Varied analysis of Acid Dissociation Constants by cyclic voltammetry and UV-visible to determine the aqueous solubility of a new diisopropylammoniumphenylsulfonate molecule

ABSTRACT

The determination of dissociation constant, solubility and thermodynamic parameters are very important physico-chemical parameters in substances and their knowledge is of fundamental importance for the validation of a pharmaceutical ingredient target. The determination of these parameters for an agent candidate target diisopropylammoniumphenylsulfonate (besylate) ($PhSO_3$-iPr_2NH_2) was determined by voltammetric and UV-visible methods.

The voltammetric method gave pKa1= 3.03±0.21 and pKa2= 10.23±0.59, while the UV-visible method determined two pKas values, pKa1= 2.21±0.04 and pKa2= 10.77±0.42 respectively. The thermodynamic parameter values calculated for the enthalpy (ΔH) and entropy (ΔS) of $PhSO_3$-$iPr_2NH_2$are of the order of ΔH= 3429.96±82.30 $KJ.mol^{-1}$and ΔS = 11.85 ± 0.26 $KJ.mol^{-1}.K^{-1}$. In addition, the Gibbs free energy of the molecule decreases with increasing temperature and the solubility shows values between 1.3 and 70 mg/mL for pH values between 2.75 and 10.5 and reaches its maximum Smax = 70mg/mL at pH equals 2.75 and 10.5.

The various physico-chemical properties of $PhSO_3$-iPr_2NH_2, which are within the range of active pharmaceutical ingredients, could make it an excellent candidate pharmaceutical ingredient.

Keywords:diisopropylammoniumphenylsufonate, voltammetric method, UV-visible method, physicochemical parameters.

1. INTRODUCTION

The majority of drugs used in therapy have low water solubility, which can reduce bioavailability and the rate of dissolution[1].So, because of their low solubility, drugs designed to be therapeutically active become a real danger to humans.Water solubility is a key parameter in drug formulation, as it has a major influence on the pharmacokinetics and pharmacodynamics of the drug [2]. Therefore, it is essential to find a way of increasing the dissociation rate and bioavailability of drugs [3,4]. The aim is to develop a less expensive, easier and faster technique for increasing the solubility of these drugs [5]. In this regard, the salification of medicines or the formation of medicinal salts is proving to be the safest and most reliable way of increasing solubility without unwanted side effects [6]. To form a drug salt, the free acid or base of the drug is combined with the base or acid of a potential counterion in specific molar ratios in a suitable solvent system.These latter participate in ionic interactions and crystallize under favorable conditions to give the solid salt [7,8,9,10]. Among potential counterions, phenylsulphonate-based salts, commonly known as besylate, are widely used as active pharmaceutical ingredients (APIs) in pharmaceutical development [11,12].These salts help to improve aqueous solubility or increase the speed of dissolution thanks to their low toxicity potential and ease of synthesis [13]. It is in this context that we have witnessed a new era in the synthesis of besylate derivatives. For example,J. H. Seo et al [14]have developed a new technique for synthesising the besylate salt of cilostazol with the aim of improving the solubility of the drug cilostazol and its physico-chemical properties such as stability, bioavailability. All these substances have acidic or basic functional groups whose ionisation coefficients (pKa) affect their physico-chemical and biological properties.However, for the validation of any phenylsulphonate-derived counter-ion product, the determination of pKas values is a key parameter for the success of the salt to be formed[15,16]. As well as influencing the choice of counterion, it provides information on the stability of the product formed and the circulation of the drug in the organism.

To determine the acid dissociation constant values, several methods were used by the researchers, namely spectroscopy [17], electrophoresis [18], potentiometric titration [19], high-performance liquid chromatography (HPLC) and capillary electrophoresis (CE) [20]. UV-Visible spectrophotometry and cyclic voltammetry [21]are the most reliable methods for determining dissociation constants thanks to their accuracy and ease of use. They have advantages over other methods in that they are relatively simple and practical, requiring a small quantity of sample for accurate measurement and covering a wide range of pKa.

In the present work we have studied the thermodynamic parameters of a diisopropylammoniumphenylsulphonate (besylate), a new crystalline molecule that we have recently synthesised for the first time [22]. This diisopropylammonium comes from diisopropylamine, a product recently used with dichloroacetate to inhibit the propagation and multiplication of liver cells [23]. The pKa values and thermodynamic parameters were determined in an aqueous medium. These parameters associated with solubility were determined by voltammetry and UV-visible method.

2. PRODUCTS, MATERIALS AND PROCEDURES

2.1 Produits

Diisopropylammonium besylate ($PhSO_3$-iPr_2NH_2) with empirical formula ($C_{12}H_{21}NSO_3$) was synthesized by SEYE et al [22]. Hydrochloric acid (HCl) and sodium dihydroxide(NaOH) were used as the titrating solution.Aqueous solutions used were prepared by an ultrapure Milli-Q (MQ 18.2 MΩcm) water

2.2 Materials

Experiments were carried out using cyclic voltammetry and U-V Visible spectrophotometry. The pH meter and electronic balance were also used.

- ✓ Cyclic voltammetry allows us to visualize the oxidation and reduction current peaks of the PhSO$_3$-iPr$_2$NH$_2$ compound. It consists of a PSTRAS device linked to a cable with three electrodes (reference, counter and working electrode) immersed in a cell containing the solution, and another cable linked to a computer.
- ✓ The Thermos Fisher Scientific U-V Visible spectrophotometer, model G10S UV-Vis serial number 2L9U285217, plots absorbance as a function of wavelength.
- ✓ The pH/ mV/oC/oFmeter is a device for measuring the pH of a solution. It consists of an electronic box that displays the pH value and an electrode that measures this value.

2.3 Procedures

A three-electrode cell was used:

- the platinum working electrode, with a diameter of 4 mm;
- TheplatinumCounter électrode ;
- the Ag/AgCl reference electrode in KCl

The experiment was performed in the electrochemical cellcontaining 10mL of solution containing 5-10^{-2}M PhSO$_3$-iPr$_2$-NH$_2$ salt. Cyclic voltammetry measurements were then carried out after the addition of NaOH or HCl to vary pH values between 2-11 values.

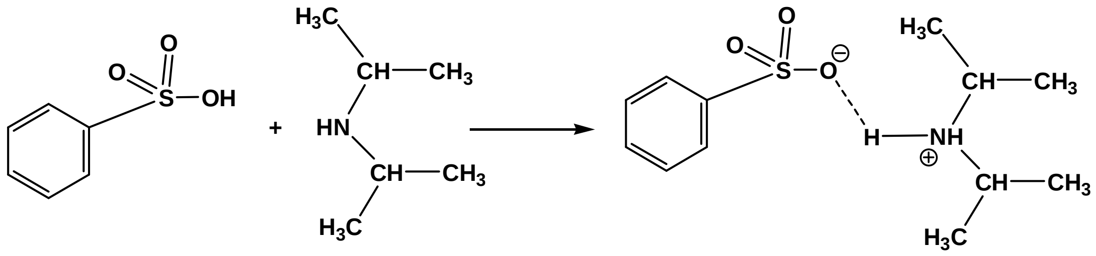
All these curves were recorded using a cyclic voltammetry in thepotential range between - 1 and 1.5 V/Ag/AgCl with a scan rate of 50mV/s.

UV-visible (UV-vis) absorption spectra of PhSO$_3$-iPr$_2$-NH$_2$ were recorded with a Thermo Fisher scientific UV-vis absorption spectrometer in 3mL of PhSO$_3$-iPr$_2$NH$_2$ of concentration 5.10^{-2} M. Experiments were repeated following the addition of titrating solution of NaOH (0.1M) and HCl (0.1M).

3. RESULTS AND DISCUSSION

3.1Proton Transfer Mechanism ofPhSO$_3$-iPr$_2$NH$_2$ salt

Diisopropylammoniumphenylsulfonate (PhSO$_3$-iPr$_2$NH$_2$) is a salt formed from a cation (or acidic site, diisopropylammonium) and an anion (or basic site, phenylsulfonate).The Scheme 1 represent the procedure for synthesis of PhSO$_3$-iPr$_2$NH$_2$ crystalline molecule and previously reported [22].PhSO$_3$-iPr$_2$NH$_2$ is capable of capturing or releasing a proton, depending on the nature of the medium. Diisopropylammoniumphenylsulfonate behaves like an amphoteric, being able to combine with both acids and bases.Neutralization of the acidic site (diisopropylammonium) takes place in a basic medium in the presence of NaOH (schema 2), followed by the formation of water, while neutralization of the basic site (phenylsulfonate) takes place in an acidic medium in the presence of HCl (schema 3).

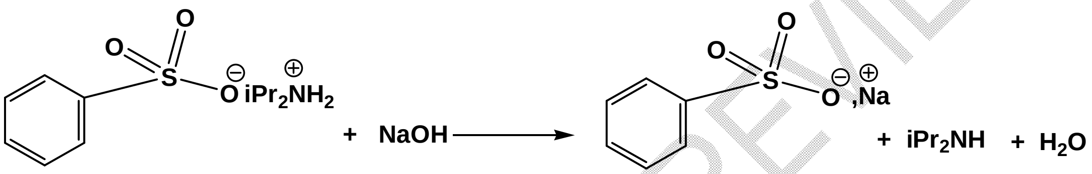


Phenylsulfonic acid ($PhSO_3$)

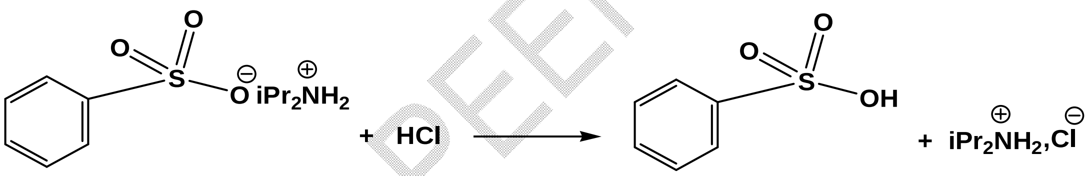
diisopropylamine (iPr_2NH_2)

Diisopropylammoniumphenylsulfonate ($PhSO_3$-iPr_2NH_2)

Schema 1. Procedure for synthesis of $PhSO_3$-iPr_2NH_2



Schema 2.$PhSO_3$-iPr_2NH_2 reaction equation in a basic medium



Schema 3. $PhSO_3$-iPr_2NH_2 reaction equation in acid medium

3.2 Study of the behavior of PhS-iPr_2NH_2

3.2.1 <u>By electrochemical method</u>

Figure 1 shows cyclic voltammetry on the bare platinum electrode immersed in 10mL of 5.10^{-2}M $PhSO_3$-iPr_2NH_2 solution. Indeed, by cyclic voltammetry, the potential between - 1 and 1.5 V/Ag/AgCl at a scan rate of 50 mV/s, we observe the presence of two oxidation peaks and two reduction peaks corresponding to the redox behavior of the $PhSO_3$-iPr_2NH_2 crystalline molecule in solution on the working electrode.

Study of the voltammetric curve (Figure 1) shows the presence of two anodic peaks at around 0.88 and -0.5 V/Ag/AgCl, and two reduction peaks at around -0.67 and 0.12 V/Ag/AgCl. The potentials of the oxidation peak at 0.88 V/Ag/AgCl and the reduction peak at 0.12 V/Ag/AgCl are very similar to those described for an aliphatic amine [24]. The presence of the second redox couple (-0.5 and - 0.67 V/Ag/AgCl) is linked to the presence of the phenylsulfonate group. These results show that the PhS-iPr_2NH_2 crystalline molecule is an electroactive compound in solution, exhibiting a reversible characteristic and confirming that it possesses two oxidizing/reducing couples.

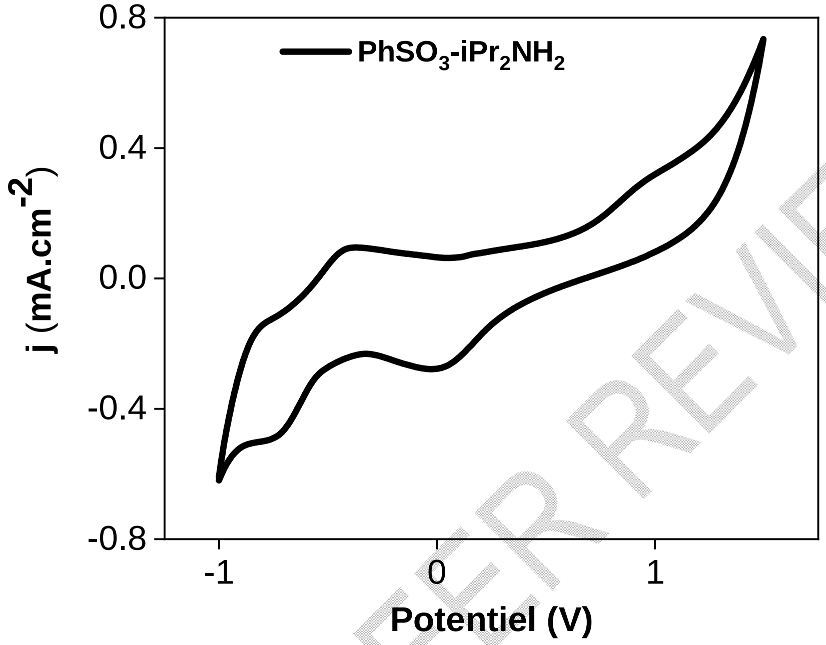


Fig. 1: Cyclic voltammetry curve for$PhSO_3$-iPr_2NH_2 (0.05M) in aqueous medium

3.2.2 <u>By UV-Visible method</u>

Figure 2 shows the UV-visible spectrum of $PhSO_3$-iPr_2NH_2 salt in aqueous medium recorded between 200-500 nm. This figure shows a maximum absorption band around 300nm.

This absorption band is due to the л→л* electronic transitions of the phenyl group. These results show that the crystalline molecule possesses luminescence properties.

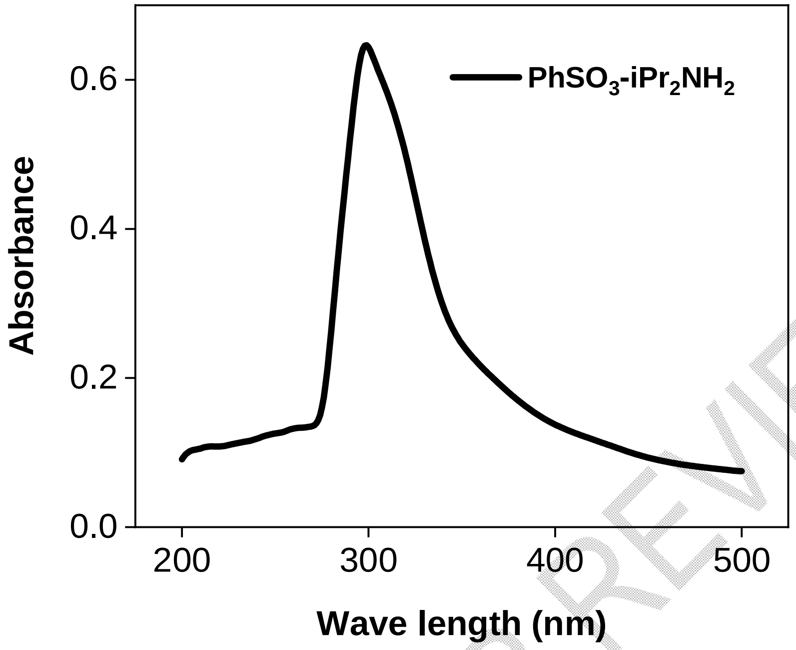


Fig.2. UV-visible curve of $PhSO_3$-iPr_2NH_2 salt (0.05M)in aqueous medium

3.3 Process for the determination of pKa values

3.3.1 By electrochemical method

The results obtained previously illustrate that the compound $PhSO_3$-iPr_2NH_2 has two oxidizing/reducing pairs as mentioned above. For pKa determination by cyclic voltammetry, oxidation and reduction currents are measured as a function of solution pH after each addition of titrant solution. After each addition of a quantity of HCl or NaOH to the electrochemical cell containing 0.05 M $PhSO_3$-iPr_2NH_2, the pH of the solution is modified, accompanied by a more intense resulting current.

For an acid-base equilibrium ($AH \rightarrow A^- + H^+$), the pKa is given by [25]:

$$pK_a = pH - \log\left(\frac{[A^-]}{[HA]}\right)$$

With [AH] and $[A^-]$ the respective concentrations of the acid and its conjugate base. The resulting total current is $I = I_{A-}[A^-] + I_{AH}[AH]$.

This is a method for determining pKa based on anodic (I_{AH}) and cathodic (I_A^-) currents by applying the equation. After combining all these data, we obtainEq1:

$pKa = pH - \log\left(\frac{I_{AH} - I_{A^-}}{I - I_{A^-}} - 1\right)$(Eq1)[26]

3.3.1.1 Effect of HCl acid addition

Figure 3 shows voltammograms of the response of the 0.05 M solution of $PhSO_3iPr_2NH_2$ in aqueous medium in the absence and presence of added HCl (0.1M). It can be seen that the addition of HCl to $PhSO_3$-iPr_2NH_2 leads to an increase in the oxidation peak and the reduction peak characteristic of the $C_6H_5SO_3H/C_6H_5SO_3$- couple (schema 2). Figure3a shows that, as a function of the progressive addition of HCl, the oxidation peak at -0.5 V/Ag/AgCl shifts towards positive potentials, while the reduction peak at -0.67 V/Ag/AgCl tends towards negative potentials. These results show that the reaction between $PhSO_3$-iPr_2NH_2 and H+ increases the charge density of the molecule [27]. In addition, the basic site of $PhSO_3$-iPr_2NH_2 undergoes neutralization, resulting in a decrease in the pH value, which tends towards acidic pH values. On the other hand, the redox couple at around 0.88 and 0.12 V/Ag/AgCl remains unchanged, suggesting that the $PhSO_3$-iPr_2NH_2 molecule is in acid-base salt form, and could be confirmed by the same technique, this time with the addition of a titrating NaOH solution.

3.3.1.2Effect of NaOH base addition

The acid-base behavior of the $PhSO_3$-iPr_2NH_2 crystalline molecule was studied by adding a NaOH (0.1M) titrating solution. For this purpose, various measurements were recorded in an aqueous solution of 0.05 M $PhSO_3$-iPr_2NH_2 after each addition of a quantity of NaOH titrating solution and represented in Figure 3b. These curves show a progressive increase in the intensity of the oxidation peaks as a function of the volume of NaOH added. Figure 3b reveals that the oxidation peak at around 0.88 V/Ag/AgCl shifts towards positive potentials, while that at - 0.5 V/Ag/AgCl remains practically unchanged.

These results confirm those previously described and indicate that the strong base NaOH attacks the acid site (Scheme 1). The addition of NaOH in solution in the presence of the $PhSO_3$-iPr_2NH_2 crystalline molecule enhances the electrochemical properties, leading to an increase in the charge density and consequently the conductivity of the medium. Electron donors can generate a high charge density. In basic media, an improvement in the rate of charge transfer is noted, leading to an increase in current intensity [28]. Furthermore, the reduction peak disappeared at around - 0.67V/Ag/AgCl in the presence of NaOH. These results indicate that the presence of NaOH in the electrolyte medium blocks the reduction process of the $PhSO_3$-iPr_2NH_2 compound in solution.

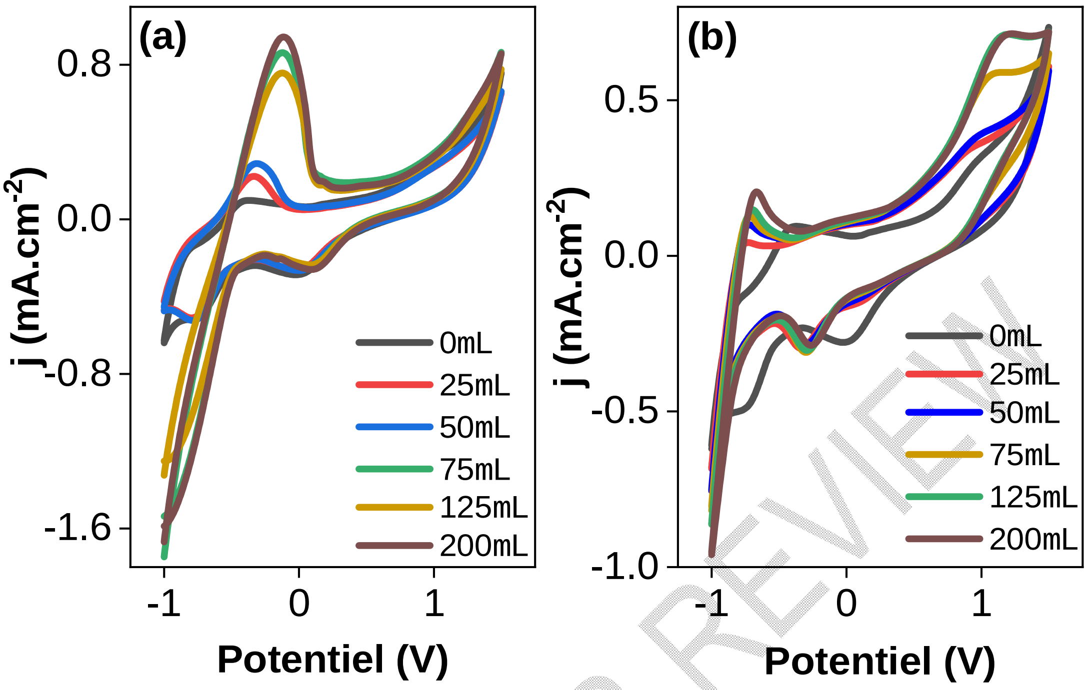


Fig. 3.Voltammetric curves for $PhSO_3$-iPr_2NH_2 saltin aqueous mediumin the absence and presence of (a) HCl (b) NaOH

3.3.2 UV-visible spectroscopy

UV-Visible spectrophotometry provides information on the excitation wavelength of compounds, and on the difference in chemical structure between compounds. To determine the acid dissociation constant pKa of a molecule accurately and reliably, UV-visible spectrophotometry is unquestionably the method of choice, especially if the substance is too insoluble. In this method, the pKa depends on the ionized forms of the molecule and the pH. The extent of a compound's ionization plays a fundamental role in characterizing its absorption, distribution, metabolism and excretion (ADME) profile [29,30].

Two methods are used to determine pKa values:

First method: this is a graphical method, plotting the pH curve as a function of $\log(A_m/A_i)$.

If $\log(A_m/A_i)$ =0, the corresponding pH value is equal to the pKa value.

Second method: this is a direct method, depending on the nature of the ionized and non-ionized forms of the compound at different pH values. It provides information on the accuracy of the first method by applying the equation below (Eq. 2).

In UV-Visible spectrophotometry, the pKa value is given by the following equation:

$$pKa = pH + \log\frac{A_m - A}{A - A_i} \qquad \text{(Eq 2)}$$

A absorbance des espèces intermédiaires
A_m absorbance de l'espèce moléculaire

A_i absorbance l'espèce ionisée
pH représente la valeur du pH du milieu intermédiaire [31,32].

For this method, the pKa value depends on the absorbance of the ionized forms, hence the need to ionize the molecule by adding acid or base.

3.3.2.1 Effect of NaOH addition

Figures4a, 4b and 4cshow the UV-visible spectra of $PhSO_3$-iPr_2NH_2 *in solution as a function of the amount of NaOH added. It can be seen that the addition of NaOH (0 – 800 µL) to the* $PhSO_3$-iPr_2NH_2 *solution first leads to a decrease in the intensity of the absorption band around 300 nm (Figure 4a). This decrease is due to the interactions of the OH- base with the proton of the cationic function of* $PhSO_3$-iPr_2NH_2 *(-O----H-), which leads to the progressive formation of the conjugated base in solution. This results in a decrease in the concentration of the cationic function, leading to a drop in absorption intensity. The latter, responsible for absorption emission at the same wavelength, increasingly loses its supremacy, leading to a hypochromic effect.*

These results correlate with those developed by A. Garcia-Leis in the case of the UV-visible study of the compound 2,2'-azino-bis (3-ethylbenzothiazoline-6-sulfonic acid (ABTS) [33].

Secondly, there is an increase in absorption intensity at the same wavelength as before, from 800 µL up to 1300 µL of added NaOH (Figure 4b). In this case, the hyperchromic effect is observed, showing that at 800 µL the conjugated base becomes the majority in solution and substitutes the cationic function in the absorption emission. And finally, with NaOH volumes above 1300 µL (figure 4c), a decrease in absorption intensities is again noted.

These results show that the entire cationic part of $PhSO_3$-iPr_2NH_2 *is completely neutralized and the concentration of the conjugated base responsible for the absorption band decreases with dilution.*

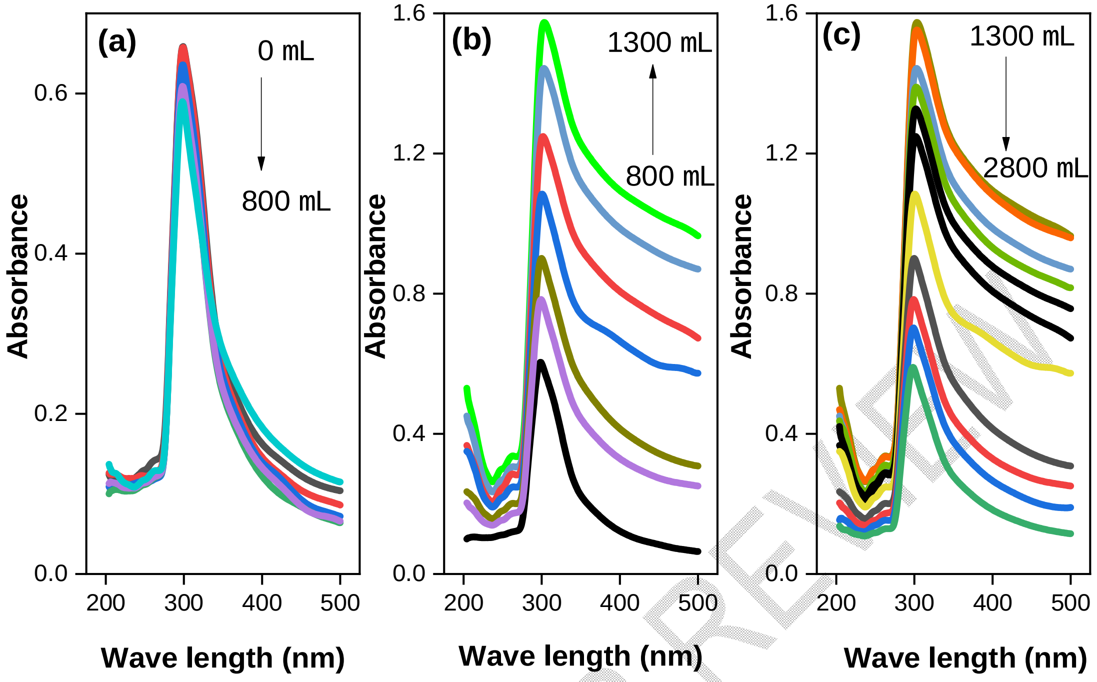


Fig. 4. Superposition curves of UV-visible spectra of $PhSO_3$-iPr_2NH_2 at different NaOH volumes (a) range from 0 to 800 µL (b) range from 800 to 1300 µL and (c) range from 1300 to 2800 µL.

3.3.2.2 Effect of HCl acid addition

The UV-visible spectra of the $PhSO_3$-iPr_2NH_2 crystalline molecule in solution as a function of the volume of HCl (0-5000 µL) added are shown in Figure 5. As a function of the amount of HCl added, the absorbance band around 300nm decreases considerably, reflecting the attachment of the H^+ proton to the O^- of the phenylsulfonate function, resulting in an interaction (-O-H-) and the formation of the conjugate acid of the $PhSO_3$-iPr_2NH_2 anion function. Moreover, these interactions lead to protonation of the anionic part of $PhSO_3$-iPr_2NH_2, followed by a reduction in the strength of the latter in solution, the concentration of which is proportional to absorption.

These results point to a hypochromic effect due to the action of HCl on $PhSO_3$-iPr_2NH_2 in solution [31].

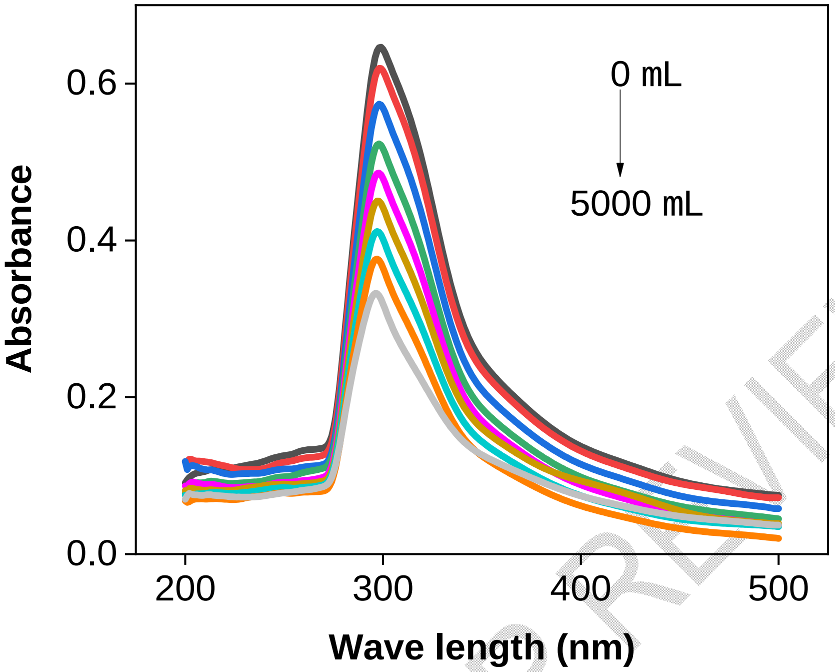


Fig. 5. UV-visible curve of $PhSO_3$-iPr_2NH_2 (0.05M) salt with HCl (0.1N) 0 - 5000 µL.

3.4Calculation of pKa values

3.4.1. By electrochemical method

Figure 6a shows the variation of anodic currents as a function of pH in acidic medium. It provides information on the range of pKa values in the region where current intensity tends towards the maximum value [34].In this case, the ($C_6H_5SO_3H/C_6H_5SO_3^-$) couple is brought into play by the amount of acid added, and is accompanied by an increase in the anodic and cathodic currents. This increase is linked to the degree of interaction of the hydrogen bond [35].In addition, the curve shows a vertical tangent inflection between 2.9 and 3.1, corresponding to neutralization of the anionic part of the PhS-iPr_2NH_2 crystalline molecule in solution. These results correlate with those found by J. Zhao [36]for cyclic voltammetry titration. Applying the formula linking pKa to pH and currents, anodic and cathodic (Eq1), we find a pKa value of the order of 3.03±0.21. The figure 6bcorresponding to the variation of anodic currents as a function of pH in basic medium shows a curve with the appearance of a classic weak acid-strong base assay, and a significant pH jump between 10.4-11.2. It provides information on the pKa value page, which lies in the pH range where the anode current approaches its maximum value. The data allow us to determine the pKa value by applying equation 2, and we find a value of pKa2= 10.23±0.59.

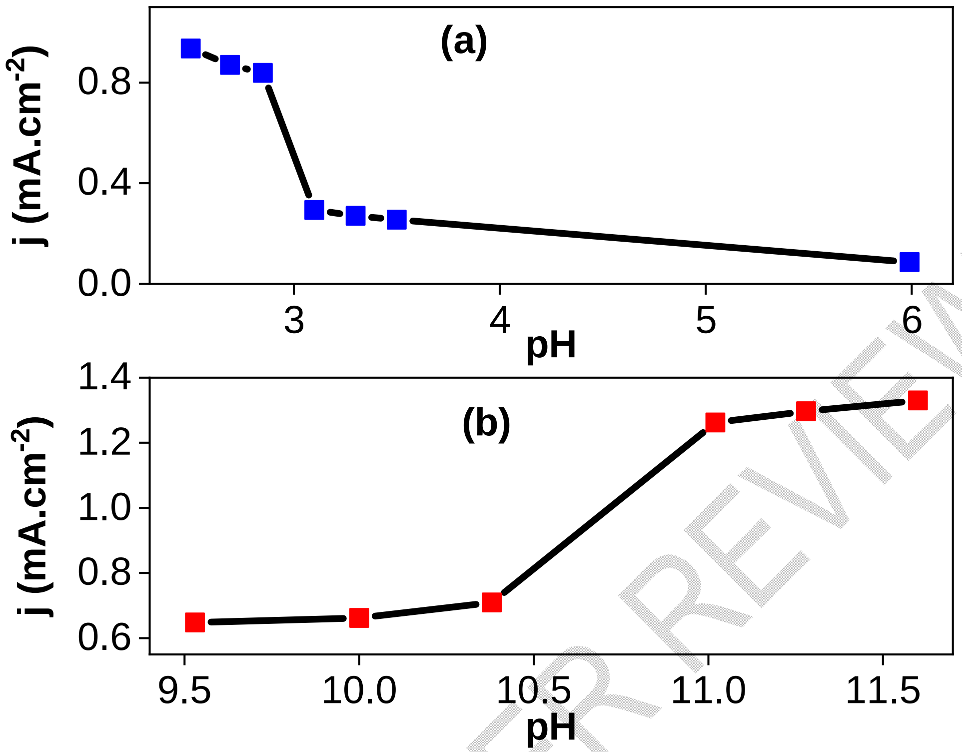


Fig. 6. Current density curve as a function of pH in (a) acidic medium and (b) basic medium

3.4.2 UV-visible analysis

To confirm the pKas values found electrochemically, the UV-visible analysis technique is commonly used. For the determination of pKas values using the first method, pH curves as a function of log(Am/Ai) are often exploited (Figures 7a and 7b).In Figure 7a, the decimal logarithm of the absorbance ratio of the molecular form to the ionized form increases with pH. Graphically, the value pKa_2=11.03 is found for $C_6H_{14}NH_2^+/C_6H_{14}NH$, corresponding to pH = pKa under conditions where $\log(A_m/A_i)$ = 0. The second method (Eq2), whose data gives a value of pKa_2=10.77±0.42. Figure 7b, on the other hand, shows the pKa value for the ($C_6H_5SO_3H/C_6H_5SO_3^-$) couple. Applying the first method corresponding to the variation of pH as a function of log(Am/Ai) gives a pKa_1 value of the order of 2.4 in the case where $\log(A_m/A_i)$ =0. The second method (Eq2) gives a pKa1 value of 2.21±0.04. These pKa values found by the two methods give a difference of around 0.3, confirming the reliability of the UV-visible technique.

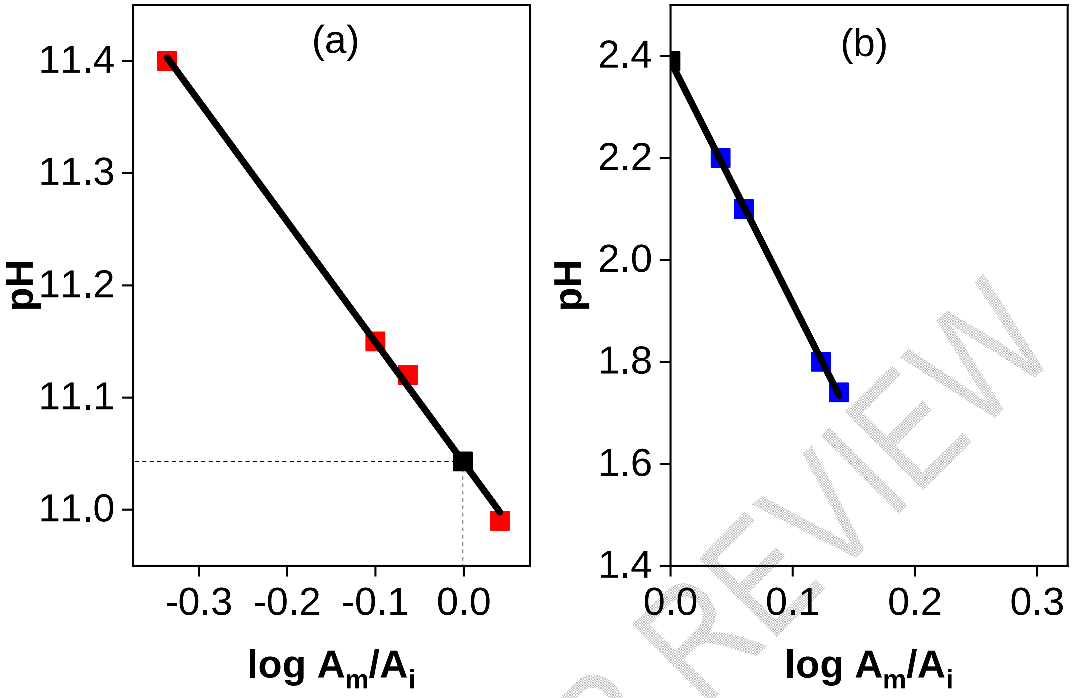


Fig. 7. pH vs. log(A_m/A_i) curve (a) in basic medium and (b) in acidic medium.

3.5Determination of thermodynamic parameters

Calculation of thermodynamic parameters: ΔG, ΔH and ΔS
Thermodynamic parameters are important in molecular chemistry. For the determination of these parameters, the Vant'Hoff equation is the most used(Eq 3).

$$\frac{\mathrm{dlnKa}}{\mathrm{dT}} = \frac{\Delta \mathrm{H}}{\mathrm{RT}^2} \text{(Eq 3)}$$

ΔH molar enthalpy,
R universal gas constant and is equal to 8.314 $J.mol^{-1}.K^{-1}$
T temperature in Kelvin (K).
The Gibbs free energy ΔG is also linked to the reaction constant by the relation(Eq 4):

$$\Delta G = -RT\ln Ka \text{ (Eq 4)}$$

It is possible to determine the Gibbs free energy from the absorbance value obtained by UV-visible spectroscopy by applying the following relationship(Eq 5):

$$\Delta G = -2.303RT\log(A) \text{ (Eq 5)}$$

WithA is the absorbance value [37].

The results obtained from equation 5 are shown in table 1.

Table 1: value of absorbances and ΔG as a function of temperature.

Température T (K)	298	303	308	313	318	323
Absorbance	0,890	0,970	1,090	1,112	1,135	1,160
$\Delta G (KJ.mol^{-1})$	288,738	76,745	-220,716	-276,308	-334,858	-398,642

The relationship between the Gibbs free energy ΔG, the entropy ΔS and the enthalpy ΔH is given by the third law of thermodynamics(Eq 6):

$$\Delta G = \Delta H - T\Delta S \text{ (Eq 6) [38].}$$

Eq 6 makes it possible to deduce the values of ΔS and ΔH by plotting the curve ΔG as a function of the temperature T (figure 8). The intercept coincides with the value of ΔH and the slope corresponds to the opposite of ΔS.

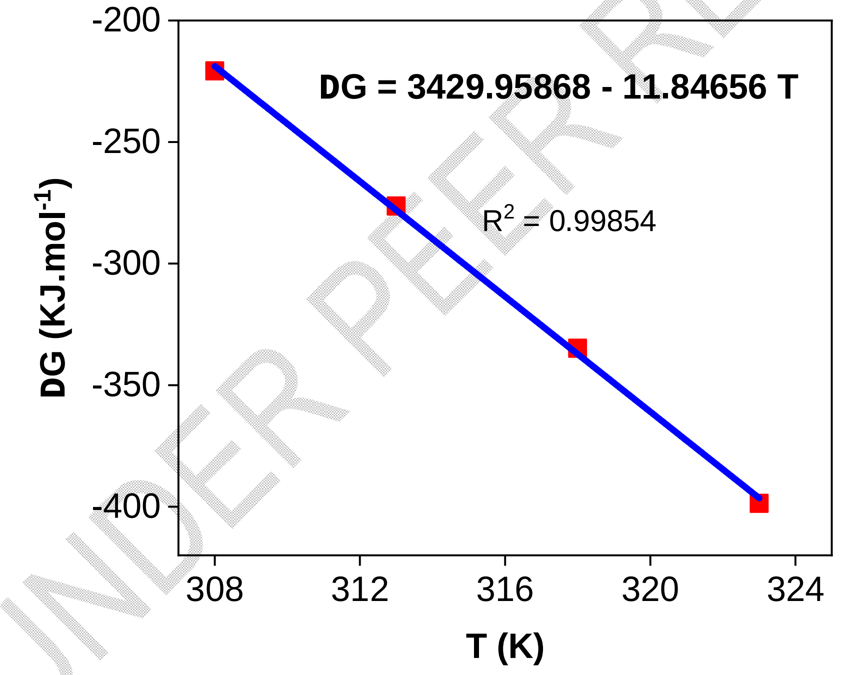


Fig. 8. Variation curve of ΔG as a function of temperature T

The obtained values of standard enthalpy, entropy and Gibbs free energy of the dissociation of $PhSO_3$-iPr_2NH_2 are given in Table 2. It seems that the values of ΔG decrease with increasing temperature. These results are similar to those of Li [39], which justifies that the dissociation process of $PhSO_3$-iPr_2NH_2 increases with increasing temperature. The positive value of ΔH= 3429.96 $kJ.mol^{-1}$ indicates that the dissociation is endothermic [40]. The positive values of ΔG indicate that the dissociation process is not spontaneous for the temperature between 298 and 303 K. The decreases of ΔG as a function of temperature showed that the dissociation of the $PhSO_3$-iPr_2NH_2 crystalline molecule is favored by the increase in temperature. However, the value of $\Delta S = 11.85 \pm 0.26$ $J.mol^{-1}.K^{-1}$ due to increased disorder resulting from dissociation processes. All these thermodynamic parameters of $PhSO_3$-iPr_2NH_2 confirm the stability of the crystalline molecule in an aqueous medium.

Table 2: The different values of ΔG, ΔH and ΔS for the PhS-iPr_2NH_2 crystalline molecule in solution

Température (K)	$\Delta G(KJ.mol^{-1})$	$\Delta H(KJ.mol^{-1})$	$\Delta S(J.mol^{-1}.K^{-1})$
298	288.74±11.43	3429,96 ± 82,30	11,85 ± 0.26
303	76.75±2.99		
308	-220.72±8.48		
313	-276.31±10.45		
318	-334.86±12.46		
323	-398.64±11.84		

3.6Determination of solubility

For the determination of the solubility of diisopropylammoniumphenylsulfonate, an amphoric compound composed of a monoacid diisopropylammonium and a monobasephenylsulfonate, the Henderson-Hasselbalch (HH) equation was used [41]

For a monoacid: The HH equation for a monoacid is given byEq 7:

$$S = S_0\ (10^{pH-pKa} + 1) \qquad (Eq\ 7) \quad [42]$$

Figure 9a shows the solubility of the $C_6H_{14}NH_2^+$ cation as a function of pH. We see that the free cationic form has a value of S_{max}=33.5 mg/mL with an intrinsic solubility value S_0=1.3mg/mL.

For a salt: the HH equation for a salt composed of acid and base is given byEq 8:

$$\log S = \log S_0 + \log\ (10^{pKa1-pH} + 10^{pH-pKa2} + 1)(Eq\ 8)\ [43]$$

Figure 9b shows the variation in solubility as a function of pH of PhS-iPr_2NH_2, an amphoteric compound. For this amphoteric the variation in pH leads to an exponential increase in solubility with an S_{max} value approximately equal to 70 mg/mL double the solubility of the salt forms of the free acid. These results show that the presence of the besylate anion increases the solubility of diisopropylammonium to a value practically equal to twice the maximum solubility of the cation alone. This once again shows the importance of the besylate anion in improving the solubility of pharmaceutical compounds. Furthermore, the salt of diisopropylammonium besylate or (diisopropylammoniumphenylsufonate) has an aqueous solubility of approximately 70 mg/mL greater than that of amlodipine besylate and bepotastine besylate with respective solubility of 2.22 mg/mL and 23.3 mg/mL [44]. This study demonstrates that diispropylammonium besylate is a potential API and could be developed in the design of pharmaceutical salts.

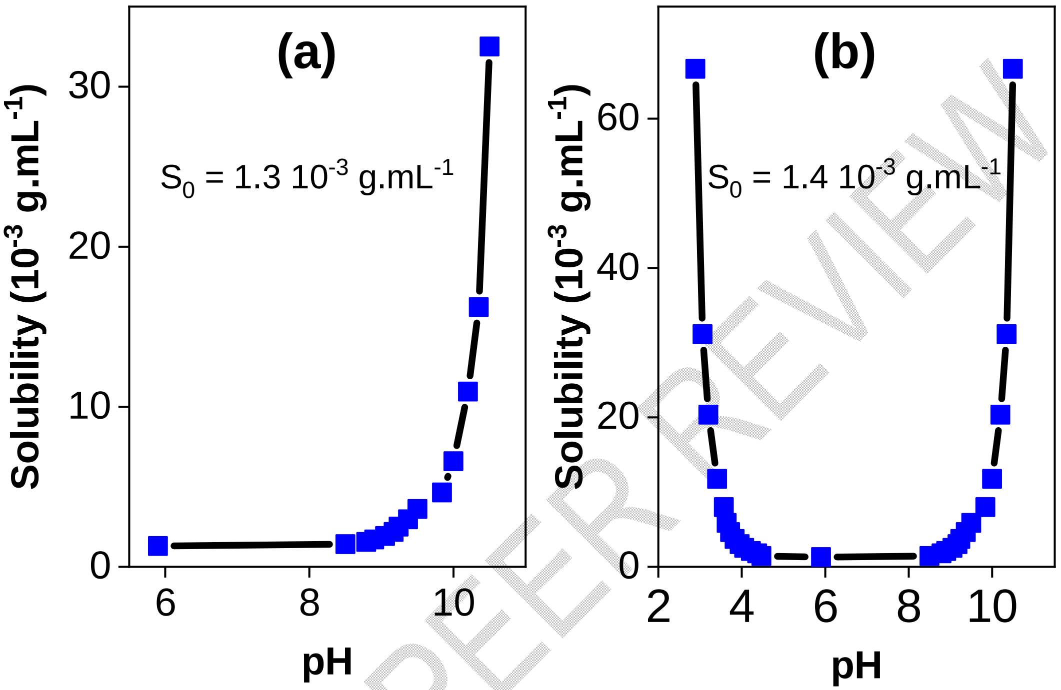


Fig. 9. pH-solubility curve (a) of the diisopropylammonium cation $C_6H_{14}NH_2^+$ (b) of $PhSO_3$-iPr_2NH_2

4. CONCLUSION

Knowledge of the characteristics of an API is an important parameter for the design of a pharmaceutical salt meeting the standard required by the World Health Organization (WHO). Our work consists of studying the physicochemical parameters of diisopropylammoniumphenylsulfonate, a potential API which could be used as a counterion for the formation of pharmaceutical salts. In our previous experiments, it has been demonstrated that diisopropylammoniumphenylsulfonate behaves like an amphoteric because it is capable of capturing a proton in an acidic medium and giving it up in a basic medium.In addition, the cyclic voltammetry curve confirmed that PhS-iPr_2NH_2 had two acid-base pairs: sulfonic acid/phenylsulfonate $C_6H_5SO_3H/C_6H_5SO_3^-$ and diisopropylammonium/diisopropylamine $C_6H_{14}NH_2^+/C_6H_{14}NH$.The thermodynamic parameters of $PhSO_3$-iPr_2NH_2 and the pKas values were determined by cyclic voltammetry and UV-visible. These two methods made it possible to obtain the pKa values of the two couples with a few errors. These methods also allowed us to calculate the solubility value of the molecule which is much higher than that developed in the literature and accepted as an active pharmaceutical ingredient.

The physicochemical parameters obtained are in the range of APIs and are satisfactory for the use of $PhSO_3$-iPr_2NH_2 as a counterion for the synthesis of pharmaceutical salts. The $PhSO_3$-iPr_2NH_2 crystalline molecule could be a potential API candidate with the aim of improving the solubility, efficacy, bioavailability and safety of drug molecules.

DISCLAIMER (ARTIFICIAL INTELLIGENCE)

Authors declare that no generative AI technologies such as large language models (chatgpt, copilot, etc.) and text-to-image generators have been used during the writing or editing of this manuscript

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