THE SULFUR DIOXIDE, OXYGEN,
SULFURIC ACID CELL

A thesis submitted in partial fulfillment of the requirements for the Degree of Master of Applied Science in the Department of Chemistry

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Accepted on behalf of the Department of Chemistry
ACKNOWLEDGEMENT.

I wish to acknowledge the helpful suggestions and assistance of Dr. W.F. Seyer, under whose direction the present work was carried out. I would also like to thank Dr. B.P. Sutherland of the Consolidated Mining and Smelting Company for supplying some much needed equipment which could not be obtained from the regular sources.
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KEY TO SYMBOLS

$\Delta F =$ free energy change, calories

$f =$ Faraday  96,500 coulombs

$M =$ mol fraction

$N =$ normality

$n =$ valence

$T =$ temperature, degrees Centigrade

$t =$ time, hours

$X =$ per cent acid

$x =$ normality of electrolyte at time "t"

$J =$ conversion factor: 1 cal. = 4.183 joules

$1 \text{ joule} = 9.87 \text{ cc atm.}$

..................
Production of sulfuric acid by present standard methods leaves little to be desired, especially when elemental sulfur, sulfide ores and some by-product gases are used. However, the utilization of the free energy of sulfur dioxide in the form of electrical energy for the production of sulfuric acid offers interesting possibilities. The cell reaction of such a process is given by:

$$\text{SO}_2 + \frac{3}{4}\text{O}_2 + \text{H}_2\text{O} = \text{H}_2\text{SO}_4$$

The free energy change $\Delta F$ involved in this reaction is made up of the free energy of formation of sulfuric acid plus the free energy of dilution of sulfuric acid. The $\Delta F$ change in forming pure sulfuric acid is 49,100 calories per mole, and for a more dilute acid it
would be even greater. Any reaction which is accompanied by a large negative free energy change is spontaneous, hence from the above free energy change we realize it would be quite possible to produce sulfuric acid in an electrolytic cell and at the same time draw off electrical energy.

By the use of sulfur dioxide in the method indicated, a number of objectives could be achieved. These include production of cheap dilute acid for industrial uses such as pickling steel or leaching copper or other ores; production of concentrated acid where economic conditions warrant, and abatement of atmospheric pollution.

OUTLINE OF THE PROBLEM

In the following experiments it was proposed to determine the electromotive force of a cell similar to

Pt, SO₂ : H₂SO₄ : O₂, Pt

at various concentrations of sulfuric acid. The above cell was studied in 1945 by T.C. Assaly, and from his work it was shown that four hydrates of sulfuric acid were stable in solution. The writer proposed to study the same cell with different electrode material to test the validity of Assaly's results.

The overall efficiency of the platinum-electrode cell was below 50 per cent over the greater part of the acid range, so a preliminary study was made to find an electrode
combination which would increase the cell efficiency. Lead electrodes were used first but were soon discarded since an oxide film formed immediately on both the anode and cathode producing a very low voltage. Tests were then carried out using tantalum metal at both electrodes and a constant voltage was obtained after about 24 hours. The efficiency of this cell was below that of Assaly's but it appeared that a combination of platinum and tantalum might give the desired results. A cell made up of a platinum anode and a tantalum cathode was then tested but this still was not satisfactory. The electrode combination was then reversed, and with tantalum as anode and platinum as cathode a great increase in efficiency was obtained. Fig. 1 shows these results graphically.

As a result of these experiments it was decided to determine the actual electromotive force of the cell:

\[ \text{Ta, SO}_2\text{, H}_2\text{SO}_4\text{, O}_2\text{, Pt} \]

at various concentrations of sulfuric acid.

**HISTORICAL**

In 1916 Messrs. M. De Kay Thompson and N. J. Thompson investigated the current efficiencies of the oxidation of sulfurous acid and they showed that this oxidation takes place with high current efficiencies even
THE VARIATION OF CELL VOLTAGE WITH TIME

USING DIFFERENT ELECTRODE MATERIAL

(0.58 N Acid)

VOLTAGE

(1) tantalum anode: platinum cathode
(2) platinum anode: platinum cathode
(3) tantalum anode: tantalum cathode
(4) lead anode: lead cathode

TIME (hours)

Fig. 1
in strong sulfuric acid solutions. Also for a given sulfuric acid concentration the current efficiency decreases with increasing current density.

Messrs. M. De Kay Thompson and A. P. Sullivan\textsuperscript{2} studied the depolarization effects of sulfur dioxide in an electrolytic cell using a platinum anode. They found that it was possible to decrease the polarization at the anode and thus decrease the equilibrium voltage of the cell by the addition of sulfur dioxide. It appeared from their experiments that the polarization was due to the accumulation of free oxygen at the anode. On addition of sulfur dioxide to the anolyte a reaction occurred between this gas and the free oxygen reducing the polarization effects.

Mr. T. C. Assaly\textsuperscript{3} studied the actual electromotive force of the cell:

\[
\text{Pt, } \text{SO}_2: \text{H}_2\text{SO}_4: \text{O}_2, \text{Pt}
\]

at various concentrations of sulfuric acid and at different temperatures. He found that in plotting cell voltage versus concentration of electrolyte that a step-wise curve was formed; each of the four steps corresponding to a hydrate of sulfuric acid. Other experimenters have shown evidence for the existence of three of these hydrates by freezing methods. The electrolytic method indicates that the sulfuric acid hydrates are stable at 25° C as well as at their freezing points.

The electromotive force of the above cell was shown to be dependent on the temperature of the electrolyte.
The voltage increased slightly with a decrease in temperature and approached a constant value at low temperatures.

**GENERAL THEORY**

**GAS ELECTRODES**

Gas electrodes commonly consist of a solid first-class conductor, in which a gas has been absorbed. Recent investigators have shown that the emf developed by so-called gas electrodes is a function of the first-class conductor in which the gas is contained as well as of the gas itself. In other words, gas electrodes are really gas-metal electrodes since metals are commonly used to absorb these gases. The potentials developed by such gas-metal electrodes are specific for the particular metal and gas, and are related to the absorbing power of the metal for the gas. The potentials are known to vary greatly with the nature of the electrolyte in which the electrodes are immersed. It has also been found that there is a somewhat different potential developed when the electrode is stationary than when it is in motion. This latter effect may be shown by moving the electrode or by keeping it stationary and moving the electrolyte. The electrolyte may be moved by stirring, causing the electrolyte to flow past the electrode, or by bubbling the gas through the electrolyte.

It has been observed that the amount of gas needed to bring about its effect on the potential of a metal is very
small indeed. In fact, the pressure or solubility of the gas has little or no effect on the potential. When a gas is being pumped out of the solution little effect is observed on the potential until the bubbling out of the gas causes a stirring effect.

HYDRATES IN AQUEOUS SOLUTION

The number of molecules of water in combination with one molecule of the dissolved substance, frequently increases from the most concentrated to the most dilute solutions, as with magnesium chloride, manganese chloride and copper chloride. With some substances the number of molecules of water held in combination by one molecule of the dissolved substance may pass through a well-defined maximum as the dilution is increased. In other cases, the number of molecules of water held in combination by one molecule of the dissolved substance may reach a maximum value as the dilution is increased; this maximum value may then remain practically constant with further increases in the dilution.

The question arises whether these hydrates are true chemical compounds or whether they represent some less stable form of combination. That they are unstable is shown by the ease with which they are broken down by heat. Most of the water can be driven off from the above solutions at a temperature only a little above 100°C. The more complex hydrates are, then, decomposed in solution at a comparatively low temperature and the water is given off in the form of vapour. In the light of these facts the hydrates can scarcely
be regarded as true chemical compounds. If however, we insist on calling them chemical compounds, we must admit that they represent a very low order of stability.

It is the general opinion that both molecules and ions combine with water, forming hydrates. It seems that the molecules are certainly capable of forming hydrates, because in very concentrated solutions where the ionization is very small, we often have considerable hydration. That ions are capable of combining with water in solution is shown by the magnitude of the hydration in many of the dilute solutions, where chiefly ions and only a few molecules are present.

**PROBABLE MECHANISM OF CELL REACTIONS**

**ANODE REACTIONS**

The sulfur dioxide first dissolves in the sulfuric acid electrolyte and then adsorbs on the surface of the electrode:

\[ \text{SO}_2 \text{(gas)} = \text{SO}_2 \text{(solution)} \quad (1) \]

The sulfur dioxide then combines with water as in equation (2)

\[ \text{SO}_2 + \text{H}_2\text{O} = \text{SO}_3^- + 2\text{H}^+ \quad (2) \]

**CATHODE REACTIONS**

Similarly, oxygen is dissolved in the electrolyte and then adsorbed on the surface of the electrode.

\[ \text{O}_2 \text{(gas)} = \text{O}_2 \text{(solution)} \quad (3) \]

The oxygen then combines with water as in equation (4)
\[ \frac{1}{2} \text{SO}_2 + \text{H}_2\text{O} + 2(-) = 20^- \]  

**OVERALL CELL REACTION**

The overall cell reaction is given by:

\[ \text{SO}_2 + \frac{1}{2} \text{O}_2 + \text{mH}_2\text{O} = \text{H}_2\text{SO}_4 \quad (N = X) \]  

The cell reaction can be divided into two parts as such:

\[ \text{SO}_2 + \frac{1}{2} \text{O}_2 + \text{H}_2\text{O} = \text{H}_2\text{SO}_4 \quad (6) \]

\[ 2 \text{H}^+ + \text{SO}_4^- + \text{mH}_2\text{O} = \text{H}_2\text{SO}_4 \quad (N = X) \quad (7) \]

\[ \text{SO}_2 + \frac{1}{2} \text{O}_2 + \text{mH}_2\text{O} = \text{H}_2\text{SO}_4 \quad (N = X) \quad (8) \]

Addition of equations (6) and (7) gives the overall cell reaction. Hence addition of the free energy of reactions of equations (6) and (7) will give the overall free energy change within the cell. The value of \( \Delta F \) for equation (6) is obtained from the difference of the free energy of formation of sulfuric acid and the sum of the free energies of formation of sulfur dioxide, water and oxygen.

\[ \Delta F_{25^\circ} \text{H}_2\text{SO}_4 = -176,000 \text{ calories} \]

\[ \Delta F_{25^\circ} \text{SO}_2 = -71,740 \text{ calories} \]

\[ \Delta F_{25^\circ} \text{H}_2\text{O} = -56,690 \text{ calories} \]

\[ \Delta F_{25^\circ} \text{O}_2 = \text{zero} \]

Therefore \( \Delta F \) change = \( -49,100 \text{ calories} \).

The number of accurate measurements from which the free energy of dilution of sulfuric acid may be calculated is limited. Tables are available on the free energy of dilution of the acid only for dilute solutions. Bronsted studied the free energy of dilution over a rather
wide range of concentrations of sulfuric acid but at
temperatures ranging only up to 9°C. Harned and Sturgis
and Lewis and Randall also studied the free energy of this
acid, but only over a small range of concentration.

A satisfactory table of the free energy of dilution
at 25°C for concentrations up to 0.2 mol fraction sulfuric
acid has been worked out by Randall and Cushman. Their
results which are the free energies $\Delta F_1$ of the reaction:

$$2H^+ + SO_4^{2-} + mH_2O = H_2SO_4 \quad (N-X) \quad (7)$$

are given in the third column of Table I, and the correspond-
ing values of mol fraction M and per cent acid X in the
first and second columns respectively.

<table>
<thead>
<tr>
<th>M</th>
<th>X(%)</th>
<th>$\Delta F_1$(cal.)</th>
<th>$\Delta F_2$(k cal.)</th>
<th>E(volts)</th>
</tr>
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<td>1.00</td>
<td>100.00</td>
<td>0</td>
<td>-49.1</td>
<td>1.065</td>
</tr>
<tr>
<td>.20</td>
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<td>5270</td>
<td>-43.8</td>
<td>.950</td>
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<td>.13</td>
<td>44.90</td>
<td>2915</td>
<td>-46.2</td>
<td>1.000</td>
</tr>
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<td>.10</td>
<td>37.70</td>
<td>1645</td>
<td>-47.5</td>
<td>1.030</td>
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<td>.08</td>
<td>31.60</td>
<td>702</td>
<td>-48.4</td>
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</tr>
<tr>
<td>.065</td>
<td>27.55</td>
<td>-85</td>
<td>-49.2</td>
<td>1.067</td>
</tr>
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<td>.05</td>
<td>22.28</td>
<td>-365</td>
<td>-50.0</td>
<td>1.084</td>
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<tr>
<td>.03</td>
<td>14.42</td>
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<td>.02</td>
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<td>-51.8</td>
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<td>.01</td>
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<td>-3702</td>
<td>-52.8</td>
<td>1.143</td>
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<tr>
<td>.002</td>
<td>1.08</td>
<td>-5613</td>
<td>-54.7</td>
<td>1.185</td>
</tr>
<tr>
<td>.0009</td>
<td>.487</td>
<td>-6547</td>
<td>-55.6</td>
<td>1.204</td>
</tr>
</tbody>
</table>

$\Delta F_2$ represents the overall cell reaction
$\Delta F_1$ represents the free energy of dilution
VARIATION IN THE THEORETICAL EMF OF THE CELL WITH CONCENTRATION OF ELECTROLYTE
By substitution in the formula:

$$E = \frac{-\Delta F}{n F}$$

$$\Delta F = \text{calories}$$

assuming a temperature of 25°C, the theoretical values of the emf that would be obtained in a completely reversible cell are obtained. The values are given in column 5 of Table I. In Fig. 2 the values of the emf given in Table I are plotted as ordinates against concentration of sulfuric acid as abscissa. A hypothetical curve was drawn from 0.2 M to 1 M sulfuric acid, the range of concentration where information on the free energy of dilution was unobtainable.

**DESIGN OF THE CELL**

A diagrammatic sketch of the cell is given in Fig. 3. It consisted of two stoppered fritted glass vessels, anode and cathode compartments, each containing sulfuric acid of the same concentration. Each of these vessels had a glass tube extending from it, about two inches above the fritted disc. These vessels were connected by means of a rubber tube containing a clay diaphragm.

Electrical contact was made by two mercury filled glass leads. One end of each lead was sealed to the stem of its particular electrode; the anode being a strip of tantalum and the cathode a platinized platinum gauze. The potential was measured by a potentiometer.

The whole cell was set in a constant temperature bath which maintained the temperature within 0.05°C of the desired value.
OPERATION OF THE CELL

PLATINIZING THE ELECTRODES

The platinum electrodes were coated with a layer of platinum black deposited electrolytically from a three per cent solution of chloroplatinic acid. The electrodes were first cleaned in warm chromic acid and then lowered into the solution. A 12-volt circuit was used and a commutator allowed the current to be reversed at desired intervals.

By means of a sliding resistance, the current was regulated so as to produce a moderate evolution of gas. The direction of the current was reversed every minute until a black and velvety coating appeared on the surface of the electrodes.

DEGASSING THE ELECTRODES

After each run the electrodes were removed from the cell, washed in distilled water, and degassed by electrolyzing them in 6 N sulfuric acid for approximately one hour. The direction of the current was reversed every ten minutes for the first fifty minutes, then every minute for the last ten minutes. The platinum electrodes were then boiled in dilute nitric acid for ten minutes to remove any poisoning agents. Finally all the electrodes were washed in boiling distilled water for fifteen minutes.

CHANGING THE CELL

The oxygen and sulfur dioxide inlet and outlet
connections to the cell were opened and the two half-cells were removed from the constant temperature bath. After flushing with warm distilled water, the half-cells were dried in a warm oven then washed with sulfuric acid electrolyte. The cell was then set up again and connections made to the sulfur dioxide and oxygen tanks. Seventy-five cubic centimeters of test sulfuric acid was then added to each half-cell and the electrodes placed in their respective half-cells.

The sulfur dioxide and oxygen were obtained from pressure tanks. The desired concentration of sulfuric acid for each run was obtained by diluting C.P. acid with distilled water. The exact strength of the acid was determined by titrating against a standard NaOH solution using phenolphthalein as indicator.

RESULTS

VARIATION OF CELL POTENTIAL WITH CONCENTRATION

Measurements were made of cell potential at various acid concentrations, and at 25°C, and these are shown in Table II. The first and second columns give the normality N and the per cent acid X and the third column the measured emf of the cell. The values of the emf given in Table II are the values obtained after the sulfur dioxide and oxygen had bubbled through the cell for nearly 24 hours and the cell potential approached a constant value.
<table>
<thead>
<tr>
<th>N</th>
<th>X(%)</th>
<th>E (volts)</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.98</td>
<td>4.69</td>
<td>0.8270</td>
</tr>
<tr>
<td>3.04</td>
<td>13.94</td>
<td>0.8120</td>
</tr>
<tr>
<td>4.80</td>
<td>21.23</td>
<td>0.8000</td>
</tr>
<tr>
<td>6.22</td>
<td>26.75</td>
<td>0.7990</td>
</tr>
<tr>
<td>8.10</td>
<td>33.60</td>
<td>0.7960</td>
</tr>
<tr>
<td>10.24</td>
<td>40.82</td>
<td>0.7927</td>
</tr>
<tr>
<td>12.20</td>
<td>46.90</td>
<td>0.7896</td>
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<tr>
<td>14.05</td>
<td>52.35</td>
<td>0.7889</td>
</tr>
<tr>
<td>17.30</td>
<td>61.11</td>
<td>0.7882</td>
</tr>
<tr>
<td>18.20</td>
<td>63.37</td>
<td>0.7690</td>
</tr>
<tr>
<td>21.35</td>
<td>70.79</td>
<td>0.7674</td>
</tr>
<tr>
<td>24.31</td>
<td>77.17</td>
<td>0.7620</td>
</tr>
<tr>
<td>25.50</td>
<td>79.50</td>
<td>0.7121</td>
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<tr>
<td>26.61</td>
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<td>29.00</td>
<td>86.13</td>
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<td>32.00</td>
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<td>33.20</td>
<td>93.90</td>
<td>0.4770</td>
</tr>
<tr>
<td>35.60</td>
<td>97.10</td>
<td>0.4657</td>
</tr>
</tbody>
</table>

The values of the emf given in Table II are plotted in Fig. 4 as ordinate using the normality of the acid as abscissa. The experimental results fall on a step-wise curve with each of the four steps corresponding to a hydrate of sulfuric acid. The first step at 30.9 N corresponds to $\text{H}_2\text{SO}_4 \cdot \text{H}_2\text{O}$; the second at 24.4 N to $\text{H}_2\text{SO}_4 \cdot 2\text{H}_2\text{O}$; the third at
THE VARIATION OF CELL VOLTAGE WITH CONCENTRATION OF ELECTROLYTE

(1) tantalum anode and platinum cathode
(2) platinum anode and cathode

NORMALITY OF H₂SO₄

Fig. 4
17.1 N to $\text{H}_2\text{SO}_4 \cdot 4\text{H}_2\text{O}$; the fourth at 5.01 N to $\text{H}_2\text{SO}_4 \cdot 20\text{H}_2\text{O}$.

It should not be at all surprising that hydrates of sulfuric acid are indicated in an electrolytic cell. Hydrates of sulfuric acid are known to exist in solution. Since the formation of a hydrate in solution is associated with a considerable free energy change, it should be associated with a marked potential change in an electrolytic cell. The potential should vary with the nature of the hydrate formed, hence it should be possible to detect the presence of hydrates of sulfuric acid in an electrolytic cell.

Let us consider the curve shown in Fig. 4. We will first consider the cell voltage at a very low acid concentration. At this concentration we have a definite cell potential, and as we increase the acid concentration we have a decrease in the overall free energy change causing the potential to drop. The drop is fairly constant up to a concentration of 5 N and over this range we have the hydrate $\text{H}_2\text{SO}_4 \cdot 20\text{H}_2\text{O}$. Beyond this concentration we do not have enough water present to form this hydrate so we have another hydrate forming, namely $\text{H}_2\text{SO}_4 \cdot 4\text{H}_2\text{O}$. From 5 N to 17.1 N the slope of our curve is constant and over this range we have the hydrate $\text{H}_2\text{SO}_4 \cdot 4\text{H}_2\text{O}$. Beyond 17.1 N we do not have enough water present to form this hydrate so that we have a change over to $\text{H}_2\text{SO}_4 \cdot 2\text{H}_2\text{O}$ accompanied by a marked change in potential. The other steps can be explained similarly.

The portion of the tantalum-platinum curve A-B shown in Fig. 4 was plotted by measuring the cell potential starting
at a low concentration of acid and working up to a concentration of 17 N. The other end of the curve was plotted by measuring the cell potential using C.P. acid then working back to 18.3 N. This gave the curve C-D-E. An apparent discrepancy in the curve is shown dotted between D and E. To check this part of the curve tests were carried out with 18.3 N, 22.1 N and 24.2 N acid. After 20 hours, values were obtained falling on the dotted portion of the curve but after another 12 hours these approached the solid portion of the curve D-E and remained constant for over 5 hours.

The curve obtained by Assaly is also shown in Fig. 4 for comparison with that obtained by the writer. The efficiency of the tantalum-platinum cell is clearly shown to be much greater than that of Assaly's over most of the acid range. The great drop in efficiency at the high acid concentration is probably due to the tantalum dissolving in the sulfuric acid.

Other experimenters have shown evidence of three hydrates of sulfuric acid by freezing methods. Pickering obtained the hydrate $\text{H}_2\text{SO}_4 \cdot 4\text{H}_2\text{O}$; Giron found the hydrate $\text{H}_2\text{SO}_4 \cdot 2\text{H}_2\text{O}$; and Donk obtained crystalline sulfuric acid monohydrate $\text{H}_2\text{SO}_4 \cdot \text{H}_2\text{O}$.

The freezing point chart of sulfuric acid hydrates is shown in Fig. 5. B, D and G are the cryohydrate points and C, E and H are the melting points of the hydrates $\text{H}_2\text{SO}_4 \cdot 4\text{H}_2\text{O}$, $\text{H}_2\text{SO}_4 \cdot 2\text{H}_2\text{O}$ and $\text{H}_2\text{SO}_4 \cdot \text{H}_2\text{O}$ respectively. From B to C, D to E
FREZING POINT CHART
of
SULFURIC ACID HYDRATES

Fig. 5
and G to H these respective hydrates solidify out. The existence of \( \text{H}_2\text{SO}_4 \cdot 2\text{OH}_2\text{O} \) is not indicated in the graph but would probably be detected between A and B on closer examination.

Measurements were made of emf with time for several concentrations of acid. The measurements for three concentrations are given in Table III. The first column gives the time, the second, third and fourth give the emf measured with 0.98 N, 27.2 N and 35.5 N acid respectively.

Time-voltage curves for these acid concentrations are shown in Fig. 6. The initial voltage in all cases was very high, dropping off to a constant value after about 24 hours. At the beginning an equilibrium existed between the tantalum metal and the sulfur dioxide in solution but as time goes on an oxide film forms on the tantalum and an equilibrium is set up between the tantalum oxide and the sulfur dioxide in solution. The initial drop in voltage occurs while the oxide film is forming then the voltage gradually approaches a constant value.
<table>
<thead>
<tr>
<th>t</th>
<th>E(0.98N)</th>
<th>E(27.2N)</th>
<th>E(35.6N)</th>
</tr>
</thead>
<tbody>
<tr>
<td>hours</td>
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<td>volts</td>
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THE VARIATION OF CELL VOLTAGE WITH TIME

(1) 0.96 N H₂SO₄
(2) 2.72 N H₂SO₄
(3) 9.76 N H₂SO₄

VOLTAGE

TIME (hours)

Fig. 6
EFFECT OF THE RATE OF SULFUR DIOXIDE AND OXYGEN INTO THE CELL ON THE EMF

In the present work a rate of flow of 60 bubbles per minute was used for both the sulfur dioxide and oxygen. Although this method of control was only qualitative it served its purpose since only voltage was being studied and not rates of formation of acid.

On increasing the rate of flow of sulfur dioxide into the cell from 60 bubbles per minute to 120 bubbles per minute the voltage increased from 0.8000 to 0.8007 after several hours. This value returned to 0.8002 on decreasing the rate back to 60 bubbles per minute. Decreasing the rate to 30 bubbles per minute cut the emf to 0.7991 which returned to 0.8001 on increasing the rate to 60 bubbles per minute.

On increasing the rate of flow of oxygen into the cell from 60 bubbles per minute to 120 bubbles per minute the voltage decreased from 0.8000 to 0.7990. On decreasing the rate to 60 bubbles per minute the emf went up to 0.7997 while shutting off the oxygen completely, caused the emf to increase slowly to 0.8015.
CONCLUSION

From the work of T. C. Assaly and the writer it has been shown that the utilization of the free energy of sulfur dioxide in the form of electrical energy has considerable possibilities.

The present work has been concentrated entirely on the voltage developed in a sulfur dioxide-oxygen cell but in future the rate of formation of sulfuric acid in such a cell should be studied. If this rate could be made appreciable then this cell would have definite possibilities for the production of sulfuric acid and electrical energy on a commercial scale.
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