I_2 water. Mix well and record the color of the cyclohexane layer.

Now you will be able to identify any of these three halogens in the cyclohexane layer If you conclude from your observations that a reaction has occurred.

D. Potassium Chlorine and Iodine

1. Put 2 mL of 0.5 M aqueous potassium chlorine solution in a test tube and add 1 mL of cyclohexane. Mix well.
2. To the test tube with of aqueous potassium bromide/cyclohexane add 1 mL of I_2 water. Mix well and record the color of the cyclohexane layer.

Now you will be able to identify any of these three halogens in the cyclohexane layer If you conclude from your observations that a reaction has occurred.

SAFETY CAUTION:
cyclohexane: extremely flammable liquid and vapor. vapor may cause flash fire. harmful or fatal if swallowed. harmful if inhaled. causes irritation to skin, eyes, and respiratory tract.
chlorine water: corrosive. causes eye and skin burns. causes digestive and respiratory tract burns.
iodine water: poison! causes severe irritation or burns to every area of contract. may be fatal if swallowed or inhaled. vapors cause severe irritation to skin, eyes, and respiratory tract. oxidizer. may cause allergic skin or respiratory reaction.
potassium bromide: harmful if swallowed or inhaled. may cause irritation to skin, eyes, and respiratory tract.
potassium iodide: may cause irritation to skin, eyes, and respiratory tract.

Results and calculation:

Part one: Potential Series for metals

1. **Reactions of Hydrogen with Copper, and Zinc**

Colors of Halogen	Observations	Net ionic equation
Copper + hydrochloric acid		
Zinc + hydrochloric acid		

2. **Reactions of Copper, silver, and Zinc**

Colors of Halogen	Observations	Net ionic equation
Copper + zinc sulfate		
Zinc + copper sulfate		

Q1: Relative oxidizing strengths, which is the stronger oxidizing agent Cu^{2+} or Zn^{2+}?

Colors of Halogen	Observations	Net ionic equation
Copper + silver nitrate		

Q2: Arrange Cu^{2+}, Zn^{2+}, Ag^{+}, and H^{+} in the decreasing order of reactivity.

<u>Part two: Potential Series for Halogens</u>

part	Colors of Halogen	Color in Cyclohexane	Observations	Net ionic equation
a				
b				
c				
d				

What is the correct order of oxidizing strength of (Cl_2, I_2)?

You don't test (F_2) in this experiment, but according to the order you observed for the other halogens, expect the oxidizing strength of F_2 and explain your answer?