# **Department of Nutrition and Food Engineering (NFE)**

**Daffodil International University**

**L A B O R A T O R Y M A N U A L**

###### **Course No.: NFE – 116 Unit: 1.0 Marks: 100**

**Course Title: Physical Inorganic & Analytical Chemistry Practical**

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**Laboratory Module A:**

1. **Experiment: Identification of Group – 1 and Group – 2 metal ions.**

**Flame tests:**

Dip the clean platinum wire (it should impart no color to the flame) into conc. HCl on a watch glass, then into a little of the substance whereby a little adheres to the wire. Introduce the wire into the base of the non-luminous Bunsen flame (air holes open). Observe the color in naked eye and also through two thickness of cobalt blue glass.

Flame coloration to Flame coloration Inference

**Naked eye through cobalt glass**

|  |  |  |
| --- | --- | --- |
| Persistent golden yellow  Violet  Brick red (yellowish-red)  Crimson (deep red)  Yellowish-green  Green  Livid blue flame  (wire slowly corroded) | Nil (colorless)  Crimson (deep red)  Light green  Purple  Bluish-green  Bluish green  ------------ | Sodium  Potassium  Calcium  Strontium  Barium  Borates, copper  Lead, arsenic antimony, bismuth. |

**After all the tests, clean the pt-wire with conc. HCl & store the wire permanently in the acid.**

1. **Experiment: Identification of Cations (Fe2+, Fe3+, Cr3+, Mn2+, Pb2+, Ni2+, Co2+, Al3+, Hg2+, Cd2+)**

THE ANALYTICAL CLASSIFICATION OF CATIONS:

For analytical purposes the cations are classified into five groups. This classification is based upon the **varying solubilities** of the chlorides, sulphides, hydroxides, phosphates and carbonates in acidic or basic mediums. A group is separated from the other by adding a reagent which precipitates the ions of one group and not the others. This reagent is called the **group reagent**. The various groups are summarized in the Table 1.

#### **TABLE 1: CLASSIFICATION OF CATIONS INTO GROUPS**

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| **Group** | **Ions in Group** | **Group Reagent** | **Group Precipitate** | **Distinguishing features** |
| **1** | Ag+, Hg2++, Pb++ | Dilute HCl | AgCl, Hg2Cl2, PbCl2 | Chlorides insoluble in dilute HCl |
| **II-A** | Ag++, Pb++, Bi+++, Cu++, Cd++ | H2S (thioacetamide) in 0.3M HCl | HgS, PbS, Bi2S3, CuS, CdS | Sulphides insoluble in 0.3M HCl |
| **II-B** | As+++, Sb+++, Sn++, Sn++++ | As2S3, Sb2S3, SnS, SnS2, |
| **III-A** | Al+++, Cr+++, Zn++, | H2S in presence of NH3 and NH4Cl | Al (OH)3, Cr(OH)3, ZnS MnS, NiS, CoS, FeS, Fe2S3 | Hydroxides or sulphides insolble in (NH4OH + H2S) |
| **III-B** | Mn+++, Ni++, Co++, Fe2+, Fe3+ |
| **IV** | Ba++, Sr++, Ca++, Mg++ | (NH4)2CO3, (NH4)2HPO4 in presence of NH3 | BaCO3, SrCO3, CaCO3, Ca3(PO4)2,  Ba3 (PO4)2, Sr3(PO4)2, MgNH4PO4, 6H2O | Carbonates & phosphates precipitated by (NH4)2CO3 & (NH4)2HPO4 from an ammoniacal solution |
| **V** | Na+, K+, NH4+ | No particular reagent. | Soluble group,  Na+, K+, NH4+ in solution. | Ions not precipitated in previous groups. |

**TABLE 2: ANALYSIS OF GROUP I CATIONS**.

White precipitate obtained in **procedure 1** may be PbCl2, AgCl and Hg2Cl2.

Place 10 mg of the salt in a 10 ml test tube. Add 6-7 drops of water and heat with stirring for 3 min in boiling water bath. Centrifuge quickly and immediately remove centrifugate while keeping mixture hot in water bath. Save both residue and centrifugate. Repeat the hot water treatment once again and combing the two centrifugates.

**Centrifugate may contain Pb++ ions.**

Add 4 drops of 1M K2CrO4.

Formation of a yellow precipitate of PbCrO4 confirms the presence of the **LEAD ION**

**Residue may contain AgCl and Hg2Cl2**.

Wash twice with hot water, as above, to remove any excess of Pb++. Add 5 drops of conc. ammonia to remaining residue. Mix thoroughly and centrifuge

**Centrifugate** may contain Ag(NH3)+2.

Add 3 M HNO3 dropwise with stirring until solution is definitely acid. The formation of a white precipitate of AgCl confirms the presence of **SILVER ION.**

**Residue** may contain Hg and

Hg NH2Cl, and some Hg2O. This combination appears black in colour and confirms the presence of **MERCUROUS ION.**

**TABLE 3: ANALYSIS OF GROUP II CATION**

Place 10 mg of the salt in a 10 ml test tube. Add 10 drops of 3M HNO3 heat in water bath for 3 min, and centrifuge, if much residue is left, repeat treatment, combining centrifugates.

**Residue** may be black HgS or white Hg(NO3)2, HgS and S. Dissolve in 4 drops of aqua regia (3 drops conc. HCl & 1 drop conc. HNO3). Add 10 drops of water and boil in water bath for 2 min, cool and add 2 drops of SnCl2 solution. White precipitate of Hg2Cl2 or gray precipitate of Hg and Hg2Cl2 confirms the presence of **MERCURIC ION**.

(**aqua regia:** conc. HCl **:** conc. HNO3 = 3:1) by volume)

**Centrifugate** contains lead, bismuth, copper, and cadmium ions. Add 10 drops of concentrated NH3, stir, and centrifuge. Save centrifugate.

**Centrifugate** contains soluble Cu(NH3)4++ & Cd(NH)++ complex ions. Divide into unequal portions (¼ and ¾)

**Precipitate** may be basic salts of bismuth and lead. Add 4 drops of 3M KOH, stir, warm in water bath and centrifuge. Save the centrifugate. Repeat KOH treatment and combine the two centrifugates.

**Large portion**.

If copper or mercury (or both) are present, add 5 or more drops of 1M KCN. Now add 10 drops of 3M KOH and 20 drops of formaldehyde. Centrifuge and wash precipitate with water three times, stirring thoroughly each time (do not perform above steps if Cu and Hg are absent) Add 5 drops of thioacetamide and allow to stand for 1 min. Yellow precipitate of CdS confirms the presence of **CADMIUM ION**. If a brown or black precipitate is obtained see note below

**Small potion**.

A blue colour indicates the presence of copper. Make solution acid with 3M CH3COOH and add 20 drops of potassium ferrocyanide solution. A reddish brown precipitate of Cu2Fe(CN)6 confirms the presence of the **CUPRIC ION**.

**Centrifugate** may contain lead as HPbO2 ion. Make barely acid with 3M CH3COOH. Add 2 drops of 1M K2CrO4 Yellow precipitate of PbCrO4 confirms the presence of the **Lead Ion**.

**NOTE**: A brown or black precipitate obtained in the Cadmium test indicates a faulty previous separation. In this case proceed as follows: Centrifuge the mixture and discard the washings. Add 10 drops of 1.5M H2SO4 stir and heat in water bath for 4 min. Centrifuge and discard any residue. Neutralize the centrifugate with concentrated NH3. Add 5 drops of thioacetamide solution and warm in water bath for 3 min. A yellow precipitate will form if Cd is present.

**TABLE 4 ANALYSIS OF GROUP IIIA CATIONS**

Place 10 mg of the salt in a 10 ml test tube. Add concentrated HCl, drop by drop, with stirring until solution is barely acid. Next, add 20 drops of saturated NH4Cl, and concentrated NH3 dropwise, with stirring until solution is barely basic. Add 4 drops of concentrated NH3 in excess. Place test tube in boiling water bath for 3 minutes. Centrifuge.

P**recipitate** may be aluminum hydroxide, wash with 10 drops of water and discard washings. Dissolve in 10 drops of 3M HCl. Divide into 2 portions: Portion I: Add freshly filtered aluminon until soln. is definitely pink. Add 1 drops of given ammonium acetate; make slightly basic with NH4OH , Heat in boiling water bath for 4-5 min. & centrifuge. Pink-red ppt indicates **Al+++.** **Portion II:** Repeat as above using alizarins (alizarin blue) instead of aluminon, Blue ppt confirms **Al+++.**

**Centrifugate** may contain CrO4 (coloured yellow if CrO4 is present: Place a few drops of centrifugates in a test tube acidify with CH3COOH & add one drop of Pb-acetate soln. Yellow ppt of PbCrO4 confirms Cr.) and Zn(NH3)4++. Add 1M barium acetate until precipitation is complete. Centrifuge.

Precipitate may be BaCrO4 and BaSO4. Add 10 drops of 3M HCl and warm 1 min in water bath. Centrifuge and discard white residue. Place 2 drops of centrifugate upon piece filter paper. Add 2 drops of H2O2. Formation of a blue colour which fades rapidly confirms presence of CHROMIUM ION: Place 10 drops of the centrifugate in a test tube: add 10 drops of Amyl alcohol and 5 drops of H2O2. Shake & allow the two layers to separate. Black colouration of the upper layer confirms **CROMIUM**.

**Centrifugate**: May contain Zn(NH3)4++ ion. Divide into 2 parts: (1) Add excess 3M KOH until any ppt. Formed is dissolved. Just acidify with dilute H2SO4 .Add 5 drops of 1% CuSO4 & 5 drops of ammonium mercury thiocyanate reagent, stir. Violett ppt. Confirms Zn. (2) Add 5 drops of thioacetamide and heat in boiling water bath for After 3min. Formation of a white precipitate indicates the presence of the ZINC ion. Add 3M HCl, dropwise until solution is acidic. If precipitate is readily soluble, with only small residue, soluble, with only small residue, the presence of **ZINC ION** is confirmed.

**TABLE 5 ANALYSIS OF GROUP IIIB CATIONS**

Place 10 mg of the salt in a 10 ml test tube. Wash with 10 drops of water and discard washings. Dissolve the precipitate in concentrated HCl. Dilute the resulting solution to approximately 2 ml and divide into four portions.

**Second portion:** Test for the ferric ion by adding 2 drops of 3M NH4CNS. The appearance of a deep red colouration indicates the presence of IRON (Normally **FERRIC ION**)

**Third portion:** Test for cobalt ion as follows: Add solid NaF and stir until undissolved NaF remains. Allow 10 or more drops of alcoholic NH4CNS to run down side of test tube. (Do not shake or stir the solution) The appearance of a blue green colour at junction of two liquids confirms presence of **COBALT ION.**

**First Portion:** Add 0.2g(a heaping microspatulafull) of solid sodium bismuthate. Stir well, then centrifuge. A reddisopurple colour due to MnO4‑ in the supernatant liquid confirms the presence of **MANGANESE ION**.

**Fourth portion** Test for nickel ion as follows: Add solid NaF and stir until undissolved NaF remains. Make the solution basic with 3M NH3, then add 4 drops (or more if necessary) of dimethylglyoxime. The formation of red precipitate confirms the presence of **NICKELT ION.**

**Valence test for iron:** If iron is confirmed, apply valence test on the original solution to determine whether it is Fe++ or Fe+++. Place 5 drops of the unknown in a test tube. Acidify with 3M HCl, then add 5 drops in excess. Add 1 ml of 3M NH4CNS.The formation of a deep red colour indicates Fe+++, Fe++ gives no colour.

1. **Experiment: Identification of Anions (Cl-, Br-, I-, CO32-, HCO3-)**

**CHLORIDE:**

Heat a mixture of 5-10 mg of the original substance and an equal weight of MnO2 with 10 drops of conc. H2SO4 on water bath yellowish-green vapors of chlorine are evolved which bleaches litmus. Paper and turns KI-starch paper blue.

**BROMIDE:**

**1**. Heat a mixture of 5-10 mg of the original substance & equal amount of MnO2 with drops of conc. H2SO4 on water bath. Reddish-brown vapour of bromine are evolved which (**a**) bleaches litmus paper, (**b**) colours moist starch paper orange-red and (**c**) fluorescein paper red.

**2**. Dissolve 10 mg of the original substance in 10 drops of water, add 10 drops of 3M HCl, 10 drops of CCl4 and then add 2-3 drops of freshly prepared chlorine water (prepared just before use by adding 10 drops of 3M HCl to 10-20 mg of solid KClO3 in a test tube. Stir and let stand 5 min.before using ). Stopper the tube and shake vigorously at least 20 times. An orange- brown colour in the CCl4 (lower) layer indicates a bromide. A colorless layer indicates that more chlorine water must be added.

**IODIDE:**

Heat a mixture of 5-10 mg of the original substance and an equal weight of MnO2 with 10 drops of conc. H2SO4 on water bath. Violet vapours of iodine are evolved which turn starch paper blue. Chlorine water and CCl4 test (as under bromide): Violet coloration of CCl4 layer.

**CARBONATE / BICARBONATE:**

Treat 5-10 mg of the substance in a small test tube with 1ml of dilute H2SO4 and note whether any reaction takes place in the cold. Heat the mixture on a water bath and observe the effect produced. Colourless gas is evolved with effervescence. Gas is odor less and renders limewater turbid. This is CO2 from carbonate or bicarbonate.

**THIOSULPHATE:**

Treat 5-10 mg of the original substance with dilute H2SO4 and place the tube in the hot water bath. Colourless gas is evolved which turns K2Cr2O7 paper green and decolorize fuchsion paper. Sulphur is deposited in the solution which is yellowish.

1. **Experiment: Identification of Anions (NH4+, NO3-, NO2-)**

**4.1. Test for the ammonium radical:**

Mix about 10 mg of the substance with about 0.5 ml (about 10 drops) of dil. NaOH solution in a test tube and heat on a hot water bath. The evolution of NH3 gas detected by its odor and its action upon red litmus paper (turns blue) and upon filter paper socked in mercurous nitrate solution (causes blackening), indicates the presence of an **ammonium salt**.

**Note: NaOH is a dangerous substance because of its destructive action upon the eyes. Keep your eyes away from the test tube containing hot NaOH solution.**

4.2. Test for nitrate or nitrite:

**If ammonium is found in test 4.1:**

Add 0.5 ml water & 0.5 ml dil. NaOH to the solution from test 4.1 and evaporate down to a volume of about 2 ml and ammonia will no longer be detected by its action on red litmus paper or mercurous nitrate paper. Add 0.5 ml of NaOH and 10 mg of Devarda’s alloy (or aluminum powder or zinc dust) to this solution. If ammonia is evolved, as detected by its action upon red litmus paper (turns blue) or mercurous nitrate paper (causes blackening), the presence of a nitrate or **nitrite** indicated.

**If ammonium is absent in test 4.1:**

Add 0.5 ml of NaOH and 10 mg of Devarda’s alloy (or aluminum powder or zinc dust) to the reaction mixture from test 4.1, heat in the water bath. If ammonia is evolved, as detected by its action upon red litmus paper (turns blue) or mercurous nitrate paper (causes blackening), the presence of a **nitrate or nitrite** indicated.

**Laboratory Module B: Quantitative Analysis (Titrimetric Analysis)**

1. **Experiment: Standardization of approximately 0.1M sodium hydroxide solution against standard 0.1M succinic acid**

**(a) Preparation of 250 ml of standard 0.1M succinic acid solution.**

The molar mass of succinic acid is 118.09. Prepare 250 mL of standard 0.1M succinic acid solution by weighing out accurately about 2.9500g of A. R. succinic acid, dissolving it in water and making up the solution to 250 mL in a volumetric flask.

Strength of succinic acid = × strength of solution to be made

**(b) Standardization of approximately** **0.1M sodium hydroxide solution against standard 0.1M succinic acid.**

**Procedure:** Take 10 mL of the supplied sodium hydroxide solution in a 250 mL. conical flask with the aid of a pipette, add a drop of phenolphthalein indicator, a pink-red solution is obtained. Place the conical flask containing the sodium hydroxide solution upon a piece of unglazed white paper beneath the burette. Now add dropwise standard succinic acid from a burette to the solution in the conical flask occasionally washing with a little distilled water from a wash bottle.

During the addition of the acid, the flask must be constantly rotated with one hand whilst the other hand controls the stopcock. Continue the titration till the end point is just reached with the addition of a single drop of acid i.e., when the pink colour just disappears. Repeat the titration with two other 10 mL. portions of the NaOH solution and take the mean. Duplicate determinations should agree within 0.05 mL of each other. Present your data in the following tabular form.

**Titration of sodium hydroxide against standard 0.1M succinic acid**

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| **Titration** | **Initial burette reading (mL)** | **Final burette reading (mL)** | **Difference = volume used (mL)** | **Mean**  **(mL)** |
| **1** |  |  |  |  |
| **2** |  |  |  |  |
| **3** |  |  |  |  |

**Calculation of the molarity of NaOH solution from the relation**.

The neutralization reaction is :

CH2-COOH CH2-COONa

+ 2 NaOH = + 2H2O

CH2-COOH CH2-COONa

2 mole NaOH ≡ 1 mole acid.

Hence



or 

or =

∴ 2×SA×VA = SB×VB , ∴ SB = 

Where

SA = Molarity of succinic acid

VA = Volume of succinic acid required for neutralization.

SB = Molarity of base (NaOH)

VB = Volume of base (NaOH)

1. **Experiment: Standardization of approximately 0.1M HCl by titration with standard 0.1M NaOH**

Take 10 mL of the previously standardized sodium hydroxide solution into a conical flask with the aid of a pipette, dilute with a little water and add 1-2 drops of methyl orange indicator, a yellow solution is obtained. Add hydrochloric acid dropwise from the burette whilist constantly rotating the liquid so as to thoroughly mix the solutions. Continue the addition until the methyl orange becomes a very faint yellow. Wash the walls of the flask down with a little distilled water from a wash bottle and continue the titration very carefully by adding acid dropwise until one drop changes the colour to red. Repeat the titrations until duplicate determinations agree within   
0.05 mL of each other. Calculate the molarity of HCl from the following relation

#### Here the neutralization reaction is

NaOH + HCl = NaCl + H2O so that 1 mole NaOH = 1 mole HCl

VASA = VBSB, ∴ SA = 

Where

VA = Volume of acid (HCl) solution required for neutralization.

VB = Volume of base (NaOH)

SA = Molarity of acid (HCl)

SB = Molarity of base (NaOH)

1. **Experiment: Standardization of 0.1M sodium thiosulphate solution with standard 0.017M K2Cr2O7**
2. **Preparation of 0.1M** **Sodium Thiosulphate Solution.**

Molecular weight of Na2S2O3-5H2O is 248.19. An approximately 0.1M sodium thiosulphate solution is prepared by dissolving 6.2 g of crystalized Na2S2O3 in water and making the volume upto the mark in a 250 mL volumetric flask.

**(b) Preparation of a** **0.017M potassium dichromate.**

The molar mass of potassium dichromate is equal to 294.21. Prepare 250 mL of standard 0.02M K2Cr2O7 solution by weighting out accurately about 1.2504 g K2Cr2O7,dissolving it in water and making up the solution upto the made in a 250 mL volumetric flask.

Strength of K2Cr2O7 solution =  × Strength of the solution to be made

**(c) Standardization of 0.1M sodium thiosulphate solution with standard 0.017M K2Cr2O7**

Place 50 mL of distilled water in a 250 mL conical flask and add 1.5g of potassium iodide and 1g of pure sodium bicarbonate and shake until the salts dissolve. Add 3 mL of conc hydrochloric acid slowly when gently rotating the flask in order to mix the liquids, run in 12.5 mL of standard 0.017M potassium dichromate solution, mix the solutions well and wash the sides of the flask with a little distilled water from the wash bottle. Cover the flask with a watch glass and allow to stand in the dark for 5 minutes in order to complete the reaction. Rinse the watch glass used as cover and dilute the solution with 150 mL of distilled water. Titrate the liberated iodine against sodium thiosulphate solution contained in a burette whilist constantly rotating the liquid so as to throughly mix the solutions. When most of the iodine has reacted as indicated by the solution acquiring a yellowish-green colour, add 1 mL drops of starch solution; the colour should change to a shade of blue. Continue the addition of thiosulphate solution dropwise until one drop changes the colour from greenish blue to light green.The end point is sharp, and is readily observed in a good light aganst a white background. Continue the titration for three times, the duplicate titrations should agree within 0.05 mL. Calculate the molarity of Na2S2O3.

### **Calculation**

6KI + Cr2O72- + 14H+ → 2Cr3+ + 3I2 + 6K+ + 7H2O

3I2 + 6 Na2S2O3 → 6NaI + 3Na2S4O6

--------------------------------------------------------------

Cr2O72- + 6 Na2S2O3 + 6KI + 14H+ → 2Cr3+ + 3Na2S2O6 + 6NaI + 6K+ + 7H2O

We can find from the equation that 1 mole of Cr2O72- reacts with 6 moles of Na2S2O3

∴ 

or 6 × No of moles of dichromate = No of moles of sodium thiosulphate

or 6 × VA × SA = VB × SB

or SB = 

Here

#### VA = Volume of dichromate

SA = Molarity of dichromate

VB = Volume of thiosulphate

SB = Molarity of sodium thiosulphate

1. **Experiment: Determination of Copper by Iodometric titration**

**Principle:** Cu(II) salts react with potassium iodide in an oxidation-reduction reaction and liberates iodine. This liberated iodine is then titrated with standard sodium thiosulphate solution.

2CuSO4 + 4KI → 2CuI + I2 + 2K2SO4

I2 + 2 Na2S2O3 → 2NaI + Na2S4O6

------------------------------------------------------------------------------

2CuSO4 + 2Na2S2O3 + 4KI → 2CuI + Na2S4O6 + 2NaI + 2K2SO4

So it is seen from the above reaction that 1 mole sodium thiosulphate reacts with 1 mole copper sulphate.

Hence

1000 mL 1M Na2S2O3 solution ≡ 1 mole Cu++ ≡ 63.54g Cu2+

∴ 1 mL 1M Na2S2O3 ≡ 0.06354g Cu2+

**Procedure:** Pipette 10 mL of the supplied copper solution into a 250 mL conical flask and add 2 mL of a 10% KI solution. Titrate the liberated iodine with 0.1M sodium thiosulphate solution until the brown colour of the iodine fades, Then add 5 drops of starch solution which will make the solution blue and continue the addition of sodium thiosulphate solution until the blue colour commence to fade. Then add 2 mL of 10% potassium or ammonium thiocyanate solution ; the blue colour will instantly become more intense. Complete the titration as quickly as possible. The blue colour removes instantly at the end point. At the end point the precipitate/solution possess a pale pink colour. Repeat the titration at least three times. Duplicate titrations should agree within 0.05 mL.

### **Calculation**

2Cu2+ + 4I− → 2CuI + I2

2 mol 1 mol

1 mol I2 ≡ 2 mol Na2S2O3

2 mol Cu2+ ≡ 1 mol I2 ≡ 2 mol Na2S2O3

1 mol Na2S2O3 ≡ 1 mol Cu2+

∴ 1000 mL 1M Na2S2O3 ≡ 1 mol Cu2+ = 63.54g Cu2+

1 mL 1M Na2S2O3 ≡ 0.06354g Cu2+

**Laboratory Module C: Quantitative Analysis (Energetics)**

1. **Experiment: Quantitative exercise to find the molar enthalpy change when ammonium chloride dissolves in water.**

**Theory:** If a hot body is kept in contact with a cold body, the heat released by the hot body is equal to the heat absorbed by the cold body i.e. heat absorbed = heat released, which is called the principle of calorimetry. Again the heat absorbed or evolved by dissolving 1 mole of solute in excessive amount of solvent is called heat of solution. The per degree Celsius change of temperature for production of 1 Kg solution of ammonium chloride by dissolving 1 mole of ammonium chloride in excess amount of water is 1 kilocalorie, i.e. 4.2 kJ. Therefore, if the total change of temperature to produce 1 mole of ammonium chloride solution is t0C, the heat of solution will be 4.2 × t kJ. Again, for producing 100g ammonium chloride solution by dissolving 0.1 mole ammonium chloride, the heat change = 0.42 × t kJ.

Therefore, heat of solution ammonium chloride = (0.42 × t)/0.1

= (0.42 × t × 10) kJ

**Procedure:**

i) The molecular mass of ammonium chloride (NH4Cl) is 53.49 g/mol. So, 0.1 mol of mashed ammonium chloride i.e. the weight of 5.35 g of mashed ammonium chloride is to be taken.

ii) 94.65 g of water is taken in 250 ml beaker. (If 5.35 g of ammonium chloride is mixed to 94.65 g of water, the solution becomes 100 g).

iii) A thermometer hanging from the stand which is tightened by a thread is dipped in water. In this condition, the temperature of water (t10 C) is recorded.

iv) 5.35 g ammonium chloride is added to water and dissolves after stirring by glass rod. As soon as the ammonium chloride completely dissolve in water, temperature of solution (t20 C) is recorded.

**Calculation:**

Initial temperature of water = t10C

Temperature of solution = t20C

Change of temperature, t0C = (t1 – t2)0C

Therefore, heat of solution of dilute of ammonium chloride,

ΔH = (0.42 × t × 10) kJ/mol

1. **Experiment: Quantitative exercise to find the molar enthalpy change when sodium thiosulphate dissolve in water.**

**Theory:** If a hot body is kept in contact with a cold body, the heat released by the hot body is equal to the heat absorbed by the cold body i.e., heat absorbed = heat released, which is called the principle of calorimetry. Again, the heat absorbed or evolved by dissolving 1 mole of solute in excessive amount of solvent is called heat of solution. The per degree Celsius change of temperature for production of 1 Kg solution of sodium thiosulphate by dissolving 1 mole of sodium thiosulphate in excess amount of water is 1 kilocalorie, i.e., 4.2 kJ. Therefore, if the total change of temperature to produce 1 mole of sodium thiosulphate solution is t0C, the heat of solution will be 4.2 × t kJ. Again, for producing 100g sodium thiosulphate solution by dissolving 0.1 mole sodium thiosulphate, the heat change = 0.42 × t kJ.

Therefore, heat of solution sodium thiosulphate = (0.42 × t)/0.1

= (0.42 × t × 10) kJ

**Procedure:**

i) The molecular mass of sodium thiosulphate (Na2S2O3) is 158.11 g/mol. So, 0.1 mol of mashed sodium thiosulphate i.e., the weight of 15.81 g of mashed sodium thiosulphate is to be taken.

ii) 84.19 g of water is taken in 250 ml beaker. (If 15.81 g of ammonium chloride is mixed to 84.19 g of water, the solution becomes 100 g).

iii) A thermometer hanging from the stand which is tightened by a thread is dipped in water. In this condition, the temperature of water (t10 C) is recorded.

iv) 15.81 g sodium thiosulphate is added to water and dissolves after stirring by glass rod. As soon as the sodium thiosulphate completely dissolve in water, temperature of solution (t20 C) is recorded.

**Calculation:**

Initial temperature of water = t10C

Temperature of solution = t20C

Change of temperature, t0C = (t1 – t2)0C

Therefore, heat of solution of dilute of sodium thiosulphate,

ΔH = (0.42 × t × 10) kJ/mol