drated salts has been demonstrated by analyses of twenty-five typical hydrates.

2. The value of the reagent in determining the water content of inorganic desiccants has been verified.

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[CONTRIBUTION FROM THE AMMONIA DEPARTMENT OF E. I. DU PONT DE NEMOURS & COMPANY, INC.]

Analytical Procedures Employing Karl Fischer Reagent. IX. Reactions with Inorganic Oxides and Related Compounds. Oxidation and Reduction Reactions

BY J. MITCHELL, JR., DONALD MILTON SMITH, E. C. ASHBY AND W. M. D. BRYANT

Karl Fischer reagent is known to provide a rapid and highly specific quantitative means of determining water in organic materials and is useful in determining the water content of many hydrated inorganic salts.[1] Its value in inorganic analytical work is potentially great although interference from side reactions is more frequent, and consequently a well-defined knowledge of the nature and stoichiometry of these reactions is essential for its successful use.

In addition to iodine, sulfur dioxide and pyridine the methanol solution of the reagent contains considerable amounts of the pyridine salts of hydroiodic and methylsulfuric acids arising from preparative side reactions.[2] These latter constituents ordinarily have no effect on the titration of water. However, hydroiodic acid, even when buffered by pyridine, has a marked solvent action upon metal oxides and in attacking these compounds normally liberates one mole of water for each oxygen atom removed. (This reaction was verified stepwise in the case of zinc oxide.) This action involves simple neutralization either with or without complex ion formation. If the metal is capable of existence in other valence states then hydrogen iodide may reduce or iodine may oxidize the ion as soon as it is formed with additional effect upon the apparent water titer. Although Karl Fischer reagent is always standardized by titration of solutions of known moisture content, this really means that the iodine content of the solution has been established. Hence it is apparent that disappearance of iodine may measure either water content or reducing ability, while the oxidizing capacity of the material is measured by an increase in iodine content of the solution.[3]

The reaction producing water by the action of hydroiodic acid upon oxygen-containing inorganic compounds is fairly general. The oxides and hydroxides of the alkali and alkaline earth metals react without difficulty. In fact, the alkali salts of weak oxygen acids behave like a mixture of the respective oxides. It is convenient to consider the over-all reaction of Karl Fischer reagent with the oxides and their derivatives rather than the preliminary step involving only hydrogen iodide. For example

$$MgO + I_2 + SO_2 + CH_3OH \longrightarrow MgI_2 + HSO_4CH_3 \quad (1)$$

$$NaOH + I_2 + SO_2 + CH_3OH \longrightarrow NaI + HI + HSO_4CH_3 \quad (2)$$

$$K_2CO_3 + I_2 + SO_2 + CH_3OH \longrightarrow 2KI + CO_2 + HSO_4CH_3 \quad (3)$$

The salt-forming function of pyridine has been ignored in order to simplify formulation. On this basis the regular reaction with water is as follows

$$H_2O + I_2 + SO_2 + CH_3OH \longrightarrow 2HI + HSO_4CH_3 \quad (4)$$

Oxides of less electropositive metals are less reactive although cuprous, zinc, silver and mercuric oxides are readily converted. Aluminum, nickel, cupric and plumbous oxides on the other hand are unreactive. Some weakly acidic oxides of nonmetals (*e. g.*, As_2O_3 and B_2O_3) are reactive but the more strongly acidic oxides do not in general interfere.

The oxidizing–reducing properties of an iodine–iodide couple are well known and it is not surprising to find this action present in Karl Fischer reagent. Reactions of this character are conditioned by the properties of the metal ion present and hence differ from the oxide reactions described above, although the gross reaction may involve both effects

$$MnO_2 + I_2 + 2SO_2 + 2CH_3OH \longrightarrow MnI_2 + 2HSO_4CH_3 \quad (5)$$

If the metal ion is associated with an anion not affected by hydroiodic acid, the reaction is

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[1] Bryant, Mitchell, Smith and Ashby, THIS JOURNAL, 63, 2924 (1941).

[2] Smith, Bryant and Mitchell, *ibid.*, 61, 2408 (1939).

[3] The iodine of the reagent is relatively inactive toward organic compounds; see reference 2.

confined to oxidation–reduction of the cation

$$CuCl_2 \cdot 2H_2O + 3/2I_2 + 2SO_2 + 2CH_3OH \longrightarrow CuI + 2HCl + 2HI + 2HSO_4CH_3 \quad (6)$$

An analogous reaction is given by copper sulfate pentahydrate and presumably by other cupric salts. The oxidizing character of elementary iodine is apparent in the reaction involving stannous chloride

$$SnCl_2 \cdot 2H_2O + 3I_2 + 2SO_2 + 2CH_3OH \longrightarrow SnI_4 + 2HI + 2HCl + 2HSO_4CH_3 \quad (7)$$

While both the dissolution of oxides and the oxidation or reduction of cations can be regarded as interfering with the normal function of Karl Fischer reagent in the determination of water in inorganic compounds, the interference is clear cut and in general stoichiometric, hence it is possible to correct the water results where the reaction is known. Also the use of Fischer reagent for the determination of reactive metal oxides or oxidation–reduction sensitive ions offers definite advantages in the solution of specific analytical problems.

Analytical Results.[4]—Some of the most useful reactions covered in the present research are those involving the hydroxides and weak acid salts of the alkali metals. The hydroxides of lithium, sodium and potassium react quantitatively with Fischer reagent as shown in Reaction 2, and as there is a mole of water formed as an intermediate from each mole of alkali metal hydroxide, the titer is a measure of the combined alkali and moisture contents. The same net titer is also obtained where an excess of glacial acetic acid of known moisture content has been added before titration. The stoichiometry of the reaction with alkali carbonates, bicarbonates, sulfites and pyrosulfites is illustrated by Equation 3 and the following

$$KHCO_3 + I_2 + SO_2 + CH_3OH \longrightarrow KI + HI + CO_2 + HSO_4CH_3 \quad (8)$$

$$Na_2SO_3 + I_2 + CH_3OH \longrightarrow 2NaI + HSO_4CH_3 \quad (9)$$

$$Na_2S_2O_5 + I_2 + CH_3OH \longrightarrow 2NaI + SO_2 + HSO_4CH_3 \quad (10)$$

The tetraborate and metarsenite exemplify a similar type of reaction except that in these cases the acidic oxide is attacked also

$$Na_2B_4O_7 + 7I_2 + 7SO_2 + 19CH_3OH \longrightarrow 2NaI + 12HI + 7HSO_4CH_3 + 4B(OCH_3)_3 \quad (11)$$

$$NaAsO_2 + 2I_2 + 2SO_2 + 2CH_3OH \longrightarrow NaI + AsI_3 + 2HSO_4CH_3 \quad (12)$$

Esterification of the boric oxide in Equation 11 is in agreement with the behavior of the boric acids (*vide infra*), but this tendency is rare among inorganic compounds. The normal and acid alkali salts of sulfuric acid are not affected by Karl Fischer reagent; any titer is derived exclusively from water of hydration. The same is true of the following related salts: sodium hyposulfite, $Na_2S_2O_4$; sodium dithionate, $Na_2S_2O_6 \cdot 2H_2O$; potassium pyrosulfate, $K_2S_2O_7$. A similar degree of stability was observed in the primary and secondary sodium orthophosphates but not in the tertiary. Trisodium phosphate gave qualitative evidence of alkali but no simple relationship could be derived.

The alkaline earth oxides and a number of others are dissolved by the hydroiodic acid of the reagent and give the type of over-all reaction shown in Equation 1. Table I is made up mainly of this type of oxide, with a few examples of the oxidizing-reducing type and also two peroxides.

TABLE I

THE ACTION OF KARL FISCHER REAGENT ON OXIDES

Oxide	Apparent moles of water formed per mole of oxide
Na_2O_2[a]	0.08
Ag_2O	1.04
MgO	0.99
CaO	1.027
BaO_2[a]	0.20
Cu_2O	1.02
CuO (pptd.)	0.006
HgO (yellow)	.986
HgO (red)	1.00
ZnO	1.00
NiO	0.05
PbO	.15
Al_2O_3	.00
As_2O_3[b]	2.78
MnO_2[c]	1.09

[a] Peroxides. [b] Very slow reaction, evidently trending toward 3.00. [c] Reduced by HI. Theoretical including reduction 1.00.

Barium hydroxide octahydrate (slightly efflorescent) showed 1.75 moles apparent water over that required for hydrate water; this is to be compared with a theoretical value of 2.00. Calcium carbonate failed to give a reaction, probably due to its insolubility in the methanol solution. This confirms a previous observation by K. Fischer.[5]

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(4) The analytical procedure and grade of materials are identical with those employed in the preceding paper (Paper VIII). The use of ethylene glycol in place of methanol although often advantageous in working with salt hydrates is not recommended as a suspending medium for oxides.

(5) K. Fischer, *Angew. Chem.*, **48**, 394 (1935).

Boric oxide and its hydrate metaboric acid[6] react as follows

$$B_2O_3 + 3I_2 + 3SO_2 + 9CH_3OH \longrightarrow 6HI + 3HSO_4CH_3 + 2B(OCH_3)_3 \quad (13)$$

and

$$HBO_2 + 2I_2 + 2SO_2 + 5CH_3OH \longrightarrow 4HI + 2HSO_4CH_3 + B(OCH_3)_3 \quad (14)$$

Reaction 15, reported earlier,[7] is included for comparison

$$H_3BO_3 + 3I_2 + 3SO_2 + 6CH_3OH \longrightarrow 6HI + 3HSO_4CH_3 + B(OCH_3)_3 \quad (15)$$

From the class of oxidation and reduction reactions, and in particular reactive anions, sodium arsenate was found to react in the following manner

$$Na_2HAsO_4 + 3I_2 + 4SO_2 + 4CH_3OH \longrightarrow 2NaI + HI + AsI_3 + 4HSO_4CH_3 \quad (16)$$

Sodium thiosulfate reacted with the iodine of the reagent in the usual way.

$$2Na_2S_2O_3 + I_2 \longrightarrow Na_2S_4O_6 + 2NaI \quad (17)$$

Sodium chromate, sodium and potassium dichromates and potassium permanganate were tried. Appreciable reaction was obtained with the chromates but the results varied considerably and no obvious stoichiometric relations could be deduced. The permanganate was relatively unreactive probably due to low solubility, but on standing a slow reaction of unknown nature occurred. Sodium nitrite, however, reacted stoichiometrically as follows

$$NaNO_2 + 1/2I_2 + SO_2 + CH_3OH \longrightarrow NaI + NO + HSO_4CH_3 \quad (18)$$

Zimmermann[8] has shown that hydrogen peroxide reacts with Fischer reagent without producing any interference, probably according to the reaction

$$H_2O_2 + SO_2 \longrightarrow H_2SO_4 \quad (19)$$

No iodine is consumed in this reaction. This behavior is out of the usual pattern according to which sulfur dioxide would be oxidized to pyridine methyl sulfate. It proved possible to confirm the presence of the simple sulfate ion in the product. Sodium peroxide behaves similarly[9] (Table I). There is no interference with barium peroxide but this is partly due to sluggish reaction.

Sodium sulfide is oxidized to a variable extent by the iodine of the reagent, no doubt due to the establishment of the well-known equilibrium

$$H_2S + I_2 \rightleftarrows 2HI + S \quad (20)$$

Greater analytical possibilities appear to be associated with a reduction or oxidation of cations. Here the reaction pattern is more uniform. Cupric sulfate pentahydrate was used by Karl Fischer in establishing the value of his reagent as applied to salt hydrates. Fischer reported a reaction in which four of the five water molecules were titrated. It was assumed that the fifth was too strongly bound to respond to the reagent and no mention was made of reduction. Experiments in this Laboratory with the pentahydrate showed beyond question that four and one-half instead of four moles of water could be titrated. It was then recalled that the cupric ion is quantitatively reduced by hydrogen iodide with the liberation of iodine. The half mole of iodine liberated exactly accounted for the half mole of water short of the theoretical five.

$$CuSO_4{\cdot}5H_2O + 5I_2 + 5SO_2 + 5CH_3OH \longrightarrow CuSO_4 + 10HI + 5HSO_4CH_3 \quad (21)$$

$$CuSO_4 + 2HI \longrightarrow CuI + 1/2I_2 + H_2SO_4 \quad (22)$$

These combined yielded

$$CuSO_4{\cdot}5H_2O + 4^1/_2I_2 + 5SO_2 + 5CH_3OH \longrightarrow CuI + 8HI + 5HSO_4CH_3 + H_2SO_4 \quad (23)$$

Qualitative confirmation of the mechanism was obtained where some reagent exactly "neutralized" by water was brought in contact with anhydrous cupric sulfate. A strong iodine color developed. The behavior of cupric chloride dihydrate was strictly analogous to that of the sulfate (*vide* equation 6). Ferric chloride hexahydrate gave a net titer equivalent to 5.60 moles of water. This is taken as evidence of reduction to the ferrous salt Stannous chloride dihydrate had a titer equivalent to 2.98 moles of water. A mole of iodine is required to oxidize it to the stannic condition, hence the extra 0.98 mole observed (equation 7).

Several oxides capable of oxidation or reduction were tried and the results, wherever conclusive, are recorded in Table I. The oxides, Fe_2O_3, Pb_3O_4 and PbO_2 showed some reaction but the results were evidently not stoichiometric. The fact that cuprous oxide readily dissolves in the reagent while cupric oxide is practically insoluble, has been of value in analyzing mechanical mixtures of the two oxides. The following data

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(6) Prepared in the crystalline condition by heating orthoboric acid with steam at 150° and 1 atm. for four to six hours. From phase diagrams by von Stackelberg, Quatram and Dressel, *Z. Elektrochem.*, **43**, 14 (1937).

(7) Mitchell, Smith and Bryant, THIS JOURNAL, **62**, 6 (1940).

(8) Zimmermann, *Fette u. Seifen*, **46**, 446 (1939).

(9) On prolonged standing with methanol before titration, oxygen is lost by sodium peroxide and the titer with Fischer reagent proportionally increased. This condition should be avoided.

obtained with mixtures of the pure oxides are of interest.[10]

TABLE II

Wt. % oxide (added)		
CuO	Cu_2O	Wt. % Cu_2O (found)
39.8	60.2	60.7
65.5	34.5	34.9
92.0	8.0	7.9

It is evident from a study of the above reactions that although Karl Fischer reagent is a less specific reagent for water in inorganic analysis than in the organic field, this circumstance is not wholly prejudicial. The great majority of interfering compounds follow relatively simple stoichiometric reaction patterns, thus permitting the use of the reagent for the quantitative determination of a variety of inorganic substances.

Summary

1. The nature and stoichiometry of reactions of Karl Fischer reagent with inorganic substances other than water have been investigated.

2. Two distinct classes of reactions have been described: (a) reactions involving the oxygen of metal oxides or salts of oxygen acids, and (b) oxidation or reduction of cations and anions by a modified iodine–iodide couple. A number of examples are given.

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[10] Since microcrystalline oxides occurring as minerals are often much less soluble than some of the less dense synthetic products, selected crystals of cuprite (Cu_2O), pyrolusite (MnO_2 + 2% H_2O) and zincite (ZnO) were ground to pass through a 200-mesh screen and were then analyzed by means of the reagent. Cuprite reacted completely in two hours; pyrolusite required six hours with shaking but reacted completely. The zincite although of good external appearance never exceeded 80% of the theoretical reaction even after forty-eight hours, hence was probably heavily contaminated.

[CONTRIBUTION FROM THE PHYSIOLOGICAL LABORATORY OF CAMBRIDGE UNIVERSITY]

The Kinetics and Rapid Thermochemistry of Carbonic Acid

BY F. J. W. ROUGHTON

In 1924 Hartridge and Roughton[1] developed a method for following the time course of the heat effects in rapid chemical changes. Our method was to drive the two solutions, whose reaction was to be studied, through separate leads into a mixing chamber and thence into an observation tube. The temperature of the running fluid at several known distances from the mixing chamber was measured thermo-electrically, and from these readings together with the speed of fluid flow down the observation tube the temperature change could be plotted against the time of the reaction. Subsequent work[2,3,4] led finally to an improved technique by which it was possible to follow, to within 1–2%, the time course of heat effects over a period of 0 to 0.01 second even when the total temperature change was no greater than 0.01°.

In the present paper I have applied the technique of Bateman and Roughton to the study of the rapid thermochemistry and kinetics of formation of carbon dioxide from bicarbonate by mixture with acid. This is usually supposed to go in two stages, of which the first

$$H^+ + HCO_3^- \longrightarrow H_2CO_3 \qquad (1)$$

should be practically instantaneous, like all similar ionic reactions in water (see Roughton, ref. 2) whereas the second

$$H_2CO_3 \longrightarrow CO_2 + H_2O \qquad (2)$$

has a definite time course, being a molecular reaction. The present research not only confirms the two-stage theory directly, but also yields for the first time the separate heats of reactions (1) and (2), the previous methods of calorimetry having given only the sum of these two heats. In addition, the velocity constant of reaction (2), *viz.*, k_0, is obtained over the range 0–37°. No direct measurements of k_0 at mammalian body temperature were hitherto available; this was an important gap in view of the role of reaction (2) in controlling the speed of elimination of carbon dioxide in the expired air.

Experimental

Preparation of Reagents

Hydrochloric Acid Solutions.—0.03 to 0.12 N by dilution of concd. hydrochloric acid, the final titer being determined by standard alkali.

Sodium Bicarbonate Solutions.—0.03 to 0.1 M by dissolving weighed amounts of sodium bicarbonate in 5 liters water and shaking the solution in a 10-liter bottle with air containing 5% carbon dioxide so as to reduce the carbonate ion content to less than 1% of the bicarbonate ion content, thus practically eliminating subsequent heat effects

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(1) Hartridge and Roughton, *Proc. Cambridge Phil. Soc.*, **22**, 426 (1924).

(2) Roughton, *Proc. Roy. Soc.* (London), **A104**, 376 (1930).

(3) La Mer and Read, THIS JOURNAL, **52**, 3098 (1930).

(4) Bateman and Roughton, *Biochem. J.*, **29**, 2622 (1935).