International Journal of Scientific Research in Chemical Sciences
Vol.**12**, Issue.**1**, pp.**06-10**, February **2025**
E-ISSN: 2455-3174
Available online at: www.isroset.org

Research Article

Determination of the Equilibrium Constant Kc by Spectrophotometry: A Comprehensive Study of the Reaction Between Iron(III) and Thyocyanate Ions

Ilias Darsaklis[1*]

[1]Dept. of Chemistry, University of Ioannina, Ioannina, Greece

*Corresponding Author:

Received: 01/Jan/2025; **Accepted**: 01/Feb/2025; **Published**: 28/Feb/2025 | https://doi.org/10.26438/ijsrcs.v12i1.182

Abstract—Iron (III) thiocyanate complex ion **[**FeSCN]$^{2+}$ is a complex with a deep red color and is formed by the reaction between trivalent iron ions (Fe^{3+}) and thiocyanate ions (SCN^{-}). In the present study, solutions of Fe^{3+} and HSCN with known concentrations were used and the equilibrium constant K_C of the reaction was calculated so that it can be estimated where the chemical equilibrium is shifted. By measuring the absorption of the compound at the maximum wavelength (λ_{max}), the concentrations of [Fe^{3+}], [HSCN], and [FeSCN]$^{2+}$ in the chemical equilibrium were determined. The results found for K_C, indicated that the chemical equilibrium is shifted to the right, meaning the amount of products exceeded the amount of the reactants in all cases.

Keywords— equilibrium constant, Le Chatelier's principle, iron (III) ions, thiocyanate ions, spectrophotometry

1. Introduction

Each chemical reaction reaches a specific equilibrium point at a certain temperature. When two reagents are combined, they begin reacting until they reach a condition where the concentrations of the reactants and products become stable and no longer convert. Once this condition is achieved under certain conditions, a chemical equilibrium will be established, and the system will remain in this state until it is disturbed in some way. The constant of the chemical equilibrium Kc is also connected with the state of chemical equilibrium [1].

Some chemical reactions are largely complementary; meaning that at least one of the reagents is consumed. Such reactions are stoichiometric since the only information we need to calculate the amounts of products formed is the amount of available reagents.

However, some reactions do not lead to completion. Let's look at the following reaction, shown in Equation (1), for example.

$$\mathrm{aA + bB \rightleftarrows cC + dD} \tag{1}$$

Where A, B, C, and D are dissolved solution substances or gases in the state of chemical equilibrium.

Whether the reactants or the products of the reaction are mixed and driven to consistent temperature and pressure conditions, the exact equilibrium will be re-established. On a molecular scale, we can say that both reactions occur at the same speed in both directions, when we are in equilibrium, that is, the equilibrium is dynamic [2].

The expressions for the equilibrium constants in the above reaction are shown below, in Equations (2) and (3) [3]:

- $$\mathrm{Kp = \frac{(pC)^c(pD)^d}{(pA)^a(pB)^b}} \tag{2}$$
  (Pressure is constant)

- $$\mathrm{Kc = \frac{[C]^c[D]^d}{[A]^a[B]^b}} \tag{3}$$
  (Concentration is constant)

The direction of the reaction depends on:
- The ratio of reactors.
- The value of the constant K_C.
- The quantities of reactants and products [4].

This study presents a comprehensive study on the determination of the equilibrium constant Kc of the chemical reaction between Fe^{3+} and SCN^{-} and its importance in understanding the behavior of the reaction, using UV-Vis spectrophotometric analysis, and evaluate the dependence of absorbance on the concentrations of reactants and products at chemical equilibrium.

In the present paper the first section involves the introduction of the equilibrium constant K_C, in the second section related work is presented, the third section explains the experimental procedure for the calculation of the equilibrium constant K_C, in the fourth section the results and data of the experimental procedure are discussed, and the fifth section presents the conclusion and future scope on the research work of this study.

2. Related Work

It is important to note that the law of equilibrium expresses the constant K_C as a relationship between the concentration of the products and the reactants, in a reaction at equilibrium, in order to determine the position of equilibrium. The value of the equilibrium constant Kc indicates the relative proportion of products and reactants and provides information about the extent of the reaction (but not about the rate of the reaction) [5].

- When $K_C > 1$, the concentration of reactants is lower than that of the products. The formation of the products is favored; meaning the position of the equilibrium is more to the right.
- When $K_C < 1$, the concentration of reactants is higher than that of the products. The equilibrium position is more to the left.
- When the constant K_C is equivalent to 1, it suggests that the concentrations of the reactants and the products are almost similar [5].

In accordance with Le Chatelier's principle, because K_P and K_C are constant, if the partial pressure or the products of the reaction change during equilibrium, a corresponding change will occur in some pressures or the concentrations of other components so that K_P and K_C remain stable [2].

The equilibrium position is affected by the following factors:

- Concentration of substances involved in the chemical equilibrium (reactants & products).
- Pressure, applied only when there are gaseous substances in the chemical equilibrium and during the reaction there is a convert in the total number of moles of the gases
- Temperature, as the chemical equilibrium constant K_C is temperature dependant.

In general, for endothermic reactions, an increase in temperature implies an increase in K_C while for exothermic reactions an increase in temperature implies a decrease in K_C. It is also obvious that the higher the value of K_C, the higher the percentage of reactants is converted into products. If substances A, B, C, and D are mixed in varying quantities, they will react until their concentrations persuade the equation of the chemical equilibrium [6].

The addition of a catalyst does not affect the position of the chemical equilibrium. The presence of a catalyst simply accelerates the restoration of chemical equilibrium. Also, the addition of inert gas, such as helium (He), while maintaining the volume and temperature of the container constant, increases the total number of gas moles in the system, thus increasing the total pressure inside the container. However, the concentrations and some pressures of the gases present in the system do not change since the inert gas does not react with any of these gases. Therefore, the addition of inert gas does not affect the position of the chemical equilibrium [6].

3. Experimental Method

The samples are prepared by mixing solutions containing known concentrations of trivalent iron (Fe^{3+}) and thiocyanic acid (HSCN). Knowing the initial composition of the solution and the concentration of ions $[FeSCN]^{2+}$ in the state of chemical equilibrium we can calculate the concentrations of the other particles in the chemical equilibrium and also calculate the constant K_C.

A solution of 5 ml $Fe(NO_3)_3$ $2 \cdot 10^{-3}$ M and 0.5M HNO_3 was added in 5 test tubes. The corresponding amount of 1 to 5 ml HSCN $2 \cdot 10^{-3}$ M and 0.5M HNO_3 was added to each test tube and the solutions were then mixed. That is, the first test tube will contain 6 ml, the second 7 ml, the third 8 ml, the fourth 9 ml, and the fifth 10 ml.

The experiment is followed by analysis with a spectrophotometer. The spectrophotometer was set to 447nm. A small amount of 0.5 HNO_3 was added to the cells of the instrument as a blank. After that, the solution from the first tube was added to the cell of the spectrophotometer, and the absorbance was recorded. The same procedure was repeated for the rest of the tubes.

Once the absorbance was measured for all the samples, a calibration curve was generated by plotting the absorbance of known concentrations of $FeSCN^{2+}$ against the concentration of the complex.

4. Results and Discussion

The reaction between the trivalent iron ions (Fe^{3+}) and the ions of thiocyanic acid [SCN^-] is shown below in Equation (4):

$$Fe^{3+}_{(aq)} + HSCN_{(aq)} \leftrightarrows [FeSCN]^{2+}_{(aq)} + H^{+}_{(aq)} \qquad (4)$$

When solutions containing trivalent iron ions (Fe^{3+}) and thiocyanate acid (HSCN) are mixed, they react as shown in Eq. (4), and $[FeSCN]^{2+}$ is formed, which exhibits a deep red colour. As a result of this reaction, the concentration of Fe^{3+} and HSCN in the state of chemical equilibrium will be lower than they would be in the absence of this reaction. This means that for each mole of $[FeSCN]^{2+}$ that is formed, one mole of Fe^{3+} and one mole of HSCN has been consumed.

According to the definition of the chemical equilibrium constant, we have:

$$\mathrm{Kc} = \frac{[\mathrm{FeSCN}]^{2+}[\mathrm{H}]^{+}}{[\mathrm{Fe}]^{3+}[\mathrm{HSCN}]} \tag{5}$$

The chemical equilibrium constant K_C for the reaction in equation (5) is conveniently sized, while the color of the $[FeSCN]^{2+}$ ions, in turn, helps to analyze the mixture at the chemical equilibrium state.

To determine the concentration of $[FeSCN]^{2+}$ ions at chemical equilibrium, a UV-Vis-1280-Shimadzu spectrophotometer was used, and the maximum absorbance of the red complex at 447 nm was measured.

Spectrophotometry is a great experimental technique that provides useful details regarding the equilibrium concentrations of a reaction mixture [7]. Spectrophotometry is distinct from potentiometry and offers advantages by allowing for the direct measurement of a characteristic of one or more reactants or products in a reaction mixture. In addition, spectrophotometry is applicable to reactions or solvent systems that may not be ideal for potentiometry [8].

Absorption A is proportional to the concentration and can be calculated from the relation A = ε· b ·C (Beer-Lambert Law). The absorption is measured in maximum wavelength (λ_{max}), for two primary reasons:

1. To achieve maximum sensitivity, since a given concentration will give the maximum absorption to this wavelength, and
2. To minimize the change in absorption due to slight variations in wavelength during measurements [9],[10].

In this study, the path length of the cell (b) was 1.00cm, and the molar absorptivity (ε) at 447nm was determined to be $2840.69 mol^{-1}\cdot L\cdot cm^{-1}$. The concentration of $[H^+]$ was maintained constant at a level of 0.5M.

The maximum absorbance values for the iron(III) thiocyanate complex $[FeSCN]^{2+}$ for each sample, at a wavelength of 447nm, are presented below in Table 1.

Table 1. Absorbance of $[FeSCN]^{2+}$ for each sample at a wavelength of 447nm

Sample	$Fe(NO_3)_3$	HSCN	Absorbance $[FeSCN]^{2+}$
1.	5 ml	1 ml	0.264
2.	5 ml	2 ml	0.342
3.	5 ml	3 ml	0.407
4.	5 ml	4 ml	0.431
5.	5 ml	5 ml	0.440

The equilibrium constant K_C of the reaction was calculated as follows:

- The initial moles of Fe^{3+} and HSCN were calculated.
- The moles of the complex $[FeSCN]^{2+}$ in the chemical equilibrium condition were determined.
- The moles of the other particles in the chemical equilibrium condition were determined.
- The constant Kc was calculated from the relation of the constant chemical equilibrium.

Sample No1:

Concentration $[FeSCN]^{2+}$ in equilibrium: C = A /εb = 0.264 / 2840,69 = $9.3\cdot10^{-5}$ M
Initial moles Fe^{3+}: $n_{Fe3+} = C_{Fe3+} \cdot V_{Fe(NO3)3} = 2\cdot10^{-3} \cdot 5\cdot10^{-3} = 10^{-5}$ moles
Initial moles HSCN: $n_{HSCN} = C_{HSCN} \cdot V_{HSCN} = 2\cdot10^{-3} \cdot 10^{-3} = 2\cdot10^{-6}$ moles
Moles $[FeSCN]^{2+}$ in equilibrium: $n_{[FeSCN]2+} = M_{[FeSCN]2+} \cdot V = 9.3\cdot10^{-5} \cdot 6\cdot10^{-3} = 5.6 \cdot10^{-7}$ moles
Moles Fe^{3+} in equilibrium = $n_{Fe3+} - n_{[FeSCN]2+} = 10\cdot10^{-6} - 0.56\cdot10^{-6} = 9.44\cdot10^{-6}$ moles
Moles HSCN in equilibrium = $n_{HSCN} - n_{[FeSCN]2+} = 2\cdot10^{-6} - 0.56\cdot10^{-6} = 1.44\cdot10^{-6}$ moles
Fe^{3+} concentration in equilibrium: C = n_{Fe3+} / V = $9.44\cdot10^{-6} / 6\cdot10^{-3} = 1.6\cdot10^{-3}$ M
HSCN concentration in equilibrium: C = n_{HSCN} / V = $1.44\cdot10^{-6} / 6\cdot10^{-3} = 0.24\cdot10^{-3}$ M
$K_C = [FeSCN^{2+}] [H^+] / [Fe^{3+}] [HSCN] = 9.3\cdot10^{-5} \cdot 0.5 / 1.6\cdot10^{-3} \cdot 0.24\cdot10^{-3} = 121.09$

Sample No2:

Concentration $[FeSCN]^{2+}$ in equilibrium: C = A /εb = 0.342 / 2840.69 = $1.2x10^{-4}$ M
Initial moles Fe^{3+}: $n_{Fe3+} = M_{Fe3+} \cdot V_{Fe(NO3)3} = 2\cdot10^{-3} \cdot 5\cdot10^{-3} = 10^{-5}$ moles
Initial moles HSCN: $n_{HSCN} = M_{HSCN} \cdot V_{HSCN} = 2\cdot10^{-3} \cdot 2\cdot10^{-3} = 4\cdot10^{-6}$ moles
Moles $[FeSCN]^{2+}$ in equilibrium: $n_{[FeSCN]2+} = M_{[FeSCN]2+} \cdot V = 1.2\cdot10^{-4} \cdot 7\cdot10^{-3} = 8.4\cdot10^{-7}$ moles
Moles Fe^{3+} in equilibrium = $n_{Fe3+} - n_{[FeSCN]2+} = 10\cdot10^{-6} - 0.84\cdot10^{-6} = 9.16\cdot10^{-6}$ moles
Moles HSCN in equilibrium = $n_{HSCN} - n_{[FeSCN]2+} = 4\cdot10^{-6} - 0.84\cdot10^{-6} = 3.16\cdot10^{-6}$ moles
Fe^{3+} concentration in equilibrium: C = n_{Fe3+} / V = $9.16\cdot10^{-6} / 7\cdot10^{-3} = 1.3\cdot10^{-3}$ M
HSCN concentration in equilibrium: C = n_{HSCN} / V = $3.16\cdot10^{-6} / 7\cdot10^{-3} = 0.45\cdot10^{-3}$ M
$Kc = [FeSCN^{2+}] [H^+] / [Fe^{3+}] [HSCN] = 1.2\cdot10^{-4} \cdot 0.5 / 1.3\cdot10^{-3} \cdot 0.45\cdot10^{-3} = 102.56$

Sample No3:

Concentration $[FeSCN]^{2+}$ in equilibrium: C = A /εb = 0.407 / 2840.69 = $1.4\cdot10^{-4}$ M
Initial moles Fe^{3+}: $n_{Fe3+} = M_{Fe3+} \cdot V_{Fe(NO3)3} = 2x10^{-3} \cdot 5\cdot10^{-3} = 10^{-5}$ moles
Initial moles HSCN: $n_{HSCN} = M_{HSCN} \cdot V_{HSCN} = 2\cdot10^{-3} \cdot 3\cdot10^{-3} = 6\cdot10^{-6}$ moles
Moles $[FeSCN]^{2+}$ in equilibrium: $n_{[FeSCN]2+} = M_{[FeSCN]2+} \cdot V = 1.4\cdot10^{-4} \cdot 8\cdot10^{-3} = 1.12\cdot10^{-6}$ moles
Moles Fe^{3+} in equilibrium = $n_{Fe3+} - n_{[FeSCN]2+} = 10\cdot10^{-6} - 1,12\cdot10^{-6} = 8,88\cdot10^{-6}$ moles
Moles HSCN in equilibrium = $n_{HSCN} - n_{[FeSCN]2+} = 6\cdot10^{-6} - 1,12\cdot10^{-6} = 4,88\cdot10^{-6}$ moles
Fe^{3+} concentration in equilibrium: C = n_{Fe3+} / V = $8.88\cdot10^{-6} / 8\cdot10^{-3} = 1.1\cdot10^{-3}$ M
HSCN concentration in equilibrium: C = n_{HSCN} / V = $4.88\cdot10^{-6} / 8\cdot10^{-3} = 0.61\cdot10^{-3}$ M
$Kc = [FeSCN^{2+}] [H^+] / [Fe^{3+}] [HSCN] = 1.4\cdot10^{-4} \cdot 0.5 / 1.1\cdot10^{-3} \cdot 0.61\cdot10^{-3} = 104.32$

Sample No4:

Concentration $[FeSCN]^{2+}$ in equilibrium: $C = A/\varepsilon b = 0.431/2840.69 = 1.5\cdot10^{-4}$ M
Initial moles Fe^{3+}: $n_{Fe3+} = M_{Fe3+} \cdot V_{Fe(NO3)3} = 2\cdot10^{-3} \cdot 5\cdot10^{-3} = 10^{-5}$ moles
Initial moles HSCN: $n_{HSCN} = M_{HSCN} \cdot V_{HSCN} = 2\cdot10^{-3} \cdot 4\cdot10^{-3} = 8\cdot10^{-6}$ moles
Moles $FeSCN^{2+}$ in equilibrium: $n_{[FeSCN]2+} = M_{[FeSCN]2+} \cdot V = 1.5\cdot10^{-4} \cdot 9\cdot10^{-3} = 1.35\cdot10^{-6}$ moles
Moles Fe3 + in balance = $n_{Fe3+} - n_{[FeSCN]2+} = 10\cdot10^{-6} - 1.35\cdot10^{-6} = 8.65\cdot10^{-6}$ moles
Moles HSCN in equilibrium = $n_{HSCN} - n_{[FeSCN]2+} = 8\cdot10^{-6} - 1.35\cdot10^{-6} = 6.65\cdot10^{-6}$ moles
Fe^{3+} concentration in equilibrium: $C = n_{Fe3+} / V = 8.65\cdot10^{-6} / 9\cdot10^{-3} = 0.96\cdot10^{-3}$ M
HSCN concentration in equilibrium: $C = n_{HSCN} / V = 6.65\cdot10^{-6} / 9\cdot10^{-3} = 0.74\cdot10^{-3}$ M
$Kc = [FeSCN^{2+}]\ [H^+]\ /\ [Fe^{3+}]\ [HSCN] = 1.5\cdot10^{-4} \cdot 0.5 / 0.96\cdot10^{-3} \cdot 0.74\cdot10^{-3} = 105.57$

Sample No5:
Concentration $[FeSCN]^{2+}$ in equilibrium: $C = A/\varepsilon b = 0{,}440/2840{,}69 = 1{,}54\cdot10^{-4}$ M
Initial moles Fe^{3+} : $n_{Fe3+} = M_{Fe3+} \cdot V_{Fe(NO3)3} = 2\cdot10^{-3} \cdot 5x10^{-3} = 10^{-5}$ moles Fe^{3+}
Initial moles HSCN: $n_{HSCN} = M_{HSCN} \cdot V_{HSCN} = 2\cdot10^{-3} \cdot 5\cdot10^{-3} = 10^{-5}$ moles
Moles $[FeSCN]^{2+}$ in equilibrium: $n_{[FeSCN]2+} = M_{[FeSCN]2+} \cdot V = 1{,}54\cdot10^{-4} \cdot 10^{-2} = 1{,}54\cdot10^{-6}$ moles
Moles Fe^{3+} in equilibrium = $n_{Fe3+} - n_{[FeSCN]2+} = 10\cdot10^{-6} – 1{,}54\cdot10^{-6} = 8{,}46\cdot10^{-6}$ moles
Moles HSCN in equilibrium = $n_{HSCN} – n_{[FeSCN]2+} = 10\cdot10^{-6} - 1{,}54\cdot10^{-6} = 8{,}46\cdot10^{-6}$ moles
Fe^{3+} concentration in equilibrium: $C = n_{Fe3+} / V = 8{,}46\cdot10^{-6} / 10^{-2} = 8{,}46\cdot10^{-4}$ M
HSCN concentration in equilibrium: $C = n_{HSCN} / V = 8{,}46\cdot10^{-6} / 10^{-2} = 8{,}46\cdot10^{-4}$ M
$Kc = ([FeSCN]^{2+}]\ [H^+]\ /\ ([Fe^{3+}]\ [HSCN]) = 1{,}54\cdot10^{-4} \cdot 0{,}5 / 8{,}46\cdot10^{-4} \cdot 8{,}46\cdot10^{-4} = 107{,}58$

The initial moles of Fe^{3+} and HSCN in the mixture, as well as the moles in each sample after chemical equilibrium has been achieved, are displayed in Table 2 below. The value of the equilibrium constant K_C for every sample individually is also shown in the same Table 2.

Table 2. Values of moles before and after chemical equilibrium and Kc for each sample

	Initial moles		moles in chemical equilibrium			K_C
	Fe^{3+}	HSCN	Fe^{3+}	HSCN	$[FeSCN]^{2+}$	
1.	10^{-5}	$2x10^{-6}$	$9{,}44\ x10^{-6}$	$1{,}44\ x10^{-6}$	$5{,}6x10^{-7}$	121,09
2.	10^{-5}	$4x10^{-6}$	$9{,}16x10^{-6}$	$3{,}16\ x10^{-6}$	$8{,}4x10^{-7}$	102,56
3.	10^{-5}	$6x10^{-6}$	$8{,}88x10^{-6}$	$4{,}88\ x10^{-6}$	$1{,}12x10^{-6}$	104,32
4.	10^{-5}	$8x10^{-6}$	$8{,}65x10^{-6}$	$6{,}65\ x10^{-6}$	$1{,}35x10^{-6}$	105,57
5.	10^{-5}	10^{-5}	$8{,}46x10^{-6}$	$8{,}46\ x10^{-6}$	$1{,}54x10^{-6}$	107,58

The calibration curve was constructed by plotting the absorbance of each sample of $FeSCN^{2+}$ against the concentration of the sample. Zero absorbance for zero concentration is also a valid point. This curve is used to determine the molar absorptivity (ε) and to calculate the equilibrium concentration of $FeSCN^{2+}$ in unknown solutions. The graph of absorbance against the known concentration of $[FeSCN]^{2+}$ is presented below in Figure 1.

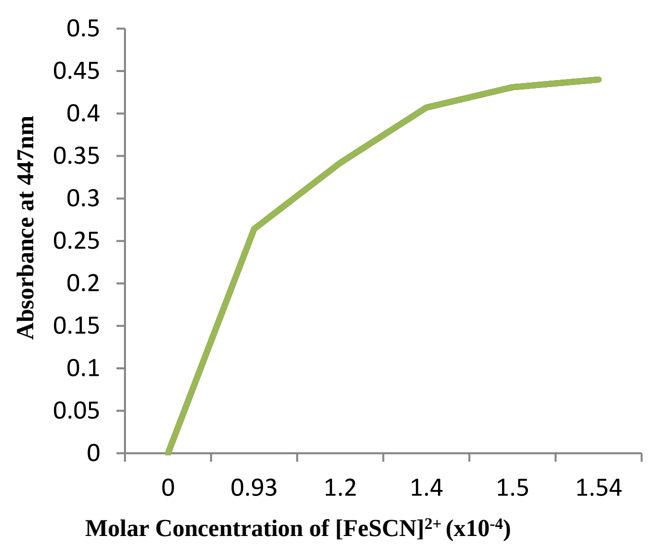

Figure 1. Graph of absorbance versus concentration of $[FeSCN]^{2+}$

The calibration curve for the $[FeSCN]^{2+}$ complex demonstrates a linear correlation between absorbance and concentration, confirming the validity of the Beer-Lambert Law for this reaction.

The average K_C value and the standard deviation were calculated from the individual K_C value obtained for each sample.

$$\text{Average Kc} = \frac{\sum \mathbf{Kc}}{\mathbf{N}} \tag{6}$$

- Where N is the number of samples

$$SD = \sqrt{\frac{\sum(Xi-\bar{X})^2}{N-1}} \tag{7}$$

- Where Xi is the value for each sample and X is the average value of all the trials.

Using Equation (6), the average K_C was calculated and had a value of K_{avg}=108.224.

The standard deviation was also calculated from Equation (7) and had a value of SD=7.4215.

A high value of Kc recommends that the equilibrium of the reaction is shifted to the right, meaning that the products dominate in the mixture at equilibrium. In contrast, a low value of Kc suggests that the equilibrium is shifted to the left.

The calculated Kc values range from approximately 102 to approximately 120 with an average value of 108. Literature values for the equilibrium constant of the reaction between Fe^{3+} and SCN^- typically range from 100 to 200, supporting the findings of the study. The small variation in Kc values suggests that the spectrophotometric method is a reliable approach for determining the equilibrium constant.

Additionally, the observed effect of concentration on the equilibrium constant aligns with Le Chatelier’s principle, which predicts that the equilibrium will shift in response to changes in the concentrations of reactants or products.

The value of the equilibrium constant Kc is significantly high across all the samples. This indicates that the products are present in greater amounts than the reactants, resulting in an equilibrium shift to the right [11].

The slight variations observed in the results of the K_C may be due to a change in temperature in the laboratory during the execution of the experiment. Also, they may be due to instrumental errors (e.g. incorrectly calibrated spectrophotometer) or personal errors (e.g. poor sampling). The concentration of reagents must be accurately determined, as insufficiency or excess may lead to deviations from Beer's law [9].

5. Conclusion and Future Scope

When Fe^{3+} ions and SCN^- ions are present together in a solution, they react and the iron(III) thiocyanate complex $[FeSCN]^{2+}$ is formed. In this study, it was demonstrated the determination of the chemical equilibrium constant K_C with succesful application of spectrophotometry by measuring the maximum absorbance of the $[FeSCN]^{2+}$ complex. The experimental results of equilibrium constant K_C were notably high in all the samples which indicates that the reaction in the chemical equilibrium is shifted to the right, suggesting that there are more products than reactants.

In this study, the equilibrium constant K_C for the reaction between iron (III) and thyocyanate ions have a range of 102 to approximately 120, which is consistent with literature values, confirming the reliability of spectrophotometry as a method.

Spectrophotometry, along with the Beer-Lambert law, has been demonstrated to be a realiable and effective method for measuring accurately the concentrations of colored substances, examining chemical equilibrium, and calculating equilibrium constants. The accuracy of the results can be further enhanced by improving the experimental parameters, including reaction time, temperature regulation, and correct calibration of the spectrophotometer. Future researche could explore the influence of temperature and ionic strength on the equilibrium constant K_C to further understand the thermodynamics of this reaction.

Data Availability
The data from the experiments conducted in this study can be found in this article.

Conflict of Interest
The author proclaims that has no conflict of interest.

Funding Source
None

Authors’ Contributions
The author carried out the research, reviewed and edited the manuscript, and approved the final version of the manuscript.

Acknowledgements
The author would like to acknowledge the Department of Chemistry of the University of Ioannina, Ioannina, Greece, for providing all the necessary laboratory facilities.

References

[1] D.G. Themelis, “Βασικές Αρχές Αναλυτικής Χημείας [Basic Principles of Analytical Chemistry]”, *5th Edition,* Ziti Publications, Greece, pp.**102-123**, **2018**

[2] N.D. Hatziliadis, “Εισαγωγή στην Ανόργανη και Γενική Χημεία [Introduction to Inorganic and General Chemistry]”, *2nd Edition,* Zorzos & SIA Publications, Greece, pp.**97-102**, **2010**

[3] J.Q. Pardo, “Do the Equilibrium Constants Have Units? A Discussion on How General Chemistry Textbooks Calculate and Report the Equilibrium Constants,” *International Journal of Physics and Chemistry Education,* Vol.**11**, No.**3**, pp.**73-83, 2019**

[4] S. Bose, “High Temperature Coating”, *1st Edition,* Butterworth-Heinemann Publisher, USA, pp. **5-16, 2007**

[5] J. Tan, K.S. Chan, “Understanding Advanced Physical Inorganic Chemistry, The learner's approach”, *1st Edition,* WS Education Publications, Singapore, pp.**229-255**, **2011**

[6] T.P.Hatziioannou, “Χημική Ισορροπία και Ανόργανη Ποιοτική Ημιμικροανάλυση [Chemical Equilibrium, and Inorganic Qualitative Semi-microanalysis]”, *1st Edition,* Dora Mavrommati Publications, Greece, pp.**64**, **69-76**, **1993**

[7] E. Keszei, M.G. Takacs, B. Vizkeleti, “a Straightforward Method to Determine Equilibrium Constants from Spectrophotometric Data,” *Journal of Chemical Education*, Vol.**77**, No.**7**, pp.**927, 2000**

[8] W.A.E. Mcbryde, “Spectrophotometric Determination of Equilibrium Constants in Solution”, *Talanta*, Vol.**21**, No.**10**, pp.**979-1004, 1974**

[9] T.P. Hatziioannou, M.A. Kouppari, “Ενόργανη Ανάλυση [Instrumental Analysis]”, *7th Edition,* D. Mavrommati Publications, Greece, pp.**20-24, 168-172**, **2010**

[10] R. Rammete, “Equilibrium Constants from Spectrophotometric Data: Principles, Practice, and Programming”, *Journal of Chemical Education,* Vol.**44**, No.**11**, pp.**647-754, 1967**

[11] M. Clugston, R. Flemming, “Advanced Chemistry,” *1st Edition,* Oxford University Press Publications, Malaysia, pp.**164-174, 2000**

AUTHORS PROFILE

Ilias Darsaklis earned his B.Sc. in Chemistry from the Department of Chemistry of the University of Ioannina, Ioannina, Greece. He later pursued his M.Sc. in «Technology and Quality of Table Olives and Olive Oil» from the Department of Food Technology of the University of Peloponnese, Kalamata, Greece.