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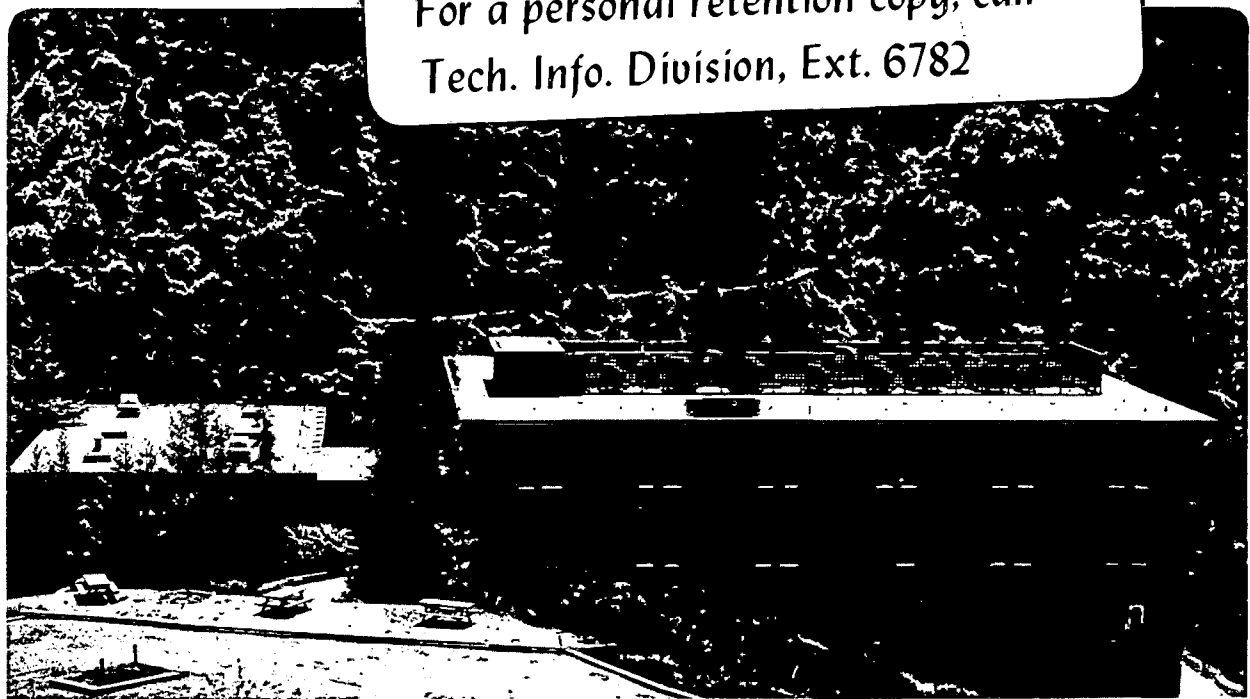
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INITIAL STUDY OF AN AMBIENT TEMPERATURE RECHARGEABLE
LITHIUM-SULFUR DIOXIDE BATTERY

William Bowden and A.N. Dey

June 1981

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Initial Study of an Ambient Temperature
Rechargeable Lithium-Sulfur Dioxide Battery

Final Report on LBL Subcontract 4507410

June 1, 1980 to November 30, 1980

Prepared for

LBL Project Manager: Philip N. Ross
Lawrence Berkeley Laboratory
University of California
Berkeley, CA 94720

Work conducted at:

DURACELL INTERNATIONAL INC.
Laboratory for Physical Science
Northwest Industrial Park
Burlington, Massachusetts 01803

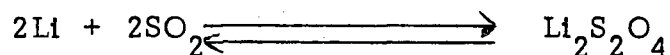
Principle Investigators:

William Bowden
A. N. Dey

This work was supported by the Assistant Secretary of Conservation and Renewable Energy, Office of Advanced Conservation Technology, Division of Electrochemical Systems Research of the U.S. Department of Energy under Contract Number W-7405-ENG-48.

Abstract

We made all-inorganic Li/SO₂ prototype rechargeable D cells with reference electrodes and demonstrated that SO₂ is deposited quantitatively in the catalytic carbon cathode as lithium dithionite (Li₂S₂O₄) during cell discharge. We demonstrated that the Li₂S₂O₄ is then quantitatively oxidized to SO₂ during the recharge portion of the cell cycle, according to the overall reaction



We found that the electrolyte salt Li₂B₁₀Cl₁₀ is oxidized in solution at slightly over the recharge potential of Li/SO₂ to give a one electron oxidation product. The oxidation product undergoes a further one electron oxidation to give a second product at potentials greater than those attained in cell recharge.

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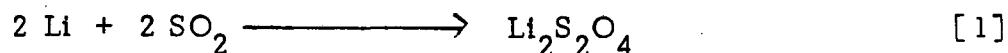
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I. Introduction

The Li/SO₂ system is commercially available in a number of cell sizes as a primary battery system (1). Commercially available cells deliver an energy density of 150 W.hr/lb (330 Whr/Kg) or 8 W.hr/in³ and power densities of 380 W/lb (836 W/Kg) or 21 W/in³. The Li/SO₂ primary system consists of a Li foil anode, a porous carbon - PTFE catalytic current collector, and a catholyte consisting of acetonitrile, SO₂ and LiBr supporting salt (1, 2). The postulated cell reaction [1] is based on qualitative tests (3).

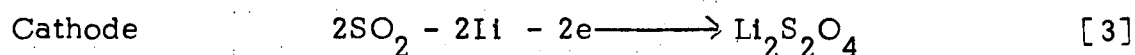
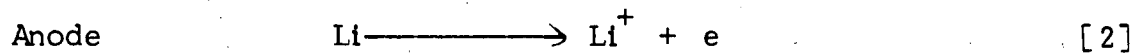


The cell reaction product is lithium dithionite, Li₂S₂O₄, an insoluble solid which precipitates in the porous carbon cathode, blocking further reduction of SO₂. During storage and discharge the lithium anode is protected from chemical reaction with SO₂ by formation of a protective film of Li₂S₂O₄.

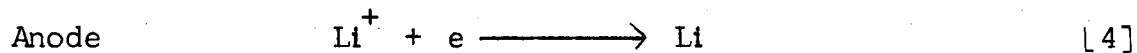
The reversibility of the cell reactions is known (4). However, the cycle life is limited by the irreversible reaction of electrodeposited lithium with the acetonitrile organic cosolvent (5) forming lithium cyanide, methane and other more complex organic compounds (6) which could not be reoxidized back to soluble lithium salts. These lithium-solvent reactions lead to an irreversible loss of lithium. This problem can be eliminated by using the salt Li₂B₁₀Cl₁₀ which is soluble in SO₂ and provides a highly conducting solution. In the Li/SO₂ system, comprised of a Li anode, porous carbon cathode and liquid SO₂ cathode active material cum solvent with Li₂B₁₀Cl₁₀ electrolyte salt the only reaction occurring is the straightforward and reversible reaction [1].

The salt Li₂B₁₀Cl₁₀ is a member of the class of polyhedral borane anions and known to be extremely stable (7). There is some evidence to indicate that the B₁₀Cl₁₀²⁻ anion may be oxidized during charging to form a stable, soluble species. This oxidation of B₁₀Cl₁₀⁻² is important in terms of providing overcharge protection of individual cells to balance the cells in a multicell battery. For example, in a lead-acid battery the electrolysis of solvent (water) provides this function.

Discharge:



Recharge:



The main goal for this six month program (LBL Contract #4507410) is to establish the validity or lack thereof of the above postulated reactions.

The work statement for this program is as follows:

Investigate the properties of a rechargeable all inorganic electrolyte lithium battery consisting of Li or Li-alloy negative electrode, porous carbon positive electrode and $\text{Li}_2\text{B}_{10}\text{Cl}_{10} - \text{SO}_2$ electrolyte.

Task I. Use hermetically sealed D cells with a Li reference electrode.

Task II. Identify the chemistry of the positive electrode reaction and verify that the discharge reaction products are rechargeable.

Task III. Report on the feasibility of this concept as a battery technology for use in electric vehicle or energy storage.

The results obtained during this contract are presented here.

II. Experimental

A. D Cell

To develop a hermetic D cell with reference electrode we selected the nickel plated cold rolled steel can used for Li/SO₂ cells that our company manufactures. This can incorporates both a filling eyelet and convoluted hermetic vent structure in the base as shown in Figure 1. The convoluted vent structure is designed to open under abusive conditions to render the cell safe. The identical vent or some modification may serve the same purpose with the rechargeable inorganic electrolyte cell. We fabricated a new top for the cell which incorporated two hermetic G/M seals instead of the usual one. We made both the G/M seals and cell tops in our laboratory as shown in Figure 2. As in a typical production cell, the porous carbon cathode is connected to the center post of a G/M seal by a welded tab while the lithium anode is connected to the can by a metal tab. The reference electrode was a Li foil which was wrapped with microporous polypropylene separator, inserted into the electrode stack and welded to the centerpost of the second G/M seal. The anode and cathode are wound together in a jelly roll structure, insulated by layers of microporous polypropylene separator as shown in Figure 3. The complete electrode stack is inserted into the can, the electrode connections welded, and the cell top is attached by TIG welding. To fill, the cell is evacuated through the fill port and filled with SO₂ electrolyte under pressure. The fill port is crimped shut and sealed by resistance welding to give a hermetically sealed cell.

B. Chemical Analysis

Chemical analyses on the cells were performed by our analytical laboratory. The selected cells were punctured in a sealed bomb specially designed and built for the task, after which the volatile contents of the cell were removed under vacuum on a vacuum line we built for the task. The electrolyte was analyzed for SO₂ by titration with NaOH in CH₃CN to a bromcresol green endpoint. The spent cells were then taken into an argon filled glove box

and the top removed with a tubing cutter. The electrode stacks were removed and disassembled. The anodes were discarded after inspection while the cathodes were analyzed for dithionite (8). The dithionite was leached from the carbon cathodes in deoxygenated pH10 buffer and determined spectrophotometrically by bleaching a Cu (II) ammine complex. Lithium was determined by atomic absorption.

C. Cyclic Voltammetry

Cyclic voltammetry and coulometry were done with a PAR 173 Potentiostat and a PAR 175 Function Generator with cells of conventional design using a platinum wire or foil working and counter electrode with a silver wire or foil pseudoreference electrode. Normal pulse and differential pulse polarography (npp and dpp) were performed with a PAR 364 polarograph. Data collection was by conventional X-Y and stripchart recorders. Solvents were Burdick and Jackson "Distilled in Glass" acetonitrile and methylene chloride, used without further purification. The supporting electrolyte salt was tetrabutylammonium perchlorate ($(\text{C}_4\text{H}_9)_4\text{ClO}_4$). (Southwestern Analytical, Electrometric Grade) kept at 70°C in a ventilated oven. Solutions were routinely degassed with Ar before use.

III. Results and Discussion

A. D Cells

We prepared six D cells with reference electrodes to a standard design of electrode length, height and thickness and filled them with 0.25M $\text{Li}_2\text{B}_{10}\text{Cl}_{10}$ in SO_2 , giving weights as shown in Table 1. The normally water white SO_2 forms a yellow solution with $\text{Li}_2\text{B}_{10}\text{Cl}_{10}$. The presence of this yellow color has been cited as evidence for extensive anion solvation by the SO_2 (9). Cell 680-2 was rejected for low electrolyte weight and not used. The electrolyte salt used in this batch of electrolyte had 0.31% H_2O by KF reagent. Later batches of $\text{Li}_2\text{B}_{10}\text{Cl}_{10}$ were more carefully dried and contained around 250 ppm H_2O . One of these cells was completely discharged at 0.5A at ambient temperature, giving 8.7 A.hr to a 2V cutoff. Two cells were selected for chemical analysis after discharge and recharge, respectively. One cell was discharged for 16 hours at 0.3A giving the voltage profile shown in Figure 4. The second cell was discharged for 16 hours at 0.3A and then recharged at 0.3A for 16 hours as shown in Figures 5 and 6. A relay was used to end the recharge at a cell voltage of 3.6V.

B. Chemical Analysis

The D cells were removed from the test area, punctured in the sealed bomb and the SO_2 removed and collected on the vacuum line. Both cell SO_2 extracts were dissolved in CH_3CN and titrated with aqueous base to a bromocresol green endpoint. Results of chemical analysis are shown in Table 2. The cell which had not been recharged contained 21.9 g SO_2 or 9.17 A.hr. The 9.17 A.hr SO_2 remaining, plus the 4.8 A.hr used in the discharge gave 13.97 A.hr of SO_2 , or 96% of the cathode active material originally found in the cell (680-6). The recharged cell contained 32.1 g SO_2 or 13.44 A.hr. This is a 96% recovery of active material from cell 680-4. The analysis of the discharged cell shows that SO_2 is consumed in the cell discharge reaction through a one electron reduction of SO_2 while the SO_2 analysis of the recharged cell shows the quantitative

regeneration of SO_2 by oxidation of the discharge product.

The punctured cells were transferred in the sealed bomb to an argon filled glove box where they were opened with a tubing cutter and the electrode stacks removed. The discharged cell (680-6) had a rather stiff cathode which was somewhat brittle with a bright, flexible anode foil. The cathode was leached of dithionite in deaerated pH10 buffer and the leachate analyzed for $\text{S}_2\text{O}_4^{2-}$. Cell 680-6 had 11.8 g or 0.092 eq. $\text{S}_2\text{O}_4^{2-}$, which is 4.94 A.hr or 102% of the calculated depth of discharge. The specific analysis for lithium dithionite confirms $\text{Li}_2\text{S}_2\text{O}_4$ as the SO_2 discharge product and demonstrates the one electron reduction of SO_2 .

The recharged cell (680-4) had a flexible cathode while the re-active anode surface tarnished even in the glove box atmosphere. There was some evidence of especially heavy Li deposition on electrode edges, but it was clear that both discharge and recharge had taken place throughout the length and height of the the electrode. The cathode of this cell contained <10 mg of $\text{Li}_2\text{S}_2\text{O}_4$, indicating that 100% of the dithionite deposited in the cathode was oxidized to SO_2 during the recharge.

The D cell is not a very satisfactory package for chemical analysis because of the difficulties in sampling and disassembling the hermetic package. The data we gained from these cells show, within the limits of the analyses, that Li/ SO_2 cells show a quantitative reduction and reoxidation of SO_2 during discharge and recharge.

C. Cyclic Voltammetry

During early experiments evaluating the $\text{Li}_2\text{B}_{10}\text{Cl}_{10} - \text{SO}_2$ electrolyte, it was observed that cells showed a voltage plateau near 3.6V vs. Li. In these experiments intensely colored blue material was formed at the porous carbon at highly oxidizing potentials. We suspected that this blue material was an oxidation product of $\text{B}_{10}\text{Cl}_{10}^{2-}$, analogous to the oxidation products of $\text{B}_9\text{Cl}_9^{2-}$ (10). Since the reaction occurred at highly oxidizing potentials, we selected acetonitrile/N $(\text{bu})_4\text{ClO}_4$ as a suitable electrolyte medium for this reaction.

A background voltammogram of the voltage region from 0 to +2V at 100 mv/sec for the $\text{CH}_3\text{CN}/\text{N}(\text{bu})_4\text{ClO}_4$ electrolyte is shown in Figure 7. A small amount of $\text{Li}_2\text{B}_{10}\text{Cl}_{10}$ was then added, the solution again degassed with Ar, and the same voltage region swept again. As shown in Figure 8; a reversible wave near +1.5V, or 3.6V vs. Li was revealed. This wave can easily be assigned to the oxidation of $\text{B}_{10}\text{Cl}_{10}^{2-}$ anion. A more sensitive cyclic voltammogram for 37.3 mg $\text{Li}_2\text{B}_{10}\text{Cl}_{10}$ in 60 ml $\text{CH}_3\text{CN}/\text{N}(\text{bu})_4\text{ClO}_4$ (1.3×10^{-3} M) at 200 mV/sec is shown in Figure 9. In this voltammogram the reversible wave near 3.6V from Li is clearly defined as well as a quasireversible wave near +1.1V which is probably due to an impurity in $\text{Li}_2\text{B}_{10}\text{Cl}_{10}$. The impurity is probably LiCl (11, 12). When the voltage limit on the scan was increased to the very high potentials of 2.5 - 2.8V vs Ag or 4.6 - 4.9V vs Li, an irreversible wave was observed very near to background. There was no indication of any chemical reversibility in this wave.

To get a better idea of the stability of the oxidized form of $\text{Li}_2\text{B}_{10}\text{Cl}_{10}$, we ran cyclic voltammograms at sweep rates from 50 to 500 mV/sec as shown in Figure 10. The result of these investigations are shown in Table 3. The ΔE_p values indicate a reversible one electron transfer slightly broadened by uncompensated solution resistance at higher sweep rates. From the increase in i_{pa}/i_{pc} from 0.86 to 0.95 with increase in V , there is some evidence for a comparatively slow reaction in solution. The oxidation of $\text{Li}_2\text{B}_{10}\text{Cl}_{10}$ is diffusion controlled at a platinum electrode in acetonitrile as demonstrated by the linear plot of i_{pa} vs. \sqrt{V} shown in Figure 11.

We hoped that the subsequent chemical reaction observed in the oxidation of $\text{Li}_2\text{B}_{10}\text{Cl}_{10}$ was due to reaction with an impurity, either in the acetonitrile or supporting electrolyte salt. We decided to investigate the one electron stoichiometry of the $\text{Li}_2\text{B}_{10}\text{Cl}_{10}$ oxidation by a constant potential electrolysis. We oxidized $\text{Li}_2\text{B}_{10}\text{Cl}_{10}$ at a Pt foil electrode at 1.62V vs the Ag

pseudoreference electrode in a stirred solution. After a few minutes of electrolysis, purple material streamed from the Pt electrode. The current declined during electrolysis to a very low level. For the first experiment n was 1.5, a second experiment gave $n = 2.5$. In a third experiment we pre-oxidized the acetonitrile solvent to destroy oxidizable impurities before electrolysis. The third experiment gave $n = 1.0$ to confirm the one electron stoichiometry of the reaction. We characterized this purple, soluble, oxidation product by the UV-VIS spectrum shown in Figure 12. The material is characterized by two absorption bands in the visible at 569 and 419 nm with more intense bands in the UV. The reaction of the purple species in solution was shown by a decline in absorbance of $\sim 10\%$ over a one hour period.

We explored the oxidation chemistry of $\text{Li}_2\text{B}_{10}\text{Cl}_{10}$ further in a series of experiments. We tried cyclic voltammetry on $\text{Li}_2\text{B}_{10}\text{Cl}_{10}$ in $\text{CH}_2\text{Cl}_2/\text{N}(\text{bu})_4\text{ClO}_4$, but found that $\text{Li}_2\text{B}_{10}\text{Cl}_{10}$ was insoluble. We then prepared $(\text{N}(\text{pr})_4)_2\text{B}_{10}\text{Cl}_{10}$ by metathesis with $\text{N}(\text{pr})_4\text{Br}$. The salt $(\text{N}(\text{pr})_4)_2\text{B}_{10}\text{Cl}_{10}$ was soluble in CH_2Cl_2 and was characterized by cyclic voltammetry. In CH_2Cl_2 the $\text{B}_{10}\text{Cl}_{10}^{2-}$ anion behaved similarly to CH_3CN except that the second, irreversible wave could not be observed because of the limited potential range of CH_2Cl_2 .

We also explored chemical methods for generating the oxidation products of $\text{Li}_2\text{B}_{10}\text{Cl}_{10}$. In both CH_3CN and aqueous solution the purple species is formed by addition of Ce^{IV} to $\text{Li}_2\text{B}_{10}\text{Cl}_{10}$. The purple product was concentrated on filter paper and dissolved in CH_2Cl_2 . This product had absorption bands at 573 and 421 nm, only slightly shifted from CH_3CN . We tried to isolate the purple product $\text{LiB}_{10}\text{Cl}_{10}$ by evaporating CH_2Cl_2 , but were unsuccessful.

The second oxidation product in $\text{Li}_2\text{B}_{10}\text{Cl}_{10}$ occurs at a very high potential $4.5 + \text{V}$ vs. Li and cannot easily be studied by cyclic voltammetry. We decided to use normal pulse (npp) and differential pulse (dpp) polarography to minimize the effect of the background current and allow the oxidation to be clearly resolved. The polarograms were taken at a platinum electrode since mercury was oxidized at the high anodic potentials required. The normal pulse polarogram for $\text{Li}_2\text{B}_{10}\text{Cl}_{10}$ in $\text{CH}_3\text{CN}/\text{N}(\text{bu})_4\text{ClO}_4$ is shown in Figure 13. As

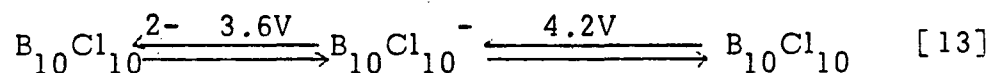
shown in the figure, the slight Cl^- impurity shows up as a polarographic maxima while both the first and second waves are quite clearly defined. In the npp the second oxidation wave is still in the rising baseline, but appears to be higher than the first oxidation. The picture presented by the oxidation product of $\text{Li}_2\text{B}_{10}\text{Cl}_{10}$ is clarified by use of differential polarography (dpp) as shown in Figure 14. The dpp reveals the initial impurity wave followed by two successive symmetric peaks, indicative of reversible processes. The peak heights, 50 and 51 μA respectively show that both are one electron oxidations. There were no further oxidation processes out to +2.8V. The apparent half wave potentials of 1.43 and 2.16V are comparable to potentials estimated by cyclic voltammetry.

In the electrochemical and chemical experiments we demonstrated that the supporting salt $\text{Li}_2\text{B}_{10}\text{Cl}_{10}$ undergoes two reversible diffusion controlled one electron oxidations. The product of the first of these oxidations, $\text{LiB}_{10}\text{Cl}_{10}$ was characterized by its UV-VIS spectrum with strong absorbances at 569 and 419 nm in CH_3CN solution.

IV. Conclusion

As required in the work statement for this contract, we developed a hermetically sealed prototype rechargeable D cell with reference electrode by modifying a conventional primary Li/SO₂ cell to incorporate a second G/M seal feedthrough for the Li reference. This allowed us to maintain the standard filling and venting arrangements for Li/SO₂ cells intact. We filled these cells with the experimental Li₂B₁₀Cl₁₀/SO₂ all inorganic electrolyte and roughly characterized primary cell performance. We used selected D cells with reference electrode for chemical analysis and demonstrated the quantitative formation of Li₂S₂O₄ in the cathode during cell discharge as well as the quantitative oxidation of Li₂S₂O₄ to SO₂ during cell recharge. This established the rechargeable nature of the positive electrode reactions.

We investigated the chemistry of the supporting salt Li₂B₁₀Cl₁₀ during recharge and found that Li₂B₁₀Cl₁₀ is oxidized in two one electron processes at high voltage [13]. Only the first oxidation process at 3.6V



vs. Li is reached during cell recharge. The initial oxidation product is a persistent material characterized by absorbances at 569 and 419 nm.

We demonstrated both the inherently high energy density and rechargeable nature of the Li/SO₂ system using the exotic electrolyte salt Li₂B₁₀Cl₁₀.

This report includes information concerning proprietary property of Duracell International Inc., and the presence of any information herein should not be interpreted to mean that the technology described was or was not developed pursuant to any government financed project.

This work was supported by the Assistant Secretary of Conservation and Renewable Energy, Office of Advanced Conservation Technology, Division of Electrochemical Systems Research of the U.S. Department of Energy under Contract Number W-7405-ENG-48.

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Table 1

Weights and Cell Capacities of Prototype Rechargeable D Cells

| Cell # | Empty Wt. (g) | Full Wt. (g) | Electrolyte Wt. (g) | SO ₂ Capacity (A.hr) |
|--------|------------------|-----------------|------------------------|---------------------------------|
| 680-1 | 48.7 | 89.6 | 40.9 | 15.0 |
| 680-2 | 50.2 | 73.2 | 23.2 | 8.5 |
| 680-3 | 48.5 | 87.1 | 38.6 | 14.9 |
| 680-4 | 50.1 | 86.4 | 36.3 | 14.0 |
| 680-5 | 48.8 | 87.7 | 38.9 | 15.0 |
| 680-6 | 48.9 | 86.6 | 37.7 | 14.5 |

TABLE 2

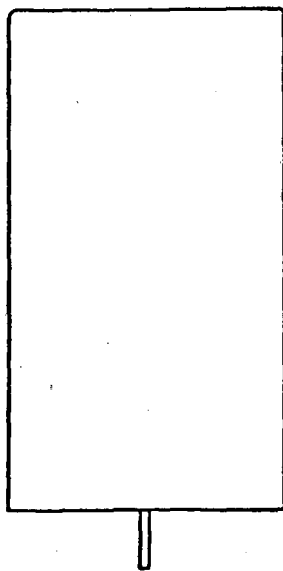
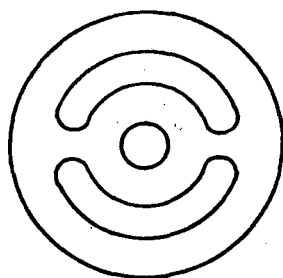
Chemical Analysis of D Cells

| Cell | Condition | SO ₂ | SO ₂ recovery | S ₂ O ₄ | S ₂ O ₄ recovery |
|-------|-----------------------|-----------------|-----------------------------|-------------------------------|---|
| 680-6 | Discharge 4.8 A.hr | 21.9g | 96% | 11.8g | 102% |
| 680-4 | Discharge 4.8 A.hr | 32.1g | 96% | <0.01g | - |
| | Recharge 4.8 A.hr | | | | |

TABLE 3

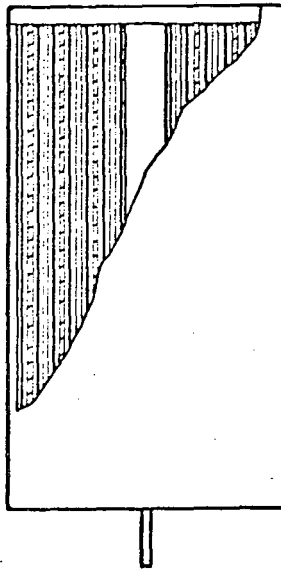
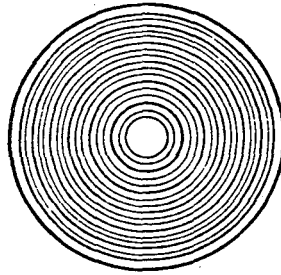
Current/Potential Data for Cyclic Voltammograms of $\text{Li}_2\text{B}_{10}\text{Cl}_{10}$

| Sweep Rate (mV/sec) | i_{pa} (μA) | i_{pc} (μA) | i_{pa}/i_{pc} | E_{ipa} (V) | E_{ipc} (V) | E_p |
|---------------------|----------------------------|----------------------------|-----------------|---------------|---------------|-------|
| 50 | 5.35 | 4.60 | 0.86 | 1.65 | 1.59 | 60 mV |
| 100 | 6.75 | 6.25 | 0.92 | 1.66 | 1.59 | 70 |
| 200 | 10.5 | 10 | 0.95 | 1.65 | 1.57 | 80 |
| 500 | 13.5 | 12.8 | 0.95 | 1.66 | 1.58 | 80 |



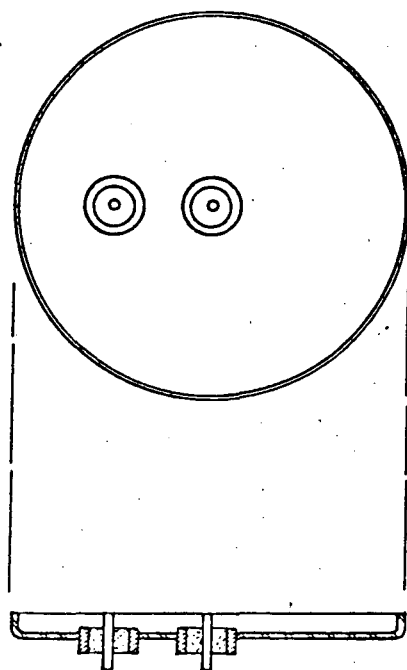
SO₂ CELL CAN

Figure 1. CRS can used for SO₂ cells showing the convoluted vent and fill eyelet



WOUND CELL

Figure 2. Side view of a standard LC26S SO₂ cell..



REFERENCE CELL TOP

Figure 3. Cell top incorporating referenced electrode feedthrough

Figure 4. Voltage vs. time behavior for a rechargeable D cell at 0.3A discharge; both anode and cathode potential vs. the lithium reference are shown.

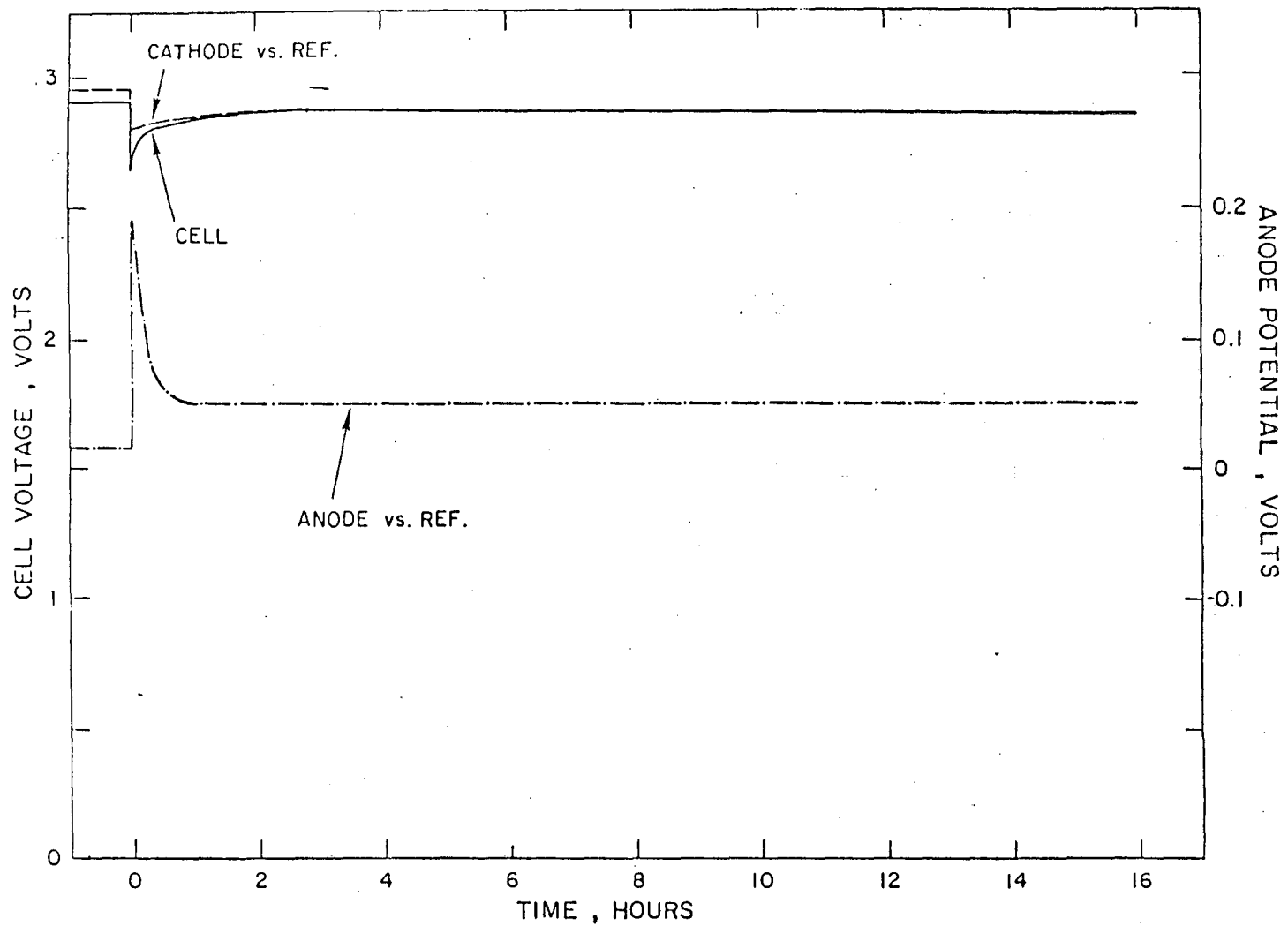


Figure 5. Voltage vs. time behavior for a rechargeable D cell at 0.3A discharge; both anode and cathode potential vs. the lithium reference are shown

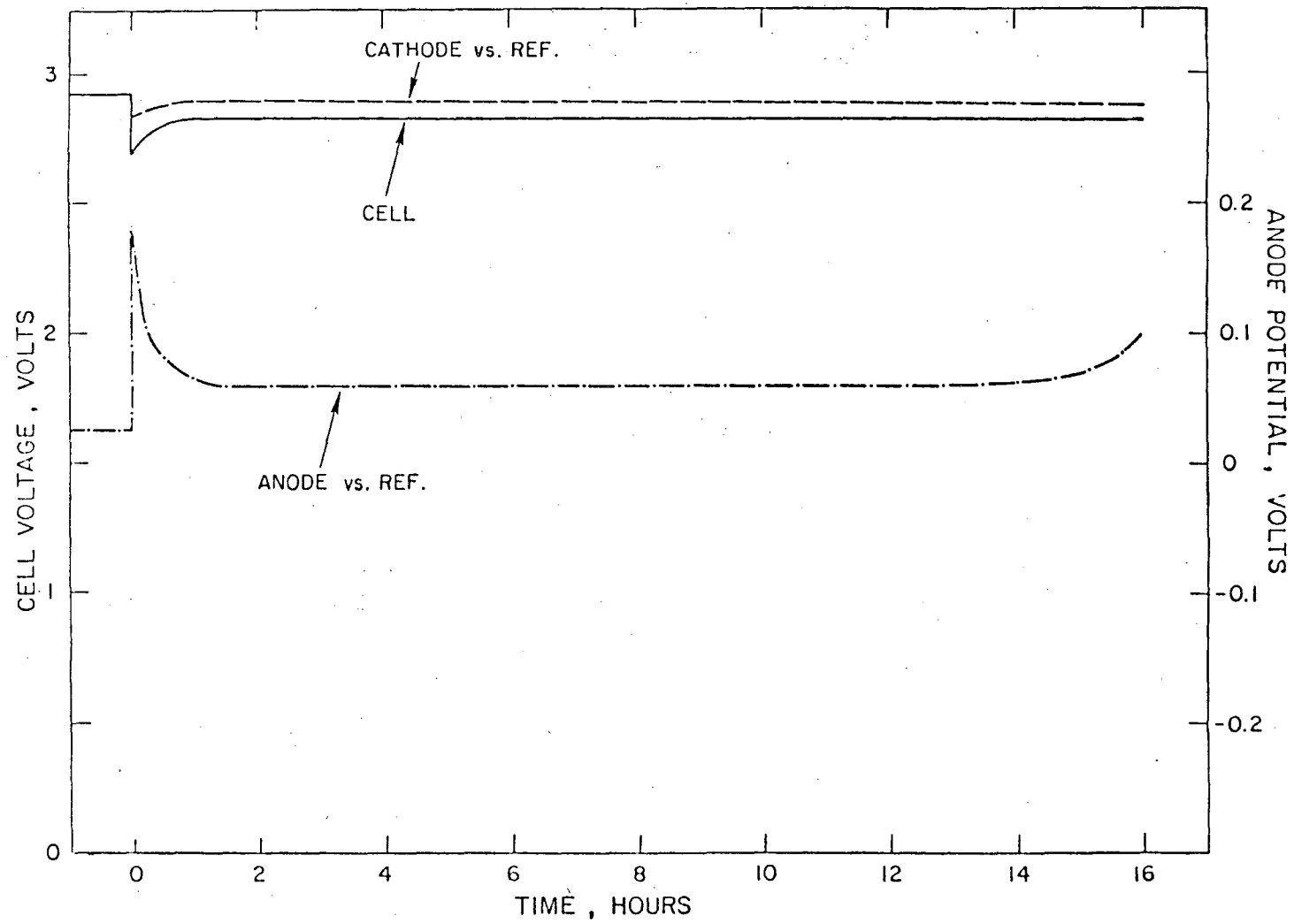
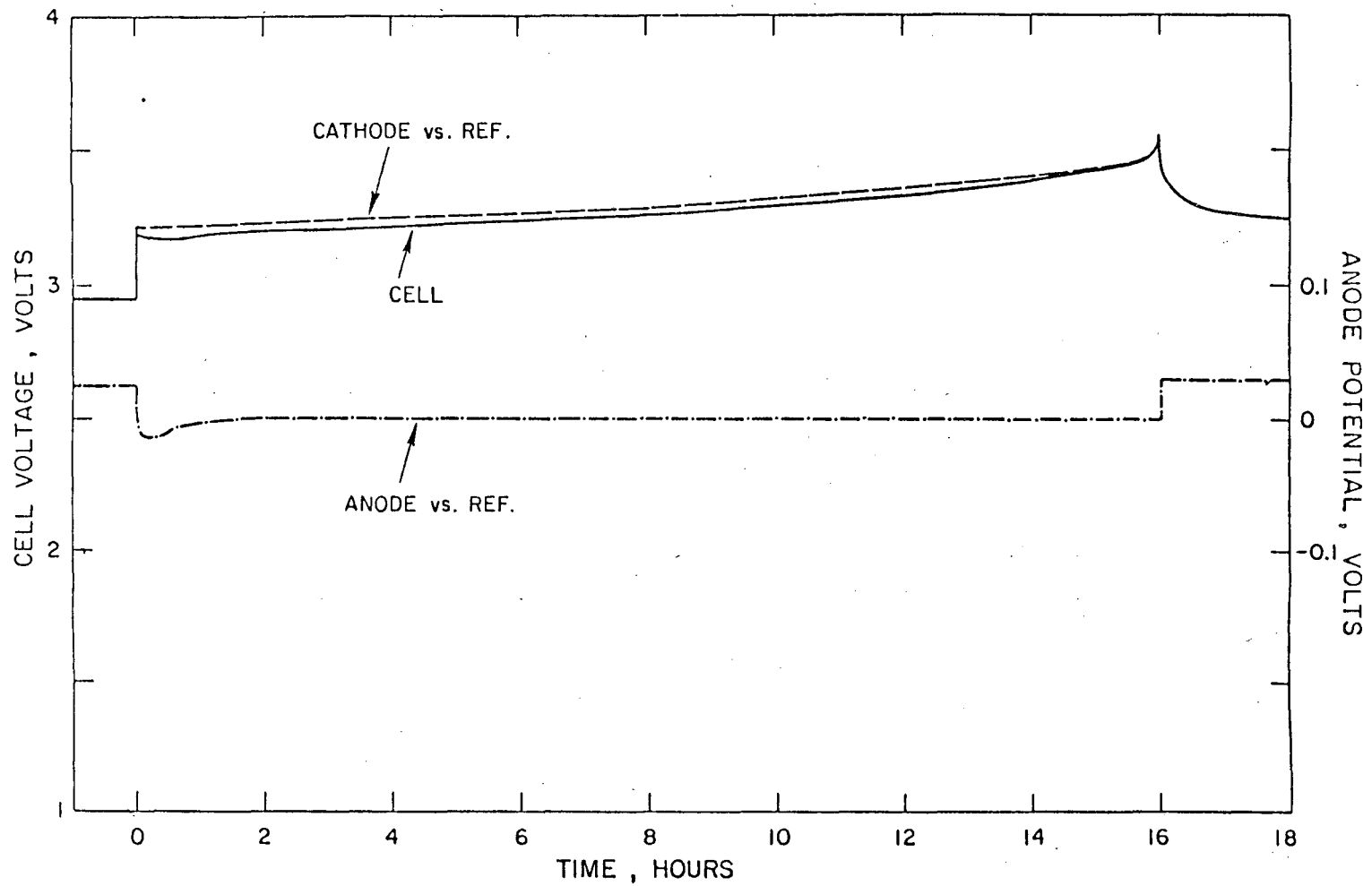


Figure 6.. Voltage vs. time behavior for the cell in Fig. 5 during recharge at 0.3A; both anode and cathode potential vs. the lithium reference are shown.



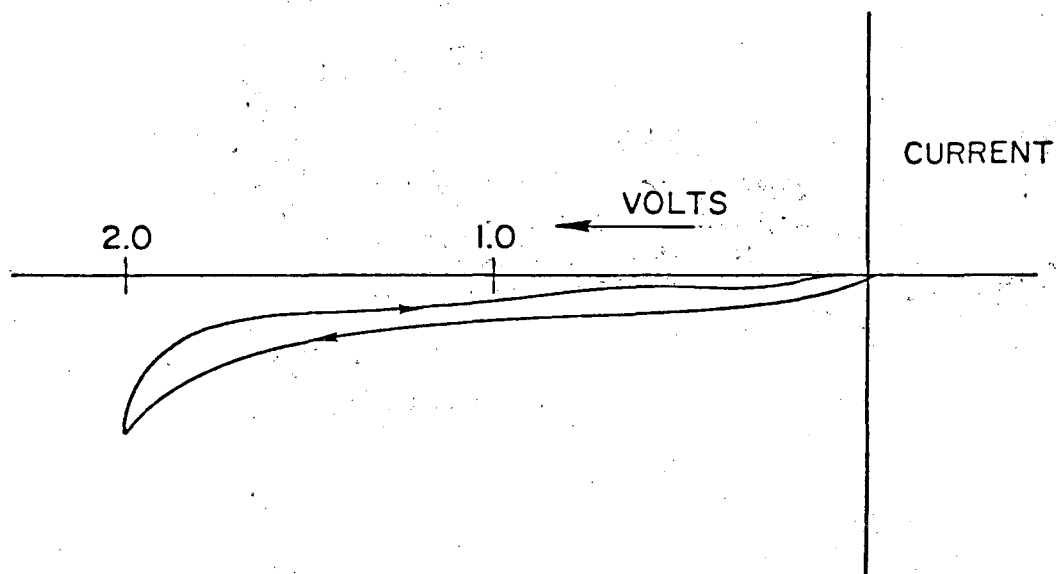


Figure 7. Cyclic voltammogram for Pt in $\text{CH}_3\text{CN}/\text{N}(\text{bu})_4\text{ClO}_4$ at 100 mV/sec.

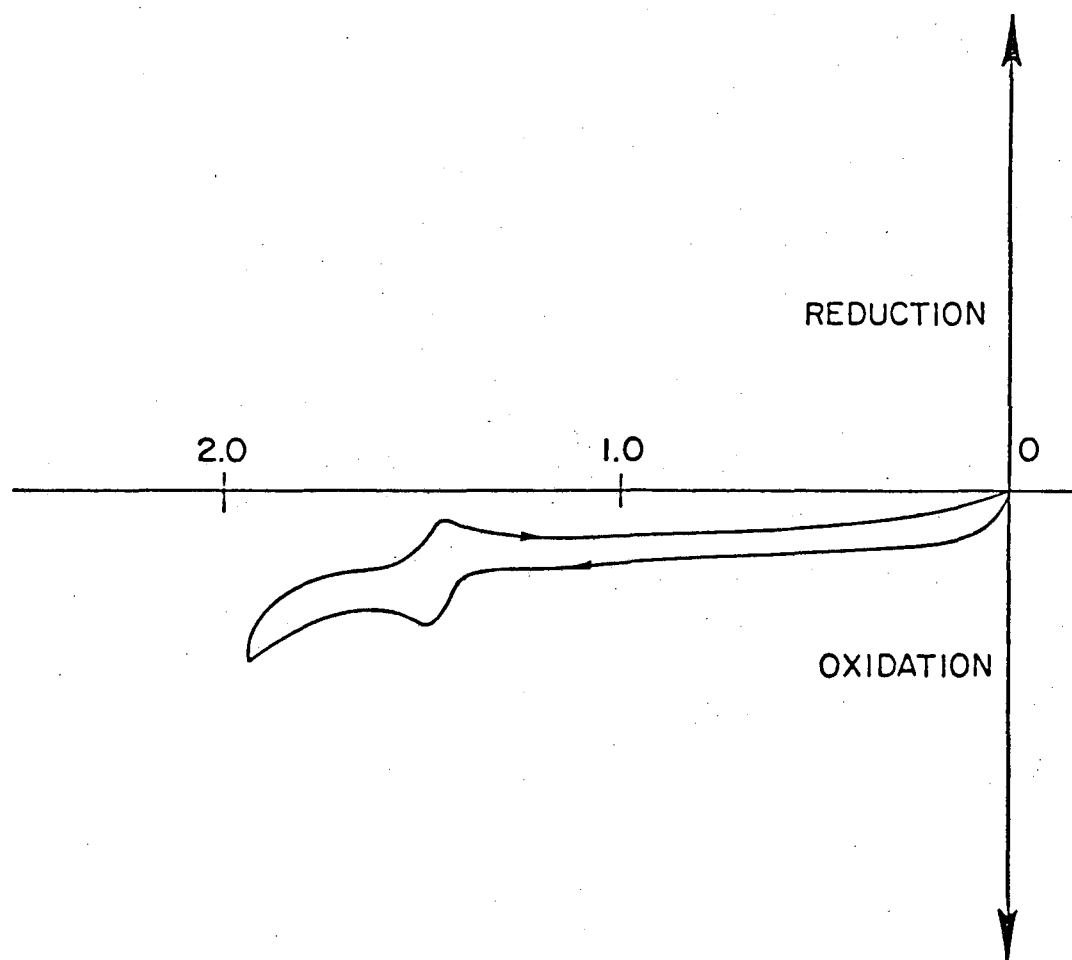


Figure 8. Cyclic voltammogram for Pt in $\text{Li}_2\text{B}_{10}\text{Cl}_{10}$ solution at 100 mV/sec.

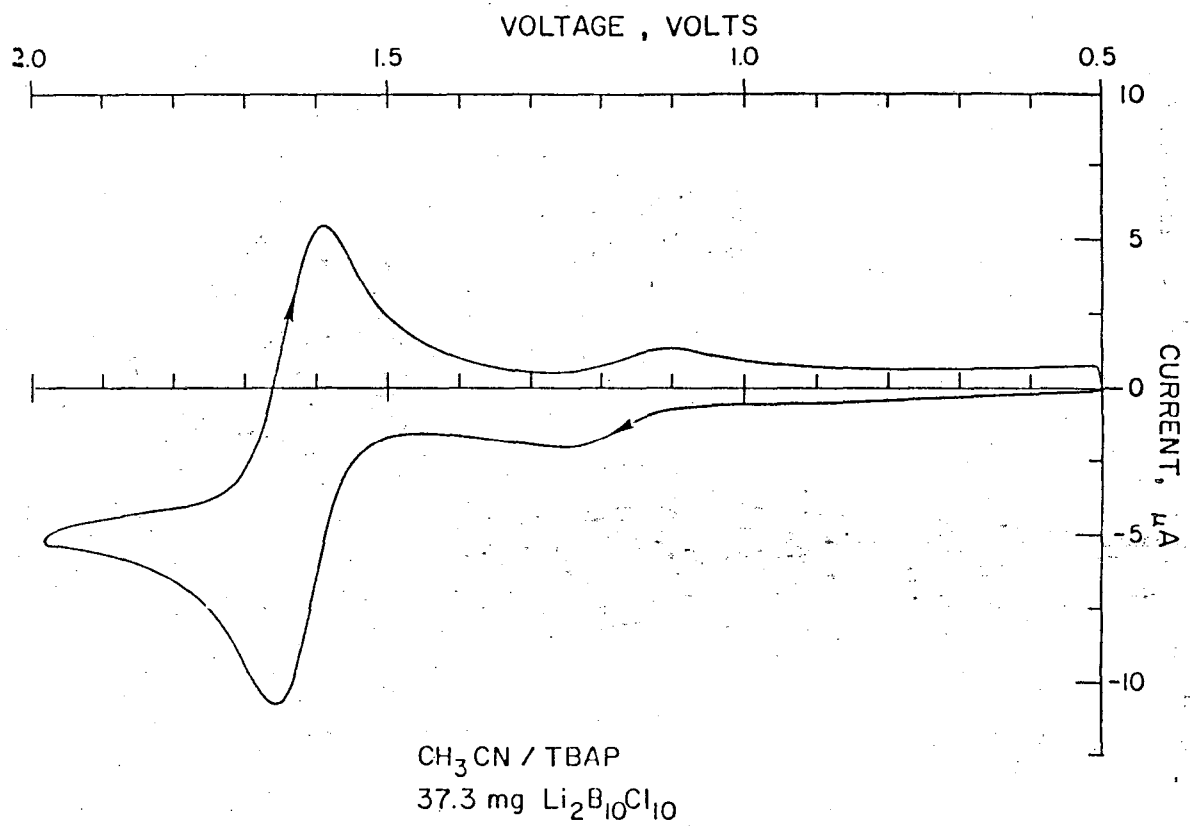


Figure 9. Cyclic voltammograms of Li₂B₁₀Cl₁₀ in CH₃CN/N(bu)₄ClO₄ at 200 mV/sec.

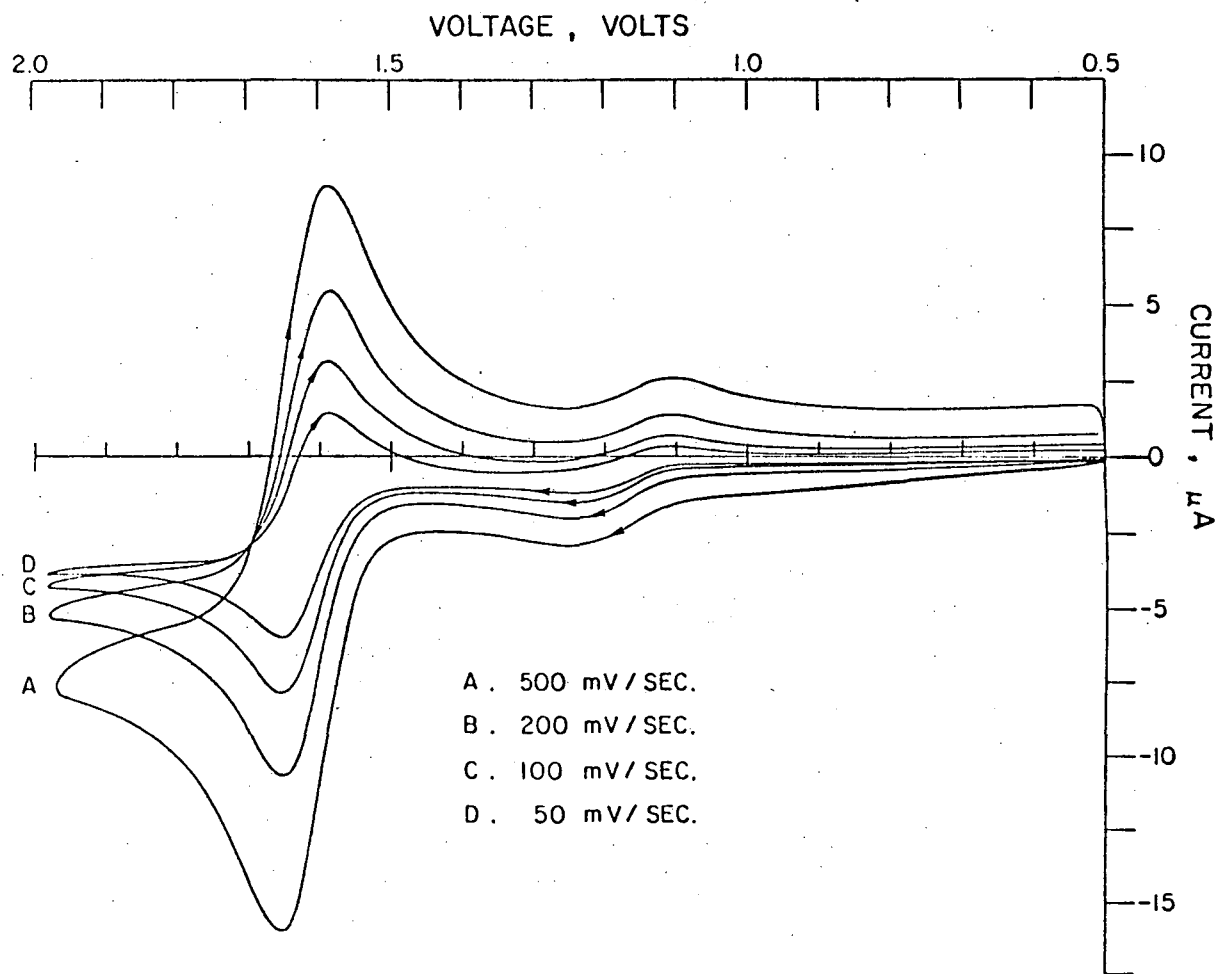


Figure 10. Cyclic voltammograms of $\text{Li}_2\text{B}_{10}\text{Cl}_{10}$ in $\text{CH}_3\text{CN}/\text{N}(\text{bu})_4\text{ClO}_4$
 a) 50 mV/sec. b) 100 mV/sec. c) 200 mV/sec. d) 500 mV/sec.

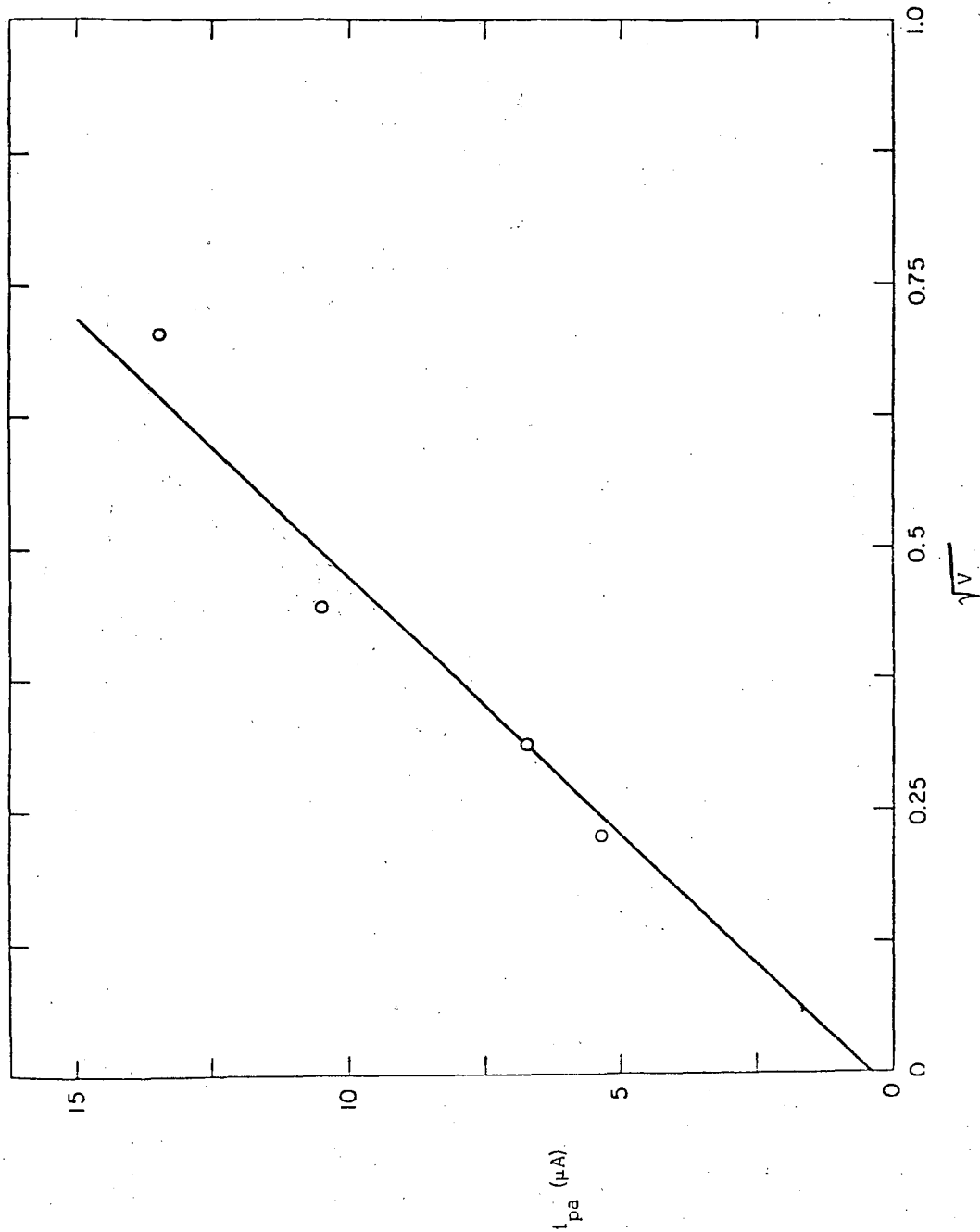


Figure 11. Plot of peak anodic current I_{pa} vs. \sqrt{V} , root of sweep rate for $Li_2B_{10}Cl_{10}$ in acetonitrile showing diffusion control.

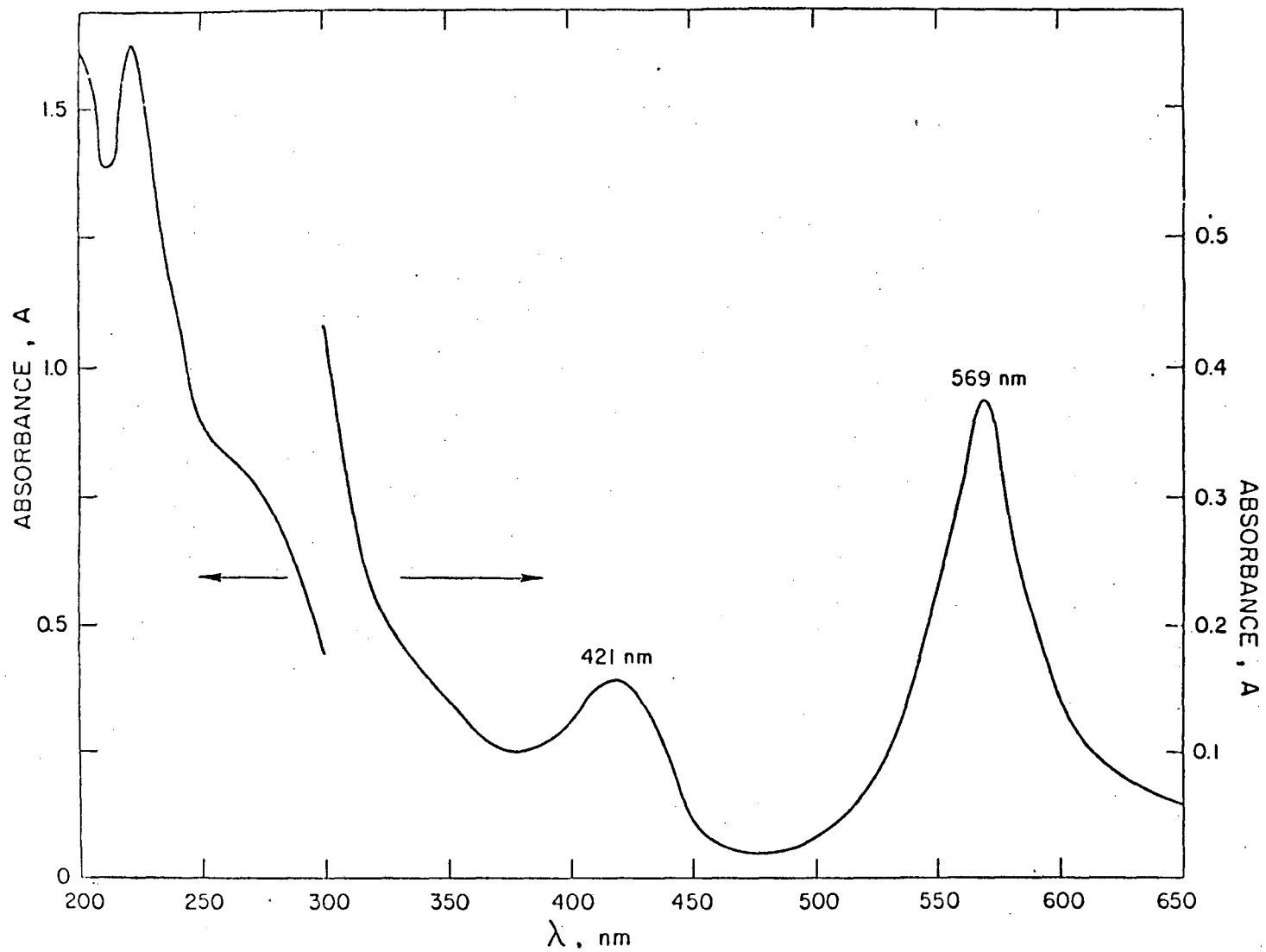


Figure 12. UV-VIS spectrum of $\text{Li}_2\text{B}_{10}\text{Cl}_{10}$ oxidation product, $\text{B}_{10}\text{Cl}_{10}$ in acetonitrile solution

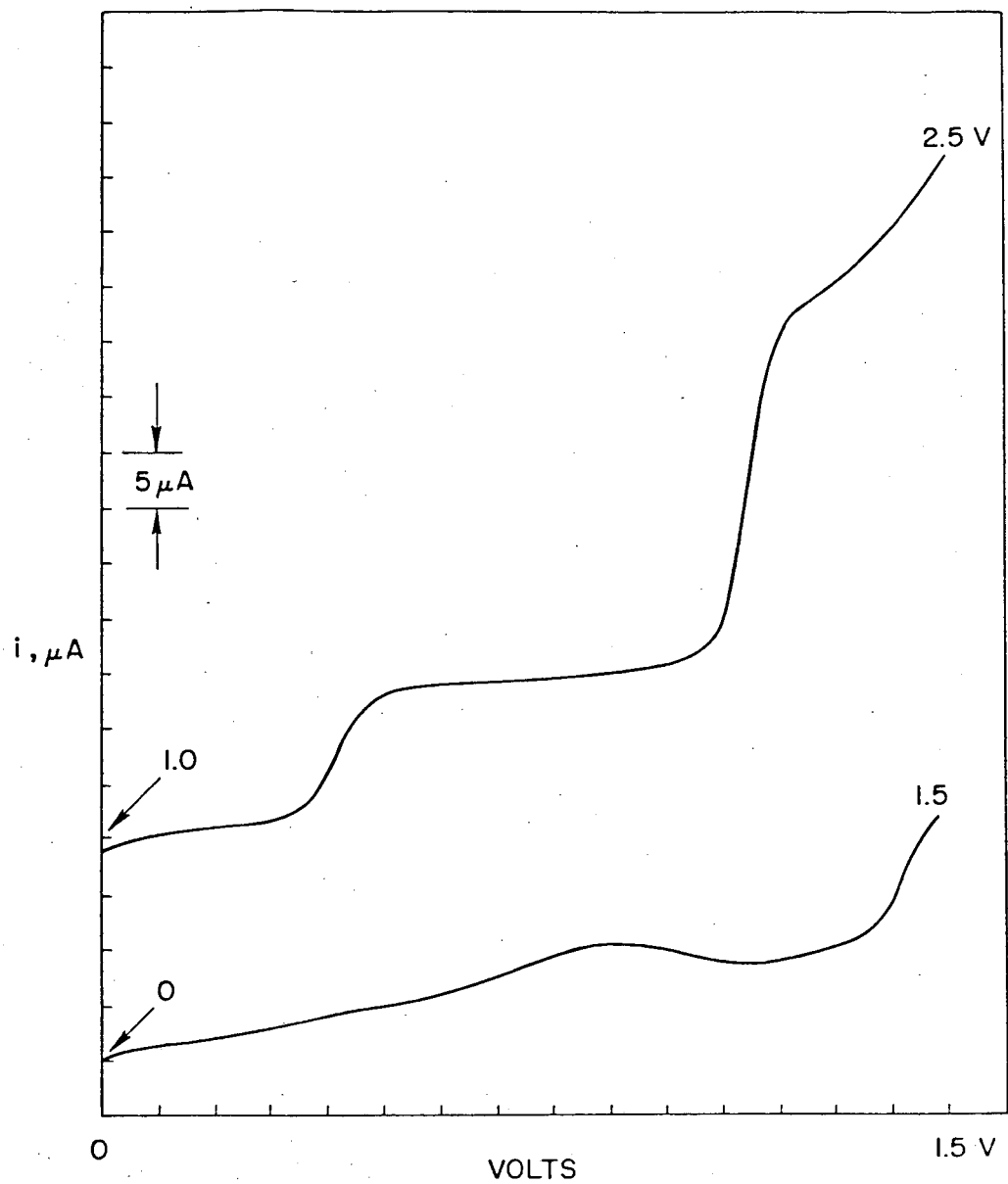


Figure 13. Normal pulse polarogram for $\text{Li}_2\text{B}_{10}\text{Cl}_{10}$ in $\text{CH}_3\text{CN}/\text{TBAClO}_4$ showing the two oxidation waves

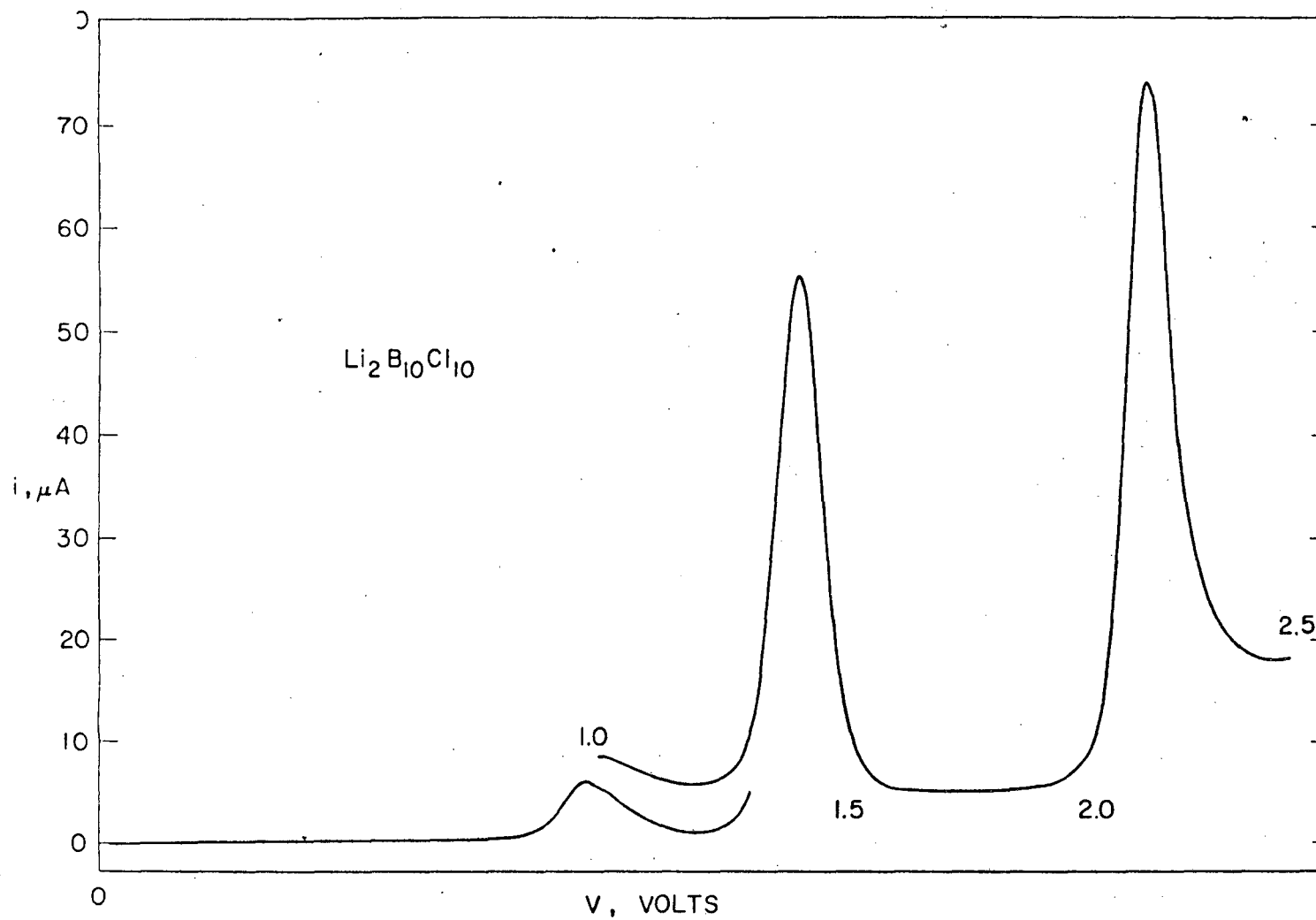


Figure 14. Differential pulse polarogram for $\text{Li}_2\text{B}_{10}\text{Cl}_{10}$ in $\text{CH}_3\text{CN}/\text{TBAClO}_4$ showing the two equal peaks for the oxidation of $\text{B}_{10}\text{Cl}_{10}^{2-}$

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