been kept in the dark was quite as heavy as in those that had been exposed to the light.

Summary.

1. Magnesium telluride and aluminum telluride were prepared, and the production of hydrogen telluride from aluminum telluride was studied.

2. Hydrogen telluride that was prepared from aluminum telluride was fractionally distilled, and the atomic weight of the tellurium in the fractions was determined by use of the ratio Te : TeO_2.

3. A new form of apparatus for the electrolytic preparation of hydrogen telluride was devised, and about forty grams of the gas were prepared.

4. Hydrogen telluride prepared by electrolysis was fractionally distilled, the fractions were decomposed by passing the gas into nitric acid, and the atomic weight of the tellurium in the several fractions was determined by titration of tellurium dioxide with potassium permanganate.

5. Solidified hydrogen telluride is a white crystalline substance that melts to a liquid of a very pale yellow color. Tellurium dissolves in liquefied hydrogen telluride, and the color of the liquefied gas seems to be due to the presence of this dissolved tellurium.

6. Determination of the melting point of hydrogen telluride gave a result of —57°; the boiling point was found to be 0°.

7. Hydrogen telluride in sealed tubes slowly dissociates. This decomposition of the gas is not materially retarded by the exclusion of light.

8. The results of the investigation indicate that the hypothetical element, dvi-tellurium, does not exist.

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SOME ERRORS IN THE DETERMINATION OF THE RARE EARTHS AS HYDROXIDES.

BY T. O. SMITH AND C. JAMES.

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The writers have observed that, when the rare earths are precipitated as the hydroxides and ignited to oxides, the results are higher than when they are precipitated as oxalates and ignited. Moreover, the former method gives results which are not concordant with varying amounts of the precipitant. Since this method has been often recommended for quantitative determinations it seemed desirable that a careful investigation should be carried out.

As a basis for determining the source of error it was thought advisable to compare the following precipitants: oxalic acid, ammonium oxalate, sodium hydroxide and ammonium hydroxide. If the cause of error were due to the formation of a basic salt, lanthanum being the most powerful base of the rare earths, would tend to give a maximum. It seemed preferable to use lanthanum chloride rather than lanthanum nitrate, since lanthanum hydroxide upon ignition would hold chlorine more readily than nitric acid.

A solution of pure lanthanum chloride was prepared by igniting a very pure lanthanum oxalate, dissolving the resulting oxide in a minimum amount of dilute hydrochloric acid, evaporating to dryness and taking up in water. The solution was made faintly acid, by a few drops of hydrochloric acid, to convert any traces of basic chloride back to the normal salt.

Oxalic Acid.—Varying amounts of the lanthanum chloride solution were measured out and precipitated in the cold by the addition of a slight excess of oxalic acid. The volumes in each case were kept constant so as to remove any doubt as to solubility. The precipitated solutions were allowed to stand over night. The oxalate was filtered off, washed with cold water, dried and ignited.

The results follow:

No.	Cc. $LaCl_3$.	Wt. of oxide.	Oxide per cc.
1	25	0.0866	0.003464
2	25	0.0864	0.003456
3	50	0.1729	0.003458
4	50	0.1732	0.003464
5	50	0.1729	0.003458
6	75	0.2594	0.003459
7	75	0.2596	0.003461
Average			0.003460

Ammonium Oxalate.—In order to avoid the formation of free hydrochloric acid and consequent possible solubility of the oxalate during precipitation, ammonium oxalate was employed as reagent. The operations were identical with the above, except for the substitution of ammonium oxalate for oxalic acid.

The results follow:

No.	Cc. $LaCl_3$.	Wt. of oxide.	Oxide per cc.
1	50	0.1725	0.003450
2	50	0.1726	0.003452
3	25	0.0864	0.003456
4	25	0.0865	0.003460
5	50	0.1729	0.003458
6	50	0.1730	0.003460
7	50	0.1728	0.003456
8	75	0.2594	0.003459
9	75	0.2592	0.003456
Average (omitting 1 and 2)			0.003458

Sodium Hydroxide.—Varying amounts of the lanthanum chloride solution were taken, diluted, precipitated at the boiling point, filtered, washed with hot water, dried and ignited.

The following figures were obtained:

No.	Cc. $LaCl_3$.	Wt. of oxide.	Oxide per cc.
1	25	0.0878	0.003512
2	50	0.1751	0.003502
3	75	0.2619	0.003492
Average			0.003502

The variation of these results with each other suggested precipitation with different amounts of the reagent. Three 50 cc. portions were taken, diluted and precipitated at the boiling point with a measured slight excess of sodium hydroxide solution. Three other 50 cc. portions were taken, diluted, heated to boiling and precipitated with double the amount of sodium hydroxide.

The data follow:

No.	Cc. $LaCl_3$.	Wt. of oxide.	Oxide per cc.
1	50	0.1751	0.003502
2	50	0.1753	0.003506
3	50	0.1755	0.003510
Average			0.003506
4	50	0.1772	0.003544
5	50	0.1771	0.003542
6	50	0.1768	0.003536
Average			0.003541

Ammonium Hydroxide.—Varying amounts of the lanthanum chloride solution were taken, heated to boiling, precipitated with a slight excess of ammonium hydroxide, filtered, washed with hot water, dried, and ignited.

The results follow:

No.	Cc. $LaCl_3$.	Wt. of oxide.	Oxide per cc.
1	25	0.0870	0.003480
2	25	0.0871	0.003484
3	50	0.1740	0.003480
4	50	0.2612	0.003484
5	75	0.2612	0.003483
6	75	0.2616	0.003488
Average			0.003483

When the results of the various precipitants are compared the following facts are brought out: Ammonium oxalate gives figures similar to oxalic acid with the exception that when an excess of ammonium oxalate is used the precipitate carries with it some ammonium oxalate. Upon washing the precipitate with water the complex, ammonium lanthanum oxalate, appears to undergo hydrolysis. When this takes place some lanthanum oxalate is produced in such a fine state of division that it passes through the filter paper. The first two results among the ammonium oxalate determinations illustrate this fact. It was observed that these filtrates were somewhat opalescent.

The amounts of oxide are high when sodium hydroxide is used. This is evidently not due to the formation of a basic salt since when a large amount of sodium hydroxide is used still higher results are obtained. The above facts indicate the carrying down of sodium. In order to prove the presence of sodium in the oxide it was dissolved in hydrochloric acid and the solution tested for sodium by means of the flame coloration and the spectroscope. Both gave very decided tests for sodium.

Ammonium hydroxide was substituted for sodium hydroxide because of its volatile nature. But in this case also high results were obtained although not as high as when sodium hydroxide was used. A qualitative analysis of the lanthanum hydroxide revealed traces of chloride which would indicate the formation of a certain amount of basic chloride. It is highly probable that ammonium hydroxide would give quantitative results if the lanthanum were present as the nitrate because upon ignition the basic nitrate, if formed, would be readily broken down to oxide.

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THE PREPARATION AND PROPERTIES OF SODIUM ZINC CYANIDE.

BY NATHANIEL HERZ.

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Previous Work.—In THIS JOURNAL,[1] W. J. Sharwood gave a summary of previous work, and the results of his own experiments on the preparation of sodium zinc cyanide. Both his results, and those of Rammelsberg, quoted by him, indicate the existence of a compound, $NaZnCy_3.2^1/_2H_2O$, which is very difficult to crystallize, while the mother liquor has the composition Na_2ZnCy_4, agreeing with the compound existing in dilute solutions, and also with the corresponding potassium compound. None of these preparations analyzed, however, gave satisfactory molecular ratios either for the salt, or for the water of crystallization, the crystals being highly efflorescent.

One of the greatest difficulties in this preparation was due to impurities in the sodium cyanide then available. The material then on the market contained appreciable quantities of potassium cyanide and of carbonates,

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[1] Vol. 25, 570–596 (1903), "A Study of the Double Cyanides of Zinc with Potassium and with Sodium."