

# **A simple spectrophotometric method for the determination of iron(II) aqueous solutions**

## **Introduction**

Iron is an absolute requirement for most forms of life, including humans and most bacterial species. Plants and animals all use iron, and it can be found in a wide variety of food sources.<sup>1</sup> The industrial uses of Fe and its compounds are numerous. <sup>2</sup> It is the major constituent in steel making. Several Fe oxide forms find use as paint pigments, polishing compounds, magnetic inks, and coatings for magnetic tapes. The soluble salts are variously used as dyeing mordant, catalysts, pigments, fertilizer, feeds, and disinfectants, and in tanning, soil  
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conditioning, and treatment of sewage and industrial wastes. <sup>3</sup> Divalent Fe is a cofactor in heme enzymes such as catalase and cytochrome C, and in non-heme

enzymes such as aldolase and tryptophan oxygenase. 2 In humans iron is an essential component involved in oxygen transport. 4,5 It is also essential for the regulation of cell growth, and differentiation of iron limits oxygen delivery to cells,6 resulting in fatigue, poor work performance, and decreased immunity. 4 On the other hand, excess amounts of iron can result in toxicity and even death.7

Toxicology considerations are important in terms of iron deficiency (anemia) and accidental acute exposure and chronic iron overload due to idiopathic hemochromatosis or as a consequence of excess dietary iron or frequent blood transfusions. The immediate cause of death from the inorganic compounds of Fe in animals is respiratory failure. Clinical signs preceding death are anorexia oligodipsia, oliguria, alkalosis, diarrhea, loss of body weight, hypothermia, and alternating irritability and depression. In human poisonings,

symptoms of iron intoxication include vomiting, cirrhosis of the liver, hemochromatosis, diarrhea, lethargy, coma, irritability, seizures, and abdominal pain.<sup>3</sup> All these findings cause great concern regarding public health, demanding accurate determination of this metal ion at trace and ultra-trace levels.

Spectrophotometry is essentially a trace analysis technique and is one of the most powerful tools in chemical analysis. Morin has been reported as a spectrophotometric reagent for aluminum,<sup>8</sup> but has not been used previously for the spectrophotometric determination of iron. This paper reports on its use in a very sensitive, highly selective spectrophotometric method for the trace determination of iron. The method possesses distinct advantages over existing methods<sup>9–19</sup> with respect to sensitivity, selectivity, range of determination, simplicity, speed, pH/acidity range, thermal

stability, accuracy, precision, and ease of operation. A comparison between existing methods<sup>9-19</sup> and the present method is shown in Table 1. The present method is also superior to even recently developed spectrophotometric methods<sup>20-26</sup> using morin with respect to sensitivity, selectivity, range of determination, simplicity, rapidity, pH/acidity range, and ease of operation. The method is based on the reaction of non-absorbing morin in slightly acidic solution (0.0001-0.0002M H<sub>2</sub>SO<sub>4</sub>) with iron(II) to produce a highly absorbent light green chelate product, followed by a direct measurement of the absorbance in an aqueous solution. With suitable masking, the reaction can be made highly selective and the reagent blank solution does not show any absorbance.

## **Experimental**

### **Apparatus**

A Shimadzu (Kyoto, Japan) (Model-1601)

double-beam UV-VIS spectrophotometer and a Jenway (UK) (Model-3010) pH meter with combined electrodes were used for the measurement of absorbance and pH, respectively.

A Hitachi polarized Zeeman (Model-Z-5000) atomic absorption spectrometer equipped with a microcomputer controlled air acetylene flame was used for comparison of the results.

### **Reagents and solutions**

All chemicals used were of analytical-reagent grade or the highest purity available. Doubly distilled deionized

water was used throughout this study. Triply distilled ethanol (from lime) was also used.

Glass vessels were cleaned by soaking in acidified solutions of  $\text{KMnO}_4$  or  $\text{K}_2\text{Cr}_2\text{O}_7$ , followed by washing with concentrated

$\text{HNO}_3$ , and were rinsed several times with high-purity deionized water. Stock solutions and environmental

water samples (1000 mL each) were kept in polypropylene bottles containing 1 mL of concentrated  $\text{HNO}_3$ .

More rigorous contamination control was applied when the iron levels in specimens were low.

### **Morin solution ( $5.05 \times 10^{-3} \text{ M}$ )**

Morin solution was prepared by dissolving the requisite amount of morin (BDH Chemicals) in a known volume

of triply distilled ethanol. More dilute solutions of the reagent were prepared as required.

### **Iron(II) standard solution ( $1.79 \times 10^{-2} \text{ M}$ )**

A 100 mL amount of stock solution (1.00 mg  $\text{mL}^{-1}$ ) of divalent iron was prepared by dissolving 497 mg of

purified-grade (Merck pro analysis grade)

$\text{FeSO}_4 \cdot 7\text{H}_2\text{O}$  in deionized water. More dilute standard solutions were

prepared by appropriate dilution of aliquots from the stock solution with deionized water as and when required.

Concentrations were checked using the standard potassium dichromate solution. 27

### **Iron(III) standard solution**

A 100 mL amount of stock solution (1.00 mg mL<sup>-1</sup>) of trivalent iron was prepared by dissolving 290 mg of ferric chloride (Aldrich A.C.S. grade) in doubly distilled deionized water containing 1-2 mL of nitric acid (1+1).

Concentrations were checked using standard potassium dichromate solution. More dilute standard solutions were prepared from this stock solution as and when required.

### **Potassium permanganate solution**

A 1.00% (w/v) potassium permanganate (Merck) solution was prepared by dissolving in deionized water.

Aliquots of this solution were standardized with oxalic acid.

### **Standard potassium dichromate solution**

A 100 mL amount of standard stock solution (0.1 M) was prepared by dissolving 1.4711 g of finely powdered K<sub>2</sub>Cr<sub>2</sub>O<sub>7</sub> (Merck) in 100 mL of deionized water.

### **Sodium azide solution**

Sodium azide solution (2.5% w/v) (Fluka purity > 99%) was freshly prepared by dissolving 2.5 g in

100 mL of  
deionized water.

### **Tartrate solution**

A 100 mL stock solution of tartrate (0.01% w/v) was prepared by dissolving 10 mg of ACS-grade (99%) potassium sodium tartrate tetrahydrate in (100 mL) deionized water.

### **Aqueous ammonia solution**

A 100 mL solution of aqueous ammonia was prepared by diluting 10 mL of concentrated  $\text{NH}_3$  (28%-30%, ACS grade) into 100 mL with de-ionized water. The solution was stored in a polypropylene bottle.

### **EDTA solution**

A 100 mL stock solution of EDTA (0.01% w/v) was prepared by dissolving 10 mg of A.C.S.-grade ( $\geq 99\%$ ) ethylenediaminetetraacetic acid as disodium salt dihydrate in (100 mL) deionized water.

### **Other solutions**

Solutions of a large number of inorganic ions and complexing agents were prepared from their Analar grade or



equivalent grade water-soluble salts (or the oxides and carbonates in hydrochloric acid).<sup>27</sup>

## **Procedure**

A volume of 0.1-1.0 mL of a neutral aqueous solution containing 0.1-100  $\mu\text{g}$  of iron(II) in a 10-mL calibrated

flask was mixed with a 1:40-1:95-fold molar excess of the morin reagent solution

(preferably 1 mL of  $5.05 \times 10^{-3}$

M) followed by the addition of 1-2 mL

(preferably 1 mL) of 0.001 M sulfuric acid. After 1 min, 4 mL of ethanol

was added and the mixture was diluted to the mark with deionized water. The absorbance was measured at

415 nm against a corresponding reagent blank.

The iron content in an unknown sample was determined using

a concurrently prepared calibration graph.

## **Results and discussion**

### **Factors Affecting the Absorbance**

#### **Absorption spectra**

The absorption spectra of the Fe(II)-morin system in 0.001 M  $\text{H}_2\text{SO}_4$  medium were recorded using the spectrophotometer.

The absorption spectra of the Fe(II)-morin were a symmetric curve with maximum absorbance at 415 nm and the average molar absorption coefficient of  $6.85 \times 10^4 \text{ L mol}^{-1} \text{ cm}^{-1}$  is shown in Figure 1. The reagent blank exhibited negligible absorbance despite having a wavelength in the same region. The reaction mechanism of the present method was as reported earlier. 28

### **Effect of solvent**

Because morin was insoluble in water, a suitable organic solvent was used for the system. Of the various solvents (acetone, isobutyl alcohol, ethanol, and 1, 4-dioxane) studied, ethanol was found to be the best solvent for the system. No absorbance was observed in the organic phase with the exception of n-butanol.

In 50 % (v/v)

ethanolic medium, however, maximum absorbance was observed; hence, a 50% ethanol solution was used in the determination procedure. It is shown in Figure 2 (other parameters were kept constant  $1 \text{ mg L}^{-1}$

of  $\text{Fe}^{2+}$  ,  
acidity 0.001 M  $\text{H}_2\text{SO}_4$  , and reagent  
concentration 1:50-fold).  
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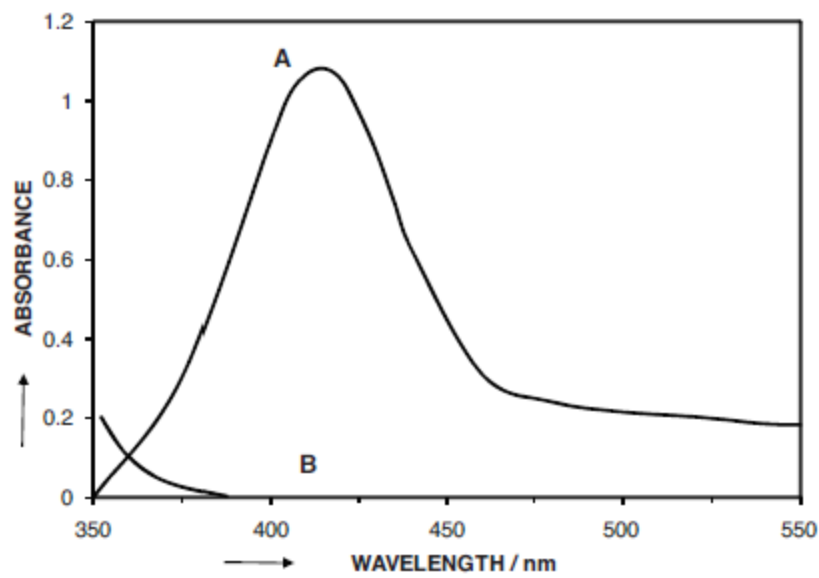
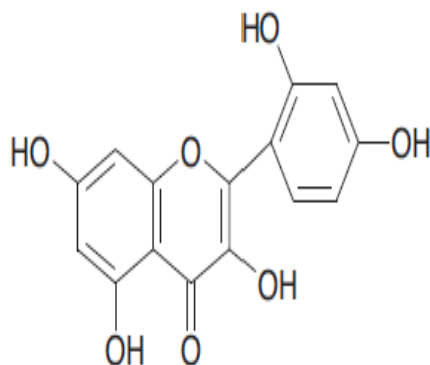


Figure 1. A and B absorption spectra of  $\text{Fe}^{II}$  – morin ( $1.5 \text{ mg L}^{-1}$ ) and the reagent blank ( $\lambda_{max} = 415 \text{ nm}$ ) in aqueous solutions.



2', 3, 4', 5, 7-Pentahydroxyflavone (morin).

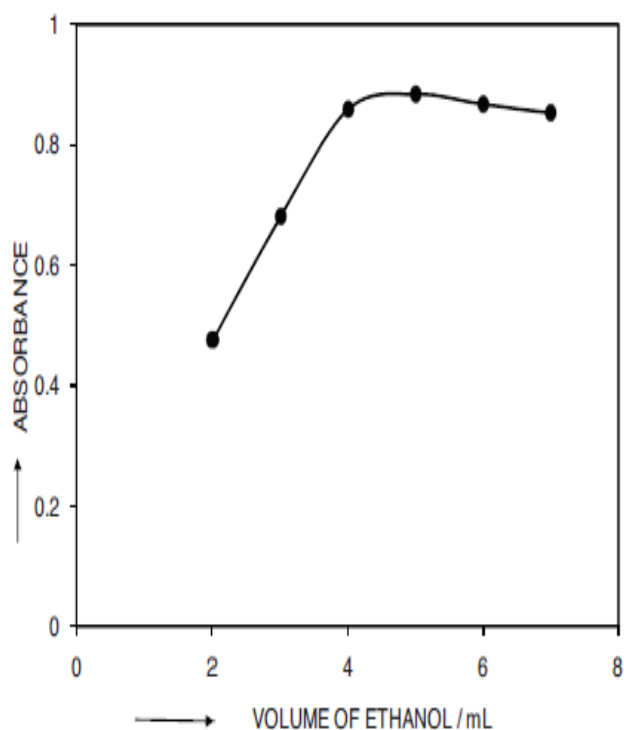


Figure 2. Effect of ethanol on the absorbance of the  $\text{Fe}^{II}$  – morin ( $1.0 \text{ mg L}^{-1}$ ) system.

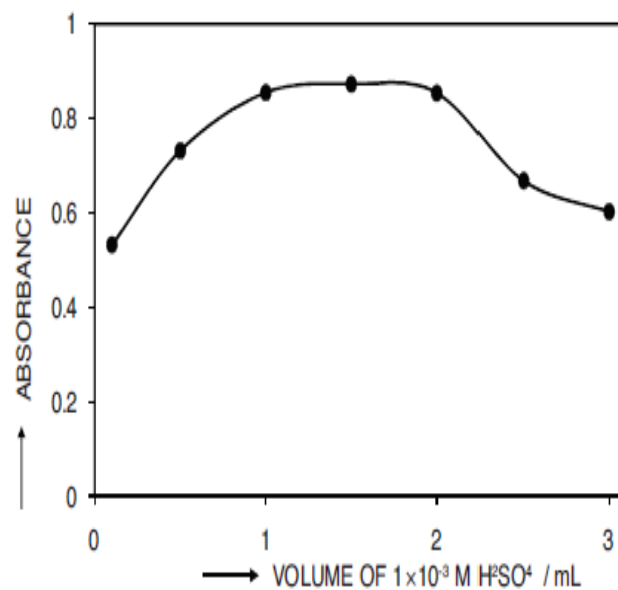


Figure 3. Effect of acidity on the absorbance of the  $\text{Fe}^{II}$  – morin ( $1.0 \text{ mg L}^{-1}$ ) system.

### **Effect of acidity**

Of the various acids (nitric, hydrochloric, sulfuric, and phosphoric) studied, sulfuric acid was found to be the best for the system. The absorbance was maximum and constant when the 10 mL of solution ( $1 \mu\text{g mL}^{-1}$ ) contained 1-2 mL of 0.001 M  $\text{H}_2\text{SO}_4$  at room temperature. Outside this range of acidity, the absorbance decreased (Figure 3) (other parameters were kept constant  $1 \text{ mg L}^{-1}$  of  $\text{Fe}^{2+}$ , 50% ethanol, and reagent concentration 1:50-fold). For all subsequent measurements, 1 mL of 0.001 M  $\text{H}_2\text{SO}_4$  was added.

### **Effect of time**

The reaction was very fast. Constant maximum absorbance was obtained within a few seconds after the dilution to volume and remained strictly unaltered for over 24 h (Figure 4) (other parameters were kept constant  $1 \text{ mg}$

$1 \text{ L}^{-1}$  of  $\text{Fe}^{2+}$ , 50% ethanol, acidity  $0.001 \text{ M H}_2\text{SO}_4$ , and reagent concentration 1:50-fold). A longer period of time was not studied.

### **Effect of temperature**

Effect of various temperatures ( $10\text{--}90^\circ\text{C}$ ) on the iron(II)-morin system was studied. The iron(II)-morin system attained maximum and constant absorbance at room temperature ( $25 \pm 5^\circ\text{C}$ ).

### **Effect of reagent concentration**

Different molar excesses of morin were added to a fixed metal ion concentration and absorbencies were measured according to the standard procedure. It was observed that at  $1 \mu\text{g mL}^{-1}$  Fe(II) metal, the analyte to reagent molar ratios of 1:40-1:95 produced a constant absorbance of the Fe-chelate (Figure 5) (other parameters were kept constant  $1 \text{ mg L}^{-1}$  of  $\text{Fe}^{2+}$ , 50% ethanol, and acidity  $0.001 \text{ M H}_2\text{SO}_4$ ). For the different Fe concentrations ( $0.1$  and  $0.5 \mu\text{g mL}^{-1}$ ) an identical effect of varying the reagent concentration was noted.

For all subsequent measurements, 1 mL of  $5.05 \times 10^{-3}$  M morin reagent was added.

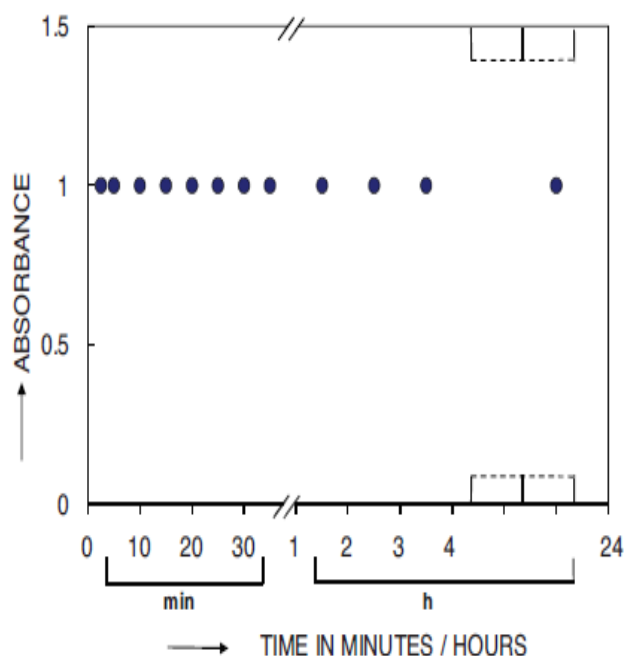


Figure 4. Effect of time on the absorbance of the  $\text{Fe}^{II}$  – morin ( $1.0 \text{ mg L}^{-1}$ ) system.

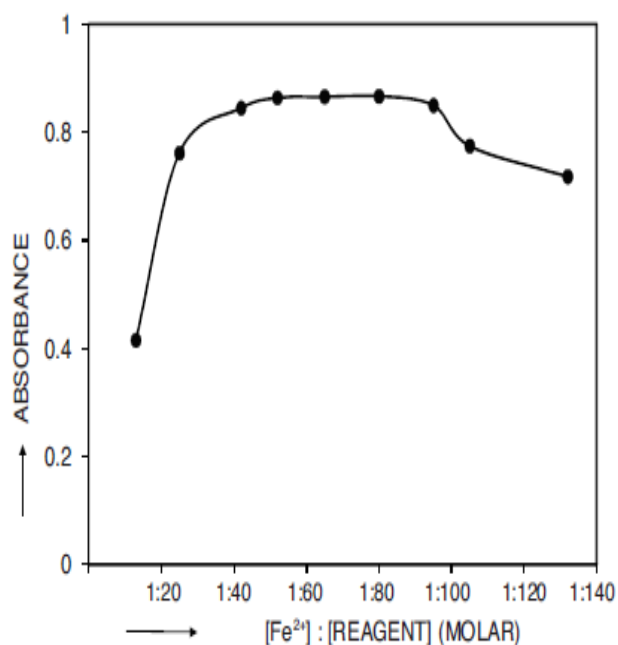


Figure 5. Effect of reagent (morin: $\text{Fe}^{II}$ ) molar concentration ratio ( $1.0 \text{ mg L}^{-1}$ ) on the absorbance of the  $\text{Fe}^{II}$  – morin ( $1.0 \text{ mg L}^{-1}$ ) system.

## Effect of metal concentration (Beer's law)

The well-known equation for spectrophotometric analysis in very dilute solutions is derived from Beer's law.

The effect of metal concentration was studied over 0.01- 0.1  $\mu$ , 0.1-1.0, and 1-50  $\mu$ g mL<sup>-1</sup> for convenience of measurement. The absorbance was linear for 0.01-10  $\mu$ g mL<sup>-1</sup> of Fe(II) at 415 nm. The molar absorptivity<sup>29</sup> was found to be  $6.85 \times 10^4$  L mol<sup>-1</sup> cm<sup>-1</sup>. Of the 3 calibration graphs, next 2 are straight-line graphs passing through the origin (Figures 6 and 7) and 1 showing the limit of the linearity is given in Figure 8.

The selected analytical parameters obtained with the optimization experiments are summarized in Table 2.



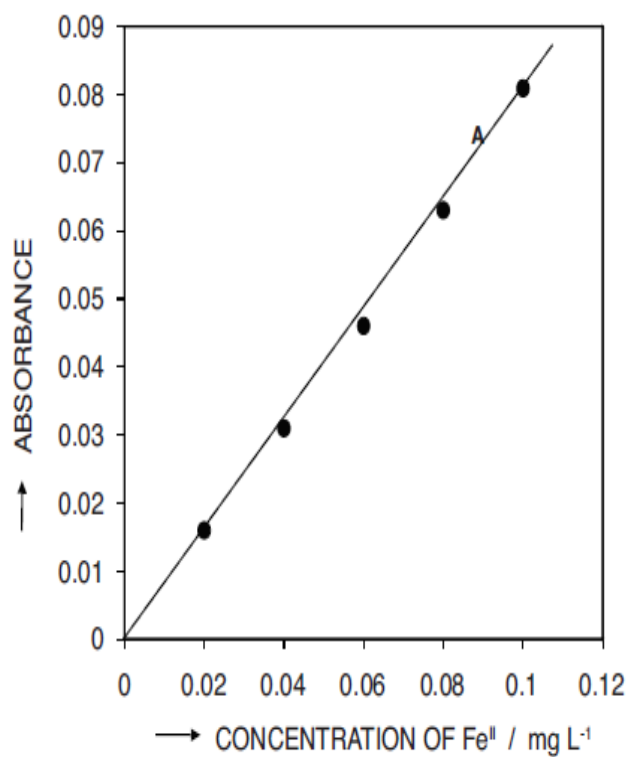


Figure 6. Calibration graph A, 0.01-0.1 mg L<sup>-1</sup> of iron(II).

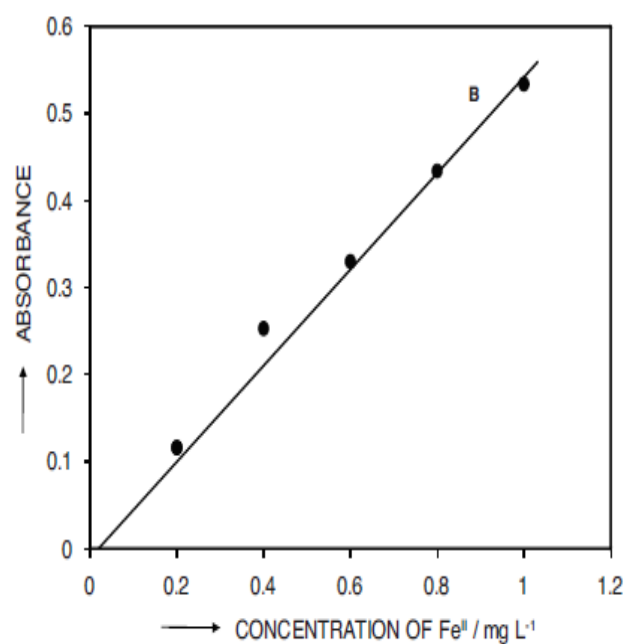


Figure 7. Calibration graph B, 0.1-1.0 mg L<sup>-1</sup> of iron(II).

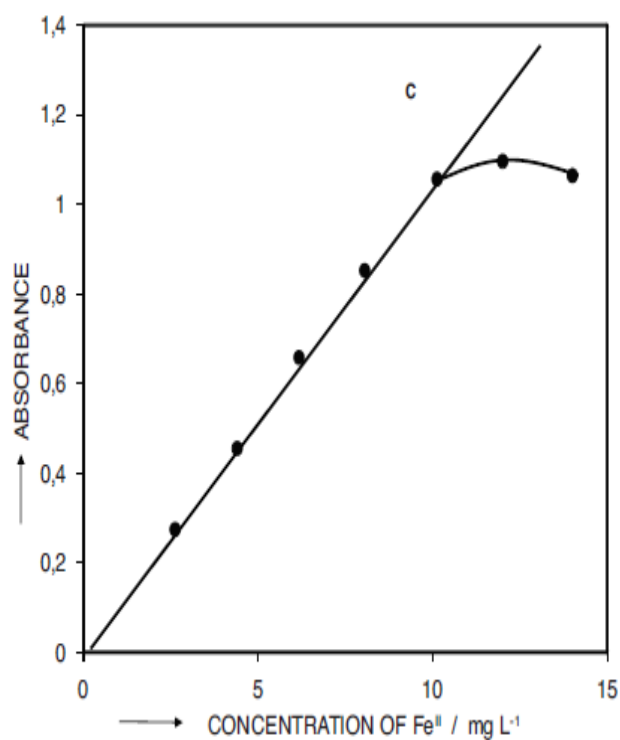


Figure 8. Calibration graph C, 1.0-10.0 mg L<sup>-1</sup> of iron(II).

**Table 2.** Selected analytical parameters obtained with the optimization experiments.

Parameters	Studied range	Selected value
Wavelength/ $\lambda_{max}$ (nm)	200-800	415
Acidity/M H <sub>2</sub> SO <sub>4</sub>	0.00001-0.003	0.0001-0.0002 (preferably, 0.0001)
pH	4.58-3.76	3.80-4.00 (preferably, 3.90)
Time/h	0-72	1 min-24 h (preferably, 5 min)
Solvent/ % Ethanol	10-100	40-60 (preferably, 50)
Temperature/°C	10-90	25 $\pm$ 5
Reagent (fold molar excess, Fe: morin)	1:5-1:132	1:40-1:95 (preferably, 1:50)
Molar absorptivity L mol <sup>-1</sup> cm <sup>-1</sup>	-	6.85 $\times$ 10 <sup>4</sup>
Linear range/mg L <sup>-1</sup>	0.001-100	0.01-1.0
Detection limit/ $\mu$ g L <sup>-1</sup>	-	0.5
Regression coefficient	-	0.9992

## Effect of foreign ions

The effect of over 50 anions, cations, and complexing agents on the determination of only 1  $\mu$ g mL<sup>-1</sup> of Fe(II) was studied. The criterion for interference was an absorbance value varying by more than

5% from the expected value for Fe(II) alone. The results are summarized in Table 3. As can be seen, a large number of ions have no significant effect on the determination of iron. The most serious interference were from V(V) and Al(III) ions. Interference from these ions is due to complex formation with morin. The greater tolerance limits (absorbance value varying by more than 5% from the accepted value for Fe(II) alone) for these ions can be achieved by using several masking methods. In order to eliminate interference of V(V) and Al(III), H<sub>2</sub>O<sub>2</sub> and EDTA were used as masking agents, respectively. During the interference studies, if a precipitate was formed, it was removed by centrifugation. The amount mentioned is not the tolerance limit but the actual amount studied. However, for those ions whose tolerance limit has been studied, their tolerance ratios are given in Table 3.

## **Composition of the absorbing complex**

Job's method<sup>31</sup> of continuous variation and the molar ratio<sup>32</sup> method were applied to ascertain the stoichiometric

composition of the complex. A Fe-morin (1:2) complex was indicated by both methods.

## **Precision and accuracy**

The precision of the present method was evaluated by determining different concentrations of iron (each analyzed  $n = 5$  times). The relative standard deviation ( $n = 5$ ) was 2%-0% for 0.1-100  $\mu\text{g}$  of iron(II) in 10 mL, indicating

that this method is highly precise and reproducible. The detection limit (3 s of the blank) was found to be 0.5

$\mu\text{g L}^{-1}$ . The method was tested by analyzing several synthetic mixtures containing iron(II) and diverse ions

(Table 4). The results for total iron were in good agreement with certified values (Table 5). The reliability of

our Fe-chelate procedure was tested by recovery studies. The average percentage

recovery obtained for addition

of iron(II) spike to some environmental water samples was quantitative as shown in Table 6.

The results of

biological analyses by the spectrophotometric method were in excellent agreement with those obtained by AAS

(Table 7). The results of soil samples (analyzed) by the spectrophotometric method were highly reproducible

(Table 8). The results of food samples analyzed by the spectrophotometric method were in good agreement

with the expected values (Table 9). The results of speciation of iron(II) and iron(III) in mixtures were highly

reproducible (Table 10). Hence, the precision and accuracy of the method were excellent.

### **Determination of iron in synthetic mixtures**

Several synthetic mixtures of varying compositions containing iron and diverse ions

of known concentrations were determined by the present method using tartrate and EDTA (0.01% w/v) as masking agent and the results were found to be highly reproducible. The results are shown in Table 4. Accurate recoveries were achieved in all solutions.

### **Determination of iron in brass, alloys and steels (Certified reference materials)**

A 0.1 g amount of a brass or alloy or steel sample containing 1.56%-34.26% of iron was weighed accurately and placed in a 50 mL Erlenmeyer flask following a method recommended by Parker et al.<sup>33</sup> To it, 10 mL of concentrated  $\text{HNO}_3$  and 1 mL of concentrated  $\text{H}_2\text{SO}_4$  were added, carefully covering the flask with a watch glass until the brisk reaction subsided. The solution was heated and simmered gently after the addition of 5 mL of concentrated  $\text{HNO}_3$  until all carbides were decomposed. The solution was carefully evaporated to dense

white fumes to drive off the oxides of nitrogen and then cooled to room temperature ( $25 \pm 5$  °C). After suitable

dilution with deionized water, the contents of the Erlenmeyer flask were warmed to dissolve the soluble salts.

Then the content of the flask was reduced to iron(II) by using freshly prepared sodium azide solution and excess

of the azide solution was removed by boiling.

The solution was then cooled and neutralized with a dilute  $\text{NH}_3$

solution. The resulting solution was filtered, if necessary, through Whatman No. 40 filter paper into a 25-mL

calibrated flask. The residue was washed with a small volume of hot (1 + 99)  $\text{H}_2\text{SO}_4$ , followed by water, and

the volume was made up to the mark with deionized water.

A suitable aliquot (1-2 mL) of the above solution was taken into a 10-mL calibrated flask and the

iron content was determined as described

above, using tartrate (0.01% w/v) and EDTA (0.01 w/v) mixture as masking agent. Based on 5 replicate analyses, the average iron concentration determined by spectrophotometric method was in close agreement with the certified values. The results are shown in Table 5.

### **Determination of iron in environmental waters**

Each filtered (with Whatman No. 40) environmental water sample (1000 mL) evaporated nearly to dryness with a mixture of 3 mL of concentrated  $H_2SO_4$  and 10 mL of concentrated  $HNO_3$  in a fume cupboard, following a method recommended by Greenberg et al. 34 and was heated with 10 mL of deionized water in order to dissolve the salts. The content of the flask was reduced to iron(II) using sodium azide. The solution was then cooled and neutralized with dilute  $NH_4OH$  solution in the presence of 1-2 mL of 0.01% (w/v) tartrate or EDTA solution.

The resulting solution was then filtered (if



necessary) and quantitatively transferred into a 25-mL calibrated flask and made up to the mark with deionized water.

An aliquot (1-2 mL) of this preconcentrated water sample was pipetted into a 10-mL calibrated flask and the iron content was determined as described above, using a mixture of tartrate and EDTA as masking agent.

The analyses of environmental water samples for iron from various sources are shown in Table 6.

Most spectrophotometric methods for the determination of iron in natural and sea-water require preconcentration of iron. 34 Preconcentration is a very cheap and easy method. The concentration of iron in natural

and sea-water is a few  $\mu\text{g L}^{-1}$  in Japan. 35 The mean concentration of iron found in UK drinking water is less

than  $1 \text{ mg L}^{-1}$  (av.  $200 \mu\text{g L}^{-1}$ ). 36

### **Determination of iron in biological samples**

Human blood (2-5 mL) and urine (20-30 mL)

were collected in polyethylene bottles from the affected persons (diseases occur due Fe poisoning such as anemia, liver cirrhosis, kidney diseases). Immediately after collection, they were stored in a salt-ice mixture and later, at the laboratory, were kept at  $-20^{\circ}\text{C}$ . The samples were taken into a 100 mL micro-Kjeldahl flask. A glass bead and 10 mL of concentrated nitric acid were added and the flask was placed on the digester under gentle heating. When the initial brisk reaction was over, the solution was removed and cooled following a method recommended by Stahr.<sup>37</sup> Then 1 mL of concentrated sulfuric acid was added carefully, followed by the addition of 2 mL of concentrated HF, and heating was continued for at least 30 min, followed by cooling. The content of the flask was reduced to iron(II) using sodium azide solution and excess azide was removed by boiling and then the content was filtered. The solution of

flask was then neutralized with dilute  $\text{NH}_4\text{OH}$  solution in the presence of 1-2 mL of a 0.01% (w/v) tartrate or EDTA solution.

The resultant solution was then transferred quantitatively into a 10-mL calibrated flask and made up to the mark with deionized water.

A suitable aliquot (1-2-mL) of the final solution was pipetted into a 10-mL calibrated flask and the iron

content was determined as described above using tartrate or EDTA as masking agent. The results of biological analyses by the spectrophotometric method were found to be in excellent agreement with those obtained by

AAS. The results are shown in Table 7.

The abnormally high value for the liver cirrhosis patient is probably due to the involvement of a high iron

concentration with Cu and Zn. Occurrence of such high iron contents is also reported in liver

cirrhosis patients  
from some developed countries. 2

### **Determination of iron in soil samples**

An air dried homogenized soil sample (100 g) was weighed accurately and placed in a 100-mL micro-Kjeldahl

flask. The sample was digested in the presence of a reducing agent (2.5% freshly prepared azide solution),

following the method recommended by Hesse.

38 The content of the flask was filtered through a Whatman No.

40 filter paper into a 25-mL calibrated flask and neutralized with dilute  $\text{NH}_4\text{OH}$  solution in the presence of 1-2

mL of a 0.01% (w/v) tartrate or EDTA solution.

Then the solution of the flask was made up to the mark with

deionized water.

Suitable aliquots (1-2 mL) were transferred into a 10-mL calibrated flask and a calculated amount of

0.001 M  $\text{H}_2\text{SO}_4$  needed to give a final acidity of 0.0001-0.0002 M  $\text{H}_2\text{SO}_4$  was added followed by 1-2 mL of a

0.01% (w/v) mixture of tartrate and EDTA solution as masking agent. The iron content was then determined by the above procedure and quantified from a calibration graph prepared concurrently. The results are shown in Table 8.

### **Determination of iron in food samples**

An air-dried food sample (arum (25 g), apple (50 g), guava (50 g), and egg (1 piece)) was taken in a 100-mL

micro-Kjeldahl flask. A glass bead and 10 mL of concentrated nitric acid were added and the flask was placed

on the digester under gentle heating. When the initial brisk reaction was over, the solution was removed and

cooled following the method recommended by Stahr.<sup>37</sup> Then 1 mL volume of concentrated sulfuric acid was

added carefully, followed by the addition of 2 mL of concentrated HF, and heating was continued for at least

30 min and then cooling. The content of the flask was reduced to iron(II) using sodium azide

solution and excess azide was removed by boiling and then the content was filtered. The solution of the flask was then neutralized with dilute  $\text{NH}_4\text{OH}$  in the presence of 1-2 mL of a 0.01% (w/v) tartrate or EDTA solution. The resultant solution was then transferred quantitatively into a 50-mL calibrated flask and made up to the mark with deionized water.

A suitable aliquot (1-2 mL) of the final solution was pipetted into a 10-mL calibrated flask and the iron content was determined as described above using tartrate or EDTA as masking agent. The high value of iron for *psidium guajava* (guava) is probably due to the involvement of a high iron concentration in the soil. The results are shown in Table 9.

### **Determination of iron(II) and iron(III) in mixtures**

Suitable aliquots (1-2 mL) of iron (II + III) mixtures (preferably 1:1, 1:2, 1:3) were taken in a 25-mL conical

flask. A few drops of freshly prepared sodium azide solution was added to reduce the trivalent iron to divalent.

A 5-mL volume of water was added to the mixtures, which were then heated on a steam bath for 10-15 min with occasional gentle shaking to remove excess azide and then cooling to room temperature. The reaction mixture was transferred quantitatively into a 10-mL calibrated flask. Then the total iron (II+III) content was determined according to the general procedure with the help of the calibration graph.

An equal aliquot of the above iron (II + III) mixture was taken into a 25-mL beaker. Then 2 mL of 0.01% (w/v) EDTA was added to mask iron(III). Then the content of the beaker was transferred into a 10-mL calibrated flask, and its iron(II) content was determined according to the general procedure. The iron concentration was calculated in  $\mu\text{g L}^{-1}$  or mg

L-1 with the aid of a calibration graph. This gives a measure of iron originally present as iron(II) in the mixture. The value of the iron(III) concentration was calculated by subtracting the concentration of iron(II) from the corresponding total iron concentration. The results were found to be highly reproducible. The results of a set of determination are given in Table 10.

### **Conclusions**

A new, simple, sensitive, selective, and inexpensive method with the Fe(II)-morin complex was developed for the determination of iron in some industrial, biological, soil, and environmental samples, for continuous monitoring to establish the trace levels of iron in different sample matrices. It offers also a very efficient procedure for speciation analysis. Although many sophisticated techniques such as pulse polarography, HPLC, AAS, ICP-AES, and ICP-MS are available for the determination of iron at trace levels in numerous complex



materials,  
factors such as the low cost of the instrument,  
easy handling, lack of requirement for  
consumables, and almost  
no maintenance have caused  
spectrophotometry to remain a popular  
technique, particularly in laboratories  
in developing countries with limited budgets.  
The sensitivity in terms of molar absorptivity  
and precision in  
terms of relative standard deviation of the  
present method are very reliable for the  
determination of iron in real  
samples down to  $\text{ng g}^{-1}$  levels in aqueous  
medium at room temperature ( $25 \pm 5^\circ \text{C}$ ).