

and the precipitate formed was dissolved in pH 4.0. This NaCl fractionation was repeated twice. The final 5.0 *m* NaCl precipitate was dissolved in water and dialysed until salt-free. The dialysed solution was adjusted to pH 5.7–5.8 and the precipitate⁴ formed was centrifuged off. The supernatant was then made alkaline and adjusted to pH 8.7–8.8, the precipitate⁴ again being removed by centrifugation. The clear fluid was then next brought to 1.65 *m* (NH₄)₂SO₄ at pH 7.0. The (NH₄)₂SO₄ precipitate was dissolved in water and dialysed. The pH and (NH₄)₂SO₄ fractionation was repeated twice. The dialysed solution of the final 1.65 *m* (NH₄)₂SO₄ precipitate was made to pH 5.7–5.8. After the precipitate was removed by centrifugation, the supernatant was adjusted to pH 8.7–8.8. The precipitate formed was centrifuged off and the supernatant fluid brought to pH 6.8–6.9. The resulting precipitate was next dissolved in slightly acid solution and the isoelectric precipitation repeated twice.

The final pH 6.8–6.9 precipitate was examined in a Tiselius electrophoresis apparatus⁵ using the scanning method of Longworth.⁶ Experiments were carried out over a range of pH values from a pH 4.0 to 9.8 at a constant ionic strength of 0.10 using the acetate and borate buffers at 1.5°. In all these experiments the material appeared as a single substance (Fig. 1) with an isoelectric point at pH 6.85.

All biological assays were performed in female rats hypophysectomized at an age of twenty-seven days. Intraperitoneal injections were begun about fourteen days later, once daily for ten days. It was found that 0.010 mg of the hormone daily caused an increase of 10 gm in body weight. On the other hand, a total dose of 5.0 mg of the product did not show lactogenic, thyrotropic, adrenocorticotropic, follicle-stimulating or interstitial-cell stimulating activities, indicating that the preparation was substantially free of other biologically active pituitary contaminants.

Further physico-chemical and biological characterizations of this protein are in progress.

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THE CHEMICAL REMOVAL OF SALTS FROM SEA WATER TO PRODUCE POTABLE WATER¹

No method appears to be known for the chemical removal of salts from sea water; in fact, such a method

⁴ Both the pH 5.8 and pH 8.8 precipitates were found low in growth activity.

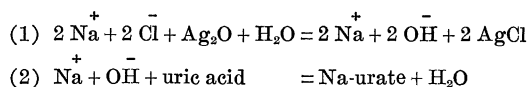
⁵ A. Tiselius, *Trans. Faraday Soc.*, 33: 524, 1937.

⁶ L. G. Longworth, *Jour. Am. Chem. Soc.*, 61: 529, 1939.

¹ The manuscript of this article was received on January 30, 1943. Publication was postponed at the request of the committee on medical research of the Office of Scientific Research and Development.

has been considered impossible by some.² The method herein described should therefore be of interest. The general principle involved is the addition of a sufficient quantity of a suitable base to precipitate the anions present followed by the addition of a sufficient quantity of a suitable acid to precipitate the cations. The method actually used is a special case of this general principle in that both the base (Ag₂O) and acid (uric acid) used as precipitants are practically insoluble; this fact obviates the necessity of accurate measurement of these reagents.

The chemical reactions in the order in which they are carried out may be illustrated as follows:



The AgCl is filtered off before the addition of the uric acid. The solution is again filtered after reaction (2), the filtrate of which is nearly free of dissolved materials (see below). The reactions above are written for the precipitation of NaCl; but most of the other important salts present in sea water appear to react in the same manner. However Ca⁺⁺ and Mg⁺⁺ would also be partly or wholly precipitated as hydroxides in reaction (1). The sulfate ion would not be expected to precipitate since Ag₂SO₄ is fairly soluble.

The following data show the experimental procedure and typical results obtained with artificial (McClendon's) sea water:³

To determine how much of the total solids was urates and uric acid, a 20 cc portion of the filtrate from (3) was acidified with 1 N HCl to pH 3. A white precipitate soon formed. After standing a few

² R. F. Braddish and others, *Jour. Am. Med. Assn.*, 120, 683, 1942.

³ C. G. Rogers, *Textbook of Comparative Physiology*, p. 154. New York: McGraw-Hill, 1927.

Solution	Amount (cc)	pH	Total solids (gms/100 cc)	Remarks
1. Art. sea water	1,000 cc	ca. 7	3.40	solution clear and colorless.
0.35 moles of Ag ₂ O added gradually over a period of about 20 minutes with stirring. Solution filtered immediately. Completion of reaction determined by persistence of brown Ag ₂ O in beaker.				
2. Filtrate from (1)	1,000 cc	10	2.17	solution clear; brown in color.
0.6 moles of uric acid added and solution stirred for about 20 minutes. Completion of reaction determined by noting fall of pH to near neutrality with indicator paper.				
3. Filtrate from (2)	790 cc	ca. 8	0.58	solution clear; sl. brownish.

minutes, the solution was filtered. The total solids determined on the filtrate was 0.45 gms 100 cc.

DISCUSSION

Water yield. About 80 per cent. water recovery was obtained in the above experiment. The unrecovered water was retained in the urate precipitate. Additional water loss would occur if dry Ag_2O was used; for in the above experiment the Ag_2O was made from AgNO_3 and was not completely dry.

Toxicity. There is no reason to believe that the final product would be toxic. The small amount of dissolved materials (0.58 per cent.) is apparently composed of urates (0.13 per cent.) and an undetermined fraction (0.45 per cent.). This latter is probably sulfate (see above); since theoretically there should be 0.41 per cent. sulfate, calculated as Na_2SO_4 , remaining in solution. In molarity, this is 0.029 moles per liter which is of sufficiently low concentration to serve as drinking water.

Taste. The water has a slightly salty taste and is not unpleasant to drink.

SUMMARY

A simple chemical method is described for the removal of most of the salts from sea water. The final product, containing 0.58 per cent. dissolved material which is apparently composed of urates and sulfates, is not unpleasant in taste and is not expected to cause toxic effects if used as drinking water.

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A PRELIMINARY ELECTRON MICROSCOPE STUDY OF THE ACTIVE DEPOSIT FROM RADIOTHORIUM

EARLY investigations of the separation of active substances by the recoil method showed that a small quantity of the parent substance is always present on the collecting plate. One explanation of this phenomenon is that the active deposit consists of aggregates of atoms. When one of the atoms disintegrates by the ejection of an alpha particle the compact aggregate of atoms recoils. Some of the recoil aggregates are deposited on the collecting plate. Lawson¹ found conclusive evidence for the theory in a study of polonium active deposit and named the phenomenon

"aggregate recoil." Chamié² proved the existence of aggregates in the active deposits of thorium, actinium and radium by her radioautograph technique. Harrington³ also found evidence of aggregates in radium active deposit.

After repeating a Chamié radioautograph with an active deposit of thorium, a new technique, that of observation with the electron microscope, was used for a visual investigation of the nature of the active deposit. The usual specimen holder, collodion film supported on 200-mesh copper screen, was the surface on which the active deposit of thorium was collected. The screens were photographed in place in the microscope. The photographs were enlarged to give a final magnification of $80,000\times$ for convenient observation and measurement of particle size.

Samples which had been exposed to the emanations from radiothorium for three, five and eight days were observed. The three-day sample exhibited diffuse spots ranging in size from 20 to 50 millimicrons in diameter. The median diameter was 29 millimicrons. The particle size range of the five-day sample was from 12 to 50 millimicrons with a median diameter of 21 millimicrons. The particle size was much smaller on the eight day sample with a range of 10 to 27 millimicrons. The median diameter was 18 millimicrons. The density of the deposit increased with increasing time of exposure to the emanations. The relative densities expressed in number of particles per square centimeter are 1, 3 and 12 for increasing exposure time. After standing for eight days, the eight-day sample was rephotographed. A deposit of uniform density was observed.

The electron microscope photographs show the aggregates in the active deposit of thorium. The change in the nature of the deposit on standing and the relation of the size distribution of the aggregates to length of exposure to thorium emanations indicates that when an atom of the aggregate disintegrates by the loss of an alpha particle the recoil force is sufficient to cause the aggregate to break up scattering the atoms in all directions.

The same phenomena, though more spectacular, have just been observed for deposits from radon. Fairly large aggregates appear after 30 minutes exposure, while after 1 hour, extraordinary small (maximum about 50 Angstrom units), sharply defined particles appear in the electron micrographs.

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¹ R. W. Lawson, *Nature*, 102: 465, 1919.

² C. Chamié, *Compt. Rend.*, 186: 1838-40, 1928.

³ E. L. Harrington, *Phil. Mag.*, 6: 685-95, 1928.