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Determination Of Stability Constants of Some Bivalent and Trivalent Metal Ions with A Biologically Active Ligand

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----------ABSTRACT----------

Amino acids, the fundamental units of proteins, perform essential roles in biological systems as metabolic intermediates, neurotransmitters, hormone precursors and regulators of osmotic balance and cellular pH. Among them, aspartic acid, a proteinogenic amino acid, is noteworthy for its ability to chelate with metal ions due to the presence of both amine and carboxylate functional groups. This chelation property is critical for modulating the bioavailability and reactivity of metal ions in biochemical systems.

In this study, we have investigated the formation and stability of complexes formed between aspartic acid and three biologically significant metal ions: Zinc (Zn^{2+}), Cobalt (Co^{2+}), and Aluminium (Al^{3+}). The determination of stability constants was done using pH-metric titration at an ionic strength of 0.02 M, applying Bjerrum's Half-Integral Method, a widely used graphical technique in coordination chemistry.

The experimental results demonstrated that, at room temperature, the stability of the complexes followed the order: $Zn^{2+} > Co^{2+} > Al^{3+}$. This sequence is consistent with the Irving–William series, which outlines the general trend of stability among divalent transition-metal complexes. The reduced stability observed for the Al^{3+} complex can be explained by its unique coordination properties combined with its high charge density, which promotes hydrolysis over stable ligand coordination under the studied conditions.

These findings provide an insight into the coordination behaviour of essential metal ions with biologically active ligands, with implications in fields such as medicinal chemistry, bioinorganic modelling and environmental chelation strategies.

Keywords: *Metal-ligand complexes, pH-metric titration, Irving–William series, transition metal ions, coordination chemistry, trivalent and divalent metal ions, bioinorganic chemistry, thermodynamic stability*

I. INTRODUCTION

A coordination compound or complex is created when a ligand binds to an acceptor through a lone pair of electrons. Coordination compounds consist of complex ions or molecules and primarily fall into two categories: neutral compounds and ionic compounds [1]. A metal complex, also known as a coordination compound, is a neutral or ionic compound in which metal atoms or ions are encircled by ligands, such as organic molecules or inorganic ions [2]. A ligand can be described as any molecule or ion that possesses at least one electron pair available for donation. Non-classical ligands are those which form compound with transition metals ions. Metals have d orbitals that can participate in bonding, while ligands possess donor capabilities as well as acceptor properties. There are different ligands such as monodentate, bidentate, tridentate, tetradentate, and multidentate — based on the number of atoms of the ligand donating an electron pair. Bidentate ligands that are fully attached to a single atom are referred to as chelation ligands. The metal complex compounds created by these ligands feature a ring structure and are referred to as chelates. The process of creating a metal chelate is called chelation. The coordination number denotes the quantity of ligands that are directly attached to the central atom or ion in a complex compound. In metal chelates, the number of ligands does not correspond to the coordination number of the metal ions. In both analytical chemistry and biological systems, the chelate effect is crucial. The coordination sphere is made up of the core metal atom and the ligands that are directly connected to it; they are encapsulated in square brackets [1]. Most credible reports indicate that aspartate is typically tridentate. Through a particular active transport mechanism, amino acids are efficiently absorbed from the intestinal lumen. In order to produce several transition metal complexes with the majority of periodic table elements, the amino acid molecules are utilized as chelating ligands and bidentate ligands [3]. There are 20

known amino acids, further classified into essential, non-essential, and semi- essential amino acids [4]. The majority of essential amino acids are thought to be vital resources for cytokine production and immune function, as well as for the production of antibodies in animals. After being absorbed, amino acids are assembled and metabolized to form proteins that are used to build various body tissues [5]. Amino acids are the basic building blocks of proteins and are essential for many physiological functions, such as muscle growth and protein synthesis. Every cell and tissue in the body depends on proteins for both structure and function. They serve as receptors, transporters, hormones, enzymes, and structural elements [6]. Amino acids are essential building components of proteins, which are required for almost all biological functions. These molecules are essential for many physiological processes, including immunological reactions, muscular contraction, enzyme functioning, and hormone and neurotransmitter synthesis. Amino acids' structure, classification, and function emphasize how crucial they are to preserving health and enabling daily functions [7]. Amino acid chelated minerals offer a significant benefit in terms of improving the effectiveness of mineral absorption and translocation within the plant. Numerous areas of human interest have seen the widespread use of metal chelating compounds. Chelators have medical uses, water softeners are utilized as ingredients in many commercial items including food preservatives and shampoos, and they also help aquacultures manage heavy metal contamination. The building blocks of proteins found in all living things are amino acids. Chelation of amino acids with minerals resulted in a significant increase in translocation and absorption efficiency [8]. D-Asp plays a part in the control of the hypothalamus and brain development. Citrullinemia, asparagine synthetase insufficiency, Canavan disease, and dicarboxylic aminoaciduria are among the genetic conditions affecting L-Asp metabolism. L-Asp contributes to the pathophysiology of neurological and mental conditions as well as changes in BCAA levels in hyperammonaemia and diabetes. It has been determined that L-Asp plays a critical role in mitochondrial activity and connects the processes of protein synthesis, PNC, glycolysis, gluconeogenesis, amino acid catabolism, and cell division. Further research is required to investigate methods of targeting L-Asp metabolism to combat cancer and the impact of increased L-Asp consumption on carcinogenesis, respectively, given the critical role of L-Asp in cell proliferation and reports of elevated cancer risk in subjects consuming higher amounts of aspartame [9]. To assess the consequences of increased aspartame intake during pregnancy, carefully planned intervention-controlled trials are required. Potential therapeutic possibilities can be created by manipulating the wide range of qualities that cobalt possesses. Research on the mechanism of cobalt medicinal compounds can yield important information about the physicochemical characteristics that can be used to build new drugs. Because Co (III) complexes can undergo bio reduction — a process in which intracellular reduction yields a bioactive agent and they can be employed as prodrugs. Labile d^7 high-spin Co (II) complexes, which quickly undergo ligand substitution, are produced by reducing inert d^6 Co (III) complexes [10]. The human immune system depends heavily on zinc, a transition metal that is necessary for all living things. Zn is an anti-inflammatory and anti-viral agent [11]. Since aluminium is not known to have any vital roles in human health, its presence in the body could be viewed as a risk without any advantages. It has been demonstrated to effectively compete with Mg (II) and block enzymatic activities that depend on Mg (II), obstructing the use of ATP (e.g. energy production processes). Neurotransmitter secretion and enzymatic activity are both disrupted by Al (III). Al (III) encourages neurofibrillary degeneration by hyper phosphorylating healthy neurofilaments. Al (III) disturbs the normal production of blood cells and competes with Cu (II) in calcifying tissues; therefore, Al is a toxic metal ion for the body [12]. This information may be useful in fields like pharmaceutical formulation, metal ion detoxification, and the creation of bioinorganic model systems. While knowledge of the weaker binding of Al^{3+} could guide methods to reduce aluminium toxicity, the strong binding affinity of Zn^{2+} and Co^{2+} to aspartic acid may be pertinent in therapeutic chelation techniques.

THEORETICAL BACKGROUND

STABILITY CONSTANTS AND FACTORS AFFECTING STABILITY CONSTANTS

The extent of interaction between a metal ion and a ligand in solution is a quantitative measure termed as the stability constants of the complex. These constants provide crucial in understanding their behaviour in various chemical and biological systems and provide information about the thermodynamic stability of metal complexes. The stability constants (k) for a metal complex formation reaction can be expressed as:

$$\mathbf{M} + \mathbf{L} \rightleftharpoons \mathrm{ML}$$

$$k = \frac{[\mathbf{ML}]}{[\mathbf{M}][\mathbf{L}]}$$

[M]= the concentration of the free metal ion [L]= the concentration of the free ligand

[ML]= the concentration of the metal-ligand complex

FACTORS AFFECTING STABILITY CONSTANTS

The intrinsic properties of the metal ion, the nature of the ligand, and the surrounding environmental conditions have a major impact on stability constants of metal–ligand complexes. The charge and size of metal ion plays an important role as higher charges lead to stronger electrostatic interactions with ligands, and smaller ionic radii allow closer approach and stronger binding. Electronic configuration effects such as ligand field stabilization, benefits transition metal ions and this helps to enhance complex stability due to the involvement of partially filled d-orbitals. Basicity and charge are major contributors in case of ligand; negative charge carrying ligand and more basic ligand tend to form more stable complexes because of their better electron-donating ability and electrostatic attraction towards metal ions. Chelate effect which includes the fact that multidentate ligands form stable ring structures with metal ions, which helps in better stability through both entropic and enthalpic contributions. One of the significant impact is of pH as it influences the protonation state of the ligand, which in turn affects its coordination ability. Electrostatic interactions between metal and ligand can be altered due to ionic strength hence, Salts like sodium nitrate which act as inert electrolytes are often employed to maintain constant ionic strength during stability constants measurements. Other factors such as temperature and solvent effects can also have an influence but they are typically controlled during experimental conditions and are considered secondary in systems.

BJERRUM'S HALF-INTEGRAL METHOD

A graphical method used in coordination chemistry to estimate stepwise formation constants of metal-ligand complexes is known as Bjerrum half-integral method. A formation curve is plotted between the average number of ligands bound to a metal ion ($\bar{n}$) versus the negative logarithm of the concentration of free ligand (pL). According to the above technique, when $\bar{n}$

= 0.5, the logarithm of the first stability constants (log k1) is about equal to pL. When $\bar{n}$ = 1.5 the pL number may then be used to estimate (log k2). The systems which have large differences in consecutive formation constants, works very well with this method. It uses pH-metric titration to get the complex formation stability constants. Complex ion formation or dissociation occurs in many steps for each molecule in the solution.

$$\bar{\mathbf{n}}_{\mathbf{H}} = \frac{\left[y + T_L^0 + \frac{[V' - V''][N_B + E^0]}{V^0 + V'} \right]}{T_L^0}$$

y = The number of titrable hydrogen ions

T^0L = Initial concentrations of the ligand reagent

V^0 = Initial Volume of solution in flask

V' = Volume of base to reach same pH in Acid blank

V'' = Volume of base to reach same pH in Ligand blank

NB = Concentration of Base

E^0 = Concentration of Acid

$$\bar{\mathbf{n}} = \left[\frac{[V''' - V''][N_B + E^0 + T_L^0(y - \bar{n}_H)]}{[V^0 + V'']\bar{n}_H \times T_m} \right]$$

$\bar{n}H$ = the average number of protons bound per non-complex-bound ligand molecule

y = The number of titrable hydrogen ions

T^0L = Initial concentrations of the ligand reagent

$\mathbf{V^0}$ = Initial Volume of solution in flask

V'' = Volume of base to reach same pH in Ligand blank

V''' = Volume of base to reach same pH in Acid + Ligand + Metal system

$$\mathbf{pL= \log \left[\left(\frac{1+10^{pKa-pH}}{T_L^0-\bar{n}T_m}\right)\left(\frac{V^0+V'''}{V^0}\right)\right]}$$

$\mathbf{E^0}$ = Concentration of Acid

Tm = The initial concentrations of the metal

$\mathbf{T^0L}$ = Initial concentrations of the ligand reagent

n̄ = the average number of ligands bound per metal atom or ion

$\mathbf{V^0}$ = Initial Volume of solution in flask

Tm = The initial concentrations of the metal

V''' = Volume of base to reach same pH in Acid + Ligand+ Metal system

Various assumptions on which Bjerrum's Half-Integral Method relies on to accurately determine the stability constants of metal– ligand complexes are as follows; First, it assumes that chemical equilibrium is rapidly established at each stage of the titration, which allow analysing of system under equilibrium conditions. A constant ionic strength throughout the titration process is also assumed, which is typically maintained by adding a sufficient concentration of an inert electrolyte, such as sodium nitrate ($NaNO_3$). Also, negligible hydrolysis of the metal ion within the pH range under investigation, or that any hydrolytic effects are appropriately corrected. Finally, the method operates under the assumption of ideal solution behaviour, whereby concentrations are used directly in calculations instead of activities. However, in more rigorous treatments, activity coefficients may be incorporated to improve accuracy.

EQUIPMENTS AND MATERIAL

Two-neck round bottom flask, Burette, Pipette, pH meter, Buffers (pH 4 and pH 7), Measuring cylinders, Standard flasks (100 mL), Combination electrode, Spatula, Weighing balance.

REAGENTS USED

Aspartic Acid, Potash Alum, Sodium Nitrate ($NaNO_3$), Nitric Acid (HNO_3), Dimethylamine ($(CH_3)_2NH$), Cobalt(ous) Sulphate ($CoSO_4$), Zinc Acetate.

EXPERIMENT

SET 01: ACID BLANK

In the first part of the experiment, the behaviour of nitric acid alone was examined through pH-metric titration using dimethylamine as the titrant. A solution was prepared by mixing 5 mL of 0.0025 M HNO_3, 2 mL of 2 M $NaNO_3$ (used to maintain ionic strength), and 32 mL of distilled water in a three-neck round-bottom flask. The initial pH was recorded with a calibrated pH meter. Dimethylamine was added gradually, two drops at a time, from a burette, and the pH was noted after each addition. As the titration progressed, a shift in the pH was observed at one point, indicating the release of a proton from the acid. The process was continued until the pH readings stabilized and no further significant change was seen.

SET 02: LIGAND BLANK

This titration aimed to study the deprotonation behaviour of aspartic acid in the absence of any metal ions. A solution containing 5 mL of 0.0025 M HNO_3, 2 mL of 2 M $NaNO_3$, 12 mL of distilled water, and 20 mL of 0.02 M aspartic acid was prepared and transferred into a clean three- neck round-bottom flask. After measuring the starting pH, dimethylamine was introduced slowly in two-drop increments. The pH was recorded after each addition. During the titration, two distinct inflexion points appeared on the pH curve, which corresponded to the release of protons by the functional groups of the aspartic acid. The titration continued until the pH value became consistent, indicating completion.

SET 03: ACID + LIGAND + METAL SYSTEMS

In the final set, complex formation between aspartic acid and metal ions was explored using the same pH-metric approach. Separate titrations were performed for each metal ion: aluminium (Al^{3+}) from potash alum, zinc (Zn^{2+}) from zinc acetate, and cobalt (Co^{2+}) from cobalt sulphate. For each titration, 5 mL of 0.02 M HNO_3, 2 mL of 2 M $NaNO_3$, 12 mL of distilled water, 20 mL of 0.02 M aspartic acid, and 1 mL of 0.02 M metal salt solution were mixed in a three- neck round-bottom flask. The initial pH of the solution was taken, and dimethylamine was again added gradually in two-drop intervals. The pH was monitored after each addition. In all three cases, two inflexion points were observed, indicating stepwise complex formation between the ligand and the respective metal ions. The titration was stopped once the pH values ceased to change significantly, and the data collected

Figure 1. Experimental Setup for pH-metric Titration

was used to assess the stability and nature of the resulting metal–ligand complex.

II. OBSERVATIONS

SYSTEM 1: ALUMINIUM ASPARTIC ACID Table 1: VALUES OF V' AND V'' FOR SAME pH

S. No.	V' (mL)	V'' (mL)	pH	n̄H
1	0.2	0.1	2.22	2.03
2	0.3	0.3	2.36	2.0
3	0.4	0.4	2.51	2.0
4	0.7	2.1	9.28	1.53
5	0.8	2.5	9.76	1.43
6	0.9	3.0	9.96	1.30
7	1.0	3.4	10.03	1.21
8	1.1	3.9	10.09	1.08
9	1.2	4.0	10.18	1.12
10	1.3	4.2	10.15	1.02
11	1.4	5.0	10.21	0.83
12	1.5	5.5	10.25	0.70

Table 2: VALUES OF V''' AND V'' AT SAME pH

S. No.	$\bar{n}H$	V‴ (mL)	V″ (mL)	pH	$\bar{n}$
1	2.03	0.2	0.1	2.22	0.16
2	2.0	0.6	0.3	2.36	0.48
3	2.0	0.8	0.4	2.51	0.64
4	1.53	2.3	2.1	9.28	0.42
5	1.43	3.0	2.5	9.76	1.12
6	1.30	3.5	3.0	9.96	1.23
7	1.21	4.0	3.4	10.03	1.58
8	1.08	5.0	3.9	10.09	3.25
9	1.12	5.3	4.2	10.15	3.97
10	1.02	5.4	4.0	10.18	3.44
11	0.83	6.2	5.0	10.21	4.59
12	0.70	6.5	5.5	10.25	4.52

Table 3: CALCULATION OF pL WITH VALUES OF $\bar{n}$

S. No.	V‴ (mL)	pH	$\bar{n}$	pL
1	0.2	2.22	0.16	4.38
2	0.6	2.36	0.48	4.28
3	0.8	2.51	0.64	4.20
4	2.3	9.28	0.42	3.96
5	3.0	9.76	1.12	3.98
6	3.5	9.96	1.23	3.99
7	4.0	10.03	1.58	4.00
8	5.0	10.09	3.25	4.02
9	5.3	10.15	3.97	4.03

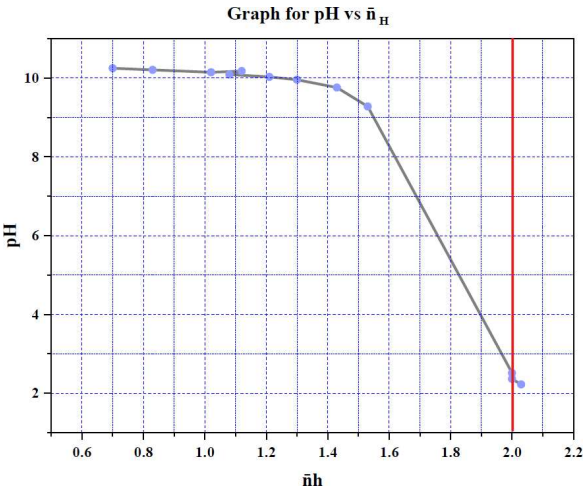


Figure 2. Plot of $\bar{n}$ vs pH for Al-Asp metal Complex

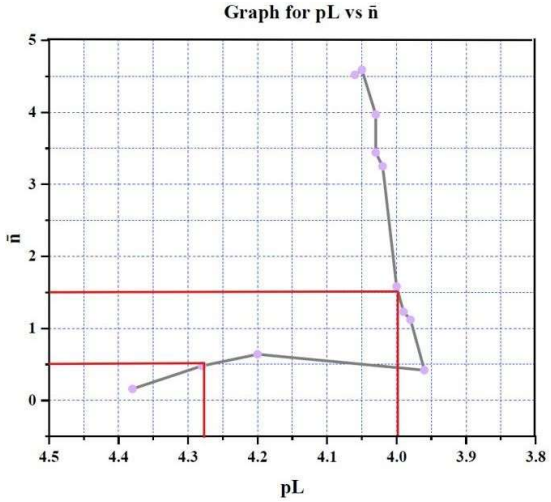


Figure 3. Plot of pL vs $\bar{n}$ for Al-Asp metal Complex

SYSTEM 2: COBALT ASPARTIC ACID

Table 4: VALUES OF V' AND V'' FOR SAME pH

S. No.	V' (mL)	V'' (mL)	pH	n̄H
1.	0	0	3.52	2.00
2.	0.8	0.8	3.57	2.00
3.	1.3	1.3	3.77	2.00
4.	1.4	2.5	8.32	1.64
5.	1.6	3.3	9.27	1.41
6.	0.6	0.7	3.54	1.09

Table 5: VALUES OF V''' AND V'' AT SAME pH

S. No.	n̄H	V''' (mL)	V'' (mL)	pH	n̄
1.	2.0	0.3	0.1	3.52	0.32
2.	2.0	0.9	0.8	3.57	0.16
3.	2.0	1.3	1.3	3.77	0.0
4.	1.64	2.6	2.5	8.32	0.19
5.	1.09	5.2	4.6	10.12	1.73

Table 6: CALCULATION OF pL WITH VALUES OF n̄

S. No.	V''' (mL)	pH	pL	n̄
1.	0.3	3.52	5.55	0.32
2.	0.9	3.57	5.57	0.16
3.	1.3	3.77	5.57	0.0
4.	2.6	8.32	5.58	0.19
5.	5.2	10.12	5.64	1.73

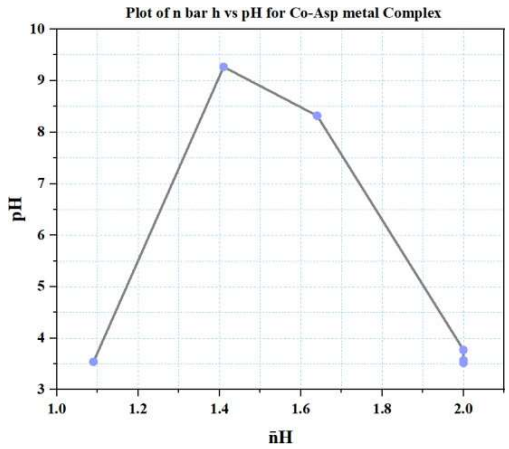


Figure 4. Plot of n̄ vs pH for Co-Asp metal Complex

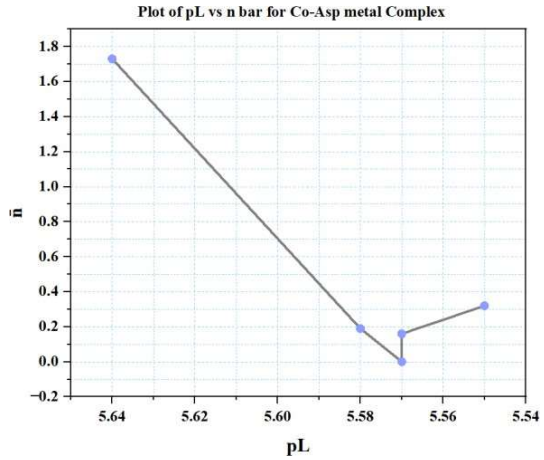


Figure 5. Plot of pL vs n̄ for Co-Asp metal Complex

SYSTEM 3: ZINC ASPARTIC ACID

Table 7: VALUES OF V' AND V'' FOR SAME pH

S. No.	V' (mL)	V'' (mL)	pH	$\bar{n}H$
1	0	0	2.28	2.00
2	3.4	4.0	11.97	1.81
3	3.2	3.9	11.95	1.78
4	2.9	3.8	11.94	1.71
5	2.5	3.7	11.90	1.62
6	2.0	3.6	11.86	1.48

Table 8: VALUES OF V''' AND V'' AT SAME pH

S. No.	V''' (mL)	V'' (mL)	pH	$\bar{n}H$	$\bar{n}$
1.	5.5	2.2	11.57	1.46	7.31
2.	5.6	2.3	11.59	1.47	7.24
3.	6.9	3.3	11.79	1.45	0.2

Table 9: CALCULATION OF pL WITH VALUES OF $\bar{n}$

S. No.	V''' (mL)	pH	$\bar{n}$	pL
1.	5.5	11.57	7.31	5.63
2.	5.6	11.59	7.24	5.67
3.	6.9	11.79	0.2	5.68

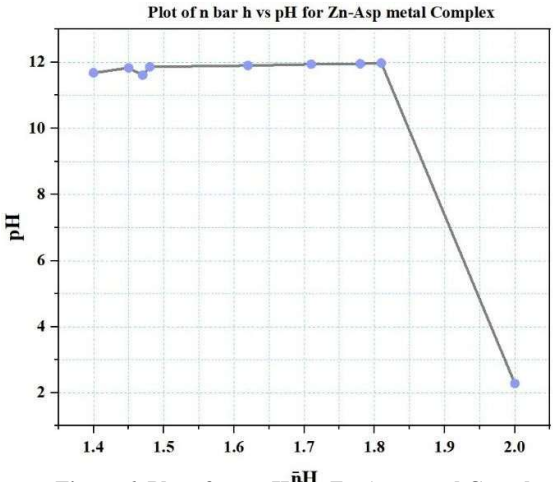


Figure 6. Plot of $\bar{n}$ vs pH for Zn-Asp metal Complex

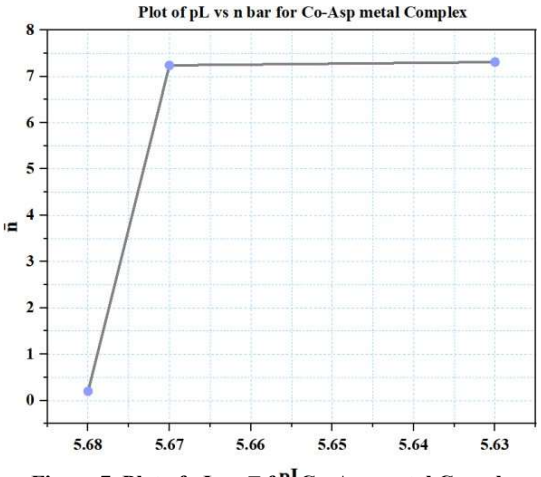
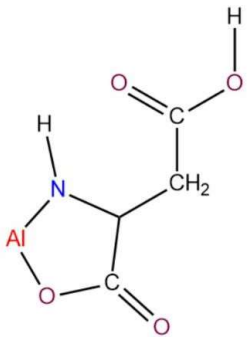


Figure 7. Plot of pL vs $\bar{n}$ for Co-Asp metal Complex

ENERGY MINIMISZED STRUCTURES OF DIFFERENT METAL ION COMLEXES WITH ASPARTIC ACID



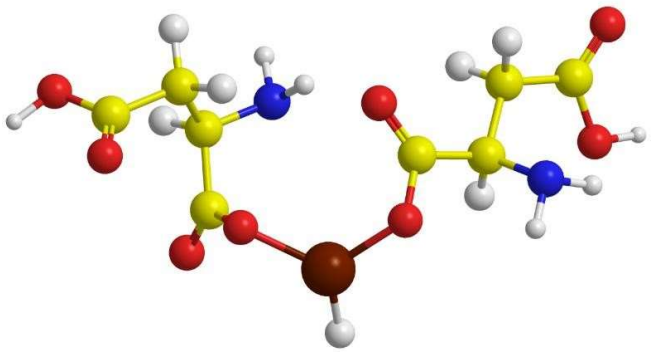
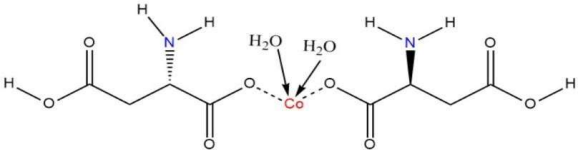


Figure 8. Al-Asp Metal Complex **Figure 9. Ball and Stick Model of Al-Asp Metal Complex**



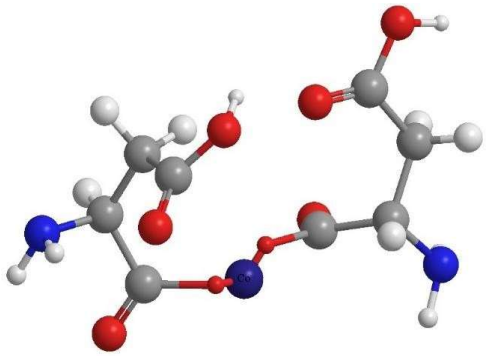
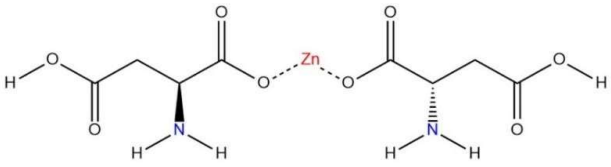


Figure 10. Co-Asp Metal Complex **Figure 11. Ball and Stick Model of Co-Asp Metal Complex**



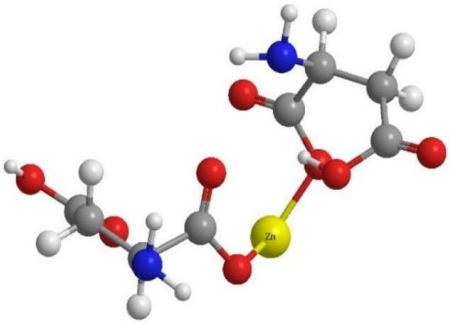


Figure 12. Zn-Asp Metal Complex **Figure 13. Ball and Stick Model of Zn-Asp Metal Complex**

III. RESULT AND DISCUSSION

The stability constants for the aspartic acid complexes of Zn^{2+}, Co^{2+}, and Al^{3+} were determined using Bjerrum's Half-Integral Method at $\bar{n} = 0.5$ and 1.5. The experimentally obtained values are summarised in Table.

S.No.	Metal Ions Used	Stability Constants at $\bar{n} = 0.5$	Stability Constants at $\bar{n} = 1.5$
1	Zinc	5.95	5.60
2	Cobalt	5.62	5.64
3	Aluminium	4.28	4.00

The order of overall stability for the complexes was found to be of the order: $Zn^{2+} > Co^{2+} > Al^{3+}$.

The Irving–William series describes a well-established pattern in the stability of divalent transition-metal complexes, which follows the order: $Mn^{2+} < Fe^{2+} < Co^{2+} < Ni^{2+} < Cu^{2+} > Zn^{2+}$ [13,14]. Although Al^{3+} does not feature in this classical sequence, our experimental results follow the anticipated behaviour for late transition metals, with Zn^{2+} showing the highest stability among the ions studied [15,16]. The slightly greater stability of Zn^{2+} over Co^{2+} can be explained by its d^{10} electron configuration, which results in a fully occupied valence shell, reducing distortion and favouring strong, stable interactions with oxygen-donor ligands such as aspartic acid [17,18].

Influence of Charge Density and Ionic Radius

A key factor affecting the stability of metal–ligand complexes is charge density [19,20]. Aluminium (III) carries a +3 charge and has an ionic radius of approximately 0.535 Å, giving it the highest charge density of the metals considered [21]. While such characteristics should promote strong attraction toward negatively charged donor atoms, in near-neutral aqueous solutions, Al^{3+} undergoes significant hydrolysis, forming hydroxo complexes that compete with ligand binding and thereby lower the measured stability constants [22,23,24]. Zinc (II) (~0.74 Å) and cobalt (II) (~0.745 Å) have similar ionic sizes [25], but differ in their ligand field effects. Zn^{2+}, with a d^{10} configuration, is free from Jahn–Teller distortion and typically forms highly symmetrical complexes [26], whereas Co^{2+} (d^7) may adopt different spin states, often resulting in slightly reduced stability [27,28].

Ligand–Metal Interaction Strength

Aspartic acid contains three potential donor sites — an amino nitrogen and two carboxylate oxygens enabling both bidentate and tridentate coordination [29,30]. In principle, hard acids such as Al^{3+} should interact strongly with oxygen donor atoms [31]. However, its high hydration energy and tendency toward hydrolysis decrease its effective coordination with the ligand [32]. Zn^{2+} and Co^{2+}, classified as borderline acids in HSAB theory, also bind well to these donor groups, but Zn^{2+} generally forms more stable complexes due to superior orbital overlap and reduced geometric strain [33,34]. The chelate effect enhances complex stability by providing an entropic advantage when ligands form multiple bonds to the same metal ion [35]. Aspartic acid can coordinate in both bidentate and tridentate modes, increasing stability for all three metals. Zn^{2+} benefits most from this effect because it can accommodate either tetrahedral or octahedral geometries without imposing significant strain on the ligand framework. In contrast, the preferred coordination geometries of Co^{2+} and Al^{3+} may introduce strain, slightly lowering their stability.

Comparison with Literature Values

The measured stability constants for Zn^{2+}–aspartate (~5.8–6.0) and Co^{2+}–aspartate (~5.5–5.7) under similar conditions are in close agreement with previously reported values [36]. For Al^{3+}– aspartate, the observed log K (~4.0–4.3) is lower than some literature values obtained at lower pH, highlighting the strong influence of hydrolysis and acidity on stability.

Application of the pH-metric method in combination with Bjerrum's Half-Integral approach produced consistent results under controlled ionic strength, aligning with findings from earlier studies. However, minor deviations may result from Al^{3+} hydrolysis at higher pH, the possible formation of mixed-ligand complexes not considered in the simple stoichiometric model, and slight fluctuations in ionic strength during titration. Nevertheless, the overall stability trend $Zn^{2+} > Co^{2+} > Al^{3+}$ is well supported by coordination chemistry principles.

IV. CONCLUSION

The present work was to identify and determine the stability constants of the complexes of Zn^{2+}, Co^{2+}, and Al^{3+} with aspartic acid at ionic strength 0.02 M by pH-metric titration and Bjerrum's Half-Integral Method. The results showed a series of stability to be $Zn^{2+} > Co^{2+} > Al^{3+}$, as coordination chemistry would predict and in general agreement with the Irving– William series for divalent transition metal ions. The apparently lower stability of the Al^{3+} complex was accounted for as due to its higher charge density and extensive hydrolysis upon dissolution in aqueous solution.

The findings of this research are significant as they offer valuable insights into the interaction properties of biologically important metal ions with amino acid ligands, which can have notable implications in areas such as metal ion detoxification, drug design, and the development of bioinorganic model systems.

The strong binding affinity of Zn^{2+} and Co^{2+} toward aspartic acid suggests potential applications in therapeutic chelation, while the observed lower binding tendency of Al^{3+} highlights its relevance in strategies aimed at mitigating aluminium toxicity.

Future work could extend this study by examining the effect of varying temperature, ionic strength,

and pH on complex stability, as well as investigating other biologically active ligands with different donor sets. Furthermore, detailed biological evaluations of these metal–aspartate complexes are proposed to be undertaken to assess their potential biochemical roles and pharmacological applications.

CONFLICT OF INTEREST

The authors declare no Conflict of Interest

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