CONTENT ENRICHMENT TRAINING PROGRAMME IN CHEMISTRY FOR JUNIOR COLLEGE LECTURERS OF APSWREI SOCIETY, HYDERABAD

(8.6.1998 to 19.6.1998)

Regional Institute of Education, Mysore 570 006
[National Council of Educational Research and Training, New Delhi]

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RIE FACULTY

Dr M S Srimathi (Coordinator)
Dr V Kesavan
Dr G T Bhandage
Dr B S Raghavendra
Dr G R Prakash
Dr R Narayanan
Dr V D Bhat
Shri D N Nagaraj

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Introduction

With the ever increasing knowledge and understanding in some of the frontier areas of science, it is imperative that a science teacher to keep in pace with these new developments. These modern advancements have to be reflected in his classroom teachings to update the students to the new concepts in the concerned areas of science. In recent years, the working teacher has been provided many opportunities to attend inservice programmes. These courses are meant mainly to update his knowledge in specific subject areas, to provide laboratory experiences and to give him exposure to the new methods of teaching.

In view of the above, the APSWREIS, Hyderabad, a leading educational organization Andhra Pradesh requested the Regional Institute of Education (RIE), Mysore, a constituent unit of NCERT, New Delhi to take up an inservice content enrichment programme for Junior College Lecturers teaching chemistry at the Intermediate level. Accordingly, the programme was planned at RIE, Mysore in June 1998.

The programme was held from 8.6.1998 to 19.6.1998 at Regional Institute of Education, Mysore. Prof S N Prasad, Principal, RIE, Mysore inaugurated the programme. In the subsequent session, the tentative schedule was presented and suggestions were invited from the participants. As per their requirements, the topics/units/laboratory work to be given greater emphasis were listed out. The programme was conducted according to the Schedule enclosed herewith (Annexure I).

A group of twenty eight Junior College Lecturers in Chemistry selected from various districts of Andhra Pradesh and deputed by APSWREIS attended the two week programme (Annexure II). The resource persons were drawn from the chemistry faculty of RIE, Mysore. In addition to them, two local resource persons were also associated. The details are furnished in Annexure III.

About the Training

There were lecture-discussion sessions in the forenoon on selected topics from the Intermediate syllabus of Andhra Pradesh. This included Atomic Structure, Chemical Bonding, Thermodynamics, Nomenclature of organic compounds, chemical equilibrium, periodic table, mechanisms of organic reactions, gaseous state, chemistry of s, p and d-block elements and kinetics of chemical reactions. There was a session on science education where education technology and psychological aspects were discussed.

In each session, the participants were encouraged to ask questions in the free environment. Overhead projector and slides were used wherever felt necessary. There were lively discussions and effective interaction. The lectures were specially designed with an inbuilt methodology and the possible difficulties arising in teaching the specific topics were also discussed.

Afternoon sessions were mainly devoted for laboratory work, wherein individual hands-on experiences were provided. These included experiments related to stoichiometry, preparation and properties of gases, thermochemistry, chemical equilibrium, chemical kinetics and electrochemistry. Some of the text book activities were also demonstrated to enable the participants to include them in their routine work. The most important aspect of the lab session was the post-lab discussions. The detailed analysis of the results obtained in the laboratory and its correlation with the theoretical aspects was done.

There were two demonstration-discussion sessions. One session dealt with the cathode ray production in a discharge tube and study of its properties. Study of Atomic Spectra of selected elements was an experience by itself. A Computer Laboratory session was arranged wherein the use of computers as a means of conveying some of the abstract aspects and upto dating the knowledge was highlighted. The ability of computer in information collection and processing was illustrated using Internet facility. Chemistry packages including Encyclopedia Britannica available on CD-ROM were also displayed.

A provision was made in the programme schedule for two problem solving sessions. In these, the participants were provided with a set of numerical problems on chemistry topics. They were encouraged to solve them individually and discuss at length with others.

Some interesting videotapes on Hydrogen spectra, Hydrogen atom – as viewed by Quantum Mechanics were screened and post-film discussion was conducted. A visit to the RIE Library was arranged on participants' request. This enabled them to get acquainted with the various journals, periodicals and textbooks on chemistry and teaching of chemistry.

The valedictory function was held on 19.6.1998 afternoon. Dr V Kesavan, Head, Dept of Chemistry gave away the certificates to the participants. Reference materials on few chapters (Annexure IV) and an Inservice Training Package on "Structure, Bonding and Reactivity" was given to the participants.

Annexure I

Training Schedule

(8.6.1998 to 19.6.1998)

Date	9.30 am to 11 am	11.15 am to 12.45 pm	2.00 pm to 5.00 pm
8.6.98	Registration and Inauguration	Discussion with participants (Dr VK, Dr GRP, Dr MSS)	Mole Concept and Stoichiometry (Dr VK, Dr GRP)
9.6.98	Chemical Bonding – I (Dr BSR)	Atomic Structure – I (Dr MSS)	Problem Solving – I (Dr GTB, Dr MSS)
10.6.98	Atomic Structure – II (Dr MSS)	Chemical Bonding – II (Dr BSR)	Inorganic Chemistry (Dr BSR, Dr GRP)
11.6.98	Thermodynamics – I (Dr GTB)	Organic Chemistry – I (Dr VK)	Thermochemical Measurements (Dr GTB, DR MSS)
12.6.98	Thermodynamics – II (Dr GTB)	Periodic Table (Dr GRP)	Problem Solving – II and Chem Study Film (Dr BSR, Dr MSS)
13.6.98	Chemical Equilibrium – I (Dr ASJ)	d-Block Elements/Metallurgy (Dr BSR)	Organic Chemistry (Dr VK, Dr BSR)
14.6.98	Field Work		
15.6.98	Chemical Equilibrium – II (Dr GTB)	s & p Block elements - I (Dr GRP)	Study of a system at equilibrium (Dr GTB, Dr KCS)
16.6.98	Organic Chemistry – II (Dr VK)	s & p Block Elements – II (Dr BSR)	Demonstration – Discharge tube experiments, atomic spectra (Dr RN)
17.6.98	Organic Chemistry – III (Dr VK)	Kinetics – I (Dr GRP)	Electrochemistry (Dr ASJ, Dr VK)
18.6.98	Science Education (Dr VDB)	Applications of Computers in Chemistry (Dr GRP, Mr DNN)	Kinetics of a Chemical Reaction (Dr ASJ, Dr KCS)
19.6.98	Gaseous State (Dr GTB)	Nuclear Chemistry / Library Work (Dr MSS)	Feedback from participants and valedictory

(M S Srimathi)
Academic Coordinator

Annexure – II

List of Participants

1. G Krishna Murthy
   Junior Lecturer in Chemistry
   A P S W R School & Junior College
   Naidupet
   Nellore Dist

2. V Hema Bhushan
   Junior Lecturer in Chemistry
   A P S W R Junior College / School
   Karempudi
   Guntur

3. N V S Lakshmi
   Junior Lecturer in Chemistry
   A P S W R Junior College
   Cheepurupalli
   Vizianagaram Dist

4. S Vidya Rani
   Junior Lecturer in Chemistry
   A P S W R Junior College
   Parkal
   Warangal

5. S Suryanarayana Reddy
   Junior Lecturer in Chemistry
   A P S W R Junior College
   Kalasamudram
   Anantapur

6. B Manohar Rao
   Junior Lecturer in Chemistry
   A P S W R Junior College
   Maluguru
   Anantapur

7. K Ramesh Babu
   Junior Lecturer in Chemistry
   A P S W R College, Atchampet
   Guntur Dist., AP

8. B Srinivasa Rao
   Junior Lecturer in Chemistry
   A P S W R Junior College
   Arugolanu
   West Godavari Dist

9. K Srinivasa Rao
Junior Lecturer in Chemistry
Dr B R A C S W Junior College
L N Puram
East Godavari Dist

10. P V S N Seetharamam
Junior Lecturer in Chemistry
A P S W R Junior College
Tiruvur 521 235

11. D Ansari Begum
Junior Lecturer in Chemistry
Dr B R A C A P S W Residential Junior College
Kurugunta 515 001

12. D Srinivasachari
Junior Lecturer in Chemistry
A P S W R Junior College
Kopperla
Vizianagaram Dist

13 K V Satyavathi
Junior Lecturer in Chemistry
A P S W R Junior College
S M Nagar, Kakinada
East Godavari Dist

14. V Laxmanjali Devi
Junior Lecturer in Chemistry
A P S W R Junior College
Adilabad

15. P Venkata Chalam
Junior Lecturer in Chemistry
A P S W R S Junior College
J P Nagar
Mahabub Nagar Dist 509 324

16 Ch Narasimha Reddy
Junior Lecturer in Chemistry
A P S W R J College
Shaikpet
Hyderabad

17. K Jayalakshmi
Junior Lecturer in Chemistry
A P S W R J College
Kota, Nellore Dist

18. K Bharathi
Junior Lecturer in Chemistry
A P S W R Junior College
Chimakurthy
Prakasam Dist

19. P Srinivasa Rao
Junior Lecturer in Chemistry
A P S W R Junior College
A R Pally
Khammam Dist

20. S N Mohan Reddy
Junior Lecturer in Chemistry
A P S W R Junior College
Armoor
Nizamabad Dist

21. M Rajendra Chary
Junior Lecturer in Chemistry
A P S W R Junior College
Jangaon
Warangal Dist

22. H Cheralu
Junior Lecturer in Chemistry
A P S W R Junior College
Ghanphur
Warangal Dist

23. T Hanumantha Rao
Junior Lecturer in Chemistry
A P S W R Junior College
Velugonda, Prakasam Dist

24. B Rupini
Junior Lecturer in Chemistry
A P S W R Junior College
Hanamkonda
Warangal Dist

25. V Udaya Sree
Junior Lecturer in Chemistry
A P S W R Junior College
Suryarpet
Nalgonda

26. T R J Vinutha
A P S W R Junior College
Chinnachowk, Cuddapah

7. G Grace
A P S W R Junior College
Polasanipalli
West Godavari Dist

8. C Santhi Visala
A P S W R Junior College
Ameenapet, Eluru, A P

9. E John Kennedy
Junior Lecturer in Chemistry
A P S R J Junior College
Rudravaram 521 001
Krishna Dist

Annexure III

LIST OF RESOURCE PERSONS

1. Dr V Kesavan
2. Dr G T Bhandage
3. Dr B S Raghavendra
4. Dr G R Prakash
5. Dr R Narayanan
6. Dr V D Bhat
7. Sri D N Nagaraj

LOCAL RESOURCE PERSONS

1. Dr A S Janardhan
2. Dr K C Srinivasa Murthy

COORDINATOR

Dr M S Srimathi, Department of Chemistry

Annexure IV

Observations and Suggestions for monitoring the training programmes :

Based on the experience gained in the present training programme, it is felt that the following points need to be considered in conducting such programmes in future and monitoring of the training.

1. The duration of the programme should be at least two weeks. Chemistry, being an experimental science and laboratory work requiring equal emphasis as theory, the number of participants should not exceed thirty.
2. The venue of the training programme should have facilities like good library, well equipped laboratory for conducting experiments at the Intermediate level. In addition, facilities for screening video films and Over Head Projector(OHP) are desirable.
3. Participating teachers should be contacted well in advance to find out their requirements as to the topics to be discussed or laboratory work with specific details.
4. Resource persons may be requested to give a write-up about the topics chosen for discussion. The write-up should include major concepts, inbuilt methodology for teaching the above and relevant evaluation items.
5. During the training period, a special emphasis may be given for solving a number of numerical problems under different topics.
6. A few sessions may be devoted exclusively for the participants to present their views on some selected topics and teaching methodology.
7. Lecture-cum-discussion sessions on psychology and science education will be useful for the participants.
8. The programme should help the participating teacher to train his/her students for competitive examinations.

In the chemistry group, the following five teachers have been identified who could be used by the APSWREIS for further training programmes.

1. Sri G Krishna Murthy
   Junior Lecturer in Chemistry
   APSWR School and Junior College
   Naidupet
   Nellore Dist

2. N V S Lakshmi
   Junior Lecturer in Chemistry
   APSWR Junior College
   Cheepurupalli
   Vizianagaram Dist

3. S N Mohan Reddy
   Junior Lecturer in Chemistry
   APSWR Junior College
   Armoov
   Nizamabad Dist

4. V Udaya Sree
   Junior Lecturer in Chemistry
   APSWR Junior College
   Suryarpet
   Nalgonda

5. T R J Vinutha
   Junior Lecturer in Chemistry
   APSWR Junior College
   Suryarpet
   Nalgonda

Annexure V

Resource/Instructional Materials given to the participants :

1. Periodic Table of elements supplied by Sargent-Welch Scientific Company, Illinois, 60089.

2. Table of Periodic Properties of Elements supplied by Sargent-Welch, Illinois, 60089

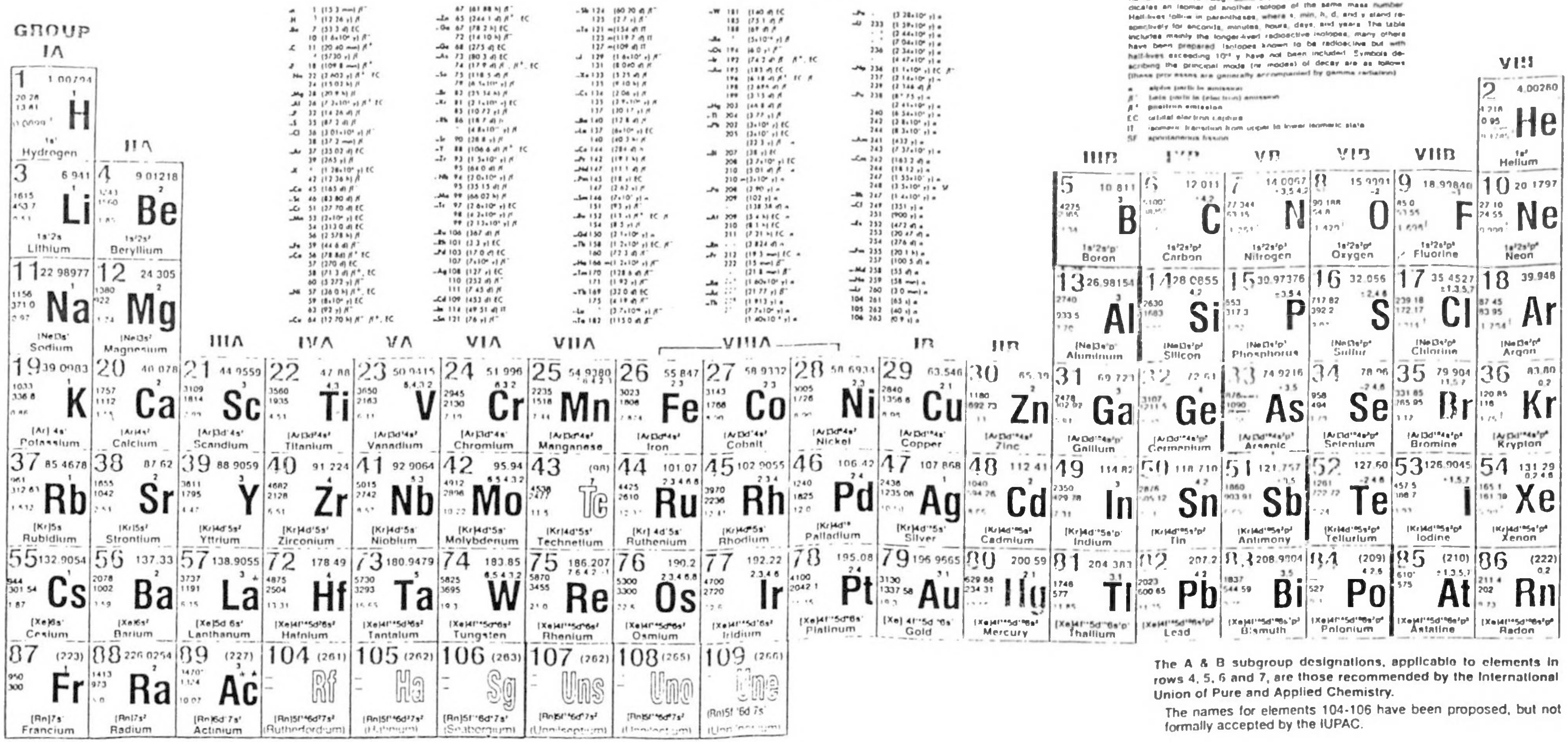
3. Pictorial representation of Electrons in Atoms – Shapes of atomic orbitals.

4. Structure, Bonding and Reactivity – Inservice Training Package Material, Chemistry Section, RIE, Mysore.

5. Laboratory procedures for conducting selected experiments in chemical equilibrium, potentiometry, thermochemical measurements, chemical kinetics.

PERIODIC TABLE OF THE ELEMENTS

Table of Selected Radioactive Isotopes



Selected Radioactive Isotopes

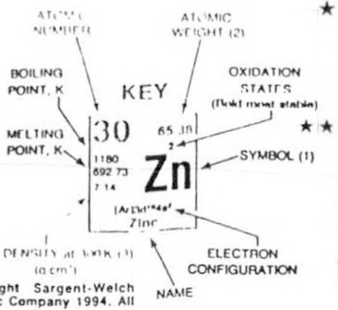
Naturally occurring radioactive isotopes are designated by a mass number in blue (although some are also manufactured). Letter m indicates an isomer of another isotope of the same mass number. Half-lives follow in parentheses, where s, min, h, d, and y stand respectively for seconds, minutes, hours, days, and years. The table includes mainly the longer-lived radioactive isotopes; many others have been prepared. Isotopes known to be radioactive but with half-lives exceeding 10^{12} y have not been included. Symbols describing the principal mode (or modes) of decay are as follows (these processes are generally accompanied by gamma radiation):

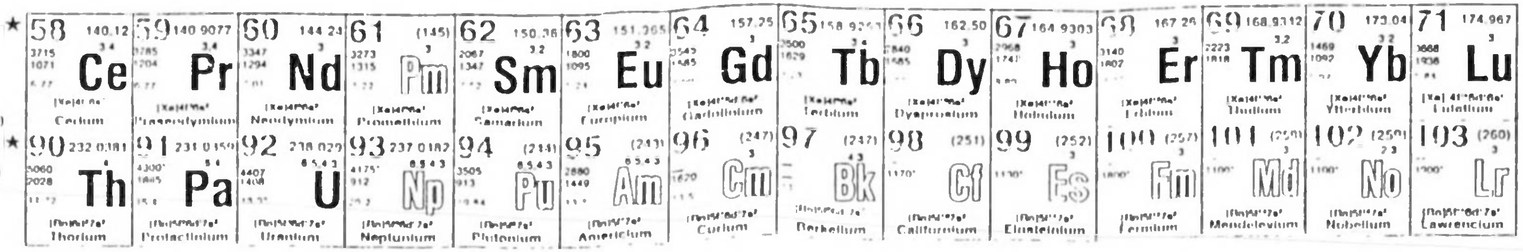
- α alpha particle emission
- β^- beta particle (electron) emission
- β^+ positron emission
- EC orbital electron capture
- IT isomeric transition from upper to lower isomeric state
- SF spontaneous fission

* Estimated Values

The A & B subgroup designations, applicable to elements in rows 4, 5, 6 and 7, are those recommended by the International Union of Pure and Applied Chemistry.

The names for elements 104-106 have been proposed, but not formally accepted by the IUPAC.





NOTES:

(1) Black — solid
Red — gas
Blue — liquid
Outline — synthetically prepared

(2) Based upon carbon-12. ( ) indicates most stable or best known isotope.

(3) Entries marked with daggers refer to the gaseous state at 273 K and 1 atm and are given in units of g/l.

Catalog Number S-18806

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VWR SCIENTIFIC
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Side 1

TABLE OF PERIODIC PROPERTIES OF THE ELEMENTS

Percent Ionic Character of a Single Chemical Bond

Difference in electronegativity	0.1	0.2	0.3	0.4	0.5	0.6	0.7	0.8	0.9	1.0	1.1	1.2	1.3	1.4	1.5	1.6	1.7	1.8	1.9	2.0	2.1	2.2	2.3	2.4	2.5	2.6	2.7	2.8	2.9	3.0	3.1	3.2
Percent ionic character %	0.5	1	2	4	6	9	12	15	19	22	26	30	34	39	43	47	51	55	59	63	67	70	74	76	79	82	84	86	88	89	91	92

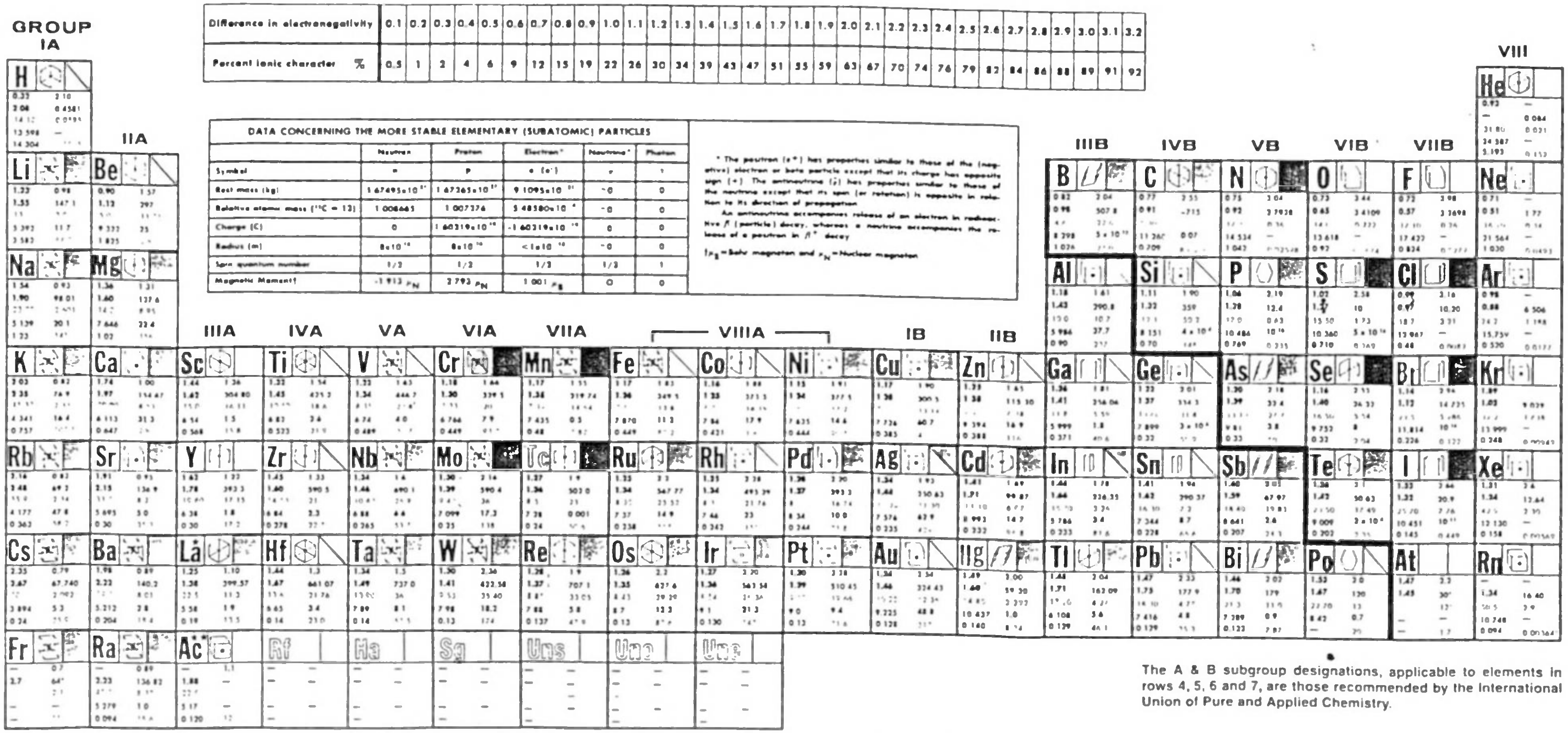
DATA CONCERNING THE MORE STABLE ELEMENTARY (SUBATOMIC) PARTICLES

	Neutron	Proton	Electron*	Neutrino*	Photon
Symbol	n	p	e (e⁻)	ν	γ
Rest mass (kg)	1.67495×10^{-27}	1.67265×10^{-27}	9.1095×10^{-31}	~0	0
Relative atomic mass (^{12}C = 12)	1.008665	1.007276	5.48580×10^{-4}	~0	0
Charge (C)	0	1.60219×10^{-19}	-1.60219×10^{-19}	0	0
Radius (m)	8×10^{-16}	8×10^{-16}	$<1\times10^{-18}$	~0	0
Spin quantum number	1/2	1/2	1/2	1/2	1
Magnetic Moment†	$-1.913\ \mu_N$	$2.793\ \mu_N$	$1.001\ \mu_B$	0	0

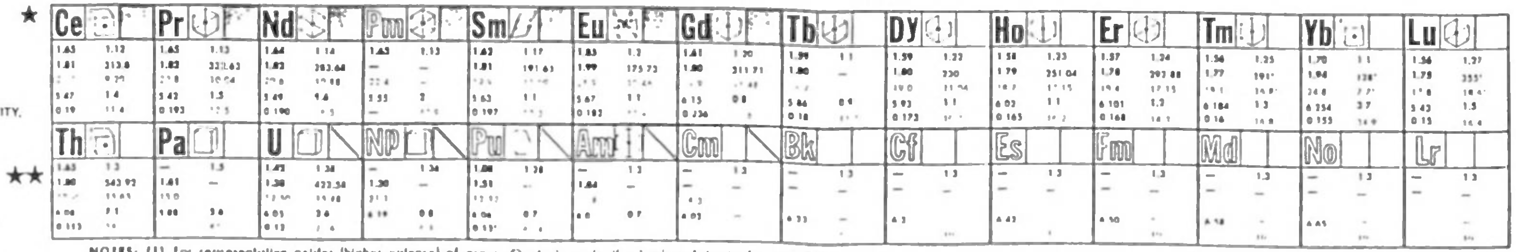
* The positron (e⁺) has properties similar to those of the (negative) electron or beta particle except that its charge has opposite sign (+). The antineutrino ($\bar{\nu}$) has properties similar to those of the neutrino except that its spin (or rotation) is opposite in relation to its direction of propagation.

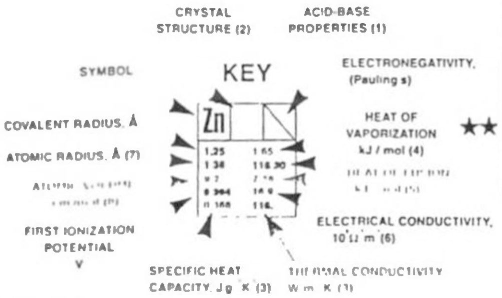
An antineutrino accompanies release of an electron in radioactive β (particle) decay, whereas a neutrino accompanies the release of a positron in β⁺ decay.

†μ_B = Bohr magneton and μ_N = Nuclear magneton



The A & B subgroup designations, applicable to elements in rows 4, 5, 6 and 7, are those recommended by the International Union of Pure and Applied Chemistry.





NOTES: (1) For representative oxides (higher valence) of group. Oxide is acidic if color is red, basic if color is blue and amphoteric if both colors are shown. Intensity of color indicates relative strength.

(2) Cubic, face centered; cubic, body centered; cubic; hexagonal; rhombohedral; tetragonal; orthorhombic; monoclinic.

(3) At 300 K (27°C)

(4) At boiling point

(5) At melting point

(6) Generally at 293K (20°C) for polycrystalline material

(7) Quantum mechanical value for free atom

(8) From density at 300 K (27°C) for liquid and solid elements; values for gaseous elements refer to liquid state at boiling point

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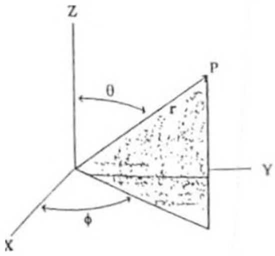
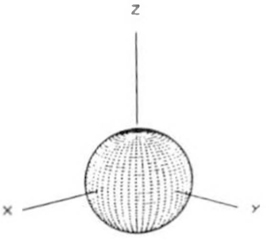
SIDE 2

Electrons in Atoms

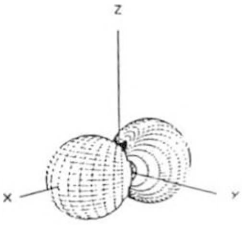
Shapes of Atomic Orbitals



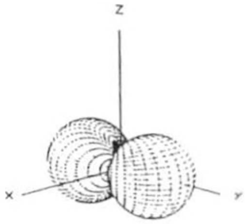
The Spherical Polar Coordinates



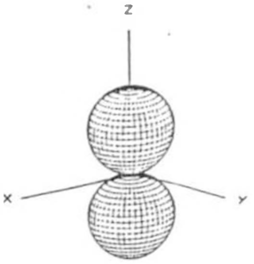
s



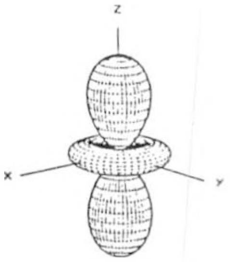
p_x



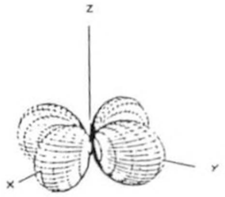
p_y



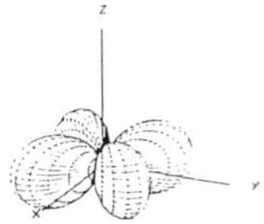
p_z



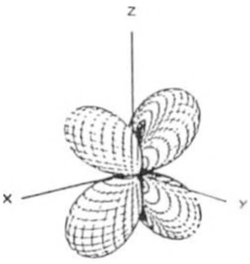
d_{z^2}



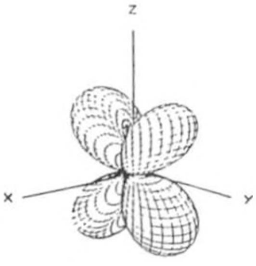
$d_{x^2-y^2}$



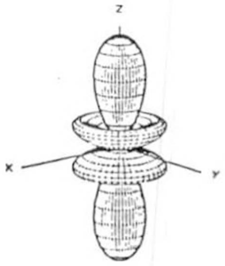
d_{xy}



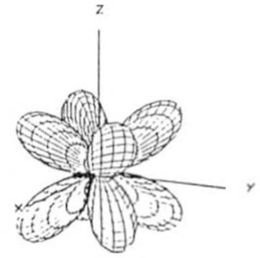
d_{xz}



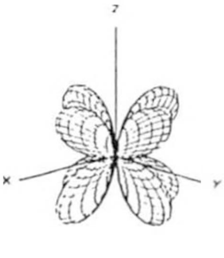
d_{yz}



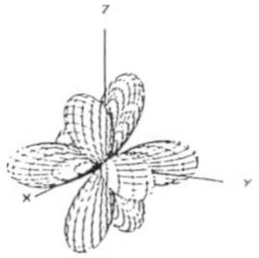
f_{z^3}



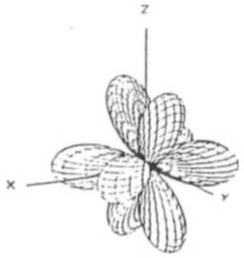
f_{xyz}



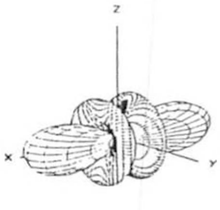
$f_{z(x^2-y^2)}$



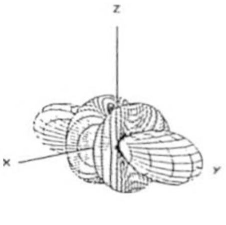
$f_{x(y^2-z^2)}$



$f_{y(z^2-x^2)}$



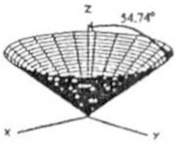
f_{x^3}



f_{y^3}

Angular Part of the Wave Functions: The figures show shapes of atomic orbitals as polar plots. The directions in which the angular part is positive (negative) is indicated in blue (red). A surface on which the angular part of the wave function vanishes is called an **angular node**. It may be a plane or the surface of a cone. The s, p, d and f orbitals have 0, 1, 2 and 3 angular nodes respectively.

The angular parts of the wave functions are listed below. The nodes are also specified using the following notations: (1) $\sigma(X,Y)$ stands for the plane formed by the X and Y axes. (2) $d(X,Y)$ is used to denote a line bisecting the angle between the positive directions of X and Y axes, while $d(-X,Y)$ is the bisector of the angle between the negative X direction and the positive Y direction. Thus, $\sigma(Z,d(X,-Y))$ would stand for the plane formed by the Z axis and a line bisecting the angle between X and -Y directions. (3) $\gamma(Z,54.74°)$ is the surface of a cone, making an angle of 54.74° with the positive Z direction (the figure below shows this surface as a grid). (4) $\phi=45°$ stands for a plane parallel to the $\sigma(X,Z)$ with this value for ϕ.



s-orbital:

$s : \sqrt{\frac{1}{4\pi}}$

p-orbitals:

$p_x : \sqrt{\frac{3}{4\pi}} \sin\theta \sin\phi ;\quad \sigma(Y,Z)$

$p_y : \sqrt{\frac{3}{4\pi}} \sin\theta \cos\phi ;\quad \sigma(X,Z)$

$p_z : \sqrt{\frac{3}{4\pi}} \cos\theta ;\quad \sigma(X,Y)$

d-orbitals:

$d_{z^2} : \sqrt{\frac{5}{16\pi}} (3\cos^2\theta - 1) ;\quad \gamma(Z,54.74°), \gamma(-Z,54.74°)$

$d_{x^2-y^2} : \sqrt{\frac{15}{16\pi}} \cos 2\phi \sin^2\theta ;\quad \phi = 45°, \phi = 135°$

$d_{xy} : \sqrt{\frac{15}{16\pi}} \sin 2\phi \sin^2\theta ;\quad \sigma(X,Z), \sigma(Y,Z)$

$d_{xz} : \sqrt{\frac{15}{16\pi}} \sin 2\theta \cos\phi ;\quad \sigma(X,Y), \sigma(Y,Z)$

$d_{yz} : \sqrt{\frac{15}{16\pi}} \sin 2\theta \sin\phi ;\quad \sigma(X,Y), \sigma(X,Z)$

f-orbitals

$f_{x^3} : \sqrt{\frac{7}{16\pi}} \sin\theta \cos\phi (5\sin^2\theta \cos^2\phi - 3) ;\quad \gamma(X,39.23°), \sigma(Y,Z), \gamma(-X,39.23°)$

$f_{y^3} : \sqrt{\frac{7}{16\pi}} \sin\theta \sin\phi (5\sin^2\theta \sin^2\phi - 3) ;\quad \gamma(Y,39.23°), \sigma(X,Z), \gamma(-Y,39.23°)$

$f_{z^3} : \sqrt{\frac{7}{16\pi}} \cos\theta (5\cos^2\theta - 3) ;\quad \gamma(Z,39.23°), \sigma(X,Y), \gamma(-Z,39.23°)$

$f_{z(x^2-y^2)} : \sqrt{\frac{105}{16\pi}} \sin^2\theta \cos\theta \cos 2\phi ;\quad \phi = 45°, \sigma(X,Y), \phi = 135°$

$f_{y(z^2-x^2)} : \sqrt{\frac{105}{16\pi}} \sin\theta \sin\phi (\cos^2\theta - \sin^2\theta \cos^2\phi) ;\quad \sigma(Y,d(Z,X)), \sigma(X,Z), \sigma(Y,d(-Z,X))$

$f_{x(y^2-z^2)} : \sqrt{\frac{105}{16\pi}} \sin\theta \cos\phi (-\cos^2\theta + \sin^2\theta \sin^2\phi) ;\quad \sigma(X,d(Y,Z)), \sigma(Y,Z), \sigma(X,d(-Y,Z))$

$f_{xyz} : \sqrt{\frac{105}{16\pi}} \sin^2\theta \sin 2\phi \cos\theta ;\quad \sigma(X,Y), \sigma(Y,Z), \sigma(X,Z)$

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Study of a System at Equilibrium

Principle : An acid and an alcohol reacts to form an ester and water while the ester is hydrolysed to give the original acid and alcohol. These opposing reactions proceed until a state of equilibrium is reached. The equilibrium constant obtained in this method is in error since the activity coefficient are neglected. The reaction is

$$C_2H_5OH + CH_3COOH \rightleftharpoons CH_3COOC_2H_5 + H_2O$$

Since the equilibrium is obtained very slowly, it is necessary to add HCl as catalyst.

Procedure : The following mixtures in the bottles are prepared, stoppered immediately and allowed to stand for a week in a warm place with occasional shaking. It is not essential to thermostat the mixture when equilibrium is established. Titrate the contents of each bottle (2 ml) with the given standardised NaOH solution.

Sample No.	Vol of HCl(cc)	Vol of ethyl acetate (cc)	Vol of H_2O (cc)	Vol of Absolute alcohol(cc)	Vol of acetic acid (cc)
1	250	0	250	0	0
2	250	250	0	0	0
3	250	200	50	0	0
4	250	100	150	0	0
5	250	200	0	50	0
6	250	200	0	0	50

For the calculation, weight of each reactant and product in the mixture is calculated. From the data, equilibrium constant (esterification/ hydrolysis) is obtained.

POTENTIOMETRY

Aim: To study the relative strengths of chemical species as oxidising agents and hence to construct electrochemical cells of different electromotive force.

Introduction :

It is known that elements differ from one another in their bility to function as oxidizing agents. A measure of this bility is their tendency to gain electrons. The noticeable vidence of this is when a metal ion from a solution of the etallic salt is deposited as the metal; or when a non-metallic lement in solution is converted to the anion. You will collect uch evidences in Part 1 of this experiment.

Since redox reactions involve transfer of electrons from one species to another, we can regard such reactions as combinations of two half reactions -- one in which a species is undergoing oxidation (giving up electrons) and a second, in which a species is undergoing reduction (gaining electrons). In Part 1, the two species were in direct contact, so that the electron transfer occurred directly. The two species could as well be physically separated from one another and the electrons generated by the oxidation of one species transferred through a metal wire to the other, which could undergo reduction. In Part 2 of this experiment you will be doing this.

The electrons that are being transferred through the metal wire may be able to perform work for us - they can light a bulb; they can drive the needle of a galvanometer and so on. In fact these latter can be taken as the observable evidences for electron transfer. The device is thus able to perform electrical work. It is able to do so because in one part of the device, we have a species which can undergo oxidation and in another part, a species which can undergo reduction. Since these essentially chemical processes are responsible for any electrical work that the device can perform, the device is called as electrochemical cell. Just as the cell reaction is a combination of two half reactions, the cell is a combination of two half cells. In each half cell is a suitable electrolyte solution in contact with a metal. The reaction in the half cell occurs at the metal-solution interface. Electrons enter and leave the solutions through the metals in contact with them. Electrical contact between the solutions in the two half cells is maintained usually through a salt bridge which consists of concentrated solution of KCl or NH_4NO_3 in a U-tube.

Since electrons can be transferred from one point to another only if there is a potential difference between the points, we can infer that the electrodes in the two half cells are at different potentials. The difference in these 'electrode potentials' is called the electromotive force, emf of the cell. There is no experimental method for determining the individual electrode potentials. The potential difference between the electrodes, that is the emf of the cells can however be measured either with a high resistance voltmeter or a potentiometer. The potentiometer is almost invariably employed for reasons to be discussed later. In this experiment, you will measure the emfs of several cells. A direct reading millivoltmeter will be used so that the measurements can be made expeditiously.

In Part 3 of this experiment, you will determine the emfs of the cells in which the calomel electrode is one of the electrodes. Since one of the electrodes is kept common, this method could serve the purpose of comparing the potentials of several electrodes with reference to calomel. We can then think of a relative scale of electrode potentials. Once we understand this, we can proceed to find out how these relative potentials of electrodes depend on the concentrations of the electrolyte solutions in the half cells of which they are a part. This will be done in Part 4.

In all these measurements, an important observation to be made is the polarity of the electrode. If the electrode is the negative terminal, an oxidation would be the spontaneous process at it. If it is positive, a reduction would be the spontaneous process at the electrode. You will notice that the electrode in different combinations may assume different ities. One could infer that whether oxidation or reduction occur in a half cell depends on the nature of the other cell. These points would become clearer during the course he experiment.

rials Needed :

Metal strips : Copper, Zinc, Lead, Nickel, etc.
100 cm^3 beakers : Two
Test tubes (large size) : Five
Measuring cylinder 20 cm^3 and 5 cm^3 : One each
Emery paper (3" square) : Two sheets
Approx. 0.1 M solutions of (about 100 cm^3 of each) : $CuSO_4$, $ZnSO_4$, $NiSO_4$, $Pb(NO_3)$, HCl, KBr, KI, etc.
Aqueous solutions of halogens : i) Chlorine water
ii) Dilute solutions of Br_2 in KBr and I_2 in KI.
Salt bridge (Agar + KCl) : One
Calomel electrode : One
Platinum foil electrode : One
Millivoltmeter and connecting wires

ocedure :

Part 1 :

1. Transfer about 10 cm^3 each of the solutions of $CuSO_4$, $ZnSO_4$, $NiSO_4$, $Pb(NO_3)_2$ and HCl into five separate test tubes.

Clean the strip of zinc with emery paper and dip in the solution of $CuSO_4$. Note the reaction, if any. Record in Table 1. (Use a X mark to indicate 'no reaction' and a ✓ mark to indicate 'reaction').

Wash the strip of zinc with water and clean with emery paper. Dip it in the solution of $NiSO_4$. Note and record reactions, if any, in Table 1 as in step 2.

Test the reaction of zinc with the other solutions provided, including HCl, as in the previous steps. Record as before.

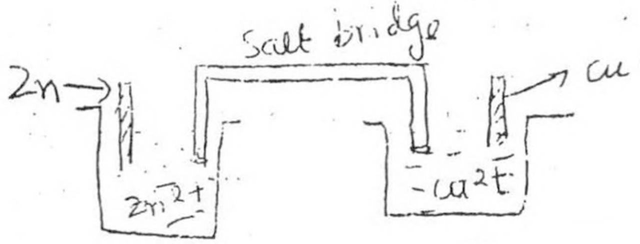
5. Repeat the above steps with the other metal strips. Record all observations in Table 1.
6. At the end of the experiments, clean the metal strips and keep them aside.
7. Take, in three separate test tubes, about 5 cm^3 each of the solutions of KCl, KBr and KI. Add a little Cl_2 water to each test tube. Observe and note changes, if any. Add two drops of CCl_4 to each test tube, shake and note observations in Table 1.
8. Repeat step 7 with Br_2 and then I_2 solutions in place of chlorine water. Record all observations in Table 1.

Part 2

1. Dip clean strips of the metals in their respective salt solutions taken in separate beakers (Copper in $CuSO_4$ and so on).

   Each of these is a half cell. Use the notation $M_{(s)}|M^{z+}_{(aq)}$ to represent the half cell (For example, $Cu_{(s)}|Cu^{+2}_{(aq)}$.
2. Connect the $Cu|Cu^{+2}$ half cell to the $Zn|Zn^{+2}$ half cell using a salt bridge as indicated in the diagram.

   This is a typical arrangement for an electrochemical cell.



3. Connect the copper and zinc strips to separate terminals on the millivoltmeter using crocodile clips and connecting wire. If the voltage indicated is positive, note the polarities of the two metal strips. Note the emf in Table 2. If the emf indicated is negative, change the polarities.
4. Combine the $Cu | Cu^{+2}$ half cell with each of the other half cells in turn and note the emf and the polarities of the metal strips in each case. Record the data in Table 2. (It is essential that the two legs of the salt bridge must first be cleaned on the outside with distilled water before dipping them in new solutions).

5. Repeat steps 2 to 4 with another pair of half cells. Work with all possible combinations of half cells of step 1. Record all the emf data in Table 2.

Part 3 :

1. Take one of the half cells constructed in Part 2 and dip the calomel electrode in it. The construction of the calomel electrode supplied to you is such that you can avoid the U-tube salt bridge in combining it with the half cell.
2. Connect the calomel to one terminal and the metal strip to the other terminal of the millivoltmeter and note the emf. Record in Table 3.
3. Repeat steps 1 and 2 with the other half cells of Part 2.
4. Wash the beakers and the metal strips. Clean the metal strips and keep them aside.

5. Take the following solutions in separate beakers.
   a) solution of Br_2 in KBr.
   b) solution of I_2 in KI.
   c) solution of $FeSO_4$ acidified with dil H_2SO_4 and partially oxidized by adding $KMnO_4$.

6. Dip a platinum foil electrode in one of these solutions. This will now be a half cell.
7. Dip the calomel electrode into the solution in step 6 to complete the cell assembly. connect the calomel and platinum foil electrodes to the terminals on the millivoltmeter. Note the emf and the polarities of the electrodes. Record in Table 3.

8. Wash the platinum foil electrode with distilled water. Similarly, wash the calomel electrode on the outside. Repeat steps 6 and 7 with each of the solutions taken in step 5. Record all data in Table 3.

Part 4 :

1. Dry one of the beakers with filter paper. Transfer, using the measuring cylinder, 20 cm^3 of one of the metallic salt solutions. (Your teacher will tell you which solution should be taken). Clean the corresponding metal strip and dip it into the solution).

2. Dip the calomel electrode into the solution above and measure the emf of the cell as before. Record the emf and the polarities of the electrodes in Table 4.

3. Use measuring cylinder to add 5 cm^3 of distilled water to the solution in the beaker. Stir for uniform concentration. Note the emf and record in Table 4.

4. Repeat step 3 with successive additions of 5 cm^3 of water and note the emf after each addition. Record all data in Table 4. Collect readings of emf at six different concentrations of the metallic salt solution.

Treatment of Data :

1. Study Table 1 and identify the metals which were able to reduce
   a) all the three other metallic ions
   b) only two of the three metallic ions
   c) only one of the three metallic ions and
   d) none of the metallic ions.

   Hence arrange the metal - metal ion half reactions

   $M^{+2}_{(aq)} + 2e \rightleftharpoons M_{(s)}$

   in the order of decreasing ease of reduction.

2. Refer to the reactions of the metals with dil HCl. Hence identify the metal which was not able to reduce hydrogen ions in the acid solution. Fit the half reaction

   $H^{+}_{aq} + e \longrightarrow H_{2(g)}$ into the above table.

3. Consider the reactions of the halogens with the solutions containing the halide ions. Prepare a table of decreasing ease of reduction of the halogens. This can be represented by the equation

   $X_{2(aq)} + 2e \longrightarrow 2X^{-}_{(aq)}$

4. Write balanced total reactions for the cases where redox reactions were observed. An example is given below.

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

5. In the electrochemical cells assembled by you in Part 2 of this experiment, an oxidation half reaction takes place at the electrode in one half cell and a reduction half reaction at the electrode in the other half cell. The former electrode is called the anode and the latter, the cathode. The electrodes can be readily identified. The one which is connected to the negative terminal on the millivoltmeter is the anode and the other is the cathode.

. In each of the half cells used by you, identify the two electrodes. Write the spontaneous half reactions occurring at these electrodes. Combine the two half reactions to write the cell reaction.

6. You have learnt the method of representation of the electrodes. The cells, which are combinations of two such electrodes can be represented as in the following example.

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

Note that the anode or the negative electrode is on the left and the positive electrode is on the right. The double vertical line indicates that a salt bridge has been used.
Represent each cell by this method and arrange the cells in the decreasing order of their emfs.

7. In the case of each half cell identify the combination in which

a) reduction was the spontaneous process in the half cell
b) oxidation was the spontaneous process in the half cell

8. In part 3 of this experiment, you have used the calomel electrode as a reference electrode. The calomel electrode can be represented as follows :

$Pt \mid Hg_{(1)} \mid Hg_2Cl_{2(s)}, KCl_{(satd\ aq)}$

Represent the cells you have constructed by combining other electrodes with the calomel electrode, bearing in mind that the negative electrode must always be on the left.

a) Which are the electrodes at which the spontaneous process is oxidation and
b) which are those at which reduction is the spontaneous process . ?

. Calculate the concentrations of the metallic salt solutions in Part 4. Enter the values in Table 4. Construct plots of

a) emf vs concentration

b) emf vs log concentration

Discussion :

1. I_2 is a stronger oxidizing agent than $Cu_{(s)}$. Use this information to construct a series of all the half reactions in part 1, in order of decreasing ease of reduction.

2. In the electrochemical cells constructed in Part 2,
   a) what is the direction of electron flow in the external circuit ?
   b) what is the direction of flow of negative ions in the solution ?
   c) why is it appropriate to designate the electrode at which oxidation occurs as the negative electrode ?

3. In a cell it is found that the calomel electrode is the cathode. What is the spontaneous reaction at this electrode ? write the balanced chemical equation for the reaction .

4. The following procedure is recommended for calculation of cell emfs from electrode potentials and for deciding the direction of the spontaneous process in a cell.

a) The cell is represented with the negative electrode on the left.

b) The cell emf is $E = E_R - E_L$, where E_R and E_L are reduction. potentials of the electrodes on the right and left side respectively. If this difference is positive, the cell representation is correct and the cell process is spontaneous. If not, the position of the electrodes are reversed in the representation.

c) The reactions at both electrodes are written as reductions.

d) The cell reaction is $R_R - R_L$, where R_R and R_L are the reduction processes at the two electrodes.

Use these conventions to answer the following questions.

1. If it is assumed that the reduction potential for the saturated calomel electrode is +0.2415 V at 25°C, calculate, from the measured emfs of the cells in Part 3, the reduction potentials of the other electrodes in combination with the calomel electrode in the various cells.

2. Using the values of the potentials calculated above, predict the emfs of the cells in Part 2. Enter the predicted values in a column adjacent to the experimentally measured values in Table 2.

3. In assigning electrode potentials, the potential of the standard hydrogen electrode (S.H.E) is arbitrarily taken as zero. Refer to the text book for a description of the S.H.E.

a) What is the meaning of the statement that the reduction potential of the saturated calomel electrode is +0.2415 V?

b) How would you represent the cell in which one of the electrodes is the saturated calomel and the other is the S.H.E. ? What will be the spontaneous cell reaction ?

c) Fit the standard hydrogen electrode potential into the series in which electrodes are arranged in the decreasing order of potentials ?

4. a) In part 4, you have observed that the emf of the cell changes as the concentration of the solution in one of the electrodes is changed. Are both the electrode potentials changing ? If not, which of these potentials is changing on dilution ?

b) Suggest an empirical relationship between the cell emf and the concentration of the solution, from the graphs you have drawn.

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Thermochemical Measurements

THE HEAT OF REACTION

In this experiment you will use a 250 ml Erlenmeyer flask as the reaction vessel and as a simple calorimeter to measure the heat evolved or absorbed during the reactions.

You may assume that the heat of reaction will be used to change the temperature of the aqueous solution and of the glass of the container. We shall neglect other small losses to the surroundings. Recall that it takes 1.0 calorie to change the temperature of one gram of water one degree centigrade. It takes 0.2 calorie to change the temperature of one gram of glass one degree centigrade.

You do not need to weigh the water used since 1.0 ml of water weighs very nearly 1.0 g and you will measure the volume to the nearest milliliter. When the reactants are added to the reaction flask you should note the change in temperature to the nearest 0.2°C. From the change in temperature and the weight of the reactants you can calculate the number of calories evolved or absorbed.

In this experiment you will measure and compare the quantity of the heat involved in three reactions.

Reaction 1: Solid sodium hydroxide dissolves in water to form an aqueous solution of ions.

$$NaOH(s) \longrightarrow Na^+(aq) + OH^-(aq) + x_1 \text{ cal}$$

$$H_1 = -x_1 \text{ cal}$$

Reaction 2 : Solid sodium hydroxide reacts with an aqueous solution of hydrogen chloride to form water and an aqueous solution of sodium chloride :

$$NaOH(s) + H^+(aq) + Cl^-(aq) \longrightarrow H_2O + Na^+(aq) + Cl^-(aq) + x_2\text{cal}$$

$$H_2 = -x_2 \text{ cal}$$

Reaction 3: An aqueous solution of sodium hydroxide reacts with an aqueous solution of hydrogen chloride to form water and an aqueous solution of sodium chloride :

$$Na^+_{(aq)} + OH^-_{(aq)} + H^+_{(aq)} + Cl^-_{(aq)} \longrightarrow H_2O + Na^+_{(aq)} + Cl^-_{(aq)} + x_3 \text{ cal}$$

$$\Delta H_3 = -x_3 \text{cal}$$

Procedure :

Determination of the Heat of Reaction 1 :

a) weigh a clean dry 250 ml Erleymeyer flask to the nearest 0

b) Put 200 ml ($\pm$ 1 ml) of cool tap water into the flask. Sti carefully with a thermometer until a constant temperature is reached (about room temperature). Record this temperature to the nearest 0.2°C.

c) weigh about 2 g of solid sodium hydroxide, NaOH to the nearest 0.01 g. Since sodium hydroxide becomes moist as it is being weighed in the open air, your teacher will give you spec instructions on weighing rapidly a prescribed number of the solid pellets which will approximately 2 g (between 1.9 g and 2.1 g).

d) Pour the weighed NaOH(s) into a water in the Erlenmeyer fl--k Swirl the flask until the sodium hydroxide is dissolved. Place the thermometer into the flask and record the extreme temperature reached.

Before proceeding to Reaction 2, rinse the 250 ml flask thoroughly with water.

Determination of the Heat of Reaction 2

Repeat steps a,b,c and d used in the determination of the heat for Reaction 1 except in step b substitute 200 ml of 0.25 M HCl for tap water. Rinse the 250 ml flask again and proceed to Reaction 3.

Determination of the Heat of Reaction 3 :

a) Measure 100 ml of 0.5 M HCl into the 250 ml flask and 100 m of 0.5 M NaOH into a 250 ml beaker. Both of these solutions should be at, or slightly below, room temperature. Check this with a thermometer. (Rinse and dry the thermometer before transferring it from one solution to another). Record the temperatures.

) Add the sodium hydroxide solution to the hydrochloric acid solution. Mix quickly and note the highest temperature reached.

Calculations :

. For each reaction, calculate

a) the change in temperature

b) the amount of heat that is absorbed by the solution,

c) the amount of heat absorbed by the beaker;

) the total amount of heat absorbed;

) the number of moles of NaOH used in each reaction

) the total amount of heat involved per mole of NaOH.

I . Express your results as heats of reaction :
ΔH_1 ΔH_2 and ΔH_3

I a) Compare ΔH_2 with $\Delta H_1 + \Delta H_3$ and explain.

) calculate the percent difference between ΔH_2 and $\Delta H_1 + \Delta H_3$ ssuming ΔH_2 to be correct.

uestions :

Write the net ionic equations for Reactions 2 and 3.

In reaction 1, ΔH_1 represents the heat of solution of $NaOH(s)$. Look at the net ionic equations for Reactions 2 and 3 and make a statement concerning the significance of ΔH_2 and ΔH_3.

Suppose you had used 4 g of $NaOH(s)$ in Reaction 1.

) what would be the number of calories evolved ?

) what effect would this have on your calculation of ΔH_1, the heat involved per mole ?

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Chemical Kinetics

Experiment :

Aim: To study the kinetics of the reaction between potassium peroxydisulphate and potassium iodide by the initial rate method. (Clock Reaction)

Introduction :

We have seen earlier that the time required for a particular fraction (x/a) of reaction to occur is independent of initial concentration a, if the reaction is first order. If $(x/a) = 1/n$, we can write that

$$k_1 t_{1/n} = \ln \; n/(n-1) \qquad (1)$$

where $t_{1/n}$ is the time for one nth of the reaction to occur. If we keep $1/n$ small, we would be confining our attention to the initial stages of the reaction and $t_{1/n}$ would be inversely related to the initial rate of the reaction.

One of the methods by which the time for one nth of a reaction to occur can be measured is illustrated by the <u>iodine clock</u>. This technique is employed here to study the reaction between $K_2S_2O_8$ and KI. A reaction mixture consisting of a known excess of KI and a relatively small concentration of $K_2S_2O_8$ and sodium thiosulphate and a little starch is prepared. The iodine liberated by the reaction between $K_2S_2O_8$ and KI should result in the appearance of a blue colour because of the starch present in the mixture. But the appearance of the colour is <u>delayed</u> till the $Na_2S_2O_3$, which is also a component of the mixtur [illegible] the blue colour theref [illegible] particular amount of iodine to be liber [illegible] This is therefore a measure of the <u>initial rate</u> of the reaction between $K_2S_2O_8$ and KI.

Once the technique for measurement of initial rate is chosen, we can use it for measurements of the rate under varying compositions of reaction mixture, temperature and so on. This is a convenient technique because a large number of measurements can be done in a short time. Moreover, when attention is confined to the initial stages of the reaction, complications due to the accumulation of products will not be significant.

Materials Needed :

1. Pipettes : 10cm^3, 10cm^3 graduated, 5cm^3, 2cm^3, 1cm^3 - one each.
2. Boiling Tubes : Two
3. Beaker 250 cm^3 : One
4. 0.5 M KI in a coloured bottle : About 250 cm^3
5. 0.015 M $K_2S_2O_8$ in a stoppered bottle : About 100 cm^3
6. 0.01 M $Na_2S_2O_3$ in a stoppered bottle : About 100 cm^3
7. 5×10^{-4} M $CuSO_4$ in a stoppered bottle : About 100 cm^3
8. 2% starch in a bottle with a dropper
9. Thermometer 0.1° sensitivity : One
10. Stopwatch : One

Procedure :

Part 1 :

1. Pipette 10cm^3 of KI followed by 1 cm^3 of $Na_2S_2O_3$ into the boiling tube labelled A.
2. Using the graduated pipette, add 9 cm^3 of distilled water to the above solution. Mix and place the tu[illegible] (thermostat).

   [illegible] cm^3 of $K_2S_2O_8$ into the boiling tube labelled B. Add 3 drops of starch, mix and place the tube in the beaker containing water. (thermostat).
4. Wait for temperature equilibriation; note the temperature of the thermostat.
5. Rapidly transfer solution B to solution A, starting the stop watch simultaneously. Pour back and forth between the tubes twice and place the mixture in the thermostat. When a definite blue colour appears, stop the watch and note the time. Record in Table 1, as time for reaction mixture 1.

Repeat the above steps with reaction mixtures 2,3 and 4 whose compositions are given at the end of these instructions. Record the corresponding times in Table 1.

at the above steps with reaction mixtures 5, 6 and 7 and in case note the time for the appearance of blue colour. Record time in Table 2. Include in the table, the time for mixture 1 and in Part 1.

Thermostat a mixture of 10 cm^3 of KI, 1 cm^3 of $Na_2S_2O_3$ and 3 cm^3 of water (mixture 8) in boiling tube A.

Thermostat a mixture of 5 cm^3 of $K_2S_2O_8$, 1 cm^3 of $CuSO_4$ and 3 drops of starch in boiling tube B.

i the two solutions rapidly, starting the stopwatch simultaneously. Pour back and forth between the tubes twice ir place the mixture in the thermostat. Stop the watch the moment a blue colour is observed. Note the time. Record in a e 3.

e at the above steps with mixtures 9, 10 and 11 and record h times in Table 3.

4

r pare a mixture of 10 cm^3 of KI, 1 cm^3 of $Na_2S_2O_3$ and 9 cm^3 ube A and place the tube i mostat).

ipette 5 cm^3 of $K_2S_2O_8$ into boiling tube B. Add 3 drops of starch, mix and place the tube in the thermostat.

i two lumps of ice to the water in the beaker. Place a mometer in the solution in tube A. When a steady ther erature is attained, rapidly transfer the solution from B A, simultaneously starting the stopwatch. Pour back and h between the tubes twice and place the mixture in the er. Stop the watch the moment a blue colour appears. t the time and the temperature. Record these in Table 4. s the tubes.

4. Pipette the solutions as given in steps 1 and 2 into the boiling tubes. Place them in the thermostat. Add some water to the water in the beaker. Stir with a glass roc

5. Place the thermometer in solution A. When a steady temperature is attained, rapidly mix the solutions, start the sto watch and note time for appearance of the blue colour. Note the temperature of the mixture and record the time and temperature in Table 4.

6. Repeat the experiment as above at two more temperatures. The change in temperature can be brought about by suitab additions of warm water to the beaker.

Treatment of Data :

1. Calculate the concentrations of each component of the reaction mixtures 1 to 11. Enter them in appropriate columns in Tables 1 to 3.

2. The concentration of $Na_2S_2O_3$ in the mixtures is related to the amount of iodine liberated when the blue colour appears. Calculate the concentration of iodine at the first appearan of blue colour. (Remember that 1 mol of $I_2 \equiv 2$ mols of $Na_2S_2O_3$). Since the concentration of $Na_2S_2O_3$ is the same in all the mixtures, the times recorded are the times for the appearance of the same amount of iodine in each case.

3. Calculate the initial rates in terms of mols dm^{-3} of iodin liberated per unit time. Record the [illegible] bles 1 to 4.

4. In all the reaction mixtures, $K_2S_2O_8$ is taken at a much lc concentration than KI. It is, therefore, the limiting reactant. With this as the reference, calculate the fract reacted (x/a) or (1/n) when the blue colour appears.

Discussion :

1. How are we justified in assuming that the time measured in these experiments is inversely related to the initial rate of the reaction ?

2. What variation was done in Part 1 ?

3. Assuming that the reaction is first order with respect to $S_2O_8^{-2}$, calculate the pseudo first order constant from the times recorded in Part 1. (Use ecn.(1) of introduction).

Record the values in Table 1. Is our assumption about the order valid ?

4 What variation was done in Part 2 ?

5 Calculate the pseudo first order rate constants in Part 2. Enter the values in Table 2.

6. Plot log k_1 against log $[I^-]$ and join the points by the st line. Find the order with respect to I^- and the true rate nstant of the reaction. Express in appropriate units.

7 One possible variable that affects the rate constant is the ionic strength of the medium. Calculate the ionic strengths in respect of the four mixtures used in Part 2 and convince yourself that it is mainly determined by the concentration of KI. Since this is maintained constant in the mixtures used in the other parts of this experiment, it can be assumed that the reactions in the other parts were carried out at constant ionic strength.

8. Plot log k_2 against $\sqrt{I}$ where k_2 is the second order rate constant calculated in step 6 and I is the ionic strength of the solution, which is calculated in step 7. Find the slope of the graph. What indication does this give about the nature of the particles involved in the formation of the activated complex for the reaction ? Identify these particles and write an equation for the formation of the complex.

9. Cu^{+2} introduced in Part 3 functions as a catalyst. What is the evidence ?

10 Calculate the pseudo first order rate constants [illegible] Enter the values in Tabl

1. Since the reaction can also occur in the absence of the catalyst, the observed rate constant k_1 is related to the oncentration of the catalyst as follows .

$$k_1 = k_0 + k_{cat} [Cu^{+2}] \qquad (2)$$

here k_0 is the rate constant for the uncatalysed path and k_{cat} is that for the catalysed path. Test this equation graphically and calculate k_{cat} from the graph. Express in proper units.

12. From Arrhenius equation and eqn(1) of introduction, show that

$\log t_{1/n} = E/ 2.303\ RT$ - a constant. (3)

13. Plot $\log t_{1/n}$ against $1/T$ in Part 4 of this experiment. Join the points by the best line and find the energy of activation. Express in proper units.

14. Calculate the second order rate constant k_2 at each temperature from the known concentration of I^- and k_1.

15. Plot $\log k_2$ against $1/T$. Join the points by the best line. Find E_a and A from the graph.

16. Calculate k_2 at 25°C from the known values of E_a and A.

17. Calculate $\Delta H^{\#}$, $\Delta G^{\#}$ and $\Delta S^{\#}$ for the reaction.

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