

Electronic Supplement¹

1 Single solute molalities and water uptake

Single solute molalities and the associated water uptake are shown in Fig. (A1a) and (A1b) for all compounds listed in Table 1, complementing Fig. (2a) and (2b).

2 Equilibrium Thermodynamics

To derive the standard definitions and relevant equations for chemical equilibrium we follow the “classical” thermodynamics (e.g. Denbigh, 1981; Seinfeld and Pandis, 1998; Wexler and Potukuchi, 1998). “Classical”, since water is generally omitted in equilibrium equations unless explicitly involved in the reaction, based on the assumption that water is neither consumed nor produced. However, when hydration is involved, water is consumed and released, which causes inconsistencies in the standard treatment.

2.1 Gibb’s free energy

The general condition for thermodynamical and chemical equilibrium is that the total free energy of the system – usually referred to as Gibb’s free energy (G) – is at a minimum, which implies that $dG = 0$ (see e.g. Denbigh, 1981).

The Gibb’s free energy expresses the sum of total free energies of the system, $G \equiv U + PV - TS$, by which the first term on the rhs denotes internal energy (U), and is always positive. The second term accounts for the energy associated with either changing a volume (V) at constant pressure (P), or by changing the pressure at constant volume; and can be either negative (e.g. by expansion) or positive (e.g. by compression). The third term accounts for the energy that is associated with the degree of order of the system at given temperature (T) and entropy (S), and is always positive.

Any differential change in G yields

$$dG = dU + PdV + VdP - TdS - SdT = 0, \quad (1)$$

by which at constant temperature ($dT = 0$) the entropy term (SdT) disappears. And at constant pressure ($dP = 0$), Eq. (1) reduces to

$$dG|_{T,P} = dU + PdV - TdS = 0. \quad (2)$$

With the definition of the internal energy, known as the Euler equation, $U = TS - PV + \mu n$, any differential change in U leads to (at constant T and P)

$$dU|_{T,P} = TdS - PdV + \mu dn, \quad (3)$$

¹provides (1) additional figures and (2) the derivation of the standard definitions of equilibrium thermodynamics referred to in: “Reformulating atmospheric aerosol thermodynamics and hygroscopic growth into haze and clouds” (Metzger and Lelieveld, 2007).

where $SdT - VdP + nd\mu = 0$ is known as the Gibbs-Duhem equation. It shows that three intensive variables are not independent (T, P , and μ are intensive, the others extensive) – if one is known, the value of the third can be determined from the Gibbs-Duhem equation.

Upon substitution of Eq. (3) in Eq. (2), i.e. eliminating dU with the Gibbs-Duhem relation, we can express dG in terms of a concentration change, which is zero when the j th – chemical reaction that forms compound j reaches equilibrium, i.e.

$$dG|_{T,P} = \sum_{i=1}^k \mu_{ij} n_{ij} = 0; \quad (4)$$

n_{ij} denotes the amount of the i th – of k – components and μ_{ij} its chemical potential.

2.2 Equilibrium constant

For the j th – chemical reaction, the amount of each component is $n_{ij} = n_{ij}^o + \nu_{ij} \varepsilon_j$, where n_{ij}^o is the initial amount of each component; ν_{ij} is their stoichiometric coefficient and ε_j their reaction coordinate.

Taking the derivative of n_{ij} , i.e. $dn_{ij} = \nu_{ij} d\varepsilon_j$ with $d\varepsilon_j = 1$, when substituted into Eq. (4) gives

$$\sum_{i=1}^k \mu_{ij} dn_{ij} = 0. \quad (5)$$

The chemical potentials μ_{ij} are usually expressed in terms of their activities a_{ij} , i.e. $\mu_{ij} = \mu_{ij}^o + RT \ln a_{ij}$. μ_{ij}^o denotes the chemical potential at a standard state and equals the partial molar Gibbs free energy g_{ij}^o with units kilo Joule per mol [kJ/mol]; R [J/mol/K] denotes the universal gas constant, T [K] the temperature. Upon substitution into Eq. (5), i.e.

$$\sum_{i=1}^k \nu_{ij} g_{ij}^o + RT \times \sum_{i=1}^k \nu_{ij} \ln a_{ij} = 0. \quad (6)$$

Eq. (6) yields upon rearranging and exponentiation, with expressing the sum of the logarithm as their product, the equilibrium constant ($K_{c,j}$) for the j th – chemical reaction at given temperature (T), i.e.

$$K_{c,j} = \exp\left(-\sum_{i=1}^k \nu_{ij} g_{ij}^o / RT\right) = \prod_{i=1}^k a_{ij}^{\nu_{ij}}. \quad (7)$$

The temperature dependency of the equilibrium constant is calculated from the van’t Hoff equation (now dropping the indices ij), i.e.

$$d \ln K_c / dT = \Delta H_f(T) / (RT^2), \quad (8)$$

where $\Delta H_f(T)$ is the standard molar enthalpy change of formation of the j th – compound at temperature T . For a small temperature change ΔH_f can be approximated by

$$\Delta H_f(T) = \Delta H_f^o(T^o) + \Delta C_p^o(T - T^o). \quad (9)$$

$\Delta H_f^o(T^o)$ [J/mol] denotes the standard molar enthalpy change of the reaction at a reference temperature T^o , (usually the standard-state temperature 298.15 K), and $\Delta Cp^o(T - T^o)$ the change of the standard molar heat capacity at constant pressure compared to T^o ; it is implicitly assumed here that ΔH_f^o and ΔCp^o are constant over the temperature range $T - T^o$.

Substituting Eq. (9) into Eq. (8) and integrating over $T - T^o$, the temperature dependency of the equilibrium constant is obtained, i.e.

$$K_c(T) = K_c^o(T^o) \times \exp \left[-\Delta H_f^o / (RT^o) \times (T^o/T - 1) - \Delta Cp^o / R (1 + \ln(T^o/T) - T^o/T) \right]. \quad (10)$$

The data on ΔG^o , ΔH_f^o , and ΔCp^o (needed to calculate equilibrium constants and their temperature dependency) can be obtained for instance from the CRC Handbook of Chemistry and Physics (2006).

Finally, the activities used in Eq. (7) of single-salt solutions consist of \pm -ion pairs, by which the \pm -ion pair activity is usually expressed as (Robinson and Stokes, 1965)

$$a_{\pm}^{\nu_{\pm}} = (\gamma_{\pm} m_{\pm})^{\nu_{\pm}} = a_{+}^{\nu_{+}} a_{-}^{\nu_{-}} = (\gamma_{+} m_{+})^{\nu_{+}} (\gamma_{-} m_{-})^{\nu_{-}} = \gamma_{\pm}^{\nu_{\pm}} m_{+}^{\nu_{+}} m_{-}^{\nu_{-}}. \quad (11)$$

ν_{+} and ν_{-} denote the number of moles of cations and anions per mole of solute (\pm -ion pair) dissociating completely into $\nu_{\pm} = \nu_{+} + \nu_{-}$ ions. m_{\pm} , m_{+} , m_{-} denote the corresponding molalities, i.e. the concentration of the ions in the solution, by which the molality is defined as mole solute per kg solvent (water) [mol/kg(H₂O)]. $\gamma_{\pm}, \gamma_{+}, \gamma_{-}$ are the corresponding mean molal binary activity coefficient [kg(H₂O)/mol] of the \pm -ion pair, and the activity coefficients of the cations and anions, respectively. The activity coefficients are introduced to correct the solution molalities for non-ideality resulting from incomplete dissociation and ion-ion interactions at higher solute concentrations.

Eq. (7) becomes in terms of activity coefficients and molalities

$$K_{c,j} = \prod_{i=1}^k (\gamma_{ij} m_{ij})^{\nu_{ij}}, \quad (12)$$

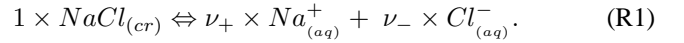
where the subscript i denotes i th- component of the j th- compound (reaction).

Since $K_{c,j}$ is a characteristic constant for the j th- chemical equilibrium reaction and usually determined with the aid of Eq. (7) from experimentally gained thermodynamic data, Eq. (12) enables to relate and calculate various solute properties that are difficult to measure, for instance solute activities (a_{ij}), molalities (m_{ij}) and activity coefficients (γ_{ij}) of \pm -ion pairs or of the individual cations and anions.

However, the summation or product over i is usually only made over the k -components that form the j th- compound, and water is usually omitted, as long it is neither consumed nor produced – even when being essential for the equilibrium reaction to take place and consumed by the solute hydration.

2.2.1 Example

For the example in Fig. 1b the stoichiometrical (i.e. on a molar basis) reaction of the dissociation of crystalline (cr) sodium chloride (NaCl) \pm -ion pair into the aqueous (aq) sodium (Na⁺) cation and chloride (Cl⁻) anion is



The corresponding equilibrium constant for this j th- reaction (R1) with its i th- components gives

$$K_{NaCl_{(cr)}} = [Na_{(aq)}^{+}]^{\nu_{+}} \times [Cl_{(aq)}^{-}]^{\nu_{-}} \times \gamma_{+}^{\nu_{+}} \gamma_{-}^{\nu_{-}}. \quad (K1)$$

2.3 Solubility

Although the above also applies to the calculation of the solubility constants, there are however some peculiarities about solubility equilibria that help apply important simplifications, despite the fact that solubility calculations can become quite complex, especially when involving complex-ion equilibria which even can dissolve insoluble salts.

First, at equilibrium the solution is saturated, i.e. it contains the maximum concentration of ions that can exist in equilibrium with its solid (crystalline) phase. In this case the ion concentrations of Na⁺_(aq) and Cl⁻_(aq) are sufficiently high so that the rate at which precipitation occurs exactly balances the rate at which NaCl_(cr) dissolves. Thus, there is no change in the concentration of these ions with time when the reaction is at equilibrium. The amount of solute that must be added to a given volume of solvent to form a saturated solution is called the solubility of the solute. In other words, if the system is at equilibrium the ion product, i.e. the product of the concentrations of the ions, equals the solubility product for the solute. If the ion product is larger or smaller than the solubility product for the solute, the system is not at equilibrium, but can rapidly adjust according to Le Chatelier's principle, so that equilibrium is restored after the excess ions precipitate from solution as a solid (NaCl_(cr)), or solids dissolves until any deficit in the ion product is compensated.

The solubility constant expression for reaction (R1) can be calculated from (K1) from its product with the concentration of the solid, and the concentration of a solid (e.g. NaCl_(cr)) is calculated from its density and its molar mass. For instance, for solid NaCl_(cr), i.e.

$$\frac{\rho_{NaCl_{(cr)}}}{1 \text{ cm}^3} = \frac{2.17 \text{ g NaCl}_{(cr)}}{1 \text{ cm}^3} \times \frac{1 \text{ cm}^3}{1 \text{ mL}} \times \frac{1000 \text{ mL}}{1 \text{ L}} \times \frac{1 \text{ mol NaCl}_{(cr)}}{58.44 \text{ g NaCl}_{(cr)}} = 37.13 \text{ mol NaCl}_{(cr)} / \text{L}. \quad (13)$$

Since the concentration of the solid is constant, which has no effect on the equilibrium, it is built into the solubility constant for the equilibrium reaction. According to (K1) we have

$$K_{sp, \text{NaCl}_{cr}} = K_{\text{NaCl}_{(cr)}} \times m_{\text{NaCl}_{(cr)}}. \quad (\text{K2})$$

(K2) is a constant and proportional to the solubility of the salt and hence called the solubility product equilibrium constant ($K_{sp, \text{NaCl}_{cr}}$). For reaction (R1), $K_{sp, \text{NaCl}_{(cr)}}(T^o) = 37.661 \text{ [mol}^2/\text{kg}^2]$ with constants $a = -\Delta H_f^o/(RT^o) = -1.56$ and $b = \Delta C_p^o/R = 16.9$, needed to calculate the temperature dependency with Eq. (10); $RT^o = 8.314 \text{ [J/mol/K]} \times 298.15 \text{ [K]} = 2.479 \text{ [J/mol]}$.

The solubility product constant requires that ν_+ cations ($\text{Na}_{(aq)}^+$) are released with ν_- anions ($\text{Cl}_{(aq)}^-$), because there is no other source of either ion in this solution, so that the concentrations of the ions at equilibrium are the same

$$a_{\pm}^{\nu_{\pm}} = a_+^{\nu_+} = a_-^{\nu_-}. \quad (14)$$

In general, it is assumed that salt solutes dissociate into their ions when dissolved in water. Ionic compounds dissolve if the energy released from the ions interactions with water molecules compensates for both a) the energy needed to break the ionic bonds of the solid and b) the energy required to separate the solvent (water) molecules.

2.4 Nernst equation

For charged species electrical forces must be considered. The potential for an electrochemical reaction is described by (Nernst, 1889)

$$E = E^o - R T / (z F) \ln(Q_c). \quad (15)$$

In the Nernst equation (15), E [V] is the electrochemical cell potential at some moment in time, E_o [V] is the cell potential for the reaction at standard-state conditions, R [J/mol/K] the ideal gas constant in joules per mole and Kelvin, T [K] the temperature in Kelvin, z [−] the number of moles of electrons transferred in the balanced equation for the reaction, $F = 96484 \text{ [C/mol]}$ the Faraday-constant that expresses the charge on a mole of electrons (calculated from Avogadro's number and the charge on a single electron ($F = 6.022045 \times 10^{23} [1/\text{mol}] \times 1.6021892 \times 10^{-19} [\text{As/e}^-]$), and Q_c [−] the reaction quotient. The magnitude of the cell potential is a measure of the driving force behind a reaction. The larger the value of the cell potential, the further the reaction is from equilibrium, while at equilibrium ($\Delta G = 0$) the overall cell potential E [V] = 0. Since then the reaction quotient equals the equilibrium constant ($Q_c = K_c$), the Nernst equation is related to the partial molar Gibbs free energy by

$$g^o = \pm z F E^o = RT \ln(K_c), \quad (16)$$

and can hence be used to measure the equilibrium constant for a reaction.

The condition that the overall cell potential is zero at equilibrium ($E = 0$) implies electroneutrality. The ability of a binary solution (one solute and solvent) to conduct an electric current by the ions resulting from dissociation must be compensated by an ion flow that results from the dissociation of the solvent (water), for which two moles of water are consumed, $2 \text{ H}_2\text{O}_{(aq)} \rightleftharpoons \text{H}_3\text{O}_{(aq)}^+ + \text{OH}_{(aq)}^-$. Since measurements suggest that the ability of pure water to conduct an electrical current at 25 °C contains $1.0 \times 10^{-7} [\text{mol/L}]$ moles per liter of each of these ions, i.e. $[\text{H}_3\text{O}_{(aq)}^+] = [\text{OH}_{(aq)}^-] = 1.0 \times 10^{-7} [\text{mol/L}]$. This product remains constant at equilibrium and yields the water dissociation equilibrium constant, $K_w = 1.0 \times 10^{-14}$; one kilogram of pure water always contains $1000 \text{ [g]}/M_{\text{H}_2\text{O}} \text{ [g/mol]} = 1000/18.015 = 55.5093 \text{ [mol]}$.

Although K_w is defined in terms of the dissociation of water, this equilibrium constant expression is equally valid for solutions of acids and bases dissolved in water. Regardless of the source of the $\text{H}_3\text{O}_{(aq)}^+$ and $\text{OH}_{(aq)}^-$ ions in water, the product of the concentrations of these ions at equilibrium at 25 °C is $1.0 \times 10^{-14} [\text{mol}^2\text{L}^{-2}]$. Thus adding an acid or base to water therefore has an effect on the concentration of both the $\text{H}_3\text{O}_{(aq)}^+$ and $\text{OH}_{(aq)}^-$ ions; they can be hence related in terms of the definition of a negative of the logarithm by $pH + pOH = 14$, with $pH = -\log [\text{H}_3\text{O}_{(aq)}^+]$ and $pOH = -\log [\text{OH}_{(aq)}^-]$.

Similarly, the discussion can be extended to acid/base-dissociation equilibria, including cation-acids and anion-bases, which yield pK_a and pK_b values that indicate the strength of the acid or base, respectively. However, for chemical reactions involving hydration the assumption that water can be omitted in equilibrium reactions does not hold, since water is actually consumed by hydration. Neglecting water in equilibrium reactions therefore introduces a conceptual problem and unnecessarily complicates the equations. In the present work we therefore reformulate chemical equilibrium to include water and exploit the implications for atmospheric aerosol modeling.

References

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Appendix A List of indices, symbols and chemical substances

Table A1. List of Indices.

Abbreviation	Name
(cr)	crystalline
(aq)	aqueous
(g)	gaseous
$+$	superscript for “cation”
$-$	superscript for “anion”
\pm	superscript for “ion pair”
o	superscript for “standard state”
i	subscript for “ i th of k -components”
j	subscript for “ j th-compound or chemical reaction”
k	subscript for “the total number of components”

Table A2. List of Symbols.

Symbol	Name	Unit
ν_{ij}	Stoichiometric coefficient: of i th-component, j th-reaction	–
ν_+	of the cation	–
ν_-	of the anion	–
ν_{\pm}	of the ion-pair	–
E	Electrochemical cell potential	V
E^o	Electrochemical cell potential at standard state	V
F	Faraday-constant	C/mol
G	Gibb’s free energy	kJ
H_f	Enthalpy of formation	kJ
$K_{c,j}$	Equilibrium constant (j th-chemical reaction)	–
K_c^o	Equilibrium constant at standard state	–
K_{sp}	Solubility product constant	(mol/kg) $^{\nu_{ij}}$
pH	Acid-base indicator (potentia Hydrogenia)	–
pK_a	Acid indicator	–
pK_b	Base indicator	–
P	Pressure	Pa
Q_c	Reaction quotient	–
R	Universal gas constant	J/mol/K
S	Entropy	J
T	Temperature	K
T^o	Standard state temperature	K
U	Internal energy	kJ
V	Volume	m ³
z	Number of moles of electrons transferred (balanced equation of j th-reaction)	–

Table A2. List of Symbols (continued).

Symbol	Name	Unit
ε_j	Reaction coordinate (j th-reaction)	–
n, n_{ij}	Number of moles	mol
n_{ij}^o	Initial amount (moles)	mol
m_+	Cation molality	mol/kg
m_-	Anion molality	mol/kg
m_{\pm}	Solute (ion-pair) molality	mol/kg
μ	Chemical potential	kJ/mol
μ_{ij}^o	Standard state chemical potential	kJ/mol
γ_+	(Molal) cation activity coefficient	kg/mol
γ_-	(Molal) anion activity coefficient	kg/mol
γ_{\pm}	Mean (molal) binary activity coeff.	kg/mol
a_{ij}	Activity of i th-component	–
a_{\pm}	Ion-pair activity of j th-compound	–
g_{ij}^o	(Partial) molar Gibb’s free energy at standard state	kJ/mol
$d\varepsilon_j$	Differential change in: reaction coordinate (j th-reaction)	–
dC_p	heat capacity	J/mol
dG	Gibb’s free energy	kJ
dH_f	enthalpy of formation	kJ
$d\mu$	chemical potential	kJ/mol
dn	number of moles	mol
dP	pressure	Pa
dS	entropy	J
dT	temperature	K
dU	internal energy	kJ
dV	volume	m ³
ΔC_p^o	Total change in: standard state molar heat capacity	kJ/mol
ΔG^o	standard state molar Gibb’s free energy	kJ/mol
ΔH_f^o	molar enthalpy of formation at standard state	kJ/mol

Table A3. List of Chemical Substances

Cations(+) and Anions(-)	Name
H ⁺	Hydrogen
NH ₄ ⁺	Ammonium
Na ⁺	Sodium
K ⁺	Potassium
Ca ²⁺	Calcium
Mg ²⁺	Magnesium
Fe ^{2+/3+}	Iron(II,III)
PO ₄ ³⁻	Phosphate
SO ₄ ²⁻	Sulfate
HSO ₄ ⁻	Bisulfate
NO ₃ ⁻	Nitrate
Cl ⁻	Chloride
Br ⁻	Bromide
I ⁻	Iodide
CO ₃ ²⁻	Carbonate
HCO ₃ ⁻	Bi-carbonate
OH ⁻	Hydroxide
CHO ₂ ⁻ = H(COO) ⁻	Formate
C ₂ H ₃ O ₂ ⁻ = CH ₃ (COO) ⁻	Acetate
C ₂ O ₄ ²⁻ = (COO) ₂ ⁻	Oxalate
C ₆ H ₅ O ₇ ³⁻ = (HO)C(COO) ₃ ⁻	Citrate

Table A3. List of Chemical Substances (continued)

Ammonium salts	Name
(NH ₄) ₃ PO ₄	Ammonium phosphate
(NH ₄) ₂ SO ₄	Ammonium sulfate
NH ₄ HSO ₄	Ammonium bisulfate
NH ₄ NO ₃	Ammonium nitrate
NH ₄ Cl	Ammonium chloride
NH ₄ Br	Ammonium bromide
NH ₄ I	Ammonium iodide
(NH ₄) ₂ CO ₃	Ammonium carbonate
NH ₄ HCO ₃	Ammonium bicarbonate
NH ₄ OH	Ammonium hydroxide
NH ₄ CHO ₂ = H(COONH ₄)	Ammonium formate
NH ₄ C ₂ H ₃ O ₂ = CH ₃ (COONH ₄)	Ammonium acetate
(NH ₄) ₂ C ₂ O ₄ = (COONH ₄) ₂	Ammonium oxalate
(NH ₄) ₂ HC ₆ H ₅ O ₇ = (HO)C(COONH ₄) ₃	Ammonium citrate

Table A3. List of Chemical Substances (continued)

Acids	Name
H ₃ PO ₄	Phosphoric acid
H ₂ SO ₄	Sulfuric acid
HNO ₃	Nitric acid
HCl	Hydrochloric acid
HBr	Hydrogen bromide
HI	Hydrogen iodide
H ₂ CO ₃	Carbonic acid
H ₂ O	Water
CH ₂ O ₂ = H(COOH)	Formic acid
C ₂ H ₄ O ₂ = CH ₃ (COOH)	Acetic acid
C ₂ H ₂ O ₄ = (COOH) ₂	Oxalic acid
C ₆ H ₈ O ₇ = (HO)C(COOH) ₃	Citric acid

Table A3. List of Chemical Substances (continued)

Sodium salts	Name
Na ₃ PO ₄	Sodium phosphate
Na ₂ SO ₄	Sodium sulfate
NaHSO ₄	Sodium bisulfate
NaNO ₃	Sodium nitrate
NaCl	Sodium chloride
NaBr	Sodium bromide
NaI	Sodium iodide
Na ₂ CO ₃	Sodium carbonate
NaHCO ₃	Sodium bicarbonate
NaOH	Sodium hydroxide
NaCHO ₂ = H(COONa)	Sodium formate
NaC ₂ H ₃ O ₂ = CH ₃ (COONa)	Sodium acetate
Na ₂ C ₂ O ₄ = (COONa) ₂	Sodium oxalate
Na ₃ C ₆ H ₅ O ₇ = (HO)C(COONa) ₃	Sodium citrate

Table A3. List of Chemical Substances (continued)

Potassium salts	Name
K_3PO_4	Potassium phosphate
K_2SO_4	Potassium sulfate
$KHSO_4$	Potassium bisulfate
KNO_3	Potassium nitrate
KCl	Potassium chloride
KBr	Potassium bromide
KI	Potassium iodide
K_2CO_3	Potassium carbonate
$KHCO_3$	Potassium bicarbonate
KOH	Potassium hydroxide
$KCHO_2 = H(COOK)$	Potassium formate
$KC_2H_3O_2 = CH_3(COOK)$	Potassium acetate
$K_2C_2O_4 = (COOK)_2$	Potassium oxalate
$K_3C_6H_5O_7 = (HO)C(COOK)_3$	Potassium citrate

Table A3. List of Chemical Substances (continued)

Calcium salts	Name
$Ca_3(PO_4)_2$	Calcium phosphate
$CaSO_4$	Calcium sulfate
$Ca(NO_3)_2$	Calcium nitrate
$CaCl_2$	Calcium chloride
$CaBr_2$	Calcium bromide
CaI_2	Calcium iodide
$CaCO_3$	Calcium carbonate
$Ca(OH)_2$	Calcium hydroxide
$Ca(CHO_2)_2$	Calcium formate
$Ca(C_2H_3O_2)_2$	Calcium acetate
CaC_2O_4	Calcium oxalate
$Ca_3(C_6H_5O_7)_2$	Calcium citrate

Table A3. List of Chemical Substances (continued)

Magnesium salts	Name
$Mg_3(PO_4)_2$	Magnesium phosphate
$MgSO_4$	Magnesium sulfate
$Mg(NO_3)_2$	Magnesium nitrate
$MgCl_2$	Magnesium chloride
$MgBr_2$	Magnesium bromide
MgI_2	Magnesium iodide
$MgCO_3$	Magnesium carbonate
$Mg(OH)_2$	Magnesium hydroxide
$Mg(CHO_2)_2$	Magnesium formate
$Mg(C_2H_3O_2)_2$	Magnesium acetate
MgC_2O_4	Magnesium oxalate
$Mg_3(C_6H_5O_7)_2$	Magnesium citrate

Table A3. List of Chemical Substances (continued)

Iron salts	Name
$FePO_4$	Iron(III) phosphate
$Fe_2(SO_4)_3$	Iron(III) sulfate
$Fe(NO_3)_3$	Iron(III) nitrate
$FeCl_3$	Iron(III) chloride
$FeBr_3$	Iron(III) bromide
FeI_3	Iron(III) iodide
$FeCO_3$	Iron(II) carbonate
$Fe(OH)_3$	Iron(III) hydroxide
$Fe(CHO_2)_3$	Iron(III) formate
$FeOH(C_2H_3O_2)_2$	Iron(III) acetate
$Fe_2(C_2O_4)_3$	Iron(III) oxalate
$FeC_6H_5O_7$	Iron(III) citrate

Table A4. List of Chemical Substances (continued)

Non salt solutes	Name
NH_3	Ammonia
$(CH_3)_2CO$	Acetone
CH_3OH	Methanol
CH_3CH_2OH	Ethanol
$C_6H_{12}O_6$	D-Fructose
$CH_2(CHOH)_4CH_2OH$	D-Mannitol
$C_{12}H_{22}O_{11}$	Sucrose