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Co(II), Ni(II), Cu(II) and Cd(II)-thiocarbohydrazone complexes: spectroscopic, DFT, thermal, and electrical conductivity studies†

 R. Fouad,^a Ibrahim A. Shaaban,^{bc} Tarik E. Ali,^{ba} Mohammed A. Assiri^b and S. S. Shenouda^d

New and stable coordinated compounds have been isolated in a good yield. The chelates have been prepared by mixing Co(II), Ni(II), Cu(II), and Cd(II) metal ions with (1*E*)-1-((6-methyl-4-oxo-4*H*-chromen-3-yl)methylene)thiocarbohydrazide (MCMT) in 2 : 1 stoichiometry (MCMT : M²⁺). Various techniques, including elemental microanalyses, molar conductance, thermal studies, FT-IR, ¹H-NMR, UV-Vis, and XRD spectral analyses, magnetic moment measurements, and electrical conductivity, were applied for the structural and spectroscopic elucidation of the coordinating compounds. Further, computational studies using the DFT-B3LYP method were reported for MCMT and its metal complexes. MCMT behaves as a neutral NS bidentate moiety that forms octahedral complexes with general formula [M(MCMT)₂Cl(OH₂)Cl·X·H₂O (M = Cu²⁺; (X = ½), Ni²⁺, Co²⁺; (X = 1)); [Cd(MCMT)₂Cl₂]·½H₂O. There is good confirmation between experimental infrared spectral data and theoretical DFT-B3LYP computational outcomes where MCMT acts as a five-membered chelate bonded to the metal ion through azomethine nitrogen and thiocarbonyl sulphur donors. The thermal analysis is studied to confirm the elucidated structure of the complexes. Also, the kinetic and thermodynamic parameters of the thermal decomposition steps were evaluated. The measured optical band gap values of the prepared compounds exhibited semiconducting nature. AC conductivity and dielectric properties of the ligand and its complexes were examined, which showed that Cu(II) complex has the highest dielectric constant referring to its high polarization and storage ability.

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1. Introduction

Thiocarbohydrazones dominate a large section of coordination chemistry owing to the flexibility of their structure, including several donor atoms for coordination with metals. The metal ion can successfully coordinate to thiocarbohydrazone through both its sulfur and nitrogen atoms, which is assumed to have medicinal activity in this class of compounds.^{1–3} Also, both atoms allow the thiocarbohydrazone an opportunity to play an important role as chelating agents for forming stable complexes either in neutral or anionic form, usually acting as monodentate or bidentate ligands through S (thione) and N (azomethine) atoms.⁴ Transition metal complexes with thiocarbohydrazone

demonstrate pharmacological activity, inclusive of antifungal, antibacterial, antimalarial,^{5–7} antitumor,^{8,9} antiviral,¹⁰ anticancer,¹¹ antioxidative,¹² antineoplastic,¹³ antiproliferative,¹⁴ and antitubercular¹⁵ properties.

A wide range of research works was concerned with the biological properties of thiocarbohydrazone complexes, and limited works presented the semiconducting nature of these compounds. Hence significant attention has been devoted to examining the conductivity and electrical properties of these compounds.

Although the electrical conductivity of some transition metal complexes has been measured,^{16,17} the electrical and dielectric properties of Co(II), Ni(II), Cu(II), and Cd(II)-(1*E*)-1-((6-methyl-4-oxo-4*H*-chromen-3-yl)methylene)thiocarbohydrazide has not been reported yet.

Depending on the above facts, the present paper is devoted to the synthesis and spectroscopic characterization of thiocarbohydrazone and its Co(II), Ni(II), Cu(II), and Cd(II) complexes. Also, quantum mechanical (QM) calculations were carried out to verify the infrared spectral interpretation and resolve the preferred chelating site. AC electrical conductivity and dielectric properties of the prepared transition metal complexes were examined. The work was embraced to

^aDepartment of Chemistry, Faculty of Education, Ain Shams University, Roxy, Cairo, Egypt. E-mail: raniaahmed@edu.asu.edu.eg; Fax: +20 22581243; Tel: +201000212207

^bDepartment of Chemistry, Faculty of Science, King Khalid University, P. O. Box 9004, Abha 61321, Saudi Arabia

^cDepartment of Chemistry, Faculty of Science (Men's Campus), Al-Azhar University, Nasr City 11884, Cairo, Egypt

^dPhysics Department, Faculty of Education, Ain Shams University, Roxy, Cairo, Egypt

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demonstrate the likelihood for their uses in electronic applications through understanding their frequency response and providing information about their storage and dissipation of the electric fields.

2. Experimental details

2.1. Materials and characterization

The starting materials were analytical grade. Thiocarbonylhydrazide and 3-formylchromone are prepared according to the published literature.^{18,19} The metal salts Co(II), Ni(II), Cu(II), and Cd(II) were used as chloride salts from BDH or Merck. Organic solvents were spectroscopic grade and were used as received.

Carbon, hydrogen, nitrogen, and sulfur microanalyses were conducted at the National Research Center in Dokki, Giza, Egypt. Complexometrically, the metal percentage in the complexes was estimated. A melting point apparatus (Stuart), England, was applied for detecting the melting or decomposition points. 10^{-3} M solutions of solid complexes in DMF were prepared for molar conductivity calculation by the corning conductivity meter NY 14831 model. A Shimadzu thermogravimetric analyzer was used to investigate TGA-DTG measurements from room temperature up to 800 °C. Using the FT-IR Nicolet IS10 spectrometer, FT-IR spectra (4000–400 cm^{-1}) of metal complexes were obtained. $^1\text{H-NMR}$ spectra were determined at room temperature on a 'Varian FT-290.90 MHz' spectrometer. Dimethyl sulfoxide and DMSO- d_6 , were used as the solvent and tetramethylsilane (TMS) as an internal reference. The UV-Vis (JASCO type V-550) spectrophotometer was used to measure the diffused reflectance of powder samples and to record electronic spectra of 10^{-3} M solid complexes, and DMF is used as a solvent. The Gouy system calculated the magnetic susceptibility of complexes at room temperature using Shelwood scientific, Cambridge Science Park, magnetic susceptibility balance (England). The effective magnetic moments were evaluated using the $\mu_{\text{eff}} = 2.828 (X_m T)^{1/2}$ B.M. relationship, where X_m is the molar susceptibility corrected for the diamagnetism of all atoms in compounds using Pascal's constants. ESR spectra were recorded on the Bruker, Model: EMX, X-band spectrometer. PHILIPS diffractometer with $\text{CuK}\alpha_1$ radiation ($k = 1.54056 \text{ \AA}$) was used to record the XRD patterns. A 40 kV accelerating voltage and a 30 mA emission current were applied. Resistance of the samples has been measured by Keithly electrometer (E-616A). The impedance, loss tangent and the capacitance of the samples have been measured using computerized PM-6304 Fluke & Phillips RCL Bridge.

2.2. Synthesis of (1E)-1-((6-methyl-4-oxo-4H-chromen-3-yl)methylene) thiocarbonylhydrazide ligand (MCMT)

(1E)-1-((6-Methyl-4-oxo-4H-chromen-3-yl)methylene) thiocarbonylhydrazide ligand (MCMT) was synthesized according to reported method.²⁰ Yield: 88%; mp ~ 193 °C; anal. found (calcd) for $\text{C}_{12}\text{H}_{12}\text{N}_4\text{O}_2\text{S}$: M_w (g): (276.31); % C: 52.22 (52.16); % H: 4.52 (4.37); % N: 20.30 (20.27); % S: 11.40 (11.60); FT-IR (KBr, ν , cm^{-1}): 3292, 3172 (NH_2 , NH), 3057 (CH)_{arom}, 1617 ($\text{C}=\text{O}$)_{γ-pyrone}, 1596 ($\text{CH}=\text{N}$), 1235, 815 ($\text{NH}-\text{C}=\text{S}$); $^1\text{H-NMR}$ (δ , ppm): 2.19 (s, 3H, CH_3), 4.19 (s, 2H, NH_2), 6.78–6.89 (m, 1H, H-8), 7.68–

7.71 (m, 1H, H-7), 8.31 (s, 1H, H-5), 8.53 (s, 1H, $\text{CH}=\text{N}$), 9.14 (s, 1H, H-2), 10.41 (s, 1H, NH), 12.09 (s, 1H, NH). $^{13}\text{C-NMR}$: 21.6 (CH_3), 116.8 (C-3), 118.0 (C-8), 125.7 (C-4a), 126.3 (C-5), 127.3 (C-6), 135.2 (C-7), 148.6 ($\text{C}=\text{N}$), 151.3 (C-8a), 154.2 (C-2), 174.1 ($\text{C}=\text{O}$), 185.0 ($\text{C}=\text{S}$); UV/Vis (DMF) λ_{max} (nm): 270, 296, 316, 345 and 410.

2.3. General method for the synthesis of transition metal complexes $[\text{M}(\text{MCMT})_2\text{OH}_2\text{Cl}]\text{Cl} \cdot \text{XH}_2\text{O}$ ($\text{M} = \text{Cu}^{2+}$; ($\text{X} = \frac{1}{2}$), Ni^{2+} , Co^{2+} ; ($\text{X} = 1$) and $[\text{Cd}(\text{MCMT})_2\text{Cl}_2] \cdot \frac{1}{2}\text{H}_2\text{O}$

An ethanolic solution of MCMT (1.6 g, 6 mmol) was added to metal salts (3 mmol) namely, $\text{CuCl}_2 \cdot 6\text{H}_2\text{O}$, $\text{NiCl}_2 \cdot 6\text{H}_2\text{O}$, $\text{CoCl}_2 \cdot 6\text{H}_2\text{O}$, and $\text{CdCl}_2 \cdot 6\text{H}_2\text{O}$. The reaction mixture was refluxed for 8 h. The precipitate was filtered and washed several times with 50% (v/v) ethanol-water to remove any traces of unreacted starting materials. Finally, the precipitate was dried in vacuum desiccators over anhydrous CaCl_2 overnight. The dried complexes are $[\text{Co}(\text{MCMT})_2(\text{Cl})(\text{H}_2\text{O})]\text{Cl} \cdot \text{H}_2\text{O}$; yield: 73%; mp > 300 °C; anal. found (calcd) for $\text{C}_{24}\text{H}_{28}\text{N}_8\text{O}_6\text{S}_2\text{Cl}_2\text{Co}$: M_w (g): (718.50); % C: 40.21 (40.11); % H: 3.75 (3.92); % N: 15.37 (15.59); % S: 8.69 (8.92); % Co: 8.11 (8.20); FT-IR (KBr, ν , cm^{-1}): 3432 (OH), 3168 (NH_2 , NH), 3024 (CH)_{arom}, 1617 ($\text{C}=\text{O}$)_{γ-pyrone}, 1545 ($\text{CH}=\text{N}$), 1212, 788 ($\text{NH}-\text{C}=\text{S}$), 544 (M-O), 475 (M-N). Molar conductance (Λ_m): 85 $\text{ohm}^{-1} \text{mol}^{-1} \text{cm}^2$; UV/Vis (DMF) λ_{max} (nm): 288, 354, 399, 434, 660. μ_{eff} (B.M.): 4.4.

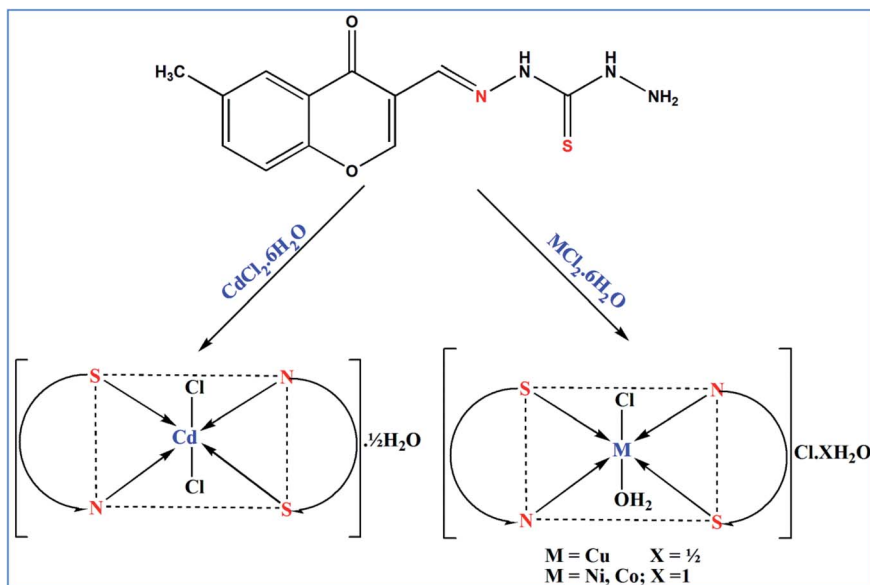
$[\text{Ni}(\text{MCMT})_2(\text{Cl})(\text{H}_2\text{O})]\text{Cl} \cdot \text{H}_2\text{O}$; yield: 70%; mp > 300 °C; anal. found (calcd) for $\text{C}_{24}\text{H}_{28}\text{N}_8\text{O}_6\text{S}_2\text{Cl}_2\text{Ni}$: M_w (g): (718.26); % C: 40.31 (40.13); % H: 3.79 (3.92); % N: 15.38 (15.59); % S: 8.77 (8.92); % Ni: 8.09 (8.17); FT-IR (KBr, ν , cm^{-1}): 3424 (OH), 3166 (NH_2 , NH), 3026 (CH)_{arom}, 1617 ($\text{C}=\text{O}$)_{γ-pyrone}, 1540 ($\text{CH}=\text{N}$), 1210, 790 ($\text{NHC}=\text{S}$), 544 (M-O), 478 (M-N). Molar conductance (Λ_m): 89 $\text{ohm}^{-1} \text{mol}^{-1} \text{cm}^2$; UV/Vis (DMF) λ_{max} (nm): 272, 354, 396, 451, 585; μ_{eff} (B.M.): 2.8.

$[\text{Cu}(\text{MCMT})_2(\text{Cl})(\text{H}_2\text{O})]\text{Cl} \cdot \frac{1}{2}\text{H}_2\text{O}$; yield: 75%; mp > 300 °C; anal. found (calcd) for $\text{C}_{24}\text{H}_{27}\text{N}_8\text{O}_{5.5}\text{S}_2\text{Cl}_2\text{Cu}$: M_w (g): (714.10); % C: 40.54 (40.36); % H: 3.77 (3.81); % N: 15.55 (15.69); % S: 8.75 (8.98); % Cu: 8.67 (8.89); FT-IR (KBr, ν , cm^{-1}): 3426 (OH), 3193, 3170 (NH_2 , NH), 3025 (CH)_{arom}, 1617 ($\text{C}=\text{O}$)_{γ-pyrone}, 1547 ($\text{CH}=\text{N}$), 1204, 789 ($\text{NH}-\text{C}=\text{S}$), 545 (M-O), 470 (M-N). Molar conductance (Λ_m): 90 $\text{ohm}^{-1} \text{mol}^{-1} \text{cm}^2$; UV/Vis (DMF) λ_{max} (nm): 274, 351, 394, 442, 580; μ_{eff} (B.M.): 1.8. EPR (at room temperature, X-band): $g_{\parallel} = 2.13$ and $g_{\perp} = 2.06$.

$[\text{Cd}(\text{MCMT})_2(\text{Cl})_2] \cdot \frac{1}{2}\text{H}_2\text{O}$; yield: 75%; mp > 300 °C; anal. found (calcd) for $\text{C}_{24}\text{H}_{25}\text{N}_8\text{O}_{4.5}\text{S}_2\text{Cl}_2\text{Cd}$: M_w (g): (744.96); % C: 38.49 (38.69); % H: 3.47 (3.38); % N: 15.10 (15.04); % S: 8.51 (8.60); % Cd: 15.30 (15.08); FT-IR (KBr, ν , cm^{-1}): 3440 (OH), 3178 (NH_2 , NH), 3030 (CH)_{arom}, 1619 ($\text{C}=\text{O}$)_{γ-pyrone}, 1540 ($\text{CH}=\text{N}$), 1209, 793 ($\text{NH}-\text{C}=\text{S}$), 477 (M-N). ($^1\text{H-NMR}$, δ , ppm): 2.36 (s, 3H, CH_3), 3.88 (s, 2H, NH_2), 6.77–6.81 (m, 1H, H-8), 6.88–6.91 (m, 1H, H-8), 7.62–7.67 (m, 2H, H-7), 8.23 (s, 1H, H-5), 8.24 (s, 1H, H-5), 8.83 (s, 1H, $\text{CH}=\text{N}$), 8.84 (s, 1H, $\text{CH}=\text{N}$), 9.14 (s, 1H, H-2), 9.15 (s, 1H, H-2), 10.49 (br, 2H, CSNHNH_2), 12.19 (s, 2H, $\text{CH}[\text{NNH}]\text{CS}$). Molar conductance (Λ_m): 15 $\text{ohm}^{-1} \text{mol}^{-1} \text{cm}^2$; UV/Vis (DMF) λ_{max} (nm): 288, 318, 417.

The chemical structures of the obtained metal complexes are represented in Scheme 1.





Scheme 1 Synthesis of transition metal complexes.

2.4. Computational methods

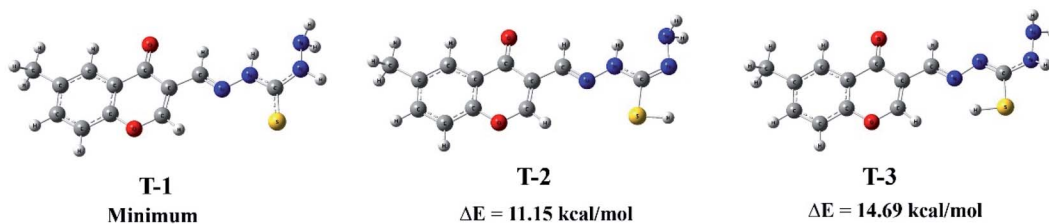
QM calculations in this work were performed based on density functional theory (DFT) using a B3LYP level,^{21,22} which is a well-established method and broadly applied for transition metal complexes.^{23–26} Standard 6-31G(d) basis set was used for all atoms except for the heavy Cd(II) ion where the effective core potential (ECP) LanL2DZ basis set was used²⁷ where the pseudo-potential of core electrons were included. Moreover, spin multiplicities of quartet, triplet, doublet, and singlet were used for Co(II), Ni(II), Cu(II), and Cd(II) complexes, respectively. All QM calculations reported herein were carried out employing the Gaussian 09 program package.²⁸ Based on the gradient method of Pulay,²⁹ full geometry optimization was achieved through simultaneous relaxation of all geometrical parameters until the convergence criteria of Gaussian 09 package was reached.²⁸ Subsequently, the vibrational frequencies were calculated to verify that the obtained optimized geometry belongs to true energy minima, which is indicated by the absence of imaginary frequencies. The *x,y,z* coordinates for all calculated structures along with their energies in atomic unit are provided in the ESI.†

3. Results and discussion

3.1. Physicochemical study

The performed synthesis involving the reaction of MCMT with metal chlorides in reflux system afforded, generally in good yield, four colorful complexes of empirical formula, $[M(\text{MCMT})_2\text{Cl}(\text{H}_2\text{O})]\text{Cl}\cdot X\text{H}_2\text{O}$; ($M = \text{Cu}^{2+}$; $X = \frac{1}{2}$; Ni^{2+} , Co^{2+} ; $X = 1$) and $[\text{Cd}(\text{MCMT})_2(\text{Cl})_2]\cdot \frac{1}{2}\text{H}_2\text{O}$. These discrete complexes are quite stable in the air and can be maintained in a desiccator for a long time without any observed decomposition. All prepared complexes are soluble in Lewis bases such as DMF and DMSO but sparingly soluble in methanol and ethanol. The elemental analysis supports the suggested chemical structures of the complexes. As shown in Scheme 1, the complexes are formed in 1 : 2 metal-to-ligand stoichiometry. Moreover, the analytical data of metal chelates showed that the ligand behaves as a neutral bidentate.

The molar conductance values of ~ 1 mM solution for all prepared complexes were measured at room temperature in DMF. The molar conductance values for Co(II), Ni(II), and Cu(II) complexes are within the characteristic range of 1 : 1 electrolytes, indicating that one anion is located out of the coordination sphere as a counter ion.^{30,31} On the other hand, the

Fig. 1 Equilibrium geometries of suggested tautomers for MCMT and their relative energies (ΔE) using B3LYP/6-31G(d) calculations.

dissolved Cd(II) complex was characterized by non-electrolytic nature.^{32,33}

3.2. Geometry structural study

After several recrystallization trials, we could not get the single crystals for the prepared metal complexes. Therefore, we took advantage of QM calculations to theoretically confirm the proposed structures for metal complexes. DFT-B3LYP calculations have been carried out to infer the preferred coordination sites of the ligand to the metal ion and to conclude the favored geometrical isomerism. The global reactivity descriptors and the order of coordination stability for the prepared metal complexes were also predicted.

Initially, we have dealt with the geometry of the MCMT ligand by exploring its tautomeric stability owing to intramolecular proton transfer from *imino* to *thiocarbonyl* group. Thus, three tautomeric structures were hypothetically proposed, *diamino-thione* (T-1) and two *imino-thiol* (T-2 and T-3). Fig. 1 shows their fully optimized geometries obtained from B3LYP/6-31G(d) calculation along with their relative energies. Accordingly, the *diamino-thione* (T-1) was identified as the most stable tautomeric structure with the lowest energy and real vibrational frequencies. The *imino-thiol* tautomers, T-2 and T-3, are higher in energy than T-1 by 11.15 and 14.69 kcal mol⁻¹, respectively. These results agree with the experimental infrared spectrum of the ligand, where the characteristic band due to the S–H stretch vibration in the 2600–2400 cm⁻¹ range was absent (Section 3.3.1). Employing B3LYP/6-31G(d) calculations, we have predicted the electrostatic potential (ESP) charges^{34,35} to obtain the charge distributions for the most stable tautomer (T-1). The calculated ESP charges for T-1 are given in the ESI, Fig. S1.† The computed ESP charges show a negative atomic charge for carbonyl O (–0.52), N_{az} (–0.45), thiocarbonyl S (–0.34), and amino N (–0.66). These values reflect the electron donating character of these atoms and can be coordinated to metal ions, *i.e.*, active coordination centers.

Owing to the presence of several active coordination centers in the ligand, five structures (S-1 to S-5) were proposed with different possibilities for chelation of the ligand to the metal ion. Firstly, the ligand can form five-membered chelates through N_{az} and S (S-1), *imino* N (S-2), or *amino* N (S-3) donors. In addition, the five-membered chelation can occur *via* the bidentate coordination sites of *amino* and *imino* Ns (S-4). The fifth structure results from coordination through N_{az} and chromone carbonyl O (S-5), forming a six-membered chelate. Consequently, geometry optimization has been performed for these proposed structures (S-1 to S-5) using Cu²⁺ as the metal ion. The equilibrium geometries for S-1 to S-5 as predicted using B3LYP/6-31G(d) calculations are shown in Fig. 2. The computational results favor S-1 with minimum energy in agreement with the infrared spectral interpretation (Section 3.3.1). For structures S-2 to S-5, high energy differences of 12.78–41.62 kcal mol⁻¹ were predicted relative to S-1.

In view of these proposed structures of the complexes, five geometrical isomers could be suggested owing to the *cis* and *trans* orientations for each pair of coordination sites. The pair of N_{az}'s, S's, and Cl/H₂O coordination sites can be located *trans* to each other (**Isomer 1, all trans**) or in *cis* configuration (**Isomer 2, all cis**). For **Isomer 3**, N_{az}'s and S's were *cis* to each other while Cl/H₂O pair arranged in the *trans* orientation. The rearrangement of N_{az}'s/S's pairs to be *cis/trans* to each other and *vice versa* results in **Isomer 4** (*cis*-N_{az}'s and *trans*-S's) and **Isomer 5** (*trans*-N_{az}'s and *cis*-S's). The geometries of **Isomers 1–5** were fully optimized using B3LYP/6-31G(d) calculations for all complexes except Cd-ion, where the ECP-LanL2DZ basis set was used. Table 1 lists the computed energy along with the energy differences (ΔE) with respect to the most stable isomer for the investigated metal complexes. Also, the optimized geometries of geometrical isomers of Co(II), Ni(II), and Cu(II) complexes are displayed in Fig. 3. For the [Cd(MCMT)₂Cl₂] complex, the equilibrium geometries for its geometrical isomers are given in the ESI (Fig. S2†). As seen from the data in Table 1, **Isomer 1** (*all trans*) for the metal complexes under investigation was found to

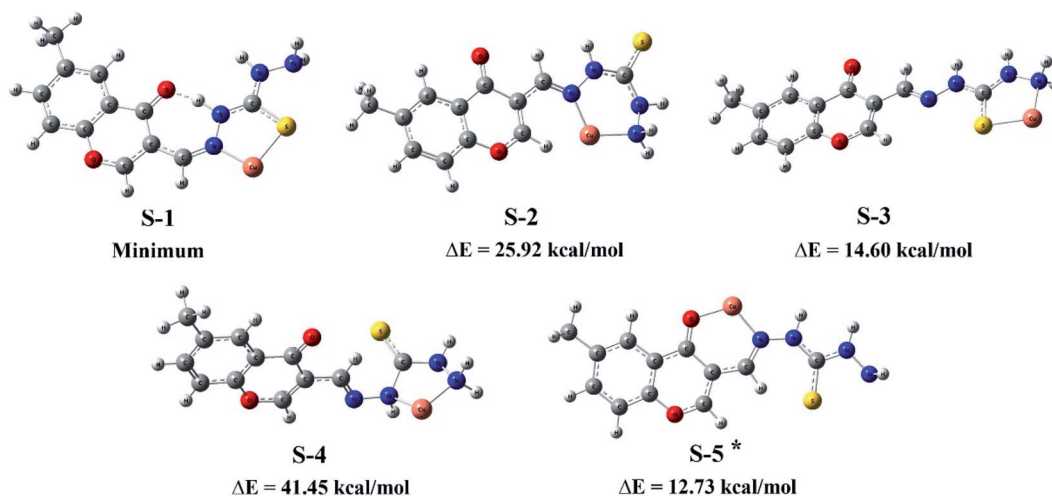


Fig. 2 Optimized geometries and relative energy for the proposed coordination sites for the ligand to Cu(II) ion. * S-5 has an imaginary frequency (*i.e.*, transition state).

Table 1 DFT-B3LYP calculated energy (E , Hartree) and relative energy (ΔE , kcal mol⁻¹) for suggested geometrical isomer of the prepared metal complexes^{a,b}

Complex	Isomer 1 ^c		Isomer 2 ^c		Isomer 3 ^c		Isomer 4 ^c		Isomer 5 ^c	
	E	ΔE	E	ΔE	E	ΔE	E	ΔE	E	ΔE
Cu(II) complex	-4640.974656	0.0	-4640.919524	34.60	-4640.92091	33.73	-4640.921366	33.44	-4640.893092	51.18
Ni(II) complex	-4508.863215	0.0	-4508.829768	20.99	-4508.83659	16.71	-4508.831065	20.17	-4508.805617	36.14
Co(II) complex	-4383.342023	0.0	-4383.309348	20.50	-4383.30149	25.43	-4383.298181	27.51	-4383.294179	30.02
Cd(II) complex	-3432.53667	0.0	-3432.454744	51.41	-3432.33548	126.25	-3432.350125	117.06	-3432.353494	114.94

^a Calculations were performed using 6-31G(d) basis set for all atoms except for Cd(II) ion where ECP-LanL2DZ basis set was used. ^b Energy differences (ΔE) are relative to the most stable **Isomer 1** (minimum energy). ^c For the structures of **Isomers 1–5**, see Fig. 3 (Co, Ni, and Cu-complexes) and Fig. S2 (Cd-complex).

be the most stable isomer with the lowest energy and real frequencies. These results were consistent with the reported crystal structure for Diaqua-bis(1,1-dimethylthiocarbazide) nickel(II) fumarate.³⁶ **Isomers 2–5** were disfavored with an energy difference of 16.71–126.25 kcal mol⁻¹ compared to **Isomer 1**, which could account for the strain and the repulsion between the *cis*-oriented coordination site pairs.

The stability of the prepared metal complexes has been anticipated by calculating the binding energy between the metal ion and ligand and the coordination energy^{37,38} as follows:

$$\Delta E = (E_{M^{2+}} + 2E_L + E_{Cl^-} + E_{H_2O}) - E_{[ML_2Cl(H_2O)]^+}, M^{2+} = Co^{2+}, Ni^{2+} \text{ and } Cu^{2+}$$

$$\Delta E = (E_{M^{2+}} + 2E_L + 2E_{Cl^-}) - E_{[ML_2Cl_2]}, M^{2+} = Cd^{2+}$$

where $E_{M^{2+}}$, E_L , E_{Cl^-} , E_{H_2O} , $E_{[ML_2Cl(H_2O)]^+}$ and $E_{[ML_2Cl_2]}$ are the computed energies using DFT-B3LYP calculations at 6-31G(d) basis set for all atoms excluding Cd atom, where LanL2DZ basis set was used. The computational results predict coordination energies of 442.24, 544.34, 581.35, and 583.02 kcal mol⁻¹ for Cd(II), Co(II), Ni(II), and Cu(II) complexes, respectively. The

positive values for coordination energy infer the highly stable form of the investigated metal complexes. Additionally, the coordination stabilities were in the order of Cu(II) \leftarrow N > Ni(II) \leftarrow N > Co(II) \leftarrow N, which agrees with the relative stabilities of complexes formed by the first row divalent metal ions reported by the Irving–Williams series.³⁹

Energies for the frontier highest occupied/lowest unoccupied molecular orbitals; (HOMO/LUMO) have been computed for **MCMT (S-1)** and their metal complexes in the most stable isomer (1, *all trans*). Afterwards, a set of electronic properties (see Table 2) describing molecular reactivity/stability such as, energy gap (E_g), chemical potential (μ), hardness (η), electronegativity (χ), and electrophilicity index (ω) were predicted.^{40,41} The HOMO's and LUMO's of the **MCMT** ligand and its Co(II) complex are displayed in Fig. 4. For **MCMT**, the HOMO was mainly delocalized over the thiocarbonyl group along with azomethine N with a low extent, which refers to strong electron donation of those moieties. Therefore, **MCMT** was preferred to chelate metal ions *via* these moieties. The HOMO of Co(II) complex has a large contribution from Cl, C=S, and H₂O coordinated ligands. However, the LUMO of **MCMT** and Co(II)

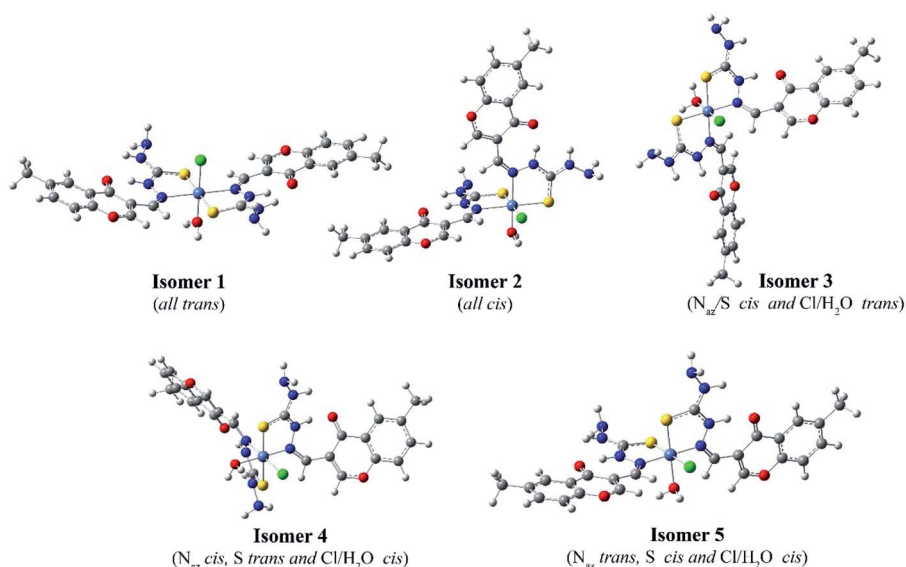


Fig. 3 DFT-B3LYP/6-31G(d) optimized geometries of the suggested geometrical isomers for Co(II), Ni(II), and Cu(II) complexes.



Table 2 Calculated molecular reactivity descriptors for ligand and its metal complexes^a

Molecular parameters	Ligand	Cu-complex	Ni-complex	Co-complex	Cd-complex
E(HOMO) a.u.	-0.19788	-0.25513	-0.24876	-0.24568	-0.13939
E(HOMO) eV	-5.385	-6.942	-6.769	-6.685	-3.793
E(LUMO) a.u.	-0.06025	-0.15521	-0.15734	-0.15725	-0.07795
E(LUMO) eV	-1.639	-4.224	-4.281	-4.279	-2.121
ΔE (LUMO-HOMO) eV	3.745	2.719	2.488	2.406	1.672
Ionization potential, IP eV	5.385	6.942	6.769	6.685	3.793
Electron affinity, EA eV	1.639	4.224	4.281	4.279	2.121
Electronegativity, χ eV	3.512	5.583	5.525	5.482	2.957
Chemical potential, μ eV	-3.512	-5.583	-5.525	-5.482	-2.957
Chemical hardness, η eV	1.873	1.359	1.244	1.203	0.836
Chemical softness, s eV	0.267	0.368	0.402	0.416	0.598
Global electrophilicity index, ω eV	3.294	11.464	12.272	12.490	5.230
Total dipole moment, μ_{tot} debye	3.623	2.252	3.239	3.102	1.747

^a Calculations were performed using B3LYP method and 6-31G(d) basis set for all atoms except for Cd(II) ion where ECP-LanL2DZ basis set was used. For metal complexes, calculations were performed for the most stable **Isomer 1** (minimum energy).

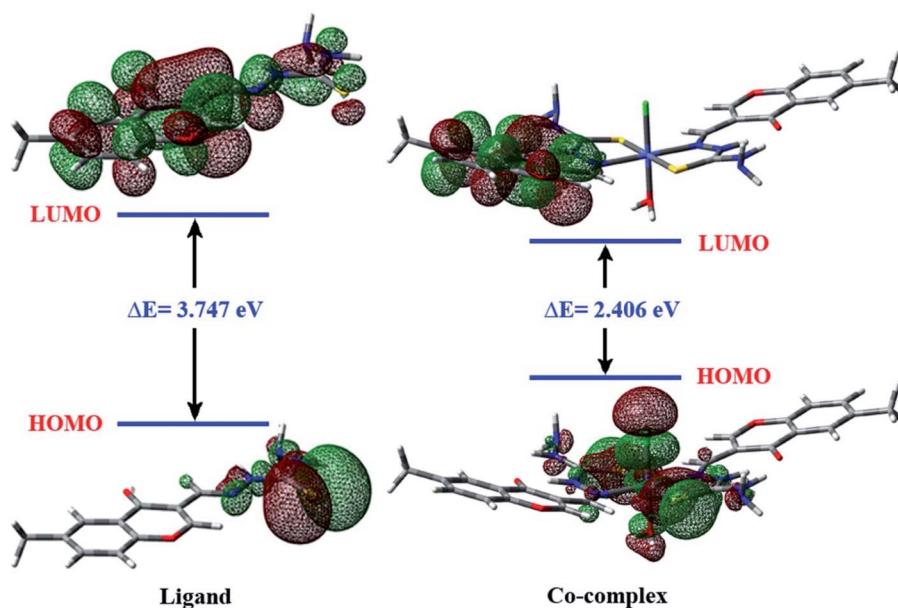


Fig. 4 HOMO and LUMO frontier molecular orbitals for MCMT ligand and its Co(II) complex as predicted using B3LYP/6-31G(d) calculations.

complex was found to be distributed over one chromone moiety.

The HOMO-LUMO energy gap (E_g) is a useful metric for predicting molecular stability and chemical reactivity, along with biological activity. Molecules with a small energy gap are anticipated to have a high chemical reactivity⁴² and result in a significant degree of intramolecular charge transfer reactions, which affect and elevate the bioactivity.^{43,44} In this work, the investigated metal complexes exhibit low values of energy gap (1.672–2.719 eV, see Table 2), implying that they are biologically potent. In addition, the predicted negative values of chemical potential indicate the stability of the complexes, *i.e.*, they do not spontaneously decompose. As shown in Table 2, the Co(II) complex exhibits the highest value of electrophilicity (ω) and reveals a high ability to accept electrons.

3.3. Thermal study

The TGA analysis of synthesized complexes offers valuable details on the stability of the complexes. The steps of thermal degradation of transition metal complexes are presented in Table 3.

A good consensus exists between thermal and analytical data suggested by the theoretical formula. The Co(II), Ni(II), Cu(II), and Cd(II) complexes lose lattice water molecules and become anhydrous at 110, 109, 94, and 106 °C, respectively. There are two routes to remove the coordinated water. In the first route, the coordinated water was removed separately as in Ni(II) and Co(II) complexes at 150 and 160 °C, respectively.

In the second route, coordinated water molecule removal was followed by releasing Cl₂ molecule as in the Cu(II) complex at 239 °C. For Ni(II) and Cd(II) complexes, the Cl₂ molecule was



Table 3 Thermal analysis of thiocarbohydrazone complexes

Compound	Temp range °C	% Total loss in weight found (calc.)	Composition of the residue	Probable composition of the expelled groups
Co(II) complex	30.0–110	2.45 (2.51)	C ₂₄ H ₂₆ N ₈ O ₅ S ₂ Cl ₂ Co	One molecule of lattice H ₂ O
	110–160	5.00 (5.02)	C ₂₄ H ₂₄ N ₈ O ₄ S ₂ Cl ₂ Co	One molecule of coordinated H ₂ O
	161–385	32.0 (31.9)	C ₂₂ H ₁₆ N ₄ O ₄ S ₂ Co	Two molecules of Cl ₂ , NH ₃ and HCN
	386–566	48.5 (48.4)	C ₂₀ H ₁₄ N ₂ O ₄ Co	Two molecules of HSCN
Ni(II) complex	30.0–109	2.50 (2.50)	C ₂₄ H ₂₆ N ₈ O ₅ S ₂ Cl ₂ Ni	One molecule of lattice H ₂ O
	110–150	4.80 (5.00)	C ₂₄ H ₂₄ N ₈ O ₄ S ₂ Cl ₂ Ni	One molecule of coordinated water
	151–253	15.0 (14.8)	C ₂₄ H ₂₄ N ₈ O ₄ S ₂ Ni	One molecule of Cl ₂
	254–364	31.5 (31.8)	C ₂₂ H ₁₆ N ₄ O ₄ S ₂ Ni	Two molecules of NH ₃ and HCN
	365–522	48.5 (48.3)	C ₂₀ H ₁₄ N ₂ O ₄ Ni	Two molecules of HSCN
Cu(II) complex	30.0–94	3.25 (1.26)	C ₂₄ H ₂₆ N ₈ S ₂ O ₅ Cl ₂ Cu	Half molecule of lattice H ₂ O
	95.0–239	24.0 (13.7)	C ₂₄ H ₂₄ N ₈ S ₂ O ₄ Cu	One molecule of coordinated H ₂ O and one molecule of Cl ₂
Cd(II) complex	240–437	47.5 (47.3)	C ₂₀ H ₁₄ N ₂ O ₄ Cu	Two molecules of NH ₃ , HCN and HSCN
	30.0–106	1.10 (1.20)	C ₂₄ H ₂₄ N ₈ O ₄ S ₂ Cl ₂ Cd	Half molecule of lattice H ₂ O
	107–232	10.9 (10.7)	C ₂₄ H ₂₄ N ₈ O ₄ S ₂ Cd	One molecule of Cl ₂
	233–269	20.0 (19.8)	C ₂₄ H ₁₈ N ₆ O ₄ S ₂ Cd	Two molecules of NH ₃
	270–306	27.0 (27.1)	C ₂₂ H ₁₆ N ₄ O ₄ S ₂ Cd	Two molecules of HCN
	307–479	43.0 (42.9)	C ₂₀ H ₁₄ N ₂ O ₄ Cd	Two molecules of HSCN

removed separately at 253 and 232 °C, respectively. On the other hand, the removal of Cl₂ molecule was eliminated with a partial decomposition of the ligand at 385 °C as in Co(II) complex.

The Cu(II) complex continues its decomposition within the temperature range of 240–437 °C with a weight loss of 47.3% assignable to the removal of two molecules of NH₃, HCN, and HSCN.

For Ni(II) complex, the partial ligand pyrolysis comprises two steps in the temperature range of 254–522 °C. The first one occurred within the range 254–364 °C with a weight loss of 31.8%, associated with the elimination of two molecules of NH₃ and HCN. The second step corresponds to a weight loss of 48.3% in the range 365–522 °C, implying the removal of two molecules of HSCN.

The partial ligand pyrolysis of the Cd(II) complex involves three steps. The first step was illustrated by a mass loss of 19.8% ascribed to the removal of two molecules of NH₃. The second thermal decomposition process took place in the temperature range of 270–306 °C with a weight loss of 27.1%. After that, two molecules of HSCN were released within the range of 307–479 °C with a weight loss of 42.9%.

The kinetic and thermodynamic parameters of decomposition processes of the complexes, namely, activation energy (ΔE), enthalpy (ΔH), entropy (ΔS), and Gibbs free energy change (ΔG), were calculated by the Coats–Redfern method^{45,46} in the following forms:

$$\ln[1 - (1 - \alpha)^{1-n}/(1 - n)T^2] = M/T + B, \text{ for } n \neq 1 \quad (1)$$

Table 4 Kinetic parameters of prepared transition metal complexes

Compound	Step	<i>n</i> order	<i>T</i> (K)	<i>A</i> (S ⁻¹)	<i>E</i> (kJ mol ⁻¹)	ΔH (kJ mol ⁻¹)	ΔS (kJ mol ⁻¹)	ΔG (kJ mol ⁻¹)
Co(II) complex	1 st	0.50	348	5.72 × 10 ⁴	36.46	33.56	-0.11	73.32
	2 nd	0.33	417	9.27 × 10 ⁶	60.28	56.82	-0.12	107.69
	3 rd	1.00	551	4.34 × 10 ³	49.83	45.25	-0.18	148.89
	4 th	1.00	751	1.63 × 10 ⁸	135.02	128.78	-0.10	206.20
Ni(II) complex	1 st	0.33	352	5.86 × 10 ²	30.90	27.97	-0.22	105.46
	2 nd	0.33	403	1.8 × 10 ⁷	59.60	56.25	-0.11	103.06
	3 rd	1.00	473	1.39 × 10 ⁶	63.92	59.99	-0.13	125
	4 th	1.00	561	3.94 × 10 ⁶	71.38	66.72	-0.11	130.5
Cu(II) complex	5 th	0.66	719	8.68 × 10 ²	56.55	50.58	-0.20	196.77
	1 st	1	338	1.34 × 10 ⁸	55.17	52.36	-0.09	85.51
	2 nd	0.50	462	1.37 × 10 ⁴	45.23	41.39	-0.17	123.17
Cd(II) complex	3 rd	1.00	548	5.84 × 10	30.90	26.35	-0.22	148.99
	1 st	1.00	339	8.62 × 10 ⁵	42.43	39.61	-0.14	87.123
	2 nd	0.33	448	2.65 × 10 ⁶	66.16	62.44	-0.13	122.04
	3 rd	0	523	5.09 × 10 ⁸	94.56	90.22	-0.09	137.64
	4 th	0.50	561	2.26 × 10 ⁶	77.69	73.02	-0.13	149.45
	5 th	0.50	666	6.92 × 10 ⁶	89.28	83.75	-0.12	169.25



$$\ln[-\ln(1 - \alpha)/T^2] = M/T + B, \text{ for } n = 1 \quad (2)$$

where $M = -E/R$ and $B = \ln AR/\Phi E$; R , A , and Φ are the universal gas constant, pre-exponential factor, and heating rate, respectively.

By plotting the left-hand side of eqn (1) or (2) versus $1000/T$ for different values of $n = 0, 0.33, 0.50, 0.66, 1.00$; the correlation coefficient, r , was calculated. The n value, which gave the best fit ($r \approx 1$), was chosen as the order parameter. The A and E values were determined from the intercept and linear slope of such stage. All other kinetic parameters, ΔH , ΔS , and ΔG , were obtained using the relationships; $\Delta H = E - RT$, $\Delta S = R[\ln(Ah/kT) - 1]$, and $\Delta G = \Delta H - T\Delta S$; k and h are the Boltzmann's constant and Plank's constant.

The thermo kinetic parameters of the decomposition steps for all complexes are collected in Table 4. The calculated data can be summarized with the following remarks:

(i) The activation energies of decomposition E are found to be in the range 30.90 – 135.02 kJ mol^{-1} . For all complexes except Ni(II) and Cu(II) complexes, the pattern of rising activation energy values is as follows: first stages, second stages, third stages. Also, the thermal stability of the complexes is demonstrated by the increasing positive values for E from one step to another.

(ii) The negative ΔS values for all complexes indicate that the activated complex is more ordered than the reactants and/or the reactions are slow.^{47–49}

(iii) The positive values of ΔH imply that the decomposition processes are endothermic.

(iv) The relatively low and positive values of ΔG reveal the metal ion autocatalytic effect on the thermal decomposition of the complexes and non-spontaneous processes.⁵⁰ These data are confirmed with the calculated chemical potential (Table 2) (Section 3.2.)

3.4. Spectroscopic and magnetic studies

The spectroscopic techniques characterize the structure of transition metal complexes, and their spectroscopic analysis data reveal consistent results with the formulated data.

3.4.1. FT-IR spectra. For transition metal complexes, the most suggestive IR bands (4000 – 400 cm^{-1}) have been identified to provide useful information about MCMT ligand bonding sites when coordinated with metal ions. FT-IR spectra of transition metal complexes are provided in Fig. S3.† An analysis and comparison of the IR spectra of MCMT and its metal complexes suggest that MCMT behaves as a neutral bidentate ligand with azomethine–nitrogen and thione–sulfur as NS ligand. The presence of bands refers to the groups of hydrazine N–H and C=S, proposing the coordination of the MCMT ligand in a neutral form to the metal center. Also, the band attributable to C=S of the MCMT ligand ($1235, 815$ cm^{-1})²⁰ was shifted to lower wavenumbers (cm^{-1}) as a result of complex formation. The stretching vibration band of the azomethine group (HC=N) was also shifted to lower frequencies upon complexation. It was thought from first sight that these shifts could indicate the bonding of the unsaturated nitrogen of the azomethine group

and thione-sulfur towards the metal ion.^{51–53} The presence of new bands due to $\nu(\text{M} \leftarrow \text{N})$ is at 470 – 478 cm^{-1} is another obvious sign of the participation of nitrogen atom of the azomethine group in the coordination.^{54–56} The IR spectral data of the Co(II), Ni(II), and Cu(II) complexes show new bands at 3424 – 3440 cm^{-1} correlating to the overlap with lattice and/or coordinated water.⁵⁷ These bands are confirmed by the presence of non-ligand bands in the region of 544 – 545 cm^{-1} in the IR data of the complexes, which correspond to $\nu(\text{M} \leftarrow \text{O})$.^{58,59}

The stretching vibrations of NH_2 , NH , and $(\text{C}=\text{O})_{\gamma\text{-pyrone}}$ groups appeared nearly at the same position in the metal complexes compared to the free ligand as an indication of the non-participation of these groups in coordination.^{60–63}

From above, the interpretation of experimental IR data is confirmed with theoretical results, which showed that the thiocarbonylhydrazone ligand during the coordination process acts as a neutral bidentate through its thione sulfur and imine nitrogen atoms. Moreover, the theoretical spectra for the studied metal complexes (in the most stable isomer; *all-trans*) were obtained using the computed vibrational frequencies along with their infrared intensities (see Fig. S3.†). As a result, the predicted spectra (dashed lines) were found to closely match the experimental spectra (solid lines), particularly in the range below 1600 cm^{-1} , with minor discrepancies in the observed/calculated intensities.

3.4.2. ¹H-NMR spectra. ¹H-NMR spectral study also supported the formation of the thiocarbonylhydrazone ligand. ¹H-NMR spectrum of the MCMT ligand (Fig. S4.†) showed the following signals: 2.19 (s, 3H, CH_3), 4.19 (s, 2H, NH_2), 6.78 – 6.89 (m, 1H, H-8), 7.68 – 7.71 (m, 1H, H-7), 8.31 (s, 1H, H-5), 8.53 (s, 1H, $\text{CH}=\text{N}$),

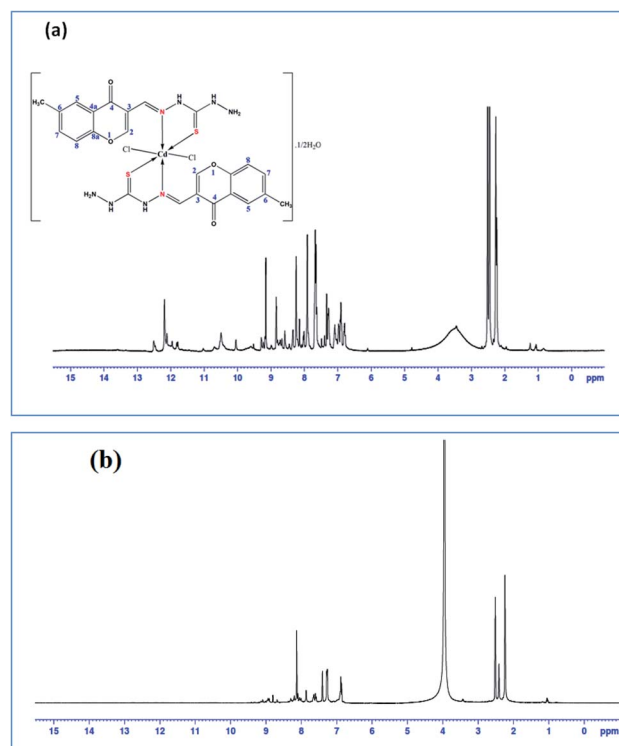


Fig. 5 ¹H-NMR spectra of Cd(II) complex in (a) DMSO- d_6 and (b) D_2O .



9.14 (s, 1H, H-2), 10.41 (s, 1H, NH), 12.09 (s, 1H, NH). In addition, the presence of a signal due to δ NH protons in the spectrum of the ligand implies the ligand was present in the keto-form in the DMSO- d_6 solution.

The $^1\text{H-NMR}$ spectra of the Cd(II) diamagnetic complex relative to TMS in DMSO- d_6 without and with D_2O are represented in Fig. 5a and b. *Via* comparing the $^1\text{H-NMR}$ spectral data of the ligand with the corresponding data in the complex, the coordinating mode of the ligand was clarified. As shown in Fig. 5a, the $^1\text{H-NMR}$ spectrum of the Cd(II) complex shows peaks at 10.49 and 12.19 ppm relative to the NH of CSNH NH_2 and CH]NNHCS groups,^{64–68} respectively, with the slightly upfield shift of two previous signals confirming the coordination of CH=N and C=S groups with metal ion and the MCMT ligand coordinating to the metal ion in neutral form. Also, the azomethine proton signals undergo an upfield shift after complexation. This is attributable to the nitrogen atoms of azomethine donating the lone pair of electrons to the central metal ion, resulting in a coordination linkage (M \leftarrow N).⁶⁹

It is significant to observe the slightly upfield shift of NH_2 protons (3.88 ppm) in the spectrum of the diamagnetic complex compared with the NH_2 signal of MCMT ligand (4.12 ppm) due to the existence of strong hydrogen bonds in the free ligand.

Also, the $^1\text{H-NMR}$ spectrum of the Cd(II) complex displays aromatic protons in the region of δ 6.77–9.15 ppm as complex multiplets.^{70,71} On adding D_2O , the signals attributable to the protons of two NH groups were removed entirely while the other signals remained at the same locations. This statement was confirmed the infrared assignment.

3.4.3. Electronic spectra and magnetic moment. The electronic absorption spectra have been considered a helpful tool in the investigation. The UV-Vis spectra of the MCMT ligand and its related complexes clarified the coordinated geometry and estimated the intra ligand, ligand-metal, and d-d electronic transitions. A good description of the important electronic absorption spectral bands of the ligand and its complexes and an examination of the studies given by magnetic moment measurements are reported. The resultant UV-visible spectra of the metal complexes are presented in Fig. 6. The electronic spectral data of MCMT is characterized by the absorption bands at 270, 296, 316, 345, and 410 nm. The highest energy bands are assigned to π - π^* transitions within aromatic rings.^{69,70} Moderate energy bands can be assigned to n - π^* transitions within C=O and C=N groups.^{72–75} Upon complexation, the π - π^* and n - π^* transitions in all the metal complexes were changed. This reveals the strong coordination between the

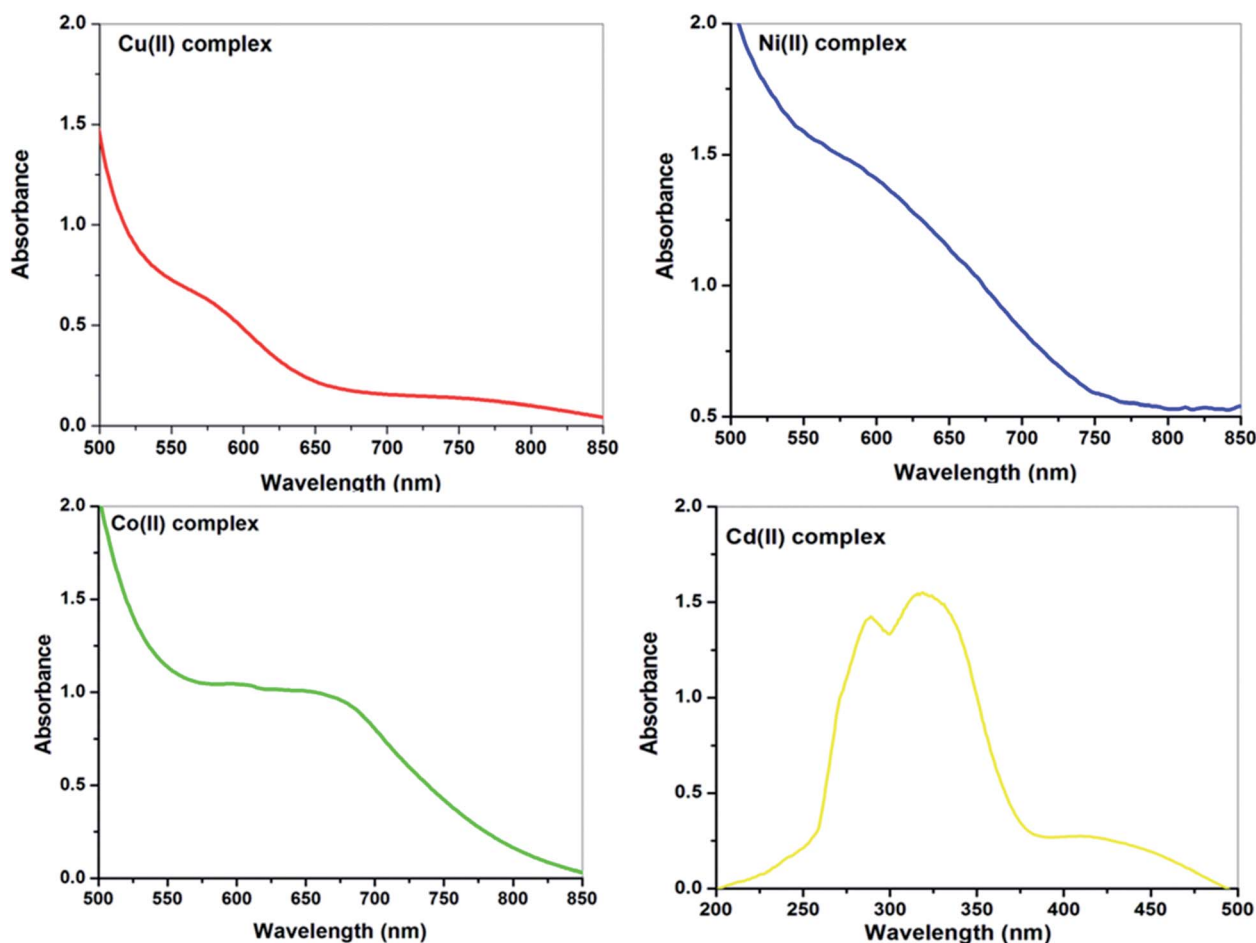


Fig. 6 UV-Vis spectra of transition metal complexes in DMF as the solvent.



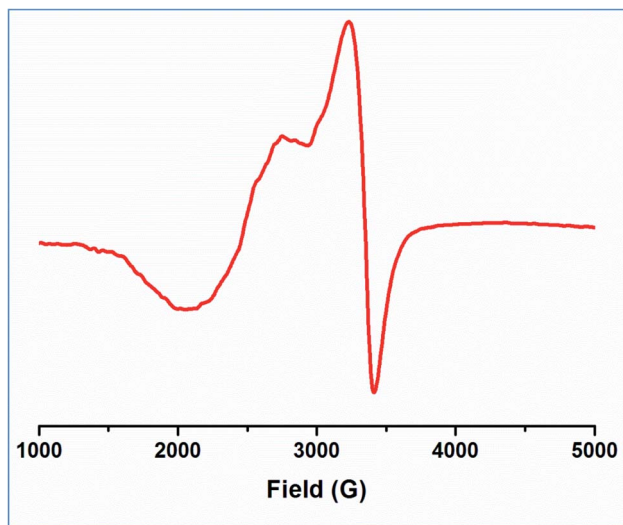


Fig. 7 ESR spectrum of Cu(II) complex.

metal ion and the MCMT ligand. The LMCT transitions are observed in the range of 415–451 nm.⁷⁵ The metal chelates showed additional absorption bands around 580–660 nm attributed to d–d transitions.³⁰

The electronic spectrum of the Co(II) complex, in addition to magnetic moment measurements, illustrated that the complex has an octahedral geometry. The electronic transitions can be successfully explained using the Tanabe–Sugano diagram. In the octahedral Co(II) complexes, the spectra usually contain three bands.⁷⁶ The first band, which is ascribed to ${}^4T_{1g}(F) \rightarrow {}^4T_{2g}(F)$ transition, was not detected due to the fact that it lies in

the near infra-red region and out of range of the used instrument. In the current work, only one broad band was observed in the visible region (660 nm), which may be due to the ${}^4T_{1g} \rightarrow {}^4A_{2g}(F)$ transition in an octahedral geometry.⁷⁷ The value of the room temperature magnetic moment of the Co(II) complex is 4.3 B.M., which agrees with the value of the octahedral geometry.⁷⁸

Generally, the spectra of octahedral Ni(II) include three bands, which are assigned as ${}^3A_{2g}(F) \rightarrow {}^3T_{2g}(F)$, ${}^3A_{2g}(F) \rightarrow {}^3T_{1g}(F)$ and ${}^3A_{2g}(F) \rightarrow {}^3T_{1g}(F)$.⁷⁹ In the electronic spectrum of the complex, the first ${}^3A_{2g}(F) \rightarrow {}^3T_{2g}(F)$ transition was not observed due to the fact that it lies in the near infra-red and it is out of range of the used instrument. The spectrum of the current Ni(II) complex (Fig. 6) shows a broad absorption band at 585 nm, which is attributed to the ${}^3A_{2g}(F) \rightarrow {}^3T_{1g}(F)$ transition, in favor of octahedral geometry.⁸⁰ Also, the magnetic moment value of the Ni(II) complex confirmed the presence of two unpaired electrons in the octahedral geometry (2.90 B.M.).⁸⁰

As a result of the Jahn–Teller distortion and the low symmetry of the environment around copper(II)-ion (d^9 -system), complete interpretations of the electronic spectra and magnetic properties are rather complicated.⁸¹ The electronic absorption spectrum of the Cu(II) complex showed a broad band in the visible region at 580 nm (Fig. 6). This band is assigned to ${}^2E_g \rightarrow {}^2T_{2g}$ transition in the distorted octahedral around Cu(II) complex.^{82,83} The magnetic moment of the Cu(II) complex appeared at 1.8 B.M., which agrees with the value of the octahedral geometry.^{84,85}

The yellow Cd(II) complex displayed absorption bands at 288, 318, and 415 nm that could be due to π - π^* , n - π^* , and CT transitions, respectively. The electronic spectrum of the Cd(II) complex does not include any specific evidence for

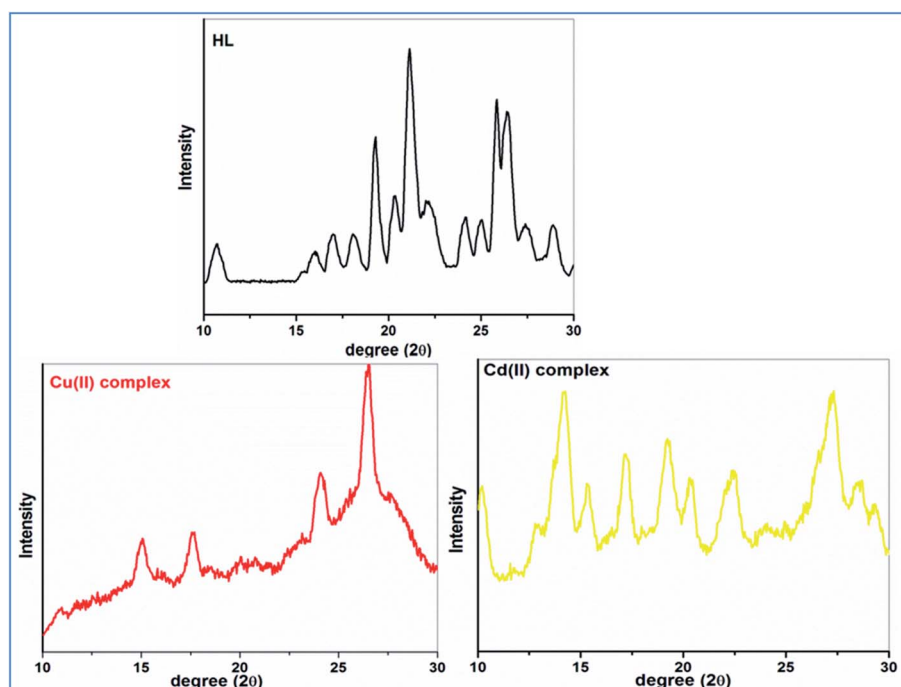


Fig. 8 XRD pattern of MCMT, Cu(II), and Cd(II) complexes.



stereochemistry, but octahedral geometry is proposed for the Cd(II) complex based on analytical, thermal, and infrared spectral evidence.

3.4.4. ESR spectra. ESR spectrum of the Cu(II) complex has been measured at room temperature, as shown in Fig. 7. The shape of ESR signals for the Cu(II) complex indicates octahedral geometry, where two anisotropic signals are observed. From the calculated g values of Cu(II) complex ($g_{\perp} = 2.06$ and $g_{\parallel} = 2.13$), it is obvious that $g_{\parallel} > g_{\perp}$ which indicates that the unpaired electron is predominantly in the $d_{x^2-y^2}$ orbital, giving ${}^2B_{1g}$ as the ground state. The observed data show that the g_{\parallel} value > 2.0023 , and hence the metal–ligand bonding essentially has a covalent character.⁸³ The geometric parameter G for the Cu(II) complex is 2.30, demonstrating a considerable exchange interaction in the solid complex. Molecular orbital coefficients, α^2 and β^2 , were calculated.⁵⁸ The lower value of β^2 (0.44) compared to α^2 (0.79) indicates that the in-plane π -bonding is more covalent than the in-plane σ -bonding.⁸³

3.4.5. X-ray diffraction (XRD) results. Powder XRD was measured for MCMT and its complexes. Fig. 8 shows the XRD patterns of the prepared MCMT and its Cu(II) and Cd(II) complexes as representative examples. XRD patterns showed that the crystals of the complexes are not ideal but lie between amorphous and crystalline characteristics. Because of this, the complexes were not found to be suitable for single crystal XRD. By observing the patterns of XRD of prepared MCMT and its Cu(II) and Cd(II) complexes, it is suggested that the Cu(II) and Cd(II) complexes exhibit crystalline character. It can be noted from the patterns that the diffraction angle and diffraction intensity of the prepared complexes differ from those of the ligand. This finding indicates that the metal ions have been bonded with the ligand to create new complexes.

The crystal sizes of the complexes were determined using the standard Scherer equation's maximal intensity peak.^{86–88} The calculated values for the complexes MCMT, Cu(II), and Cd(II) were 16.6, 15.5, and 12.2 nm, respectively, suggesting nano-size of the crystals. Because of the extremely amorphous pattern of Ni(II) and Co(II) complexes, their crystal size could not be determined.

3.5. Optical study

The measured diffused reflectance (DR) was analyzed using the Kubelka–Munk (K–M) model, which illustrates the behavior of scattered light.⁸⁹ The product of the absorption coefficient α and sample thickness t is the K–M function $F(\text{DR})$, which can be calculated as follows:^{90,91}

$$F(\text{DR}) = (\alpha t) = \frac{(1 - \text{DR})^2}{2 \text{DR}}$$

To determine the energy gap E_g , the absorption coefficient has been analyzed using the following equation:^{92,93}

Table 5 Values of optical energy gaps of MCMT ligand and its transition metal complexes

Compound	Energy gap, [eV]
MCMT	1.65, 2.64, 2.96
Co(II) complex	1.54
Ni(II) complex	1.72, 2.06
Cu(II) complex	1.39, 2.00
Cd(II) complex	1.62, 2.37, 2.73

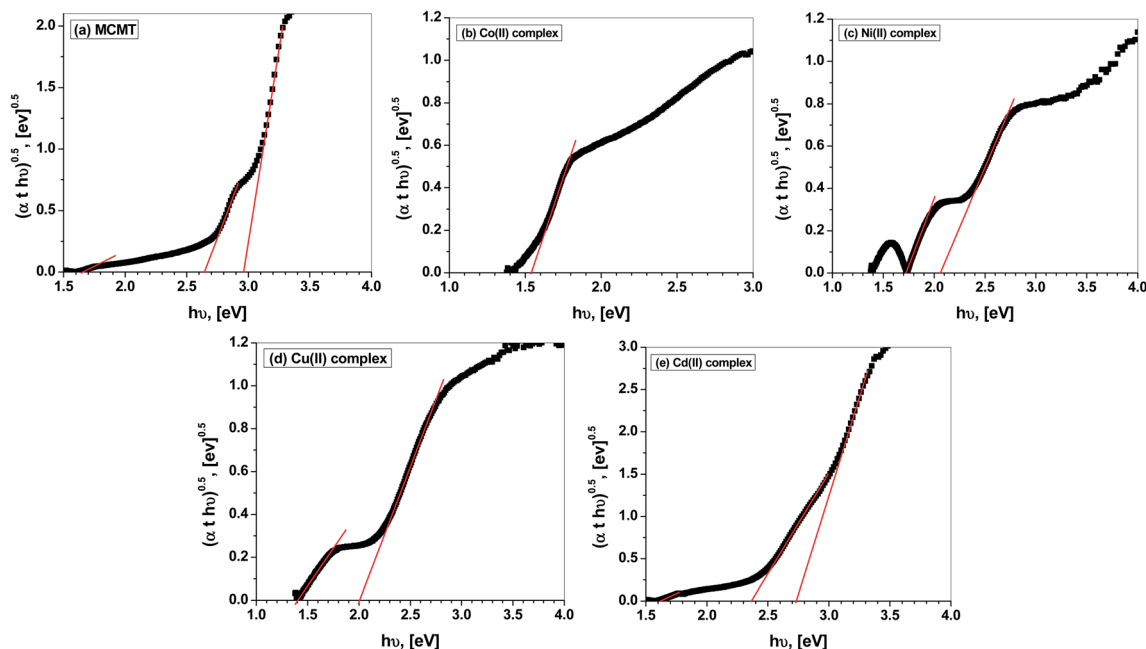


Fig. 9 $(\alpha t h\nu)^{0.5}$ versus $h\nu$ of: (a) MCMT ligand; (b) Co(II) complex; (c) Ni(II) complex; (d) Cu(II) complex; (e) Cd(II) complex.



$$(\alpha hv) = A(hv - E_g)^m$$

where A is a constant and m is a number determining the type of optical transitions. $m = 2$ for indirect allowed transition while $m = 0.5$ for direct allowed transition. Fig. 9a–e show that $(\alpha hv)^{0.5}$ versus hv have regions with linear fitting. This indicates the indirect allowed optical transitions in the ligand and its complexes. The energy gaps have been determined from the extrapolation to the hv axis. Values of the energy gaps are presented in Table 5, showing the semi-conductive and proficient photovoltaic nature of the ligand and complexes.^{94,95} The observed optical energy gap of 1.62 eV for the Cd(II) complex is in excellent agreement with that obtained theoretically (Table 2). On the other hand, the calculated values for Cu(II), Ni(II), and Co(II) complexes were found to be overestimated by 0.428–0.866 eV (Tables 2 and 5). These deviations could be attributed to the fact that the experimental value was obtained for the sample in the solid phase while the theoretical values were predicted for an isolated molecule of the complex ignoring the interactions in the solid state.

3.6. Electrical and dielectric properties

The ac electrical conductivity σ_{ac} of the ligand and its complexes has been calculated using the following equation:⁹⁶

$$\sigma_{ac} = \frac{d}{ZA} - \frac{d}{RA}$$

where Z is the impedance, R is the resistance, d is the thickness, and A is the area. The frequency dependence of ac conductivity is presented in Fig. 10. It is clear that σ_{ac} for all samples increases linearly with frequency. This is a common feature for semiconducting materials where increasing the frequency promotes electronic jumps between localized states and increases the mobility of charge carriers.^{96,97} Clearly, the Cu(II) complex has higher ac conductivity than the ligand while Ni(II),

Co(II), and Cd(II) complexes have lower ac conductivities. This behavior is similar to the Cu(II) complex derived from acyclovir (ACV).⁹⁸ The higher ac conductivity of the Cu(II) complex can be assigned to: (i) the Cu(II) complex has higher coordination energy based on theoretical calculations, and thus strong coordination power between ligand and Cu(II) ion, consequently, implying that the Cu(II) metal ion in the Π -electron delocalization of the organic molecule during complexation leads to decreasing the energy gap and increment the electrical conductivity; (ii) the high crystallinity of the Cu(II) complex as discussed in the XRD section. The ligand has higher ac conductivity than that of Ni(II) and Co(II) complexes due to its high crystallinity while Ni(II) and Co(II) complexes have an amorphous structure. Also, keto–enol transformation is known to occur in such ligand, which implies the movement of Π -electrons along the ligand (*i.e.*, delocalization), which prompts the increment of electrical conductivity. On the other hand, the low ac conductivity of the Cd(II) complex may be attributed to the absence of free unpaired electrons. Also, the Cu(II) complex has the highest ac conduction response to the frequency. This is a promising feature to be a potential applicant for electronic applications. The frequency dependence has been analyzed using the following Jonscher's universal power equation:^{99,100}

$$\sigma_{ac} = B(2\pi f)^S$$

where B is a constant. The frequency exponent S determines the mechanism of the ac conduction. Fig. 11 presents $\ln \sigma_{ac}$ versus $\ln(2\pi f)$ for all complexes. The slope of both lines (S) is close to unity, as seen in Table 6. This behavior contradicts the Quantum Mechanical Tunneling (QMT) model,^{101,102} which predicts the value for S to be about 0.81. Thus, the ac conduction in these complexes could be explained as correlated barrier hopping (CBH).

The CBH model considered a simultaneous hopping of two electrons between close pairs of centers (negatively and

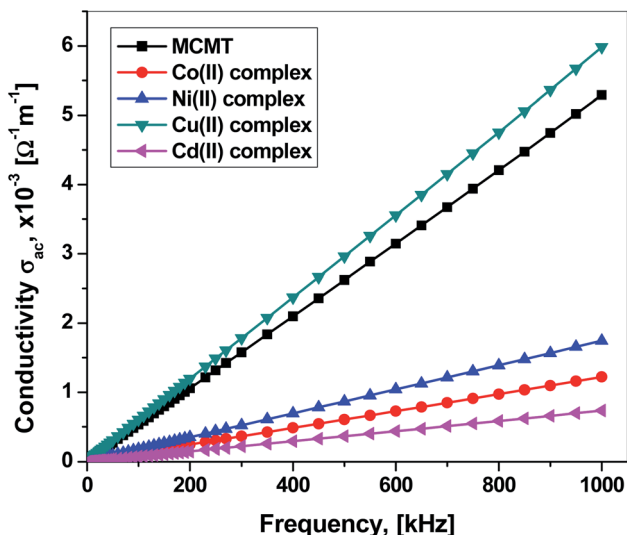


Fig. 10 Frequency dependence of the ac conductivity for MCMT and Co(II), Ni(II), Cu(II), and Cd(II) complexes.

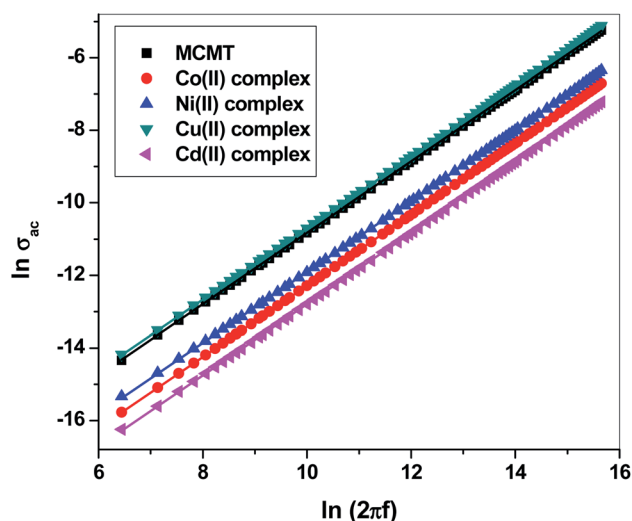


Fig. 11 $\ln \sigma_{ac}$ versus $\ln(2\pi f)$ for MCMT and Co(II), Ni(II), Cu(II), and Cd(II) complexes.



Table 6 Values of the frequency exponent and maximum barrier height (W_m)

Compound	S	W_m
MCMT	0.9824	2.9850
Co(II) complex	0.9815	2.8399
Ni(II) complex	0.9764	2.2311
Cu(II) complex	0.9819	2.9010
Cd(II) complex	0.9809	2.7507

positively charged centers) over a barrier height, and these two centers are correlated with inter-site separation (5–10 Å apart) *via* the coulombic interaction. When all the charged, spin-paired defect centers in the materials are associated into close pairs, the pairs could be with intimate or random distribution. So, the CBH could happen between intimate valence alternation pairs (IVAP) or non-intimate valence alternation pairs (NVAP), respectively. The frequency exponent S is temperature dependent with values less than 1, while S is very close to unity in IVAP. Thus, the ac conduction in these complexes could be explained as correlated barrier hopping (CBH) between intimate valence alternation pairs (IVAP).^{103,104} The maximum barrier height (W_m) in this model can be calculated using the following equation:¹⁰⁵

$$W_m = \frac{2k_B T}{1 - S}$$

Values of W_m are presented in Table 6. Clearly, the Ni(II) complex has the lowest maximum barrier height.

Fig. 12 shows the frequency dependence of the dielectric constant ϵ_1 for the studied complexes. Clearly, the Cu(II) complex has a higher dielectric constant than the ligand, while Ni(II), Co(II), and Cd(II) complexes have a lower dielectric constant. At low frequencies, ϵ_1 decreases sharply with

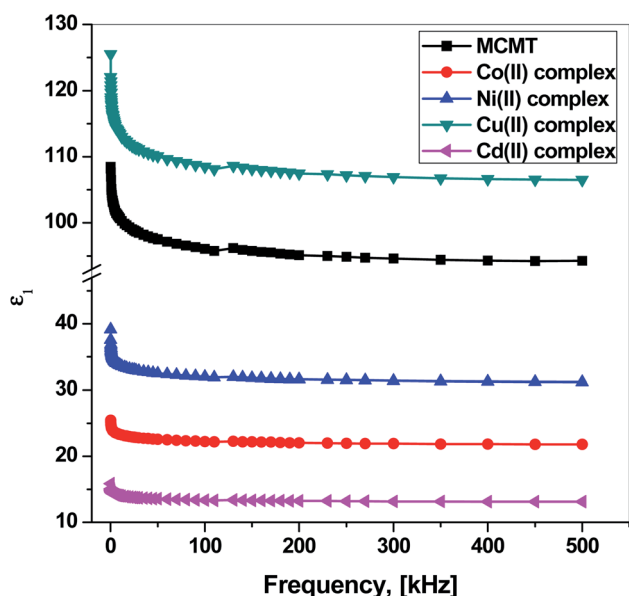


Fig. 12 Frequency dependence of the dielectric constant ϵ_1 for MCMT, Co(II), Ni(II), Cu(II), and Cd(II) complexes.

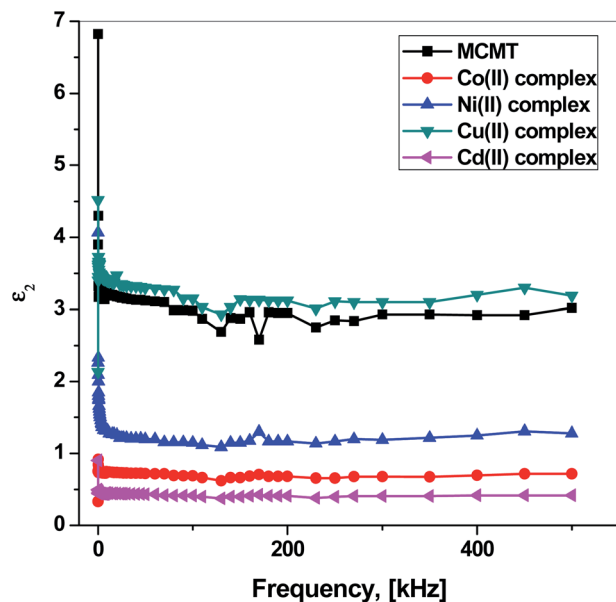


Fig. 13 Frequency dependence of the dielectric loss ϵ_2 for MCMT, Co(II), Ni(II), Cu(II), and Cd(II) complexes.

increasing frequency for all the complexes. Then, ϵ_1 decreases gradually with increasing frequency. At high frequencies, ϵ_1 decreases slightly. When the sample is placed in an alternating electric field, the dielectric properties can be interpreted by considering a set of dipoles. The dielectric constant represents the degree of polarization of the material and its ability to store the electric field.¹⁰⁶ Thus, Cu(II) complex is considered to have the highest polarization and storage ability.

At low frequencies, the dielectric constant ϵ_1 for the polar material is due to the contribution of multicomponent of polarizability, deformational polarization (electronic and ionic polarization), and relaxation polarization (orientation and interfacial polarization).¹⁰⁷ When the frequency is increased, the dipoles will no longer be able to rotate rapidly, so that their oscillations begin to lag behind those of the field. So, the orientation and ionic contribution decreases, leading to decreasing ϵ_1 .¹⁰⁸ As the frequency is further increased, the dipole will be unable to follow the field completely, stopping the orientation and ionic polarization. Thus, ϵ_1 approaches approximately to a constant value at higher frequencies due to the interfacial contribution only.⁹⁷

Fig. 13 shows the frequency dependence of the dielectric loss ϵ_2 , which represents the dissipation of energy.¹⁰⁶ Clearly, the Cu(II) complex has the highest dissipation of energy, while Cd(II) complex has the lowest dissipation of energy. At low frequencies, ϵ_2 is high and decreases with increasing frequency. Then, it starts to decrease slightly with some fluctuation. At lower frequencies, the high values of ϵ_2 is due to many sources such as ion jump, conduction loss of charge carriers, ion polarization, and ion vibration. At higher frequencies, the ion vibrations could be the only source for the dielectric loss.⁹⁷



4. Conclusion

The current paper introduced the synthesis, structure, and physicochemical properties of the new complexes of Co(II), Ni(II), Cu(II), and Cd(II) ions with (1E)-1-((6-methyl-4-oxo-4H-chromen-3-yl)methylene)thiocarbonohydrazide ligand (MCMT). The MCMT was able to form complexes with metal ions by coordination through N and S atoms. The complexes are formed in 1 : 2 (ML₂) ratio and exhibit octahedral geometry. The crystal study indicates that the complexes are in nanodomain. The results of DFT-B3LYP calculations are found to be in good agreement with the infrared spectral analysis, and the complexes prefer the *all trans* geometrical isomer with coordination stability in the order of Cu(II) > Ni(II) > Co(II). The ligand and its complexes have indirect allowed optical transitions. Frequency response of the AC conductivity of the ligand and its complexes refers to correlated barrier hopping between intimate valence alternation pairs. Cu(II) complex has the highest response for the AC conductivity to the frequency and the highest polarization and storage ability.

Conflicts of interest

The authors declare that they have no known competing financial interests or personal relationships that could have appeared to influence the work reported in this paper.

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