

# Effects of the Composition of Solution on the Potential of Pyrite Electrode

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*With 9 Text-figures*

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## I. Introductory

R. C. WELLS, well known for his investigations on the potential of several mineral electrodes, classifies the potential of pyrite electrode as either "ferric-ferrous" or "sulphide."<sup>1)</sup>

The subject was thoroughly studied from a leptological point of view by Prof. A. MATSUBARA,<sup>2)</sup> who came to the conclusion that the potential of pyrite electrode is proportional to the surface density of "free crystalline bond" of a surface in action, and that the electrical phenomena exhibited by pyrite are chiefly due to the action of its sulphur atoms.

Under his suggestion, the present writer made an attempt to determine the effects of hydrogen ion concentration of a solution on the potential of a sulphide mineral electrode. So far as the ex-

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1) R. C. WELLS, "Electric Activity in Ore Deposits," *Bulletin U. S. G. S.*, 548 (1914).  
2) A. MATSUBARA, These Memoirs, Ser. B, Vol. I, No. 4.

perimental results are concerned, it is quite possible to say that the potential of pyrite electrode can be expressed by a general equation such as

$$E = E_0 + k \cdot \text{Log} [\text{H}^+]$$

where  $E_0$  denote the  $k$  are constants which are almost independent of the concentration but depend on the nature of the solution, and  $[\text{H}^+]$  the hydrogen ion concentration.

## II. The Potential of Pyrite against Acid and Alkaline Solutions

In order to corroborate the effects of change in the concentration of acid and alkaline solution upon the potential of pyrite, the following experiments were carried out. Sulphuric acid and sodium hydroxyde solutions of various concentrations, ranging from 6 normal to  $10^{-5}$  normal, were respectively prepared. About 50 c.c. of each of them was placed in a glass jar having a capacity of ca. 200 c.c. The air was completely expelled from the jar by introducing a stream of knal gas through the solution for an hour or so before the measurements. For the electrode particular care was taken to select a specimen having a perfectly smooth and shining crystal face, free from all mechanical and chemical defects such as striation, tarnish and minute etched graving. The piece of pyrite thus selected was fastened to the end of a suitable glass tube by means of sealing wax, which covered the whole except the crystal face which was to serve as an electrode. It was immersed in the solution by means of a plug fitted to the mouth of the jar. This half cell was connected to a saturated potassium chloride calomel electrode<sup>1)</sup> in conjunction with a saturated salt bridge. The latter was preferred in order to eliminate or minimize the contact potential brought about by the interdiffusion at the conjunctions. The whole was kept at 18°C in a thermostat.

All measurements were made by a methode of compensation with a galvanometer of fairly high sensibility. The electromotive force of the cell underwent some change during about half an hour or so after the immersion of the electrode, but after it had settled

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1) LOOMIS and ACREE, *Amer. Chem. J.*, **46**, 621 (1911).

FALES and VOSBURGH, *J. Amer. Chem. Soc.*, **40**, 1291 (1918).

FALES and MUDGE, *ibid.*, **42**, 2434 (1920).

nearly to a constant, the reading was taken. The potential of pyrite electrode was obtained from the electromotive force of the combined cell and the potential of saturated potassium chloride calomel electrode, namely 0.5252 volts at 18° C<sup>1)</sup>; but the diffusion potentials were not taken into account. The following tables show the results thus obtained.

Table I  
Electrode: (100) of pyrite from Tsunatori.

Solutions \ Electrodes	I (volt)	II (volt)	III (volt)	IV (volt)
6n-H <sub>2</sub> SO <sub>4</sub>	0.928	0.946	0.924	0.950
n „	0.951	0.952	0.932	0.950
10 <sup>-1</sup> „	0.940	0.943	0.927	0.935
10 <sup>-2</sup> „	0.902	0.912	0.898	0.908
10 <sup>-3</sup> „	0.846	0.857	0.855	0.855
10 <sup>-4</sup> „	0.790	0.812	0.799	0.805
10 <sup>-5</sup> „	0.748	0.753	0.745	0.752

Table II  
Electrode: (100) of pyrite from Tsunatori.

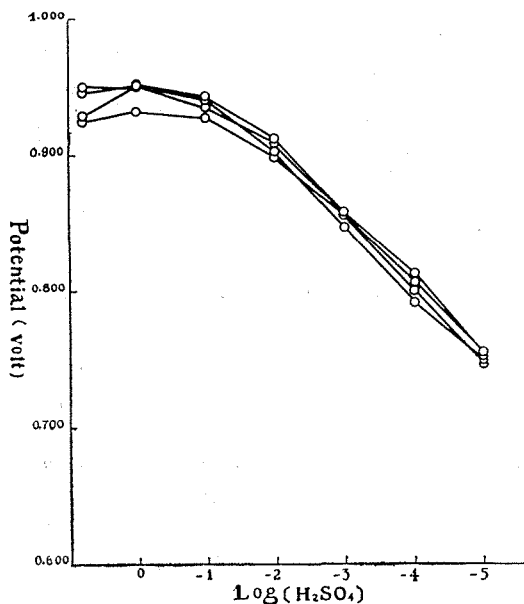
Solutions \ Electrodes	I (volt)	II (volt)	III (volt)	IV (volt)
6n-NaOH	0.310	0.290	0.322	0.310
n „	0.314	0.319	0.334	0.314
10 <sup>-1</sup> „	0.375	0.357	0.389	0.361
10 <sup>-2</sup> „	0.432	0.402	0.424	0.404
10 <sup>-3</sup> „	0.482	0.461	0.466	0.449*
10 <sup>-4</sup> „	0.538	0.521	0.507	0.511
10 <sup>-5</sup> „	0.580	0.596	0.600	0.573

The relations between the potential and the concentration are diagrammatically represented in Fig. 1, for the concentration of sulphuric acid, and in Fig. 2 for the sodium hydroxyde.

In the case of sodium hydroxyde, the potentials vary almost in linear manner with the logarithm of the concentration, as may be seen in Fig. 2; whereas in the case of sulphuric acid the same relation holds only under the concentration of 10<sup>-2</sup> normal, as may be seen in Fig. 1.

1) FALES and MUDGE, *loc. cit.*

Fig. 1



When the pyrite is immersed in the sulphuric acid, the hydrogen sulphide, in all possibility, would be produced on its surface, even though it may merely be a trace, as the result of some chemical action of pyrite itself. If so, the potential of pyrite should assume rather a lower value than otherwise especially in highly concentrated acid. Probably, this is the reason why the potentials against 6*n*-H<sub>2</sub>SO<sub>4</sub> are always lower than those against the *n*-H<sub>2</sub>SO<sub>4</sub>.

In order to examine whether or not the potential of pyrite is quite independent of the sodium or sulphate ion concentration, the measurements were carried out employing the solution of Na<sub>2</sub>SO<sub>4</sub> of various concentrations. Results are cited in Table III.

Regardless of the change in concentration, the results are practically constant, the total mean being 0.665 volts.

Fig. 2

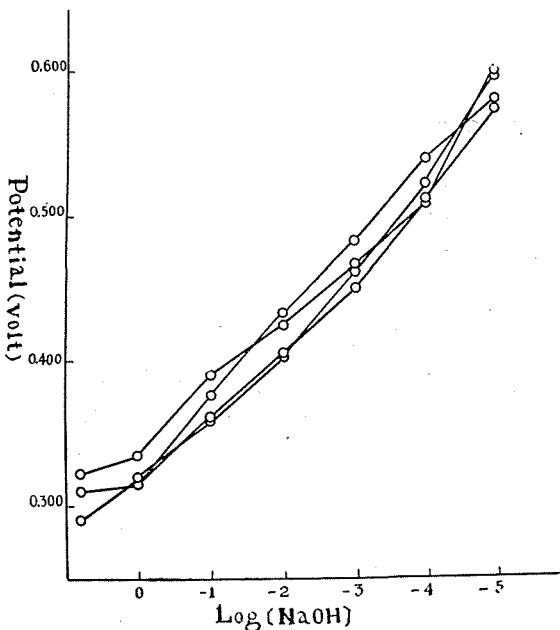


Table III  
Electrode: (100) of pyrite from Tsunatori.

Solutions \ Electrodes	I (volt)	II (volt)	III (volt)	IV (volt)
Sat. Na <sub>2</sub> SO <sub>4</sub>	0.665	0.651	0.665	0.668
n     ,,	0.671	0.659	0.670	0.672
10 <sup>-1</sup> ,,	0.666	0.662	0.666	0.672
10 <sup>-2</sup> ,,	0.665	0.652	0.663	0.672
10 <sup>-3</sup> ,,	0.671	0.657	0.652	0.672
10 <sup>-4</sup> ,,	0.670	0.657	0.659	0.667
Mean	0.668	0.656	0.663	0.671
Total mean . . . . .				<b>0.665</b>

### III. Effects of Hydrogen Ion Concentration upon the Potential of Pyrite

After the previous experiments, it is evident that the potential of pyrite against the acid or alkaline solution is remarkably responsible for the hydrogen ion concentration of the solution. The magnitude of the latter in solutions which were used in the previous experiments may be taken approximately as follows.

Table IV

#### a) Sulphuric acid

Solutions	Hydrogen ion conc.	pH	Remarks (Various calculated values of hydrogen ion conc.)
n-H <sub>2</sub> SO <sub>4</sub>	0.52	0.28	0.5336 <sup>3)</sup> 0.517 <sup>3)</sup>
10 <sup>-1</sup> ,,	0.059	1.23	0.0585 <sup>1)</sup> 0.0625 <sup>2)</sup> 0.0555 <sup>3)</sup> 0.05859 <sup>1)</sup> 0.0587 <sup>2)</sup>
10 <sup>-2</sup> ,,	0.0082	2.09	0.00825 <sup>1)</sup> 0.0081 <sup>2)</sup> 0.00804 <sup>3)</sup>
10 <sup>-3</sup> ,,	0.00094	3.03	0.000943 <sup>3)</sup>
10 <sup>-4</sup> ,,	0.0001	4.00 <sup>3)</sup>	
10 <sup>-5</sup> ,,	0.00001	5.00 <sup>3)</sup>	

1) NOYES and STEWART, *J. Amer. Chem.*, **32**, 1133 (1910).

2) These values were obtained from conductance data by NOYES and STEWART.

Table IV (Continued)

b) Sodium hydroxyde. The pH-values were calculated from the hydroxyde ion concentrations assuming the dissociation constant of the water as  $0.72 \times 10^{-14}$  at  $18^\circ\text{C}$ .<sup>7)</sup>

Solutions	Hydroxyde ion conc.	pH	Remarks (Various calculated values of hydroxyde ion conc.)
n-NaOH	0.72	14.00	0.722 <sup>8)</sup> 0.725 <sup>11)</sup>
$10^{-1}$ „	0.090	13.10	0.0899 <sup>8)</sup> 0.0903 <sup>11)</sup>
$10^{-2}$ „	0.0094	12.12	0.00935 <sup>11)</sup> 0.00936 <sup>9)</sup> 0.00962 <sup>10)</sup> 0.00939 <sup>11)</sup>
$10^{-3}$ „	0.001 <sup>12)</sup>	11.14	
$10^{-4}$ „	0.0001 <sup>12)</sup>	10.14	
$10^{-5}$ „	0.00001 <sup>12)</sup>	9.14	

3) T. ONODA, *ZS. anorg. Chem.*, **165**, 137 (1927).

These values were calculated from the results of the equivalent conductivity and the specific viscosity at  $30^\circ\text{C}$  according to the following equation:

$$C_H = C_n \left( \frac{\Lambda}{\Lambda_\infty} \cdot \frac{\eta}{\eta_\infty} \right)$$

where  $C_H$  stands for the hydrogen ion concentration,  $C_n$ , the normality of sulphuric acid,  $\Lambda$  and  $\Lambda_\infty$  the equivalent conductivity at the given and the maximum dilution,  $\eta$  and  $\eta_\infty$ , the viscosity of the given solution and the water respectively.

4) J. E. ENKLAAR, *ZS. phy. Chem.*, **80**, 617 (1912).

5) The writer had calculated only from the following data of the equivalent conductivity. According to KOHLRAUSCH [Landolt-Börnstein phy.-chem. Tabellen, II, 5 Aufl. Berlin, S. 1081 (1923)],

$$\left. \begin{array}{l} \Lambda = 198 \text{ (Normal)} \\ = 225 \text{ (0.1 Normal)} \\ = 308 \text{ (0.01 Normal)} \\ = 361 \text{ (0.001 Normal)} \end{array} \right\} \text{ at } 18^\circ\text{C}.$$

According to NOYES and co-workers [*ZS. phy. Chem.*, **70**, 362 (1910)]

$$\Lambda_\infty = 383 \quad \text{at } 18^\circ\text{C}.$$

6) Sulphuric acid was assumed to be in a completely dissociated state.

7) R. LORENZ and BÖHI, *ZS. phy. Chem.*, **66**, 748 (1909).

8) ABEGG's *Handbuch der anorganischen Chemie*, II, Abteil. 1, S. 246 (1908).

9) This value was obtained from the dissociation degree which was interpolated between the two, 94.1 and 93.0% at 0.009458 and 0.01058 normal reported by H. C. JONES [*ZS. phy. Chem.*, **12**, 639 (1893)].

10) A. A. NOYES, A. C. MELCHER, H. C. COOPER and G. W. EASTMAN, *ZS. phy. Chem.* **70**, 362 (1910).

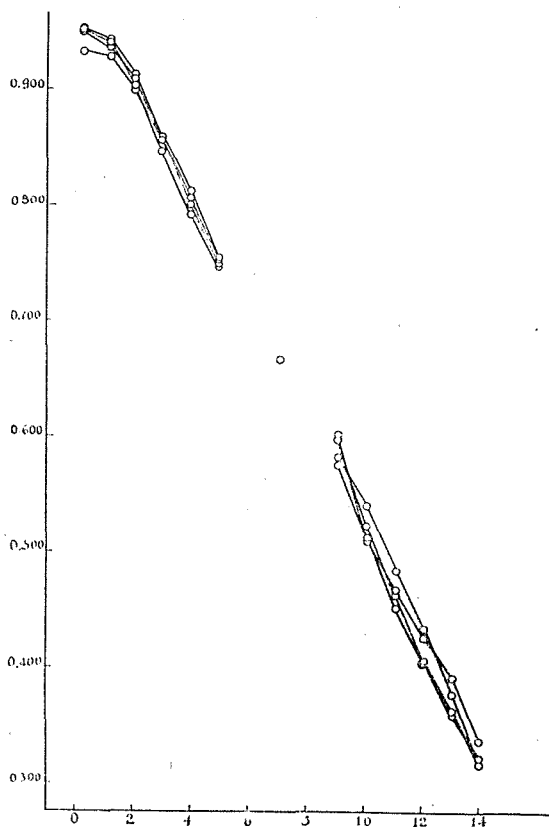
11) The present writer had calculated only from the data of the equivalent conductivity. According to KOHLRAUSCH [Landolt-Börnstein phy.-chem. Tabellen, II, 5 Aufl. Berlin, S. 1082 (1923)],

$$\left. \begin{array}{l} \Lambda = 157 \text{ (Normal)} \\ = 195.4 \text{ (0.1 Normal)} \\ = 203.4 \text{ (0.01 Normal)} \end{array} \right\} \text{ at } 18^\circ\text{C}.$$

According to NOYES and co-workers

$$\Lambda_\infty = 216.5 \quad \text{at } 18^\circ\text{C}.$$

Fig. 3



The relation between the potential and pH-value of the solution is diagrammatically represented in Fig. 3. There we adopted 0.665 volts, which was obtained as the mean value of pyrite potential against the solution of  $\text{Na}_2\text{SO}_4$  in the foregoing experiment, as that against a neutral solution whose pH-value can be calculated as 7.07 from the dissociation constant of water.

As may be distinctly seen in Fig. 3, the potential variation is nearly accordant with that of pH of the solution in the linear manner.

To demonstrate the above statement with more certainty, the following experiment was undertaken. 100 c.c. of  $n/10\text{-H}_2\text{SO}_4$  was taken as the initial solution and the air was completely expelled from it as illustrated before. The pyrite electrode was then immersed in this solution and electrically connected to a calomel electrode. The electromotive force of the combined cell was measured when it had settled nearly to a constant. A suitable amount of  $n/10\text{-NaOH}$  was added to the initial solution from time to time and the measurements were repeated at each step of addition. The potential of pyrite electrode was calculated from the electromotive force of the cell and the potential of calomel electrode. The results

12) Sodium hydroxyde was assumed to be in the state of complete dissociation. According to H. C. JONES, the dissociation degree of the solution of 0.001058 normal is 100%.

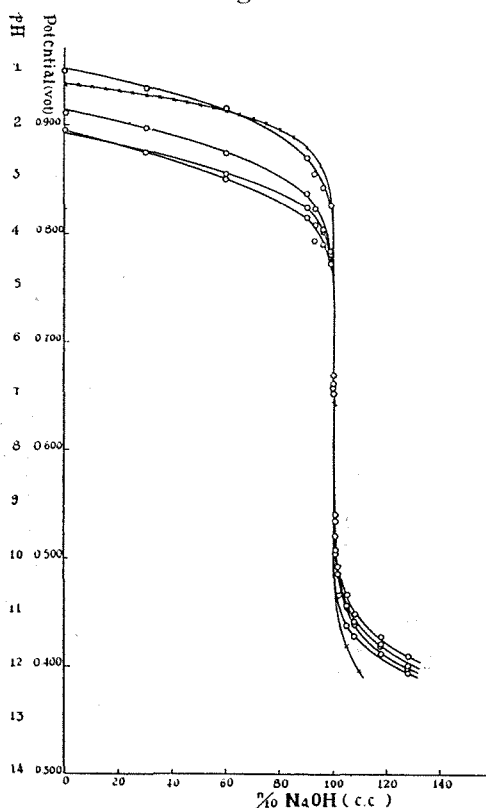
Table V  
 Initial solution: 100 c.c. of n/10-H<sub>2</sub>SO<sub>4</sub>  
 Electrode: (100) of pyrite from Tsunatori.

n/10-NaOH added (c.c.)	Potential of pyrite electrode (volt)			
	I	II	III	IV
0	0.911	0.949	0.895	0.895
30	0.898	0.934	0.875	0.876
60	0.872	0.915	0.851	0.856
90	0.837	0.856	0.793	0.808
93	0.823	0.856	0.793	0.808
96	0.802	0.843	0.791	0.804
99	0.781	0.826	0.773	0.782
100	0.661	0.669	0.657	0.652
101	0.503	0.534	0.520	0.540
102	0.465	0.492	0.486	0.507
105	0.437	0.454	0.455	0.466
108	0.427	0.438	0.440	0.448
118	0.411	0.417	0.419	0.427
148	0.392	0.397	0.399	0.408

Fig. 4

are summarized in the above table, and the relation between the potential and the volume of n/10-NaOH is graphically represented in Fig. 4.

Under the same operation, J. E. ENKLAAR<sup>1)</sup> had already determined the pH-value at each step of the addition of n/10-NaOH. The so called titration curve he thus obtained is shown in Fig. 4 drawn through the points marked by crosses to distinguish it from the writer's curves. As one might expect, the dispositions of the writer's curves are quite similar with the titration curve of ENKLAAR, i. e. the variations of the pyrite potential and pH of the solution



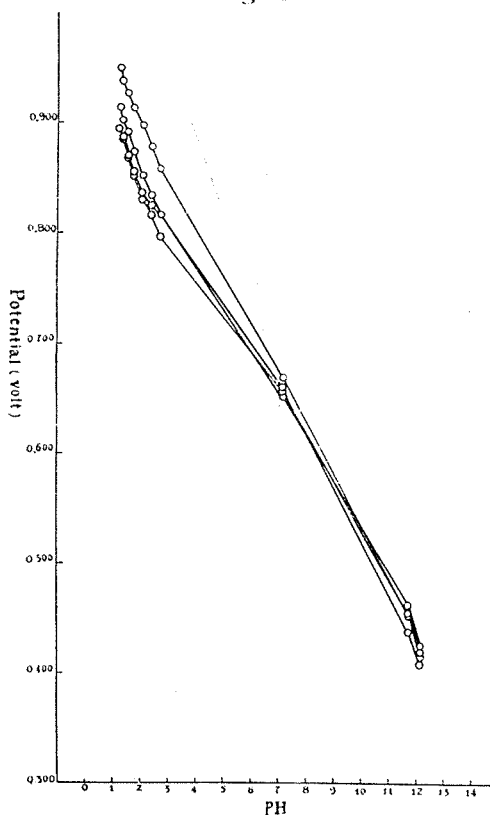
1) J. E. ENKLAAR, *loc. cit.*



Table VI

n/10-NaOH added (c.c.)	pH	Potential of pyrite electrode (volt)			
		I	II	III	IV
0	1.2322	0.914	0.950	0.890	0.895
20	1.3414	0.902	0.938	0.885	0.887
40	1.5182	0.889	0.927	0.868	0.870
60	1.7261	0.873	0.914	0.852	0.856
80	2.0589	0.852	0.897	0.830	0.837
90	2.3969	0.834	0.879	0.817	0.825
95	2.7262	0.817	0.858	0.797	0.817
100	7.1543	0.661	0.667	0.657	0.652
105	11.669	0.438	0.454	0.455	0.463
110	12.085	0.408	0.415	0.419	0.425

Fig. 5



similarly show a rapid change at the neutral point. Therefore it is safe to assume that the variation of the pyrite potential is based upon the hydrogen ion concentration of the solution.

On each curve in Fig. 4, the pyrite potential at each step of the addition of n-NaOH is interpolated and the values thus obtained were correlated to the corresponding pH-values which were actually obtained by ENKLAAR. The results are shown in Table VI.

According to the above data, the relation between the potential and pH-value could be expressed by almost linear curves as is shown in Fig. 5.

## IV. Pyrite Potential against Sulphide Solution

In order to determine the effects of change in the concentration of sulphide solution upon the potential of pyrite, the following ex-

Table VII  
Electrode: (100) of pyrite from Tsunatori.

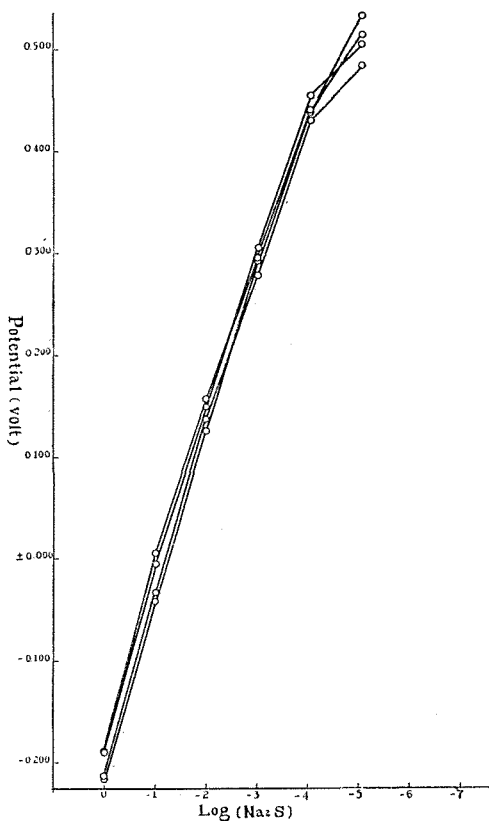
Solutions	Electrodes	I (volt)	II (volt)	III (volt)	IV (volt)	Mean
	n-Na <sub>2</sub> S	-0.190	-0.216	-0.213	-0.189	-0.202
	10 <sup>-1</sup> „	-0.006	-0.042	-0.033	-0.006	-0.022
	10 <sup>-2</sup> „	0.149	0.125	0.137	0.157	0.142
	10 <sup>-3</sup> „	0.306	0.296	0.279	0.294	0.294
	10 <sup>-4</sup> „	0.455	0.439	0.406	0.440	0.435
	10 <sup>-5</sup> „	0.505	0.546	0.460	0.514	0.506

periment was undertaken. Sodium sulphide solutions of various content ranging from normal to 10<sup>-5</sup> normal by tenfold dilution were prepared. The pyrite potentials against these solutions were respectively determined in the order of concentration under the same procedure as before. Results are shown in the above table.

The relations between the potential and the logarithm of concentration are given in Fig. 6.

The pyrite potential against n-Na<sub>2</sub>S is considerably lower than that against n-NaOH, their difference being ca. 0.5 volts. This fact may be ascribed to the presence of hydrosulphide

Fig. 6



or sulphide ion as their hydrogen ion concentrations are considered to be nearly the same.

Whereas the hydrosulphide or sulphide ion concentrations are decreased with the dilution, the hydrogen ion, on the contrary, is increased: but those variations are by no means restricted by the relation of solubility product of hydrogen sulphide. In this circumstance, the potential variations are far greater than those in the case of the solution of sodium hydroxyde for the same interval of change of the concentration.

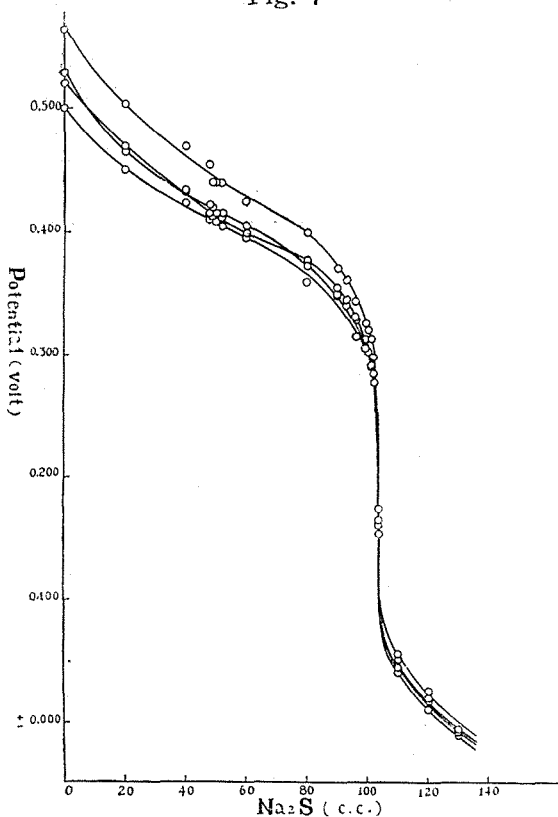
To determine the effect of hydrogen ion concentration on the pyrite potential, the following experiments were undertaken. 100 c.c. of 1.028 n-H<sub>2</sub>SO<sub>4</sub> was syphoned into a glass jar with a capacity of about 200 c.c. The pyrite electrode was immersed in the solution by means of the gum stopper in which the electrode was inserted. After perfect saturation of the solution with hydrogen sulphide, a

Table VIII

Initial solution: 100 c.c. of 1.028 n-H<sub>2</sub>SO<sub>4</sub>  
Electrode: (100) of porite from Tsunatori.

n-Na <sub>2</sub> S added (c.c.)	Potential of pyrite electrode (volt)			
	I	II	III	IV
0	0.976	0.950	0.991	0.976
Satd. w. H <sub>2</sub> S	0.522	0.530	0.564	0.502
20	0.470	0.465	0.503	0.451
40	0.435	0.433	0.470	0.424
48	0.416	0.423	0.455	0.411
49	0.413	0.421	0.441	0.410
50	0.408	0.416	0.440	0.405
52	0.412	0.415	0.440	0.405
60	0.399	0.405	0.425	0.396
80	0.377	0.372	0.399	0.369
90	0.354	0.349	0.371	0.348
93	0.345	0.344	0.353	0.339
96	0.328	0.327	0.343	0.330
99	0.312	0.308	0.325	0.314
100	0.302	0.300	0.319	0.306
101	0.296	0.290	0.313	0.301
102	0.284	0.277	0.298	0.291
103	0.160	0.153	0.174	0.164
110	0.044	0.040	0.056	0.051
120	0.017	0.011	0.025	0.018
130	-0.012	-0.012	-0.007	-0.012

Fig. 7



suitable quantity of  $n\text{-Na}_2\text{S}$  was added to the solution. The pyrite potentials at each step of the addition were determined just as before. In this case, the variation of hydrogen and hydrosulphide or sulphide ion concentrations are always subjected to the relation of the solubility product of hydrogen sulphide.

By introducing hydrogen sulphide into the initial solution of  $1.028\ n\text{-H}_2\text{SO}_4$ , the pyrite potential underwent a considerable change, but after perfect saturation it showed ca. 0.53 volts.

The results are tabulated in Table VIII, and the relation between the potential and the volume of  $n\text{-Na}_2\text{S}$  is diagrammatically represented in Fig. 7.

The curves which again resemble the titration curve, show us well that the potential behaviors depend on the hydrogen ion concentration. To represent this relation more certainly, the relation between the potential of pyrite and pH of the solution is diagrammatically shown in Fig. 8 employing the figures in Table IX. That the potential stands in a linear relation with pH is quite clear.

Table IX

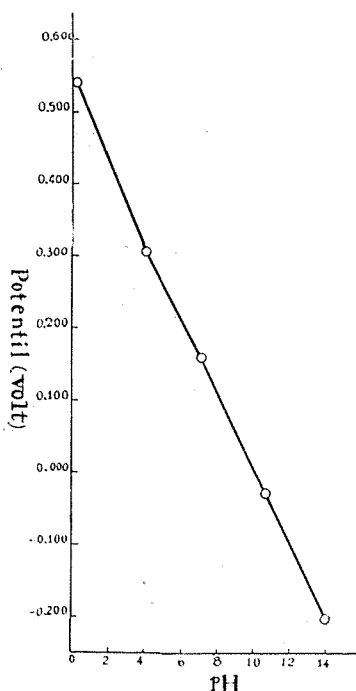
Solutions	Hydrogen ion conc.	pH	Potential of pyrite electrode (volt)					
$n\text{-H}_2\text{SO}_4$ satd. w. $\text{H}_2\text{S}$	0.53	0.28	0.522	0.530	0.564} Mean 0.502	0.588	0.548}	0.542
$n/10\text{-Na}_2\text{SO}_4$ satd. w. $\text{H}_2\text{S}$	$0.9 \times 10^{-11}$	4.05	0.303	0.313	0.303} Mean 0.302			0.305

Table IX (Continued)

Solutions	Hydrogen ion conc.	pH	Potential of pyrite electrode (volt)
Neutral soln. satd. w. H <sub>2</sub> S	$0.85 \times 10^{-7}$ 2)	7.07	0.16 <sup>3)</sup>
n-NaSH	$3.3 \times 10^{-11}$ 4)	10.48	-0.03 <sup>5)</sup>
2n-Na <sub>2</sub> S	$1.3 \times 10^{-14}$ 6)	13.89	-0.237 -0.196 -0.230 -0.216 } Mean -0.220

### V. Experimental Formulas for the Potential of Pyrite Electrode

Fig. 8



According to the previous experimental results, the potential of pyrite could be regarded as the linear function of pH of the solution as shown in Fig. 9. Therefore we may write the general expression of the pyrite potential as follows:

$$E = E_0 - k.pH$$

where  $E_0$  and  $k$  show the constants which are quite independent of the concentration but depend on the nature of solution. Here the constant  $E_0$ , which corresponds to the potential against a solution whose pH is zero, could be obtained by the graphical extrapolation with the help of the nearly linear function as 1.007 volts in the case of the acid and alkaline solution. Hence the equation becomes

$$E = 1.007 - k.pH.$$

By substituting  $E$  and pH in the

1) According to FALES [*Inorganic Quantitative Analysis*, p. 159 (1928)], the hydrogen ion concentration of hydrogen sulphide water is  $0.9 \times 10^{-4}$ .

2) This value was calculated from the dissociation constant of water.

3) From the curves given in Fig. 7, this value was approximately obtained as the potential corresponding to that at neutral state.

4) FALES, *loc. cit.*

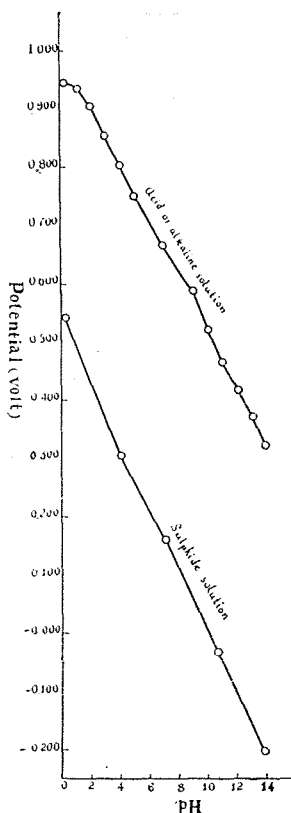
5) R. C. WELLS, *loc. cit.*

6) FALES, *loc. cit.*

Table X

Solutions	pH	$E$	$k$
n-H <sub>2</sub> SO <sub>4</sub>	0.28	0.946	—
10 <sup>-1</sup> "	1.23	0.936	0.058 (?)
10 <sup>-2</sup> "	2.09	0.905	0.049
10 <sup>-3</sup> "	3.03	0.853	0.051
10 <sup>-4</sup> "	4	0.802	0.051
10 <sup>-5</sup> "	5	0.750	0.051
n/10-Na <sub>2</sub> SO <sub>4</sub>	7.07	0.665	0.048
10 <sup>-2</sup> n-NaOH	9.14	0.587	0.046
10 <sup>-1</sup> "	10.14	0.519	0.048
10 <sup>-3</sup> "	11.14	0.465	0.049
10 <sup>-2</sup> "	12.12	0.416	0.049
10 <sup>-1</sup> "	13.10	0.371	0.049
n "	14.00	0.320	0.049
Mean . . . . .			0.050

Fig. 9



above equation with the corresponding experimental results, we get the constant  $k$  according to the following equation :

$$k = \frac{1.007 - E}{\text{pH}}$$

For the calculation of  $k$ , we adopted the mean values of pyrite potentials given in Tables I, II and III as  $E$ . The results are summarized in Table X.

There is a fair approach to constancy in the value of  $k$ , its mean value being 0.050. This is quite sufficient to show that the potential of pyrite electrode in this case may be expressed by the following experimental formula :

$$E = 1.007 - 0.050 \cdot \text{pH},$$

$$\text{or } E = 1.007 + 0.050 \cdot \text{Log} [\text{H}'].$$

In the case of the sulphide solution, the potential variation and pH change of the solution is in the linear manner but with opposite sense just as in the case of the acid and alkaline solution. Therefore the following equation should also be applicable for the potential behaviors in this case :

$$E = E_0 - k \cdot \text{pH}.$$

The value of  $E_0$  is easily obtained by means of the graphical extrapolation as 0.557 volts. By substituting  $E$  and pH with the results shown in Table IX we get  $k$  at various values of pH. The results thus obtained are as follows.

Table XI

Solutions	pH	$E$ (volt)	$k$
n-H <sub>2</sub> SO <sub>4</sub> satd. w. H <sub>2</sub> S	0.28	0.542	0.054
n/10-Na <sub>2</sub> SO <sub>4</sub> satd. w. H <sub>2</sub> S	4.05	0.305	0.062
Neutral solution satd. w. H <sub>2</sub> S	7.07	0.16	0.056
n-NaHS	10.48	-0.03	0.056
2n-Na <sub>2</sub> S	13.89	-0.220	0.056
Mean . . . . .			0.057

As there is fair approach to constancy in the value of  $k$ , whose mean value is precisely identical with  $\frac{RT}{F} = 0.055$  at 18°C, so we may write the general experimental formula for the pyrite potential as:

$$E = 0.557 - 0.056 \cdot \text{pH}$$

or 
$$E = 0.557 + 0.057 \cdot \text{Log} [\text{H}^+]$$

## VI. Conclusion

- 1) The potential of a pyrite electrode depends to a remarkable degree on the hydrogen ion concentration of the solution. But it is quite independent of the sodium or sulphate ion concentration.
- 2) In the case of acid or alkaline solution, the pyrite potential corresponding to that against the solution whose pH-value is zero is 1.007 volts; whereas in the case of sulphide solution, it is only 0.557 volts.

3) The following two experimental formulas are proposed for the calculation of potential of pyrite electrode;

a) 
$$E = 1.007 + 0.050 \text{ Log} [\text{H}^+]$$

b) 
$$E = 0.557 + 0.057 \text{ Log} [\text{H}^+]$$

a) is applicable for the acid or alkaline solution; and b) for the sulphide solution.

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