bases in chloroform and other solvents. Such an investigation is now approaching completion. All of our data (which we hope soon to publish) support the views previously stated.

For the sake of brevity, we give here only the data on certain organic salts of quinine in chloroform solution. In so far as solubility conditions permit, a uniform concentration of quinine has been maintained, and (where possible) the amount of acid has been varied from $^1/_{10}$ to 20 moles per mole of base. Some of the curves are reproduced in diagram 1.

Our conclusions from the data there given are as follows: (1) The curves are of such diversified character that we do not feel justified in calling any one of them normal or anomalous. (2) The curves for the diphenate and the 4,4′-dinitrodiphenate so closely resemble those of certain other salts (*e. g.*, the 2,4-dinitrobenzoate) that there are no grounds for assuming that quinine dissymmetrizes one of these acids but not the others. That all these acids exist in pairs of mirror image forms is highly improbable.

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ANOTHER INTERPRETATION OF EXPERIMENTS WHICH ARE SUPPOSED TO INDICATE THE EXISTENCE OF OXYACIDS OF FLUORINE

Sir:

Dennis and Rochow [THIS JOURNAL, **54,** 832 (1932); *ibid.*, **55,** 2431 (1933)] found that the action of fluorine produced, in an alkaline solution, an oxidizing agent, a small part of which could not, under certain circumstances, have been a peroxide or ozonate. They concluded, therefore, that the salt of an oxyacid of fluorine had been formed. However, there is another plausible interpretation of their observations. Of their results, the most indicative was that the passage of fluorine through cold 50% potassium hydroxide left in the solution an oxidizing agent, a part of which was not destroyed either by standing at room temperature for seventy hours or by evaporation of the liquid and subsequent fusion of the alkali in a beaker. Surely a peroxide or ozonate could not have resisted such severe treatment.

Attempts of the author to repeat this experiment have usually met with failure. The substance of considerable oxidizing strength has been decomposed by standing at room temperature for several hours or by boiling the solution for a minute. The destruction has been so complete that an acidified sample of the solution did not give a blue color with starch and potassium iodide for several minutes.

Fusion of potassium hydroxide in a Pyrex beaker has been found to produce a soluble substance capable of liberating iodine from a solution of potassium iodide containing sulfuric acid. This makes it difficult to test for an oxidizing agent which was present in the alkali before fusion.

At times it has been possible to reproduce the observations of Dennis and Rochow, but in each case of this sort the solution was found to contain a small amount of chloride ion. It therefore seems quite possible that the stable oxidizing agent was an oxy salt of chlorine which could result from a trace of chloride in either the potassium hydroxide, the potassium acid fluoride used for the preparation of fluorine or the sodium fluoride employed to remove hydrogen fluoride from the fluorine. Since this type of experiment should be a rather sensitive test for chlorine, it is surprising that any negative results have been obtained.

By means of the customary tests with titanium sulfate and with chromic acid, it has been shown that the oxidizing agent, which is so unstable in a strongly alkaline solution, is actually a peroxide. This confirms the work of Moissan, Fichter and Bladergroen [H. Moissan, "Le Fluor," 1900, p. 228; Fichter and Bladergroen, *Helv. Chim. Acta*, **10,** 549 (1927)].

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BINDING AND INTERCHANGE OF WATER MOLECULES IN A SALT HYDRATE

Sir:

The statement of Fresenius ["Quant. Anal.," 1897, p. 799] that cupric sulfate pentahydrate, when heated near 140° in a current of dry air, yields a residue of monohydrate of constant weight is substantially correct for analytical practice, in spite of the fact that the equilibrium