Abstract of D. Phil. Thesis.

Kinetic and Equilibrium Studies of Positive Halogen Compounds.

by

E. Gelles.

(Balliol College).

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Kinetic studies of the halogenation of ketones indicate that there is a considerable variation in the position of equilibrium of some halogenation reactions.

A quantitative measure of the position of these equilibria has now been obtained from the effect of ketones and other classes of organic compounds on the oxidation potential of halogen-halide electrodes.

Electromotive force measurements have been made on cells of the type

$$Pt,\ X_2,\ X^-\ (Buffer)\ /\ 3.5\ N\ KCl\ /\ (Buffer)\ X^-,\ X_2,\ RH,\ Pt$$

where X_2, X^- represent iodine or bromine and the corresponding anion

RH an organic compound with a reactive C-H or N-H bond.

These measurements have yielded data from which the thermodynamic equilibrium constants of the reaction

$$(1) \qquad RH + X_2 = RX + X^- + H^+ \quad (aqueous)$$

have been calculated.

Full allowance is made for the formation of the trihalide anion, hydrolysis of the halogen, etc. Most measurements were carried out at 25°C , but some were also made at other temperatures, thus yielding heats as well as free energies of halogenation.

The substances studied included 10 ketones or keto-esters, ranging in reactivity from acetone to acetylacetone, 4 mono-halogenated ketonic substances, 4 nitroparaffins, cyanacetic ester, and acetamide.

From these data, the difference in the bond energies of the carbon-hydrogen and carbon-halogen bonds, and of the nitrogen-hydrogen and nitrogen-halogen bonds, has been obtained in the following manner.

The Gibbs free energy $\triangle F$ of the halogenation reaction(1) is the sum of the free energy changes of the steps

$$RH_{aq} \xrightarrow[1]{} RH_g \xrightarrow[2]{} R_g + H_g \qquad R_g + X_g \xrightarrow[3]{} RX_g \xrightarrow[4]{} RX_{aq}$$

$$X_{2_{aq}} \xrightarrow[5]{} X_{2_g} \xrightarrow[6]{} 2X_g \qquad H_g + X_g \xrightarrow[7]{} H^+_{aq} + X^-_{aq}$$

Similarly, the heat of halogenation $\triangle H$ is the sum of the heat content changes of steps 1-7.

The thermochemical quantities associated with all but steps 2 and 3 are accurately known. $\triangle H_2 + \triangle H_3$ is

equal to the difference in dissociation energy of the hydrogen and halogen bonds. This difference can therefore be evaluated from the heat of halogenation. It can also be derived from the free energies of halogenation by calculating the entropy changes in steps 2 and 3.

These bond energy terms are discussed in relation to the molecular structures of the organic compounds.

Calculations on analogous thermochemical cycles have been made in connection with the equilibria of hydrolysis of the halogens in water, and oxygen-halogen bond dissociation energies have been derived.

The equilibrium constants of halogenation have also been considered in terms of the ionisation constants of the hydrogen and halogen bonds.

$$K_H = \frac{(R^-)(H^+)}{(RH)} \qquad \text{and} \qquad K_X = \frac{(R^-)(X^+)}{(RX)}$$

and of the ionisation constant of the halogen itself

$$K_{X_2} = \frac{(X^-)(X^+)}{(X_2)}$$

In this connection the evidence for the existence of halogen cations in aqueous solution is reviewed. The free energies of formation of these cations have been calculated, and their significance in both kinetic and equilibrium problems is discussed.

Kinetic and Equilibrium Studies of Positive Halogen Compounds.

Thesis submitted for the Degree of Doctor of Philosophy in the University of Oxford.

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CONTENTS.

PAGE.

Part II.

Introduction.

Advances in organic structural theory have led to an interest in the more detailed correlation of chemical behaviour with structural changes. This theory is based on both kinetic and equilibrium studies. The former have been made on a large variety of reactions; equilibrium measurements, however, have been almost entirely confined to acid-base systems. The study of the relation of structure to equilibrium properties clearly requires an extension of thermodynamic measurements to non-prototropic reactions. The work described in the following pages in an attempt in this direction.

The acid-base catalysed halogenation of ketonic substances has been the subject of many kinetic studies [1-5] . These are summarised in the latest of a series of papers, in which some further measurements by the present author are recorded [5].

Indications were obtained in this work, that the bromination of ketones went to completion, whereas there appeared to be a considerable variation in the position of equilibrium of some iodinations.

This thesis is primarily concerned with the exploration of equilibria in halogenation reactions; the measurement of equilibrium constants in the cases of numerous organic substrates and for different halogens, and their interpretation in terms of the varying structural factors involved.

The investigation was ultimately extended beyond ketones, to other compounds containing a reactive carbon-hydrogen bond, and also to nitrogen and oxygen compounds.

Part I.

Equilibrium Measurements.

Experimental Method.

The type of equilibrium which is under consideration can be represented as

$$RH + X_2 = RX + X^- + H^+$$

where X_2, X^- is iodine or bromine and the corresponding anion

RH an organic compound with a reactive C-H or N-H bond.

An electrochemical method was chosen for studying these halogenation equilibria ; namely, the measurement of the effect of organic substrates on the oxidation potential of the halogen-halide system.

It is known that even in extremely dilute iodine-iodide and bromine-bromide solutions an accurate redox potential is attained at a platinum electrode (e.g. Jones & Baeckstrom 1934, Jones & Kaplan 1928 [6 & 7]), and the hope was later borne out that the halide electrodes would behave in a reversible manner in presence of the organic substrates.

A possible alternative method would have been the measurement of the colour change of the halogen-halide solutions. The range of halogen concentration which could be studied by this method would be much smaller, and consequently only the more incomplete iodinations would have been suitable. In contrast, the E.M.F. method proved capable

of measuring equilibrium constants over a range of 20 powers of 10. Moreover, the iodination measurements, at least, proved to be of greater accuracy than might have been attained by light absorption measurement with the usual spectrophotometer.

A third possible method, which was however inapplicable to the cases studied, would have been the titration of halogen in systems where the rate of the reaction

$$RX + X^- + H^+ \rightarrow RH + X_2$$

is sufficiently slow to allow estimation of the equilibrium concentration.

The first task in developing the method of measurement was to see whether the halogen electrode would behave reversibly on addition of various ketones. For this purpose preliminary experiments were conducted, using the following system at room temperature

$$Hg \,/\, Hg_2Cl_{2\,(s)}\,,\ 0.1\ KCl.\ /\ KCl\ Bridge/\ I_{2\,(aq.)}\,,\ I^-,\ Pt$$

Three 0.1 N Calomel electrodes were prepared in the usual Ostwald vessels. No special precautions were taken : the electrodes agreed to within 0.2 mv. when placed against each other.

A platinum gauze electrode was inserted into standard iodine solutions in a beaker. The E.M.F. was measured by

means of a Universal Tinsley potentiometer using a standard Weston cell. Standard solutions of malonic ester, acetoacetic ester, and acetylacetone, and of chloracetate and acetate buffers were prepared. E.M.F. measurements were made after addition of known volumes of buffers and ketones to the iodine solutions.

In this manner the important qualitative results were established that steady E.M.F.'s were attainable in these systems, and that the redox potential at the halogen electrode could be used to determine the position of halogenation equilibrium.

After a few preliminary experiments the calomel electrodes were abandoned and the final cell arrangement adopted.

Solution 1. Solution 2.

Pt, X_2X^-, (Buffer) / Salt Bridge / (Buffer), X_2, X^-, RH, Pt

The halogen electrodes are more accurate and reproducible than the calomel electrodes. The measurements is also more accurate because it now depends on the difference between two similar solutions of different concentration.

Both electrode solutions are made up to 0.1 ionic strength, and in view of their similarity the liquid junction potential is negligible. With the present cell arrangement, this potential, if any, is independent of the nature or form

of the liquid junctions, as the components of both electrode solutions are the same.

The two electrode solutions can be made up from equal quantities of stock halogen solution. The E.M.F. of such a cell would be zero. When a known quantity of substrate is added to solution 2. and both solutions are made up to the same volume,

$$\text{E.M.F.} = \frac{RT}{2F}\left[\ln\frac{(X^-)^2_2}{(X_2)_2} - \ln\frac{(X^-)^2_1}{(X_2)_1}\right]$$

where the subscripts refer to solutions 1 and 2.

From the known value of the trihalide equilibrium constant, the concentrations $(X^-)_1$ and $(X_2)_1$ can be calculated.

If the initial concentrations of RH, $(X_2 + X^-_3)$, and $(X^- + X^-_3)$ are a, b, and c respectively, (in Soln 2)

(1) $RH + X_2 \rightleftharpoons RX + X^- + H^+$

$(a-x)\quad (b-x-y)\qquad x\quad (c+x-y)$

(2) $X_2 + X^- \rightleftharpoons X^-_3$

$(b-x-y)\quad (c+x-y)\qquad y$

The E.M.F. of the concentration cell with two halogen solutions at different concentrations, gives the ratio of the halide to

the halogen concentrations, i.e. $\frac{(c + x - y)^2}{(b - x - y)}$

The second equation connecting the unknowns x and y is provided by equilibrium (2) and hence one can solve for x and y .

The equilibrium constant $K = \frac{(H^+)(c + x - y).x}{(a - x)(b - x - y)} \cdot f_{H^+} \cdot f_{I^-}$

The method is fully illustrated on p. 22 .

In order to obtain thermodynamic equilibrium constants from these measurements, the activity coefficients of the various species have to be taken into account. Therefore, all electrode solutions were made up to 0.1 ionic strength by means of an inert salt. The mean activity coefficients are then constant throughout, and can be calculated from the extended Debye-Huckel expression. $f_{\pm}$ is taken as 0.64 at this ionic strength.

Potassium chloride was chosen for this purpose in preference to potassium nitrate etc., because it can be obtained pure more easily. It was established that any specific interaction of the chloride ion with any of the species of the other halogens was negligible at $I = 0.1$.

For example, at 25° C.

Pt, $(I_2 + I^-_3)$	0.0112 M	3.5 M KCl	$(I_2 + I^-_3)$ 0.0028 M ,Pt
$(I^- + I^-_3)$	0.0516 M		$(I^- + I^-_3)$ 0.0204 M

Made up to 0.5 M with inert salt.

E.M.F. (KCl) = 0.0400 volts .

E.M.F. (KNO_3) = 0.0408 volts.

From the difference at this very high ionic strength, it can be ~~xxx~~ concluded that potassium chloride may be used at I = 0.1

Design and Operation of Cell.

The first type of cell considered was a simple arrangement without a salt bridge. It consisted of two electrode compartments connected by a short length of capillary tubing and a three way tap, which allowed each compartment to be separately connected to a central outlet tube. They could thus be filled in turn by suction as far as the tap, with removal of all air pockets via the central tube. A cell of this design was built and given preliminary tests.

However, all accurate measurements were made with a cell of different type. It was of a somewhat similar design to that employed by Everett and Wynne-Jones [8]. The cell is shown in Figs. 2-3.

It consists essentially of the two electrode vessels A, A' connected by a salt bridge compartment. Platinum gauze electrodes C, C' fit into the electrode vessels at 'Quickfit' joints E, E'. These electrode compartments are connected with the salt bridge compartment by thin tubes ending in jets T, T', and controlled by taps R, R' which are of specially wide bore. The jets end in the compartments U, U' which are provided with outlet tubes Y, Y' and are connected to the salt solution reservoir W via tap V.

Figs. 1 (3) are drawn to scale. Two cells of this type were made from monax glass. In one of them, the electrode vessels were provided with outlet tubes X, X'.

The cell is operated as follows. After cleaning, the electrode compartments are filled separately. The electrode solution is poured in, the outlet tubes Y, Y' provided with screw clips are closed, and tap R is opened. Suction applied at W (with tap V open) fills the electrode compartment to the tip of the jet. The tap R is then closed, and the platinum gauze electrode placed in position. The same procedure is then applied to the second compartment.

Tubes Y, Y' are opened and the salt bridge solution now placed in reservoir W is allowed to run in until compartment U, U' are all but filled.

All outlet tubes are then closed and with a head of salt solution in W, tap V and tap R are opened. The salt solution rises up the jet tube and a well-defined and stable liquid junction is established halfway up tube T. Tap R is closed and the procedure repeated with R'. Tap V is then closed.

The cell, which is held in a secure metal casing, can then be lowered into the thermostat to just below the level of the joints E, E'. Taps R, R' are opened while electromotive force readings are taken.

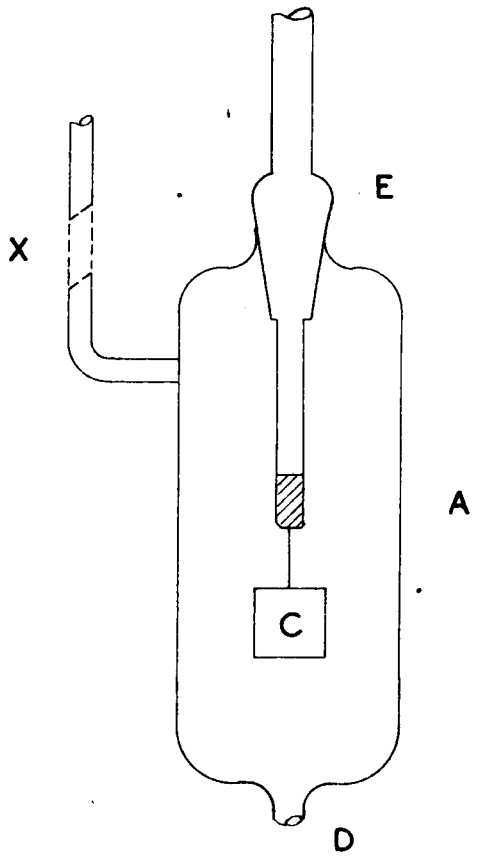
For all measurements, the electrode compartments were washed first with water and then with the solution to be used. All taps were greased with 'Silicone' grease, which is not attacked by halogens. The cells were filled in such a way that the levels in A, A' were equal. The levels in U, U' were also the same and slightly higher than those in the

half-cells. In this way syphoning effects were eliminated.

The platinum gauze electrodes were washed with distilled water and heated in an alcohol flame before use. No variation in E.M.F. was ever observed on interchange of electrodes.

FIGURE 1

HALF CELL. FRONT VIEW



D: TO R AND SALT BRIDGE

1 CM

FIGURE 2

SIDE VIEW

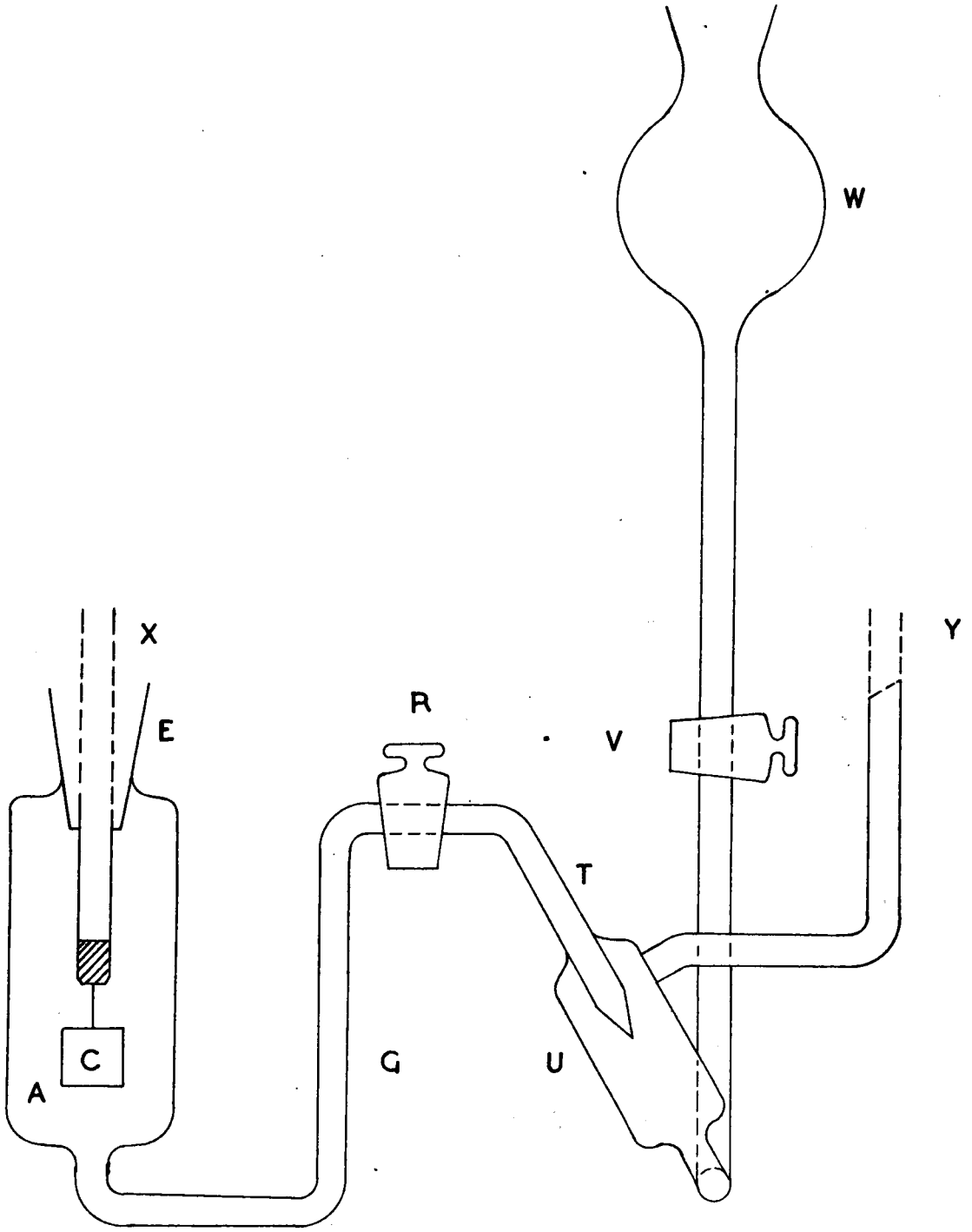
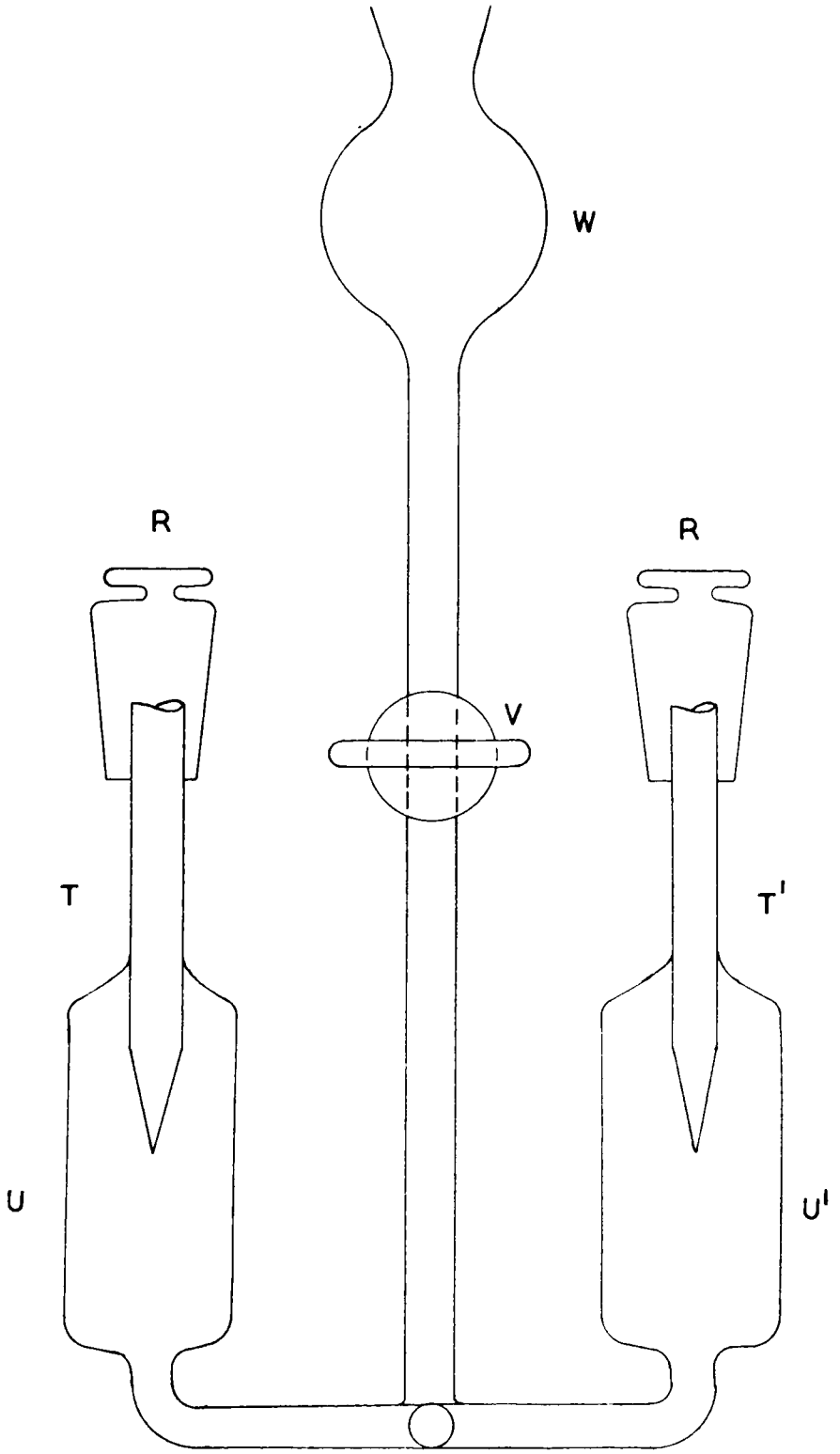


FIGURE 3

BACK VIEW. SALT BRIDGE



Other Apparatus and Materials.

The electromotive force was measured with a Tinsley Potentiometer Type 3387 B by comparison with a recently standardised Weston Cadmium Cell, using a Cambridge galvanometer (450 Ω) Type L 92855 and Universal Shunt. The sensitive galvanometer (1.1 metre deflection at 1 metre for 1 μamp.) was required because of the high resistance of the E,M.F. cell. Readings were taken to 0.1 millivolt.

All measurements were carried out in a water thermostat controlled to ± 0.01°C.

All solutions were made up with water obtained by distillation of laboratory distilled water through a Phoenix glass conductivity water still. Grade A volumetric apparatus was used throughout.

Constant boiling hydrochloric acid was used as the acid-base standard, while for halogen standards pure resublimed iodine and Analar potassium iodate were used.

Thiosulphate was employed for estimation of bromine. Solutions of Analar acetic acid and chloracetic acid and of sodium hydroxide, prepared from pure washed sticks of caustic soda, were used for the preparation of buffers. They were checked against the above standards. Phosphate buffers were prepared from Analar potassium dihydrogen phosphate and standard sodium hydroxide.

Potassium iodide, bromide, and chloride were Analar products.

Ketones, esters, and nitroparaffins were dried and redistilled before use. (Sym-dichloracetone and acetamide were twice recrystallised).

Kinetics of Halogenation of Ketones.

The results of the kinetic studies of the halogenation of ketones [1-5] have a direct bearing on the equilibrium studies in indicating the rate of attainment of the equilibrium.

(1) $RH + X_2 \rightleftharpoons RX + X^- + H^+$

They show that in acid and/or base catalysed halogenations, the rate determining prototropy, e.g.

(2) $RH + Base \rightarrow R^- + Acid.$

is succeeded by the fast halogenation

(3) $R^- + X_2 \rightarrow RX + X^-$

where,

(4) $Acid \rightleftharpoons Base + H^+$

The rate is proportional to the catalyst concentration and independent of the halogen concentration. The catalytic rate constants are related to the dissociation constants of the catalytic species by a linear free energy relation of the Brønsted type. For most systems basic catalysis is more important than acid catalysis.

The rate of reaction (2) can therefore be calculated for a given system, and from (1)-(4) this is seen to be equal to the rate of the forward reaction in (1)

$$RH + X_2 \quad = \quad RX + X^- + H^+$$

Since the backward reaction is always much faster, this forward rate indicates the rate of attainment of equilibrium.

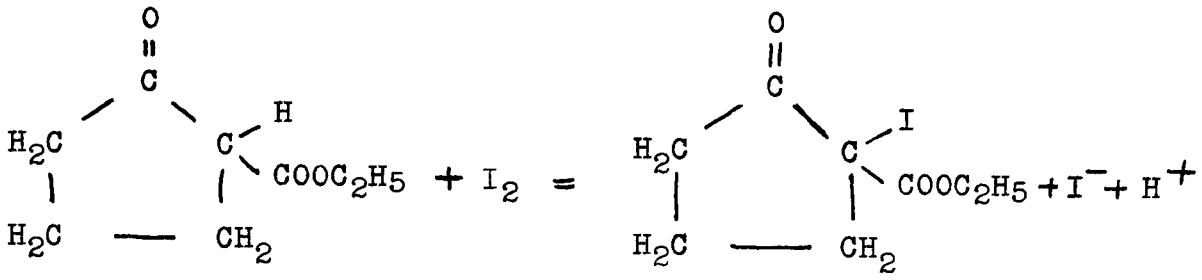
Equilibrium Measurements at 25° C.

The first substances investigated were ketones, as a great deal was already known about their halogenation from kinetics.

Other classes of organic compounds, in which the carbon-hydrogen bond is activated for halogenation by carbonyl, carboxyl, carbethoxy, nitro, and cyanogroups were then studied. The investigation was finally extended to nitrogen compounds.

The Iodination of Ketones.

2- Carbethoxycyclopentanone.



This ketone was selected for the first experiments because it has only one reactive C-H bond, and because a study of its kinetics of bromination indicated that equilibrium would be attained rapidly in acetate or chloracetate buffers [4].

It was prepared according to the method of Linstead and Mead [9] and submitted to chemical purification to remove the adipic ester present to the extent of about 2%. Organic Syntheses Coll. Vol. 2, p.116. Distilled under reduced pressure b. 116-117°/20 mm.

The purity of this ketone and of similar ones was determined by bromination, which goes to completion, and estimation of the bromoketone via iodine

$$R\,Br + 2I^- + H^+ = RH + Br^- + I_2$$

This cannot, however, be done directly, because of the back-iodination of the ketone RH. If the iodide or hydrogen ion concentrations were raised sufficiently to make the above reaction go to completion, considerable atmospheric oxidation

would take place. The procedure adopted for the estimation of the bromoketone was therefore similar to that described elsewhere [4]. It involved the addition of iodide and excess standard thiosulphate; after a sufficient interval a little acid was added and a back-titration carried out with standard iodate.

This method showed the purity of the ketone to increase from 97.4% to 99.3% after chemical purification.

A series of equilibrium measurements was made with this ketone, varying the hydrogen ion concentration over more than four powers of 10, and also varying the ketone, iodine, and iodide concentrations.

The procedure evolved is best illustrated by a Specimen Experiment.

Electrode Soln. 1. Made up from standard iodine-iodide solution and standard potassium chloride to 100 cc.

Electrode Soln. 2. Made up from the same quantity of iodine-iodide solution; chloracetic acid and sodium hydroxide; standard potassium chloride and ketone solution.

Soln. 1. $(I_2 + I_3^-) = 0.00288$ M. Ionic Strength= 0.1

$(I^- + I_3^-) = 0.0204$ M.

Soln. 2. Initial iodine-iodide concentrations as in Soln. 1.

Initial (RH) = 0.00460 M.

Ionic Strength = 0.1

$$pH = 2.66 - \log \frac{(A^-)}{(HA)} = 2.65$$

(corrected for acid produced during halogenation).

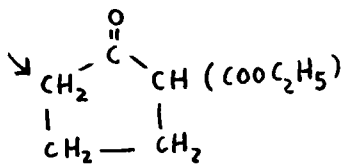
Ketone added to Solution 2 and made up to 100 cc. at 9.00 h. Cell filled and placed in thermostat at 9.05 h.

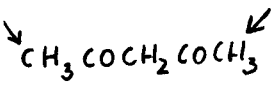
Time	E.M.F. (volts)	Time	E.M.F. (volts)
9.10	0.0357	9.45	0.0353
9.15	0.0356	9.55	0.0352
9.25	0.0355	10.05	0.0350
9.35	0.0354	10.15	0.0349

E.M.F. (t = 0) = 0.0359 volts.

Calculation of the rate of reaction from the known buffer concentrations and catalytic constants shows that equilibrium will be established in less than a minute; electrode equilibrium is almost instantaneous.

The slow change of E.M.F. with time is due to further reaction, such as hydrolytic fission or further slow iodination. Fission of halogenated ketones has been observed in kinetic studies. Further slow iodination can occur at the less active methylene groups in this ketone or at methyl-groups in some other ketonic substances.





In general, the change of E.M.F. with time after the initial mixing of solutions can be divided into two parts : -

(a) the change during the time required for the attainment of the initial equilibrium,

(b) that due to further slow reaction.

The time required for (a) was always consistent with that estimated from kinetic data. The further slow change of E.M.F. (b) , when it occurred at all, was found to be at a rate proportional to that of (a). In confirmation of the view that it is due to further reaction of this kind, it was found that in the case of malonic ester a steady E.M.F. was always reached after the time required for (a). In this substance further reaction (b) cannot of course occur.

For the compounds in which it does occur, the E.M.F. corresponding to the primary iodination equilibrium must be obtained by back-extrapolation.

The expression for the electromotive force of the cell in the above specimen experiment is :

$$E.M.F. = \frac{RT}{2F}\left[\ln\frac{(I^-)_2^2}{(I_2)_2} - \ln\frac{(I^-)_1^2}{(I_2)_2}\right]$$

the subscripts refer to solutions 1 and 2 , and the symbols in round brackets refer to the concentration of species in mols/litre.

At $25^oC.$ $\qquad (I_3^-) = 714\ (I_2)(I^-) \qquad [6]$

Hence, in Solution 1

$$(I^-) = 0.0178 \qquad \text{and} \qquad (I_2) = 0.000210$$

$$(1)\quad 0.0359 = 0.0296\left[\log_{10}\frac{(0.0204+x-y)^2}{(0.00288-x-y)} - \log_{10}\frac{(0.0178)^2}{(0.000210)}\right]$$

In Solution 2

$$(2) \qquad y = 714\ (0.00288 - x - y)(0.0204 - y+x)$$

where $\qquad x = (RI) \qquad$ and $\qquad y = (I_3^-)$

and the concentrations in the equilibrium of halogenation are as shown .

$$\underset{(0.00460 - x)(0.00288-x-y)}{RH} + \underset{}{I_2} \rightleftharpoons \underset{x\ (0.0204-y+x)}{RI} + I^- + H^+$$

From equations (1) and (2) we have

$$x = 0.00252 \qquad \text{and} \qquad y = 0.00029$$

hence,

$$(R\ I) = 2.52.10^{-3}$$

$$(R\ H) = 2.08.10^{-3}$$

and $\qquad (I^-) = 2.26.10^{-2}$

Using the value for (I^-) in equation (1)

gives $(I_2) = 2.10.10^{-5}$

$(H^+) = 2.22.10^{-3}$

The product of the activity coefficients $f_{H+} \cdot f_{I^-} = 0.64$ at ionic strength $I = 0.1$.

The equilibrium constant of iodination

$$K = \frac{(H^+)(I^-)(RI)}{(RH)(I_2)} \cdot f_{H+} \cdot f_{I^-} = \underline{1.86}$$

An experiment in an acetate buffer is illustrated below.

Solutions 1 and 2. Initial $(I_2 + I_3^-) = 0.00288$

$(I^- + I_3^-) = 0.0204$

Final $I = 0.1$

Solution 2 Initial $(RH) = 0.00460$

$(CH_3COO^-) = 0.0533$

$(CH_3COOH) = 0.0567$

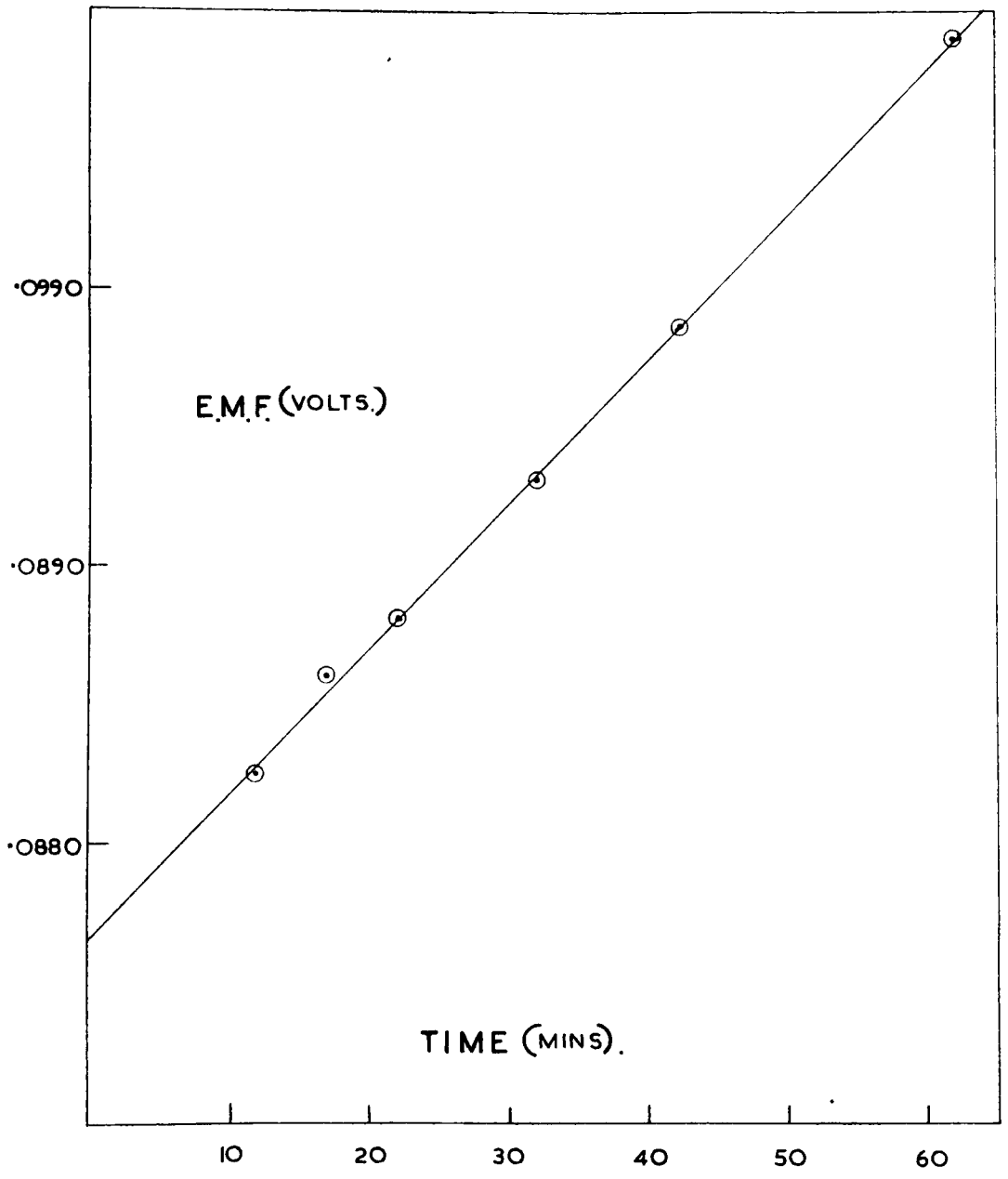
$pH = 4.51 = 4.56 + \log(A^-)/(HA)$

Ketone added to Solution 2 and made up to 100cc 15.50 h.

Cell in thermostat 15.57 h.

FIGURE 4

IODINATION OF 2-CARBETHOXYCYCLOPENTANONE



Time.	E.M.F.	Time.	E.M.F.
16.02	0.0882_5	16.22	0.0893
16.07	0.0886	16.32	0.0898_5
16.12	0.0888	16.52	0.0909

E.M.F. (t = 0) = 0.0877 volts.

Solution of the two simultaneous equations as in the chloracetate buffer experiment gives

$$x = (RI) = 0.00287 \qquad y = (I_3^-) \leq 0.00001$$

hence $(I^-) = 0.0232$

From equation (1) $(I_2) = 3.90.10^{-7}$

(2) $(I_3^-) = 6.46.10^{-6}$

$$K = \frac{(H^+)(I^-)(RI)}{(RH)(I_2)} \cdot f_{H^+}.f_{I^-} = \frac{(3.06.10^{-5})(2.32.10^{-2})(2.87.10^{-3})}{(1.73.10^{-3})(3.90.10^{-7})}$$

$$\times 0.64 = \underline{1.94}$$

Owing to the lower hydrogen ion concentration, practically all the initial iodine is converted into iodoketone, and the tri-iodide concentration is very small. At even lower hydrogen ion concentrations,

$(RI) = (I_2 + I_3^-)$ initial

$(I^-) = (I^- + I_3^-)$ initial + (RI)

whence the calculation reduces to the solution of the E.M.F. equation (1) for one unknown .

Under the conditions of all experiments, iodine hydrolysis is negligible [38] .

The equilibrium is unaffected by the exclusion of oxygen or light.

Table 1 gives some specimen results which illustrate the accuracy of the measurements and the constancy of the above expression over wide ranges of hydrogen ion concentration etc.

Table 1 : The Iodination of 2-Carbethoxycyclopentanone.

Initial Concs.		E.M.F.	Final Concentrations (mols/litre)					K
$(I_2 + I_3^-)$	$(I^- + I_3^-)$	(volts)	(H^+)	(I^-)	(RI)	(RH)	(I_2)	
.00290	.0204	.1141	$3.98.10^{-6}$	$2.33.10^{-2}$	$2.90.10^{-3}$	$1.70.10^{-3}$	$5.15.10^{-8}$	1.96
.00290	.0204	.1053	$7.62.10^{-6}$	$2.33.10^{-2}$	$2.90.10^{-3}$	$1.70.10^{-3}$	$1.02.10^{-7}$	1.91
.00288	.0204	.0877	$3.06.10^{-5}$	$2.32.10^{-2}$	$2.87.10^{-3}$	$1.73.10^{-3}$	$3.90.10^{-7}$	1.94
.00290	.0204	.0509	$5.83.10^{-4}$	$2.31.10^{-2}$	$2.78.10^{-3}$	$1.82.10^{-3}$	$6.89.10^{-6}$	1.91
.00288	.0204	.0359	$2.22.10^{-3}$	$2.26.10^{-2}$	$2.52.10^{-3}$	$2.08.10^{-3}$	$2.10.10^{-5}$	1.86
.00492	.0338	.0383	$2.43.10^{-3}$	$3.77.10^{-2}$	$4.40.10^{-3}$	$7.14.10^{-3}$	$1.93.10^{-5}$	1.88
.00290	.0204	.0294	$4.31.10^{-3}$	$2.22.10^{-2}$	$2.33.10^{-3}$	$2.27.10^{-3}$	$3.40.10^{-5}$	1.86
.00288	.0204	.0148	$2.67.10^{-2}$	$2.05.10^{-2}$	$1.47_5 10^{-3}$	$3.12_5 10^{-3}$	$8.97.10^{-5}$	1.85

Average value = 1.90.

The Iodination of Acetoacetic Ester.

In many ketones more than one reactive C-H bond is present. For example, malonic ester, acetylacetone and acetoacetic ester have a reactive methylene group activated by carboxyl or carbethoxy groups.

Acetoacetic ester was selected for the first attempt to study the successive equilibria.

(1) $RH_2 + X_2 \rightleftharpoons RHX + X^- + H^+$

(2) $RHX + X_2 \rightleftharpoons RX_2 + X^- + H^+$

It was hoped it would be possible to determine the first equilibrium constant in experiments in which the ketone was initially in large excess of the iodine concentration.

If a,b,c are the initial ketone, iodine, and iodide concentrations there will be the following three equilibria

(1) $\underset{(a-x-z)(b-x-y-z)}{RH_2} + \underset{}{I_2} \rightleftharpoons \underset{(x-z)(c+x+z-y)}{RHI} + I^- + H^+$

(2) $\underset{(x-z)\ (b-x-y-z)}{RHI} + I_2 \rightleftharpoons \underset{(\ z\ )(c+x+z-y)}{RI_2} + I^- + H^+$

(3) $\underset{(c+x+z-y)(b-x-y-z)}{I^-} + I_2 \rightleftharpoons \underset{(y)}{I_3^-}$

with concentrations as shown. If the first equilibrium constant could be measured independently as suggested above,

then the three unknowns x, y, z can be obtained from experiments at various ketone and iodine concentrations by solving three simultaneous equations : one involving the measured electromotive force and the other two equations from equilibria (1) and (3). Thus K_2 , the second equilibrium constant could be obtained.

The first experiments were designed to be over a range of hydrogen ion concentration and with an initial ketone-iodine ratio varying from 1-20.

It was found that over this entire range the results calculated on the basis of mono-iodination according to equilibrium (1) gave a constant value for the expression

$$\frac{(RHI)(I^-)(H^+)}{(RH_2)(I_2)} \cdot f_{H^+} \cdot f_{I^-}$$

which is equal to the first equilibrium constant.

A solution of the simultaneous equations for equilibria (1), (2), and (3) with the experimental data showed that

$$K_1 / K_2 > 10^2$$

The second iodination evidently has very much less tendency to occur than the first and, as will be shown later, it was therefore possible to measure it independently.

Experiment with initial ketone greatly in excess of iodine concentration.

Solns. 1 and 2 Initial $(I_2 + I_3^-) = 0.00123$

$(I^- + I_3^-) = 0.00845$

Soln 2 Initial (Ketone) = 0.0242

pH = 5.15_5 $\mathcal{I} = 0.1$

Ketone added to Solution 2 at 14.40 h.

Solution made up to 100 cc. in the usual manner.

Cell in thermostat 14.45 h.

Time	E.M.F.
14.55	0.1353_5
15.10	0.1355_5
15.25	0.1357
15.40	0.1358_5

E.M.F. (t = 0) = 0.1352 volts

Table 2 gives the results of this and other experiments on the first iodination. The hydrogen ion concentration is varied over several powers of 10 and the ketone-iodine ratio from just over 1 to 20.

Table 2 : The Iodination of Acetoacetic Ester.

Initial Concs.		E.M.F.	Final Concentrations (mols/litre)					K
$(I_2 + I_3^-)$	$(I^- + I_3^-)$	(volts)	(H^+)	(I^-)	(RI)	(RH)	(I_2)	
.00123	.00845	.1352	$7.00.10^{-6}$	$9.68.10^{-3}$	$1.23.10^{-3}$	$2.30.10^{-2}$	$8.93.10^{-9}$	.248
.00246	.0169	.0800	$7.45.10^{-6}$	$1.94.10^{-2}$	$2.46.10^{-3}$	$1.17.10^{-3}$	$7.66.10^{-9}$	.254
.00246	.0169	.0994	$3.02.10^{-5}$	$1.94.10^{-2}$	$2.46.10^{-3}$	$2.17.10^{-2}$	$1.69.10^{-7}$	.252
.00123	.00845	.0611	$2.14.10^{-3}$	$9.64.10^{-3}$	$1.21.10^{-3}$	$2.30.10^{-2}$	$2.84.10^{-6}$	.244
.00246	.0169	.0445	$2.33.10^{-3}$	$1.90.10^{-2}$	$2.29.10^{-3}$	$2.19.10^{-2}$	$1.17.10^{-5}$	.250
.00246	.0169	.0350	$2.40.10^{-3}$	$1.87.10^{-2}$	$2.12.10^{-3}$	$9.98.10^{-3}$	$2.39.10^{-5}$	.255

Average value = 0.250.

The Independent Measurement of the Second Iodination Constant.

The evident wider difference in the position of the two equilibria made this independent measurement possible, which is very much simpler than the analysis originally envisaged.

A solution is taken in which there is initially an excess of iodine over ketone. From a knowledge of the first equilibrium constant, it is possible to calculate the pH corresponding to completion of reaction (1)

$$RH_2 + I_2 = RHI + I^- + H^+$$

If this or a higher pH is used, all the ketone is converted into mono-iodoketone, whose further iodination according to reaction (2)

$$RHI + I_2 = RI_2 + I^- + H^+$$

can thus be studied independently.

Specimen Experiment.

Solution 1. $(I_2 + I_3^-) = 0.00420$

$(I^- + I_3^-) = 0.00948$

$I = 0.1$

Solution 2. Initial $(I_2 + I_3^-) = 0.00420$

" $(I^- + I_3^-) = 0.00948$

" (Ketone) = 0.00391

$(H^+) = 6.31.10$ $\qquad I = 0.1$

From the first equilibrium constant $K_1 = 0.250$ we have at the above pH

$$0.250 = \frac{(RHI)}{(RH_2)} \cdot \frac{(I^-)(H^+)}{(I_2)} \cdot 0.64$$

hence

$$\frac{(RHI)}{(RH_2)} \sim 10^3$$

for practical purposes all the ketone can be considered converted to iodoketone

Initial $(RHI) = 0.00391$ $\qquad (I_2 + I_3^-) = 0.00029$

$$(I^- + I_3^-) = 0.0133_9$$

3 experiments gave E.M.F.'s of 0.0712, 0.0711, and 0.0710 volts..

Calculation on the basis of equilibrium (2)

$$\underset{(.00391-x)}{RHI} + I_2 = \underset{(x)}{RI_2} + I^- + H^+$$

Initial $(I_2 + I_3^-) = (RI_2) + (RHI) + (I_2) + (I_3^-)$

gives final concentrations

$(RI_2) = 0.000133$ $\qquad (RHI) = 0.00378$ $\qquad (I_2) = 0.0000149$

$(I_3^-) = 0.000142$ $\qquad (I^-) = 0.0133_8$

$$K_2 = \frac{(RI_2)(I^-)(H^+).0.64}{(RH_2)(I_2)} = \underline{1.31.10^{-4}}$$

The second equilibrium constants were measured in this way for acetoacetic ester, malonic ester, and acetylacetone. They are not as accurate as the first equilibrium constants, for reasons which will be discussed when these measurements are described in more detail.

The Iodination of Other Ketones.

2-Carbethoxycyclohexanone.

At the time of these experiments the kinetics of halogenation of this ketone had not been studied.

The ketone had been prepared by a Dieckmann reaction on Pimelic ester and distilled under reduced pressure.

The equilibrium measurements are shown in Table 3. The E.M.F. reached a steady value after the period required for establishment of halogenation equilibrium. This ketone is therefore less reactive to the further slow change, which was observed with 2-Carbethoxycyclopentanone.

The variation of E.M.F. with time is recorded for a specimen ? experiment, because the rate constant of halogenation can be estimated from the time required for a steady E.M.F. to be reached, and from the known buffer concentrations.

Initial $(I_2 + I_3^-)$ = 0.000861 (Acetate) = 0.0469

$(I^- + I_3^-)$ = 0.00461 (Ketone) = 0.00516

Ketone added at 16.40 h.

Time	E.M.F.	Time	E.M.F.
16.50	0.031	17.05	0.1233
16.55	0.1175	17.10	0.1235
17.00	0.1226	17.15	0.1235
		17.45	0.1235

E.M.F. = 0.1235 volts.

Comparison with carbethoxycyclopentanone shows that the rate of iodination in acetate buffers must be over 100 times faster with that ketone than with carbethoxycyclohexanone. This was later confirmed by independent kinetic measurements in this laboratory.

Table 3 : The Iodination of 2-Carbethoxycyclohexanone.

Initial Concs.		E.M.F.	Final Concentrations (mols/litre)					K
$(I_2 + I_3^-)$	$(I^- + I_3^-)$	(volts)	(H^+)	(I^-)	(RI)	(RH)	(I_2)	
$.000860_6$	.00461	.1235	$2.06.10^{-5}$	$5.47.10^{-3}$	$8.61.10^{-4}$	$4.30.10^{-3}$	$2.86.10^{-8}$	0.505
.000861	.00461	.1184	$1.24.10^{-5}$	$5.47.10^{-3}$	$8.61.10^{-4}$	$1.72.10^{-3}$	$4.25.10^{-8}$	0.512
.000861	.00461	.1195	$2.06.10^{-5}$	$5.47.10^{-3}$	$8.61.10^{-4}$	$3.19.10^{-3}$	$3.90.10^{-8}$	0.500
.000861	.00461	.0182	$2.0\ .10^{-2}$	$4.84.10^{-3}$	$5.05.10^{-4}$	$7.83.10^{-4}$	$7.96.10^{-5}$	0.508

Average value = 0.506.

Malonic Ester.

With this substance, the E.M.F. always reached a perfectly steady value after a time corresponding to the known rate of halogenation. This supports the evidence that the further slow change in E.M.F. observed with some ketones is due to hydrolytic fission of the halogenated ketone or slow iodination of the less reactive methyl or methylene groups. These reactions are impossible in the case of malonic ester.

Measurement of the second equilibrium constant is described below.

Table 4 : The Iodination of Malonic Ester.

Initial Concs.		E.M.F.	Final Concentrations (mols/litre)					K
$(I_2 + I_3^-)$	$(I^- + I_3^-)$	(volts)	(H^+)	(I^-)	(RI)	(RH)	(I_2)	
.00116	.01316	.1007	$4.95.10^{-6}$	$1.43.10^{-2}$	$1.16.10^{-3}$	$1.52.10^{-2}$	$6.65.10^{-8}$	.0522
.00116	.01316	.0906	$4.95.10^{-6}$	$1.43.10^{-2}$	$1.16.10^{-3}$	$7.01.10^{-3}$	$1.46.10^{-7}$	.0514
.00116	.01316	$.0779_5$	$2.92.10^{-5}$	$1.43.10^{-2}$	$1.16.10^{-3}$	$1.52.10^{-3}$	$3.90.10^{-7}$	.0525
.00116	.01316	.0331	$5.08.10^{-4}$	$1.40.10^{-2}$	$1.02.10^{-3}$	$7.15.10^{-3}$	$1.23.10^{-5}$	.0530

Average value = 0.0523.

Malonic acid anion.

A rough estimate was also made of the equilibrium of iodination of this substrate.

$$COO^-CH_2COOH + I_2 = COO^-CH(I)COOH + I^- + H^+$$

The first and second dissociation constants of malonic acid have been measured by Gane and Ingold [10] . They are $1.49.10^{-3}$ and $2.03.10^{-6}$ respectively at 25° C.

Experiments were conducted with malonic acid in acetate buffers.

Specimen Experiment : -

E.M.F.	Final Concentrations (mols/litre)					K
(volts)	(H^+)	(I^-)	(RI^-)	(RH^-)	(I_2)	
.1475	$2.57.10^{-6}$	$6.80.10^{-3}$	$5.58.10^{-4}$	$3.81.10^{-3}$	$1.55.10^{-9}$	1.0

Initial Conc. $(I_2 + I_3^-) = .000558$

$(I^- + I_3^-) = .00624$

Acetylacetone.

The enol contents of the various ketones studied are given in table 22 p. 108a. Only in the cases of carbethoxy-cyclohexanone and acetylacetone do they exceed 0.5%. So far as is known this also applies to the enol content of halogenated ketones.

The equilibrium constants of halogenation are for the equilibrium mixture of (ketone + enol). In the majority of cases, therefore, they are in effect equal to the constants for the ketonic form.

Table 5 shows the equilibrium constant of iodination of acetylacetone to be 1.12. The enol content of acetylacetone is 17% [12] and that of α-monoiodoacetylacetone is estimated to be 7.5% [13]. Hence the equilibrium constant of iodination for the ketonic form of acetylacetone is 1.26.

Table 5 : The Iodination of Acetylacetone.

Initial Concs.		E.M.F.	Final Concentrations (mols/litre)					K
$(I_2 + I_3^-)$	$(I^- + I_3^-)$	(volts)	(H^+)	(I^-)	(RI)	(RH)	(I_2)	
.00116	.01316	.1410	$4.95.10^{-6}$	$1.43.10^{-2}$	$1.16.10^{-3}$	$1.67_5 10^{-2}$	$2.89.10^{-9}$	1.12
.00116	.01316	.1183	$2.92.10^{-5}$	$1.43.10^{-2}$	$1.16.10^{-3}$	$1.67_5 10^{-2}$	$7.69.10^{-9}$	1.10
.00116	.01316	$.0826_5$	$4.84.10^{-4}$	$1.43.10^{-2}$	$1.16.10^{-3}$	$1.67_5 10^{-2}$	$2.71.10^{-7}$	1.12
.00116	.01316	.0632	$2.24.10^{-3}$	$1.43.10^{-2}$	$1.15.10^{-3}$	$1.67_5 10^{-2}$	$1.23.10^{-6}$	1.15

Average value = 1.12.

Sym-Dichloracetone.

No kinetic measurements have been reported for this ketone. In the equilibrium experiments, a steady E.M.F. was attained in each case, and from the time required for this and the known buffer concentrations, the rate of halogenation was estimated to be approximately the same as that of 2-Carbethoxycyclohexanone.

A small drift in the value of the equilibrium constant with pH was observed. This behaviour was not observed with any other substance and no satisfactory explanation can be advanced at present. A steady value is however reached as the pH is lowered.

Table 6 : The Iodination of Sym-Dichloracetone (Preliminary)

Initial Concs.		E.M.F.	Final Concentrations (mols/litre)					K
$(I_2 + I_3^-)$	$(I^- + I_3^-)$	(volts)	(H^+)	(I^-)	(RI)	(RH)	(I_2)	
.00215	$.0115_2$	.1295	$1.01.10^{-5}$	$1.37.10^{-2}$	$2.15.10^{-3}$	$3.20.10^{-2}$	$2.31.10^{-8}$	0.256
.00215	.0115	.1255	$1.32.10^{-5}$	$1.37.10^{-2}$	$2.15.10^{-3}$	$3.20.10^{-2}$	$3.16.10^{-8}$	0.246
.00215	.0115	.1182	$2.22.10^{-5}$	$1.37.10^{-2}$	$2.15.10^{-3}$	$3.20.10^{-2}$	$5.58.10^{-8}$	0.235
.00215	.0115	.1204	$3.53.10^{-5}$	$1.37.10^{-2}$	$2.15.10^{-3}$	$3.20.10^{-2}$	$4.70.10^{-8}$	0.214
.00215	.0115	.1046	$5.77.10^{-5}$	$1.37.10^{-2}$	$2.15.10^{-3}$	$3.20.10^{-2}$	$1.61.10^{-7}$	0.211
.00215	.0115	.0965	$1.08.10^{-4}$	$1.37.10^{-2}$	$2.15.10^{-3}$	$3.20.10^{-2}$	$3.02.10^{-7}$	0.209
.00215	.0115	.0909	$1.61.10^{-4}$	$1.37.10^{-2}$	$2.15.10^{-3}$	$3.20.10^{-2}$	$4.66.10^{-7}$	0.203

Average value = 0.203. ×

× Extrapolated.

Acetone.

The kinetics of the iodination of acetone have been studied in great detail. [e.g. 14] .

From the point of view of the planning of equilibrium measurements, the essential results were that the hydrogen ion was a far more active catalyst than the undissociated acid or anion base catalysts, that especially in buffers the iodination was very slow, and that it seemed to go to completion. It also appeared that in acid solution with a large excess of acetone over iodine, only the monohalogenated derivative would be formed.

It was found that the equilibrium could only be studied in solutions containing an appreciable hydrogen ion concentration i.e. in dilute solutions of mineral acid. In buffers the attainment of equilibrium was so slow as to introduce a number of complications. As even in the acid solutions equilibrium was not established for some days, a change in experimental procedure was adopted.

Hitherto, freshly mixed solutions had immediately been placed in the E.M.F. cell and the electromotive force followed with time. For very slow halogenations it was found both convenient and desirable to make up the experimental mixtures, and to allow them to stand for the time required to reach equilibrium, before being placed in the cell. The necessary time was estimated from initial experiments and from the previously determined rate constants. In this way evaporation and diffusion of solutions, in the cell over long

periods, are avoided.

In connection with these measurements in mineral acid another point must be mentioned. When the hydrogen ion concentration is low, as in the chloracetate, acetate, and phosphate buffers, it is immaterial whether the balancing iodine solution, against which the buffered ketone-iodine mixture is compared, is made up with or without buffer. Exactly the same E.M.F. is obtained in either case.

In mineral acid, however, where there is an appreciable hydrogen ion concentration, the balancing iodine solution always had to be made up to the same hydrogen ion concentration as that existing in the ketone-iodine mixture.

It was found that the cell

$$\text{Pt},\ X_2\ X^-\ (m)\ ,\ 0{\cdot}05\ M\ H^+ \ /\ 3.5\ N KCl\ /\ X_2\ ,\ X^-\ (m)\ ,\ \text{Pt}$$
$$\mu = 0.1 \qquad\qquad \mu = 0.1$$

had an E.M.F. of 0.0011 volts at 25° C., at the same iodine and iodide concentrations as used in the acetone experiments.

This may be due to the incompleteness of dissociation of HI_3 or to a liquid junction potential arising out of the high concentration of hydrogen ion with its very high mobility, whose electrical transport is therefore not inappreciable compared with that of the salt bridge solution.

Acetone from sodium iodide compound was dried before use. Table 7 gives a summary of equilibrium experiments for the reaction

$$CH_3COCH_3 + I_2 = CH_3COCH_2I + I^- + H^+$$

Table 7 : The Iodination of Acetone.

Initial Concs.		E.M.F.	Final Concentrations (mols/litre)					K
$(I_2 + I_3^-)$	$(I^- + I_3^-)$	(volts)	(H^+)	(I^-)	(RI)	(RH)	(I_2)	
.00420	.00947	.1279	$5.12.10^{-2}$	1.36_710^{-2}	$4.20.10^{-3}$	$3.39.10^{-1}$	$1.92.10^{-7}$	28.9
.00383	.00942	.1165	$2.78.10^{-2}$	1.32_510^{-2}	$3.83.10^{-3}$	$8.22.10^{-2}$	$3.59.10^{-7}$	30.0
.00383	.00942	.1050	$3.41.10^{-2}$	1.32_510^{-2}	$3.82.10^{-3}$	$4.42.10^{-2}$	$8.76.10^{-7}$	28.5
.00383	.00942	.1042	$3.41.10^{-2}$	$1.32.10^{-2}$	$3.82.10^{-3}$	$3.92.10^{-2}$	$9.40.10^{-7}$	29.9
.00420	.00947	.1422	$2.77.10^{-2}$	1.36_710^{-2}	$4.20.10^{-3}$	$5.80.10^{-1}$	$6.31.10^{-8}$	27.8
.00420	.00947	.1321	$2.77.10^{-2}$	1.36_710^{-2}	$4.20.10^{-3}$	$2.48.10^{-1}$	1.38_510^{-7}	29.6
.00420	.00947	.1245	$2.77.10^{-2}$	1.36_710^{-2}	$4.20.10^{-3}$	$1.41.10^{-1}$	$2.50.10^{-7}$	28.9
.00420	.00947	.1568	$1.36.10^{-2}$	1.36_710^{-2}	$4.20.10^{-3}$	$8.17.10^{-1}$	$2.04.10^{-8}$	29.9

Average value = 29.2

Cyclohexanone and Cyclopentanone.

The kinetics of halogenation of these ketones have been studied in this laboratory since the equilibrium measurements were made. The latter were essentially similar to the acetone experiments. The case of cyclopentanone requires further investigation. Table 8 shows that for this substance the equilibrium constants vary in an irregular manner.

Table 8 : The Iodination of Cyclopentanone.

Initial Concs.		E.M.F.	Final Concentrations (mols/litre)					K
$(I_2 + I_3^-)$	$(I^- + I_3^-)$	(volts)	(H^+)	(I^-)	(RI)	(RH)	(I_2)	
.000558	.00624	.1225	$3.10.10^{-2}$	$6.80.10^{-3}$	$5.58.10^{-4}$	$6.34.10^{-3}$	$1.09.10^{-8}$	1090
.000558	.00624	.1362	$1.17.10^{-2}$	$6.80.10^{-3}$	$5.58.10^{-4}$	$1.06.10^{-2}$	$3.76.10^{-9}$	720
.000558	.00624	.1194	$4.27.10^{-2}$	$6.80.10^{-3}$	$5.58.10^{-4}$	$1.06.10^{-2}$	$1.31.10^{-8}$	745
.000558	.00624	.1280	$2.28.10^{-2}$	$6.80.10^{-3}$	$5.58.10^{-4}$	$1.06.10^{-2}$	$7.12.10^{-9}$	730
.000558	.00624	.1277	$3.10.10^{-2}$	$6.80.10^{-3}$	$5.58.10^{-4}$	$2.18.10^{-2}$	$7.29.10^{-9}$	475
.000558	.00624	.1316	$3.10.10^{-2}$	$6.80.10^{-3}$	$5.58.10^{-4}$	$4.74.10^{-2}$	$5.38.10^{-9}$	295
.000558	.00624	.1414	$1.28.10^{-2}$	$6.80.10^{-3}$	$5.58.10^{-4}$	$1.14.10^{-2}$	$2.51.10^{-9}$	1090
.000558	.00624	.1452	$6.67.10^{-3}$	$6.80.10^{-3}$	$5.58.10^{-4}$	$2.34.10^{-2}$	$1.87.10^{-9}$	371

Table 9 : The Iodination of Cyclohexanone.

Initial Concs.		E.M.F.	Final Concentrations (mols/litre)					K
$(I_2 + I_3^-)$	$(I^- + I_3^-)$	(volts)	(H^+)	(I^-)	(RI)	(RH)	(I_2)	
.00332	.00963	.1570	$4.15.10^{-2}$	$1.29_5 10^{-2}$	$3.32.10^{-3}$	$9.29.10^{-2}$	$1.07.10^{-8}$	1150
.00332	.00963	.1606	$4.15.10^{-2}$	$1.29_5 10^{-2}$	$3.32.10^{-3}$	$1.33.10^{-2}$	$7.51.10^{-9}$	1150
.00332	.00963	.1631	$2.24.10^{-2}$	$1.29_5 10^{-2}$	$3.32.10^{-3}$	$8.46.10^{-2}$	$6.18.10^{-9}$	1180
.00332	.00963	.1586	$2.24.10^{-2}$	$1.29_5 10^{-2}$	$3.32.10^{-3}$	$5.96.10^{-2}$	$8.78.10^{-9}$	1180
.00168	.00379	.2011	$1.11.10^{-2}$	$5.47.10^{-3}$	$1.68.10^{-3}$	$1.44.10^{-1}$	$3.84.10^{-10}$	1180

Average value = 1170.

The Iodination of Mono-iodoketones.

The equilibria of the type

$$RHI + I_2 \rightleftharpoons RI_2 + I^- + H^+$$

were studied in the cases of iodo-acetoacetic ester, iodo-malonic ester, and iodo-acetylacetone. The method has already been outlined for acetoacetic ester. p. 31.

The unhalogenated ketonic substance was iodinated under such conditions that it was entirely converted to the mono-iodo compound, whose further equilibrium with the di-iodo compound can be measured.

Owing to the instability of the di-iodinated substances, it has not been found possible to obtain very consistent results. They were not improved by the exclusion of oxygen and light.

In some experiments in acetate buffers, only a small proportion of the mono-iodoketone is iodinated. Some experiments were therefore performed in phosphate buffers. In these buffers, however, the greatly increased catalytic power of the anion bases lead to a more rapid secondary drift of E.M.F. with time, and the error in back-extrapolation is much greater.

Specimen Experiments.

1/ The Iodination of Iodoacetoacetic ester in Phosphate Buffer.

Initial Concs. $(I_2 + I_3^-)$	Initial Concs. $(I^- + I_3^-)$	E.M.F. (volts)	Final Concentrations (mols/litre) (H^+)	(I^-)	(RI_2)	(RHI)	(I_2)	K_2
.00111	.0275	.0164	$1.70.10^{-7}$	$2.70.10^{-2}$	$2.02.10^{-4}$	$1.01.10^{-3}$	$4.50.10^{-5}$	$1.30.10^{-5}$
.00111	.0275	.0174	$5.89.10^{-8}$	$2.70.10^{-2}$	$2.56.10^{-4}$	$9.54.10^{-3}$	$4.23.10^{-5}$	$0.65.10^{-5}$
2/ The Iodination of Iodo-acetylacetone.								
.00116	.0132	.0290	$2.95.10^{-6}$	$2.79.10^{-2}$	$1.50.10^{-4}$	$1.64.10^{-3}$	$1.82.10^{-5}$	$2.65.10^{-4}$
.00116	.0132	.0350	$7.24.10^{-7}$	$2.81.10^{-2}$	$2.85.10^{-4}$	$1.51.10^{-3}$	$1.16.10^{-5}$	$2.12.10^{-4}$
.00204	.0112	.0333	$4.00.10^{-6}$	$1.07.10^{-3}$	$7.12.10^{-4}$	$1.08.10^{-3}$	$1.53.10^{-4}$	$1.19.10^{-4}$
.00420	.00947	.0490	$9.33.10^{-6}$	$1.19.10^{-2}$	$1.53.10^{-4}$	$1.80.10^{-3}$	$9.78.10^{-5}$	$6.20.10^{-3}$
3/ The Iodination of Iodomalonic ester.								
.00116	.0132	.0215	$3.60.10^{-8}$	$1.37.10^{-2}$	$3.4.10^{-5}$	$7.83.10^{-4}$	$2.88.10^{-5}$	$4.8.10^{-7}$
.00311	.0102	.0112	$3.66.10^{-8}$	$7.81.10^{-3}$	$9.4.10^{-5}$	$7.23.10^{-4}$	$4.50.10^{-4}$	$5.3.10^{-8}$

The experiments in acetate buffers are more reliable than in phosphate buffers, where there is an appreciable drift of E.M.F. with time, and when fission of the di-iodo-compound or its rearrangement may be more important.

The equilibrium constants are of the order of 10^{-3}, 10^{-4}, and 10^{-7} for iodo-acetylacetone, acetoacetic ester, and malonic ester respectively.

These rough values show that the difficulty in obtaining consistent results cannot be due to disproportion.

$$RI_2 + RH_2 = 2\,RHI$$

for $\dfrac{(RHI)^2}{(RI_2)(RH_2)} = \dfrac{K_1}{K_2} > 10^3$ in each case.

The Bromination of Ketones.

The study of the equilibria of bromination was planned along the same lines as that of iodination.

Differences in experimental detail included the necessity of standardising solutions of bromine (Analar) with thiosulphate at frequent intervals.

Bromine has a lower tendency than iodine to form the trihalide ion.

At 25° C., $(I_3^-) = 714\ (I_2)(I^-)$ and

$(Br_3^-) = 16\ (Br_2)(Br^-)$ [6 & 7]

Hydrolysis of bromine is also more appreciable

At 25° C., $\frac{(H^+)(I^-)(HIO)}{(I_2)} = 3.10^{-13}$ [38-40]

while the corresponding constant for bromine is $5.8.10^{-9}$ [41] However, even at the highest pH studied the hydrolysis is of no importance because of the very low bromine concentrations. The balancing bromine solution 1 was not buffered.

Specimen Experiment : Bromination of 2-Carbethoxycyclohexanone.

Solution 1. $(Br_2 + Br_3^-) = 0.00180_4$

$I = 0.1$

$(Br^- + Br_3^-) = 0.0112_2$

Solution 2. Initial $(Br_2 + Br_3^-) = 0.00180_4$

" $(Br^- + Br_3^-) = 0.0112_2$

" (Ketone) = 0.00324

$(H^+) = 1.04.10^{-4}$ (Acetate Buffer) $I = 0.1$

Ketone added to Solution 2 at 14.15 h.

Cell in thermostat 14.20 h.

Time.	E.M.F.	Time.	E.M.F.
14.25	0.0025	16.30	.4210
14.45	0.0088	16.40.	.4377
15.30.	0.0273	16.50	.4475
16.05	0.3280	17.00	.4510
16.15	0.3963	17.10	.4510
16.25	0.4155	17.30	.4510

E.M.F. = 0.4510 volts.

The time required for attainment of a steady E.M.F. is that expected from the ~~xxx~~ iodination experiments.

In Solution 1. $(Br_3^-) = 16\left(.001804 - (Br_3^-) \right)\left(.01122 - (Br_3^-) \right)$

$(Br_2) = 0.00153$ $(Br^-) = 0.01095$ $(Br_3^-) = 0.000268$

from the E.M.F. :

$$0.4510 = 0.0296 \left[\log \frac{(0.01122 + x - y)^2}{(0.00180_5 - x - y)} - \log \frac{(0.01095)^2}{(0.00153_6)} \right]$$

where $x = (RBr)$ and $y = (Br_3^-)$

as y must be negligible, and $x = 1.80.10^{-3}$

$$\frac{451}{296} - 1.1073 = \log \frac{(0.01122 + x)^2}{(0.00180_5 - x)} .$$

$$\log (Br_2) = \log (0.01302)^2 - 14.1291$$

$$(Br_2) = 1.26.10^{-18}$$

and $(RBr) = 1.80.10^{-3}$ $(RH) = 1.44.10^{-3}$ $(Br^-) = 1.30.10^{-2}$

$$K = \frac{(H^+)(Br^-)(RBr)}{(Br_2)(RH)} \cdot f_{H^+} f_{Br^-}$$

$$\frac{(1.04.10^{-4})(1.30.10^{-2})(1.80.10^{-3}) . 0.64}{(1.26.10^{-18})(1.44.10^{-3})} = \underline{8.6.10^{11}}$$

The reproducibility of experiments in acetate buffers (pH = 4-5) is illustrated by the bromination of acetoacetic ester.

Table 10 : The Bromination of Acetoacetic Ester.

Initial Concs.		E.M.F.	Final Concentrations (mols/litre)					K
$(Br_2 + Br_3^-)$	$(Br^- + Br_3^-)$	(volts)	(H^+)	(Br^-)	(RBr)	(RH)	(Br_2)	
.00180	.0112	.4370	$1.02.10^{-4}$	$1.30.10^{-2}$	$1.80.10^{-3}$	$1.33.10^{-2}$	$3.73.10^{-18}$	$3.08.10^{1}$
.00180	.0112	.4455	$5.62.10^{-5}$	$1.30.10^{-2}$	$1.80.10^{-3}$	$1.33.10^{-2}$	$1.93.10^{-18}$	$3.28.10^{10}$
.00180	.0112	.4510	$3.45.10^{-5}$	$1.30.10^{-2}$	$1.80.10^{-3}$	$1.33.10^{-2}$	$1.26.10^{-18}$	$3.09.10^{10}$
.00180	.0112	.4575	$2.15.10^{-5}$	$1.30.10^{-2}$	$1.80.10^{-3}$	$1.33.10^{-2}$	$7.60.10^{-19}$	$3.18.10^{10}$

The Bromination of 2-Carbethoxycyclopentanone.

This ketone was selected for a detailed study as in the case of the iodination experiments.

Table 11 contains a summary of the experiments extending over a pH range of 4. The "equilibrium constants" vary over $1\frac{1}{2}$ powers of 10 over this range.

The expression $\frac{(H^+)^{4/3}(Br^-)(RBr)}{(Br_2)(RH)}$ seems to have a more constant value over the range studied.

if $$K' = \frac{(H^+)^{4/3}\ (Br^-)(RBr)}{(Br_2)(RH)}\ f_{H^+} \cdot f_{Br^-}$$

K	K'	K	K'
$1.50.10^{13}$	$3.39.10^{12}$	$1.12.10^{14}$	$3.47.10^{12}$
$1.83.10^{13}$	$3.09.10^{12}$	$1.09.10^{14}$	$3.09.10^{12}$
$2.46.10^{13}$	$3.16.10^{12}$	$1.44.10^{14}$	$3.71.10^{12}$
$3.78.10^{13}$	$3.47.10^{12}$	$2.60.10^{14}$	$5.01.10^{12}$
$1.06.10^{14}$	$4.07.10^{12}$	$3.32.10^{14}$	$5.89.10^{12}$

It is seen from Table 11 and Figure 5 that the variation depends on the hydrogen ion concentration. It is possible that at very low bromine concentrations another electrode reaction involving hydrogen ions becomes apparent, and that this reaction gradually increases in importance with decreasing pH and bromine concentration.

In these circumstances a different experimental method, namely the measurement of pH at constant bromine concentration, may be more successful.

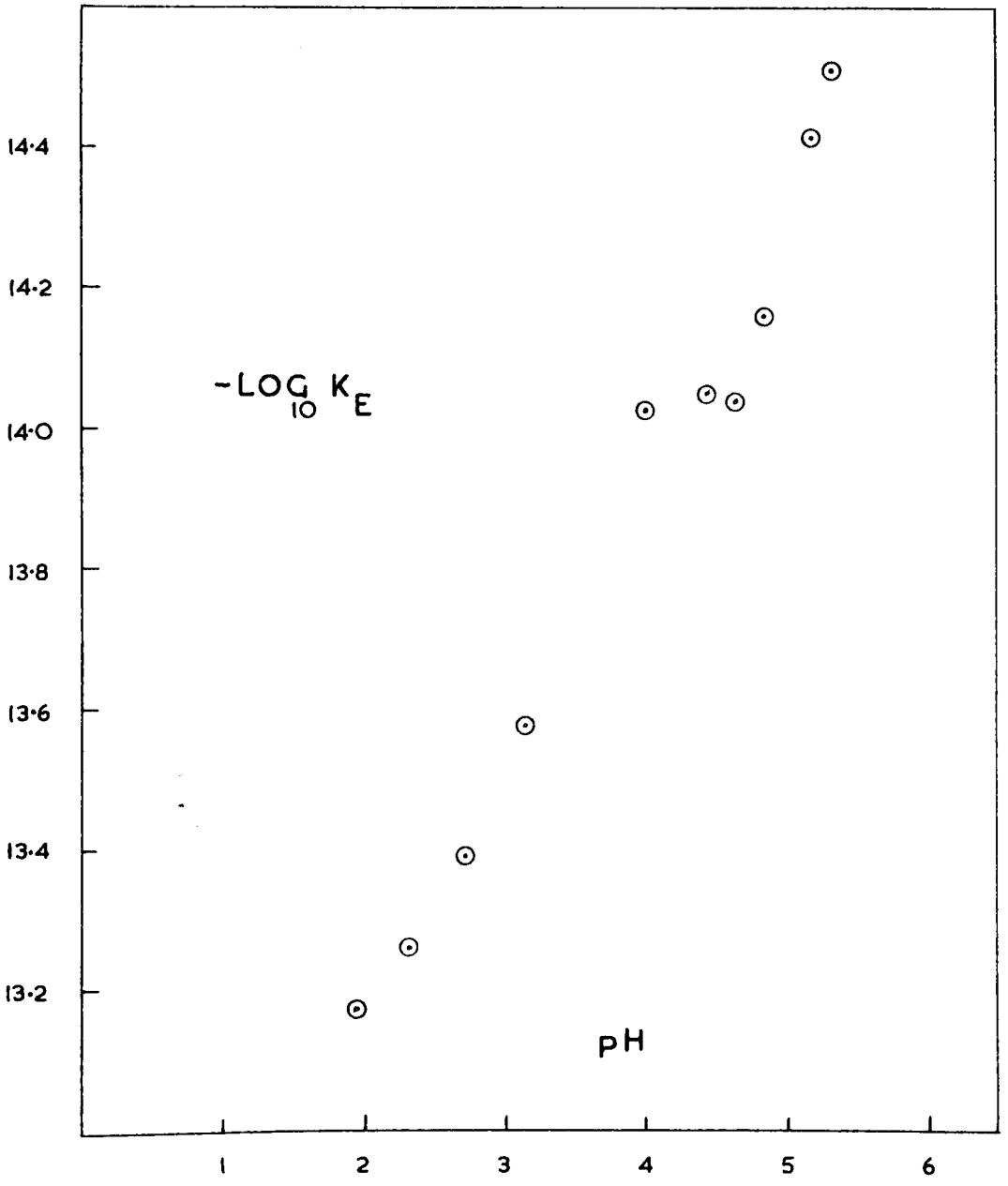
Other bromination experiments were made with malonic ester, bromomalonic ester, acetylacetone, and acetone.

Table 11 : The Bromination of 2-Carbethoxycyclopentanone.

Initial Concs.		E.M.F.	Final Concentrations (mols/litre)					K
$(Br_2 + Br_3^-)$	$(Br^- + Br_3^-)$	(volts)	(H^+)	(Br^-)	(RBr)	(RH)	(Br_2)	
.00178	.0112	.4588	$1.12.10^{-2}$	$1.30.10^{-2}$	$1.78.10^{-3}$	$1.65.10^{-2}$	$6.70.10^{-19}$	$1.50.10^{13}$
.00192	.0113	.4703	$4.90.10^{-3}$	$1.32.10^{-2}$	$1.92.10^{-3}$	$1.43.10^{-2}$	$3.04.10^{-19}$	$1.83.10^{13}$
.00178	.0112	.4886	$1.94.10^{-3}$	$1.30.10^{-2}$	$1.78.10^{-3}$	$1.75.10^{-2}$	$6.65.10^{-20}$	$2.46.10^{13}$
.00192	.0113	.5042	$7.24.10^{-4}$	$1.32.10^{-2}$	$1.92.10^{-3}$	$1.43.10^{-2}$	$2.18.10^{-20}$	$3.78.10^{13}$
.00178	.0112	.5527	$5.64.10^{-5}$	$1.30.10^{-2}$	$1.78.10^{-3}$	$1.75.10^{-2}$	$4.52.10^{-22}$	$1.06.10^{14}$
.00178	.0112	.5595	$3.45.10^{-5}$	$1.30.10^{-2}$	$1.78.10^{-3}$	$1.65.10^{-2}$	$2.77.10^{-22}$	$1.12.10^{14}$
.00178	.0112	.5648	$2.15.10^{-5}$	$1.30.10^{-2}$	$1.78.10^{-3}$	$1.65.10^{-2}$	$1.77.10^{-22}$	$1.09.10^{14}$
.00192	.0113	.5707	$1.44.10^{-5}$	$1.32.10^{-2}$	$1.92.10^{-3}$	$1.32.10^{-2}$	$1.23.10^{-22}$	$1.44.10^{14}$
.00192	.0113	.5900	$6.31.10^{-6}$	$1.32.10^{-2}$	$1.92.10^{-3}$	$1.43.10^{-2}$	$2.75.10^{-23}$	$2.60.10^{14}$
.00192	.0113	.5970	$4.67.10^{-6}$	$1.32.10^{-2}$	$1.92.10^{-3}$	$1.43.10^{-2}$	$2.28.10^{-25}$	$3.32.10^{14}$

FIGURE 5

BROMINATION OF CARBETHOXYCYCLOPENTANONE



The second equilibrium of bromination of malonic ester, which was studied by brominating bromomalonic ester, does not interfere with the measurement of the first bromination equilibrium.

The same difficulties were encountered with all the above ketonic substances as with 2-Carbethoxycyclopentanone. The equilibrium constants are therefore only reliable to the nearest power of 10.

Malonic Ester : Specimen Experiments.

E.M.F.	Final Concentrations (mols/litre)					K
(volts)	(H^+)	(Br^-)	(RBr)	(RH)	(Br_2)	
.441	$2.15.10^{-5}$	$1.30.10^{-2}$	$1.80.10^{-3}$	$1.11.10^{-2}$	$3.29.10^{-18}$	$8.8.10^9$
.452	$2.15.10^{-5}$	$1.30.10^{-2}$	$1.80.10^{-3}$	$2.41.10^{-2}$	$1.16.10^{-18}$	$11.5.10^9$
.409	$1.02.10^{-4}$	$1.30.10^{-2}$	$1.80.10^{-3}$	$2.41.10^{-2}$	$3.30.10^{-17}$	$1.9.10^9$

Initial Conc. $(Br_2 + Br_3^-) = .00180$
$(Br^- + Br_3^-) = .0112$
throughout

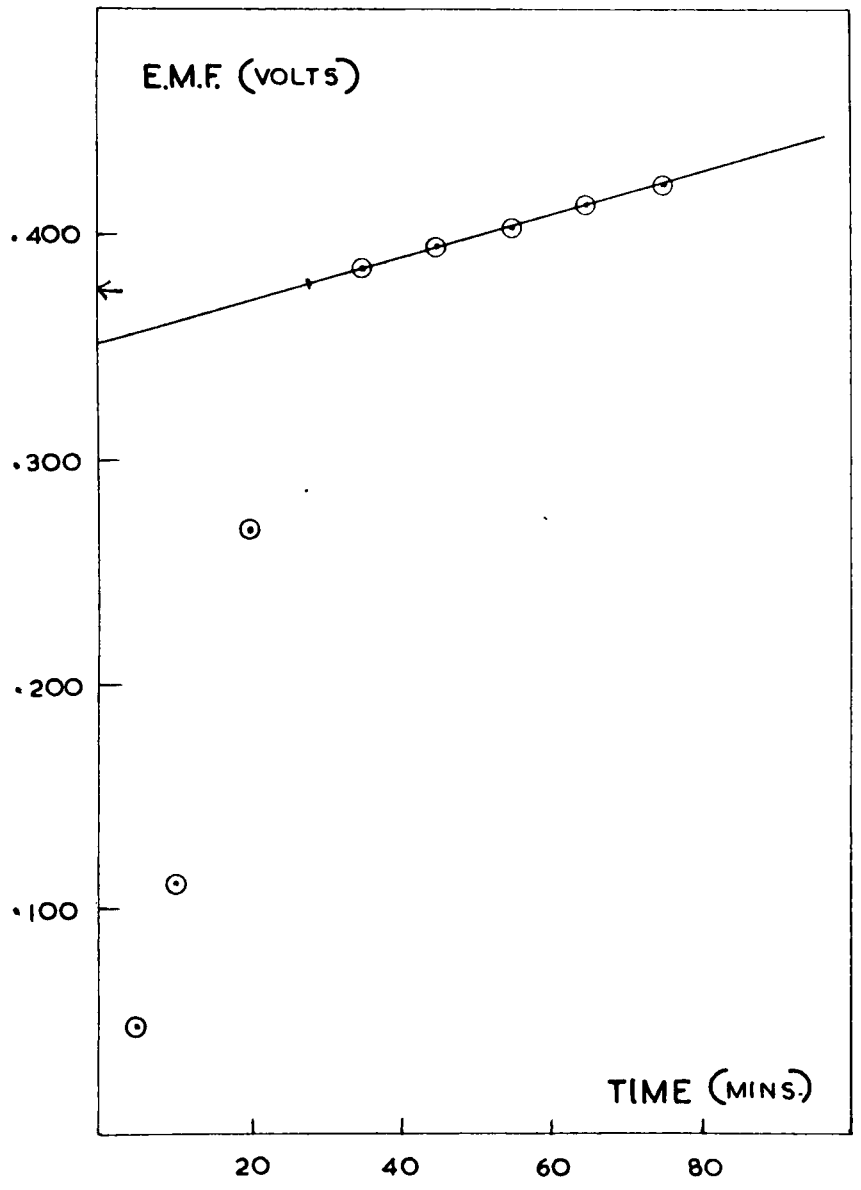
Bromomalonic Ester.

The ester was prepared as described in Organic Syntheses (Coll. Vol. 1, p. 240) b 121-122/14 mm.

The experiments carried out both in acetate and chloracetate buffers yielded results of a qualitative order. The E.M.F.'s did not become steady; there is little doubt that the dibromo compound is unstable. With this substance, the difficulties common to all bromination experiments, as well

FIGURE 6

BROMINATION OF BROM-MALONIC ESTER



as those previously met with in di-iodination experiments were encountered.

E.M.F. (volts)	Final Concentrations (mols/litre) (H^+)	(Br^-)	(RBr_2)	(RHBr)	(Br_2)	K
.352	$9.33.10^{-5}$	$6.57.10^{-3}$	$9.25.10^{-4}$	$2.05.10^{-4}$	$1.23.10^{-17}$	$1.4.10^{7}$
.292	$9.87.10^{-4}$	$6.57.10^{-3}$	$9.25.10^{-4}$	$1.33.10^{-3}$	$1.63.10^{-15}$	$1.4.10^{5}$

Initial Conc. $(Br_2 + Br_3^-)$ = .000925 throughout

$(Br^- + Br_3^-)$ = .00565

Acetylacetone.

As in the case of acetoacetic ester and 2-Carbethoxycyclopentanone, constant K values were obtained over relatively small pH ranges in acetate buffers, but there is a perceptible fall of equilibrium constant with pH.

Specimen Experiments.

E.M.F. (volts)	Final Concentrations (mols/litre) (H^+)	$(B\bar{r})$	(RBr)	(RH)	(Br_2)	K
.563	$2.15.10^{-5}$	$1.30.10^{-2}$	$1.78.10^{-3}$	$4.05.10^{-2}$	$2.04.10^{-22}$	$3.86.10^{13}$
.568	$3.45.10^{-5}$	$1.30.10^{-2}$	$1.78.10^{-3}$	$9.28.10^{-2}$	$1.39.10^{-22}$	$3.96.10^{13}$
.573	$2.15.10^{-5}$	$1.30.10^{-2}$	$1.78.10^{-3}$	$9.28.10^{-2}$	$9.23.10^{-23}$	$3.71.10^{13}$
.562	$5.60.10^{-5}$	$1.30.10^{-2}$	$1.78.10^{-3}$	$9.28.10^{-2}$	$2.23.10^{-23}$	$4.01.10^{13}$

Initial Conc. $(Br_2 + Br_3^-)$ = .00178 $(Br^- + Br_3^-)$ = .0112 throughout.

Acetone.

Buffer mixtures took of the order of a fortnight to come to equilibrium.

Specimen Experiment.

E.M.F.	Final Concentrations (mols/litre)					K
(volts)	(H^+)	(Br^-)	(RBr)	(RH)	(Br_2)	
.512	$1.86.10^{-3}$	$1.30.10^{-2}$	$1.76.10^{-3}$	$3.30.10^{-1}$	$1.06.10^{-20}$	$7.1.10^{12}$

Initial Concs. $(Br_2 + Br_3^-) = .00176$ $(Br^- + Br_3^-) = .0112$

Table 12.

Equilibrium Constants of Bromination.

(to the nearest power of 10)

2-Carbethoxycyclopentanone	10^{13}
2-Carbethoxycyclohexanone	10^{12}
Malonic Ester	10^{9}
Acetoacetic Ester	10^{10}
Acetylacetone	10^{13}
Bromomalonic Ester	10^{6}
Acetone	10^{13}

The Halogenation of Nitroparaffins.

The kinetics of bromination of some nitroparaffins have been studied. [15, 16]. The rate constants were used to calculate the time required for the various experimental mixtures to come to equilibrium. The dissociation constants of normal and aci forms have also been determined see p.108a

The nitroparaffins used were supplied by I.C.I.(Billingham) Ltd. They were dried and distilled at reduced pressure.

Nitromethane.

A series of experiments with varying initial nitromethane -iodine ratios were conducted. Although the equilibrium constant calculated for the equilibrium

$$CH_3NO_2 + I_2 = CH_2(I)NO_2 + I^- + H^+$$

showed some variation, this was not related to the initial nitroparaffin-iodine ratio. There is no doubt that the free energy of iodination of the second iodination equilibrium differs sufficiently from that of the first equilibrium, as in the case of ketones, to allow independent study of both equilibria.

Specimen experiment.-

Solution 1. Initial $(I_2 + I_3^-) = 0.000861$

" $(I^- + I_3^-) = 0.00461$

Solution 2. Initial iodine-iodide concentrations as in 1.

" (CH_3NO_2) = 0.0563

(Acetate) = 0.0695

(Acetic acid) = 0.0107 (H^+) = $4.27.10^{-6}$ I = 0.1

Nitromethane added to Solution 2 at 10.05 h. Solution made up to 100 cc.

Cell fitted and placed in thermostat at 10.10 h.

Time.	E.M.F.	Time.	E.M.F.
10.20	0.0042	12.30	0.1458
10.50	0.0350	12.50	0.1460
11.40	0.1416	13.50	0.1460
12.10	0.1453	14.50	0.1460

E.M.F. = 0.1460 volts.

hence, (RI) = 0.000861 (H^+) = $4.27.10^{-6}$

(RH) = 0.0554 (I_2) = $4.97.10^{-9}$

(I^-) = $5.47.10^{-3}$

$$K = \frac{(RI)\ (I^-)\ (H^+)\ .\ 0.64}{(RH)\ (I_2)} = \underline{0.046_8}$$

Table 13 gives a summary of experiments with this substance. The variation in the values for the equilibrium constant is apparently random. It is probably connected with further change of the halogenated compounds. The variation is only of the order of a factor of 2 over a pH range of 3, so that

the average value quoted is a good measure of the iodination equilibrium.

The two values marked with an asterisk were obtained in experiments in which solutions used in the previous experiment were mixed with a fresh buffer. In this way a fresh equilibrium is set up and a test was made whether the whole system behaved in a reversible manner.

Table 13 : The Iodination of Nitromethane.

Initial Concs. $(I_2 + I_3^-)$	Initial Concs. $(I^- + I_3^-)$	E.M.F. (volts)	Final Concentrations (mols/litre) (H^+)	(I^-)	(RI)	(RH)	(I_2)	K	
.000861	.00461	.1366	$7.46.10^{-6}$	$5.47.10^{-3}$	$8.61.10^{-4}$	$5.54.10^{-2}$	$1.03.10^{-8}$	0.0395	
.000861	.00461	.1460	$4.27.10^{-6}$	$5.47.10^{-3}$	$8.61.10^{-4}$	$5.54.10^{-2}$	$4.97.10^{-9}$	0.0468	
.000861	.00461	.1117	$3.16.10^{-5}$	$5.47.10^{-3}$	$8.61.10^{-4}$	$5.54.10^{-2}$	$7.05.10^{-8}$	0.0245	
.000861	.00461	.0974	$9.34.10^{-5}$	$5.47.10^{-3}$	$8.61.10^{-4}$	$5.54.10^{-2}$	$2.25.10^{-7}$	0.0226	
.000861	.00461	.0930	$2.27.10^{-4}$	$5.47.10^{-3}$	$8.66.10^{-4}$	$5.54.10^{-2}$	$3.12.10^{-7}$	0.0396	
.000344	.00184	.1154	$1.59.10^{-5}$	$2.18.10^{-3}$	$3.44.10^{-4}$	$2.22.10^{-2}$	$8.66.10^{-9}$	0.0398	*
.00124	.00643	.1053	$3.16.10^{-5}$	$7.67.10^{-3}$	$1.24.10^{-3}$	$5.51.10^{-2}$	$1.41.10^{-7}$	0.0249	
.000620	.00321	.1176	$7.59.10^{-6}$	$3.83.10^{-3}$	$6.20.10^{-4}$	$2.75.10^{-2}$	$1.33.10^{-8}$	0.0316	*
x .00124	.00643	.1705	$2.67.10^{-7}$	$7.67.10^{-3}$	$1.24.10^{-3}$	$9.36.10^{-2}$	$8.75.10^{-10}$	0.0198	

Average value = 0.03(2)

x Phosphate Buffer.

Some bromination experiments were also made with nitromethane.

E.M.F.	Final Concentrations (mols/litre)					K
(volts)	(H^+)	(Br^-)	(RBr)	(RH)	(Br_2)	
.4836	$4.27.10^{-5}$	$1.21.10^{-2}$	$9.15.10^{-4}$	$3.45.10^{-2}$	$4.29.10^{-20}$	$2.04.10^{11}$

Initial Concs. $(Br_2 + Br_3^-) = .000915$

$(Br^- + Br_3^-) = .0112$

The iodination of nitroethane, 1-Nitropropane, and 2-Nitropropane was studied in a similar manner. With these nitroparaffins, a slow secondary drift with E.M.F. was observed.

Table 14 A : The Iodination of Nitroethane.

Initial Concs.		E.M.F.	Final Concentrations (mols/litre)					K
$(I_2 + I_3^-)$	$(I^- + I_3^-)$	(volts)	(H^+)	(I^-)	(RI)	(RH)	(I_2)	
.00124	.00643	.1036	$2.27.10^{-4}$	$7.67.10^{-3}$	$1.24.10^{-3}$	$4.75.10^{-2}$	$1.59.10^{-7}$	.0804
.00124	.00643	.1520	$4.36.10^{-6}$	$7.67.10^{-3}$	$1.24.10^{-3}$	$7.43.10^{-2}$	$3.69.10^{-9}$	.0968
.00124	.00643	.1831	$2.67.10^{-7}$	$7.67.10^{-3}$	$1.24.10^{-3}$	$7.43.10^{-2}$	$3.29.10^{-10}$	.0663
.00124	.00643	.1453	$7.59.10^{-6}$	$7.67.10^{-3}$	$1.24.10^{-3}$	$5.92.10^{-2}$	$6.36.10^{-9}$	.0122

Average value = 0.09

Table 14 B : The Iodination of 1-Nitropropane.

Initial Concs.		E.M.F.	Final Concentrations (mols/litre)					K
$(I_2 + I_3^-)$	$(I^- + I_3^-)$	(volts)	(H^+)	(I^-)	(RI)	(RH)	(I_2)	
.00124	.00643	.1347	$7.59.10^{-6}$	$7.67.10^{-3}$	$1.24.10^{-3}$	$2.13.10^{-2}$	$1.42.10^{-8}$	0.151
.00124	.00643	.1518	$7.59.10^{-6}$	$7.67.10^{-3}$	$1.24.10^{-3}$	$8.20.10^{-2}$	$3.75.10^{-9}$	0.150
.00124	.00643	.1472	$7.59.10^{-6}$	$7.67.10^{-3}$	$1.24.10^{-3}$	$5.97.10^{-2}$	$5.36.10^{-9}$	0.144
.00124	.00643	.1487	$7.59.10^{-6}$	$7.67.10^{-3}$	$1.24.10^{-3}$	$6.73.10^{-2}$	$4.77.10^{-9}$	0.144
.00124	.00643	.1580	$4.36.10^{-6}$	$7.67.10^{-3}$	$1.24.10^{-3}$	$6.23.10^{-2}$	$2.31.10^{-9}$	0.184
.00124	.00643	.1829	$2.67.10^{-7}$	$7.67.10^{-3}$	$1.24.10^{-3}$	$4.80.10^{-2}$	$3.34.10^{-10}$	0.101
.00124	.00643	.1824	$2.67.10^{-7}$	$7.67.10^{-3}$	$1.24.10^{-3}$	$4.75.10^{-2}$	$3.47.10^{-10}$	0.099
.00124	.00643	.1242	$3.20.10^{-5}$	$7.67.10^{-3}$	$1.24.10^{-3}$	$4.96.10^{-2}$	$3.07.10^{-8}$	0.128

Average value = 0.13

Table 14 C : The Iodination of 2-Nitropropane.

Initial Concs. $(I_2 + I_3^-)$	Initial Concs. $(I^- + I_3^-)$	E.M.F. (volts)	Final Concentrations (mols/litre) (H^+)	(I^-)	(RI)	(RH)	(I_2)	K
.000942	.00829	0.0888	$5.89.10^{-6}$	$9.23.10^{-3}$	$9.42.10^{-4}$	$3.16.10^{-2}$	$2.24.10^{-7}$	$.0046_2$
.000942	.00829	0.0783	$1.02.10^{-5}$	$9.23.10^{-3}$	$9.42.10^{-4}$	$3.16.10^{-2}$	$5.07.10^{-7}$	$.0035_2$

Average value = 0.004.

The Halogenation of Cyanides.

Experiments with alkyl cyanides showed that the -CN group did not activate the carbon-hydrogen bond sufficiently for halogenation to occur. For instance, no halogenation could be detected in solutions of methyl cyanide in mineral acid or acetate buffers even after one month.

In the case of ketonic substances or nitroparaffins, halogenation proceeds through the mesomeric ion in which there is a stabilising distribution of negative charge.

e.g. $CH_3COCH_2COCH_3 \rightarrow CH_3 - \overset{\ominus}{C} - CH - C - CH_3$ (O below each C) Mesomeric ion.

$CH_3NO_2 \rightarrow CH_2 - N\begin{matrix}O\\O\end{matrix}$ " "

This mesomerism is much less effective in the cyanides.

When the effect of the cyano group reinforces the carbethoxy group in activation of the C-H bond as in cyanacetic ester rapid halogenation is found to occur.

$$N \equiv C - CH_2 - COOC_2H_5 \longrightarrow N \equiv C - \overset{\ominus}{CH} - C(=O) - OC_2H_5 \rightarrow NC - \underset{I}{CH} - COOC_2H_5$$

Specimen Experiment.-

Solution 1. Initial $(I_2 + I_3^-) = 0.000558$

" $(I^- + I_3^-) = 0.00624$

$\mathcal{I} = 0.1$

100 cc.

Solution 2. Initial iodine-iodide concentrations as in Soln. 1

" (Cyanacetic ester) = 0.0271

$(H^+) = 4.53.10^{-6}$ $I = 0.1$

(Acetate) = .0669 (Acetic acid) = .0110

Ester added at 15.15 h. and solution 2 made up to 100 cc.

Time.	E.M.F.	Time.	E.M.F.
15.20	.1026	15.35	.1034
15.25	.1032	15.40	.1035
15.30	.1033	15.45	.1036
		16.45	.1048

E.M.F. = 0.1030 volts.

Experiments in chloracetate buffers confirmed that in the above experiment iodination would be complete after 10 minutes. The rate of halogenation is therefore intermediate between that for malonic and acetoacetic esters. There is a slow change of E.M.F. with time after the primary equilibrium is reached, as in the case of several ketones. The value of the E.M.F. used in calculation is obtained by back-extrapolation.

Table 15 : The Iodination of Cyanacetic Ester.

	Initial Concs.		E.M.F.	Final Concentrations (mols/litre)					K
	$(I_2 + I_3^-)$	$(I^- + I_3^-)$	(volts)	(H^+)	(I^-)	(RI)	(RH)	(I_2)	
	.00558	.00624	.0775	$2.29.10^{-5}$	$6.80.10^{-3}$	$5.58.10^{-4}$	$1.83.10^{-2}$	$3.62.10^{-7}$	.00842
	.00558	.00624	.0700	$4.90.10^{-5}$	$6.80.10^{-3}$	$5.58.10^{-4}$	$2.18.10^{-2}$	$6.49.10^{-7}$	.00840
	.00558	.00624	.0934	$1.02.10^{-5}$	$6.80.10^{-3}$	$5.58.10^{-4}$	$2.73.10^{-2}$	$1.05.10^{-7}$	.00864
	.00558	.00624	.1030	$4.53.10^{-6}$	$6.80.10^{-3}$	$5.58.10^{-4}$	$2.65.10^{-2}$	$4.94.10^{-8}$	.00842
X	.00558	.00624	.0475	$4.31.10^{-4}$	$6.80.10^{-3}$	$5.36.10^{-4}$	$3.32.10^{-2}$	$3.69.10^{-6}$	.00819

Average value = 0.00841

X Chloracetate Buffer.

The Halogenation of Amides.

Experiments have now been described on compounds containing a carbon-hydrogen bond activated by carboxyl, carbethoxy, carbonyl, nitro, and cyano groups.

An attempt was made to extend this investigation to the halogenation equilibria of nitrogen compounds.

$$-NH- + X_2 = -NX- + X^- + H^+$$

The simplest organic compounds containing a nitrogen hydrogen bond are amines. These are not sufficiently reactive however; moreoever, their basic character would introduce complications. The reactivity of the N-H bond is greatly increased by the carbonyl group in amides and it was found possible to study the halogenation equilibria of acetamide.

In acetate buffers the halogenation was found to be very rapid. In the case of bromine, further disappearance of bromine, after the initial equilibrium is established, is sufficiently slow to make back-extrapolation reliable. This is not the case with iodine. Iodo-acetamide is even less stable than the bromocompound, so that in addition to the secondary change previously observed, there is likely to be a decomposition of the halogenated derivative. It is possible that a Hofmann type degradation occurs. The equilibrium constants are therefore only of a semi-quantitative character.

Specimen experiments : -

Bromination

E.M.F. (volts)	Final Concentrations (mols/litre) (H^+)	(Br^-)	(RBr)	(RH)	(Br_2)	K
.381	$2.32.10^{-5}$	$1.21.10^{-2}$	$9.15.10^{-4}$	$1.24.10^{-1}$	$1.25.10^{-16}$	$10.6.10^{6}$
.353	$1.31.10^{-4}$	$1.21.10^{-2}$	$9.15.10^{-4}$	$1.24.10^{-1}$	$1.11.10^{-15}$	$6.8.10^{6}$

Initial Concs. $(Br_2 + Br_3^-) = .000915$

$(Br^- + Br_3^-) = .0112$

throughout

$$K = 9.10^{6}$$

Iodination

E.M.F. (volts)	Final Concentrations (mols/litre) (H^+)	(I^-)	(RI)	(RH)	(I_2)	K
.060	$5.32.10^{-6}$	$9.23.10^{-3}$	$9.42.10^{-4}$	$6.24.10^{-1}$	$2.24.10^{-6}$	$3.3.10^{-5}$
.070	$5.93.10^{-6}$	$9.23.10^{-3}$	$9.42.10^{-4}$	$6.24.10^{-1}$	$1.63.10^{-6}$	$3.5.10^{-5}$

Initial Concs. $(I_2 + I_3^-) = .000942$

$(I^- + I_3^-) = .00829$

throughout

$$K = 3.10^{-5}$$

Measurements at temperatures other than 25°C.

The previous chapters record the determination of free energies of halogenation at 25°C. It will be seen later, when certain bond energy quantities are evaluated from these data, that a knowledge of the heat of halogenation of one or two of the simpler substrates would be desirable. The equilibrium constants of iodination of acetone and acetoacetic ester were therefore measured at several temperatures, and the heats of iodination calculated from the Gibbs-Helmholtz relation.

Measurements were made at 0°, 20°, 25°, and 30° C. In order to calculate the equilibrium constants, the relevant constants for the iodine-triiodide equilibrium are required. These have been measured directly at 0°C. and 25°C. [6] where they are 1390 and 714 respectively.

$$I_2 + I^- = I_3^-$$

0°C	ΔF = 3919 cal.	K = 0.00072	$K_{I_3^-}$ = 1/K = 1390
25°C	ΔF = 3885 cal.	K = 0.00140	$K_{I_3^-}$ = 1/K = 714

The heat content change can be obtained from the Gibbs-Helmholtz relation. Since the entropy change is very small, the free energies at 20°C and 30°C can be calculated very accurately.

20°C	ΔF = 3892 cal.	K = 0.00123	$K_{I_3^-}$ = 812
30°C	ΔF = 3878 cal.	K = 0.00158	$K_{I_3^-}$ = 633

Acetoacetic ester.

The measurements at 0°C were carried out with the cell immersed in a thermally insulated bath of ice and water, which was stirred by hand. Acetoacetic ester mixtures in acetate buffers come to equilibrium in about a minute. Measurements at 25°C also showed that there was practically no drift of E.M.F. due to secondary reaction. The E.M.F. fell for a few minutes after immersion of the cell and remained constant when temperature equilibrium was reached.

For experiments at 20°, 25°, and 30° C it was therefore possible to use the same mixture throughout. The cell was placed in turn in the water thermostat controlled to the above three temperatures to 0.02°C until a steady E.M.F. was obtained in each case. The procedure was then repeated in the reverse order, when almost identical readings were obtained.

Specimen Experiments.

The Iodination of Acetoacetic Ester.

	$K_{I_3^-}$	RT/2F	Initial Concs.		E.M.F.	Final Concentrations (mols/litre)					K
			$(I_2 + I_3^-)$	$(I^- + I_3^-)$	(volts)	(H^+)	(I^-)	(RI)	(RH)	(I_2)	
0°C	1390	0.0271	.000942	.00829	.1021	$2.21.10^{-5}$	$9.23.10^{-3}$	$9.42.10^{-4}$	$2.92.10^{-2}$	$2.20.10^{-8}$	.204
20°C	812	0.0291	.000942	.00829	.1200	$1.95.10^{-5}$	$9.23.10^{-3}$	$9.42.10^{-4}$	$2.92.10^{-2}$	$1.52.10^{-8}$	.242
25°C	714	0.0296	.000942	.00829	.1237	$1.95.10^{-5}$	$9.23.10^{-3}$	$9.42.10^{-4}$	$2.92.10^{-2}$	$1.48.10^{-8}$	.251
30°C	635	0.0301	.000942	.00829	.1278	$1.95.10^{-5}$	$9.23.10^{-3}$	$9.42.10^{-4}$	$2.92.10^{-2}$	$1.41.10^{-8}$	.264

In calculating the equilibrium constants a very small allowance had to be made for the variation of the activity coefficients from 30° to 0° C.

Ketone added at 11.30 h. Cell placed in thermostat :-

20°C at 11.35 h.		25°C at 11.55 h.		30°C at 12.15 h.	
Time.	E.M.F.	Time.	E.M.F.	Time.	E.M.F.
11.45	.1200	12.00	.1237	12.25	.1278
11.50	.1200	12.05	.1237	12.30	.1278
11.55	.1200	12.15	.1237	12.45	.1278

Table 16.

	K	ΔF (cals)	ΔF/T	1/T
0°C	.204	860	3.15	.00366
20°C	.242	824	2.81	.00341
25°C	.251	817	2.74	.00335
30°C	.264	810	2.68	.00330

Table 16 gives a summary of the equilibrium constants. The free energies of halogenation are obtained from them by means of the Van't Hoff isotherm, and the heat of halogenation is obtained from the Gibbs-Helmholtz relation.

Graph 7 shows the plot of $\Delta F/T$ against the reciprocal of the absolute temperature. The slope gives the heat of halogenation ΔH in cals.

$$\Delta H = \Delta F - T \cdot \frac{d \Delta F}{dT}$$

For the reaction

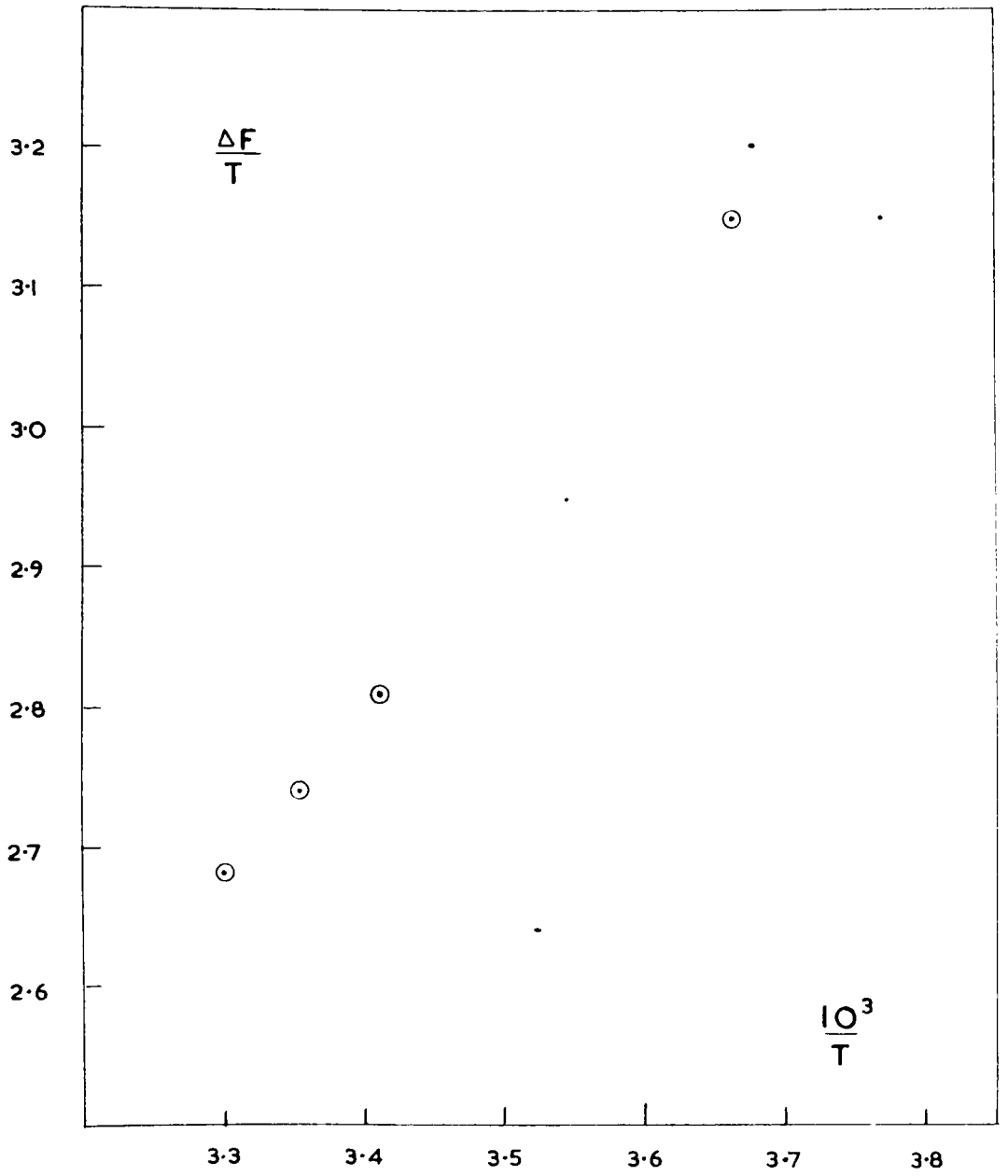
$$CH_3COCH_2COOC_2H_3 + I_2 = CH_3COCH(I)COOC_2H_3 + I^- + H^+$$

at 25°C. $\underline{\Delta H = 1.2 \text{ k.cal.}}$ (± 0.2 k.cal.)

FIGURE 7

IODINATION OF ACETOACETIC ESTER AT SEVERAL TEMPERATURES.

ΔH (SLOPE) = 1·2 - 1·3 K.CAL.



Acetone.

As the iodination of acetone is very slow, the technique of studying the same solution successively at several temperatures cannot be adopted.

Specimen Experiments.-

E.M.F. (volts)	Final Concentrations (mols/litre) (H^+)	(I^-)	(RI)	(RH)	(I_2)	K
20°C.						
.1001	$2.76.10^{-2}$	$9.23.10^{-3}$	$9.42.10^{-4}$	$7.00.10^{-2}$	$7.34.10^{-8}$	30.0
25°C.						
.1031	$2.76.10^{-2}$	$9.23.10^{-3}$	$9.42.10^{-4}$	$7.00.10^{-2}$	$7.37.10^{-8}$	29.9
30°C.						
.1063	$2.76.10^{-2}$	$9.23.10^{-3}$	$9.42.10^{-4}$	$7.00.10^{-2}$	$7.29.10^{-8}$	30.1

Initial Concs. $(I_2 + I_3^-)$ = .000942

$(I^- + I_3^-)$ = .00829 throughout

Table 17.

	K	ΔF (cal)
20°C	30.0	- 1973
25°C	29.9	- 2013
30°C	30.1	- 2047

$$\Delta H = \Delta F - T \frac{d(\Delta F)}{dT}$$

For the reaction

$$CH_3COCH_3 + I_2 = CH_3COCH_2I + I^- + H^+$$

at 25°C $\underline{\Delta H = 0.2 \text{ k.cal.}}$ (± 0.5 k.cal.)

Table of Results.

Table 18 presents the equilibrium constants of halogenation whose measurement is described in Part I. The values marked with an asterisk are accurate to ± 2-3 % .

Table 18.

At 25°C.	Iodine.	Bromine.
Acetone	29.2 *	10^{13}
Cyclopentanone	700	
Cyclohexanone	1170 *	
Sym-Dichloracetone	0.203	
2-Carbethoxycyclohexanone	0.505 *	10^{12}
Malonic Ester	0.0523 *	10^{9}
Acetoacetic Ester	0.250 *	10^{10}
2-Carbethoxycyclopentanone	1.90 *	10^{13}
Acetylacetone	1.12 *	10^{13}
Malonic acid anion $COO^{-}CH_2COOH$	1	
Iodomalonic Ester	10^{-7}	
α -Iodoacetoacetic Ester	10^{-4}	
α -Iodoacetylacetone	10^{-3}	
Brommalonic Ester		10^{6}

At 25°C.	Iodine.	Bromine.
Nitromethane	0.03	2.10^{11}
Nitroethane	0.09	
1:Nitropropane	0.13	
2:Nitropropane	0.004	
Cyanacetic Ester	0.00841 *	
Acetamide	3.10^{-5}	9.10^{6}

Iodination of	at 0°C.	20°C.	25°C.	30°C.
Acetoacetic Ester	.204 *	.242 *	.250 *	.264 *
Acetone	--	30.0 *	29.2 *	30.1 *

Discussion.

The preceding chapters record the development of a method of studying halogenation equilibria, and the measurement of equilibrium constants of iodination or bromination for 20 organic compounds.

It is now intended to discuss the factors governing the position of equilibrium of such a halogenation reaction, and to account for their variation from one reaction to another.

The halogenation equilibrium can in general be represented as follows : -

$$RH_{aq} + X_{2\,aq} = RX_{aq} + X^-_{aq} + H^+_{aq}$$

The energetics of this system can be analysed in several ways.

For instance, the reaction may be considered as the sum of the reactions

$$RH_{aq} \xrightarrow[1]{} RH_g \xrightarrow[2]{} R_g + H_g$$

$$R_g + X_g \xrightarrow[3]{} RX_g \xrightarrow[4]{} RX_{aq}$$

$$X_{2\,aq} \xrightarrow[5]{} X_{2\,g} \xrightarrow[6]{} X_g + X_g$$

$$H_g + X_g \xrightarrow[7]{} HX_g \xrightarrow[8]{} H^+_{aq} + X^-_{aq}$$

In this way attention is focussed on the dependence of the equilibrium on the dissociation energies of the hydrogen and halogen bonds, as well as solvation energies. It will be seen

later that this approach provides a method of evaluating the difference in the dissociation energies of the hydrogen and halogen bonds in the organic compounds RH and RX i.e. the quantity corresponding to the sum of reactions 2 and 3.

A thermodynamically equivalent approach would be to consider the halogenation reaction in terms of the ionisations which occur in aqueous solution. This brings out the analogy between the release of a proton by an acid and the production of a halogen cation by a positive halogen compound. The hydrogen and halogen bonds ~~and~~ of the organic compound, which are broken and formed in halogenation, both have a thermodynamic tendency to ionise. The halogenation can be considered as the sum of the processes :-

$$RH_{aq} = R^-_{aq} + H^+_{aq}$$

$$X_{2\ aq} = X^-_{aq} + X^+_{aq}$$

$$X^+_{aq} + R^-_{aq} = RX_{aq}$$

Both approaches are valuable in illuminating different aspects of the energetics of the halogenation equilibrium, and they will now be discussed in more detail.

The Equilibrium and Bond Energies.

The Gibbs free energy ΔF of the reaction at $T^{o}K$

$$RH + X_2 = RX + X^- + H^+$$

is the sum of the free energies of the steps

1. $RH_{aq} \longrightarrow RH_g$
2. $RH_g \longrightarrow R_g + H_g$
3. $R_g + X_g \longrightarrow RX_g$
4. $RX_g \longrightarrow RX_{aq}$
5. $X_{2\,aq} \longrightarrow X_{2\,g}$
6. $X_{2\,g} \longrightarrow 2X_g$
7. $H_g + X_g \longrightarrow H^+_{aq} + X^-_{aq}$

Similarly, the heat of halogenation ΔH is the sum of the heat content changes in steps 1-7.

Free Energy Measurements at $25^{o}C$.

$\Delta F_1 + \Delta F_4$

This is the difference in the free energies of vaporisation of the organic compound and its halogen derivative. This quantity is of the order of 0.2 k.cal./mole [19].

$\Delta F_2 + \Delta F_3$

Let D_H and D_X represent the dissociation energies of the C-H or N-H and the C-X or N-X bonds respectively, at $25^{o}C$.

Then, $$\Delta H_2 + \Delta H_3 = D_H - D_X$$

and $$\Delta F_2 + \Delta F_3 = (D_H - D_X) - 298 (\Delta S_H - \Delta S_X)$$

where ΔS is the entropy increase on dissociation of RH or RX.

ΔF_5

It is obtained from vapour pressure [20, 21] and solubility measurements [22, 23] for the respective steps

$$X_2 \text{ (standard)} \longrightarrow X_2 \text{ g}$$

$$X_2 \text{ (standard)} \longrightarrow X_2 \text{ aq.}$$

At 25°C, ΔF_5 = 0.7 k.cal. for iodine and - 0.2 k.cal. for bromine.

ΔF_6

The values for this and other thermochemicalquantities , whose sources are not specifically stated, are taken from the tables of the National Bureau of Standards, Washington, D.C., 1947.

ΔF_6 = 28.9 k.cal. for iodine and 38.6 k.cal. for bromine.

ΔF_7

This quantity involves the ionisation potential of hydrogen and the electfon affinity of the halogens, as well as the free energies of solution of the ions.

The most accurate measure can be obtained from E.M.F. measurements on the cell [6, 7] .

$$H_2 \text{ (gas)} \quad / \quad HX \text{ (aqueous)} \quad / \quad X_2 \text{ (standard)}$$

This gives the free energy of the reaction

(a) $\frac{1}{2} H_2$ (gas) + $\frac{1}{2} X_2$ (standard) $\longrightarrow$ H^+_{aq} + X^-_{aq}

$$\Delta F_7 = \Delta F_a + \Delta F_b + \Delta F_c + \Delta F_d$$

where

(b) $\frac{1}{2} X_2$ (gas) $\longrightarrow$ $\frac{1}{2} X_2$ (standard)

(c) H (gas) $\longrightarrow$ $\frac{1}{2} H_2$ (gas)

(d) X (gas) $\longrightarrow$ $\frac{1}{2} X_2$ (standard)

ΔF_7 = - (12.35 + 2.31 + 48.57 + 14.45)

= - 77.68 k.cal. for iodine

and

ΔF_7 = - (24.57 + 0.37 + 48.57 + 19.30)

= - 92.81 k.cal. for bromine

The free energy of halogenation ΔF is the sum of the free energies $\Delta F_1 + \Delta F_2 \ldots\ldots + \Delta F_7$

If D_H, D_I, and D_{Br} are the dissociation energies in k.cal./mole for the reactions at 25°C.

$$RH_g \rightarrow R_g + H_g$$
$$RI_g \rightarrow R_g + I_g$$
$$RBr_g \rightarrow R_g + Br_g$$

we have for iodination at 25°C : -

$\Delta F = 0.2 + (D_H - D_I) - 298 (\Delta S_H - \Delta S_I) + 0.7 + 28.9 - 77.68$

$D_H - D_I = \Delta F + 298 (\Delta S_H - \Delta S_I) + 48.0$ k.cal.

and similarly for bromination at 25°C : -

$\Delta F = 0.2 + (D_H - D_{Br}) - 298(\Delta S_H - \Delta S_{Br}) - 0.2 + 38.6 - 92.81$

$D_H - D_{Br} = \Delta F + 298 (\Delta S_H - \Delta S_{Br}) + 54.2$ k.cal

The above equations give the difference in the bond dissociation energies in terms of the free energy of halogenation.

The calculation of the entropy terms will now be described.

The entropy term at a temperature $T^{o}K$

$$\Delta S_H - \Delta S_X = S(H) - S(X) - S(RH) + S(RX)$$

This is equal to the entropy change in the reaction

$$RH + X = RX + H$$

There is one polyatomic and one monatomic species before and after reaction. If complete excitation of the rotational levels of both RH and RX is assumed, the entropy change will be independent of temperature.

Translational and rotational partition functions give the following expression for the entropy change [25]. Vibrational contributions at $25^{o}C.$ are neglected (v.i.)

$$\Delta S_H - \Delta S_X = S(H) - S(X) - S(RH) + S(RX)$$

$$= \sum_i 2.287\,(3 \log_{10} M - 2 \log_{10} \sigma + 2 \log_{10} n_i + \log_{10} ABC_i)$$

(The algebraic sum of the entropies of the four species)

M is the molecular or atomic weight

σ is the symmetry number of RH or RX

n is the number of stereoisomerides of RH or RX

ABC is the product of the principal moments of inertia of RH or RX in atomic weight - Angström units.

The Iodination of Ketones.

The entropy term can be divided into translational, and symmetry terms and the term contributed by the moments of inertia.

The translational term is shown in table 19. The entropy term involving σ and n makes allowance for the fact that in the halogen equilibrium there are often different numbers of hydrogen and iodine bonds involved.

Thus, for acetone there are 6 C-H bonds and 1 C-I bond to be considered.

σ (Iodoacetone) = 3 σ (Acetone) = 18

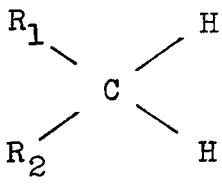
A factor of $\log_{10}6$ is introduced into the entropy term, and the equilibrium constant must be divided by a factor of 6, as might have been expected.

When the groups attached to the reactive carbon atom are not identical, however, stereoisomers are possible. Thus, in the case of acetoacetic ester the ratio of the symmetry numbers of RH and RX is unity, but

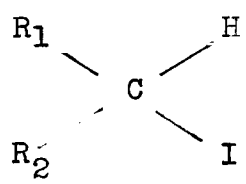
n (Iodoacetoacetic ester) = 2 n (Acetoacetic ester) = 1

and hence K_e must be divided by a factor of 2.

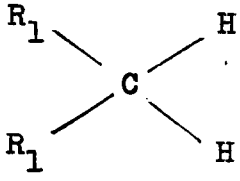
In general,



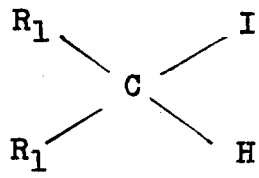
$\sigma \propto 1$, n = 1



$\sigma \propto 1$, n = 2



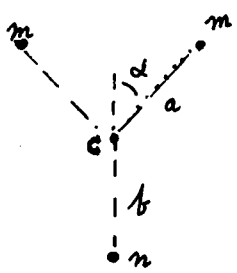
$\sigma \propto 2, \quad n = 1$



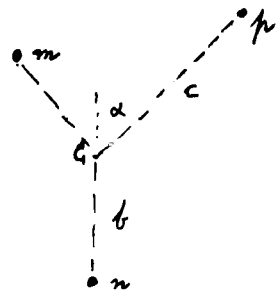
$\sigma \propto 1, \quad n = 1$

The entropy terms due to the symmetry factors are recorded in table 19.

Moments of inertia were calculated for acetone and iodoacetone using the following interatomic distances : -
C-H = 1.10 Å , C = O = 1.22 Å , C-I = 2.09 Å
The calculation can be simplified by taking the methyl groups as point masses .



Acetone.



Iodoacetone.

m = 15, n = 16, p = 141 , $\alpha = 54^{\circ}\ 44'$
and from the geometry of the tetrahedron a = 1.77 Å, c = 3.21 Å
The simplifications introduced are justified in view of the ~~ef~~ effect of the iodine atom which outweighs other factors.
Moments were taken about the most convenient point within the molecules, namely the central ~~x~~ carbon atom C, and moments of inertia calculated after the method of Hirschfelder [26; also 25].

$$\frac{\text{ABC (Iodoacetone)}}{\text{ABC (Acetone)}} = 10^3$$

This gives rise to an entropy term of 6.9 cal./°C. and $T\Delta S_{ABC} = 2.06$ k.cal. at 25°C.

The ratio of the moments of inertia of other ketones can be estimated in the following manner. The ratio will obviously decrease with increasing size of ketone.

If an inverse proportionality between this ratio and the molecular weight of the iodoketone is assumed, the entropy terms of the various ketones can be obtained. (When $1/M = 0$, the ratio of the moments of inertia is unity and the entropy term is zero.) The entropy terms are found to range over 0.5 k.cal. An average value of 1.8 k.cal. has been adopted.

Table 19 : The Iodination of Ketones : $D_H - D_I$ at 25°C.

The substances are arranged in order of increasing reactivity, i.e. acidity. Values are quoted in k.cal./mole.

Substance	ΔF	←— T ($\Delta S_H - \Delta S_I$) —→				$D_H - D_I$
		Transl.	Rot.	σ+n	Sum.	$48.0 + \Delta F + T(\Delta S_H - \Delta S_I)$
Cyclopentanone	- 3.88	- 3.49	1.8	0.40	- 1.3	42.8
Acetone	- 1.99	- 3.27	2.0_6	1.06	- 0.2	45.8
Cyclohexanone	- 4.18	- 3.57	1.8	0.40	- 1.4	42.4
Sym-Dichloracetone	0.87	- 3.69	1.8	0.82	- 1.1	47.8
2-Carbethoxycyclohexanone	0.40_5	- 3.82	1.8	0	- 2.0	46.4
Malonic Ester	1.73	- 3.79	1.8	0.40	- 1.6	48.1
Acetoacetic Ester	0.82	- 3.72	1.8	0.40	- 1.5	47.3
2-Carbethoxycyclopentanone	$- 0.39_6$	- 3.78	1.8	0	- 2.0	45.6
Acetylacetone	- 0.13	- 3.58	1.8	0.40	- 1.4	46.5

Table 19 shows the free energies of halogenation, the calculated entropy terms, and the values for the bond energy quantity ($D_H - D_I$) derived from them.

A check on the entropy calculations is provided by the values obtained for the heat of halogenation ΔH. It was seen above that this quantity is the sum of the heat content changes of steps 1 - 7.

$\Delta H_1 + \Delta H_4$ is the difference in the heats of vaporisation of the organic compound and its halogen derivative.

It appears from the work of Butler [19] and from available solubility data that the average value for the substances studied is $(\Delta H_1 + \Delta H_4) = -1$ k.cal. (~~See Graph 7~~).

$$\Delta H_2 + \Delta H_3 = D_H - D_X$$

ΔH_5, ΔH_6, and ΔH_7 are obtained from a study of the same reactions as the corresponding free energy changes.

ΔH_5 = 9.9 k.cal. for iodine and 8.4 k.cal. for bromine

ΔH_6 = 36.1 l.cal. for iodine and 46.1 k.cal. for bromine

ΔH_7 = $-(13.37 + 7.44 + 52.09 + 18.03) = -90.93$ k.cal. for iodine , and

= $-(28.90 + 3.67 + 52.09 + 23.04) = -107.70$ k.cal. for bromine (see p. 88).

$$\Delta H = \Delta H_1 + \Delta H_2 \ldots\ldots + \Delta H_7$$

Hence, we have for iodine at 25°C.

$$\Delta H = (\Delta H_1 + \Delta H_4) + (D_H - D_I) + 9.9 + 36.1 - 90.93$$

$$D_H - D_I = \Delta H - (\Delta H_1 + \Delta H_4) + 44.9 \text{ k.cal.}$$

and for bromine at 25°C. :

$$\Delta H = (\Delta H_1 + \Delta H_4) + (D_H - D_{Br}) + 8.4 + 46.1 - 107.70$$

$$D_H - D_{Br} = \Delta H - (\Delta H_1 + \Delta H_4) + 53.2 \text{ k.cal.}$$

In particular, the heats of iodination of acetone and acetoacetic ester have been measured (p. 80, 82) .

For acetoacetic ester ,

$\Delta H = 1.2$ k.cal.

and hence,

$$D_H - D_I = \underline{47.1 \text{ k.cal.}}$$

and for acetone,

$\Delta H = 0.2$ k.cal.

and hence,

$$D_H - D_I = \underline{46.1 \text{ k.cal.}}$$

in excellent agreement with the values obtained from the free energies of iodination at 25°C., and entropy calculations. (47.3 and 45.8).

It can be seen from the above considerations that the equilibrium constants would be almost the same for all substrates, were it not for the variation in the dissociation

energies of the hydrogen and halogen bonds.

A treatment of the variation in the dissociation energy of bonds of a given type in terms of the resonance energies of both radical and undissociated molecules has been attempted [34 - 36] .

However, the carbon - hydrogen bond dossociation energies for the above ketones are not known. It cannot even be said whether they are likely to increase or decrease with the general reactivity of the ketone.

The discussion will therefore have to be confined to differences in dissociation energies , i.e. $D_H - D_X$. The average value of $D_H - D_I$ for the ketones listed in table 19 is 45.8 k.cal. This is very close to the values obtained for the alkyl halides by different methods, such as pyrolytic and thermochemical measurements (41 - 47 k.cal.) [28]

The dissociation energies at 25°C obtained in the above treatment are equal to the heat content changes at 0°K plus kinetic energies. The latter terms do not necessarily cancel when the difference of $D_H - D_I$ is taken. It has therefore been held that for an investigation of the electrical effects of substituents in a compound, free energies of reaction are a better measure than heat content changes [32] . However, it is seen from table 19 that the variation of $D_H - D_I$ is in the same order as the free energies of halogenation ΔF, except in two instances where a difference of

less than 6.3 k.cal. is involved, which is about equal to the error of calculation. The equilibrium constants are therefore inversely proportional to the difference in the bond energies of the C-H and C-I bonds in RH and RI.

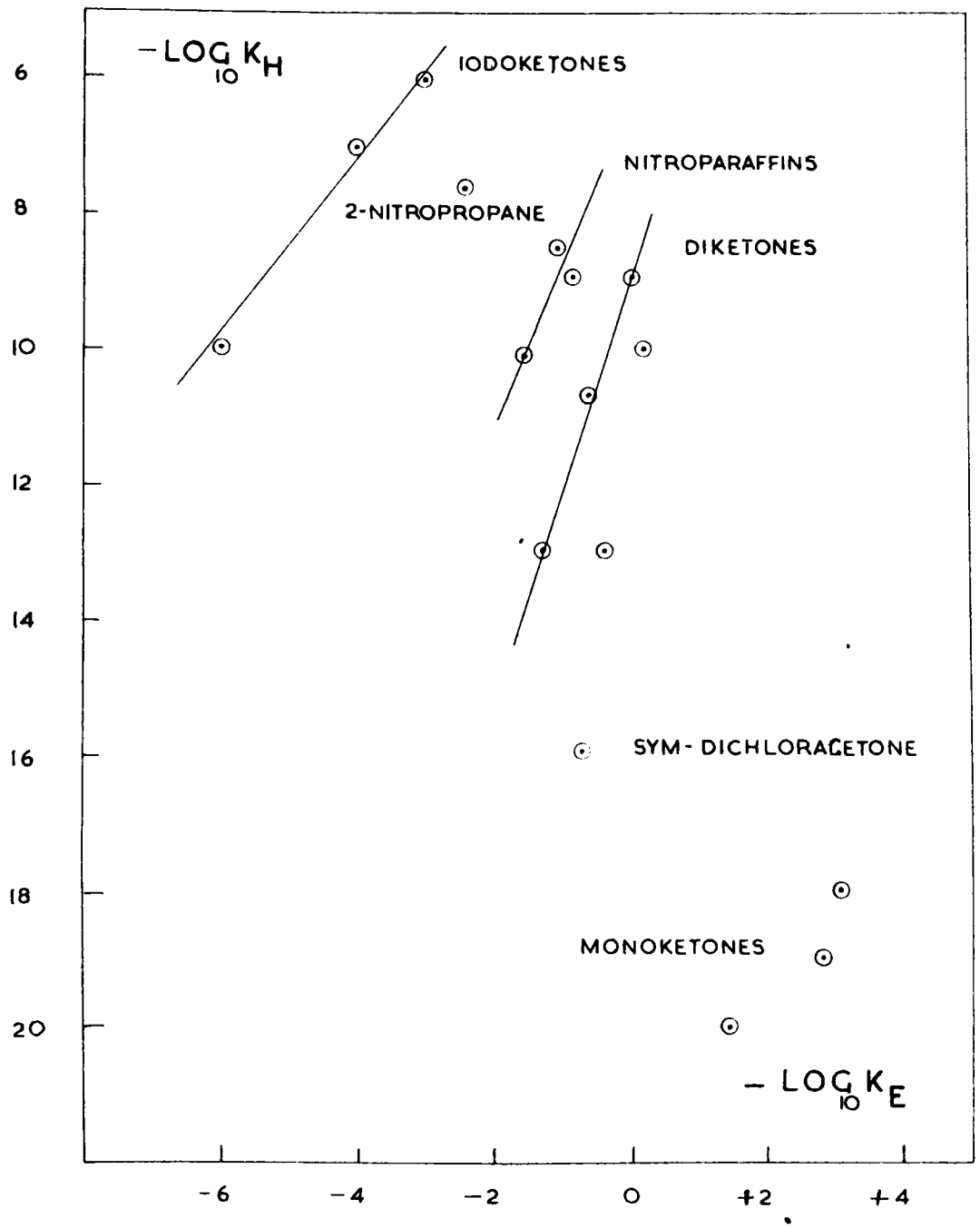
A decrease in the value of the equilibrium constant signifies an increase in $D_H - D_I$, and must mean a relative weakening of the carbon-iodine bond.

For a series of compounds in which steric effects are constant in both RH and RI, the variation in the energy of RH and RI might be supposed to be related to the reactivity of RH. When these conditions are obeyed, it is indeed found that as the reactivity increases, $D_H - D_X$ decreases, and hence the equilibrium constant increases. As a measure of reactivity the acid dissociation constant of the ketone K_H, its rate of halogenation R, or the Brønsted exponent α may be used. (See p. 103-).

Graph 8 shows the linear variation of log K_e (equilibrium constant) with log K_H, for the following series of structurally similar compounds in which the reactivity increases in the order given (1) malonic ester, acetoacetic ester, acetylacetone (2) iodomalonic ester, iodoacetoacetic ester, iodo-acetylacetone (3) nitromethane, nitroethane, 1:nitropropane. In these series the reactivity changes because of alteration of the inductive and tautomeric effects acting on the C-H and C-I bonds.

FIGURE 8

ELECTRICAL AND STERIC EFFECTS ON EQUILIBRIUM PROPERTIES



When there is a change in steric effect, however, no such relation is found. Thus, scale models show that there is appreciable steric hindrance in di-iodomalonic ester and other di-iodoketones and in α-iodo-2-nitropropane. The C-I bonds in these compounds are weakened by several k.cal. $D_H - D_I$ increases and the equilibrium constants are appreciably lower. No conformance to the above relationships based on electrical effects can then be expected.

Equilibria in Bromination.

Bromination experiments have not yet yielded sufficiently accurate data for the evaluation of $D_H - D_{Br}$ in individual cases. However, an average value of 10^{12} (cf. Table 12) for the equilibrium constant gives, according to the equation on p. 89

$$D_H - D_{Br} = 37.5 \text{ k.cal.}$$

since the average value for $D_H - D_I$ was noted to be 46 k.cal.

$$D_{Br} - D_I = \underline{8.5. \text{ k.cal.}}$$

This is considerably lower than the values ($12\frac{1}{2}$ k.cal. for Me & Et halides) reported for the alkyl halides [28] .

The difficulties encountered in the study of bromination equilibria make it unlikely that accurate data could be obtained for chlorinations. By analogy with the iodination and bromination equilibria (and the bond energies of the alkyl halides [28])

$$D_{Cl} - D_{Br} = 10 \text{ k.cal.}$$

$$D_H - D_{Cl} = 26.5 \text{ k.cal.}$$

and the equilibrium constants would be of the order of 10^{25}.

Dissociation Energies of Carbon, Nitrogen, and Oxygen Bonds.

From the experiments with acetamide the following more approximate values have been obtained : -

$D_H - D_I = 52$ k.cal. $\quad D_H - D_{Br} = 45$ k.cal.

The values of $D_H - D_I$ for methane is known to be 48 k.cal. The value for acetone has been found to be 45.8 k.cal. As acetamide is the nitrogen analogue of acetone, we deduce $D_H - D_I$ for NH_2I to be 54 k.cal., and similarly $D_H - D_{Br}$ for the bromination of ammonia to be 46 k.cal.

At 25^oC , the C-H dissociation energy in methane is 104 k.cal. [29] , the N-H dissociation energy in ammonia is 106 k.cal. [30] , and the O-H dissociation energy in water is 120 k.cal. [31] .

Hence,

$CH_3I \rightarrow CH_3 + I$	56 k.cal.	$CH_3Br \rightarrow CH_3 + Br$	69.5 k.cal
$NH_2I \rightarrow NH_2 + I$	52 k.cal.	$NH_2Br \rightarrow NH_2 + Br$	60 k.cal.
$HOI \rightarrow HO + I$	57.4 k.cal.	$HOBr \rightarrow HO + Br$	55.3 k.cal

The dissociation energies of the hypohalous acids are calculated in part II, p. 11-.

The differing trend in these energies which in the case of hydrogen is $C < N < O$, of bromine $C > N > O$, and of iodine $C > N < O$, is due to two factors : - overlapping of orbitals and electronegativity difference.

This is illustrated by the empirical method of Pauling, which relates the bond energy of a bond AB to the arithmetic mean of the bond energies of AA and BB, and the electronegativity difference between atoms A and B [38] .

$$D(A\text{-}B) = \tfrac{1}{2}\left[D(A\text{-}A) + D(B\text{-}B)\right] + 23.06\,(x_A \sim x_B)^2$$

The quantities calculated by this method are not dissociation energies, and depend on the values used for the C-C, N-N, and O-O bonds, but they illustrate the trend in the dissociation energies found experimentally.

Iodine Bonds :	C , N , O	60, 54, 58.5 k.cal.
Bromine Bonds :	C , N , O	67.5, 54, 52 k.cal.

Equilibrium Constants and Ionisation Constants.

The equilibrium constants of halogenation

$$K_e = \frac{(RX)(X^-)(H^+)}{(RH)(X_2)} \qquad \text{(activities)}$$

can also be considered in terms of the ionisation constants for the hydrogen and halogen bonds K_H and K_X and of the ionisation constant of the halogen itself K_{X_2}

$$K_H = \frac{(R^-)(H^+)}{(RH)} \qquad K_X = \frac{(R^-)(X^+)}{(RX)} \qquad K_{X_2} = \frac{(X^-)(X^+)}{(X_2)}$$

$$K_e = \frac{K_H K_{X_2}}{K_X}$$

As was explained on p. 86 this is an equivalent approach to the interpretation of the equilibrium data. The ionisation constants of the halogens are calculated in Part II, p. 116 They are found to have the following values at 25°C :

$$K_{I_2} = 10^{-8} \quad \text{and} \quad K_{Br_2} = 10^{-21}$$

(Whatever the exact values of K_{X_2} , the relative values of K_X are correct) .

The acid dissociation constants of ketones can be estimated from the rates of halogenation [1, 2] .

For each ketone, there exists a relation between rate constants k and dissociation constants K of catalysts.

$k = GK^{\alpha}$ for a given substrate G, α are constants.

The general reactivity of each ketone is conveniently expressed by the rate constant R for a hypothetical anion catalyst $K = 10^{-4}$. Values of R and of the Brønsted exponent α are given in Table 22.

For a series of substrates with the same catalyst, a logarithmic relationship between the rate constants of halogenation R and the acid dissociation constants of the substrates, K_H

$$\log_{10} K_H = \int \frac{1}{\alpha} d \log_{10} R + A$$

is used [2], where α, the exponent of the Brønsted relation, can be expressed in terms of R.

The integral can be evaluated graphically and the values of A can also be obtained, since K_H has been measured directly for acetoacetic ester and acetylacetone [12]. The values estimated for K_H by this method agree well with the directly measured values for acetylacetone and acetoacetic ester at one end of the series, and with the experimentally estimated value for acetophenone ($K_H = 10^{-19}$) [37] at the other end of the range.

The acid dissociation constants of the nitroparaffins have been measured directly [17, 18].

The ionisation constants for the halogen bonds can now be calculated

$$K_X = K_{X_2} \cdot K_H / K_e$$

Table 20 shows that the ionisation constants of the carbon-iodine bonds alter in the same order as the carbon-hydrogen ionisation constants.

Table 20.

Substance.	K_e	K_H	K_I
Cyclopentanone	700	10^{-20}	10^{-31}
Acetone	29.2	10^{-20}	10^{-30}
Cyclohexanone	1170	10^{-18}	10^{-29}
Sym-Dichloracetone	0.203	10^{-16}	10^{-24}
2-Carbethoxycyclohexanone	0.505	10^{-13}	10^{-21}
Malonic ester	0.0523	10^{-13}	10^{-20}
Acetoacetic ester	0.250	$2.11.10^{-11}$	10^{-18}
2-Carbethoxycyclopentanone	1.90	10^{-10}	10^{-18}
Acetylacetone	1.12	$1.17.10^{-9}$	10^{-17}
Nitromethane	0.03	$6.15.10^{-11}$	10^{-17}
Nitroethane	0.09	$3.50.10^{-9}$	10^{-16}
1:Nitropropane	0.14	$1.1 .10^{-9}$	10^{-16}
2:Nitropropane	0.004	$2.12.10^{-8}$	5.10^{-14}

Graph 9 illustrates this proportionality. The data on bromination are not complete or very accurate, but the same trend may be observed.

FIGURE 9

IONISATION OF C-H AND C-I BONDS

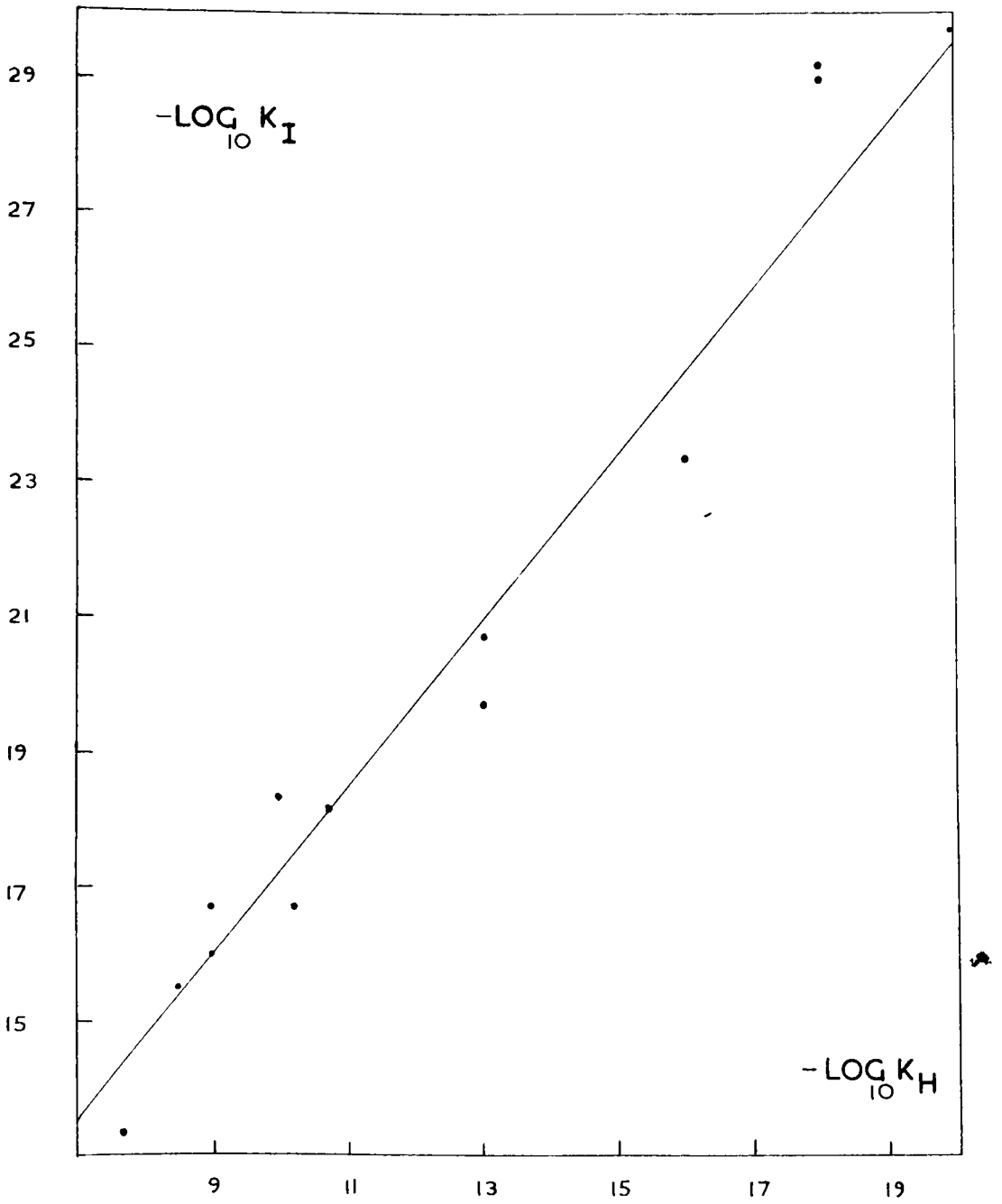


Table 21.

Substance	K_e	K_H	K_{Br}
Acetone	10^{13}	10^{-20}	10^{-54}
2.Carbethoxy cyclohexanone	10^{12}	10^{-13}	10^{-46}
Malonic ester	10^{9}	10^{-13}	$\sim 10^{-43}$
Acetoacetic ester	10^{10}	$2.11.10^{-11}$	
2.Carbethoxy-cyclopentanone	10^{13}	10^{-10}	
Acetylacetone	10^{13}	$1.17.10^{-9}$	
Nitromethane	10^{11}	$6.15.10^{-11}$	

The relation of K_H and K_x can best be seen from a consideration of the free energies of ionisation and free energies of the various steps in the following thermochemical cycle (Z represents H, I, or Br)

$$RZ_{aq} \xrightarrow{\Delta F_2} R^-_{aq} + Z^+_{aq}$$

$$RZ_{aq} \xrightarrow[1]{} RZ_g \qquad RZ_g \xrightarrow[2]{} R_g + Z_g$$

$$R_g \xrightarrow[3]{} R^-_g \xrightarrow[4]{} R^-_{aq} \qquad Z_g \xrightarrow[5]{} Z^+_g \xrightarrow[6]{} Z^+_{aq}$$

$$\Delta F_z = - RT \ln K_z = \Delta F_1 + \Delta F_2 \ldots + \Delta F_6$$

The free energy of ionisation is seen to depend on

(a) the free energy of vaporisation of the organic compound;

(b) the dissociation energy of the carbon-hydrogen, or carbon-halogen bond;

(c) the electron affinity and solvation free energy of the organic radical;

(d) the ionisation potential of the hydrogen or halogen atom and the free energy of solvation of the ions.

(a) is not known for the individual compound, but is only of the order of a few k.cal.

(b) It was seen that the difference in the C-H and C-X energies can be obtained for individual compounds. As yet, only average values are available for the separate bond energies. (Entropies can be calculated from the formulae of statistical mechanics).

(d) The ionisation potentials are accurately known for hydrogen and the halogens. The solvation energy of the hydrogen ion has been estimated, while those for the halogen cations have been calculated in Part II of this Thesis, p. 116-120.

It is therefore seen that if individual bond energies were available, (c), the (electron affinity + solvation energy) of organic radicals could be obtained.

The magnitude of the above quantities indicates that the greater tendency to ionisation of the C-H bond compared with the halogen bonds is primarily due to the large solvation energy of the hydrogen ion; that the larger degree of ionisation of the C-I bond compared with the C-Br bond

is due to the lower ionisation potential of iodine rather than the lower bond dissociation energy, and that in general the solvation energies and electron affinities and work functions can be of greater influence than the bond energies.

Table 22.

Substance	K_H	$\log_{10}R$	% enol	α	$K_e(I_2)$
Cyclopentanone	10^{-20}		$4.8.10^{-3}$		700
Acetone	10^{-20}	-6.78	$2.5.10^{-4}$	.88	29.2
Cyclohexanone	10^{-18}	-6.20	$2.4.10^{-2}$		1170
Sym-dichloracetone	10^{-16}	-2			0.203
2-Carbethoxy-cyclohexanone	10^{-13}	-0.98	7.3	.63	0.505
Malonic ester	10^{-13}	-0.76	0	.79	0.0523
Acetoacetic ester	$2.11.10^{-11}$	+0.72	.33	.59	0.250
2-Carbethoxy-cyclopentanone	10^{-10}	+1.18	.25	.58	1.90
Acetylacetone	$1.17.10^{-9}$	+1.54	17	.48	1.12
Malonate ion	$2.03.10^{-6}$				1
Iodomalonic ester					10^{-7}
α-Iodoacetoacetic ester					10^{-4}
α-Iodoacetylacetone					10^{-5}
Bromomalonic ester		+0.72			
Nitromethane	$6.15.10^{-11}$		$1.1.10^{-5}$		0.03
Nitroethane	$3.50.10^{-9}$		$8.9.10^{-3}$		0.09
1-Nitropropane	$1.1 .10^{-9}$		$7.0.10^{-3}$		0.13
2-Nitropropane	$2.12.10^{-8}$		$2.75.10^{-1}$		0.004
Cyanacetic ester	10^{-11}				0.00841
Acetamide					3.10^{-5}

Part II.

Aqueous Solutions of the Halogens.

Thermochemical Calculations.

In Part I of this thesis the measurement of halogenation equilibria has been described. It was seen that allowance has to be made for any equilibria set up by reaction of the halogen with halide ion or with water. It was also seen in the Discussion p. 86 , that there is an analogy between the tendency of a positive halogen compound to yield a halogen cation, and of an acid to give rise to a proton.

It is therefore intended to give further consideration to the various species present in aqueous solutions of the halogens and to the equilibria which are established between them.

In aqueous solutions of the halogens, the following equilibria have been investigated : -

(1)	$X_2 + H_2O$	=	$H^+ + X^- + HOX$	[38-42]
(2)	$X_2 + X^-$	=	X_3^-	[6,7,43,44]
(3)	$X_2 + n(H_2O)$	=	$X_2 (H_2O)n$	[45]
(4)	$2X_2 + X^-$	=	X_5^-	[7]
(5)	HOX	=	$H^+ + OX^-$	[46,47]
(6)	X_2	=	$X^+ + X^-$	[57,58]
(7)	HOX	=	$X^+ + OH^-$	[57,58]

In the E.M.F. measurements reported, only reaction (2) had to be allowed for.

The Hydrolysis of the Halogens.

The hydrolysis equilibrium constants are known at several temperatures for bromine [41] and chlorine [42] . In the case of iodine the only accurate value is for 25°C. It was obtained by conductivity measurements [38] as for the other halogens, and recently confirmed by glass electrode studies [40] . The results of Popescu, who used a doubtful titration technique, must be rejected in view of the method, and the apparent neglect of reaction (2) . [39]

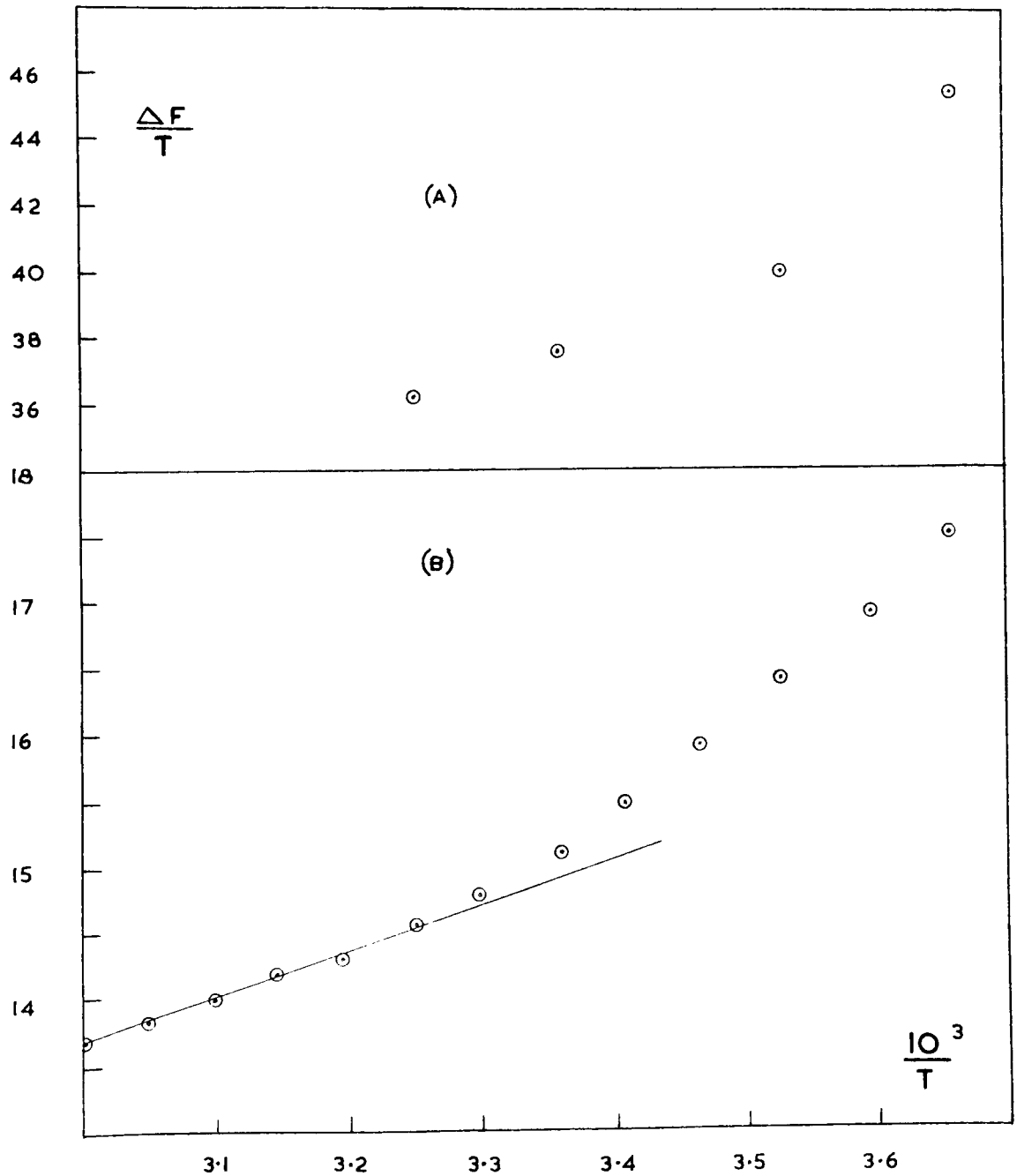
The hydrolysis equilibrium constants at 25°C are 3.10^{-13} for iodine, $5.8.10^{-9}$ for bromine, and $4.8_4.10^{-4}$ for chlorine. Equilibrium measurements at other temperatures are shown in Graph 10, in which $\frac{\Delta F}{T}$ is plotted against the reciprocal of the absolute temperature. (ΔF is the free energy of hydrolysis reaction (1)). According to the Gibbs-Helmholtz equation, the slope of this plot should give the heat of hydrolysis ΔH.

The anomalous temperature variation of the hydrolysis equilibrium constants has been interpreted in terms of reaction (3) [45] . It seems probable, in any case, that the reaction is a composite one . It is therefore uncertain whether the values of heats of reaction, derived from the data for higher temperatures, correspond to the heats of hydrolysis of reaction (1) at 25°C. The values agree however with the ones derived from thermochemical measurements at 25°C. (p. 113)

FIGURE 10

HYDROLYSIS OF THE HALOGENS

(A) BROMINE (LIEBHAVSKY [41])
(B) CHLORINE (JAKOWKIN [42])



Iodine $\Delta H(l) = 15.1$ k.cal.

Bromine $\Delta H(l) = 9.4$ k.cal.

Chlorine $\Delta H(l) = 3.$ k.cal.

The value for iodine is based on the free energy change at 25°C, and an entropy change estimated from the data on bromine and chlorine.

An appreciable ionisation according to reactions (6) or (7) would of course affect equilibrium (1), but it will be seen later that this is not sufficiently important.

Several methods have been used to investigate equilibria (2) and (3) at several temperatures. They have been mainly partition and distribution measurements. The equilibrium constants for reaction (2) at 25°C are 714 for iodine [6], 16.0 for bromine [7] and 0.01 for chlorine [43]. Ions of the type X_5^- are not important - the equilibrium constant for reaction (4) for bromine is 40 [7].

The acid ionisation constants for reaction (5) have been measured by electrometric methods. The values of at 25°C are 2.10^{-11} for iodine [47], $1.4.10^{-9}$ for bromine [46] and 3.10^{-8} for chlorine [46]. The heats of this reaction could be estimated roughly from the free energy data, on the plausible assumption that reactions (1) and (5) involve similar entropy changes.

(1) $X_2 + 2H_2O = X^- + H_3O^+ + HOX$

(5) $HOX + H_2O = OX^- + H_3O^+$

At 25°C, $\Delta H_5 = 12$, 9.5, and 7 k.cal. for iodine, bromine, chlorine respectively.

Apart from direct equilibrium studies at several temperatures, indirect methods of estimating the heats of these various reactions are also available.

Thus, the heat of hydrolysis (1) is by Hess' Law equal to the sum of the heat of ionisation of water (13.7 k.cal.) and the heat of reaction.

(8) $$X_2 + OH^- = XOH + X^-$$

The latter has been obtained in the case of iodine, from a study of kinetics. [48] . This leads to the temperature coefficients for the forward and back reaction; and by applying the Van't Hoff isochore to these data, we obtain $\Delta H_{(1)}$ =14.4 k.cal.

The heat of hydrolysis can also be obtained indirectly from the thermochemical measurements of the heat of reaction [49-51] .

(9) $$X_2 + 2NaOH = NaOX + NaX + H_2O$$

$$\Delta H(1) = 13.7 + \Delta H(8)$$

and $$\Delta H(8) = \Delta H(9) - \Delta H(5) + 13.7$$

For example, in the case of bromine $\Delta H(9)$ - 8.7 k.cal. hence $\Delta H(8)$ = - 4.5 k.cal. and $\Delta H(1)$ = 9.2 k.cal.

From these direct and indirect methods, the heats of hydrolysis at 25°C are taken to be

Iodine	14.7	k.cal.
Bromine	9.2	k.cal.
Chlorine	3.0	k.cal.

The Dissociation Energy of the Oxygen-Halogen Bonds.

The heat of the hydrolysis reaction (1) is the sum of the heat content changes of

$$X_2\,aq \xrightarrow[1]{} X_2\,g \xrightarrow[2]{} 2X_g \qquad H_g + X_g \xrightarrow[3]{} HX_{aq}$$

$$H_2O_l \xrightarrow[4]{} H_2O_g \xrightarrow[5]{} H_g + OH_g \qquad X_g + OH_g \xrightarrow[6]{} XOH_g$$

$$XOH_g \xrightarrow[7]{} XOH_{aq}$$

$$\Delta H = \Delta H_1 + \Delta H_2 \ldots\ldots + \Delta H_7$$

Using the values for ΔH selected above, the best thermochemical data from the National Bureau of Standards, and for $\Delta H_{\underline{5}}$ the value obtained by Dwyer & Oldenburg

we have for the hydrolysis of iodine at 25°C

$$14.7 = 9.9 + 36.1 - 90.9 + 10.6 + 120.7 - (D + Q)_I$$

for the bromine hydrolysis :

$$9.2 = 8.4 + 46.1 - 107.7 + 10.6 + 120.7 - (D + Q)_{Br}$$

and for the chlorine hydrolysis :

$$3.0 = 6.9 + 58.0 - 121.0 + 10.6 + 120.7 - (D + Q)_{Cl}$$

where D is the heat of formation of the O-X bond in step 6

and Q is the heat of solution of HOX (gas) in step 7 .

$(D + Q)$ = 71.7, 68.8, and 72.2 k.cal./mole. respectively.

The heats of solution of HOX can only be estimated theoretically

As their value is only a small fraction of the value of (D + Q), the following expression is probably sufficiently accurate

$$Q_{HOX} = \tfrac{1}{2}(Q_{X_2} + Q_{H_2O}) + 4.0 \quad \text{k.cal.}$$

Q_{X_2} is the heat of solution of the corresponding halogen, and Q_{H_2O} the heat of condensation of water. The last factor of 4.0 k.cal. is the contribution from the additional hydrogen bond.

Q_{HOI} = 14.3 k.cal., Q_{HOBr} = 13.5 k.cal., Q_{HOCl} = 12.8 kcal/mole

and hence we have for the dissociation energies of the oxygen-halogen bonds

D_{HO-I} = 57.4, D_{HO-Br} = 55.3, and D_{HO-Cl} = 59.4 k.cal./mole.

These values are accurate to about 3 k.cal. No reliable values seem to have been reported for any oxygen-halogen bonds previously. The value of 49.3 k.cal. for the O-Cl bond dissociation energy in Cl_2O depends on an uncertain interpretation of optical data [52].

The size of these dissociation energies and their variation is discussed in connection with the carbon and nitrogen-halogen bonds on p. 101.

The Ionisation of the Halogens.

The general chemical properties of iodine, its basic character, and electrochemical behaviour [54] have led to numerous suggestions that the iodine cation was formed in aqueous solution by ionisation. Very recently, qualitative electrophoretic experiments have established the presence of bromine cations in aqueous solutions of hypobromous acid [56]. The role of halogen cations as intermediates in kinetic mechanisms has been discussed by several authors [55, 67-69].

There has been no reliable measurement of the ionisation constants of the halogens. The interpretation of Murray's electrochemical measurements on the ionisation of iodine [57] and of Winther's distribution experiments [58] appears to be unsatisfactory. So are the conlusions of Sidgwick on the upper limit of ionisation of HOX [59].

$$HOX = HO^- + X^+$$

and

$$HOX = H^+ + OX^-$$

He attempts to obtain the limit of the concentrations of the above species (set by the ionic product of water) by treating the above equilibria separately; this is unjustifiable.

The following thermochemical calculations are an attempt to obtain these ionisation constants from first principles and to illustrate the importance of halogen cations in equilibrium studies in aqueous solution.

In the gas phase, the reaction

(6a) $$X_2 = X^+ + X^-$$

has a Gibbs free energy ΔF = 195.4, 231.1, and 263.4 k.cal. for iodine, bromine, and chlorine respectively. This is calculated from the free energy of dissociation, the ionisation potential, and the electron affinity. It is seen that the tendency to ionise in the gas phase is extremely small. It can become appreciable in aqueous solution, as will be shown below, because of the large heats of solution of the ions.

Consider the reaction in aqueous solution

(6b) $$X_2 = X^+ + X^-$$

which is the sum of the following steps

$$X_{2\,aq} \xrightarrow[1]{} X_{2\,g} \xrightarrow[2]{} 2X_g \xrightarrow[3]{} X_g^+ + X_g^-$$

$$X_g^+ \xrightarrow[4]{} X_{aq}^+ \qquad\qquad X_g^- \xrightarrow[5]{} X_{aq}^-$$

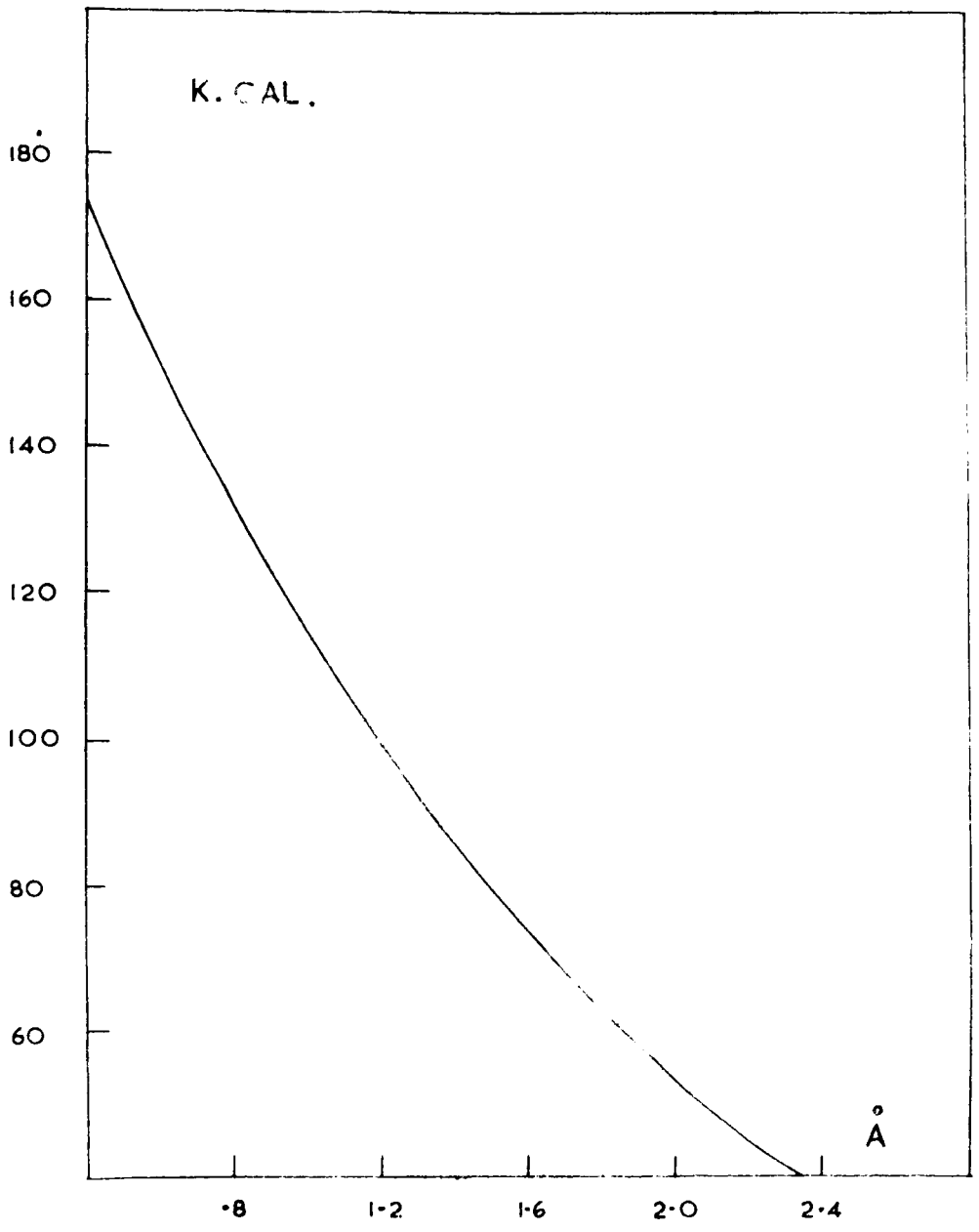
The free energies for steps 1 to 3 are well known. Those for step 4 and 5 are obtained by the method of Eley & Evans [60]. These authors have calculated the heats and entropies of solution of ions from first principles; thermochemichal quantities for other ions can be estimated from their radii with the help of these data. Graph 11 shows the variation of the heat of solution of spherical monovalent ions with radius.

The radii of I^+, Br^+, and Cl^+ are estimated to be 1.12 Å, 0.95 Å, and 0.83 Å respectively from the screening constants given by Pauling [61.]. These positive ions

FIGURE 11

HEAT OF SOLUTION OF MONOVALENT POSITIVE IONS IN RELATION TO THEIR RADIUS.

(ELEY & EVANS [60]).



have an incomplete electron shell and it is very probable that the octet is completed by a water molecule giving $(H_2OX)^+$

It is however of interest to estimate the ionising tendency if the cation is the bare ion X^+

$$\Delta F = \Delta F_1 + \Delta F_2 \quad \ldots\ldots + \Delta F_5$$

For iodine, (25° C),

$\Delta F = 0.7 + 28.9 + (242.2 - 75.7) - 99 - 43 = 54$ k.cal.

for bromine,

$\Delta F = -0.2 + 38.6 + (274.6 - 82.1) - 109 - 53 = 69$ k.cal.

and for chlorine,

$\Delta F = -1.0 + 50.4 + (300.3 - 87.3) - 117 - 58 = 87$ k.cal.

The values for the free energy of solution of the positive ions are only very approximate, because they have been interpolated from data calculated for ions with complete electron shells.

The tendency to ionisation is greatly increased if the positive ion has the structure $(H_2OX)^+$

(6c) $$H_2O_l + X_{2\ aq} \longrightarrow (H_2OX)^+_{aq} + X^-_{aq}$$

$$H_2O_l \xrightarrow[1]{} H_2O_g \qquad X_{2\ aq} \xrightarrow[2]{} X_{2\ g} \xrightarrow[3]{} 2X_g \xrightarrow[4]{} X_g^+ + X_g^-$$

$$X_g^+ + H_2O_g \xrightarrow[5]{} (H_2OX)^+_g \xrightarrow[6]{} (H_2OX)^+_{aq}$$

$$X^-_g \xrightarrow[7]{} X^-_{aq}$$

There is little evidence to show whether the dissociation energy of the O-X bond in $(H_2OX)^+$ is greater or smaller than in HOX . Some support for the view that it is a few per cent lower comes from the values of force constants of pairs of molecules AH_n and $\overset{+}{A}H_{n+1}$ [62] . However, the symmetry of these molecules is different, and it is in any case dangerous to draw definite conclusions about bond energies from force constant data. [63]

The values of the entropy terms $T\Delta S$ for the dissociation

$$HOX \longrightarrow HO + X$$

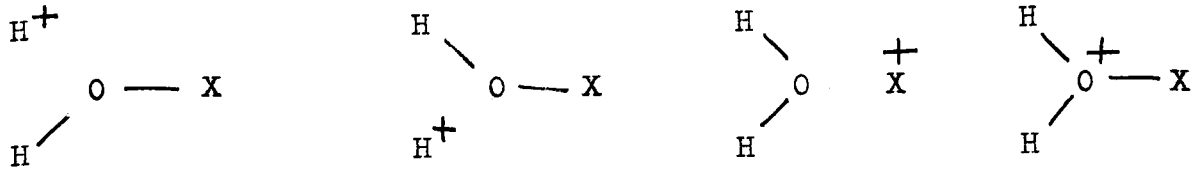
have already been quoted. Taking the dissociation energy in $(H_2OX)^+$ equal to that in HOX, we have

$$\Delta F_5 = \begin{cases} -46 \text{ k.cal. for iodine} \\ -45 \text{ k.cal.. for bromine} \\ -49 \text{ k.cal. for chlorine} \end{cases}$$

These values are, if anything, a few k.cal. too high.

The further solvation of the cation by coulombic interaction with water molecules can be treated approximately by the method of Eley & Evans. In this treatment the coordination of four water molecules with a monovalent ion is considered.

The positive charge of the ion $(H_2OX)^+$ may be regarded as distributed according to the structures



with due regard for the different electro-negativities of the atoms

$$Q_{(H_2OI)^+} = \left(0.5\ Q_{H_3O^+} + 0.2\ Q_{I^+}\right) = 103 \text{ k.cal.}$$

and similarly,

$$Q_{(H_2OBr)^+} = 111 \text{ k.cal.}, \qquad Q_{(H_2OCl)^+} = 121 \text{ k.cal.}$$

ΔF_6 = 98, 105, and 115 k.cal. respectively.

We now have for the ionisation of

Iodine

$$\Delta F = \Delta F_1 + \Delta F_2 + \Delta F_3 + \Delta F_4 + \Delta F_5 + \Delta F_6 + \Delta F_7$$

$$= 2.0 + 0.7 + 28.9 + (242.2 - 75.7) - 46 - 98 - 43$$

$$= \underline{11 \text{ k.cal.}}$$

Bromine

$$\Delta F = 2.0 - 0.2 + 38.6 + (274.6 - 82.1) - 45 - 106 - 53$$

$$= \underline{29 \text{ k.cal.}}$$

Chlorine

$$\Delta F = 2.0 - 1.0 + 50.4 + (300.3 - 87.3) - 49 - 115 - 58$$

$$= \underline{42 \text{ k.cal.}}$$

The Van't Hoff isotherm gives the following equilibrium constants of ionisation, K_{X_2} (Reaction 6, p. 118)

for Iodine : 10^{-8}

Bromine : 10^{-21}

Chlorine : 10^{-31}

The non-systematic errors in the above calculation are smaller than the systematic ones, The latter cannot be very considerable in a direction leading to greater ionisation, as an ionisation of iodine higher than that calculated would have been observed experimentally. The above discussion illustrates the factors which can lead to appreciable ionisation.

It may be noted that other reactions producing the cation

e.g. $H^+ + HOX \rightarrow (H_2OX)^+$

have free energies and heats connected with those of the other equilibria by the additive laws of thermochemistry.

From the above calculations it appears that the basic ionisation constant of HOI is greater than the acid one ($K_5 = 2.10^{-11}$)

(7) $$HOX = X^+ + OH^-$$

$$K_7 = \frac{K_w K_6}{K_1} = \frac{10^{-14} \cdot 10^{-8}}{10^{-13}} = \frac{10^{-9}}{\quad}$$

A consideration of the various equilibria occurin-g in an aqueous solution of iodine shows that if there is an appreciable iodide concentration, ionisation has a very small effect in altering the concentrations of the species involved in the other equilibria. Even when no extra iodide is added, the error in the equilibrium constant of hydrolysis (as determined for example by conductivity measurement) is within experimental error. On the other hand, in acid solution in the presence of

silver nitrate, almost all the iodine is present in the form of the cation. (It would be interesting to study the absorption spectrum under these conditions [64]). A very useful method of halogenation is based on the removal of halide anions by silver ions [66] .

These comments may be illustrated by the following examples in which the approximate concentrations of the various species have been calculated from the equilibrium constants quoted above, and with the accepted values for the solubility products of the silver halides [65] .

Table 23.

I.	M/1000 X_2 ,	M/10 $AgNO_3$,	N H_2SO_4		
		X_2	$(H_2OX)^+$	X^-	HOX
Iodine		10^{-10}	10^{-3}	10^{-15}	10^{-8}
Bromine		10^{-5}	10^{-16}	10^{-10}	10^{-3}
Chlorine		10^{-9}	10^{-31}	10^{-9}	10^{-3}
II.	M/1000 X_2				
Iodine		10^{-3}	10^{-6}	10^{-5}	10^{-6}
Bromine		10^{-3}	10^{-21}	10^{-4}	10^{-4}
Chlorine		10^{-5}	10^{-53}	10^{-3}	10^{-3}
III.	M/1000 X_2 ,	M/10 X^-			
Iodine		10^{-3}	10^{-10}	10^{-1}	10^{-10}
Bromine		10^{-3}	10^{-23}	10^{-1}	10^{-5}
Chlorine		5.10^{-4}	10^{-34}	10^{-1}	5.10^{-4}

The Kinetic Significance of Halogen Cations.

Evidence for the participation of halogen cations in halogenation reactions has only recently been obtained [66-69] It comes from studies of the halogenation of organic compounds by hypohalous acids in aqueous solution.

Derbyshire and Waters [66] found rates of bromination which showed a marked dependence on the hydrogen ion concentration. The rate determining step is the reaction of the halogenating species with the substrate, and from the kinetics under varying conditions it is possible to deduce that a considerable fraction of halogenation is due to the halogen cation.

De la Mare, Hughes, and Vernon [69] have recently been studying chlorination in hypochlorous acid solution. They find that with some reactive substrates, a first order reaction occurs, which is independent of the substrate studied, and which they suppose to be the formation of a chlorine cation

$$Cl - Y = Cl^{+} + :Y \qquad (Y = OH, \overset{+}{O}H_2)$$

The rate determining step is considered to be the dissociation of the oxygen-chlorine bond. With less reactive substrates there is a change to a second order mechanism, the rate determining step now being the reaction between the substrate and the halogenating species.

These studies have been carried out under conditions which favour the presence of halogen cations.

Thermochemical calculations of the type described above have an indirect bearing on these kinetic problems.

In solutions of a hypohalous acid we have the equilibrium

(10) $$HOX + H_3O^+ = (H_2OX)^+ + H_2O$$

The equilibrium constants of this reaction $K_{10} = K_6/K_1 = 10^5$, 10^{-12}, and 10^{-27} for iodine, bromine, and chlorine respectively.

Since there are equal numbers of reactants and products in this reaction, there will be only a small entropy change. Although there may be no appreciable potential barrier in this prototropy, the great increase in the endothermicity of the reaction on passing from bromine to chlorine makes it seem quite possible that reaction (10) is sufficiently slow to be rate determining in some chlorinations.

$$HOX + H^+ = (H_2OX)^+$$

1

$$(H_2OX)^+ + RH = H_2O + RX + H^+$$

2

The reactivity of the halogen cations increases from iodine to chlorine. Reaction 2 becomes faster and reaction 1 slower, so that a change from second to first order kinetics can occur.

The transition state in halogenation may be represented as

$$Y \ldots\ldots X \ldots\ldots R^{-}$$

where $Y = X^{-}$, HO^{-}, or H_2O

The concentrations of the various species in aqueous solutions of the halogens have been calculated above (p. 122).

The insertion of these concentrations into the simple Arrhenius rate expressions shows that, if halogenation is to be preponderantly due to halogen cations, the activation energy of the reaction where $Y = X^{-}$ (free halogen) must be greater than the activation energy where $Y = H_2O$ (halogen cation) by the following amounts

I.

Iodine	$E_{I^-} - E_{H_2O}$	> 1	k.cal.	approx.
Bromine	$E_{Br^-} - E_{H_2O}$	> 4	k.cal.	"
Chlorine	$E_{Cl^-} - E_{H_2O}$	> 7	k.cal.	"

The halogen cations can be considered kinetically significant, if the differences are greater than the above values. The actual differences can be estimated by means of the semi-empirical method of Hirschfelder, which relates free energies of reaction to the dissociation energies of the bonds broken in reaction [70] .

e.g. (a) $X_2 + RH = X^- + RX + H^+$

(b) $(H_2OX)^+ + RH = H_2O + RX + H^+$

$$E_{X^-} - E_{H_2O} = 0.28 \quad (D_{X^-} - D_{H_2O})$$

where D_{X^-} is the heat of the reaction

$$X_2 = X^- + X^+ \text{(bare)} \qquad \text{aq.}$$

and D_{H_2O} is the heat of the reaction

$$(H_2OX)^+ = H_2O + X^+ \qquad \text{aq.}$$

The heats of the latter reaction may be obtained from the cycle

$(H_2OX)^+$ aq $\rightarrow$ $(H_2OX)^+$ g $\rightarrow H_2O_g + X_g^+ \rightarrow$ H_2O aq + X^+ aq

but are in any case not greater than the dissociation energies of HOX (p. 115) .

The heats of ionisation of the halogens may be calculated by means of the cycles used for deriving the free energies of ionisation (p. 118) .

For analogous calculations with the hypohalous acids, the heats of dissociation D_{HO^-} can be obtained from the cycles

HOX aq $\rightarrow$ HOX g $\rightarrow HO_g + X_g \rightarrow HO_g^- + X_g^+ \rightarrow HO^-_{aq} + X^+_{aq}$

using the values calculated for the heats of solution of HOX (p. 115) the dissociation energies of HOX (p. 115) and the

heats of solution of X^+ (p. 118), as well as the value for the (electron affinity + heat of solvation) of OH^- given by Evans and Uri [71].

		D_{HO^-}	D_{X^-}	D_{H_2O}	$E_{X^-} - E_{H_2O}$
	Iodine	62	62	57	$1\frac{1}{2}$ k.cal.
II.	Bromine	82	76	55	6 k.cal.
	Chlorine	103	94	59	10 k.cal.

These calculations are approximate regarding equilibrium data, and because of the uncertainty in the charge distribution and solvation effects in the transition state.

It is seen that the values calculated for $(E_{X^-} - E_{H_2O})$ in II exceed the minimum values for this quantity (I), p. 125 required by the Arrhenius rate expression. Subject to the reservations as to the accuracy of the above calculation, the kinetic significance of the halogen cations has thus been illustrated.

Summary.

Equilibria in the halogenation of compounds containing C-H , N-H , and O-H bonds have been studied.

1. Electromotive force measurements have been made on cells of the type

Pt, X_2 , X^-, (Buffer) / 3.5 N KCl / (Buffer), X^-, X_2 , RH, Pt

where X_2 , X^- represent iodine or bromine and the corresponding anion

RH an organic compound with a reactive C-H or N-H bond.

From these data the thermodynamic equilibrium constants of the reaction

$$RH + X_2 = RX + X^- + H^+$$

have been obtained for 20 organic compounds, in some cases at several temperatures.

2. From the free energies and heats of halogenation together with other thermochemical data, the differences in the dissociation energy of the C-H and C-X bonds or the N-H and N-X bonds in the compounds RH and RX have been obtained. Their variation is interpreted in terms of the structural changes in the organic compounds.

3. Thermochemical calculations on the hydrolysis of the halogens have yielded dissociation energies of oxygen-halogen bonds.

4. The equilibrium constants of halogenation have also been considered in terms of the ionisation constants of the hydrogen and halogen bonds and of the halogen itself : -

$$K_e = \frac{(H^+)(X^-)(RX)}{(RH)(X_2)} = \frac{(R^-)(H^+)}{(RH)} \cdot \frac{(RX)}{(R^-)(X^+)} \cdot \frac{(X^-)(X^+)}{(X_2)}$$

5. In this connection the evidence for the existence of halogen cations in aqueous solution is reviewed. Their free energies of formation have been calculated, and their significance in both kinetic and equilibrium problems has been discussed.

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