

DISSOCIATION CONSTANTS OF INORGANIC
ACIDS AND BASES IN AQUEOUS SOLUTION

CONSTANTES DE DISSOCIATION
DES ACIDES ET DES BASES INORGANIQUES
EN SOLUTION AQUEUSE

DIVISION DE CHIMIE ANALYTIQUE
COMMISSION DE CHIMIE ÉLECTROANALYTIQUE*

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ANALYTICAL CHEMISTRY DIVISION
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INTRODUCTION

Most of the existing tables of dissociation constants of inorganic acids and bases in aqueous solution are fragmentary in character, include little or no experimental details, and give few references. Easily the most comprehensive of the previous collections is *Stability Constants of Metal-Ion Complexes*, compiled by L. G. Sillén and A. E. Martell, and published as Special Publication No. 17 of the Chemical Society, London, in 1964. However, because of the nature of this compilation, the pK values in it tend to be overlain by the much greater bulk of the stability constant data. In many cases, also, it is difficult to decide by inspection which of the pK values should be taken from the wide range sometimes given for a particular substance.

The present Table follows the pattern of the similar Tables for organic acids and organic bases, which were also prepared at the request of the International Union of Pure and Applied Chemistry as part of the work of the Commission on Electrochemical Data. The Table of organic acids, compiled by Kortum, Vogel, and Andrussov was published in *Pure and Applied Chemistry*, **1**, 187–536 (1960), and also separately as a book*. The Table of organic bases, by the present author, was published in 1965 as a supplement to *Pure and Applied Chemistry*.†

For convenience, the dissociation constants of inorganic acids and bases have been given, in most cases, in the form of pK_a values, and the classes of compounds include not only conventional acids and bases such as boric acid and magnesium hydroxide, but also hydrated metal ions (which behave as acids when they undergo hydrolysis) and free radicals, such as the hydroxyl radical, $\cdot\text{OH}$. All of these reactions have in common the gain or loss of a proton or a hydroxyl ion. On the other hand, the hydrolyses of metal-complex ions such as the cobaltammines have been excluded, as being more appropriate to the stability constant compilation mentioned above.

In general, and largely because of the difficulties attending pK measurements on inorganic species, it is not possible to offer a critical assessment of most of the published values. In particular cases, such as water, highly precise constants are available over a range of temperatures, and the uncertainty is only of the order of 0.001 pH unit. More commonly, only a few, often widely discordant, values have been reported.

This is partly because of the chemical reactivity of the materials themselves. For example, nitrous acid readily decomposes to dinitrogen trioxide. At concentrations above 0.01 M, boric acid is appreciably polymerised to polyboric acids; molybdic acid solutions contain $\text{Mo}_7\text{O}_{24}^{6-}$ and higher species; bisulphite ion is in equilibrium with pyrosulphite ion, $\text{S}_2\text{O}_5^{2-}$; and many transition and higher-valent metal ions form polynuclear species on hydrolysis.

* G. Kortum, W. Vogel and K. Andrussov. *Dissociation Constants of Organic Acids in Aqueous Solution*. Butterworth & Co. Ltd., London, 1961.

† D. D. Perrin. *Dissociation Constants of Organic Bases in Aqueous Solution*. Butterworth & Co. Ltd., London, 1965.

Often, too, unsatisfactory methods of determination have been used. Thus, pH titration measurements are seldom satisfactory if pK values lie below 2 or above 12, and in such circumstances can give quite misleading results. Again, pK values for the hydrolysis of metal ions have often been obtained from measurements of the pH values of solutions of their purified salts in water. As Sillén has pointed out (*Quart. Rev.*, **13**, 146 (1959)), inorganic salts often adsorb tenaciously onto their surfaces traces of acidic or basic impurities, which persist even on repeated recrystallization, so that the measured pH values of their solutions may be much higher or lower than expected.

Even with experimentally accurate results, extrapolation to thermodynamic pK values at $I = 0$ is not always possible. The usual basis of such extrapolation is the Debye-Hückel equation,

$$-\log f_{\pm} = \frac{Z_i^2 AI^{\frac{1}{2}}}{1 + \kappa a I^{\frac{1}{2}}} - bI$$

which is used to calculate the activity coefficient term. For precise work, values of a (the "mean distance of nearest approach" of the ions) and b are chosen to fit the data over a range of ionic strengths, so that the value of the pK , extrapolated to $I = 0$, can be obtained. At low ionic strengths and where moderate accuracy (say ± 0.05 pH unit) is sufficient some simplifying assumptions can often be made. Thus, Davies' equation (*J. Chem. Soc.* **1938**, 2093) is obtained by taking $Ka = 1$, $b = 0.2$; Güntelberg's equation (*Z. physik. Chem. Leipzig*, **123**, 199 (1926)) sets $Ka = 1$, $b = 0$; and the approximation $Ka = 0$, $b = 0$ (*i.e.* $-\log f = Z_i^2 AI^{\frac{1}{2}}$) is also used. However, with moderately strong acids and bases (pK values less than 2 or greater than 12), the numerical values of the thermodynamic pK constants depend in part on the assumptions made in deriving them, including the ion-size parameter a used in the extended Debye-Hückel equation (see, for example, R. G. Bates, V. E. Bower, R. G. Canham and J. E. Prue, *Trans. Faraday Soc.*, **55**, 2062 (1959); A. K. Covington, J. V. Dobson and W. F. K. Wynne-Jones, *Trans. Faraday Soc.*, **61**, 2057 (1965), E. A. Guggenheim, *Trans. Faraday Soc.* **62**, 2750 (1966)). Thus, the pK of bisulphite ion at 25° varies from 1.927 to 1.967 as K_a is varied from 1.0 to 1.7. In the same way, pK_b for $\text{Ca}(\text{OH})_2$ varies from 1.14 to 1.27 at 25°, depending on the choice of parameters.

A distinction must also be made between true and apparent pK values. The first pK of carbon dioxide in water as measured is about 6.4 at 20°, whereas the true pK of carbonic acid (H_2CO_3) is 3.8. The difference between the apparent and the true pK values is due to the slight extent to which carbon dioxide is covalently hydrated in water. Similarly, periodic acid exists as H_5IO_6 and HIO_4 (mainly as the latter), so that its measured second pK (8.3) is very much higher than its first one (about 2).

In the absence of experimental values, especially for some of the oxyacids, attempts have been made to predict pK values, usually from similarities of structure. The more commonly used methods are those of J. E. Ricci (*J. Am. Chem. Soc.*, **70**, 109 (1948)), L. Pauling (*General Chemistry*, Freeman, San Francisco, 1947, p. 394), and A. Kossiakoff and D. Harker (*J. Am. Chem. Soc.*, **60**, 2047, (1938)). Even in apparently simple cases, there may be considerable uncertainty. For example different values would be predicted

INTRODUCTION

for germanic acid depending on whether it existed mainly as $\text{GeO}(\text{OH})_2$ or $\text{Ge}(\text{OH})_4$.

Because of the many different kinds of uncertainties inherent in the present $\text{p}K$ compilation, no attempt has been made to assess the accuracy of each entry. Nevertheless, where possible, I have attempted to select what appear to be the best available values. The results for hydrogen sulphide illustrate this. Thus, several methods have indicated that the second $\text{p}K$ of hydrogen sulphide is about 14, which is too high for potentiometric titration methods to be applicable. Hence the $\text{p}K_2$ values that have been obtained by potentiometric titration are not set out in the Table. Instead, references to the papers where they are given are included under "other measurements". This heading also covers results where insufficient experimental details are given.

I. HOW TO USE THE TABLE

GENERAL ARRANGEMENT

The Table summarizes data recorded in the literature up to the end of 1967 for the dissociation constants of inorganic acids and bases in aqueous solution. It also includes references to acidity functions for strong acids and bases, and details about the formation of polynuclear species where this is relevant. The substances are listed alphabetically, with chemical formulae, so that the entries are self-indexing.

Column 1 gives the name of the substance and the negative logarithm of the dissociation constant (pK_a). Wherever possible, these values are thermodynamic ones obtained by extrapolation to ionic strength $I = 0$, generally by using some form of the Debye-Hückel equation such as that due to Davies. In all cases, pK values are listed in decreasing extent of protonation.

Column 2 gives the temperature of measurements in °C.

Column 3 lists details such as:

$I = \frac{1}{2}\sum C_i Z_i^2$ = ionic strength

c = concentration in mole/l, or

m = concentration in mole/1000 g. of water.

It also records any other details relating to the pK value quoted. Designation of a constant as "practical" implies that it includes both the activity of the hydrogen ion (usually as measured by pH meter) and the concentrations of the other species.

Column 4 summarises the method of measurement, the procedure used in evaluating the constants, and any corrections that were taken into consideration; the symbols have the meanings set out under "Methods of Measurement", page 7. Because different investigators rarely use identical procedures, these symbols can only serve as guides: for fullest details the original papers should be consulted.

Column 5 gives the literature references which are listed alphabetically at the end of the Table.

II. METHODS OF MEASUREMENT AND CALCULATION

The abbreviations in *Column 4* of the Table are, with only minor differences, the same as those used in "Dissociation Constants of Organic Bases in Aqueous Solution".

CONDUCTOMETRIC METHODS

- C1 Measurements in solutions of salt and acid
C2 Measurements in solution of base only

ELECTROMETRIC METHODS

[i] Cells without diffusion potentials

- Ela Method of Harned and Ehlers (*J. Am. Chem. Soc.* **54**, 1350 (1932))
(Cell of type Pt (H₂)|B, BCl, NaCl||B, BCl, NaCl|AgCl|Ag, for which $E - E_0 + (RT/F) \ln [\text{BH}^+] [\text{Cl}^-]/[\text{B}] = - (RT/F) \ln K'$, and extrapolate to $I = 0$)
- Elb Method of Harned and Owen (*J. Am. Chem. Soc.* **52**, 5079 (1930)) ,Pt(H₂)|B, NaCl|AgCl|Ag, where molality of B is M , $E = E_0 - (RT/F) \ln ([m_{\text{H}^+}] [m_{\text{Cl}^-}] f_{\pm}^2)$. Extrapolate to $I = 0$ at constant M , then to $M = 0$)
- Elcg Determination of [H⁺] from cells of the type, Glass|solution, Cl-|AgCl|Ag
- Elch Determination of [H⁺] from the cell, Pt(H₂) solution, Cl-|AgCl|Ag
- Eld Method of Bates (*J. Am. Chem. Soc.* **70**, 1579 (1948)). Determination of K_1 and K_2 for dibasic acids
- Ele Method of Bates and Pinching (*J. Res. Natl. Bur. Std.* **43**, 519 (1949)). A particular case of method Elcg in which the solution is a buffer comprising a weak base and a weak acid

[ii] Approximately symmetrical cells with diffusion potentials

- E2a Method of Owen (*J. Am. Chem. Soc.* **60**, 2229 (1938))
- E2b Method of Larsson and Adell (*Z. Physik. Chem.* **156**, 352, 381 (1931)) (Uses cell Pt(H₂)|B, NaCl|sat. KCl|NaOH, NaCl|(H₂)Pt and an approx. K to adjust to equal ionic strengths in the half-cells. From E obtain [H⁺] and hence K' : extrapolation to $I = 0$ gives K)
- E2c Method of Everett and Landsman (*Proc. Roy. Soc. London*, **A215**, 403, (1952))
(This is like E2b but uses a second weak base of known pK instead of a strong base. The method gives the ratio of the two constants)

[iii] Unsymmetrical cells with diffusion potentials

E3ag	pH measurements in buffer solutions of weak electrolytes using glass electrodes
E3ah	Similar measurements using hydrogen electrodes
E3bg	Measurements of pH changes during titrations using glass electrodes
E3bh	Similar measurements using hydrogen electrodes
E3b, quin	Similar measurements using quinhydrone electrodes
E3c	Differential potentiometric methods
E3d	pH measurements at equal concentrations of salt and base

OPTICAL METHODS

01	Direct determination of the degree of dissociation by extinction coefficient measurements in solutions of weak bases and salts
02	Colorimetric determination with an indicator of known pK
03	Colorimetric determination with an indicator calibrated with a buffer solution of known pH
04	Method of von Halban and Brüll (<i>Helv. Chim. Acta</i> 27 , 1719 (1944)) (Solutions of the base being studied, plus indicator, are compared with similar solutions containing alkali and indicator)
05	Light absorption measurements combined with electrometric measurements
06	Light absorption measurements using solutions of mineral acids of known concentrations and (usually) Hammett's acidity function, H_0
07	Similar to 06 but using solutions of alkalis

OTHER METHODS

ANALYT	Constants derived from chemical analysis
CALORIM	Calorimetric measurements
CAT	Constants estimated from catalytic coefficients
CRYOSC	Cryoscopic measurements
DISTRIB	Distribution between solvents
FP	Constants derived from freezing-point data
ION	Ion-exchange studies
KIN	Constants estimated from kinetic measurements
NMR	Nuclear magnetic resonance measurements
POLAROG	Polarographic measurements
RAMAN	Measurements of Raman spectra
REDOX	Oxidation-reduction potentials
SOLY	Solubility measurements
VAP	Vapour pressure measurements

CALCULATIONS

[i] Conductance measurements

R1a Method of Davies (*The Conductivity of Solutions*, Chapman Hall, London 1930)

(By successive approximations, f_A is calculated from the Debye-Hückel-Onsager equation in the form

$$f_A = 1 - A(\alpha c_0)^{\frac{1}{2}}/\Lambda_0$$

which assumes that Λ_0 can be obtained from Kohlrausch's law of independent ionic mobilities)

R1b Method of MacInnes (*J. Am. Chem. Soc.* **48**, 2068 (1926))

(The quantity $\Lambda_e = f_A \Lambda_0$ is determined directly, where Λ_e is the conductance of the weak electrolyte if it were completely dissociated at the ionic strength studied: it is necessary to know Λ for strong electrolytes as a function of I)

R1c Method of Fuoss and Krauss (*J. Am. Chem. Soc.* **55**, 476 (1933))

(The Debye-Hückel-Onsager equation is used in the form, $\Lambda_c = \alpha(\Lambda_0 - A(\alpha c_0)^{\frac{1}{2}})$ to derive an equation relating Λ_0 , c and K , which is solved by successive approximation until Λ_0 is constant at all values)

R1d Method of Shedlovsky (*J. Franklin Inst.* **225**, 739 (1938))

(This is like R1c but a different equation is used)

R1e Method of Fuoss (*J. Am. Chem. Soc.* **79**, 3301 (1957))

[ii] Differential potentiometric measurements

R2a Method of Kilpi (*Z. Physik. Chem.* **173**, 223, 427 (1935); **175**, 239 (1936) (at point of inflection).

III. TABLE

Nos. 1-2

Name, Formula and pK value	T(°C)	Remarks	Methods	Reference
1. (Aquo) Aluminium ion, Al³⁺				
3.00	15	pK for hydrolysis of Al ³⁺ , $\epsilon = 0.0005 - 0.01$ M in AlCl ₃ , extrapolated against <i>I</i> ‡	E3ag	S17
5.15	20			
4.98	25	pK for hydrolysis of Al ³⁺ , <i>I</i> varied from 0.0025 to 0.019, extrapolated to <i>I</i> = 0	E3ag	H41
4.96	25	pK for hydrolysis of Al ³⁺ , $\epsilon = 10^{-5} - 10^{-2}$ M in Al(ClO ₄) ₃ , extrapolated to <i>I</i> = 0	E3ag	F33
5.02	25		E3,quin	K12
4.96	25	pK for hydrolysis of Al ³⁺ ; also log <i>K</i> = 7.55 for 2Al(OH) ₂ ⁺ ⇌ Al ₂ (OH) ₂ ⁴⁺ , and log <i>K</i> = 6.89 for 2Al ₂ (OH) ₄ ²⁺ + 2H ₂ O ⇌ Al ₄ (OH) ₁₀ ²⁺ + 2H ⁺	E3ag	K67
5.03	25	pK for hydrolysis of Al ³⁺	E3AG	I19
5.10	25	pK for hydrolysis of Al ₂ ³⁺	KIN	H64a
4.49	25	pK for hydrolysis of Al ³⁺ , from dissociation field effect relaxation times	KIN	K69
2.88	100	pK for hydrolysis of Al ³⁺	C1	M9
11.22	25	pK for Al(OH) ₃ + H ₂ O ⇌ Al(OH) ₄ ⁻ + H ⁺ Hydrolysis of Al ³⁺ in 2 M NaClO ₄ at 40° gives, mainly, one or more polynuclear complexes Other measurements: B94, D23, F5, I20, L1, T7, W25.		B97
2. Amidophosphoric acid, NH₂PO₃H₂				
3.00	25	<i>I</i> = 0.2 (KCl), "practical" constants	E3bg	C13
8.15	25	<i>I</i> = 1(NMMe ₄ Br), concentration constants, <i>f</i> _± assumed same as for HBr	E3bg	I12
3.3		Titration of 0.1 M solution; pK of +NH ₃ PO ₃ H ₂ given as 2.1		
3.8	20		E3b	R41
2.92			E,Sb	K29
7.86			E,h	M26
2.8	10			H15
4.6	25	Other measurements: C19		R10

Name, Formula and pK value	$T(^{\circ}\text{C})$	Remarks	Methods	Reference	
3. Aminodisulphonic acid, $\text{NH}(\text{HSO}_3)_2$	25	$\text{p}K_3; I = 1.0 (\text{NaCl})$	E3ag	D37a	
4. Aminophosphazenes, see Hexaminotriphosphazene, Octaminotetraphosphazene.					
5. Aminophosphoric acid, see Amidophosphoric acid.					
6. Ammonia, NH_3					
10-081	0	Equal concentrations of NH_3 and KH phenol sulphonate, c varied from 0.011 to 0.104 M, activity coefficients calculated from Debye-Hückel equation, $\text{p}K$ plotted against I	E1ch	B20	
9-903	5				
9-730	10				
9-564	15				
9-401	20				
9-246	25				
9-093	30				
8-947	35				
8-805	40				
8-671	45				
8-540	50				
10-081	0			E1a	B19
9-904	5				
9-731	10				
9-564	15				
9-400	20				
9-245	25				
9-093	30				
8-947	35				
8-805	40				
8-670	45				
8-539	50				
9-555	15	Thermodynamic quantities are derived from these values. I varies from 0.06 to 0.20. Extrapolated to zero concentration of NH_4^+ at each I , then extrapolated against I	E2b	E23	
9-240	25				
8-946	35				
8-670	45				
		Thermodynamic quantities are derived from these values.			

Name, Formula and pK value	$T(^{\circ}C)$	Remarks	Methods	Reference
32.49 27.66 29.8 40	-33.2 24.8 25	Self-ionization of liquid ammonia, from thermodynamic data		C32
		Self-ionization of liquid ammonia, from thermodynamic data		J13
		Approximate pK of NH_2^- , theoretical calculation		S29
		A value of 4.20 at 25° has been claimed from high field conductance measurements to be the true pK_b of $NH_4^+ + OH^- \rightleftharpoons NH_4OH$		B35
		A similar value, 4.28 at 20°, has been estimated from published data		M43
		For pK values in methanol-water mixtures, see E26, P1.		
		Other measurements: B51, F41, H26, H37, K3, K26, L49, M44, N25, O16, P13, S31, W22.		
7. (Aquo) Antimony III ion, Sb^{3+}				
1.4 11.8 0.87 11.0	25 25 25 25	pK for $SbOH^{2+} \rightleftharpoons SbO^+ + H^+$ pK for $SbO^+ + H_2O \rightleftharpoons HSbO_2 + H^+$ pK for $SbO^+ + H_2O \rightleftharpoons HSbO_2 + H^+$ pK for $HSbO_2 + 2H_2O \rightleftharpoons Sb(OH)_4^- + H^+$	SOLY SOLY SOLY SOLY	K5 P29
8. Antimony pentoxide, Sb_2O_5 See also Dodeca-antimonic acid.				
2.55	25	pK for $HSb(OH)_6 \rightleftharpoons Sb(OH)_6^- + H^+$; $I = 0.5(NM_{C_2}Cl)$; Sb concentration $\leq 10^{-3} M$; at higher concentrations polynuclear complexes are also formed.	E3b	L17
Aquo metal ion, See entry under appropriate metal ion				
9. Arsenic acid, H_3AsO_4				
2.089 2.114 2.138 2.163 2.194 2.223 2.265 2.296 2.332	7.054 7.032 7.015 6.999 6.990 6.980 6.974 6.973 6.973	I varied from 0.007 to 0.096 (for K_1) and 0.010 to 0.21 (for K_2); extrapolated to $I = 0$	E1a,quin	A10

Thermodynamic quantities are derived from the results.

2-383	6-973	45		
2-420	6-980	50		
	$pK_1 = 2.014 + 5 \times 10^{-5} (t - 40.0)^2$			
	$pK_2 = 6.971 + 5 \times 10^{-2} (t - 39.4)^2$, t in °C.			
2-49	7-05	10		F19
2-19	6-94	25		
1-95	6-87	35		
2-15	6-80	50		
2-301	7-08	25		
		25	Taking pK_2 of H_3PO_4 as 7.16	E3ag E3ag
			For values of pK_1 in D_2O/H_2O mixtures, see S3	H78 S3
			Other measurements: B58, B86, C23, K48, L50, M8, S54, W4, W5	

10. Arsenious acid, H_3AsO_3 (HA_3O_2)

9-295	15	Molal scale; $\epsilon = 0.008$, $I = 0.1$ (KCl)	E3dg	A29
9-265	20			
9-18	25			
9-09	30			
8-97	35			
8-885	40			
8-81	45			
9-294	25	In KCl solutions, extrapolated to $I = 0$	E	A28
9-22	25	Taking pK of boric acid as 9.19	E3ag	H78
9-26	18	"Practical" constant, titration of 0.017 M H_3AsO_3	E3bg	B86
9-08	25		E3ag	I15
9-4	Room	pK_2 obtained from ultraviolet spectra	O	G33
	32	Other measurements: B58, C11, G5, K36, K48, T15, W5, W25, Z2	GRYOSC	S61

11. Azido-dithiocarbonic acid, $HSCSN_3$

1-67	25	Free acid readily decomposes	C1	S55
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12. (Aquo) Barium ion, Ba^{2+}

0-62	5	pK_1 of $BaOH^+$; $I = 0.1$; f_{\pm} calculated by Davies' equation, for		
0-60	15	extrapolation to $I = 0$; from c.m.f. data of H.S. Harned and		

Name, Formula and pK value	T(°C)	Remarks	Methods	Reference
0.64	25	C. G. Geary, <i>J. Am. Chem. Soc.</i> 59 2032 (1937)		
0.69	25			
0.72	45			
0.64	25	Thermodynamic quantities are derived from the results. pK _b of BeOH ⁺ ; $I = 0.04$ to 0.17; using Davies' equation and activity measurements of H. S. Harned and C. M. Mason, <i>J. Am. Chem. Soc.</i> 54 , 1441 (1932)		D7
0.85	25	$\epsilon = 0.02 - 0.05$ (Ba(OH) ₂), $I = 0.23$ to 0.6 (Ba(OH) ₂ + BaCl ₂); extrapolation to $I = 0$, using Davies' equation	KIN	B30
0.62	25	$I = 0.1$ to 0.45	CAT, KIN	B31
0.72	25	Concentration constant; 0.2 - 1 N BaCl ₂ ; salt effect on indicator	O3	K39
0.00	25	$I = 3$ (NaClO ₄) Other measurements: B32, K64	E2ah	C7
13. (Aquo) Beryllium ion, Be²⁺				
5.7	~7	Beryllium ions readily hydrolyze in solution and form condensed species containing more than one beryllium atom. See, for example, C8 and K1.	E3ag	S30
6.5	25	Successive pK values for hydrolysis of Be ²⁺ ; $I = 0.1$ (NaClO ₄); rapid-reaction measurements; BeOH ⁺ quickly forms trimer Be ₃ (OH) ₃ ³⁺	E3bg	M21
>6.1	25	pK for Be ²⁺ + OH ⁻ ⇌ BeOH ⁺ + H ⁺ ; $I = 1$ (NaClO ₄); Be ₂ OH ³⁺ also formed		H56
10.82		pK for Be ²⁺ + 2OH ⁻ + H ⁺ ; $I = 3$ (NaClO ₄); recalculation of data from refs. C8 and K1 using a computer; also -log K = 10.87 for Be ²⁺ + 2H ₂ O ⇌ Be(OH) ₂ + 2H ⁺ ; constants given for Be ₃ (OH) ₃ ³⁺ and Be ₂ OH ³⁺		A9
10.46		pK _b for Be(OH) ₂ ⇌ BeOH ⁺ + OH ⁻ ; $\epsilon = 0.01$; between pH 6.2 - 5.4; at lower pH values di- and tri-nuclear complexes are formed; constants are given	E3b	
		pK for Be(OH) ₂ + H ₂ O ⇌ Be(OH) ₃ ⁻ + H ⁺ ; tracer concentrations; also -log K = 13.65 for Be ²⁺ + 2H ₂ O ⇌ Be(OH) ₂ + 2H ⁺	DISTRIB	G39

14. (Aquo) Bismuth(III) ion, Bi³⁺

1-58

25

E3bg

O6

pK for $Bi^{3+} \rightleftharpoons Bi(OH)^{2+} + H^+$; $I = 3(NaClO_4)$; $[Bi^{3+}]$ determined by Bi-Hg electrode; main equilibrium is $6Bi^{3+} + H_2O \rightleftharpoons Bi_6(OH)_{12}^{6+} + 12H^+$, with $\log K = 0.33$
 Hydrolysis of Bi^{3+} gives $Bi_6O_6^{6+} + 12H^+$, with $-\log K = 0.53$ at 25° and $I = 1(NaClO_4)$, and at higher pH values $Bi_6O_6(OH)_3^{3+}$, with $\log K = -8.1$
 Hydrolysis of $Bi_6O_6^{6+}$ ($= Bi_6(OH)_{12}^{6+}$) gives $Bi_9(OH)_{20}^{7+}$, $Bi_9(OH)_{21}^{6+}$ and $Bi_9(OH)_{22}^{5+}$; constants are listed

$-\log K = 8.81$ for $3Be^{2+} \rightleftharpoons Be_3(OH)_3^{3+} + 3H^+$; $-\log K = 3.24$ for $2Be^{2+} \rightleftharpoons Be_2OH_3^{3+} + H^+$; $-\log K = 11.0$ for $Be^{2+} \rightleftharpoons Be(OH)_2 + 2H^+$; all for $I = 0.5(NaClO_4)$, c 0.001 to 0.08 M in Be^{2+}
 $-\log K = 10.9$ for $Be^{2+} \rightleftharpoons Be(OH)_2 + 2H^+$, at 25° and $I = 3(NaClO_4)$; constants also given for di- and tri-nuclear species.
 Other measurements: L40, W26.

O7

K1

B37

15. Boric acid, H₃BO₃

9-5078
 9-4374
 9-3785
 9-3255
 9-2780
 9-2340
 9-1947
 9-1605
 9-1282
 9-1013
 9-0766
 9-0537
 9-0310

0
 5
 10
 15
 20
 25
 30
 35
 40
 45
 50
 55
 60

Molal scale; equimolar concentrations (0.003 to 0.03 M) of NaCl, borax and boric acid; extrapolated to $I = 0$ using extended Debye-Hückel equation

M11

E1ch

O18

$pK = 2237.94/T + 0.016883T - 3.305$ (T in °K)
 Thermodynamic quantities are derived from the results. Molal scale. I varied from 0.02 to 3 by adding NaCl; extrapolated to zero boric acid concentration at constant I , then to $I = 0$

9-440
 9-380
 9-327
 9-280

5
 10
 15
 20

E1ch

O18

Name, Formula and p <i>K</i> value	T(°C)	Remarks	Methods	Reference
9·237	25			
9·198	30			
9·164	35			
9·132	40			
9·080	50			
9·380	10	Thermodynamic quantities are derived from the results		
9·327	15	<i>I</i> varied from 0·01 to 0·12; constants corrected using Debye-Hückel equation and extrapolated to <i>I</i> = 0	E1a	O15
9·280	20			
9·236	25			
9·197	30			
9·132	40			
9·080	50			
9·21	20	$pK = 9·023 + 8 \times 10^{-5} (76·7 - t)^2$ (<i>t</i> in °C)		
8·98	25	<i>I</i> = 0·04. The second p <i>K</i> of boric acid is greater than 14	E3ah	F31
9·00	25	<i>I</i> = 0·1 (NaClO ₄)	E3bh	I4
	25	<i>I</i> = 3(NaClO ₄)		
9·00	25	At boric acid concentrations above 0·4 M, higher than trimeric complexes are also formed		
		<i>I</i> = 3(NaClO ₄); boric acid concentrations varied from 0·01 to 0·60 M. Other equilibria were:		
		3H ₃ BO ₃ ⇌ H ₄ B ₃ O ₇ ⁻ + H ⁺ + 2H ₂ O, log <i>K</i> = - 6·84,		
		3H ₃ BO ₃ ⇌ H ₅ B ₃ O ₈ ²⁻ + 2H ⁺ + H ₂ O, log <i>K</i> = - 15·44		
		Polymeric species are important at concentrations above about 0·01 M		
		Other measurements: B71, B88, E6, F9, H10, H25, I5, I6, K42, K48, L15, L41, M24, O17, P39.	E3bh	I9
16. (Aquo) Cadmium ion, Cd²⁺				
10·2	25	p <i>K</i> for hydrolysis of Cd ²⁺ ; <i>I</i> = 3(NaClO ₄ + Cd(ClO ₄) ₂); <i>c</i> = 0·1 to 1·45 (Cd(ClO ₄) ₂); Cd ₂ OH ³⁺ and Cd ₄ (OH) ₄ ⁴⁺ are also formed	E3bg, quin	B44
9·0	25	p <i>K</i> for hydrolysis of Cd ²⁺ ; <i>I</i> = 3(NaClO ₄ + Cd(ClO ₄) ₂); <i>c</i> = 0·01 to 0·9 (Cd(ClO ₄) ₂)	E3bg	M12

KIN K69
SOLY G14
DISTRIB D46
POLAROG L3

pK for hydrolysis of Cd^{2+} ; $c = 0.02$ (CdCl_2)
pK for hydrolysis of Cd^{2+}
pK_b for $\text{HCdO}_2^+ + \text{H}_2\text{O} \rightleftharpoons \text{Cd}(\text{OH})_2 + \text{OH}^-$
 $I = 3(\text{NaClO}_4)$, pK_b for $\text{CdOH}^+ \rightleftharpoons \text{Cd}^{2+} + \text{OH}^-$
pK_b for $\text{Cd}(\text{OH})_2 \rightleftharpoons \text{CdOH}^+ + \text{OH}^-$
pK_b for $\text{Cd}(\text{OH})_3^- \rightleftharpoons \text{Cd}(\text{OH})_2 + \text{OH}^-$
pK_b for $\text{Cd}(\text{OH})_4^{2-} \rightleftharpoons \text{Cd}(\text{OH})_3^- + \text{OH}^-$
on assumption that $\log K_a = \frac{1}{4} \log K_1 K_2 K_3 K_4 + ((5 - 2n)/2)$
 $\log (K_n/K_{n+1})$
 $\log K$ for $\text{Cd}^{2+} + 4\text{OH}^- \rightleftharpoons \text{Cd}(\text{OH})_4^{2-}$ is about 9.7 at 25°
Other measurements: C12, G38, L39

17. Caesium hydroxide, CsOH

For alkalinity function for CsOH solutions, see L12a, M40.

18. (Aquo) Calcium ion, Ca^{2+}

9.49	100	pK _b for CaOH^+ ; $m = 0.002 - 0.02$ $\text{Ca}(\text{OH})_2$ in 0.003 - 0.01 M CaCl_2 or 0.006 - 0.02 M KCl; values of pK _b depend on choice of $\gamma_{\text{Cl}}/\gamma_{\text{OH}}$ used to evaluate molality of hydroxyl ion.	E1b	B17
9.3	10	pK _b for CaOH^+ ; $I = 0.02$ to 0.1 ($\text{Ca}(\text{OH})_2 + \text{CaCl}_2$); $f \pm$ calculated assuming Davies' equation	E1ch	G26
0.7	25	pK _b for CaOH^+ ; $I = 0.007$ to 0.08	SOLY	B29
4.30	25	$I = 0.02$ to 0.08		
3.44	40	$I = 0.04$ to 0.10		
2.58	15	extrapolated to $I = 0$ assuming Davies' equation;		
1.72	25	$\text{Ca}(\text{IO}_3)_2$ in KOH solutions.	SOLY	D9
	0	pK _b for CaOH^+ ; $\text{Ca}(\text{IO}_3)_2$ in $\text{Ca}(\text{OH})_2$ solutions; extrapolated using Davies' equation.	KIN	B32
	0	pK _b for CaOH^+ ; $I = 0.18$ to 0.30; value sensitive to choice of activity coefficient	SOLY	D6
	25	$I = 0.025$ to 0.08		
	25	calculation of data of G. Kilde, <i>Z. Anorg. Allgem. Chem.</i> , 218 113 (1934)		
	25	Recalculation of data of F. M. Lea and G. E. Bessey, <i>J. Chem. Soc.</i> 1937 1612	C2	B30

Name, Formula and pK value	T(°C)	Remarks	Methods	Reference
1.46	25	$I = 0.13$ to 0.24 ($\text{Ca}(\text{OH})_2 + \text{CaCl}_2$); $\epsilon = 0.02 - 0.03$	KIN	B30
1.29	25	$\text{Ca}(\text{OH})_2$; extrapolated using Davies' equation $I = 0.02$ to 0.05 ; f \pm calculated from Guggenheim's equation (<i>Phil. Mag.</i> , 19 , 588 (1935))	CAT, KIN	B31
1.03		Concentration constant; $0.2 - 1N$ CaCl_2 ; from salt effect on indicator	O3	K39
0.64	25	$I = 3(\text{NaClO}_4)$ For the acidity function of $\text{Ca}(\text{OH})_2$ solutions from 0.95° and $I = 0.01$ to 0.20 , see B18 Other measurements; G42	E2ah	C7
19. Carbonic acid, H_2CO_3				
6.577	0	Apparent pK values; double extrapolation procedure to eliminate effect of added NaCl and to obtain values at zero bicarbonate concentration	E1ch	H29
6.517	5			
6.465	10			
6.420	15			
6.382	20			
6.351	25			
6.327	30			
6.309	35			
6.296	40			
6.289	45			
6.287	50			
$pK_1 = 3404.71/T - 14.8435 + 0.032786T$ (T in °K) Thermodynamic quantities are derived from the results. Apparent pK values; $I = 0.004 - 0.2$, extrapolated to $I = 0$				
6.579	0			
6.517	5			
6.464	10			
6.419	15			
6.381	20			
6.352	25			
6.327	30			
6.309	35			
6.298	40			
6.290	45			
			E1ch	H31

6-285	50					
6-514	5	Apparent p <i>K</i> values; $I = 0.003 - 3$; extrapolated to $I = 0$ by fitting to an extended Debye-Hückel equation		E3bh	N7	
6-421	15					
6-349	25					
6-310	35					
6-294	45					
6-583	0	$pK_1 = 6.572 - 0.012173t + 0.00013329t^2$ (t in °C)				
6-429	15	Apparent p <i>K</i> values; 0.001 N in KHCO ₃ , KCl, HCl, and saturated CO ₂ solutions		C1,R1d	S40	
6-366	25					
6-317	38					
6-35	25	Apparent p <i>K</i> value		E2b, quin	A45	
6-35	25	Apparent p <i>K</i> value		E1c, quin	A44	
6-38	25	Apparent p <i>K</i> values, molal scale, I varied from 0.0001 to 0.1,	1035 atmosphere, 2050 atmosphere, 2930 atmosphere, 1 atmosphere, 1030 atmosphere, 2035 atmosphere, 2930 atmosphere, 1 atmosphere, 1015 atmosphere, 2010 atmosphere, 3000 atmosphere, 1 atmosphere, 1020 atmosphere, 2010 atmosphere, 2950 atmosphere, 1 atmosphere, 1050 atmosphere, 2060 atmosphere, 2800 atmosphere,	C1	E10	
5-90						
5-48						
5-15						
6-32	35					
5-85						
5-45						
5-12	45					
6-32						
5-89						
5-50						
5-16						
6-30	55					
5-86						
5-49						
5-17	65					
6-31						
5-88						
5-51						
5-26						
10-625	0	I varied from 0.02 to 0.16; extrapolated to $I = 0$ using extended Debye-Hückel equation		E3ah	H39	
10-557	5					
10-490	10					
10-430	15					
10-377	20					
10-329	25					

Name, Formula and pK value	T(°C)	Remarks	Methods	Reference
10.290	30	$pK_2 = 2902.39/T - 6.4980 + 0.02379T$ (T in °K) Thermodynamic quantities are derived from the results. I varied from 0.005 to 0.1; extrapolated against I	E.g	C45
10.250	35			
10.220	40			
10.195	45			
10.172	50			
10.179	60	$pK_2 = 2909.10/T - 6.119 + 0.02272T$ (T in °K) Double extrapolation, first to values in pure aqueous NaCl solutions, then against I to $I = 0$ $I = 0$ $I = 1(\text{NaCl})$	E3ah	W9
10.153	70			
10.142	80			
10.140	90			
10.641	0			
10.397	18	I varied from 0.01 to 0.2; extrapolated against I^{\ddagger} I varied from 0.01 to 0.2; extrapolated against I^{\ddagger} True pK for $\text{H}_2\text{CO}_3 \rightleftharpoons \text{H}^+ + \text{HCO}_3^-$; high field conductivity measurements	VAP	N6 E9
10.32	25			
10.33	25			
6.35	50			
6.29	100			
6.24	100			
6.33	150			
6.55	200			
6.42	100			
6.77	150			
7.27	200			
7.89	250			
8.70	300			
6.46	100			
6.81	156			
7.14	200			
10.96	218	I varied from 0.01 to 0.2; extrapolated against I^{\ddagger} I varied from 0.01 to 0.2; extrapolated against I^{\ddagger} True pK for $\text{H}_2\text{CO}_3 \rightleftharpoons \text{H}^+ + \text{HCO}_3^-$; high field conductivity measurements	CI	R48
6.34	25			
6.31	38			
3.81	5			
3.75	15			
3.76	25			
3.78	35			

3.80	38				
3.80	45				
3.68	0.5				D81
3.88	25				B34
3.75	23.5				S13
3.82	30.2				
3.89	35.6				
3.80	4				L38
20. Caro's acid, see Peroxymonosulphuric acid					
21. (Aquo) Cerium(III) ion, Ce³⁺					
	25				M38
					B49
22. (Aquo) Cerium(IV) ion, Ce⁴⁺					
	5				R20a
	15				
	25				
	35				
	25				O1a
	1.6				B8
	25				
	25				S42
23. Chloramine, see Monochloramine					

Name, Formula and pK value	T(°C)	Remarks	Methods	Reference
4-1 ~5.6	20	Successive pK values for hydrolysis of Cr ³⁺ ; I = 0.1 (NaClO ₄); rapid-flow measurements	E3ag	S30
4-10 5.55	25	Successive "practical" pK values for hydrolysis of Cr ³⁺ ; I = 0.04 to 0.4	E3ag	E18
3-96	25	Other measurements: B54, C15, D23, L4	O5	
29. (Aquo) Cobalt(II) ion, Co²⁺				
9-96	15	pK for hydrolysis of Co ²⁺ at I = 0.25 and 0.75 (NaClO ₄)	E3bg	B64
9-85	25			
9-62	35			
9-50	45			
8-9	30			
8-7	100	pK for hydrolysis of Co ²⁺ ; I = 0.1 (KCl) pK for hydrolysis of Co ²⁺ Other measurements: A7, D23, G12, F31	E3bg K1N	C12 K68
30. (Aquo) Cobalt(III) ion, Co³⁺				
2-10	12.5	pK for hydrolysis of Co ³⁺ ; I = 1 (NaClO ₄)	O6	S79
1-98	18.5			
1-78	23.6			
1-71	28.2			
The above values are uncertain because high cobaltic concentrations and low acidities favour formation of polynuclear species				
31. (Aquo) Copper(II) ion, Cu²⁺				
8-0	25	The pK for Cu ²⁺ is not known; hydrolysis of Cu ²⁺ gives almost entirely polynuclear complexes of the type, Cu _n (OH) _{2n-2} ²⁺ ; formation constants for Cu ₂ (OH) ₂ ²⁺ from 15-42° are given. pK for Cu ²⁺ ⇌ CuOH ⁺ + H ⁺ ; I = 3 (NaClO ₄); the major species formed is Cu ₂ (OH) ₂ ²⁺ , with -log K = 10.6 pK for Cu ²⁺ ⇌ CuOH ⁺ + H ⁺ ; the major species formed is Cu ₂ (OH) ₂ ²⁺ , with -log K = 10.89 Hydrolysis of Cu ²⁺ gives Cu ₂ (OH) ₂ ²⁺ , with -log K at 25° ranging from 10.5 to 10.9	E3bg	B33
7-97	18			
				H4

Name, Formula and p <i>K</i> value	<i>T</i> (°C)	Remarks	Methods	Reference
14-784	30			
14-468	40			
14-182	50			
15-439	10	Molar scale		
15-049	20			
14-869	25			
14-699	30			
14-385	40			
14-103	50	Mole fraction scale		
17-224	10			
16-834	20			
16-653	25			
16-482	30			
16-166	40			
15-880	50			
15-08	15	$pK_m = 4913.14/T - 7.5117 + 0.0200854T$ (<i>T</i> in °K) Thermodynamic quantities are derived from the results. Molar scale; <i>I</i> = 0.04 to 0.10, extrapolated to <i>I</i> = 0	E2bh	W35
14-71	25			
14-37	35			
14-807	20	Molal scale, taking $pK_w = 14.073$ for H ₂ O	E3ah	S31
14-812	25	Molal scale	ANALYT	K24
14-81	25	Using 0.01 M Ba(OH) ₂ in H ₂ O/D ₂ O mixtures; molal scale.	E3ag	S3
14-86	25	Molar scale, <i>I</i> = 0	E3bh	G31
14-856	25	Molar scale, <i>I</i> = 0	E3bh	G31a
14-80	25	<i>I</i> = 0	E1cd	A2
		For p <i>K</i> values of H ₂ O/D ₂ O mixtures, see G31, G31a		
37. Deuterio-ammonia, ND₃				
9-757	20	<i>I</i> = 0; in D ₂ O; taking p <i>K</i> of NH ₃ in H ₂ O as 9.265	E3ah	S31
4-9		p <i>K</i> _b ; rough estimate		L24
38. Deuterio-arsenic acid, D₃AsO₄				
2-596	25	p <i>K</i> ₁ in D ₂ O; <i>I</i> = 0; from measurements in D ₂ O/H ₂ O mixtures	E3ag	S3

39. Deuteriocarbonic acid, D₂CO₃								
6-77	10-96	25	Apparent p <i>K</i> in D ₂ O; taking p <i>K</i> ₁ for CO ₂ in H ₂ O = 6.35	E3a,quin	C40			
	10-93	25	In D ₂ O; meter standardized in H ₂ O; p <i>K</i> ₂ for H ₂ CO ₃ in H ₂ O taken as 10.33	E3ag	G29			
		25	In D ₂ O; taking p <i>K</i> ₂ for H ₂ CO ₃ in H ₂ O as 10.33	E3ag	C41			
40. Deuteriodisulphuric acid, D₂S₂O₇			For p <i>K</i> _a in concentrated H ₂ SO ₄ , see F21					
41. Deuteriohydrazine, N₂D₄								
	9-08	17	“Practical” constant; in D ₂ O; <i>I</i> = (KCl)	E3dg	F12			
	8-69	30						
	9-11	18	“Practical” constant, in D ₂ O; <i>I</i> = 1 (KCl)	E3dg	B102			
42. Deuteriohydrazoic acid, DN₃								
	5-01	20	In D ₂ O; <i>I</i> = 0	E3bg	B108			
43. Deuterio-iodic acid, DIO₃								
	1-15	25	In D ₂ O; taking p <i>K</i> of HIO ₃ in H ₂ O = 0.85	C1,R1d	M2			
44. Deuteriophosphoric acid, D₃PO₄								
	2-350	25	In D ₂ O; <i>I</i> = 0; extrapolated from measurements in D ₂ O/H ₂ O mixtures	E3ag	S3			
	2-362	25	In D ₂ O; taking p <i>K</i> ₁ for H ₃ PO ₄ in H ₂ O as 2.128	C	M2			
	2-31	25	In D ₂ O; meter standardised in H ₂ O; taking p <i>K</i> ₁ for H ₃ PO ₄ in H ₂ O as 2.11	E3ag	G29			
	7-8846	5	In D ₂ O; using KD ₂ PO ₄ /Na ₂ DPO ₄ mixtures from 0.005 to 0.025 m, and NaCl 0.005 m; extrapolated to <i>I</i> = 0	E1a	G8			
	7-8499	10						
	7-8233	15						
	7-7986	20						
	7-7796	25						
	7-7667	30						
	7-7547	35						
	7-7484	40						
	7-7433	45						
	7-7435	50						

$pK_2 = 2202.11/T - 5.9823 + 0.021388T$ (T in °K)
Thermodynamic quantities are derived from the results.

~1	~2.2	3.24	6.80	9.50	37	Values of pK_1 and pK_2 could be seriously in error because of experimental limitations $I = 0.1$; same remarks as above $I = 0.1$	E3bg	I11
~1	~2.4	3.59	7.02	9.28	50			
49. Diperosmic acid, See Osmic(VIII) acid								
50. Diperruthenic acid, (hydrated RuO_4)								
11.17					20?	Distribution between CCl_4 and water		DISTRIB M15
14.24						pK_b for $H_2RuO_5 \rightleftharpoons HRuO_4^+ + OH_2^-$		S48
11.9								
51. Disulphuric acid, $H_2S_2O_7$								
-12	-8					Theoretical predictions (Ricci's method) of pK_1 and pK_2		G24
-13	-8					Theoretical predictions (Pauling's method) of pK_1 and pK_2		B13
1.85					10	pK in concentrated H_2SO_4 ; molal scale	FP	B74
2.52					20	pK in concentrated H_2SO_4 ; molal scale		
52. Dithionic acid, $H_2S_2O_6$								
-3.4	-0.2							K52
53. Dithionous acid, see Hyposulphurous acid								
54. Dodeca-antimonic acid, $H_{12}(Sb(OH)_6)_{12}$								
<1.55	<1.55	<1.55			25	pK_1, pK_2, pK_3		L17
1.55	2.95	4.35				pK_4, pK_5, pK_6	E3b	
5.75	7.15					pK_7, pK_8 , all at $I = 0.5$ (NMe ₄ Cl); this acid exists in equilibrium with mononuclear antimony species at Sb(V) concentrations above 10^{-3} M		
55. Dodecahydrododecaboric acid, $H_2B_{12}H_{12}$								
For acidity function, see M50								
56. Dodecatungstic acid, $H_{10}W_{12}O_{41}$								
~3.6	5.27	6.28			20	$pK_8, pK_9, pK_{10}; I = 0.1$ (NaCl); rapid-reaction technique	E3ag	S83

Name, Formula and pK value	T ($^{\circ}\text{C}$)	Remarks	Methods	Reference
57. (Aque) Dysprosium(III) ion, Dy^{3+} 8-10	25	pK_a for hydrolysis of Dy^{3+} ; titration of 0.004-0.009 M $\text{Dy}(\text{ClO}_4)_3$ with 0.02 M $\text{Ba}(\text{OH})_2$; $I = 0.3(\text{NaClO}_4)$	E3b	F33a
58. (Aque) Erbium(III) ion, Er^{3+} 7-99	25	pK_a for hydrolysis of Er^{3+} ; titration of 0.004-0.009 M $\text{Er}(\text{ClO}_4)_3$ with 0.02 M $\text{Ba}(\text{OH})_2$; $I = 0.3(\text{NaClO}_4)$	E3b	F33a
59. (Aque) Europium(III) ion, Eu^{3+} 8-31	25	pK_a for hydrolysis of Eu^{3+} ; titration of 0.004-0.009 M $\text{Ba}(\text{OH})_2$; $I = 0.3(\text{NaClO}_4)$	E3b	F33a
~8-8	25	pK for hydrolysis of Eu^{3+} ; hydrolysis of "pure" salt; $c = 0.001$ - 0.01 M $\text{Eu}_2(\text{SO}_4)_3$	E3ag	M38
60. Ferric ion, see Iron(III) ion				
61. Ferricyanic acid, $\text{H}_3\text{Fe}(\text{CN})_6$ <1	25	pK_3 ; $c \approx 10^{-3}$	E3bg	J14
62. Ferrocyanic acid, $\text{H}_4\text{Fe}(\text{CN})_6$ 2.2 4.17 2.57 3 4.3 4.17 2.3 4.28 4.25	25 25 17 25 25 25	pK_3, pK_4 ; $c \approx 10^{-3}$; $I = 0.01$ to 0.5; extrapolated to $I = 0$ pK_3, pK_4 ; $I = 0$ pK_3, pK_4 ; $I = 0$ pK_4 ; $I = 0.001$ to 0.25; extrapolated against I pK_3, pK_4 variation of redox potential with pH; extrapolated to $I = 0$ pK_4 ; variation of redox potential with pH Other measurements: K37	E3bg E3bg E3bg E3bg REDOX REDOX	J14 H74 N14 L7 H20a K46
63. Ferrous ion, see Iron(II) ion				
64. Fluorophosphoric acid, $\text{H}_2\text{PO}_3\text{F}$ ~0.5 4.80 5.12	40 25	"Practical" constants	E3b E	D32 R47

65. Fluorosulphuric acid, HF₂SO₃

For p*K*_a in sulphuric acid, see B11

66. (Aquo) Gadolinium(III) ion, Gd³⁺

25	p <i>K</i> for hydrolysis of Gd ³⁺ ; hydrolysis of "pure" salt; <i>c</i> = 0.001–0.01 M Gd ₂ (SO ₄) ₃	E3ag	M38
25	p <i>K</i> _a for hydrolysis of Gd ³⁺ ; titration of 0.004–0.009 M Gd(ClO ₄) ₃ with 0.02 M Ba(OH) ₂ ; <i>I</i> = 0.3(NaClO ₄)	E3b	F33a

67. (Aquo) Gallium(III) ion, Ga³⁺

25	p <i>K</i> for hydrolysis of Ga ³⁺ ; <i>c</i> = 0.004–0.25 M in Ga ³⁺	E3bg	M39
25	p <i>K</i> for hydrolysis of Ga ³⁺ ; <i>I</i> = 0.5(NaClO ₄) ₄	O6	W20
18	Successive p <i>K</i> values for hydrolysis of Ga ³⁺	E3ah	F32
18	Successive p <i>K</i> values for hydrolysis of Ga(OH) ₄ [−] to Ga(OH) ₅ ^{2−} and Ga(OH) ₆ ^{3−}		I22
20	p <i>K</i> for Ga(OH) ₃ + H ₂ O ⇌ Ga(OH) ₄ [−] + H ⁺ ; <i>c</i> = 0.005–0.025 M in Ga(OH) ₄ [−] ; from pH of hydrolysed alkali-metal salts		
20	Successive p <i>K</i> values for hydrolysis of Ga ³⁺ to GaOH ²⁺ , Ga(OH) ₂ ⁺ and Ga(OH) ₃ ; <i>I</i> = 1(NaCl)	DISTRIB	A17

68. Germanic acid, H₄GeO₄ (H₂GeO₃)

9.02	p <i>K</i> ₁ , p <i>K</i> ₂ ; <i>I</i> = 0.5(NaClO ₄)	E3bg	H1
8.98	<i>I</i> = 1(NaClO ₄)		
9.03	Below 0.004 M, germanic acid is mainly monomeric; at higher concentrations an octagermanic anion is also formed		
8.98	p <i>K</i> ₁ , p <i>K</i> ₂ ; <i>I</i> = 0.5(NaCl); log <i>K</i> = 29.14 for 8Ge(OH) ₄ + 3OH [−] ⇌ (Ge(OH) ₄) ₈ (OH) ₃ ^{3−}	E3bg,h	I7
8.92	<i>I</i> = 0	E3cg	A25
8.73			
8.62			
9.1	<i>I</i> = 2(KCl)	E3b	C9
9.08	In 0.5 M Na ₂ SO ₄	E	L42
12.43	<i>I</i> = 3(NaCl)	E3bh	I10
12.31	In saturated Na ₂ SO ₄ solution	CRYOSC	K63
	Other measurements: G15, G52, P43, R35, S26		

69. Gold(III) hydroxide, Au(OH)₃

<11.7	Successive <i>K</i> _p values for ionization to H ₂ AuO ₃ ^{2−} , HauO ₃ ^{3−} and AuO ₃ ^{3−} ,	SOLY	J11
13.36	>15.3		

Name, Formula and pK value	T(°C)	Remarks	Methods	Reference
70. (Aquo) Hafnium(IV) ion, Hf^{4+} -0.12 0.23 0.42 0.52	25	Successive pK values for hydrolysis of Hf^{4+} ; $I = 1(\text{HClO}_4)$; using low (radio-isotope) Hf^{4+} concentrations; at concentrations above 10^{-3} M polymers (mainly trimers and tetramers) are also formed	DISTRIB	P24
71. Heptamolybdic acid, $\text{H}_6\text{Mo}_7\text{O}_{24}$ ~3.7 4.33	25	$\text{p}K_5, \text{p}K_6; I = 3(\text{NaClO}_4)$; concentration constants; also log $K = 57.7$ for $7\text{MoO}_4^{2-} + 8\text{H}^+ \rightleftharpoons \text{Mo}_7\text{O}_{24}^{6-} + 4\text{H}_2\text{O}$	E1cg,h	S8
72. Hexadecapolyphosphoric acid, $\text{H}_{13}\text{P}_{16}\text{O}_{49}$ ~2 2.92 6.48 6.08 ~2 2.64 6.48 8.02 ~2 2.52 6.50 8.18	25 37 50	Concentration constants: $I = 1(\text{NMe}_4\text{Br})$; f_{\pm} assumed same as for HBr ; first two of these pK values may be seriously in error because of experimental difficulties	E3bg	I11
73. Hexametaphosphoric acid, $\text{H}_6\text{P}_6\text{O}_{18}$ 2 5.60 7.82		No details		K3a
74. Hexaminotriphosphazene, $\text{N}_3\text{P}_3(\text{NH}_2)_6$ <3.2 7.65 7.70	25 25	$\text{p}K_1, \text{p}K_2; I = 0$ $\text{p}K_3; I = 0$	E E	F10 F11
75. Hexapolyphosphoric acid, $\text{H}_6\text{P}_6\text{O}_{19}$ ~2.1 2.19 5.98 8.13 ~1.3 2.22 5.83 8.02 ~1.3 2.22 5.81 8.00	25 37 50	Concentration constants; $I = 1(\text{NMe}_4\text{Br})$; f_{\pm} assumed same as for HBr ; lowest two pK values uncertain because of experimental difficulty Concentration constants, as above	E3bg E3bg	I12 I11
76. (Aquo) Holmium(III) ion, Ho^{3+} 8.04	25	$\text{p}K_9$ for hydrolysis of Ho^{3+} ; titration of $0.004\text{-}0.009$ M $\text{Ho}(\text{ClO}_4)_3$ with 0.02 M $\text{Ba}(\text{OH})_2$; $I = 0.3(\text{NaClO}_4)$	E3b	F33a

77. Hydrazine, N₂H₄

-0.88	8.11	O2.E	S28
0.27	7.94	E3bg	Y15
	8.24	E3dg	W2
	7.99		
	7.82		
	8.60	E3dg	B102
	8.40		
	8.20		
	8.15	E3bg	J5
	8.07	E3dg	H59

$I = 0$
 $I = 0$
 I from 0.01 to 0.15; extrapolated to $I = 0$

“Practical” constant; $I = 1$ (KCl)

“Practical” constant; $I = 0.3$ (NaClO₄)

“Practical” constant; 0.02–0.05 M hydrazine

Other measurements; B78, G21, H78

For H₂ acidity function of hydrazine see D26, F8, S12, S73

78. Hydrazinosulphuric acid, ⁺NH₃NHSO₃⁻
 3.85

E3bg A39

79. Hydrazoic acid, HN₃

4.72		E1cg	Y16
4.65, 4.68		E3bg	
4.62			
4.59		O5	B110
4.68		E3ag	H78
4.692		E3bg	B108
4.686		E3cg	S36
4.684			
4.682			
4.680			
4.680			
4.680			
4.70			
4.64		E3bg	B67a
4.58			
4.55		E3b	Q4
-6.21		DISTRIB	B6
-10.1			

0.0075 M solution

$I = 0$

$I = 0.01$ to 0.03

$I = 0.03$ to 1.0; p*K* corrected using Debye-Hückel equation and extrapolated against I

$I = 0$

$I = 0.02$

$I = 0.01$ to 0.04, extrapolated to $I = 0$

$I = 0.1$ to 1.3 (KCl); extrapolated against $I^{\frac{1}{2}}$

p*K* of monocation, H₂N₃⁺, using H₀ function for H₂SO₄

p*K* of dication, H₃N₃²⁺, using H₀ function for H₂SO₄ and data

by A. Hantzsch (*Ber.*, **63B** 1782 (1930))

Other measurements: B88, H21, H64, O9, W15

Name, Formula and pK value	T(°C)	Remarks	Methods	Reference
80. Iodiodic acid, HI				
~-9	25	Using Raoult's law	VAP	B25
~-9.5	25	Calculation from thermodynamic data		MI
81. Hydrobromic acid, HBr				
~-8	25	Using Raoult's law	VAP	B25
~-9	25	Calculation from thermodynamic data For Hammett acidity function of HBr, see P11, V8		MI
82. Hydrochloric acid, HCl				
-7.3	0		VAP	R21
-6.8	10			
-6.4	20			
-6.1	25			
-5.9	30			
-5.4	40			
-5.1	50			
~-7.4	0			
~-7	25	Assuming free HCl is like free HCN	VAP	W33
~-7	25	Calculation from thermodynamic data	B25	B25
~-7	25	Calculation from thermodynamic data		MI
~-7		Assuming solubilities of free HCl and RCl (where R = CH ₃ , C ₂ H ₅ , etc.) in water fall in regular sequence		E2
~-7		Assuming $K_{HF}/K_{HCl} \approx K_{H_2O}/K_{H_2S} \approx 10^{-9}$		S22
~-6		I = 0; in superheated steam, density 0.525 g/ml	C1	P12
3.26	360			
3.42	373			
3.47	378			
4.11	370			
4.14	373			
4.24	378			
4.32	383			
4.61	373			
4.74	378			

pK of H_2Cl^+ ; theoretical prediction
 pK of HCl; theoretical prediction
 For pK values of HCl in superheated steam between 400 and 700°, with densities from 0.3 to 0.8 g/cm³, see F26
 For pK value in absolute ethanol, see S22
 For Hammett acidity function of HCl see B27, B75, D21 (in the presence of LiCl and NaCl), G17 and G18 (temperature range), P10 (in the presence of added salts), P11, V8.
 For H_A acidity function, see Y5
 For H_L acidity function, see P27
 For H_0 , H_0'' , H_R and H_R' acidity functions of HCl, see A36

83. Hydrocyanic acid, HCN

9-216	25	Taking pK of <i>m</i> -bromophenol as 9.004; 0.01-0.05 M borax buffers; extrapolation to $I = 0$, using extended Debye-Hückel equation; freshly prepared cyanide solutions	O2	A22
9-63	10	$I = 0.002$ to 0.024; extrapolation to $I = 0$ using extended Debye-Hückel equation, freshly prepared cyanide solutions	E3bg	I23
9-49	15			
9-36	20			
9-21	25			
9-11	30			
8-99	35			
8-88	40			
8-78	45			
9-36	20	Thermodynamic quantities are derived from the results	E3bg	B67a
9-19	26	$I = 0.01$ to 0.04, extrapolated to $I = 0$		
9-05	33			
9-30	28	Other measurements: A20, B84, B87, B88, H28, K36:	E3b	G10

84. Hydrofluoric acid, HF (H_2F_2)

3-18	25	for $I = 0$; 0.01-0.1 M in HF, 0.002-0.01 M in KF; over these temperatures, K_1 for $F^- + HF \rightleftharpoons HF_2^-$ is 3.4, 4.0, 4.7, 4.8, 4.9, 5.7, 5.8, 8, respectively	C1, R1b	E11
3-40	50			
3-64	75			
3-85	100			
4-09	125			
4-34	150			

Name, Formula and pK value	T(°C)	Remarks	Methods	Reference
4.58	175	Data fit $pK = 2.75 + 295/T - 1.91 \log T + 0.014 T$ (T in °K); Thermodynamic quantities are derived from the results for $I = 0$; 0.001 M in NaF; dilute HF solutions for $I = 0$; using Pb-Hg/PbF ₂ instead of Ag/AgCl. 0.001-1.0 M in HF, K_1 for $F^- + HF \rightleftharpoons HF_2^-$ was 3.94, 3.86, 4.32 at 15, 25, 35°		
4.89	200			
3.21	25	Recalculation of data by E. Deussen (<i>Z. Anorg. Allgem. Chem.</i> , 44 312 (1905)); K_1 for $F^- + HF \rightleftharpoons HF_2^-$ was 2.43, 2.70 at 0, 25°		
3.10	15	Taking $a_{HF_2^-}/(a_{HF} a_{F^-}) = 5.4$		
3.17	25	pK of H ₂ F ⁺ ; theoretical prediction		
3.25	35	Other measurements: A15, A41, B67, B98, B99, C24, C30, C31, D10, D41, F3, F29, P28, R33, R34, R50 (at 100, 156, 218°), S78		
2.96	0	For Hammett acidity function of HF, see B26, H80, N2 (in ethanol-water mixtures), P11		
3.16	25			
3.16	25			
~9				
85. Hydrogen peroxide, H₂O₂				
11.86	15	$I = 0.05$ to $4.8(NaClO_4)$; extrapolated against $I^{1/2}$; $c = 0.55$ M H ₂ O ₂	E3ag	E22
11.75	20			
11.65	25			
11.55	30			
11.45	35			
11.81	20			
11.92	10			
11.62	25			
11.34	35			
11.21	50			
12.11	0			
12.23	0			
12.19	0			
11.85	19			
			E3bg CALOR	K4 S2
			KIN DISTRIB CI	J22
			O1	J17

11-58		30	$I = 0.1$ (phosphate buffers) corrected to $I = 0$ by Debye-Hückel equation In strong hydrogen peroxide solutions (above several per cent H_2O_2 in water), superacidity is observed, giving lower values of pK which pass through a flat minimum (8.7) near 50% Other measurements: E24	O1	M51 M34, K32
86. Hydrogen polysulphide 3-8 6-3		20	For H_2S_4 ; $I = 0.1$ ($NaClO_4$); rapid-flow measurements; 3.4 and 5.6 for H_2S_5	E3ag	S30
87. Hydrogen selenide, H_2Se 15.0		22	Estimated uncertainty ± 0.6 pH units; the direct titration of H_2Se with KOH gives low pK_2 values because of aerial oxidation	SOLY	W27
	14	25	Value needed to fit experimental E_4/pH plot	POLAROG	L35
	3-89 3-73 3-77	25 25-9	$I \sim 0.03$; titration of H_2Se in the dark $c = 0.008-0.1 M H_2Se$	E3bg C1	H9 H61 B104
88. Hydrogen sesquioxide, H_2O_3 9-10			Theoretical prediction		C46
89. Hydrogen sulphide, H_2S 7-33 7-24 7-13 7-05 6-97 6-90 6-79 6-69 6-62 7-57 7-06 6-82 7-02		5 10 15 20 25 30 40 50 60 0 25 50 25		C1	W30
		20	$c = 0.001$ to $0.017 M$ in H_2S	C1,R1b	L44
		25	I varied from 0.01 to 0.17, phosphate buffers	O1	E14
	14.0 14.7 13-85	20 0 30	Extrapolation of measured pK versus alkali concentration	O7 FP,C O5	J3a M54

Name, Formula and pK value	T(°C)	Remarks	Methods	Reference
7-26	10	Values of pK_2 (obtained from titrations) given in this reference are probably too low	E3bg	T20
7-07	25			
6-99	35			
6-91	50			
6-96	18	$\epsilon = 0.001-0.04$ M in H ₂ S	E3ag, VAP	G35
6-87	20			
6-79	25			
6-66	35			
6-54	45			
6-91	25			
6-81	25	1 atmosphere pressure	E3bg	Y17
6-68	25	500 atmosphere pressure	Cl	E13
6-56	25	1000 atmosphere pressure		
6-45	25	1500 atmosphere pressure		
6-37	25	2000 atmosphere pressure		
	0	Calculated from thermodynamic data and potential measurements		M13
	14-75			
	13-90			
6-88	25	$I = 1(KCl)$; Hg electrode versus calomel		W18
6-99	20	Calculated from published thermodynamic data	E	P32
6-89	25			
12-24	40			
6-81	80			
6-54	10-68			
6-52	90			
6-59	9-27			
	120			
	8-55			
90. Hydrogen telluride, H₂Te				
2-64	18	$\epsilon = 0.003-0.09$ M H ₂ Te	Cl	H61
2				BI04
11	25	Value of pK_2 needed to fit E ₁ /pH plot		POLAROG L35
12-16	25	Value of pK_3 needed to fit E ₁ /pH plot		POLAROG P7

91. Hydroperoxy radical, HO₂					
4-4	23	pK for HO ₂ ⇌ H ⁺ + O ₂ ⁻ ; from pH-dependence of reaction with tetramethane; species generated by electron irradiation	KIN	C46	
4-45	23	pH-dependence of rate of reaction with tetranitromethane	O	R1a	
4-5		pulsed radiolysis experiments	O	C46a	
~2		estimate		UI	
~6	20	estimate		W11	
92. Hydrosulphuric acid, see Hydrogen sulphide					
93. Hydroxylamine, NH₂OH					
6-186	15	$I = 0.25, 1, 2.25(\text{NaClO}_4)$; extrapolated to $I = 0$	E3bg	L47	
6-063	20	using Debye Hückel equation			
5-948	25				
5-730	35	$pK = 2775.7/T - 5.8899 + 0.0084782T$ (T in °K)			
		Thermodynamic quantities are calculated from the results for $I = 0$; taking pK of 3,4-dinitrophenol as 5.46, 5.42, and 5.38	O2	R22	
6-04	20	For $I = 0$	E3ag	H8	
5-96	25	$I = 0.0023$ to 0.023; extrapolated against I^\ddagger	DISTRIB		
5-84	30	"Practical" constant; $I = 1(\text{KCl})$	E3bg	B51	
5-98	25	"Practical" constant; $I = 1(\text{KCl})$; in D ₂ O	E3ag	B103	
5-93	25	Other measurements: F28, I15, M16, M52, M53, R28, S83, W21	E3ag		
5-97	30				
6-04	30				
6-49	30				
94. Hydroxylamine-<i>N,N</i>-disulphonic acid, HON(HSO₃)₂					
11-85	25	$pK_3; I = 1.6(\text{K}_2\text{SO}_4?)$	E3ag	A8a	
95. Hydroxylamine-<i>N</i>-sulphonic acid, HO.NH.OSO₂H					
~12-5	Room	$pK_3; I = 1.5(\text{K}_2\text{SO}_4)$	E3ag	A8a	
12-38	64-2	$I = 1.6(\text{Na}_2\text{CO}_3 \text{ or } \text{Na}_2\text{SO}_4)$			
12-20	73-8				
12-10	83-5				
96. Hydroxylamine O-sulphonate, ⁺NH₃OSO₃⁻					
1-48	45	$I = 1(\text{NaClO}_4)$	E3ag	C3	

Name, Formula and pK value	T(°C)	Remarks	Methods	Reference
97. Hydroxyl radical, ·OH	23	Pulse radiolytic method		R1
11·9	~23	Pulse radiolysis; pK obtained from pH-dependence of rate of formation of radical ion, ·CO ₃ ⁻		W9a
11·8				
98. Hypobromous acid, HOBr	10		O	F 18
8·66	25			
8·49	25			
8·23	50		E3bg	K11
8·80	15·65			
8·60	25·28			
8·47	35·55			
8·36	45·55			
8·68	22	$I = 0·02$ to $0·1$	E3bg	S44
8·69	20	$c = 0·01-0·02$ M BrO ⁻	E3bg	S45
		Other measurements: C14, F2, K16, L29, S53, S64		
99. Hypochlorous acid, HOCl	0	Measured relative to pK ₂ of HgPO ₄ ;	O3	M46a
7·825	5	$I = 0·05$ to $0·2$; extrapolated to $I = 0$		
7·754	10			
7·690	15			
7·633	20			
7·582	25			
7·537	30			
7·497	35			
7·463	10	For $I = 0$; $c = 0·01$ M HOCl	E3bg	F16
7·49	25			
7·30	35			
7·18	50			
7·05	10	For $I = 0$; $c = 0·003$ M HOCl	O5	O13
7·50	25			
7·31	35			
7·19	50			
7·06				

7.82	0	Extrapolated to zero time, and to $I = 0$ using Debye-Hückel equation	E3ag	C4
7-72	10			
7-65	15			
7-53	25			
7-49	35			
7-46	45			
7-53	25		E3bg	H7
7-50	20		E3bg	S45
7-49	20		E3bg	M59
7-66	0-6	For $I = 0$; using Debye-Hückel equation	E3bg	A38
7-55	20			
7-42	27	"Practical" constant; $c = 0.25$ M HOCl Other measurements: B85, D4, G2, G27, H65, H73, I2, K17, S4, S46, S52, S53, S58, Y11	E3bg, R2a	L36
100. Hypoiodous acid, HOI				
10-64	25	Also $pK = 14.48$ for $I_2OH^- \rightleftharpoons I_2O^{2-} + H^+$	E,h	C21
9-7	22		E3bg	J19
~11	25		KIN	F39
12.4	20		E	S51
9-49	25	pK_b for $HOI \rightleftharpoons I^- + OH^-$; iodine electrode	O5	M57
1-35	25	pK of H_2OI^+	E	A18
1-54	25	pK of H_2OI^+ ; cells of type $Pt, I_2, Ag^+, H^+/Sat.KNO_3/I^-, I_2, H^+, Pt$	E	B28
101. Hyponitrous acid, $H_2N_2O_2$				
7-51	0	$I \leq 0.06$; from rates of decomposition	KIN	P34
7-22	20			
7-09	30			
	25			
11-35	50			
11-09	55			
10-97	18			
11-1	15			
	25			
7-32	25	For $I = 0$	E3ag	H79
7-21	35	Using borate buffers in determining K_2	KIN	
7-17	45	Using NaOH solutions		
6-92	45			

Name, Formula and p <i>K</i> value	<i>T</i> (°C)	Remarks	Methods	Reference
7.05	25	For <i>I</i> = 0	E3bg	L10
11.4	25		O3	
6.75	25	<i>I</i> = 1 Other measurements: A3, P33	KIN	B106
102. Hypophosphoric acid, H₄P₂O₆				
<2	20	"Practical" constants; <i>I</i> = 0.1 (KCl) Concentration constants; titration of 0.01 M Na ₄ P ₂ O ₆ in 0.049 M HCl with 0.1 M NaOH	E3bh E3bh	S35 T16
103. Hypophosphorous acid, H₃PO₂				
1.23	25	For <i>I</i> = 0	C1, R1c	P8
1.07	18	"Practical" constant; titration of 0.11 N HgPO ₂ with 0.11 N NaOH	E3bg	M48
1.02	16	<i>I</i> = 0.16	E3a	G44a
1.12	30	<i>I</i> = 0.16		
1.2	45	<i>I</i> = 0.57		
1.03	16	<i>I</i> = 1.13 (KCl); concentration constant Other measurements: B89, G43, K41, M33, N26		
104. Hyposulphurous acid, H₂S₂O₄				
0.35	25		C1	J3
105. Imidodiphosphoric acid, H₂O₃P.NH. PO₃H₂				
~1.5	2.66	7.32	10.22	25
~2	2.85	7.08	9.72	25
~2	2.81	7.05	9.77	25
~1.5	3.05	7.62	10.36	25
~1.8	2.60	7.16	9.79	37
~1.8	2.68	6.99	9.52	37
~1.8	2.81	6.90	9.41	50
~1.8	2.83	6.88	9.32	50
Concentration constants; <i>I</i> = 0.1 (NM ₄ Br); <i>f</i> ± assumed same as for HBr; p <i>K</i> ₁ may be seriously in error because of experimental difficulties				
		<i>I</i> = 0.2		
		<i>I</i> = 0.3		
		<i>I</i> = 1.0		
		<i>I</i> = 0.1; as above		
		<i>I</i> = 0.3		
		<i>I</i> = 0.1		
		<i>I</i> = 0.3		
			E3bg	I12
				I11

106. (Aquo) Indium(III) ion, In^{3+}					
4.43	3.9		25	Successive pK values for hydrolysis of In^{3+} to $\text{In}(\text{OH})_2^+$ and $\text{In}(\text{OH})_3$; $I = 3(\text{NaClO}_4)$; using In-Hg electrode; above 0.001 M, indium forms $\text{In}[(\text{OH})_2\text{In}]_n^{(3+n)+}$	B41
4.4	4.4		25	Successive pK values for hydrolysis of In^{3+} ; $I = 3(\text{NaClO}_4)$; tracer amounts of In^{3+}	DISTRIB R31
11.89	11.55	11.32	20 ± 2	Successive pK _b values for hydrolysis of In^{3+} to $\text{In}(\text{OH})_2^+$ and $\text{In}(\text{OH})_3$; $I = 1$	DISTRIB H19
6.95			25	pK for hydrolysis of In^{3+} to form a mixed hydroxy-chloro complex; $I = 3(\text{NaCl})$; $\epsilon = 0.001-0.04$ M In^{3+} ; a binuclear $\text{In}_2(\text{OH})_2$ chloro complex is also formed Other measurements: H43, H51, M36, M37	E3ag B48
107. Iodic acid, HIO_3 (H_5IO_6)					
0.804			25	Obtained by three independent methods, taking ion-size parameter of 5Å; value depends on ion-size assumed; cmf method due to A. K. Covington and J. E. Prue, <i>J. Chem. Soc.</i> 1955 , 3701	Cl, E, and KIN P26a
0.785			25	Solubility of AgIO_3 in HNO_3 and KNO_3 , extrapolated against $I^{\frac{1}{2}}$; $I = 0.008 - 0.5$	SOLY L25
0.815			30		
0.84			35		
0.788			25	Solubility of $\text{Ba}(\text{IO}_3)_2$ in 1:1 electrolyte solutions, extrapolated against $I^{\frac{1}{2}}$; $I = 0.0025$ to 1	SOLY N3
0.773			25	$I = 0.0026$ to 0.01; extrapolated against $I^{\frac{1}{2}}$	O4 H14
0.773			25	Calculated from data of C. A. Kraus and H. C. Parker, <i>J. Am. Chem. Soc.</i> , 44 , 2429 (1922)	Cl, R1c F38
0.807			25	Calculated from data of C. A. Kraus and H. C. Parker, <i>J. Am. Chem. Soc.</i> 44 , 2429 (1922), taking an ion size of 3Å	Cl, R1c L17a
0.58			0		FP A4
0.72			18	For $I = 0$	Cl
0.74			30	Other measurements: D44a, K18, K60, M2, O12, R36. For H_0 acidity function of aqueous HIO_3 , see D18	NMR H67
108. (Aquo) Iron(II) ion, Fe^{2+}					
6.93			20	pK for hydrolysis of Fe^{2+} ; $I = 0.5-2(\text{NaClO}_4)$	E3bg B65
6.74			25		
6.49			35		

Name, Formula and p <i>K</i> value	T(°C)	Remarks	Methods	Reference
6-34	40			
7-15	20	p <i>K</i> for hydrolysis of Fe ²⁺ ; <i>I</i> = 1(NaClO ₄)	E3bg	B63
6-8	25	p <i>K</i> for hydrolysis of Fe ²⁺	SOLY	L19
8-3	25	p <i>K</i> for hydrolysis of Fe ³⁺ ; <i>c</i> = 0.02-0.08 M FeCl ₃ ; hydrolysis of "pure" salts	E3ag	G13
7-9	25	Concentration constants; <i>I</i> = 0.5(KCl)		
7-2	25	p <i>K</i> for hydrolysis of Fe ²⁺ , from rate of H ₂ O ₂ decomposition as function of pH in presence of Fe ²⁺ ; <i>I</i> = 1(NaClO ₄)	KIN	W14
3-3	25	log <i>K</i> for Fe ²⁺ + 3OH ⁻ ⇌ Fe(OH) ₃ ; is estimated from polarography to be 7.85 in 1.375 N NaOH		S21
		Other measurements: H45, I34		
109. (Aquo) Iron(III) ion, Fe³⁺				
2-71	15	p <i>K</i> for hydrolysis of Fe ³⁺ ; concentration constant; <i>I</i> = 0.01	O5	T21
2-46	25			
2-29	35			
2-30	20	p <i>K</i> for hydrolysis of Fe ³⁺ ; <i>I</i> = 0.025 to 0.15 (NaClO ₄ , HClO ₄); extrapolated to <i>I</i> = 0	O6	R16
2-34	25			
2-38	18	p <i>K</i> for hydrolysis of Fe ³⁺ ; <i>I</i> = 0.01 to 0.03; extrapolated to <i>I</i> = 0	O6	M30
2-19	25			
2-02	32			
2-96	18	<i>I</i> = 1(NaClO ₄); constants are also given for 2FeOH ²⁺ ⇌ Fe ₂ (OH) ₄ ⁺		
2-79	25			
2-61	32			
2-17	25	p <i>K</i> for hydrolysis of Fe ³⁺ ; <i>I</i> = 0.015 to 3.0; extrapolated to <i>I</i> = 0 using Debye-Hückel equation; constants are also given for 2FeOH ²⁺ ⇌ Fe ₂ (OH) ₄ ⁺	O6	M31
2-19	25	For <i>I</i> = 0		
2-63	20-22	<i>I</i> = 0.1(KNO ₃)	O6	S47
2-80	25	<i>I</i> = 0.5(NaClO ₄)	O5	P35
2-92	21	<i>I</i> = 0.55; in D ₂ O	O6	W20
2-74	20	Successive p <i>K</i> values for hydrolysis of Fe ³⁺ ; <i>I</i> = 1(NaClO ₄); -log <i>K</i> = 2.85 for 2Fe ³⁺ + 2H ₂ O ⇌ Fe ₂ (OH) ₂ ⁴⁺ + 2H ⁺	REDOX	H75 P18
2-83	25	Successive p <i>K</i> values for hydrolysis of Fe ³⁺	E	I18
2-83	25		C	

3-05	3-26	25	Successive pK values for hydrolysis of Fe^{3+} ; $I = 3(NaClO_4)$; also $-\log K = 2.91$ for $2Fe^{3+} + 2H_2O \rightleftharpoons Fe_2(OH)_2^{4+} + 2H^+$. Values of $-\log K$ for $2Fe^{3+} + 2H_2O \rightleftharpoons Fe_2(OH)_2^{4+} + 2H^+$, from 15-41, are estimated from magnetic measurements. Other measurements: A31, B10, B52, B53, B76, B91, B94, C30, I20, L5, L33, O10, S81	REDOX	H46 M56
110. Isohyppophosphoric acid, $H_4P_2O_6$					
4-5	8-5	25	pK_2 , pK_3 ; $c = 0.02 M$	E3bg	B60
1-67	6-26		pK_2 , pK_3 ; $I = 0.1$ to $1.0 (Et_4NCl)$; extrapolated to $I = 0$; pK_1 estimated as 0.6	E3bg	C10a
111. (Aquo) Lanthanum(III) ion, La^{3+}					
~10		25	pK for hydrolysis of La^{3+} ; from hydrolysis of "pure" salt; $c = 0.001-0.01 M La_2(SO_4)_3$	E3ag	M38
9-06		25	pK_6 for hydrolysis of La^{3+} ; titration of $0.004-0.009 M La(ClO_4)_3$ with $0.02 M Ba(OH)_2$; $I = 0.3 (NaClO_4)$	E3b	F33a
8-98		25	ditto, using $0.02 M NaOH$	E3bg	B43
10-1		25	pK for hydrolysis of La^{3+} ; $I = 3(LiClO_4)$; $c = 0.1-1.0 M La(ClO_4)_3$; also $-\log K = 9.95$ for $2La^{3+} + H_2O \rightleftharpoons La_2OH^{5+} + H^+$; other species include $La_5(OH)_9^{6+}$ and $La_6(OH)_{10}^{6+}$	E3bg	W17
5-6		20	pK_6 ; $c = 0.01 M LaCl_3$	E3bg	D8
3-3		25	pK_6 ; estimated from solubility measurements of I. M. Kolthoff and R. Elmquist, <i>J. Am. Chem. Soc.</i> , 53 , 1217 (1931)		
~5		18	pK_6	DISTRIB	V6
112. (Aquo) Lead(II) ion, Pb^{2+}					
7-78		18	pK for hydrolysis of Pb^{2+} ; for $I = 0$; $c = 0.005-0.4 M (PbNO_3)_2$; also $\log K = -7.30$ for $2Pb^{2+} + H_2O \rightleftharpoons Pb_2OH^{3+} + H^+$, and $\log K = -20.93$ for $4Pb^{2+} + 4H_2O \rightleftharpoons Pb_4(OH)_4^{4+} + 4H^+$	E3bg	P16
8-66		20	pK for hydrolysis of Pb^{2+} ; $Pb_4(OH)_4^{4+}$ is also formed	E3bg	F4
7-93		25	pK for hydrolysis of Pb^{2+} ; $I = 2(NaClO_4)$; also $\log K = -19.35$ for $4Pb^{2+} + 4H_2O \rightleftharpoons Pb_4(OH)_4^{4+} + 4H^+$	H76	
8-84		25	pK for hydrolysis of Pb^{2+} ; $I = 2(NaNO_3)$; also $\log K = -7.11$ for $2Pb^{2+} + H_2O \rightleftharpoons Pb_2OH^{3+} + H^+$; $\log K = -21.72$ for $4Pb^{2+} + 4H_2O \rightleftharpoons Pb_4(OH)_4^{4+} + 4H^+$	E3ag	H77
7-1	10-1	25	pK values for stepwise hydrolysis of Pb^{2+} to $PbOH^+$, $(PbOH)_2$ and $(Pb(OH))_3$; $I = 1 (KNO_3)$		POLAROG G38

1.42
1.59
1.76

182
227
271

For alkalinity function of LiOH solutions, see L12a, M40, S73

114. (Aquo) Lutecium(III) ion, Lu³⁺

20 pK_b for LuOH²⁺ ⇌ Lu³⁺ + OH⁻; *c* = 0.01 M LuCl₃
 25 pK_a for hydrolysis of Lu³⁺; titration of 0.004–0.009 M Lu(ClO₄)₃
 with 0.02 M Ba(OH)₂; *I* = 0.3(NaClO₄)
 25 ditto, using 0.02 M NaOH

E3bg

F33a

115. (Aquo) Magnesium ion, Mg²⁺

25 pK_b; for *I* = 0; *c* = 0.03 M MgCl₂
 18 pK_b
 25 pK_b; *I* = 0
 18 pK_b; concentration constant; *c* = 0.1–0.5 N MgCl₂; salt effect
 on indicator
 25 pK for hydrolysis of Mg²⁺; *I* = 3(NaCl, MgCl₂)
 30 pK for hydrolysis of Mg²⁺; *I* = 0.1(KCl); *c* = 0.01 M
 100 pK for hydrolysis of Mg²⁺; taking pK_w = 12.88; *c* = 0.06 M
 MgCl₂; rate of inversion of sucrose

E3bg

S74

G28

H71

K39

E3bg, h

L22

Cl2

K69

116. (Aquo) Manganese(II) ion, Mn²⁺

15 pK for hydrolysis of Mn²⁺; *I* = 0.002 to 0.04; extrapolated to
I = 0 by fitting extended Debye-Hückel equation

E3bg

P20

25
20
25
30
10.19
10.10
30
100

pK for hydrolysis of Mn²⁺; *I* = 0.1(KCl)
 pK for hydrolysis of Mn²⁺

E3bg

Cl2

K69

117. (Aquo) Manganese(III) ion, Mn³⁺

25 pK for hydrolysis of Mn³⁺; *I* = 4(Mn(ClO₄)₂.HClO₄)
 23 *I* = 5.3 to 6.1 (Mn(ClO₄)₂.HClO₄)
 23 *I* = 6(HClO₄, NaClO₄)

O6

W13

F1

D34

118. Manganic acid, H₂MnO₄

35 pK₃; *I* ≈ 0.1

KIN

L37