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# Spectrophotometric determination of micro amount of mercury (II) using an (Azo) derivative in the presence of surfactant: Study of thermodynamic functions and their analytical application

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***Abstract*---**Imidazole azo ligand (DPIDA), one of the azo compounds, was prepared and tested in the sample. In addition to the regular azoth protocol. Because of their ability to form stable complexes with metal ions in some oxidation states, imidazole azo ligands are important heterocyclic azo ligands in coordination chemistry. The complexation of the imidazole azo ligand (DPIDA) was accomplished by adding 4,5-diphenyl imidazole and N1, N1-dimethylbenzene1,4-diamine dihydrochloride. Measurements of 1HNMR, IR, UV-Vis, and molarity Their conductivity demonstrates their octahedral geometry with a bidentate ligand coordinated from an (N3) atom of the imidazole ring and one nitrogen atom of the azo group. The thesis also looked at using this reagent was used to do spectroscopy with the mercury (II) ion in binary water solution, and it was revealed that the reagent is difficult with the ion and has the highest sensitivity. The exact wavelength of absorption is (554) nm (pH = 6). Even though the pH was higher in the concentration range of (0.1 – 6.0 g /mL) with adherence to Beer's law, it was revealed to be a dynamic mercury pair reagent with the stability of more than (24 hours). The influence of reagent concentration, reaction duration, masking agent, Works investigated sophisticated stoichiometry with reagent-metal molar ratios, as well as various parameters such as the effect of cations and anions on absorption, as well as the effect of ionic strength and temperature (the constant changes) According to the mallard equation, the metal reincarnationthe proportion is (1:2), the ability to absorb substances at a molar level (3.4x104 L. $mol^{-1}$. cm $^{-1}$). Detection and evaluation limits are (0.019 g/mL) and (0.627 g/mL)

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correspondingly. As a part of the complex the preparation of the solid, as well as some of its physical properties including solubility and molar conductivity, as well as the moment at which the complex melts away, were all investigated. Spectroscopic methods [FT.IR., UV-Vis] were used to classify both substances. As compared to the free reagent, the bathochromic transition is visible in (UV-Vis) absorption spectra, demonstrating the exactness and exactitude of the tool used to make educated guesses about the element mercury the standard deviation from the mean (RSD per cent), although the relative error (E per cent) differed from (4.011 per cent, 0.803 per cent). The thermodynamic functions (ΔH,ΔG,ΔS) were determined as well by investigating the influence of temperature. The approach was extended to both models of the environment and industry, producing results that were surprisingly precise and informative.

***Keywords*---**Azo dye (DPIDA), Mercury (II), Spectrophotometry.

## Introduction

Azo dyes are a rapidly expanding family of organic compounds within the known dye families due to their wide range of applications in textile dyeing, metallochromic markers, pharmaceuticals, cosmetics, inkjet printing, and food coloring [1,2]. Azo dyes having a heterocyclic moiety have also made strides in electronics, such as storage, optoelectronic, and photoswitchable devices [3]. Furthermore, as compared to simple aromatic chemicals [4], heterocyclic azo dyes have a significant batho effect.

Furthermore, azo dyes dependent on heterocyclic compounds were discovered to have potent antioxidant, anticancer [5,6], antituberculosis, antibacterial, insecticide, antidiabetics, antiviral, and other chemotherapeutic characteristics [7,8,9,10], the shape of a stable five-member ring with each ion passing through the (N3) atom of the imidazole ring and one of the nitrogen atoms of the azo group [8], In addition to the contribution of the imidazole molecule in the preparation of certain ligands that have a conjugated system that enhances their stability and follows the colour shift of them before and after coordination with metal ions [9], Since most azo dyes have acid-base properties and a fixed number of isosbestic points (which reflect the number of equilibriums in such an azo dye)[10,11], they are used as acid-base indicators.

Mercury may occur in a variety of ways [12], all of which are poisonous. Metallic mercury is the most common form of mercury found naturally in the atmosphere [13]. Microorganisms and natural processes may convert these mercury types from one to the other [14]. The most popular source of mercury released by these natural processes is methyl mercury. This compound is of special interest because it has the potential to bioaccumulate and bio amplify in the food chain [15].

Acute organic mercury intoxication causes blindness, deafness, tremors, sleeplessness, memory loss, neuromuscular abnormalities, headaches [16],

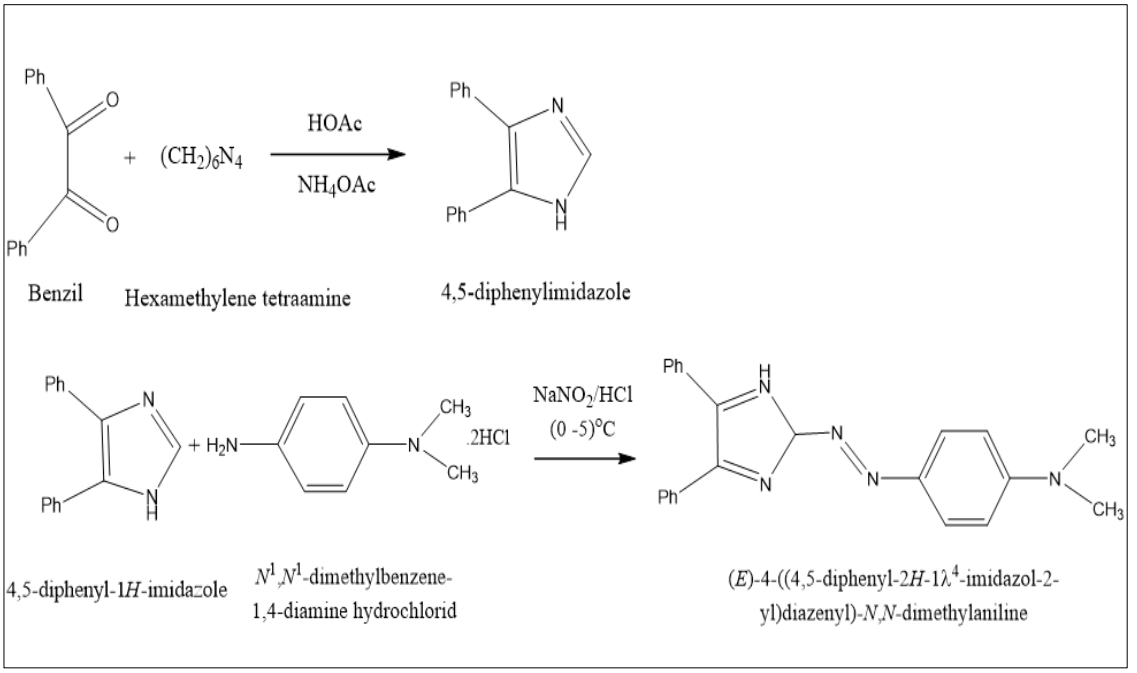
sluggish visual and slow motor nerve activity, and a decline in cognitive function [17]. When inorganic mercury compounds are eaten, they cause a metallic taste in the mouth, nausea, vomiting, and severe stomach discomfort [18].

Chronic mercury exposure can result in a number of disorders, including acrodynia (pink disease), neurological diseases, Hunter-Russell syndrome, crazy hatters' disease, leukemia, and Minamata sickness [19]. In comparison, some mercury compounds are typically non-toxic and have been used as medications, such as to treat syphilis. Human activity contributes a substantial quantity of mercury to the environment, mostly through chlor-alkali plants, plastics industries, electrical, paint, and pharmaceutical sectors.[20].

**Experimental**
**Preparation of (DPIDA) ligand**

A/ preparation of starting materials (4,5 Diphenyl imidazole)
The alpha Di carbonyl condensation method was adopted with ammonia and aldehyde to form a derivative imidazole, and the precursor was prepared with benzyl with both hexamethyl tetra amine and amino acetate in the presence of ice acetic acid to obtain the product, it was added (50 ml of ice acetic acid to a mixture of (2.1 g,0.01 mole) of benzene, (0.015 moles, 21 hexamethyl tetraamine) and (0.07 mole 3.7 g) of ammonium acetate in a100 ml) glass flask and raised the mixture for a while. One hour with continues stirring, then transfer the solution after cooling it to a Baker s capacity (1 L) and dilute by adding (400 ml) of distilling water and after adding ammonium hydroxide precipitate the imidazole derivative the To clear the residue, it was filtered and rinsed with distilled water.excess base and dried in the air and was recrystallized from ethanol and obtain on a white precipitate it dried and its melting point was measured (229 - 230 C) and the yield ratio was (75 %).

B/Coupling reaction between 4,5 diphenyl imidazole and N 1, N 1 dimethyl benzene 1,4 diamine dihydrochloride Two and thirty-one gram of N1, N1dimethylbenzene1,4 diamine dihydrochloride was dissolved in twenty-five ml of distilled water then one ml of hydrochloric acid added gradually to this solution which cooled in an ice bath 0 5 o C the formation of diazonium salt occurred by addition the solution of sodium nitrate which prepared by dissolved0.70 gm of it in Drop by drop, 10 ml of distilled water with stirring, This solution left in the ice bath for 30 minute then coupled with alcoholic solution of 4,5diphenyl imidazole which prepared by dissolving 2.21 gm of imidazole derivative and 0.44gm of sodium hydroxide in 25 ml of ethanol, the orange precipitate was appeared after the completing of addition, f and dried then recrystallize d from ethanol yield percentage 71 % as shown in scheme 1.




Scheme 1: Preparation of (DPIDA) ligand

**Materials and Methods**

This study contains the following materials and methods. All of the instruments and supplies for scientific investigationand the preparation's ingredients and solutionsare of the highest quality.

**Standard Solution Preparation:**

- 1-Mercury (II) solution (1000 g/mL): - was made by dissolving (0.135g) of [$HgCl_2$] in 100 mL of purified water.
- 2-Solution of sodium hydroxide (NaOH) (0.1M): This solution was made by adding 100 mL of distilled water to 0.4 g of sodium hydroxide.
- 3-Hydrochloric acid solution (0.1M): This solution was made by diluting 0.40 mL of concentrated hydrochloric acid (38 percent, 1.19 g/mL) in 50 mL of purified water.
- 4-Solution for reagents (DPIDA) (0.001 M): -was made by dissolving the necessary weight (0.091g) in absolute ethanol and filling the remaining capacity (250mL) with ethanol.

**Interferences**

Cation solutions of ($Cd^{2+}$, $Ni^{2+}$, $Fe^{3+}$, $Ba^{2+}$, $pb^{2+}$, $Cr^{3+}$, $Co^{2+}$) ions (0.1mg/mL) were made by means of dissolution (0.163g) of $CdCl_2$, (0.221g) of $NiCl_2$, (0.226g) of $FeCl_3.6H_2O$, (0.190g) of $BaCl_2$, (0.159g) of $pbCl_2$, (0.403g).

**Preliminary study**

A prepared Mercury (II) solution containing 100 g/ml of mercury was placed in a test tube, and 1ml of a ligand (DPIDA) solution (0.001M) was added dropwise with shaking to the test tube, noting the color formation of a precipitate. Drops of hydrochloric acid (0.1 M) were added to one part of this mixture, and drops of NaOH (0.1 M) were added to the other part.

**Results and Discussion**

Figures 2 and 3 show the absorption spectra of the reagent and mercury (II) complex. A considerable rise in absorbance and bathochromic shift are both associated with complex formation in the reagent solution spectrum (max=476 nm), whereas the mercury (II) complex produced at pH=6 has an absorption maximum (max=554 nm).

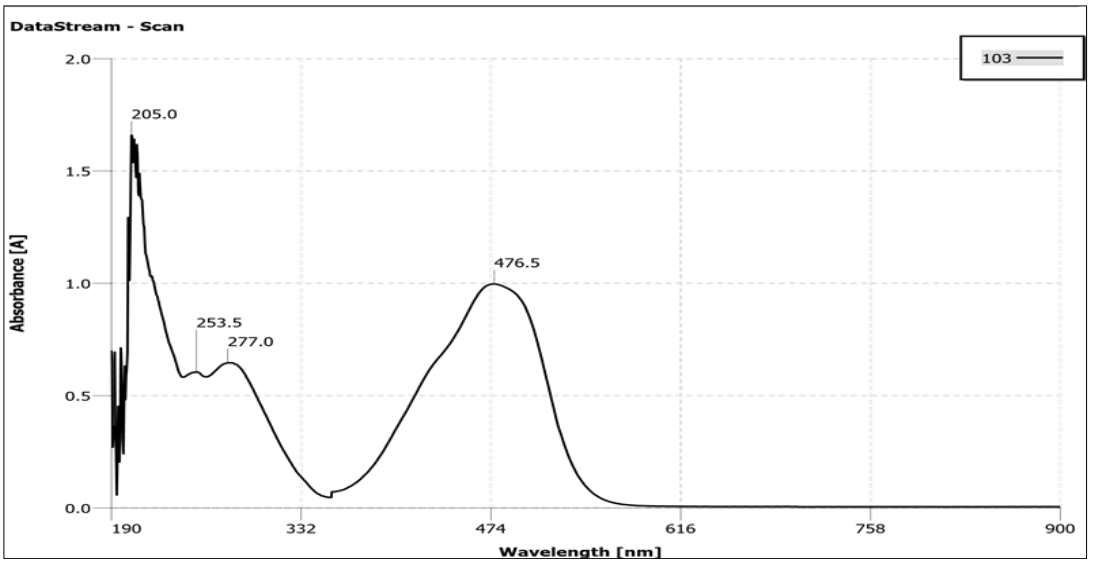



Figure 1. The reagent's molecular spectrum (DPIDA)

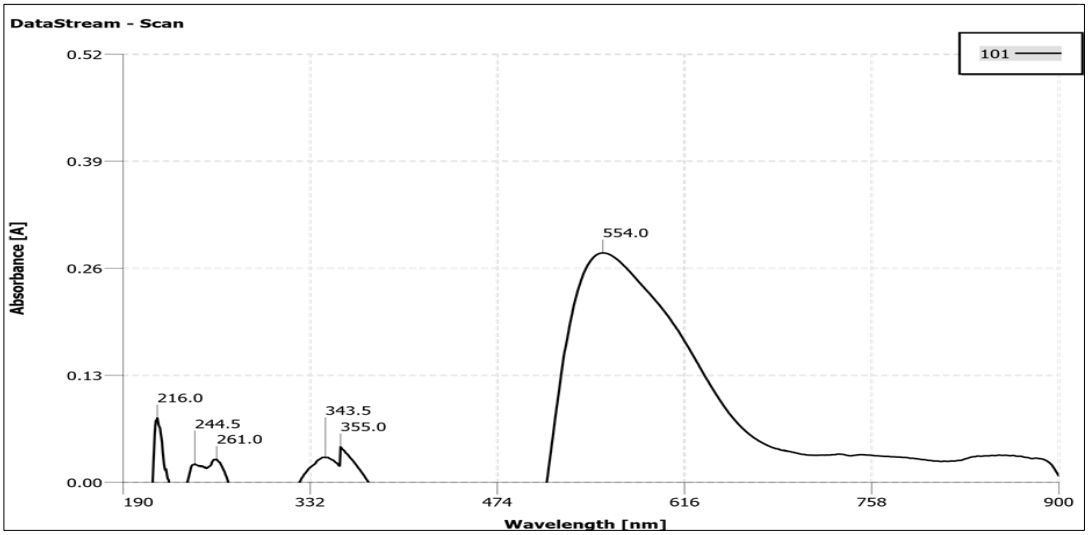



Figure 2. The Hg (II) complex (DPIDA) spectral profile

**The effect of adding surfactants (micelle medium)**

Among the surfactants had a great impact on spectrophotometry and chromatography, especially in the development of new methods for estimating ions. Numerous studies have given the issue of the formation of(metal-ligand) complexes in the systems containing the (Micellar) system, the great importance for reasons including the formation of the complex may be more stable in these systems, so the effect of adding (Triton X-114) was studies. and (Triton X-100) and (Tween-80) with different concentrations in the absorbance of mercury (II), as shown in table (1).

Table 1
Absorption values of mercury complex in different mediums

| Concentration of surfactant (v/v %) | Absorption values of Hg complex in (TWEEN-80) | Absorption values of Hg complex in (TRITON-114) | Absorption values of Hg complex in (TRITON-100) |
|---|---|---|---|
| 1% | 0.549 | 0.490 | 0.412 |
| 2% | 0.601 | 0.514 | 0.450 |
| 3% | 0.615 | 0.430 | 0.477 |

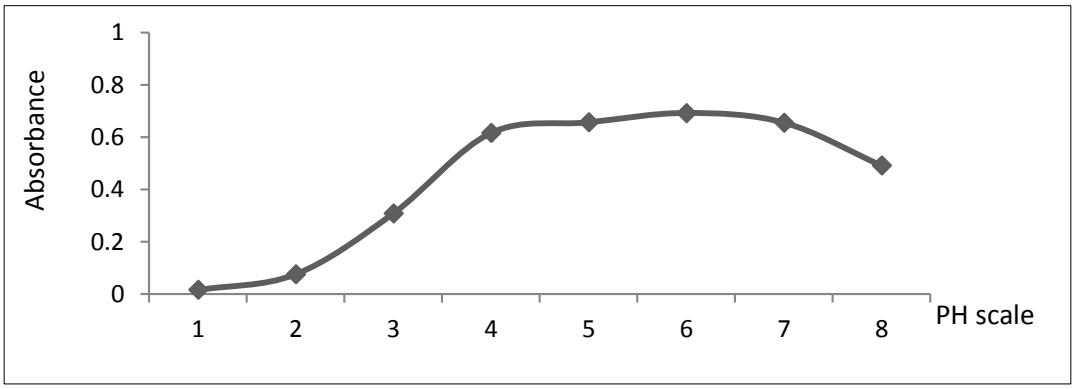



Figure 3: - Effect of pH value

**Optimization of Reaction Conditions**
**Effect of PH value**

How pH affects things The final pH of each solution was calculated using a pH-meter and the absorbance at (560 nm) at 20 Co was measured using a standard amount of Mercury (II) and reagent (DPIDA). The buffered solution included a standard amount of Mercury (II) and reagent (DPIDA).

Absorbance increased smoothly from (1.0 – 8.0) to (10.0), but then rapidly decreased (see Figure 3). (above pH 6.0), In these conditions, an increase in the reagent's sensitivity at this pH value might explain the rise in mercury complex solution absorbance.
The Time Effect: The Impact of the Passage of Time on the Stability of Complex Systems When the circumstances are ideal, as shown in table (1), the reagent reacts with the ion to form a complex, and these results show the composition of the mercury (II) ion, they remain stable for 90 minutes after the experiment starts.

Table 1
The effect of time

| Time/Min. | Abs. |
|---|---|
| 1 | 0.661 |
| 10 | 0.685 |
| 20 | 0.698 |
| 30 | 0.712 |
| 40 | 0.707 |
| 50 | 0.705 |
| 70 | 0.702 |
| 100 | 0.698 |
| 24h | 0.653 |
| 48 h | 0.653 |

Table 2 shows the effect of reagent concentration on mercury complex absorbance at (pH=6) as shown in the experiment. The results showed that when the reagent concentration was raised, the absorbance increased.

Table2
How reagent amount affects reaction rates

| **Volume Conc. of L. x 10-3** | **0.25** | **0.5** | **0.75** | **1.0** | **1.25** | **1.5** |
|---|---|---|---|---|---|---|
| **Absorbance** | **0.222** | **0.309** | **0.350** | **0.430** | **0.115** | **0.101** |

**Effect of Sequence**

To study the sequence of the reaction content in a complex absorbance, the three arrangements of addition was depending, and the result given in table (3)

Table 3
The effect of sequence

| Sequence of Number | Sequence of Addition | Abs. of Cu Complex |
|---|---|---|
| 1 | M+L+PH | 0.677 |
| 2 | L+M+PH | 0.621 |
| 3 | M+PH+L | 0.562 |

M = mercury ion, L = ligand, pH = hydrogen ion functions the results in table (3) show that the first arrangement is the best one, while the other sequence results in a decrease in complex absorbance, which may be attributed to the effect of acid, base inions with metal, so the addition of the first sequence was relied on to determine the mercury ion complex in this method.

**Temperature Affects**

The influence of temperature variation a comprehensive investigation was conducted on complicated structure. The findings of this investigation are shown in Table (4), which demonstrate that the complex's absorption values on the creation of a complex maximum and give the optimum hue and saturation at 30 C0, after which the absorption values decline, presumably due to a diminution of complex's stability. The complex should be prepared at a temperature no greater than 30 degrees Celsius.

Table 4
Temperature Fluctuation

| Temperature /c | Abs. of $Hg^{2+}$ complex |
|---|---|
| 10 | 0.673 |
| 20 | 0.679 |
| 30 | 0.684 |
| 40 | 0.621 |
| 50 | 0.591 |
| 60 | 0.537 |

**Calibration Curve**

Beverages with low concentrations obeyed in the concentration range (0.1 - 6.0%) with an absorbance of (3.4x104Lmol cm 1), Figure (3) showed the calibration curve of mercury ion and the table (4) provided analytical data to quantify mercury ion by reagent (DPIDA).

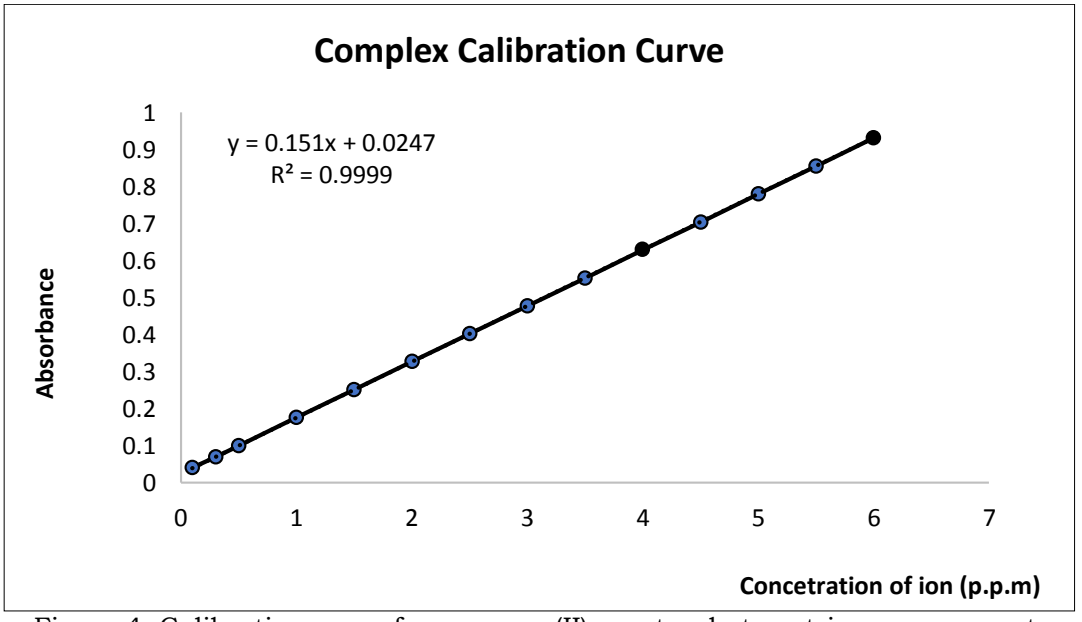



Figure 4. Calibration curve for mercury (II) spectrophotometric measurement

Table 5
Data collected for mercury (II) analysis

| Analytical Data | Value |
|---|---|
| linear equation | Y=0.151x |
| Linear range [M] | 0.1- 6.0 |
| Detection limit(µg/mL) [a] | 0.019 |

| | |
|---|---|
| Limit of quantification (µg/mL) [b] | 0. 627 |
| Molar Absorptivity ($Lmol^{-1}cm^{-1}$) | $3.4\times10^4$ |
| Correlation coefficient | 0.9989 |
| $\lambda_{max}$ | 554nm |
| Temp. | $30^0C$ |
| Time | 90 min |
| Colour of product | purple |

LOD detection = (SD/S) *3.3 (Limit of quantification) LOQ = (SD/S) *10, where SD is the standard deviation and S is the calibration curve's slope.

**Stoichiometry and Formation Constant Estimation**

1- Method of the Mole Ratio
The mole ratio approach and the addition of Job's methodof using a known and constant concentration from continuous variations were chosen to study the mercury (II) ion (1.994 x 10-5 M) with the increasing composition of the complex formed, with results shown in concentration from reagent (DPIDA) (1.994x10-5 - Figs 7 and 8). Both methods revealed that the ratio of 13.958x10-5M), The method demonstrates that mercury ion metal ion to reagent molecules (M:L) was (1:2) at pH=6, forming a (1:2) (metal -L) complex with reagent...

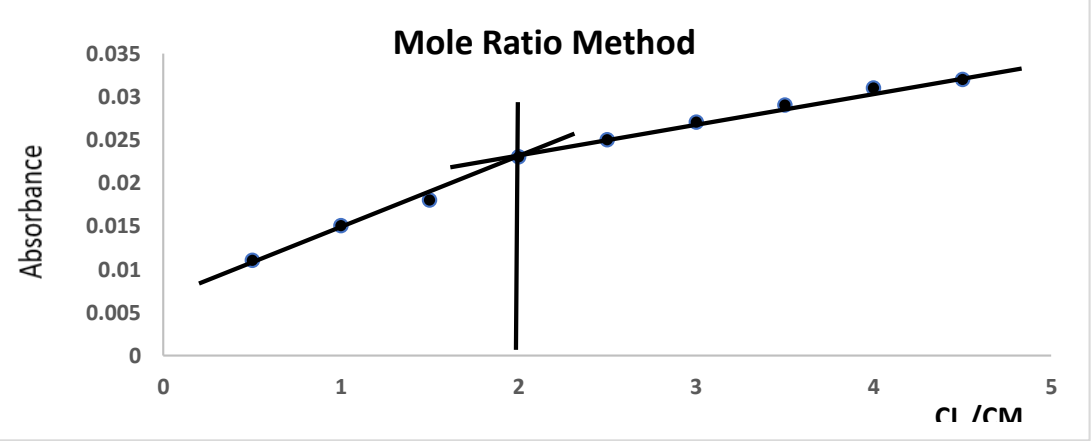



Figure 5. Mole ratio method

The equilibrium reaction of the colored complex was utilized to derive the stability constant using the mole ratio technique. In the following table, the calculated results are shown (6).

$$M^{+n} + nL \leftrightarrow MLn \dots\dots\dots\dots\dots\dots (4-3)$$
$$\alpha + nc\alpha \leftrightarrow (1-\alpha)c$$

$$K = \frac{[MLn]}{[M^{+n}][L]} \dots\dots\dots\dots\dots (5-3)$$

$$K = \frac{(1-\alpha)c}{\alpha c(n\alpha c)^n} \dots\dots\dots\dots\dots (6-3)$$

$$K = \frac{1\_\alpha}{n^n\alpha^{n+1}c^n} \dots\dots\dots\dots\dots (7-3)$$

$$\alpha = \frac{Am\_As}{Am} \dots\dots\dots\dots\dots (8-3)$$

Where (Am) denotes to the greatest extent possibleand (As) denotes absorption at the stoichiometric concentration.

Table 6
The complex's equilibrium stability constant

| Complex | Am Value | As Value | degree of dissociation α | Stability Constant K |
|---|---|---|---|---|
| [Hg (DPIDA)$_2$] | 0.023 | 0.032 | 0.281 | 1.3x10$^8$ |

The results in Table (6) demonstrate that the complex has great stability, allowing the ligand (DPIDA) to be used in the spectrum estimate of mercury ions.

**Job's approach**

In this procedure, a combination of various same concentrations in different amounts of solution (1x10$^{-3}$ M) from both the ion ($Hg^{+2}$) and the ligand was prepared.

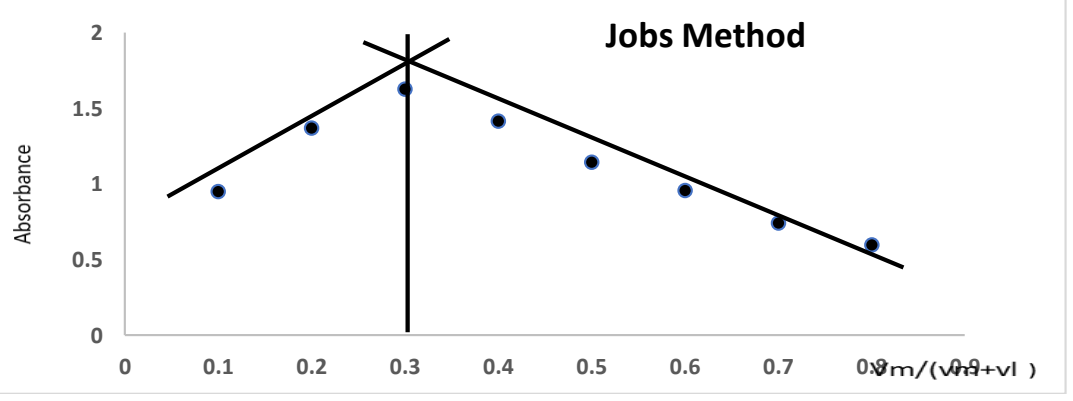



Figure 6. Job`s method of continuous variations

**Addition of buffer solutions and their results**

To investigate how a buffer solution affects things type on the absorbance of the mercury (II) complex, They experimented with four different kinds of buffer solutions, and the difference in the absorption values of the mercury (II)ion

complex with the reagent (DPIDA) making use of the ideal circumstances is shown in Table (7).

Table 7
A study on the impact of buffer solutions

| Type of emulsifying agent | Amount of mercury absorbed by complex |
|---|---|
| Phosphate | 0.629 |
| Citric acid | 0.420 |
| Acetate | 0.617 |
| Ascorbate | 0.579 |
| Absorbance value without addition buffer solution of Hg complex = 0.677 | |

As a consequence of these findings, the mercury (II) complex absorbs less light than hydrochloric acid and dilute sodium hydroxide, therefore the acidic function was modified to provide excellent sensitivity and accuracy for mercury (II) ion measurement by utilizing simply an acid or a base.

Strong ions have a different effect. To find out how much of an impact adding different amounts of Na2SO4 and NaNO3 to the complex solution has, the results are provided in Table (8).

Table 8
Ionic strength solutions and their effects

| Added salt | The molar concentration of added salt | Absorption after adding salt | Added salt | The molar concentration of added salt | Absorption after adding salt |
|---|---|---|---|---|---|
| NaNo3 | 0.0005<br>0.005<br>0.05<br>0.5 | 0.602<br>0.621<br>0.647<br>0.673 | Na2SO4 | 0.0005<br>0.005<br>0.05<br>0.5 | 0.622<br>0.668<br>0.674<br>0.683 |
| Absorption before adding salt to the Hg complex at 560 nm=0.677 | | | | | |

**Foreign Ion Effect**

To investigate the interference effect, a specific concentration of several cation and anion solutions employed as though there were alien ions mixed with an iron solution:
The cation effect A- The ion effect B- The ion effect

Table 9
Cation's Effect

| Interfering ion | Ion salt formula | **4**µg/ml | | 10µg/ml | |
|---|---|---|---|---|---|
| | | The absorption values after adding the ions | Mistake percentage E% | The absorption values after adding the ions | Mistake percentage E% |
| _ | _ | 0.677 | _ | 0.677 | _ |
| $Fe^{3+}$ | FeCl3.6H2O | 0.406 | - 40.029 | 0.486 | - 28.212 |
| $Cu^{2+}$ | CuCl2.6H2O | 0.061 | - 90.098 | 0.084 | - 87.592 |
| $Cr^{3+}$ | CrCl3.6H2O | 0.738 | 9.010 | 0.593 | - 12.407 |
| $Co^{2+}$ | CoCl2.6H2O | 0.169 | - 75.036 | 0.075 | - 88.921 |
| $Ba^{2+}$ | BaCl2.2H2O | 0.649 | - 4.135 | 0.652 | - 3.692 |
| $Ni^{2+}$ | NiCl2.6H2O | 0.043 | - 93.648 | 0.192 | - 71.639 |
| $Cd^{2+}$ | CdCl2 | 0.379 | - 44.017 | 0.131 | - 80.649 |

Table10
Inion's impact

| Interfering ions | Ion Formulary Structure | **4**µg / ml | | 40µg/ml | |
|---|---|---|---|---|---|
| | | The absorption values after adding the ions | Mistake percentage E% | The absorption values after adding the ions | Mistake percentage E% |
| _ | _ | 0.677 | _ | 0.677 | _ |
| $Cr2O7^{2-}$ | K2Cr2O7 | 0.885 | 30.723 | 0.716 | 5.760 |
| $SCN^{1-}$ | KSCN | 0.555 | - 18.020 | 0.541 | - 20.088 |
| $CO3^{2-}$ | K2CO3 | 0.712 | 5.169 | 0.309 | - 54.357 |
| $SO4^{2-}$ | K2SO4 | 0.434 | - 35.893 | 0.557 | - 17.725 |
| $IO3^{2-}$ | KIO3 | 0.565 | - 16.543 | 0.636 | - 6.056 |
| $CrO4^{2-}$ | K2CrO4 | 0.464 | - 31.462 | 0.445 | - 34.268 |
| $Br^{1-}$ | KBr | 0.552 | - 18.463 | 0.493 | - 27.178 |

Some ions were chosen to explore the influence of Hg (II) ion interferences (Table 9). Some ions enhanced absorbance while others lowered absorbance, which was owing to rivalry between these in order to form a compound with Hg (II), which reduced competition and increased the sensitivity of this technique to the Hg (II) ion. Hg-specific but yet being delicate response (II).
The ability to be discerning in the response is possible validated by applying appropriate agents of omission and substitution.

**Precision and Accuracy of the Methods Explained**

The implemented method's accuracy and precision were measured in terms of recovery and relative standard deviation (RSD percent). The table shows the recovery and RSD percent results (11).

Table 11
Studies on precision and accuracy

| Conc.of $Hg^{+2}$ present[M] | Conc. of $Hg^{+2}$ found[M] | RSD% | Recovery% | Error% |
|---|---|---|---|---|
| $9.971\times10^{-6}$ | $9.752\times10^{-6}$ | 1.651 | 100.609 | -0.609 |
| $19.941\times10^{-6}$ | $19.473\times10^{-6}$ | 1.004 | 100.596 | -0.596 |
| $29.911\times10^{-6}$ | $29.693\times10^{-6}$ | 0.803 | 100.105 | -1.057 |

Table (11) shows that the devised approach was exact, with a relative standard deviation of 0.4 percent.

The influence of temperature on the stability constant of the [Hg(DPIDA)2] complex:

Table 12
The levels of stability remain constant as temperature rises

| T (C) | T (k) | $\alpha$ | $k \times 10^8$ |
|---|---|---|---|
| 10 | 283.15 | 0.299 | 1.0146 |
| 15 | 288.15 | 0.289 | 1.1282 |
| 20 | 293.15 | 0.282 | 1.1937 |
| 25 | 298.15 | 0.277 | 1.2163 |
| 30 | 303.15 | 0.271 | 1.2739 |

12 demonstrated that temperatures had a minimal influence on the complex's stability.

Thermodynamic Function of the Complex:

Thermodynamic functions H, G, and S were computed, and the results are shown in Figure (6) and Table (12).

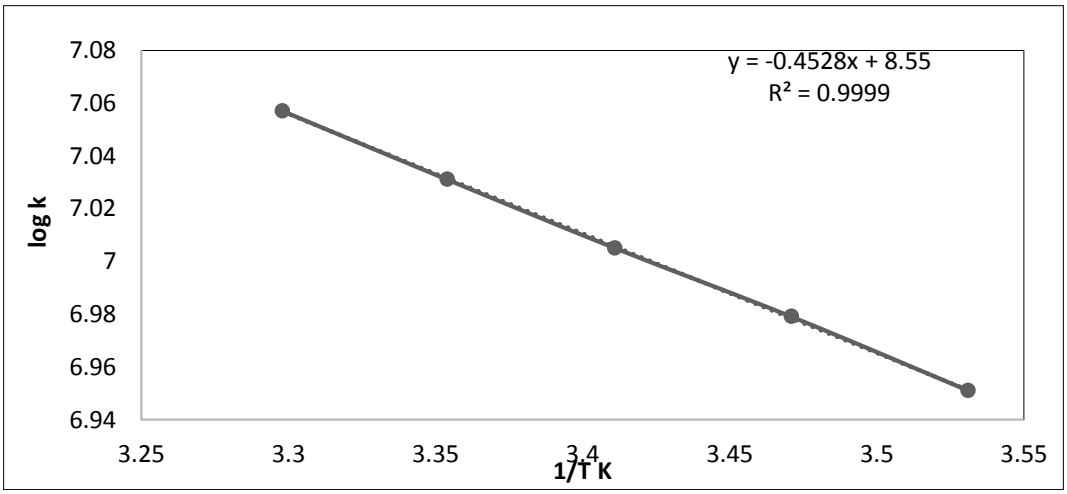



Figure 7 depicts the relationship between Log K and 1/T values for the mercury (II) complex.

The influence of temperature on the thermodynamic function of the mercury (II) complex is shown in Table 13.

| T(K) | (1/T) | Log K | ΔH Kj. Mol⁻¹ | ΔG Kj. Mol⁻¹ | ΔS Kj. Mol⁻¹.k⁻¹ |
|---|---|---|---|---|---|
| 283.15 | 3.531 | 6.951 | 8.669 | - 43.397 | 0.1838 |
| 288.15 | 3.470 | 6.979 | | - 44.416 | 0.1842 |
| 293.15 | 3.411 | 7.005 | | - 45.324 | 0.1841 |
| 298.15 | 3.354 | 7.031 | | - 46.145 | 0.1838 |
| 303.15 | 3.298 | 7.057 | | - 47.033 | 0.1837 |

A positive value of enthalpy stated that the reaction was endothermic, which may be observed by increasing the temperature. The complex formation will be reduced, and the reaction was spontaneous based on the negative sign of free energy.

## FT-IR Spectroscopy of Ligands and Complexes Research

To better understand the results of the FT-IR analysis and absorption frequencies of the complex and reagent, have a look at the figures and tables (DPIDA).

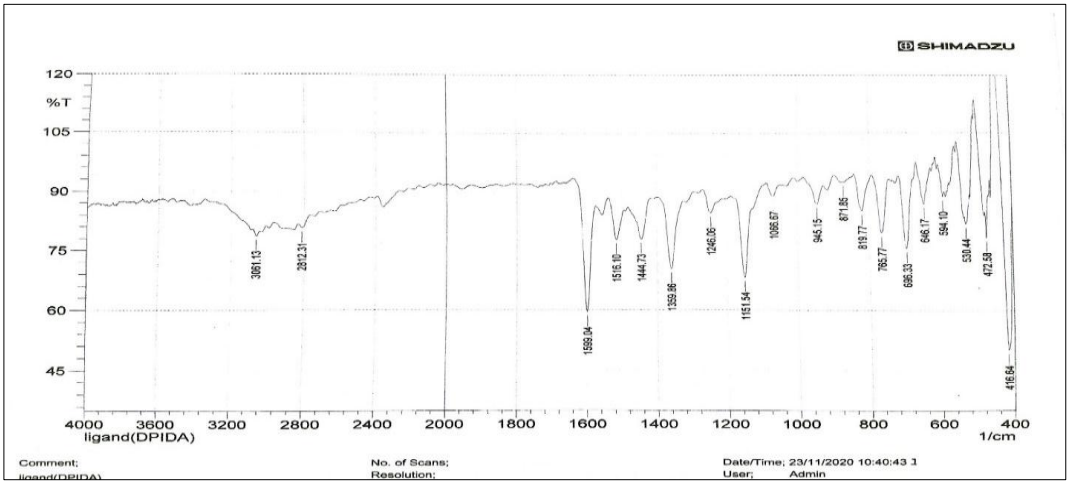



Figure 8: - ligand FT-IR spectral analysis

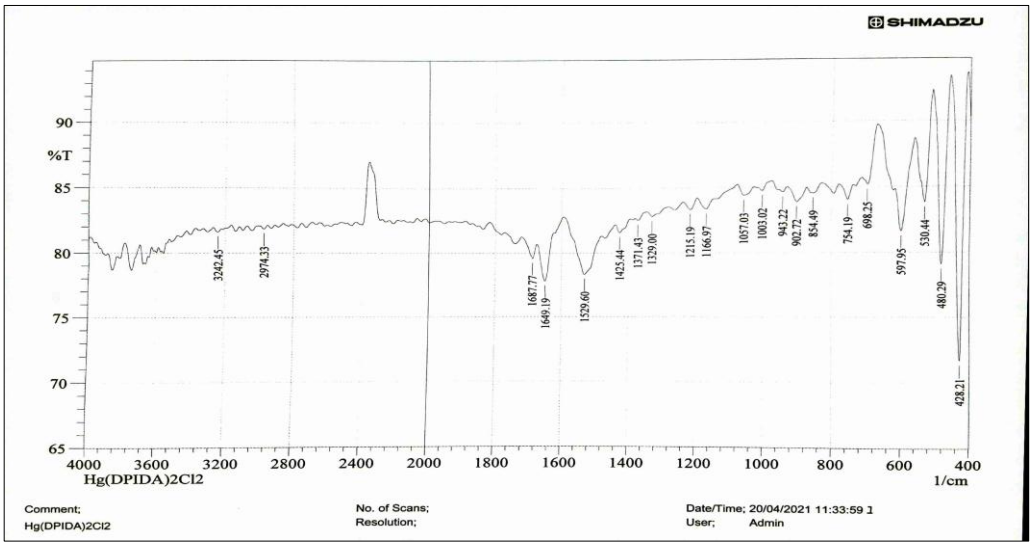



Figure 9: - ligand FT-IR spectroscopy

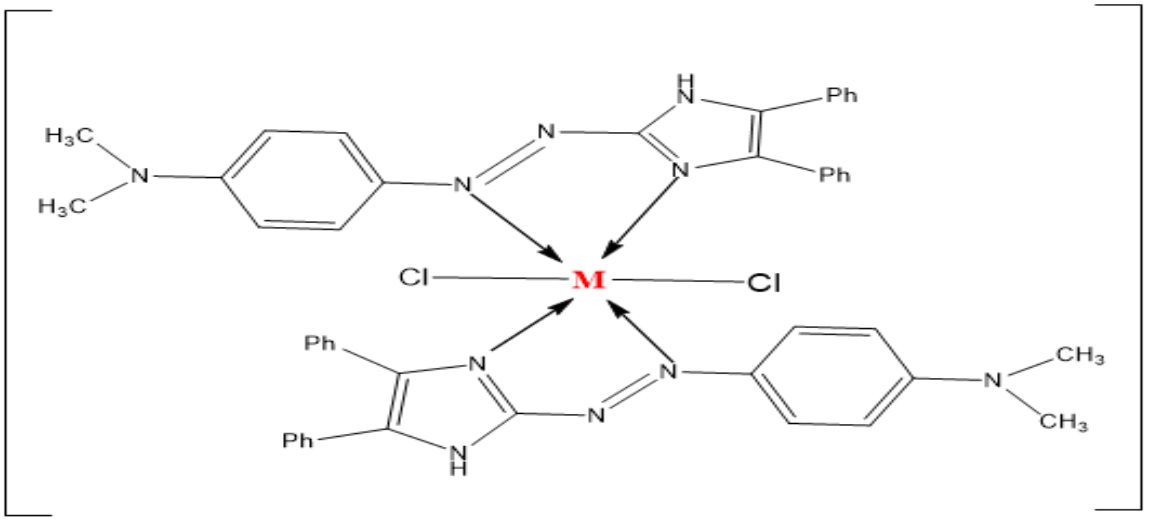
Table 14
Absorptivity and fluorescence measurements using Fourier transform infrared imaging (DPIDA)

| Compound | DPIDA | [Hg(DPIDA)2C12] |
|---|---|---|
| Compound | DPIDA | [Hg(DPIDA)2C12] |
| (N-H) | 3061 m | 3242 w |
| (C-H) Ar | 2812 m | 2974 w |
| (N=N) | 1444 m | 1425m |
| (C=N) | 1599 m | 1529 m |
| (C=C) | 1516 m | 1529 m |
| (C-O) | 1200 m | 1215 s |
| (M-O) | ..... | 480 s |
| (M-N) | ..... | 597 s |

**Compound Suggestion Chart**

The complicated construction as seen in the accompanying diagram (10) is owing to analysis of the FT-IR spectra and stoichiometry of the Job and Mole ratio approaches.

**S=strong, M=medium, W=weak**




M= Mn(II), Co(II), Ni(II) , Cu(II), Zn(II), Cd(II) , Hg(II)

Figure 7: Suggested structure of [M(DPIDA)$_2$ Cl$_2$]

**Application**

Samples of tap water and well water were produced and then put to Ligand (DPIDA) to detect mercury concentrations here are some examples of, the findings are reported in Table (15).

Table 15
The outcome of the application of mercury (II) in samples

| Sample | Spectrophotometric method μg/mL | Flam atomic absorption μg/mL |
|---|---|---|
| Well water (west of the city of karbala) | $4.52x10^{-5}$ | $6.22x10^{-5}$ |
| Well water (east of the city of karbala) | $1.85x10^{-6}$ | $2.62x10^{-6}$ |
| Well water (north of city of karbala) | $2.39x10^{-6}$ | $3.02x10^{-6}$ |
| Well water (south of city of karbala) | $3.38x10^{-6}$ | $1.99x10^{-6}$ |
| Tap water | $9.02x10^{-6}$ | $8.22x10^{-6}$ |

**Conclusion**

a technique for assessing small changes with extreme sensitivity mercury in several samples, as well as a low-cost approach for measuring Hg, have been devised (II). Verification and application investigations revealed that the proposed approach is capable of quantifying Hg (II). The results showed that the reagent can detect mercury (II) in a variety of materials. Analytical criteria as an example: identification; detection limit; precision; and recovery show that this technique is capable of determining Hg (II).

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