

## CONCLUSIONS.

1. **Unguentum Iodi**, U. S. P. is readily absorbed through the skin.
2. **Unguentum Iodi Denigrescens**, N. F. is apparently not absorbed through the skin as shown by absence of iodides in the urine.
3. **Petroxolinum Iodidi**, N. F. is absorbed through the skin.
4. **Unguentum Iodi**, U. S. P. possesses antiseptic properties.
5. **Unguentum Iodi Denigrescens**, N. F. apparently does not possess antiseptic properties.

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- (3) Herzfeld and Elin, *Zentr. Bioch. Bioph.*, 13 (1912), 251.<sup>1</sup>
- (4) Bartenbach, *Bioch. Bioph.*, 10 (1909), 278.<sup>1</sup>

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THE EFFECT OF SUNLIGHT ON CITRIC ACID IN THE PRESENCE OF  
FERRIC SALTS.

WITH SPECIAL REFERENCE TO THE ELIXIR OF IRON QUININE AND STRYCHNINE  
AND ELIXIR OF IRON QUININE AND STRYCHNINE PHOSPHATES.\*

BY JOSEPH B. BURT.

This study was suggested by the investigations of Chilson<sup>1</sup> and of Mueller.<sup>2</sup> The principal reason for modification of the pharmacopœial formula for the elixir of iron, quinine and strychnine phosphates was because of the unstable character of the preparation manifested by precipitation and color changes upon standing. Chilson made a search of the literature for suggested changes and new formulas for the elixir and prepared samples in accordance with each method or modification found, without obtaining a single sample of the elixir which was entirely satisfactory. A review of the literature on the elixir of iron, quinine and strychnine shows that this preparation has also been subjected to the same criticism, and chiefly for the same reasons. In attempting to classify the suggested remedies for the incompatibility in the elixir of iron, quinine and strychnine phosphates, it

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<sup>1</sup> From "Manual of Pharmacy," Sollman, as original articles could not be obtained.

\* Scientific Section, A. Ph. A., St. Louis meeting, 1927.

<sup>1</sup> Thesis, "Suggested Modifications of the United States Pharmacopœial Formula for Elixir of Iron, Quinine and Strychnine Phosphates," University of Wisconsin, 1914.

<sup>2</sup> Thesis, "The Action of Light on Solutions of Soluble Ferric Phosphate," University of Wisconsin, 1915.

becomes evident that the study of this elixir has been almost entirely empirical with the result that many of the suggestions are contradictory, and fail to overcome the difficulty. In only one case<sup>1</sup> is an explanation given for the formation of the precipitate, which is assumed to be a mixture of the phosphates of the alkaloids, quinine and strychnine. Some of the suggestions offered for preventing or redissolving the precipitate are to leave the elixir slightly acid in reaction,<sup>2</sup> to have the reaction exactly neutral,<sup>3</sup> to change the order of mixing,<sup>4</sup> to use heat while preparing,<sup>5</sup> to avoid heating,<sup>6</sup> and to employ heat to redissolve the precipitate after its formation<sup>7</sup>. Others suggest modification of the formula by reducing the quantity of ammonium carbonate and acetic acid used,<sup>8</sup> by omitting these constituents<sup>9</sup> and the substitution of potassium citrate,<sup>10</sup> or sodium citrate,<sup>11</sup> by the addition of citric acid,<sup>12</sup> and by addition of lactic acid.<sup>13</sup> Others suggest the use of pyrophosphate of iron<sup>14</sup> instead of the soluble phosphate of iron. Another suggestion recommends the use of sodium hydroxide<sup>15</sup> instead of ammonia water in neutralizing the finished elixir. In two instances<sup>16</sup> writers say that the U. S. P. formula is perfectly satisfactory, if sufficient care is used in following the directions for preparation.

Little reference is made to the change in color. One writer<sup>17</sup> recommends storing the elixir in small, well-filled bottles, since it is prone to become dark. Another<sup>18</sup> gives a formula for which it is claimed the green color will be retained, even when exposed to light. It is suggested that the color may be retained indefinitely<sup>19</sup> by rendering the preparation acid with hydrochloric acid. Only four references are made to the fact that preparations of this type are sensitive to light.<sup>20</sup>

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<sup>1</sup> P. Thielke, *Drug. Circ.*, 51 (1907), 305 and 423.

<sup>2</sup> H. C. Hughes, *Am. J. Pharm.*, 78 (1906), 417; T. A. Egan, *Drug. Circ.*, 55 (1911), 691.

<sup>3</sup> R. J. Fritzing, *Drug. Circ.*, 55 (1911), 685.

<sup>4</sup> P. Thielke, *Drug. Circ.*, 51 (1907), 305 and 423; C. H. Ware, *Drug. Circ.*, 55 (1911), 408.

<sup>5</sup> T. Latham, *JOUR. A. PH. A.*, 2 (1913), 202; "Question Box Editor," *Pharm. Era*, 35 (1906), 146.

<sup>6</sup> C. A. Bachman, *Proc. Wis. Ph. A.*, 7 (1886), 44; A. F. Marquier, *Proc. N. J. Ph. A.*, 39 (1911), 47; T. A. Egan, *Drug. Circ.*, 55 (1911), 691.

<sup>7</sup> G. W. Koch, *Drug. Circ.*, 51 (1907), 305.

<sup>8</sup> S. K. Sass, *JOUR. A. PH. A.*, 1 (1912), 1251.

<sup>9</sup> C. A. Bachman, *Proc. Wis. Ph. A.*, 7 (1886), 44; Anon., *Natl. Drug.*, 40 (1910), 115; A. F. Marquier, *Proc. N. J. Ph. A.*, 39 (1911), 47; O. J. Cloughly, *JOUR. A. PH. A.*, 2 (1913), 1004; T. Latham, *JOUR. A. PH. A.*, 2 (1913), 202; G. W. Beringer, *Am. J. Pharm.*, 85 (1913), 303; Editor, "Notes & Queries," *Drug. Circ.*, 57 (1913), 461.

<sup>10</sup> G. W. Beringer, *Am. J. Pharm.*, 85 (1913), 303; G. C. Diekman, *Drug. Circ.*, 55 (1911), 631; S. A. Sharp, *PROC. A. PH. A.*, 58 (1910), 73.

<sup>11</sup> C. A. Bachman, *Proc. Wis. Ph. A.*, 7 (1886), 44; G. Bachman, *Drug. Circ.*, 57 (1912), 62.

<sup>12</sup> Anon., *Natl. Drug.*, 40 (1910), 115.

<sup>13</sup> W. L. Cliffe, *Drug. Circ.*, 57 (1913), 199.

<sup>14</sup> C. A. Bachman, *Proc. Wis. Ph. A.*, 7 (1886), 44; Anon., *Natl. Drug.*, 40 (1910), 115.

<sup>15</sup> O. J. Cloughly, *JOUR. A. PH. A.*, 2 (1913), 1004.

<sup>16</sup> F. Hemm, *Proc. Mo. Ph. A.* (1906), 52.

<sup>17</sup> W. C. Mittelbach, *Am. J. Pharm.*, 71 (1889), 495.

<sup>18</sup> A. F. Marquier, *Proc. N. J. Ph. A.*, 39 (1911), 47.

<sup>19</sup> W. R. White, *JOUR. A. PH. A.*, 2 (1913), 939.

<sup>20</sup> W. C. Mittelbach, *Am. J. Pharm.*, 71 (1889), 495; A. F. Marquier, *Proc. N. J. Ph. A.*, 39 (1911), 47; W. R. White, *JOUR. A. PH. A.*, 2 (1913), 939; M. B. Chittick, *Pharm. Era*, 57 (1923), 635.

One writer<sup>1</sup> reports the formation of a gas in the elixir of iron, quinine and strychnine.

In the course of the investigation by Mueller,<sup>2</sup> a study of the behavior of solutions of soluble ferric phosphate was made. In order to simplify the problem, and at the same time approach it in a more scientific manner, solutions or suspensions of citric acid, of sodium phosphate and of ferric hydroxide (the ingredients entering into the manufacture of soluble ferric phosphate) were prepared, and also solutions or suspensions of the following combinations were made: (1) ferric hydroxide and citric acid; (2) ferric hydroxide and sodium phosphate; (3) citric acid and sodium phosphate and (4) ferric hydroxide, citric acid and sodium phosphate. Since the effect of light was believed to have some influence upon the precipitation and color changes which occurred, portions of these solutions were exposed to sunlight. In each case where citric acid and ferric hydroxide occurred in the same mixture, the evolution of a gas was observed, and the iron was reduced. Samples of the gas were collected and tests revealed the presence of carbon dioxide. A similar test was made using ferrous hydroxide and citric acid, but no gas was produced. It was the opinion of Mueller that the fundamental difficulty in the elixir of iron, quinine and strychnine phosphates must be looked for in the mixtures of chemicals known as soluble ferric phosphate, and that a part of the difficulty is due to the action of sunlight upon citric acid in the presence of ferric iron.

A review of the literature covering the effect of sunlight on citric acid and other acids in the presence of metallic compounds which are easily reduced, reveals that this type of reaction was first reported by Liebig<sup>3</sup> in 1859, who observed that when tartaric acid was decomposed by manganese dioxide and sunlight "carbonic acid and a copper-reducing substance" were formed, and that under similar conditions malic acid was changed to "acetaldehyde and carbonic acid" and that citric acid was converted into "acetone and carbonic acid."

W. Seekamp<sup>4</sup> found in 1893 that a similar decomposition was produced by sunlight if uranic oxide were added to solutions of citric acid. He explained the decomposition by assuming that acetone, oxalic acid and carbon dioxide were first formed, and that the oxalic acid was subsequently oxidized, producing carbon dioxide and water.

Other investigators have observed similar effects of sunlight upon citric acid, but they are not entirely in agreement as to the products formed. C. Neuberg<sup>5</sup> exposed solutions of a number of organic compounds to sunlight in the presence of uranic oxide and its salts. In the case of citric acid he observed that the uranium compound had been reduced and reported that the solution had a reducing effect upon Fehling's solution and had developed a characteristic fruit-like odor. A. Benrath,<sup>6</sup> in 1912, reported that in the presence of ferric nitrate and nitric acid, citric acid was decomposed by sunlight, and that the reaction mixture was found

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<sup>1</sup> R. E. Terry, *Midland Drug.*, 49 (1915), 108.

<sup>2</sup> Thesis, "The Action of Light on Solutions of Soluble Ferric Phosphate," University of Wisconsin (1915).

<sup>3</sup> *Annalen der Chemie*, 113, 1.

<sup>4</sup> *Ibid.*, 278, 373.

<sup>5</sup> "Catalytic Reactions of Sunlight," *Biochem. Zeitschrift*, 27 (1910), 305.

<sup>6</sup> *Chemisches Centralblatt* (1912), II, 1342.

to contain acetone, acetic acid, hydrocyanic acid, formaldehyde and nitrous oxide. Later, in 1917, the same investigator<sup>1</sup> reported the formation of acetone and some formaldehyde, where ferric alum had been used as the ferric salt. G. Ciamucian and P. Silber<sup>2</sup> reported in 1913 that a 5% solution of citric acid, when placed in contact with oxygen and exposed to sunlight all summer, produced acetone and carbon dioxide, and that they were unable to find any other products of the reaction. According to J. Plotnikow,<sup>3</sup> if oxygen is present and an iron salt is used, citric acid produces acetone, carbon dioxide and water.

#### EXPERIMENTAL.

*I.*—In order to determine whether citric acid would completely dissolve ferric hydroxide, when the two were in molecular proportions, producing a solution of normal ferric citrate, a solution of ferric chloride was prepared which upon analysis was found to contain 0.21244 Gm. per cc. A solution of citric acid was also prepared, containing 0.20636 Gm. of the crystalline acid per cc. A sufficient quantity of the solution of ferric chloride to represent 12.2061 Gm. of ferric hydroxide (which is the chemical equivalent of 24 Gm. of citric acid) was precipitated with ammonia water, and washed until free from chlorides and the washings failed to give an alkaline reaction. A sufficient quantity of the standard solution of citric acid to represent 24 Gm. of citric acid was then added to the magma of ferric hydroxide, and the mixture diluted to approximately 1500 cc. After standing 24 hours, a considerable quantity of undissolved ferric hydroxide remained. Exactly 1 Gm. of citric acid was added each day until the magma finally was completely dissolved, and the solution finally contained 27 Gm. of citric acid instead of the 24 Gm. originally added. This shows that ferric hydroxide does not dissolve completely in solutions of citric acid, when the two are present in exactly molecular proportions, and that an excess of the acid must be present to effect solution.

*II.*—In order to determine the effect of the excess of citric acid upon the reaction of the solution when exposed to sunlight, a solution was prepared containing 24 Gm. of citric acid per liter, and a quantity of ferric hydroxide considerably less than the weight chemically equivalent to the citric acid. (The ferric hydroxide was obtained from approximately 12 Gm. of solid ferric chloride, which was roughly weighed. Because of the original moisture contained in the salt, and that absorbed during the weighing, the quantity of ferric hydroxide represented was correspondingly less than the weight chemically equivalent to the citric acid used.) Six portions of 25 cc. each of the solution were divided into three groups of two each. The first two were exposed to sunlight in the original condition. Another two portions were rendered neutral in reaction by the addition of ammonia water, and to the third group, just half of the ammonia water required for neutralization was added. The six portions were each held over chloroform in 100-cc. Lunge nitrometers, and exposed to sunlight for two days.

<sup>1</sup> *Jour. für prak. Chem.*, 204, 190.

<sup>2</sup> *Chemisches Centralblatt* (1913), II, 136.

<sup>3</sup> *Lehrbuch der Photochemie*, 1920, 538.

		Cc. of gas.	Appearance of solution.
Acid solutions	I	18.5	Colorless
	II	20.3	Colorless
Half-neutralized solutions	I	9.7	Colorless
	II	9.8	Colorless
Neutral solutions	I	none	Some color retained, but lighter
	II	none	Some color retained, but lighter

This shows that the extent of the reaction is dependent upon the excess citric acid in the solution, since the quantity of gas collected varies almost in direct proportion with the quantity of free acid, and a neutral reaction the decomposition of the citric acid.

*III.*—When the gas collected in Experiment II was passed through lime water, the liquid became cloudy, showing the presence of carbon dioxide. When another portion of the gas was passed into a strong solution of potassium hydroxide, it was completely absorbed, showing that carbon dioxide was the only constituent of the gas. Upon distillation of the exposed solution, the distillate was found to give a pronounced iodoform reaction, indicating the presence of acetone. This solution was also tested for formaldehyde, since Benrath<sup>1</sup> had reported the formation of formaldehyde in a somewhat similar reaction mixture. The resorcinol test and the ring test, employing sulphuric acid and a 1% aqueous solution of phenol, were used, and both gave negative results.

The undistilled portion of the exposed solution was to be examined principally for oxalic acid and other acids, since Seekamp<sup>2</sup> explained the reaction by assuming that oxalic acid was produced as an intermediate product as a result of the decomposition of the citric acid. Upon treatment of the solution with hydrogen sulphide to precipitate the iron, the resulting sulphide was found to pass directly through the filter and appeared to be colloidal in character. Evaporation of the mixture to a small volume failed to overcome this difficulty, and it was found to be impossible to obtain a clear solution free from iron for these tests. However, other experiments to be described bear out the assumption that carbon dioxide, acetone and water are the only products formed by the decomposition of the citric acid.

*IV.*—Four 100-cc. portions, A, B, C and D, of the original solutions used in the experiment described in Section II were placed in flasks so arranged that any gas evolved would displace an equal volume of water which would flow into a graduated cylinder and could be roughly measured. Portions A and B were exposed in 300-cc. Pyrex glass flasks, and 240-cc. quartz flasks were used for C and D. After exposure to sunlight for two days, they appeared to be completed, as no further increase in the volume of gas was observed upon additional exposure. The total volumes obtained were:

A (glass)	B (glass)	C (quartz)	D (quartz)
110 cc.	70 cc.	100 cc.	105 cc.

At this point the solutions were colorless, indicating the reduction of the iron. It was observed, however, that after opening the flasks and allowing the solutions to

<sup>1</sup> *Chem. Centralblatt* (1912), II, 1342.

<sup>2</sup> *Annalen der Chemie*, 278, 373.

stand in diffused light for several days they gradually regained a brown color. In order to determine whether such solutions, containing ferric iron as a result of reoxidation upon contact with air, would react a second time, the solutions were exposed again. While the total volumes of gas show a wide variation, it is clear that all four portions reacted. This would seem to indicate that the end-point of the reaction was brought about by the complete reduction of the iron, rather than by the exhaustion of the excess citric acid.

Cc. gas.	Cc. gas.	Cc. gas.	Cc. gas.
A (glass)	B (glass)	C (quartz)	D (quartz)
790	315	170	370

Continued exposure failed to give higher readings. It is probable that the apparatus used for Solutions B and C were defective, and permitted escape of some of the gas during the early stages of the reaction.

V.—In order to determine the ratio of carbon dioxide to the acetone produced, a solution was prepared containing 49.2645 Gm. of citric acid per liter and a quantity of ferric hydroxide considerably less than that weight chemically equivalent (in molecular proportions). Four portions of 240 cc. each (the capacity of the quartz flasks) were exposed to sunlight in the apparatus described in Section IV, two portions, A and B in Pyrex glass flasks and the other two, C and D in quartz flasks. At the end of a week the reaction had apparently gone to completion. The volumes of gas collected were as follows:

Cc. gas.	Cc. gas.	Cc. gas.	Cc. gas.
A (glass)	B (glass)	C (quartz)	D (quartz)
1510	1500	1180	1190

The solutions were then distilled and the acetone determined by Goodman's variation of the Messenger Method.<sup>1</sup> For the purpose of comparison, the theoretical volume of carbon dioxide has been calculated from the acetone, assuming that three molecules are produced for each molecule of acetone. These calculations are based upon a pressure of 760 mm. of mercury, and at the temperature of 30° C., which was assumed to be the average.

	Gm. citric acid by analysis.	Gm. acetone by analysis.	Eq. to citric acid, Gm.	Eq. to CO <sub>2</sub> at 30° C., cc.	CO <sub>2</sub> actually collected.
A	11.8245	1.1015	3.9846	1415.28	1510
B	11.8245	1.0885	3.9425	1398.58	1500
C	11.8245	0.8937	3.2371	1148.28	1180
D	11.8245	0.9037	3.2732	1032.13	1190

In each case the actual quantity of carbon dioxide obtained exceeds slightly the theoretical quantity. This may be explained by the fact that these experiments were undertaken during very hot weather, and the temperature in the direct sunlight was no doubt at times higher than the assumed 30° C. Any expansion of the gas would cause a corresponding volume of water to be displaced and in this way increase the reading. Since the theoretical volume of carbon dioxide is calculated with the assumption that three molecules are produced for each molecule of acetone, the results obtained give confirmation of the correctness of this view.

<sup>1</sup> *J. Am. Chem. Soc.*, 42, 39.

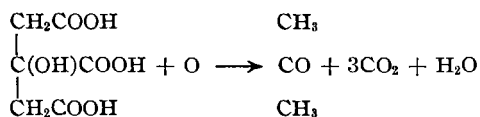
VI.—An attempt was made to check the accuracy of the gasometric determination of the carbon dioxide by absorbing it in a strong solution of potassium hydroxide as it was produced by the reaction. The solution was to be titrated with acid, using phenolphthalein and methyl orange as indicators. A series of twelve solutions were prepared and exposed as before, but in all cases the results obtained in the determination of the carbon dioxide showed such wide variation and were so low that they have been omitted here. The method chosen is apparently applicable only where small quantities of carbonate are present in the solution of the alkali. However, the quantity of acetone produced was carefully determined for each sample, and the equivalent quantity of citric acid calculated by the factor  $\frac{210.11}{58.06}$  or 3.6188. In the solutions used for these experiments, the iron content was known. The assumption that the iron is reduced to the ferrous condition and forms ferrous citrate,  $\text{Fe}_3(\text{C}_6\text{H}_5\text{O}_7)_2$ , permits a comparison of the quantity of citric acid left unchanged to the quantity of hydroxide contained in the solution, in order to determine whether the amount of ferric hydroxide is sufficient, when reduced, to hold the citric acid in combination as ferrous citrate. The ratio  $\frac{3\text{Fe}(\text{OH})_3}{2(\text{H}_3\text{C}_6\text{H}_5\text{O}_7 - \text{H}_2\text{O})}$  is  $\frac{320.58}{420.22}$  or 0.76287. This factor permits calculation from the citric acid of the corresponding weight of ferric hydroxide, which, when reduced, would be required to combine with a given quantity of citric acid as ferrous citrate. These values, together with the quantity of citric acid originally present, the quantity decomposed, and the quantity unchanged are compared with the quantity of ferric hydroxide actually present in the following table. As before, the samples labelled A and B were exposed in Pyrex glass containers, and those labelled C and D were exposed in quartz flasks.

	Gm. total citric acid in 100 cc.	Gm. citric acid decomposed per 100 cc.	Gm. citric acid left unchanged per 100 cc.	$\text{Fe}(\text{OH})_3$ required to combine as $\text{Fe}_3(\text{C}_6\text{H}_5\text{O}_7)_2$ per 100 cc.	$\text{Fe}(\text{OH})_3$ actually present per 100 cc.
I. A	2.7000	1.1671	1.5329	1.1693	1.1421
B	2.7000	1.4266	1.2744	0.9722	1.1421
C	2.7000	1.2883	1.4117	1.0769	1.1421
D	2.7000	1.2713	1.4287	1.0899	1.1421
II. A	1.4000	0.9340	0.4660	0.3555	0.6103
B	1.4000	0.9488	0.4512	0.3442	0.6103
C	1.4000	1.1110	0.2890	0.2204	0.6103
D	1.4000	1.0947	0.3053	0.2329	0.6103
III. A	1.4000	0.6119	0.7881	0.6012	0.6103
B	1.4000	0.5898	0.8102	0.6180	0.6103
C	1.4000	0.5714	0.8286	0.6321	0.6103
D	1.4000	0.5677	0.8323	0.6349	0.6103

It will be noted that in samples IA, IIIB, IIIC and IIID, the reaction has apparently gone beyond the end-point indicated by the formation of ferrous citrate. The only explanation which may be offered for this is the formation of basic salts, while in the other samples, where the quantity of ferric hydroxide present is more than sufficient to form ferrous citrate, it is probable that acid salts are formed.

## SUMMARY AND CONCLUSIONS.

As a result of the experiments conducted, certain conclusions may be drawn. The reaction which occurs appears to be a photochemical decarboxylation and oxidation of the citric acid at the expense of the ferric iron, in accordance with the equation:



In all cases there was a reduction of the iron to the ferrous condition. There was no precipitation observed as reported by Mueller,<sup>1</sup> who found that metallic iron had been produced.<sup>2</sup> The end-point of the reaction is not a very definite one, since standing of the solution in diffused light results in a partial re-oxidation of the iron, which permits resumption of the reaction if the solution be again exposed to sunlight. The extent of the reaction depends upon the excess of citric acid in the solution, and a neutral solution is not affected by sunlight. This seems to make it plausible that only the free acid is attacked, although it should be noted that the reduction of the iron, which results in a lowering of its combining power, liberates free acid in excess of that present as free acid in the original solution.

The results of the experiment described in Section V serve as a basis for the conclusion that three molecules of carbon dioxide are produced for each molecule of acetone. Failure to find formaldehyde in the distillate confirms the belief that carbon dioxide, acetone and water are the only products resulting from the decomposition of the citric acid.

Evidently the light required for this reaction need not necessarily be rich in the ultraviolet. A comparison of the results shows that in a majority of the experiments, the solutions exposed in glass containers produced slightly more acetone and carbon dioxide, when compared volume for volume and for the same length of exposure, than the solutions exposed in quartz.

In so far as this reaction has a bearing upon the elixir of iron, quinine and strychnine, and the elixir of iron, quinine and strychnine phosphates, while the reaction is only one phase of the incompatibility occurring in these preparations, yet certain generalizations may be formulated. The color change in these preparations indicates that the iron has been reduced, and if there is any appreciable quantity of citric acid in excess, and the solution be exposed even to diffused light for a long period of time, there is a possibility of the citric acid gradually breaking down into acetone, carbon dioxide and water. Certainly rendering these preparations acid in order to stabilize them, as has been suggested, would have the opposite effect. On the other hand a slight alkaline reaction will result in the precipitation of the alkaloids and some of the iron. Apparently these preparations require as nearly neutral a reaction as can be obtained. These elixirs should be stored in

<sup>1</sup> Thesis, University of Wisconsin (1915).

<sup>2</sup> It is possible that the reduction of the iron to the metallic state was the result of side reactions induced by the chloroform, or more probably by the mercury in the trap.



amber glass containers, and it is recommended that they be dispensed on prescriptions in the same type of container, instead of the usual prescription bottle.

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A METHOD FOR THE QUANTITATIVE ESTIMATION OF SULPHUR  
COMBINED AS SULPHIDES IN POTASSA SULPHURATA  
(LIVER OF SULPHUR).\*

BY CHARLES H. ROGERS.

"Sulphurated Potassa is a mixture composed chiefly of potassium polysulphides and potassium thiosulphate and contains not less than 12.8 per cent of sulphur (S) in combination as sulphides."—U. S. Pharmacopœia, Tenth Revision.

PREPARATION AND DESCRIPTION OF POTASSA SULPHURATA.

This product is made by thoroughly triturating one part of sublimed sulphur with two parts of potassium carbonate, placing the mixture in a covered crucible, and heating until the mass is in a state of perfect fusion. The melt is then poured upon a slab and when cool is broken into pieces of convenient size and stored in a well-stoppered bottle.

The composition of the product thus obtained varies according to the amount of heat employed. Any of the following reactions may occur.<sup>1</sup>

1.  $3K_2CO_3 + 6S_2 = 2K_2S_5 + K_2S_2O_3 + 3CO_2$ .
2.  $3K_2CO_3 + 4S_2 = 2K_2S_3 + K_2S_2O_3 + 3CO_2$ .
3.  $4K_2CO_3 + 5S_2 = 3K_2S_3 + K_2SO_4 + 4CO_2$ .

Therefore, it will be noted that sulphurated potassa is not a definite chemical compound but a mixture of the polysulphides, the sulphate and the thiosulphate of potassium.

The synonym, liver of sulphur, alludes to the liver-brown color of the freshly prepared material. Upon exposure to air sulphurated potassa absorbs moisture and carbon dioxide and is changed from a brown to a greenish yellow color, and eventually to a gray mass, which consists of potassium carbonate, potassium thiosulphate and potassium sulphate. This decomposition renders it nearly worthless as a therapeutic agent.

ASSAY OF POTASSA SULPHURATA.

U. S. P. X METHOD.

"Dissolve 1 Gm. of crystallized copper sulphate accurately weighed, in 15 cc. of distilled water and 1 Gm. of Sulphurated Potassa accurately weighed in 10 cc. of distilled water, in a flask. Add the copper sulphate solution to the sulphurated potassa solution in a stoppered flask, shake for a few minutes and filter. Acidulate the filtrate with acetic acid, refilter if necessary and add an equal volume of hydrogen sulphide T. S. No black precipitate is produced."

The result obtained by this method of assay indicates only whether the compound is above or below the official requirement. This test does not give results

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\* Scientific Section, A. Ph. A., St. Louis meeting, 1927.

<sup>1</sup> "Principles of Pharmacy," H. V. Army, Second edition, page 412.