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Schiff base pyrazolone complexes of iron (III): synthesis, characterization, antimicrobial and antioxidant activity

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Abstract A series of novel Schiff base tetradentate ligands and its iron (III) coordination compounds were synthesized, characterized and evaluated for antibacterial and antioxidant activity. All the synthesized ligands and complexes were characterized by spectroscopic and crystallographic techniques. Electronic spectra, Mössbauer spectra, magnetic moment and conductance study evidence the fact of octahedral arrangement around iron (III). The in vitro antimicrobial Schiff base complexes bear polar and nonpolar properties together; this makes them suitable for permeation to the cells and tissues, it enhances the activity against the bacteria. The antioxidant properties were measured with DPPH (2,2'-diphenyl-2-picrylhydrazyl). These properties were due to the unique feature of ligands such as highest lipophilicity, lowest electron withdrawing power and highest polarisability.

Keywords Schiff base pyrazolone \cdot Iron (III) complexes \cdot Antimicrobial \cdot Antioxidant

Introduction

Recently, one important branch of pharmaceutical chemistry-synthesis, and research of the Schiff base and their metal complexes has attracted many attention. (Wang et al.,

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2007; Ali et al., 2002; Sabry, 2006; Zitouni et al., 2001; Casas et al., 2008; Youseef and Abbady, 2014) Pyrazolone, as a prominent structural motif is found in numerous biologically active compounds (Qi and Ready, 2007; Kalluraya et al., 2014). Typical pyrazole derivatives are pyrazole-3ones (pyrazolones) whose synthesis, reactivity and numerous application are well-documented (Varvounis et al., 2001). Due to the trouble-free preparation and its rich biological activity of broad-spectrum antibacterial action, antisepsis (Wang et al., 2007; Yang et al., 2001; Kimata et al., 2007; Costa et al., 2006), antipyretic and analgesics (Uramaru et al., 2010), antitumor (Casas et al., 2008; Camacho et al., 2013), antituberculosis (Botta et al., 2008), antidepressant, and anticonvulsant activities (Aziz et al., 2009), pyrazolone and its complexes have both received considerable attention in coordination chemistry (Surati and Thaker, 2006; Thaker et al., 2007) and medicinal chemistry (Huang et al., 2012; Ruf et al., 2012).

However, to best our knowledge, very little has appeared on the antioxidant (Yang, 2002; George et al., 2014) properties of Schiff base pyrazolone coordination compounds, and its suitability against bacteria and fungus. Many reports available about the antioxidant and antimicrobial activity of antipyrine based derivatives (Vieirera et al., 2010; Nakagawa et al., 2006).

Iron have also a critically important metal for a wide variety of cellular events, indeed no life form is possible without the presence of this element (Bailar et al., 1975). Iron holds this central position by virtue of its facile redox chemistry and its high affinity for oxygen (Liu and Hider, 2002; Andrews, 1999).

Due to the large structural variability, biological importance and our interest in this area is focused for a considerable time on the investigation of coordination chemistry of transition metal with using pyrazolone-based ligands, (Surati and

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Tanker, 2010; Surati et al., 2008; Surati, 2011; Vyas et al., 2011) initiate us to design iron (III) Schiff base complexes and evaluate systematically their antimicrobial and antioxidant activity. The aim of this work is to evaluate and come across the suitability of these coordination compounds for antioxidant and antimicrobial activities.

Experimental

Materials and methods

The solvents were used after purification using the standard method described in literature (Perrin et al., 2008; Gordon et al., 1978). 1-Phenyl-3-methyl-2-pyrazoline-5-one (E-Merck); ethylenediamine (BDH); *o*-phenylenediamine (Fluka AG, Switzerland) and FeCl₃·6H₂O, (Aldrich) were used as received.

Instruments

Elemental analyses (C, H, N) were performed at CDRI, Lucknow. Infrared spectra were recorded with a Perkin-Elmer IR spectrophotometer (4000–50 cm⁻¹) using KBr pellets. ¹H-NMR spectra of the compound was recorded with Bruker-400 using CDCl₃ as a solvent and Tetramethylsilane as an internal reference. Mass spectra (EI) obtained on a JEOL D-300 mass spectrometer at SAIF, IIT Madras, Chennai. The FAB mass spectra were recorded on a JEOL SX 102 mass spectrometer using Argon/Xenon (6kV, 10mA) as the Fast atom bombardment (FAB) gas. The accelerating voltage was 10kV and the spectra were recorded at room temperature with using *m*-nitrobenzyl alcohol as matrix. Electronic spectra in the 200-800 nm range were obtained in acetone on a "SHIMADZU" UV 160 A using quartz cell of 1 cm³. Magnetic measurements were carried out at room temperature by the Gouy method using Hg[Co(SCN)₄] as calibrant. Molar conductance of the Schiff base complexes was determined on Systronics direct reading conductivity meter type CM-82T. Thermal gravimetric analysis (TGA) was carried out on universal V3.0G Thermal Ananlysis (TA) instrument in the range 0-700 °C at a heating rate of 10 °C/min under nitrogen. Differential scanning calorimetry (DSC) was carried out on universal V3.0G TA instrument in the range 0-300 °C. The Mössbauer spectra were collected using a Cryophysics MS-1 microprocessor-controlled spectrometer operating in the constant acceleration mode. The source was 25 mCi ⁵⁷Co/Rh and spectra were fitted using a standard nonlinear least square package. X-ray structure analyses Powder diffraction diagrams were obtained on a Philips PW 1710 diffractometer (Cu/Ka; scan rate 1 min^{-1}) with a graphite monochromator.

Synthesis of ligands

3-Methyl-5-oxo-1-(p-tolyl)-4,5-dihydro-1H-pyrazole-4-carbaldehyde (**1a**) and p-(4-formyl-4,5- dihydro-3-methyl-5oxopyrazol-1-yl)benzenesulphonic acid (**1b**) were prepared by literature (David and Straley, 1961).

Synthesis of Schiff base ligands H_2L^1 , H_2L^2 , H_2L^3 , and H_2L^4

The Schiff bases were synthesized by refluxing for two and half hour in methanolic solution of ethylenediamine (0.60 g, 0.01 mol) or *o*-phenylenediamine (1.08 g, 0.01 mol) with (1a) (4.32 g, 0.02 mol) or (1b) (5.64 g, 0.02 mol) in 20 cm³ methanolic solution. A solid mass separated, was collected and washed by ether. Crystallization was done with ethanol and dried over CaCl₂ (Surati and Tanker, 2010).

Characterization of H_2L^1 , H_2L^2 , H_2L^3 , and H_2L^4 (See Supplementary Materials)

Synthesis of Schiff base complexes of iron (III) with ligands H_2L^1 , H_2L^2 , H_2L^3 , and H_2L^4

All the iron (III) complexes of the Schiff base were prepared by the following general method. The salt $\text{FeCl}_3 \cdot 6\text{H}_2\text{O}$ (2.70 g, 0.01 mol) was dissolved in water (15 mL) and the solution was added to a warm (~65 °C), stirred solution of the corresponding Schiff bases H_2L^1 (4.56 g, 0.01 mol), H_2L^2 (5.88 g, 0.01 mol) or H_2L^3 (5.04 g, 0.01 mol), or H_2L^4 (6.36 g, 0.01 mol) in 1:1 mixed solvent (DMF + methanol) (25 mL). The reaction mixture was refluxed for 3–4 h and then concentrated to half of its volume. The resulting precipitate was filtered by suction and dried in vacuum over anhydrous CaCl₂.

Results and Discussion

The newly synthesized tetradentate ONNO donor Schiff base complexes of iron (III) are stable at atmospheric temperature and pressure. They are insoluble in organic solvents like methanol, ethanol, and acetone. However, Schiff base ligand and its complexes are soluble in coordinating solvents like DMF and DMSO, and mixed solvents (DMF + methanol), (DMF + ethanol) etc. The analytical data are in good evidence with the proposed structure of Schiff base ligands as well as complexes (Scheme 1 and 2). All the analytical data are summarized in Table 1.

¹H-NMR Spectra

The ¹H-NMR spectra of Schiff base ligands H_2L^1 , H_2L^2 , H_2L^3 , and H_2L^4 were carried out in DMSO-d₆ at room

Scheme 1 Synthetic pathway for tetradentate Schiff base ligands H_2L^1 , H_2L^2 , H_2L^3 , and H_2L^4



temperature. Ligands H_2L^1 and H_2L^2 show triplet at ~1.5 ppm for methylene proton of ethylenediamine. Ligands H_2L^1 and H_2L^3 show two singlets with integration of six protons near the δ ~2.35 and ~2.79 ppm for methyl group attached to pyrazolone ring and phenyl ring, respectively (Silverstein et al., 1991). The enolic nature of all Schiff base ligand shows broad singlet at δ ~10.52 ppm, due to rapid exchange interaction of keto-enol tautomerism (Surati and Tanker, 2010). The sharp singlet at δ ~8.91 ppm, observed due to aldehydic proton in all Schiff base ligand. The phenyl multiplet observed in the range of δ ~7.31-7.98 ppm.

Infrared Spectra

Characteristic frequency of FT-IR spectra were summarized as above in Table 2. The infrared spectra of the ligand shows a broad band in a region $3100-3600 \text{ cm}^{-1}$, which may be due to v (OH). Free v (OH) is generally observed between $3500-3600 \text{ cm}^{-1}$. The low value of this band is due to intermolecular or intramolecular H-bonding (Silverstein et al., 1991; Jadeja et al., 2004), which suggests the presence of keto-enol tautomeric form, at least in the solid state. The same is also inferred from its NMR signal at about $\delta \sim 10.52$ ppm due to an enolic proton. The phenyl group shows v (CH) at 3060 cm⁻¹ and v (C=C) at 1540 cm⁻¹. The bands at ~1640, ~1594, and ~1300 cm⁻¹ may be assigned to v (C=N) (azomethine), v (C=N) (pyrazoline ring), and v(C–O), respectively (Surati and Thaker, 2006; Thaker et al., 2007; Silverstein et al., 1991).

The ligands as well as its corresponding complexes show absorptions in the regions $3000-2800 \text{ cm}^{-1}$, which may be due to v(C-H). The Schiff base ligand of the present study shows a strong band due to δ (O-H) in the region $1210-1270 \text{ cm}^{-1}$ Although all of the metal complexes show this band, its intensity is found to be lower than that of ligand (Nakamoto, 1978; Raj et al., 1992). This may be due to the deprotonation of the 5-OH group of the ligand (Nakamoto, 1978; Surati et al., 2008). All of the metal chelates show bands in the region $3100-3600 \text{ cm}^{-1}$, which may be due to the presence of water molecules (Dwyer and Mellor, 1962). As shown by TGA of all the metal chelates. this band may be assigned to the one coordinated water molecules, which can be also inferred from bands at \sim 710 cm^{-1} and may be due to the bending modes of vibrations of the water (Dwyer and Mellor, 1962).

A characteristic band of intermediate intensity is found at $\sim 1300 \text{ cm}^{-1}$, which may be due to v(C-O) (Figgis and Lewis, 1960). On coordination, this band is shifted toward higher frequency indicating that the oxygen of the 5-OH group of the pyrazoline ring of the ligand has taken part in the coordination (Figgis and Lewis, 1960; Surati and

Scheme 2 a Schiff base complexes [Fe(L¹)ClH₂O]; R=CH₃ and [Fe(L²)ClH₂O]; R=SO₃H, b Schiff base complexes [Fe(L³)ClH₂O]; R=CH₃ and [Fe(L⁴)ClH₂O]; R=SO₃H



Tanker, 2010; Surati et al., 2008). The observed low-energy shift of v (C=N) (azomethine) in the metal chelates suggests nitrogen coordination. While v (C=N) (cyclic) observed at the same energy in the chelates indicates the nonparticipation of the cyclic nitrogen in the coordination (Figgis and Lewis, 1960; Dwyer and Mellor, 1962; Surati and Tanker, 2010; Surati et al., 2008). The Fe–N stretching vibration for complexes was observed at 258 cm⁻¹ and that for Fe–Cl is seen at 381 cm⁻¹ in the anion species. The stretching vibrations of Fe–N in complexes are consistent with the high-spin character of these complexes. Low spin frequencies for Fe–N appear around 50–70 cm⁻¹ higher frequency to that of high-spin complexes (Raj et al., 1992).

Electronic Spectra

UV-visible spectra of iron (III) Schiff base complexes were carried out in DMF. The data are summarized in Table 2.

The electronic spectra of Schiff base complexes of iron (III) exhibit *d-d* absorption bands at 427-445 and 523–599 nm. The first band may be assigned to the ${}^{6}A_{1} \rightarrow {}^{4}A_{1}$, ${}^{4}E(G)$ transition, while the second band would be due to the ${}^{6}A_{1} \rightarrow {}^{4}T_{2}(G)$ transition, suggesting an octahedral arrangement around the iron (III) ion (Lever, 1968). Also the charge transfer bands observe at higher energy such as ~276 and ~367 nm. The band observed at ~550 for [Fe (L¹)·(Cl)·(H₂O)] and [Fe·(L³)·(Cl)·(H₂O)], ~556 for [Fe·(L²)·(Cl)·(H₂O)] and [Fe·(L⁴)·(Cl)·(H₂O)]. Second band observed at 613 nm for [Fe·(L³)·(Cl)·(H₂O)] and ~648 nm for other iron (III) complexes.

Mass Spectra

Mass spectra of tetradentate Schiff base ligands H_2L^1 , H_2L^2 , H_2L^3 , and H_2L^4 were good agreements with their proposed structure. Melting point of all the ligands is higher (above 200 °C), as results of this, the mass spectra we carried out by EI. Cleavage take place in the Schiff base ligands and

fragments observed were good agreement with structure of ligands (Johnstone, 1972; Mclafferty, 1973).

Molecular mass of all the metal complexes performed with FAB mass spectrometer. A FAB mass spectrum of representative complex [Fe(H₂L¹)·Cl·H₂O] were recorded and the largest m/z isotropic peak appeared at 547 which corresponding to the composition [Fe(L¹)·Cl] (563.88 molecular mass $-H_2O$). Fragmentation of a weakly coordinated molecule/ion in FAB mass spectrum is not unlikely. To confirm the presence of weakly coordinating H₂O molecules in the complexes, we carried out thermogravimetric analyses of complexes. However, the coordinated ion Cl was strongly coordinated to metal that confirmed with M+2 peak at half intensity at 549 m/z (Johnstone, 1972; Mclafferty, 1973).

Thermal Study

Thermal analyses of all the complexes were carried out by the Thermogravimetry (TG), Differential thermal analysis (DTA), and Differential scanning calorimetry (DSC) techniques. The experimental results revealed that the degradation occurred in multiple stages, following a complex mechanism. For each stage the kinetic parameters and the thermogravimetric characteristics have been estimated.

Thermal behavior of all complexes are explained below.

$[Fe(L^1) \cdot Cl \cdot H_2O]$

The TG and DTA curves of the $[Fe(L^1)\cdot Cl\cdot H_2O]$ are given following degradation scheme. The compounds slowly started to decompose at 201 °C. The first mass loss (3.13 % est.; 3.19 % calc) up to 231 °C is in good agreement with loss of one coordinated water molecule. The DTA curve shows an endothermic peak at 226 °C. This is immediately followed by exothermic process at 319 °C with mass loss (7.0 % est.; 6.47 % calc) between 294–328 °C, due to the removal of HCl. In the temperature range 356–671 °C with

Complexes	Color	Yield (%)	Melting point (°C)	FW /	Analysis found per	cent (Calc	ulated %)		Mass peak m/e	\wedge (ohm ⁻¹ cm ² mol ⁻¹)	μ _{eff} (B.M.)
					н	z	s	Μ			
$[Fe(L^{1})\cdot(CI)\cdot(H_{2}O)]$ $C_{26}H_{30}N_{6}O_{3}CIFe$	Reddish brown	82	<250 ^a	563.88 5	55.35(55.33) 5.35(5.32) 14.8	86(14.89) —	10.0(10.09)	547	25	5.41
[Fe(L ²)·(Cl)·(H ₂ O)] C ₂₄ H ₂₆ N ₆ O ₉ S ₂ Cl Fe	Reddish brown	76	<250 ^a	695.55 4	1.43(41.40) 3.72(3.73) 12.()4(12.07) 9.12(9.20)	12.42(12.45)) 679	29	5.43
$[Fe(L^3)\cdot(CI)\cdot(H_2O)]$ $C_{30}H_{30}N_6O_3CIFe$	Reddish brown	79	<250 ^a	611.66 5	58.86(58.85) 4.88(4.90) 13.7	1(13.73) —	10.96(10.95)) 594	42	5.42
$[Fe(L^4).(CI)\cdot(H_2O)]$ $C_{28}H_{26}N_6O_9S_2CIFe$	Reddish brown	83	<250 ^a	745.09 4	15.08(45.09) 3.45(3.48) 11.2	25(11.27) 8.58(8.63	13.34(13.34)) 729	32	5.42
^a Decomposition point											

rable 1 Analytical data and physical properties of complexes

mass loss (71.20 est.; 72.18 % calc), due to the pyrolysis of organic molecule. The DTA curve shows an exothermic peak at 556 °C. The final product of the thermal decomposition Fe_2O_3 was determined by metal analysis (Furman, 1962).

[Fe(L²)·Cl·H₂O]

The TG and DTA curves of the $[Fe(L^2)\cdot Cl\cdot H_2O]$ are given following degradation scheme. The compound slowly started to decompose at 211 °C. The first mass loss (2.96 % est.; 2.58 % calc) upto 236 °C is in good agreement with loss of one coordinated water molecule. The DTA curve shows an exothermic peak at 246 °C. This is immediately followed by exothermic process at 564 °C with mass loss (75.62 % est.; 77.43 % calc) between 302–650 °C, due to the removal of HCl and pyrolysis of organic molecule. The final product of the thermal decomposition Fe₂O₃ was determined by metal analysis (Furman, 1962).

$[Fe(L^3) \cdot Cl \cdot H_2O]$

The TG and DTA curves of the $[Fe(L^3) \cdot Cl \cdot H_2O]$ are given following degradation scheme. The compound slowly started to decompose at 235 °C. The first mass loss (2.96 % est.; 2.58 % calc) up to 257 °C is in good agreement with loss of one coordinated water molecule. The DTA curve shows an endothermic peak at 248 °C. This is immediately followed by exothermic process at 329 °C with mass loss (4.59 % est.; 5.96 % calc) between 330-365 °C, due to the removal of HCl. In the temperature range 379–699 °C observed with mass loss (72.90 est.; 74.31 % calc), due to the pyrolysis of organic molecule. The DTA curve shows an endothermic peak at 515 °C. The final product of the thermal decomposition Fe₂O₃ was determined by metal analysis (Furman, 1962).

$[Fe(L^4) \cdot Cl \cdot H_2O]$

The TG and DTA curves of the $[Fe(L^4)\cdot Cl\cdot H_2O]$ are given following degradation scheme. The compound slowly started to decompose at 221 °C. The first mass loss (2.96 % est.; 2.41 % calc) up to 232 °C is in good agreement with loss of one coordinated water molecule. The DTA curve shows an endothermic peak at 253 °C. This is immediately followed by exothermic process at 329 °C with mass loss (4.59 % est.; 5.02 % calc) between 285–342 °C, due to the removal of HCl. In the temperature range 345–710 °C observed with mass loss (76.90 est.; 78.73 % calc), due to the pyrolysis of organic molecule. The DTA curve shows an exothermic peak at 521 °C. The final product of the thermal decomposition Fe₂O₃ was determined by metal analysis (Furman, 1962).

Complexes	ν H ₂ O	δ (H ₂ C	$\nu C = N (Coor$	dinated)	$\nu C = N (c$	yclic)	νC-0) (Str.)) <i>d</i> -	d (nm)	Charge transfer	r band (nm)
[Fe(L1) (Cl) (H2O)]	3486	713	1624		1593		1345		44	2, 552	296,367	
$[Fe(L^2) (Cl) (H_2O)]$	3470	710	1624		1594		1345		42	7, 543	290, 367	
$[Fe(L^3) (Cl) (H_2O)]$	3470	710	1620		1594		1349		44	5, 558	295, 366	
[Fe(L ⁴)(Cl)(H ₂ O)]	3470	_	1624		1598		1348		44 66	1, 546, 5	293, 365	
Table 3 The thermo	dynamic		mpley	Sten	$C (KI \sigma^{-1})$	<u>т</u>	Т.	Т.	T	<u>лн</u> *(н		
and decomposition parameters for complexes		C	mplex	Step	$C_p(\mathbf{K}, \mathbf{g})$ C^{-1}	1 _s	11	- 2	- p	mol^{-1})		
		Δ.5	$\int^{*} (JK^{-1} \operatorname{mol}^{-1})$		ΔE^* (kJ mol ⁻¹)							
		[F	$e(L^{1})$ (Cl) (H ₂ O)]	Ι	1.35	201	223	308	252	115.35	0.457	7.94
				II								21.36
				III								45.62
		[F	$e(L^2)$ (Cl) (H ₂ O)]	Ι	2.48	211	214	301	253	119.26	0.471	8.41
				II								20.60
				III								49.68
		[F	$e(L^3)$ (Cl) (H ₂ O)]	Ι	1.38	235	225	309	253	116.23	0.459	10.29
				II								26.69
				III								53.62
		[F	$e(L^4)$ (Cl) (H ₂ O)]	Ι	1.30	221	221	309	251	114.89	0.457	9.94
				II								22.36
				III								49.29

Table 2 FT-IR and electronic spectral data of complexes and their assignments

From above results it is confirmed that iron (III) complexes contain one coordinated water molecule. Also one coordinated ion chloride and acetate ions are present in iron (III). Schiff base complexes of iron (III) shows similar behavior except [Fe(L¹)·Cl·H₂O], in this complexes degradation of chloride and organic ligand take place in single step, while in other complexes of iron (III) shows three step degradation. The thermal stability of these complexes is mentioned in Table 3. From the above results, order of thermal stability of these complexes can be given as follows.

$$\begin{split} [\text{Fe}(L^4).\text{Cl}.\text{H}_2\text{O}] > [\text{Fe}(L^3).\text{Cl}.\text{H}_2\text{O}] \\ > [\text{Fe}(L^2).\text{Cl}.\text{H}_2\text{O}] > [\text{Fe}(L^1).\text{Cl}.\text{H}_2\text{O}] \end{split}$$

Differential scanning calorimetry (DSC)

Differential scanning calorimetry (DSC) curves are show exothermic process. The complexes of iron (III) exhibited exothermic process at temperature ~279 °C corresponds to the melting point of the complexes (Bell, 1969). The area of the endothermic peak corresponding to the heat of fusion and the peak temperature corresponds to the melting point. The melting (T_p), transition temperature (T_1 , T_2), heat of reaction (Δ H), and entropy (Δ S) of the complexes were calculated from DSC results and are given in Table 3. The heat capacities C_p of the complexes were calculated from DSC results and are summarized in Table 3.

Mössbauer study

Mössbauer spectra were carried out in solid sample to study the environment surrounding of the nucleus. This technique now finds wide application in diverse fields, such as solid state physics, metallurgy, chemistry, and biochemistry. Mössbauer studies of iron containing complexes attended more interest and are important to know the metal environment in metal complexes. (Greenwood and Gibb, 1971; Gibb, 1976) Mössbauer spectra of representative complexes $[Fe(L^3)\cdot Cl\cdot H_2O]$ and $[Fe(L^4)\cdot Cl\cdot H_2O]$ at room temperature and liquid nitrogen temperatures have been studied. The room temperature spectrum of [Fe(L³)·Cl·H₂O] (Supplementary Materials fig. 1S) shows two lines with isomer shift, $\delta = 0.46 \text{ mm s}^{-1}$ and quadrupole splitting, $\Delta E_q = 0.64 \text{ mm}$ s⁻¹. The LNT spectrum also shows two lines with isomer shift, $\delta = 0.58 \text{ mm} \text{ s}^{-1}$ and quadrupole splitting, $\Delta E_{\rm q} = 0.73 \,\rm mm \, s^{-1}$ (Supplementary Materials fig. 2S). Similar mössbauer behavior is observed for complex [Fe(L³)·Cl·H₂O] with isomer shift, $\delta = 0.43 \text{ mm s}^{-1}$,

quadrupole splitting, $\Delta E_q = 0.61 \text{ mm s}^{-1}$ at room temperature and $\delta = 0.55 \text{ mm s}^{-1}$, $\Delta E_q = 0.71 \text{ mm s}^{-1}$ at LNT. The results indicate that iron (III) complexes are high-spin with octahedral geometry (Long, 1984).

X-ray diffraction study

From X-ray diffraction analysis we have successfully calculated the unit cell parameters and crystal systems of Schiff base complexes of iron (III) by using Ito's method (Azaroff and Buerger, 1958). The *d*-spacing are listed in Table 4 in a decreasing order of values. Iron (III) complexes shows simple cubic crystal system with six coordinates the a=b=c; $\alpha=\beta=\gamma=90$. Relative intensity, *d*-spacing, angle (2 θ) and miller indices were calculated. All the data are well summarized in Table 4.

Molar conductance

The observed molar conductances of the iron (III) complexes in 10^{-3} molar DMF solution are in the range 25–42 ohm⁻¹ cm² mol⁻¹ summarized in Table 1, which are consistent with the nonelectrolytic nature of all the complexes (Geary, 1971).

Magnetic measurements

The magnetic moment (μ_{eff}) per metal atom of the iron (III) complexes ~5.61 B.M. falls in the range reported for high-spin octahedral iron (III) complexes (Table 1). This statement is also supported by the mössbauer results (Long, 1984; El-tabl et al., 2005).

Antibacterial activity

Antibacterial activity of Schiff base $(H_2L^1-H_2L^4)$ and the corresponding iron (III) complexes $[Fe(L^1)(Cl)(H_2O)]$ -[Fe $(L^4)(Cl)(H_2O]$ were studied against two Gram positive bacteria B. subtilis, S. aureus (MTCC-619 and MTCC-69, respectively), two Gram negative bacteria E. coli, K. pneumonia (MTCC-722 and MTCC-109, respectively). The reason behind the choice of selection of Gram positive bacteria named S. aureus and B. subtilis, are normally present in respiratory tract and skin of human plays important role in nitrate reduction and catalase mechanism, also ideal choice for chromosome replication in bacteria, enzyme production, and Gram negative bacteria's like Escherichia coli who are harmless but have a significant role in food poisoning by its serotypes, because according to its serotypes they are categorized into enterohemorrhagic or pathogenic form and probiotic or nonpathogenic form. Probiotic forms are those who are unable to produce enterotoxins or cytotoxins, whereas pathogenic forms are

Table 4 Powder X-ray diffraction data of the [Fe(L₂) (Cl) (H₂O)]

20	Relative I	d	d^2	d_I^2/d^2	hkl	<i>N</i> *	$a_{\rm calc}^{**}$
18.02	36.20	4.6362	21.494	1	111	3	8.030134
34.0	43.5	2.8577	8.1664	2.6320	_	8	8.082796
36.09	26.3	2.4890	6.1951	3.4695	310	10	7.870909
46.90	100	2.4465	5.9853	3.5911	311	10	7.736512
48.60	20.8	2.0212	4.0852	5.2614	400	16	8.0848
56.09	46.2	1.6520	2.7291	7.8758	422	24	8.093114

* $N=h^2 + k^2 + l^2$

** $a=d(N)^{1/2}$

those that play role in food poisoning, e.g., Nissle studied the probiotic strains of *Escherichia coli* in 1917 and referred EcN who are non-uropathogenic and noninvasive as well (Sonnenborn and Schulze, 2009). As well as serotypes 0157:H7 who are among the six pathotypes of *E. coli* bacterial strain and can produce Verotoxins (VTEC) and Shiga like toxin (SLTs) that are related to the vast range of human infections like hemorrhagic colitis and hemolytic uremic syndrome (Broseta, 2001), similarly *K. pneumoniae* are important pathogens in nosocomial infections.

The susceptibility of certain strains of bacteria to the ligand and their corresponding Schiff base complexes were evaluated by agar diffusion method (Singh and Kushawaha, 2001; Sakıyan et al., 2004); so depending upon the area or zone of inhibition, with respect to the selected range of concentration of our designed molecule to determine the minimum inhibitory concentration aspects, at minimum concentration our designed molecule showed considerable increase in antibacterial activity. Table 5 shows the results for all compounds and comparison of MIC value (in $\mu g \text{ cm}^{-3}$) of iron (III) complexes and standard drug streptomycin against those of different bacteria. The tetradentate Schiff base ligand shows antibacterial activity starting from $150-350 \,\mu g \, cm^{-3}$ against four strains of bacteria. It is very interesting that iron (III) complexes show higher activity as compared with Schiff base (azomethine) ligands, because iron (III) Schiff base complexes bear polar and nonpolar properties together; this makes them suitable for permeation to the cell and tissue (Rehder, 2003). Changing the hydrophilicity and lipophilicity probably leads to bring down the solubility and permeability barrier of cell, which in turn enhances the bioavailability of chemotherapeutics on one hand and potentiality at another. (Surati and Tanker, 2010; Surati, 2011; Rehder, 2003)

Antioxidant activity

There are different methods used to evaluate the in vitro antioxidant capacity of isolated compounds, mixtures of **Table 5** Antibacterial activity (concentration 2.5–7.0 µg/mL) of ligand and Iron (III) complexes

Compounds	Gram positive	e	Gram negativ	/e	Scavenging activity (%)
	MTCC-619	MTCC-96	MTCC-722	MTCC-109	
H_2L^1	26	24	27	32	59
H_2L^2	28	29	30	34	41
H_2L^3	38	34	35	33	66
H_2L^4	41	46	42	48	48
$[Fe(L^1)(Cl)(H_2O)]$	52	55	57	65	57
$[Fe(L^2)(Cl)(H_2O)]$	50	54	52	58	48
$[Fe(L^3)(Cl)(H_2O)]$	56	56	58	60	69
$[Fe(L^4)(Cl)(H_2O)]$	57	53	54	58	51
Ref.	27	33	38	44	49

MTCC-619 – Bacillus subtilis, MTCC-96 – Staphylococcus aureus, MTCC-722 – Escherichia coli, MTCC-109 – Klebsiella pneumonia, Ref. Streptomycin used as antibacterial standard

compounds, biological fluids, tissues, and synthetic compounds that involve different mechanisms of determination of antioxidant activity, e.g., chemical methods based on scavenging of ROS or RNS such as peroxynitrite (Heijnen et al., 2001), the hydroxyl radical and superoxide. Other methods measure the disappearance of free radicals using spectrophotometry, such as ABTS + (2,2'-azinobis-(3-ethyl-benzothiazoline-6-sulphonate) cation radical (Van den Berg et al., 2000; Manojkumar et al., 2009) or DPPH (2,2'-diphenyl-1-picrylhydrazyl) (Benzie and Strain, 1996). Other assays to determine the total antioxidant power include techniques such as the ferric reducing/antioxidant power method (Benzie and Strain, 1996) or use the in situ electrochemically generated bromine (Abdullin et al., 2001). The results in the measurement of antioxidant capacity depend on the method used.

In the present paper, we have carried out scavenging effect on DPPH. 1.5 mL methanolic solution of the synthesized compounds (0.2 mM) was added to 1.5 mL (0.2 mM) solution of DPPH radical in methanol (final concentration of DPPH and synthesized compounds was 0.1 mM). The mixture was shaken vigorously and allowed to stand for 30 min. After this the absorbance at 517 nm was determined and the percentage of scavenging activity was calculated using the formula shown below. Ascorbic acid was used as the reference compound. All tests and analyses were undertaken on three replicate and the results were averaged. The results are given Table 5.

Scavenging activity $(\%) = \{[(Ab + As) - Am]/Ab\}X 100\%$

Ab: absorbance of 0.1 mM methanolic solution of DPPH at 517 nm, As: absorbance of 0.1 mM methanolic solution of test compounds at 517 nm, Am: absorbance of methanolic mixture of the drug and DPPH at 517 nm. All the ligands and iron (III) complexes shows good antioxidant activity though this activity are not change significantly

with complexation in case of $[Fe(L^4) \cdot Cl \cdot H_2O]$ and $[Fe(L^2) \cdot Cl \cdot H_2O]$. While in case of $[Fe(L^3) \cdot Cl \cdot H_2O]$ and $[Fe(L^1) \cdot Cl \cdot H_2O]$ the activity were remarkable improve. The effect of substituents on the phenyl ring of pyrazolone-based ligands H_2L^1 , H_2L^2 , H_2L^3 , H_2L^4 in producing antioxidant activity in the descending order were found to be $H_2L^3 > H_2L^1 > H_2L^4 > H_2L^2$. These trends were observed due to substituents having highest lipophilicity, lowest electron withdrawing power and highest polarisability. In the same way after complexation the antioxidant activity in the decreasing order were found $[Fe(L^3) \cdot Cl \cdot H_2O] > [Fe(L^1) \cdot Cl \cdot H_2O] > [Fe(L^4) \cdot Cl \cdot H_2O] > [Fe(L^2) \cdot Cl \cdot H_2O]$, it indicates that the antioxidant properties is largely due to the Schiff base ligands (Van den Berg et al., 2000).

Conclusion

On the basis of above studies, the general structure of the Fe (III) complexes are proposed as shown in Scheme 2. The Schiff base ligands are behaving as O, N, N, O donor tetradentate ligands. Also the chloride and coordinated water is coordinated with metal ion confirmed from FT-IR, thermal study. Electronic spectra, Mössbauer spectra, magnetic moment, and conductance study evidence the fact of octahedral arrangement around iron (III). Moreover, biological screening state this Schiff base complexes enhance the activity against the bacteria and fungi due to Schiff base complexes bear polar and nonpolar properties together, this makes them suitable for permeation to the cells and tissues. The antioxidant properties of ligands are due to highest lipophilicity, lowest electron withdrawing power, and highest polarisability.

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Compliance with ethical standards

Conflict of interest The authors declare that they have no conflict of interest.

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