all the analyses except those mentioned above. While not as rapid as method 1, it was found to be entirely trustworthy, and in addition proved itself far more convenient and rapid than 2. It is therefore recommended as the best of the three methods.

In all the determinations, after adding the sodium bismuthate, the solution was slowly brought to the boiling point, allowed to settle a moment, and then decanted from the residue through a gooch. The residue must be thoroughly washed with 2 per cent. sulphuric acid in order to remove all the cerium. The filtrate is then titrated with the ferrous sulphate solution until an excess has been run in, the excess being titrated back with the permanganate solution.

Duplicate determinations were made on each sample with the following results:

No.	Material.	Source.	Per cent. CeO_2.
1*a* 1*b*	Cerite	Unknown	22.08 21.97
2*a* 2*b*	Monazite	Brazil	28.60 28.74
3*a* 3*b*	Monazite	Brazil	22.84 22.56
4*a* 4*b*	Monazite	Brazil	23.13 23.16
5*a* 5*b*	Monazite	Brazil	23.93 24.01
6*a* 6*b*	Monazite	Brazil	27.25 27.09
7*a* 7*b*	Monazite	Brazil	24.31 24.28
8*a* 8*b*	Monazite	North Carolina	22.19 21.99
9*a* 9*b*	Monazite	North Carolina	22.54 22.49
10*a* 10*b*	Monazite	North Carolina	20.47 20.55

The above results show that the new volumetric method for cerium yields accurate results when applied to monazite and cerite. In addition to its accuracy, the method recommends itself on account of its ease of manipulation and its rapidity.

QUANTITATIVE LABORATORY.

THE REACTION BETWEEN HYDRIODIC ACID AND BROMIC ACID IN THE PRESENCE OF A LARGE AMOUNT OF HYDROCHLORIC ACID.

BY D. L. RANDALL.

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Andrews[1] has shown that when a solution of potassium iodide, strongly

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acidified with hydrochloric acid, is titrated with a standard solution of potassium iodate the reaction takes place according to the following equation:

$$2KI + KIO_3 + 6HCl = 3KCl + 3ICl + 3H_2O$$

When a small amount of chloroform is present the completion of the reaction can be very easily recognized by the fading of the violet color in the chloroform and the amount of iodide present accurately determined from the quantity of iodate used. Andrews also used this method for the determination of free iodine, chlorates, chromates, arsenites and ferrous salts.

At the summer meeting of the American Chemical Society, at Detroit, it was mentioned that very few uses were now made of bromates, in spite of the fact that the salts of bromic acid can easily be made in a condition of great purity. It seemed possible that potassium bromate might be substituted for the more expensive potassium iodate for the determination of the iodides and the other substances worked upon by Andrews.

Experiments were made in which standard potassium bromate solution was added to potassium iodide in a solution that contained about fifty per cent. of strong hydrochloric acid, and it was soon found that a less amount of potassium bromate solution was used than that required by the equation

$$2KI + KBrO_3 + 6HCl = 3KCl + 3H_2O + 2ICl + BrCl.$$

The amount of bromate used was almost exactly two-thirds the quantity required by the above equation. That seemed to indicate that three molecules of potassium iodide instead of two were oxidized by one molecule of bromate, so that the reaction probably goes as follows:

$$3KI + KBrO_3 + 6HCl = 4KCl + 3H_2O + 2ICl + IBr.$$

That the bromate and iodide act in this way is shown by the experiments in which the following solutions were used:

1. A standard solution of potassium bromate made by dissolving 5.5673 grams of the recrystallized salt in one liter of water.
2. An approximately tenth-normal solution of potassium iodide prepared by dissolving 16.607 grams of the commercial C. P. salt in a liter of water.
3. An approximately tenth-normal solution of sodium thiosulphate.

The sodium thiosulphate solution was standardized by treating a known amount of the standard bromate solution with two grams of potassium iodide and five cc. of concentrated hydrochloric acid and titrating the iodine set free with the thiosulphate. The strength of the potassium iodide solution was determined by treating 20 cc. of the iodide solution in a distilling flask with an excess of ferric alum and 5 cc. of sulphuric acid (1 : 1) and distilling over the iodine set free into potassium iodide solution and titrating with the thiosulphate.

In each of the following experiments 20 cc. of the potassium iodide solution in a 250 cc. glass-stoppered bottle were treated with different amounts of hydrochloric acid and five cc. of chloroform. The bromate solution was run in until, after violent shaking, the chloroform had lost its pink color. The results obtained are given in the following table:

A.

KI soln. cc.	HCl (1 : 1). cc.	H_2O. cc.	$KBrO_3$ soln. cc.	I taken. Gram.	I found. Gram.	Error.
20	40	0	19.92	0.2523	0.2530	+0.0007
20	30	10	19.94	0.2523	0.2532	+0.0009
20	20	20	20.00	0.2523	0.2539	+0.0016

B.[1]

KI soln. cc.	Conc. HCl. cc.	H_2O. cc.	$KBrO_3$ soln. cc.	I taken. Gram.	I found. Gram.	Error.
20	40	0	19.87	0.2508	0.2523	+0.0015
20	40	0	19.87	0.2508	0.2523	+0.0015
20	40	0	19.81	0.2508	0.2516	+0.0008
20	20	20	19.87	0.2508	0.2523	+0.0015
20	30	10	19.90	0.2508	0.2527	+0.0019
20	40	0	19.80	0.2508	0.2514	+0.0006
20	40	0	19.85	0.2508	0.2521	+0.0013
20	40	0	19.80	0.2508	0.2514	+0.0006

In all cases a slight positive error could be observed, probably due to the fact that the end point was less sharp with the iodine and chloroform than with the starch iodine indicator.

If a bromide is substituted for an iodide, while a reaction takes place and the chloroform is colored, still complete bleaching is not obtained even by adding an excess of the bromate; so we cannot determine in this way what reaction takes place between bromides and bromates in the presence of strong hydrochloric acid. With ferrous salts there is no coloration of the chloroform at all.

Summary.

This work has shown that like the iodate the bromate will oxidize the iodide, but the bromate is a more powerful oxidizer in that the bromine itself will react with one atom of iodine. Further, iodides may be easily and rapidly determined by titrating the iodide solution with standard bromate solution in the presence of twice its volume of hydrochloric acid.

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COMPOSITION OF SEA WATERS NEAR BEAUFORT, NORTH CAROLINA.

BY ALVIN S. WHEELER.

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Clarke, in his "Data of Geochemistry," reports no analyses of sea waters along the coasts of North America. Possibly no careful analyses have

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[1] Solutions of different strengths than those used in A were taken.