GLOSSARY OF TERMS RELATED TO SOLUBILITY

(IUPAC Recommendations 2008)

Prepared for publication by

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Glossary of terms related to solubility

(IUPAC Recommendations 2008)

Abstract: Phenomena related to the solubility of solids, liquids, and gases with one another are of interest to scientists and technologists in an array of disciplines. The diversity of backgrounds of individuals concerned with solubility creates a potential for confusion and miscommunication and heightens the need for an authoritative glossary of terms related to solubility. This glossary defines 166 terms used to describe solubility and related phenomena. The definitions are consistent with one another and with recommendations of the International Union of Pure and Applied Chemistry for terminology and nomenclature.

Keywords: solubility; phase equilibria; solution equilibria; solid-state equilibria; molten salt equilibria; thermodynamics; IUPAC Analytical Chemistry Division.

1. INTRODUCTION

Disciplines concerned with solubility and related phenomena extend well beyond the traditional branches of chemistry to a wide range of biomedical, environmental, and industrial fields including, for example, mineralogy, pharmacology, oceanography, and petroleum engineering. The diversity in the technical backgrounds and training of individuals concerned with solubility heightens the potential for confusion and miscommunication of both concepts and data related to solubility. This glossary seeks to reduce such confusion and miscommunication by presenting a set of solubility-related terms that are consistent with one another and with IUPAC recommendations for chemical terminology (the online “Gold Book”) [1], as well as with specific recommendations for quantities, units, and symbols (the “Green Book”, 3rd ed.) [2].

This glossary defines 166 terms which were selected by the authors as the central set related to solubility. Within each definition, terms defined elsewhere in the glossary are indicated by italics upon their first use. Inevitably, some users will seek terms that are not defined here or will wish to consult definitions of terms used in these definitions. In both cases, the authors recommend the IUPAC Compendium of Chemical Terminology (the “Gold Book”), particularly the online version [1]. The authors also recommend the IUPAC publication Quantities, Units and Symbols in Physical Chemistry, 3rd ed. (the “Green Book”) [2] as an authoritative and consistent guide to the presentation of chemical data.

Definitions of particular terms have, in many cases, been expanded or modified compared to the corresponding definitions given in the Gold Book [1], which means that such definitions have been modified compared to the original IUPAC recommendations used to prepare entries in the Gold Book. The objective has been to make the definitions as clear and as useful as possible within the general format of the glossary, but without changing significantly the meaning of the Gold Book entries. Definitions taken from another source conclude with a notation “from [ref]” to indicate that source; where a definition in any source has been modified to produce the definition given here, the notation is “modified from [ref]”. Where synonymous terms are in common usage, the recommended term (“main term”) is followed on the second line of the glossary entry (and on subsequent lines where necessary) by the synonym or synonyms which are then followed by the definition and notes, if any. The synonyms are also listed separately in their proper alphabetical position, with the annotation “See (main term)”. This approach can, for instance, be observed by consulting the entries for “mole ratio” (a main term) and “amount ratio” (its synonym).
2. GLOSSARY OF TERMS

absorption coefficient, $\beta^*$ (in gas solubility)

Volume $V^g$ of an amount $n^B_1$ of a dissolved gas at a given standard temperature, usually $T^\circ = 273.15$ K, and total standard pressure $p^\circ$ divided by the volume of the solvent of volume $V^1$ that contains an amount $n^A_1$ of solvent at the same temperature $T$ and pressure $p$.

**Note 1:** There is only one absorption coefficient, as compared to the Bunsen, Ostwald, and Kuenen coefficients, because the solvent in the definition is always the pure solvent, not the gaseous solution. The mathematical definition is

$$\beta^*_B = \frac{V^g(T, p^A_B + p^B_B = p^\circ, n^B_1)/V^1(T, p^\circ, n^A_1)}{V^1(T, p^\circ, n^A_1)}$$

where $p^A_B, p^B_B$ are the partial pressures of solvent and gas.

**Note 2:** For an ideal gas, the absorption coefficient and Bunsen coefficient $\alpha^*$ are related by

$$\beta^*/\alpha^* = p^B_B/p^\circ = (1 - p_A/p^\circ)$$

since $V^g$ is inversely proportional to pressure.

**Note 3:** The relations between the molality $m^B_B(p^\circ)$ or mole fraction $x^B_B(p^\circ)$ of dissolved gas and the absorption coefficients are

$$\frac{1}{x^B_B(p^\circ)} = 1 + \frac{1}{m^B_B(p^\circ) M^A} = 1 + \frac{RT Z^B_B \left(1 + p_A / p^\circ\right)}{p^\circ V^m,A_B \beta^*_B}$$

where $V^A, V^m,A$ are the respective partial molar volume and molar volume of the solvent and $Z^B_B$ is the compression factor of the gas.

**Note 4:** The absorption coefficient and the related quantities for expression of gas solubility, the Bunsen, Kuenen, and Ostwald coefficients, appear frequently in the older literature of gas solubility determination. However, the modern practice, recommended here, is to express gas solubility as molality, mole fraction, or mole ratio.

From [3].

activity, $a$

relative activity
Factor in relation between chemical potential and composition of a mixture or solution. For a substance B,

$$a^*_B = \exp\left(\frac{\mu^*_B - \mu^*_B^\circ}{RT}\right)$$

where superscript $\circ$ denotes a standard chemical potential.

**Note 1:** The choice of standard state for the chemical potential must be specified.

**Note 2:** An equivalent definition is

$$a^*_B = \lambda^*_B / \lambda^*_B^\circ$$
where

$$
\lambda_B^\alpha = \exp \left( \frac{\mu_B^\alpha}{RT} \right)
$$

with $\lambda$ the absolute activity.

Note 3: Appears only as cross-references between activity and relative activity in [1].

From [2].

**activity coefficient, $f$, $\gamma_m$, $\gamma_c$, $\gamma_x$**

Dimensionless correction factor that multiplies the quantity used to express the composition of the substance (usually mole fraction, molality, or amount concentration) to produce the (relative) activity of the substance.

(a) Referenced to Raoult’s law, mole fraction basis. For a substance B in a liquid or solid mixture containing mole fractions $x_A, x_B, x_C, \ldots$ of the substances A, B, C, $\ldots$: a dimensionless quantity $f_B$ defined in terms of the chemical potential $\mu_B$ of B in the mixture by

$$
RT \ln (x_B f_B) = RT \ln \alpha_B = \mu_B (T, p, x) - \mu_B^* (T, p)
$$

where $x$ denotes the set of mole fractions $x_A, x_B, x_C, \ldots$ and $\mu_B^*$ is the standard chemical potential, defined as the value for pure B. The activity coefficient has the value 1 for pure substance B.

(b) Referenced to Henry’s law, molality basis. For a solute B in a solution (especially a dilute liquid solution) containing molalities $m_B, m_C, \ldots$, of solutes B, C, $\ldots$ in a solvent A: a dimensionless quantity $\gamma_{m,B}$ defined in terms of the chemical potential $\mu_B$ of B in the solution by

$$
RT \ln (\gamma_{m,B} m_B/m^\alpha) = RT \ln \alpha_B = \mu_B - \mu_{m,B}^\alpha
$$

where $m^\alpha = 1$ mol kg$^{-1}$ is the standard molality and $\mu_{m,B}^\alpha$ is the standard chemical potential, defined as the value at infinite dilution of all solutes. The activity coefficient has the value 1 at infinite dilution of all solutes.

(c) Referenced to Henry’s law, amount concentration basis. For a solute B in a solution (especially a dilute liquid solution) containing amount concentrations $c_B, c_C, \ldots$ of solutes B, C, $\ldots$ in a solvent A: a dimensionless quantity $\gamma_{c,B}$ defined in terms of the chemical potential $\mu_B$ of B in the solution by

$$
RT \ln (\gamma_{c,B} c_B/c^\alpha) = RT \ln \alpha_B = \mu_B - \mu_{c,B}^\alpha
$$

where $c^\alpha = 1$ mol dm$^{-3}$ is the standard amount concentration and $\mu_{c,B}^\alpha$ is the standard chemical potential, defined as the value at infinite dilution of all solutes. The activity coefficient has the value 1 at infinite dilution of all solutes.
(d) Referenced to Henry’s law, mole fraction basis. For a solute B in a solution (especially a dilute liquid solution) containing mole fractions \(x_B, x_C, \ldots\) of solutes B, C, \ldots in a solvent A: a dimensionless quantity \(\gamma_{x,B}\) defined in terms of the chemical potential \(\mu_B\) of B in the solution by

\[
RT \ln(\gamma_{x,B}x_B) = RT \ln a_B = \mu_B - \mu_{x,B}^\ominus
\]

\[
\mu_{x,B}^\ominus = \lim_{x_B \to 0} \left[ \mu_B - RT \ln \left( \gamma_{x,B}x_B \right) \right]
\]

where \(\mu_{x,B}^\ominus\) is the standard chemical potential, defined as the value at infinite dilution of all solutes. The activity coefficient has the value 1 at infinite dilution of all solutes.

See also activity coefficient at infinite dilution.
Modified from [1,2].

**activity coefficient at infinite dilution, \(f^\infty\)**

For a substance B, activity coefficient \(f_B\) extrapolated to infinite dilution:

\[
\ln f_B^{\infty} = \lim_{x_B \to 0} \left( \frac{\mu_B - \mu_B^\ominus}{RT} - \ln x_B \right)
\]

*Note 1:* Useful for dilute mixtures as an alternative to the standard chemical potential on a molality basis.

*Note 2:* The relation between the activity coefficient at infinite dilution and the standard chemical potentials is, for a solute B in a solvent A,

\[
\ln f_B^{\infty} = \frac{\mu_B^\ominus - \mu_B^\bullet}{RT} - \ln \left( M_A m_B^\bullet \right)
\]

where \(M_A\) is the molar mass of the solvent.

See also activity coefficient.

**amount**

See amount of substance.

**amount concentration, \(c\)**

amount-of-substance concentration
concentration
substance concentration (in clinical chemistry)
molarity (in older literature).

Amount of a constituent in a mixture divided by the volume of the mixture.

*Note 1:* For constituent B, amount concentration is often denoted [B].

*Note 2:* Concentration alone may be used where there is no ambiguity about its meaning.

*Note 3:* The common unit is mole per cubic decimeter (mol dm\(^{-3}\)) or mole per liter (mol L\(^{-1}\)) sometimes denoted by (small capital) \(M\).

Modified from [1].

amount fraction, \( x \)
See mole fraction.

amount of substance, \( n \)
amount
chemical amount
Base quantity in the SI system of quantities. It is the number of elementary entities divided by the Avogadro constant.

Note 1: Amount of substance is proportional to the number of entities, the proportionality constant being the reciprocal Avogadro constant. Therefore, amount of substance, just as number of entities, must be accompanied by a specification of the type of entities.

Note 2: The words “of substance” may be replaced by the specification of the entity, for example: amount of chlorine atoms, \( n(\text{Cl}) \), amount of chlorine molecules, \( n(\text{Cl}_2) \). No specification of the entity might lead to ambiguities [amount of sulfur could stand for \( n(\text{S}) \), \( n(\text{S}_8) \), etc.], but in many cases the implied entity is assumed to be known: for molecular compounds, it is usually the molecule [e.g., amount of benzene usually means \( n(\text{C}_6\text{H}_6) \)]; for ionic compounds, it is the simplest formula unit [e.g., amount of sodium chloride usually means \( n(\text{NaCl}) \)]; and for metals, it is the atom [e.g., amount of silver usually stands for \( n(\text{Ag}) \)].

Note 3: In some derived quantities, the words “of substance” are also omitted, e.g., amount concentration, amount fraction. Thus, in many cases the name of the base quantity is shortened to amount, and to avoid possible confusion with the general meaning of the word, the attribute chemical is added. Chemical amount is hence the alternative name for amount of substance. In the field of clinical chemistry, the words “of substance” should not be omitted and abbreviations such as substance concentration (for amount of substance concentration) and substance fraction are in use. The quantity had no name prior to 1969 and was simply referred to as the number of moles.

From [1,2].

amount-of-substance fraction
See mole fraction.

amount ratio, \( r \)
See mole ratio.

analytical method (in determination of solubility)
Class of experimental procedures for solubility determination in which a saturated solution is prepared and then analyzed to determine composition.
See also synthetic method.
aqumolality, $m^{(C)}$

solvomolality

For a solute B in a system containing solvent A and a reference solvent C,

$$m_B^{(C)} = \frac{m_B}{M - M_C}$$

where $m_B$ is the molality of B, $M_C$ is the molar mass of a reference component, and $\bar{M}$ is the average molar mass of a mixed solvent, defined for a system containing two solvents as

$$\bar{M} = x_{v,A} M_A + (1 - x_{v,A}) M_C$$

with $x_{v,A}$ the solvent amount fraction of solvent A.

Note 1: Used most frequently in discussing comparative solubilities in water (C) and heavy water (A) and their mixtures, but applies to any reference component, where solvomolality is a more appropriate term.

Note 2: For history and equivalence with older definitions, see [5].

azetotropic point

azeotrope

For a mixture, temperature and pressure (the azeotropic temperature and pressure) at which the compositions of the liquid and vapor phases become equal (the azeotropic composition), but the intensive properties of the two phases (such as molar volume) are different.

See also critical point.

From [6].

binary system

System containing two components.

See also unary system, ternary system, higher-order system.

binodal curve

See coexistence curve.

Bunsen coefficient, $\alpha$

Volume $V^g$ of an amount $n_B^1$ of a gas dissolved at a given standard temperature $T^o$ (usually 273.15 K) and given standard (partial) pressure $p_B^o$ (1 bar = 0.1 MPa or, in older literature, 1 atm) divided by the volume of the solvent $V^l$ containing an amount $n_A$ of solvent at temperature $T$ and the given total pressure $p^o$.

Note 1: There are two Bunsen coefficients, depending on whether the liquid is the equilibrium solution or the pure liquid, with mathematical definitions:

Bunsen coefficient, solution reference

$$\alpha_B = \frac{V^g(T^o, p^o, n_B^1)}{V^l(T^o, p^o, n_A, n_B^1)}$$

Bunsen coefficient, pure solvent reference

$$\alpha_B^* = \frac{V^g(T^o, p^o, n_B^1)}{V^l(T^o, p^o, n_A)}$$

where $n_B^1$ is the amount of dissolved gas in the liquid solution.

Note 2: The relations between the molality $m_B(p^o)$ or mole fraction $x_B(p^o)$ of dissolved gas and the Bunsen coefficients are
\[
\frac{1}{x_B(p^\bullet)} = 1 + \frac{1}{m_B(p^\bullet)M_A} = 1 + \frac{RT^\alpha Z_B^\alpha(T^\alpha)}{p^\alpha V_A^\alpha \alpha_B^\alpha}
\]
\[
\frac{1}{x_B(p^\bullet)} = 1 + \frac{1}{m_B(p^\bullet)M_A} = 1 + \frac{RT^\alpha Z_B^\alpha(T^\alpha)}{p^\alpha V_{m,A} \alpha_B^\alpha}
\]

where \(V_A\), \(V_{m,A}\) are the respective partial molar volume and molar volume of the solvent and \(Z_B\) is the compression factor of the gas.

**Note 3:** The Bunsen coefficient and the related quantities for expression of gas solubility; absorption coefficient, Kuenen coefficient, and Ostwald coefficient appear frequently in the older literature of gas solubility determination. However, the modern practice, recommended here, is to express gas solubility as molality, mole fraction, or mole ratio.

From [5].

catatatic reaction
See metatetic reaction.

chemical amount
See amount of substance.

chlorinity, \(Cl\), \(w_{Cl}\)
Mass of dissolved halides (reported as chloride) in sea water, brackish water, brine, or other saline solution divided by the mass of the solution.

**Note 1:** For a sample of sea water, 0.328 534 times the mass of pure reference silver, “Atomgewichtssilber”, necessary to precipitate the halides (chloride + bromide + iodide, but expressed as chloride) contained in a sample of sea water divided by the mass of this sample.

**Note 2:** Used (especially before 1978) in calculating salinity, and based on the assumption that the concentration ratios of the principal salts in sea water are constant throughout the oceans. This assumption is only approximate.

**Note 3:** The symbol \(Cl\) is recommended in [8].

**Note 4:** Before 1978, the usual unit for \(Cl\) expressed as a mass fraction was permil (no longer recommended). Now it is usually expressed as \(g\ kg^{-1}\) with or without the units given explicitly.

Example: the chlorinity of a sample of sea water is \(w_{Cl} = 19.375\ 0\ %\) or 19.375 0 g kg\(^{-1}\) or 19.375 0 × 10\(^{-3}\) or 19.375 0.
See also salinity.
From [7,8].

Clarke–Glew equation
Semi-empirical equation describing the temperature dependence of the standard Gibbs energy of solution,

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\[ \Delta_{\text{mix}} G^\circ(T, p^0)/RT = A_0 + A_1(T/T_{\text{ref}}) + A_2 \ln(T/T_{\text{ref}}) + A_3(T/T_{\text{ref}})^2 + \ldots \]

where the \( A_i \) are constants that can be related to thermodynamic quantities (but subject to uncertainty because of statistical correlations among the fitting coefficients) and \( T_{\text{ref}} \) is a reference temperature. See also Clarke–Glew–Weiss equation. Modified from [5,9].

**Clarke–Glew–Weiss equation**

*Clarke–Glew equation* with \( T_{\text{ref}} = 100 \) K.

Modified from [5,10].

**cloud point**

Critical solution point, particularly when used as an end-point of a turbidometric titration to determine solubility.

See also synthetic method.

**coexistence curve**

binodal curve

conodal curve

Boundary of stable phase separation (limits of solubility) in a liquid or solid system of two or more components.

*Note 1*: The locus of the compositions of two co-existing phases on a phase diagram.

*Note 2*: In a binary system, a plot of experimental variables such as temperature or pressure, or of theoretical variables such as Gibbs energy, against mole fraction displays the co-existence curve. In ternary or higher-order systems, coexistence curves are displayed on a plot of composition at constant temperature or pressure, usually on a ternary diagram.

See also critical solution point, spinodal curve, conjugate phases. Modified from [11,12].

**common ion effect**

Decrease in solubility of a salt when a second non-saturating salt with one ion in common with the salt is added to its saturated solution.

*Note*: Restricted in practice to composition regions where common ions do not form soluble complexes with the saturating salt and to salts with sufficiently low solubility and at sufficiently low ionic strengths that the activity coefficient of the salt is close to unity. At higher ionic strengths, or for salts with high solubility, the activity coefficient usually decreases with addition of non-saturating salt, resulting in a higher solubility.

**component**

thermodynamic component

independent component

*Constituent* whose amount or concentration can be varied independently in a mixture.

*Note*: The number of components in a given system is the minimum number of independent species necessary to define the composition of all phases of that system. Relations among the constituents, such as charge balance or chemical equilibria, must be taken...
into account in determining the number of components. The term “component” is also often used in the more general sense defined under constituent.

Modified from [1].

**compression factor, Z**
compressibility factor
Product of pressure and molar volume divided by the gas constant and thermodynamic temperature. For an ideal gas it is equal to 1.
From [1].

**congruent transition**
Transition in which the two-phase equilibrium of melting, vaporization, or allotropism of a compound involves phases of the same composition.
From [1,12].

**conjugate phases**
Two phases of variable composition in mutual thermodynamic equilibrium.
See also coexistence curve.
Modified from [1].

**conodal**
connodal
See tie line.

**conodal curve**
connodal curve
See coexistence curve.

**constituent**
Chemical species present in a system.
See also component.
Modified from [1].

**critical index, β**
Index in the basic equation describing a coexistence curve

\[
\beta = \lim_{T \to T_c} \left( \frac{\partial \ln (x^" - x')}{\partial \ln |T - T_c|} \right)
\]

where \(x^"\), \(x'\) are the mole fractions of one component in the respective concentrated and dilute phases and \(T_c\) is the critical solution temperature.
critical point
For a pure phase, temperature and pressure (the critical temperature, \( T_c \), and pressure, \( p_c \)) at which intensive properties of liquid and vapor (density, heat capacity, etc.) become equal.

For a mixture, temperature, pressure, and composition (the critical temperature, pressure, and composition) at which the compositions of the liquid and vapor phases, as well as intensive properties of the liquid and vapor phases, become equal.

Note: For a pure phase, the highest temperature (critical temperature) and pressure (critical pressure) at which both a gaseous and a liquid phase can exist.

See also azeotropic point, critical solution point.
Modified from [6].

critical solution point
cloud point
consolute point
plait point (ternary and higher-order systems)
In a binary system, point with coordinates critical solution temperature (CST) or critical solution pressure (CSP) and critical composition on a temperature-composition or pressure-composition phase diagram at which the distinction between coexistent phases disappears.

In ternary and higher-order systems, composition below or above which, on a ternary or higher composition phase diagram at constant temperature and pressure, the distinction between coexistent phases disappears, and tie-lines between coexisting phases become tangential to the coexistence curve.

Note 1: Disappearance of the distinction between phases corresponds to disappearance of a miscibility gap; see mutual solubility.

Note 2: The locus of the plait point composition against temperature is called the plait point curve.

Note 3: In solid–solid, solid–liquid, and liquid–liquid systems, both upper and lower critical solution temperatures (UCST, LCST) or upper and lower critical solution pressures (UCSP, LCSP) can occur. In some systems, both can be observed.

Note 4: In principle, there is no distinction between critical solution points in liquid–liquid or solid–solid systems and critical points in liquid–vapor systems. In binary systems, both types of points are determined by the derivatives of the Gibbs energy with respect to mole fraction

\[
\begin{align*}
\left( \frac{\partial^2 G_m}{\partial x^2} \right)_{T, p} &= 0, \quad \left( \frac{\partial^3 G_m}{\partial x^3} \right)_{T, p} = 0, \\
\left( \frac{\partial^4 G_m}{\partial x^4} \right)_{T, p} &> 0
\end{align*}
\]

See critical point.
Modified from [1,12,13].

crystallization
Formation of a crystalline solid phase from either (a) a solid, liquid, or gaseous mixture or solution, or (b) a pure liquid or gas.

Note: Crystallization usually occurs under laboratory conditions by altering the temperature or pressure of a system, or by evaporation of a solvent.

Modified from [1,12].
data
Experimental results, often numerical.

Note 1: The term is plural; its singular form is datum.

Note 2: Data directly obtained from experimental apparatus may be referred to as raw data. Raw data may be subjected to numerical operations (unit conversions, etc.) to give transformed data, which maintain a point-for-point correspondence with the raw data. However, the use of the term “data” to refer to the results of numerical modeling operations or other elaborate calculations is controversial and discouraged. The latter are better referred to as “results” and, when presented graphically, are usually represented by smooth lines.

density, \( \rho \)
Mass density
Mass of a pure substance, mixture, or solution divided by its volume.
From [2].

dissolution
Process of mixing of two or more phases with the formation of one new homogeneous phase (i.e., the solution).
Modified from [1,12].

dystectic reaction
congruent melting
indifferent melting
Isobaric, reversible melting or dissolution with either complete or partial dissociation on heating of a solid compound, \( \text{A}_a\text{B}_b \), formed by components A and B

\[
\text{A}_a\text{B}_b (s) \rightleftharpoons \text{A}_a\text{B}_b (l) \rightleftharpoons a\text{A}(l) + b\text{B}(l)
\]
where the forward arrows indicate the direction of heating.

Note 1: The dystectic temperature is a maximum \((dT/dx = 0)\) of the melting temperature-composition curve. The dystectic point is the isobarically invariant maximum at the dystectic composition and dystectic temperature, where the compositions of the liquid and solid phases are equal. The composition of a non-stoichiometric compound is also equal to that of the melt at the dystectic temperature and composition.

Note 2: From Greek \(\delta\omega\tau\epsilon\kappa\tau\sigma\omega\gamma\), difficult (or highest) melting.

Note 3: Occasionally, a similar reaction takes place in the subsolidus region. Quite appropriately, it is called a dystectoid reaction.
Examples: systems \(\text{SO}_3 + \text{H}_2\text{O} \) [14], \(\text{Mg} + \text{Sn} \) [15].
From [14,16].

dystonic reaction
Reversible dissolution in an isothermal, isobaric system of three or more components characterized by dissolution and saturation with a stoichiometric compound consisting of two or more of these components. The equilibrium process is, for example,
\[ AB \cdot pH_2O \rightleftharpoons A(aq, \text{sat}) + B(aq, \text{sat}) + pH_2O(l) \]

**Note 1:** The chemical potential of solvent and consequently its partial pressure reaches a maximum value at the dystonic composition or point, which is isothermally and isobarically invariant.

**Note 2:** Dystonic points have been detected in aqueous media only.

**Note 3:** From Greek δύστονος, difficult (highest) tension (vapor pressure).

Examples: systems \( \text{Na}_2\text{SO}_4 + \text{ZnSO}_4 + \text{H}_2\text{O} \), \( \text{Na}_2\text{SO}_4 + \text{CdSO}_4 + \text{H}_2\text{O} \), \( \text{Na}_2\text{SO}_4 + \text{CuSO}_4 + \text{H}_2\text{O} \).

See also eutonic reaction, peritonic reaction.

**equilibrium**

State of a system in which the macroscopic properties of each phase of the system become uniform and independent of time. If the temperature is uniform throughout the system, a state of thermal equilibrium has been reached; if the pressure is uniform, a state of hydrostatic equilibrium has been reached; and if the chemical potential of each component is uniform, a state of chemical equilibrium has been reached. If all these quantities become uniform, the system is said to be in a state of complete thermodynamic equilibrium.

**Note 1:** Complete thermodynamic equilibrium can be expressed in many ways, depending on which variables are of interest. For solubility purposes, the important variables are \( T, p \) and the chemical potentials \( \mu_i \) of the \( C \) components. Equilibrium conditions involving these variables may be obtained by the condition that the variation of the Gibbs energy, expressed in terms of \( T, p \) and the amount of substance \( n_i \) of the \( C \) components, is zero, \( \delta G(T, p, n_i) = 0 \) (i.e., the Gibbs energy is a minimum). If the variation is negative, an irreversible change of the system can occur.

**Note 2:** Solubility equilibrium is an example of a state of complete thermodynamic equilibrium. For example, a saturated solution of a solid in a liquid at a fixed temperature and pressure is in a state of complete thermodynamic equilibrium. If the system is subjected to a small increase in temperature, a small portion of solid will dissolve to restore the equilibrium (if the solubility increases with temperature), while if there is a small decrease in temperature, a small portion of solid will precipitate. This is the basis for determining accurate values of solubility by approaching the equilibrium solubility from both supersaturation and undersaturation directions.

**Note 3:** In [1], chemical equilibrium is defined. The definition here is more inclusive.

Modified and extended from [1].

**eutectic reaction**

Isothermal reversible reaction of a liquid phase \( l \) which is transformed into two (or more) different solid phases \( \alpha \) and \( \beta \) during the cooling of a system. In a binary system,

\[ l \rightleftharpoons \alpha + \beta \]

where \( l \) is a liquid phase, \( \alpha, \beta \) are solid phases, and the forward arrow indicates the direction of cooling. The **equilibrium** reaction occurs along the eutectic line at the eutectic temperature. At the eutectic composition, the compositions of the liquid and solid phases are equal, and intermediate to the compositions of the solid phases of the system.

**Note 1:** The solid phases may be pure phases, solid mixtures, or binary compounds.

Note 2: The eutectic line and composition (hence point) are isobaric invariants of the system, and represent the composition and the minimum melting temperature along the two intersecting melting curves.

Note 3: From Greek ευτηκτος, easy (or lowest) melting.

Note 4: The definition in [1] has been extended, but with eutectic temperature and composition included in the definition.

Example: system Ag + Cu [15] and many salt + water systems.
Modified from [1,12,16].

eutectoid reaction
Isothermal reversible reaction of a solid mixture phase $\gamma$ which is transformed into two (or more) different solid (pure, binary compound or mixture) phases $\alpha$ and $\beta$ during the cooling of a system. In a binary system,

$$\gamma \rightleftharpoons \alpha + \beta$$

where the forward arrow indicates the direction of cooling. The equilibrium reaction occurs along the eutectoid line at the eutectoid temperature. At the eutectoid composition, the compositions of the high-temperature solid phase and the mechanical mixture of the low-temperature solid phases are equal, and intermediate to the compositions of the pure solid phases $\alpha$ and $\beta$. This composition and the eutectoid temperature, which are isobaric invariants of the system, define the eutectoid point.

Note: Derived from Greek and Latin, meaning “resembling a eutectic”.

Example: system Fe + C [15].
See eutectic reaction, Note 2.
Modified from [1,12,16].

eutonic reaction
Isothermally invariant reaction (isothermally and isobarically invariant) reaction of double saturation
Reversible dissolution in a system of three or more components characterized by the composition of a solution that is simultaneously saturated with respect to all (at least two or more) dissolved solutes. The two simultaneous equilibrium processes are, for example,

$$A \cdot p\text{H}_2\text{O} \rightleftharpoons A(\text{aq, sat}) + p\text{H}_2\text{O}(l)$$

$$B \cdot q\text{H}_2\text{O} \rightleftharpoons B(\text{aq, sat}) + q\text{H}_2\text{O}(l)$$

Note 1: The eutonic point is invariant in a ternary system at a given temperature and pressure.

Note 2: The chemical potential of solvent decreases along the two solubility curves that intersect at the eutonic point, and thus reaches a minimum value at that point. Thus, colligative properties (that depend on solvent activity) reach a minimum value at the eutonic point. For example, if the solutes are nonvolatile and the solvent is volatile, the vapor pressure reaches a minimum value.

Note 3: Eutonic points have been detected in aqueous media only.

Note 4: From Greek ευτονος, easy (or lowest) tension (or vapor pressure).

Examples: systems Na$_2$SO$_4$ + ZnSO$_4$ + H$_2$O, Na$_2$SO$_4$ + CdSO$_4$ + H$_2$O, Na$_2$SO$_4$ + CuSO$_4$ + H$_2$O.
Glossary of terms related to solubility

See also eutonic reaction, peritonic reaction.
Introduced in [17].

**fitting equation**
smoothing equation
Theoretically based or empirical equation for interpolation of data over a range of temperature, pressure, composition, or other variable.

*Note:* The use of a fitting equation for extrapolation of data beyond the equation’s known range of validity is a potential source of extreme error.

**fugacity, \( f_B, \tilde{p}_B \)**
For a substance B in a gaseous mixture,

\[
f_B = \lambda_B \lim_{p \to 0} (p_B / \lambda_B)_T
\]

where \( p_B \) is the partial pressure of B and \( \lambda_B \) its absolute activity.
SI unit: Pa.
From [1,2].

**fugacity coefficient, \( \phi \)**
fugacity divided by the partial pressure of a gaseous constituent.
Modified from [1].

**Gibbs–Duhem equation**
Equation relating the intensive variables \( T, p \) and the \( C \) chemical potentials \( \mu_i \) in a phase

\[
SdT - Vdp + \sum_{i=1}^{C} n_i d\mu_i = 0
\]

where \( C \) is the total number of components \( i \) in a phase.

*Note 1:* Note that the variables in this equation are the intensive quantities \( T, p \) and \( \mu_i \).

*Note 2:* The Gibbs–Duhem equation may be written in terms of intensive quantities,

\[
\sum_{i=1}^{C} n_i \left( S_i dT - V_i dp + d\mu_i \right) = 0
\]

where \( S_i, V_i, x_i \) are the respective partial molar entropy, partial molar volume, and mole fraction of component \( i \).

*Note 3:* There is a Gibbs–Duhem equation for each phase in a system exhibiting multiphase equilibria. Application of the conditions for an equilibrium state leads to the phase rule as one example of the application of this equation. When equilibrium conditions are applied, \( T, p \) and \( \mu \) are equal in all phases of an equilibrated system, while \( S_i, V_i \) and \( x_i \) are not.

From [14].

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Gibbs energy of transfer
Change in Gibbs energy at a given temperature and pressure for transfer of a substance between two different phases.

Gibbs–Konovalov equations
van der Waals’ equations
Pair of equations for a binary mixture of components A and B that relate the variables \( T, p \), in one phase of variable composition, \( \alpha \), to the variables \( T, p \), in a coexisting equilibrium phase of variable composition, \( \beta \):

\[
- \left( 1 - x_B^\beta \right) \frac{\Delta^\beta H_A}{T} + x_B^\beta \frac{\Delta^\beta H_B}{T} \quad dT + \left[ \left( 1 - x_B^\beta \right) \frac{\Delta^\beta V_A}{T} + x_B^\beta \frac{\Delta^\beta V_B}{T} \right] \quad dp +
\]

\[
\left( x_B^\alpha - x_B^\beta \right) \left( \frac{\partial^2 G_m}{\partial x_B^2} \right)_{T, p} \quad dx_B^\alpha = 0
\]

\[
\left( 1 - x_B^\alpha \right) \frac{\Delta^\alpha H_A}{T} + x_B^\alpha \frac{\Delta^\alpha H_B}{T} \quad dT + \left[ \left( 1 - x_B^\alpha \right) \frac{\Delta^\alpha V_A}{T} + x_B^\alpha \frac{\Delta^\alpha V_B}{T} \right] \quad dp +
\]

\[
\left( x_B^\beta - x_B^\alpha \right) \left( \frac{\partial^2 G_m}{\partial x_B^2} \right)_{T, p} \quad dx_B^\beta = 0
\]

where \( \Delta^\alpha H_A = H_A^\beta - H_A^\alpha \), \( \Delta^\alpha V_A = V_A^\beta - V_A^\alpha \) are the enthalpy and volume of transfer of component A from phase \( \alpha \) to phase \( \beta \), and similarly for component B.

Note 1: (condition for diffusional stability). This quantity may also be expressed in terms of the derivatives of the chemical potentials, using

\[
\left( \frac{\partial^2 G_m}{\partial x_B^2} \right)_{T, p} = \left\{ \left[ \frac{\partial^2 \mu}{\partial x_B^2} \right]_{T, p} - \frac{1}{x_B} \left( \frac{\partial \mu_A}{\partial x_B} \right)_{T, p} \right\} = \frac{1}{x_A} \left( \frac{\partial \mu_B}{\partial x_B} \right)_{T, p}
\]

Note 2: These equations show that an extremum occurs for each phase equation when the compositions of the two phases are equal, and that the slope of the \( T \)-composition or \( p \)-composition curve is zero for each phase equation at the extremum.

Note 3: Sometimes the German transliteration Konovalow is found.

From [13,14].

Henry’s law
Fugacity (\( \bar{p}_B \)) of a solute (B) in a solution is directly proportional to the activity (\( a_B \)) of the solute

\[
\bar{p}_B = a_{x_B} p^0/K_{x_B} = \gamma_{x_B} x_B^B K_{x_B}
\]

where \( K_{x_B} \) is the solubility coefficient for infinite dilution, i.e., for pure solvent, \( \gamma_{x_B} \) is the activity coefficient (referenced to Henry’s law, mole fraction basis), \( x_B \) is the equilibrium mole fraction of dissolved gas and \( K_{H,B} \) is the Henry’s law constant.
Note 1: Obsolete terminology (rational activity, chemical activity) from [1] has been updated.

Note 2: The solubility constant is the standard equilibrium constant for the equilibrium \( B(g) ⇌ B(\text{satd l}) \), while the Henry’s law constant is the standard equilibrium constant for the equilibrium \( B(\text{satd l}) ⇌ B(g) \). Thus, the two constants are reciprocals of one another, to within a factor of standard pressure. The definition of the solubility constant \( K_{x,B} \) as a dimensionless quantity is

\[
K_{x,B} = \frac{γ_{x,B}ρ^A}{\bar{p}_B} = \exp \left[ \frac{\mu^\circ_B (g,T) - \mu^\circ_{x,B} (1;T,p)}{RT} \right]
\]

Note 3: For the solvent (A), the above relation is called Raoult’s law, and the proportionality factor is the fugacity of the pure solvent, \( \bar{p}_A^* \)

\[
\bar{p}_A = \bar{p}_A^* a_A = f_A x_A
\]

where \( f_A \) is the activity coefficient of A referenced to Raoult’s law at mole fraction \( x_A \).

Note 4: Henry’s law is used often as a limiting law for converting solubility data from the experimental pressure to standard partial pressure, provided the mole fraction of the gas in the liquid is small and that the difference in pressures is small.

See activity coefficient, solubility coefficient, Henry’s law constant, Raoult’s law, Poynting correction.

Modified from [1].

**Henry’s law constant, \( k_H \)**

Henry coefficient

Henry’s law coefficient

Henry constant

For a gas B,

\[
k_{H,B} = \lim_{x_B \to 0} \left( \frac{f_B}{x_B} \right) = \left( \frac{\partial f_B}{x_B} \right)_{x_B \to 0}
\]

where \( f_B \) is the fugacity of B.

Note 1: At low pressures, \( f_B \) becomes equal to the partial pressure \( p_B \).

Note 2: Henry’s law constant in terms of mole fractions has units Pa, but is sometimes expressed in terms of molalities or amount concentrations, with corresponding units Pa kg mol\(^{-1}\), Pa m\(^3\) mol\(^{-1}\), respectively.

Note 3: Henry’s law constant is sometimes given as the reciprocal of the forms defined above, so its exact definition must always be given. It is recommended to use the definition above, and to refer to its reciprocal as a solubility constant.

See also Henry’s law.

From [2,5].

**higher-order system**

multicomponent system

System containing more than three components.

See also unary system, binary system, ternary system.
immiscibility
Inability of two or more substances to form a homogeneous mixture or solution.

incongruent reaction
See peritectic reaction.

independent component
See component.

infinite dilution
Limiting composition or other variable in a solution obtained by extrapolating to a value of zero for the variable describing the composition of the solution.

infinite miscibility
miscibility in all proportions
Property of a system of forming a single phase at all relative proportions of its components.
See also mutual solubility.

initial complex method
See wet residue method.

ionic mole fraction, $x_+$, $x_-$
ionic amount fraction
species mole fraction
species amount fraction

(a) For an ionized salt B in solution,

$$x_{+B} = \frac{v_{+B}x_B}{1 + \sum_{i=1}^{s} (v_i - 1)x_i}, x_{-B} = \frac{v_{-B}x_{+B}}{v_{+B}}$$

where the summation is over all $s$ solute components $i$ and $v_i = v_{+i} + v_{-i}$ is the sum of the stoichiometric numbers of the ions formed from salt $i$.

(b) For a single solvent A in an ionic solution,

$$x'_A = \frac{x_A}{1 + \sum_{i=1}^{s} (v_i - 1)x_i}$$

Note 1: These are generalizations of terms defined in [18], and are used in formulating fitting equations for solubility of salts, in defining activity coefficients on the mole fraction scale, and in discussing salt effects on solubilities of gases.

Note 2: Note that

$$\sum_{i=1}^{s} (x_{+i} - x_{-i}) + x'_A = 1$$

From [5, 18].

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**Ionic strength**
Measure of effective *molality* or *amount concentration* of ionic species in solution:

(a) on molality basis: \( \text{Im} = \frac{1}{2} \sum z_i^2 m_i \)
(b) on amount concentration basis: \( \text{Ic} = \frac{1}{2} \sum z_i^2 c_i \)

where \( z_i \) is the charge number of species \( i \), and the summation is over all ionic species.

From [2].

**Ising model**
Theory of *coexistence curves* or other discontinuities in the properties of *phases* (such as order-disorder or magnetic transitions) in *binary systems*, based on a one-, two-, or three-dimensional statistical mechanical nearest-neighbor lattice theory.

*Note:* The theory predicts phase separation in a wide variety of systems, including binary liquid or solid systems that possess critical solution points.

From [19].

**Isobar**
Line joining points of equal pressure on a *phase diagram*.

**Isopleth**
Line joining points of equal composition on a *phase diagram*.

**Isotherm (solubility)**
Line joining points of equal temperature on a *phase diagram*.

**Isothermal method**
See *synthetic method*.

**Jänecke coordinates**
solute mole (or mass) fraction coordinates
Coordinates used mainly in representation of multicomponent *phase diagrams* that distinguish *solute s* and a single *solvent*. The *mole* or *mass fractions* of the *s* solutes are chosen as the primary composition variables and the mole or mass fraction of solvent as a secondary variable, so that the solute amount (mass) fractions may be represented, for example, in *ternary* and *quaternary systems*, as linear (ternary) or planar triangular or square planar (quaternary) diagrams, with the solvent amount or mass fraction along an axis perpendicular to the chosen diagram of solute variables. Quantitatively,

\[
x_{s,B} = x_B / \sum_{i=1}^{s} x_i
\]

where the summation is over the *s* solute substances. Note that \( \sum_{i=1}^{s} x_{s,B} = 1 \).

*Note 1:* The mole fraction of solvent component \( A \) in a mixture containing \( p - s \) solvents (total species \( p \)) is replaced by a special case of the *solvent mole fraction* \( x_{s,A} = x_{A} / \sum_{i=s+1}^{p} x_i \), which becomes infinite for pure solvent.

Note 2: This quantity can be scaled to a finite value by adding an arbitrary constant to the denominator.

See also solvent mole (mass) fraction.
From [5,20].

Kuenen coefficient, $S$
Volume $V_g$ of an amount $n_B^l$ of a dissolved gas at a given standard temperature $T^o$ (usually 273.15 K) and given standard pressure $p^o$ (1 bar = 0.1 MPa or, in older literature, 1 atm) divided by the mass $m_l$ of the dissolving liquid containing an amount $n_A$ of solvent at temperature $T$ and the given pressure $p^o$.

Note 1: There are two Kuenen coefficients, depending on whether the liquid is the equilibrium solution or the pure liquid, with mathematical definitions:

- Kuenen coefficient, solution reference $S_B = V_g(T^o, p^o, n_B^l)/m_l(T, p^o, n_A, n_B^l)$
- Kuenen coefficient, pure solvent reference $S_B^* = V_g(T^o, p^o, n_B^l)/m_l(T, p^o, n_A)$

Note 2: The relations between the molality $m_B(p^o)$ or mole fraction $x_B(p^o)$ of dissolved gas and the Kuenen coefficients are

$$\frac{1}{x_B(p^o)} = 1 + \frac{1}{m_B(p^o) M_A} = 1 + \frac{RT^o Z_B^*}{p^o M_A S_B^*}$$

where $M_A$ is the molar mass the solvent and $Z_B$ is the compression factor of the gas.

Note 3: The Kuenen coefficient and the related quantities for expression of gas solubility: absorption coefficient, Bunsen coefficient, and Ostwald coefficient appear frequently in the older literature of gas solubility determination. However, the modern practice, recommended here, is to express gas solubility as molality, mole fraction, or mole ratio.

From [5].

lower critical solution temperature
See critical solution point.

mass concentration, $\gamma, \rho$
mass density
Mass of a constituent in a mixture divided by the volume of the mixture.
From [2].

mass density, $\rho$
See density.
mass fraction, $w$
Mass of a particular constituent divided by the sum of the masses of all system constituents.

\[ w_B = \frac{m_B}{\sum_{i=1}^{c} m_i} \]

*Note 1*: $m$ is used here as a symbol for mass, not molality.

*Note 2*: Solubility is often expressed in units ppm, which is equivalent to the SI units mg/kg if the physical quantity involved is mass fraction, or mmol/mol if mole fraction, etc., or even mmol/L, etc., for aqueous solutions where the density of the solution is approximately 1 g/cm$^3$. However, it is recommended that the term “parts per million, ppm” be avoided since, as noted, “parts” may be measured on any one of a mass, amount, or volume basis. If it is used, the physical quantity to which it refers must be specified.

From [2].

mass percent
Deprecated term that includes both a quantity and a unit. Replace by mass fraction expressed as percent.

Examples: Deprecated: the solubility of B is 2.5 mass %. Recommended: the solubility of B is $w_B = 0.025$, or $w_B = 2.5 \%$, or (especially in table headings and labels on axes of plots) 100 $w_B$, with entry 2.5 in table or on axis.

mass ratio, $\zeta$
Mass of one constituent A divided by the mass of a second constituent B in the same system.

\[ \zeta_{A,B} = \frac{m_A}{m_B} \]

mass solubility, $C_w$
Obsolete term for molality of a saturated solution of a gas.

melt
Liquid state of system that is a solid at room temperature.

metastable state
metastability
State of a system in which a perturbation of any one of its defining variables may cause a change to a more stable state.

*Note 1*: A system in a metastable state is in a state of metastable equilibrium, and so can be described consistently by thermodynamic methods.

*Note 2*: Although a driving force for the transition of a metastable state to a stable state exists ($\Delta_{\text{trs}} G < 0$ at constant $T$ and $p$) the transition is retarded, i.e., the transition to the more stable state does not occur significantly during the time of observation.

Modified from [1,11].
metatectic reaction
catatectic reaction
Isothermal reversible reaction of a solid mixture phase \( \beta \) which is transformed into a different solid phase \( \alpha \) plus a liquid phase \( l \) during cooling of a system. For a binary system,

\[ \beta \rightleftharpoons + l \]

where the forward arrow indicates the direction of cooling. The equilibrium transformation occurs along the metatectic line characterized by the metatectic temperature. The metatectic composition and temperature, isobaric invariants of the system, define the metatectic point, which lies between the compositions of liquid \( l \) and solid phase \( \alpha \).

Note 1: From Greek \( \mu \varepsilon \tau \alpha \), referring to position of melting relative to eutectectic and monotectic + -\( \tau \varepsilon \kappa \tau \zeta \), fusible.

Note 2: The alternate name cataratetic reaction, from Greek \( \kappa \alpha \tau \alpha \), down + -\( \tau \varepsilon \kappa \tau \zeta \), fusible has been suggested to emphasize the phenomenon of melting during cooling [20].

Note 3: Occasionally, a similar reaction takes place in the subsolidus region. Quite appropriately, it is called metatectoid reaction.

Example: system Fe + Zr [15].
From [16].

miscibility
Ability of two liquids to undergo spontaneous mixing to form a homogeneous mixture.
See antonym immiscibility.

miscibility gap
See mutual solubility.

mixture
Gaseous, liquid, or solid phase containing more than one substance, when all the substances are treated in the same way.
See also solution.
From [2,13].

molality, \( m, b \)
Amount of a solute divided by the mass of the solvent.
Modified from [1].

mole fraction, \( x \)
amount-of-substance fraction
amount fraction
Amount of substance of a constituent divided by the total amount of all constituents in a mixture.

Note: Usually, \( x \) is used for constituents in solid or liquid phases, \( y \) in gas phases.
From [1].
mole percent
Deprecated term that includes both a quantity and a unit. Replace by mole fraction expressed as percent.
Example: Deprecated: the solubility of B is 2.5 mole %. Recommended: the solubility of B is $x_B = 0.025$, or $x_B = 2.5 \%$, or (especially in table headings and labels on axes of plots) $100x_B$, with entry 2.5 in table or on axis.

mole ratio, $r$
amount ratio
*Amount of one constituent* divided by the amount of a second constituent in the same *system*. For constituents 1 and 2,
$$r_{12} = \frac{n_1}{n_2}$$
From [4].

monotectic reaction
Isothermal reversible reaction of a liquid *phase* to form a solid phase and a second liquid phase during cooling of a *system*. In a *binary system*,
$$l_1 \rightleftharpoons \alpha + l_2$$
where the forward arrow indicates the direction of cooling. The *equilibrium* transformation occurs along the monotectic line, characterized by the monotectic temperature. The monotectic composition and temperature, isobaric invariants of the system, define the monotectic point, which lies between the compositions of liquid $l_2$ and solid phase $\alpha$.

*Note:* From Greek $\muονος$ = one and $τηκτος$ = fusible.
Examples: systems Pb + Zn [16], methanol + cyclohexane [22].
Modified from [1,12,16].

monotectoid reaction
Isothermal reversible reaction of a solid *phase* $\alpha_1$ which is transformed into two different solid phases $\alpha_2$ and $\beta$ during the cooling of a *system*. In a *binary system*,
$$\alpha_1 \rightleftharpoons \alpha_2 + \beta$$
where the forward arrow indicates the direction of cooling. The *equilibrium* process occurs along the monotectoid line at the monotectoid temperature. The monotectoid composition and temperature, isobaric invariants of the system, define the monotectoid point, which lies between the compositions of phases $\beta$ and $\alpha_2$.

*Note 1:* Derived from Greek, meaning “resembling a monotectic”.

*Note 2:* Because the monotectoid temperature is an isobaric invariant, the statement in [1,11], that it is the maximum temperature at which the monotectoid reaction can occur is incorrect.
Example: system Al + Zn [15].
Modified from [1,12,16].

multicomponent system
See higher-order system.
mutual solubility
In a system of two or more liquid or solid components, solubility of all components in all phases.

Note: If mutual solubility is limited over a range of temperature and composition, the liquids or solids are said to exhibit partial miscibility and the system possesses a miscibility gap.

non-saturating solute
Solute which forms an unsaturated solution.

Note: Together with the term saturating solute used to distinguish among solutes in ternary and higher-order systems.

number concentration, C
Number of entities (N) of a constituent in a mixture divided by the volume (V) of the mixture.
See also amount concentration.
From [1].

osmotic coefficient, \( \varphi_x, \varphi_m \)
Factor to correct for non-ideal behavior of the solvent in a solution.
(a) Mole fraction basis. For solvent A,
\[
\varphi_{x,A} = \frac{(\mu_A - \mu_A^*)}{RT} \ln x_A
\]
where \( \mu_A^* \) is the standard chemical potential of A, i.e., the chemical potential of pure liquid A.
(b) Molality basis. For a solvent A in a solution of total molality \( \Sigma m_B \),
\[
\varphi_{m,A} = \frac{(\mu_A - \mu_A^*)}{RTM_A \Sigma m_B}
\]
where \( M_A \) is the molar mass of the solvent.

Note 1: The coefficient \( \varphi_x \) was previously called the rational osmotic coefficient.

Note 2: For a single salt with sum of stoichiometric coefficients of its ions \( v \) and molality \( m_B \),
\[
\varphi_{m,A} = \frac{(\mu_A - \mu_A^*)}{\nu RTM_A m_B}
\]
From [1,2], index B corrected from [1] and sign corrected from [2].

Ostwald coefficient, \( L \)
Volume \( V^g \) of an amount \( n_B \) of a dissolved gas calculated at given temperature \( T \) and pressure \( p \) divided by the volume of the dissolving liquid of volume \( V^l \) and containing an amount \( n_A \) of solvent at the same temperature \( T \) and pressure \( p \).

Note 1: There are two Ostwald coefficients, depending on whether the liquid is the equilibrium solution or the pure liquid, with mathematical definitions:

Ostwald coefficient, solution reference \( L_B = V^g(T, p, n_B^1)/V^l(T, p, n_A, n_B^1) = c_B^1/c_B^g \)

Ostwald coefficient, pure solvent reference \( L_B^* = V^g(T, p, n_B^1)/V^l(T, p, n_A) \)

Note 2: The relations between the molality \( m_B(p) \) or mole fraction \( x_B(p) \) of dissolved gas and the Ostwald coefficients are
\[
\frac{1}{x_B(p)} = 1 + \frac{1}{m_B(p) M_A} = 1 + \frac{RTZ_B}{V_A p_B L_B}
\]

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where \( V_A, V_{m,A} \) are the respective partial molar volume and molar volume of the solvent and \( Z_B \) is the compression factor of the gas.

**Note 3:** A discussion of the Ostwald coefficient from a historical perspective is available [22].

**Note 4:** The Ostwald coefficient and the related quantities for expression of gas solubility, the absorption, Bunsen, and Kuenen coefficients, appear frequently in the older literature of gas solubility determination. However, the modern practice, recommended here, is to express gas solubility as molality, mole fraction, or mole ratio.

From [5].

**Ostwald ripening**

Dissolution of small crystals or sol particles and the redeposition of the dissolved species on the surfaces of larger crystals or sol particles.

**Note:** The process occurs because smaller particles have a higher surface energy, hence higher total Gibbs energy, than larger particles, giving rise to an apparent higher solubility.

From [24] as modified from [1].

**partial miscibility**

See mutual solubility.

**partial pressure, \( p_B \)**

In a mixture of ideal gases, \( p_B = y_B p \), where \( y_B \) is the mole fraction of constituent B and \( p \) is the total pressure.

**Note:** In real (non-ideal) gases, there is a difficulty defining partial pressure [1,2].

**percent, %**

One part in a hundred parts.

Example: The mole fraction \( x = 2.8 \times 10^{-2} = 2.8 \% \).

From [1].

**peritectic reaction**

incongruent reaction

Isothermal, reversible reaction between two phases, a liquid and a solid, that results, on cooling of a binary, ternary, or higher-order system in one, two, … \((n - 1, \text{ where } n \text{ is the number of components})\) new solid phases. For example, in a binary system

\[
1 + \alpha \rightleftharpoons \beta
\]

where the forward arrow indicates the direction of cooling. The equilibrium process occurs along the peritectic line, characterized by the peritectic temperature. The peritectic composition and temperature, isobaric invariants of the system, define the peritectic point, which lies between the compositions of phases 1 and \( \alpha \).
Note: From Greek περι-, around and τηκτος, fusible.
Example: alloy system Cu + Zn [15], salt-water system Na₂SO₄ + H₂O.
Modified from [1,12,16].

peritectoid reaction
Isothermal, reversible reaction in the solid state, that, on cooling of a binary, ternary, or higher-order system, results in one, two, ... (n – 1) new solid phases. For a binary system.

\[ \alpha + \gamma \rightleftharpoons \beta \]

where the forward arrow indicates the direction of cooling. The equilibrium process occurs at the peritectoid point, characterized by the peritectoid line at the peritectoid temperature. The peritectoid composition and temperature, isobaric invariants of the system, define the peritectoid point, which lies between the compositions of phases \( \alpha \) and \( \beta \).

Note 1: Derived from Greek, meaning “resembling a peritectic”.

Note 2: Metatectoid, meaning “resembling a metatectic”, is not an acceptable synonym for peritectoid.

Example: systems Al + Cu [15], hexacosane (C₂₆H₅₄) + octacosane (C₂₈H₅₈) [25].
Modified from [1,12,16].

peritonic reaction
transition reaction (in phase equilibria)
Isothermal, isobaric reversible reaction between two phases, a saturated liquid and a solid, that results, on removal of the solvent component of a ternary system in one new solid phase. For example,

\[ l + \alpha \rightleftharpoons \beta \]

where the forward arrow indicates the direction of removal of the solvent component. The system is in equilibrium along the peritonic line, on which is found the peritonic composition or point, an isothermal, isobaric invariant of the system which lies between the compositions of phases \( l \) and \( \alpha \).

Example: In an ionic system,

\[ AX·pH₂O(s) + B^{2+}(aq, sat) + (q−p)H₂O(l) \rightleftharpoons BX·qH₂O(s) + A^{z+}(aq, sat) \]

The two salts may also contain the same ions.

Note 1: Named from a combination of eutonic and peritectic.

Note 2: The very general terms transition point or composition are deprecated, as they can refer to many other types of equilibria. Historically, the use of transition point in this case dates back to at least ref. [17].

Example: system KCl + MgCl₂ + H₂O
See also dystonic reaction, eutonic reaction, transition point.

permil, ‰
per mille
permille
promille
One part in a thousand parts.
Example: The mole fraction \( x = 2.8 \times 10^{-3} = 2.8 \text{‰} \).
From [1].
See also chlorinity, salinity.

**phase**
System or portion of a system which is uniform in chemical composition and physical state.

*Note:* At equilibrium, all intensive variables (temperature, pressure, electric field, magnetic induction, chemical potential, etc.) are uniform within a phase.

Modified from [1,12].

**phase diagram**
Graphical representation, by use of points, lines, and surfaces, of the phases present at chemical equilibrium in unary, binary, ternary, and higher-order systems containing two or more phases.

*Note:* Phase diagrams may employ any pairs of the variables temperature, pressure, and compositions of various phases. Thus, temperature-composition and pressure-composition phase diagrams are possible, as well as diagrams showing only compositions under isothermal or isobaric conditions.

See also isobar, isopleths, isotherm.

**phase rule**
Gibbs’ phase rule
Relation connecting number of possible stable phases \( P \) in an equilibrium system with the number of components, \( C \), and the number of degrees of freedom, \( F \), i.e., the number of variables that can be assigned free values:

\[
F = C + 2 - P
\]

*Note 1:* The phase rule as stated holds under the conditions: (a) negligible surface contributions (unless the curvature is constant); (b) uniform normal pressure (perpendicular to the phase surface) over all phases is the only external force; (c) interphase surfaces are deformable, heat-conducting, and permeable to all components.

*Note 2:* Sometimes, the phase rule is written as \( F = (C - r) + 2 - P \) where \( r \) is the number of chemical reactions which can reach chemical equilibrium in the system. This form requires that \( C \) be the number of species assumed for the system, rather than the number of components.

*Note 3:* For the phase rule in the presence of surface phases, see [26].

Modified from [1,12,26].

**plait point**
See critical solution point.

**polythermal method**
See synthetic method.

Poynting correction, $I_p$
Factor used to convert experimentally determined Henry's law constants at the saturation vapor pressure of the solvent A, $p^*_A$, to a reference pressure $p$

$$I_p(p) = \int \frac{V_B^{\infty}}{p^*_A RT} dp$$

$$k_{H,B}(p) = k_{H,B} \left( p^*_A \right) \exp \left[ I_p(p) \right]$$

where $V_B^{\infty}$ is the partial molar volume of liquid B at infinite dilution in solvent A. The second equation gives the relation between the two Henry's law constants.

See Henry’s law constant.
From [5].

precipitation
Formation of a solid material (a precipitate) from a liquid solution in which the material is present in amounts greater than its solubility in the liquid.

Note: The solid material may eventually sediment due to the action of gravity.
Modified from [1,12].

primary data
Data reported in peer-reviewed scientific reports of original research which allow an assessment of data quality.

Note: Primary data are distinguished from secondary data appearing in reviews, handbooks, compendia, etc. which hold the possibility of error and bias through transcription error, incomplete coverage of the primary literature, etc.

Raoult’s Law
Fugacity ($\tilde{p}_B$) of a gaseous component (B) in equilibrium with a liquid or solid mixture containing B is directly proportional to the activity ($a_B$) of the component in the mixture

$$\tilde{p}_B = \tilde{p}_B^{\infty} f_A^B x_B$$

where $f_A^B$ is the activity coefficient of A referenced to Raoult’s law at mole fraction $x_A$ and $\tilde{p}_B^{\infty}$ is the fugacity of pure B.

Note: An ideal mixture is defined by replacing fugacities with partial pressures and setting the activity coefficient equal to unity

$$p_B = p_B^{\infty} x_B = y_B p$$

The total pressure, $p$, is then simply the sum of the partial pressures of all components.
See activity coefficient, Henry’s law.

reciprocal salt system
System containing $n$ ionic species with or without a single solvent. Because of electrical neutrality, the system has $n$ components in an aqueous system with a single solvent and $n - 1$ components in a molten salt system.

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Note 1: There is a special notation for reciprocal salt systems that involves use of double vertical bars. All ionic species are specified. For example,

$$K^+, Na^+ \parallel Cl^-, NO_3^- + H_2O$$

denotes a four-component (quaternary) aqueous reciprocal salt system in which each cation can combine with each anion to produce anhydrous, hydrated, binary, or ternary solid components; i.e., the solution may become saturated with respect to any one of KCl, NaCl, KNO3, NaNO3, their hydrates, or anhydrous or hydrated binary compounds of the simple salts.

Note 2: Ternary aqueous systems containing a common ion can be considered as a limiting case of this class of system, where there is only one cation or one anion. For example,

$$Na^+, Zn^{2+} \parallel SO_4^{2-} + H_2O$$

reduced pressure, $p_r$
Pressure divided by the critical pressure.

reduced temperature, $T_r$
Temperature divided by the critical temperature.

reference data
Data that are critically evaluated and verified, obtained from an identified source, and related to a property of a phenomenon, body, or substance, or a system of components of known composition and structure.

relative activity
See activity.

retrograde solubility
(1) Solubility that decreases with an increase in conditions such as temperature or added component, where an increase is expected as the usual case.

(2) In ternary liquid systems with coexisting phases, where the critical solution point is not a maximum on a ternary diagram, passing from a homogeneous mixture in which the composition of a given component is greater than that of the critical mixture to a heterogeneous mixture and back to a homogeneous mixture by altering the relative amounts of the two other components.
From [27].

salinity, $S$
Mass of dissolved salts in sea water, brackish water, brine, or other saline solution divided by the mass of the solution.

Note 1: In practice, this quantity cannot be measured directly in sea water or other natural waters because of the difficulty of drying the salts from these waters. Salinity is usually calculated from another property (e.g., chlorinity, electrical conductivity) whose relationship to salinity is well known.

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In oceanography, where precise and reproducible determination of sea water density is of interest, practical salinity, $S$, is defined in terms of the ratio $k_{15}$ of the electrical conductivity of the sea water sample at 15 °C and 1 atm (1.01325 kPa) to that of a potassium chloride (KCl) solution, in which the mass fraction of KCl is $32.453 \times 10^{-3}$ at the same temperature and pressure.

The Practical Salinity Scale, established in 1978, relates $S$ to $k_{15}$ by the equation

$$S/\text{g kg}^{-1} = \sum_{i=0}^{4} a_i k_{15}^{i/2} \text{ range } 2 \leq S \leq 42$$

where $a_0 = 0.0080$, $a_1 = -0.1692$, $a_2 = 25.3851$, $a_3 = 14.0941$, $a_4 = 7.0261$, $a_5 = 2.7081$, $\sum_{i=0}^{5} a_i = 35.0000$.

Especially before 1978, practical salinity was also determined from chlorinity, $w_{\text{Cl}}$ as $S = 1.80645 w_{\text{Cl}}$.

Before 1978, the usual unit for $S$ expressed as a mass fraction was permil (no longer recommended), now usually in units g kg$^{-1}$ or expressed as g kg$^{-1}$ without the units given explicitly.

Example: the practical salinity of a sample of artificial sea water is $S = 35.0000 \%$ or 35.0000 g kg$^{-1}$ or 35.0000 $\times 10^{-3}$ or 35.0000.

See also chlorinity.

From [7,8].

**Salt effect** (in solubility)

Change in solubility of a solute in aqueous solution on addition of a salt that does not possess a common ion with the original solute.

*Note:* If the solubility increases on addition of a salt, the addition is said to cause salting-in; if the opposite, it causes salting-out.

See also Sechenov equation.

**Saturated solution**

Solution that has the same composition of a solute as one that is in equilibrium with undissolved solute at specified values of temperature and pressure.

*Note:* Ternary and higher-order systems can be saturated with respect to one component while being unsaturated with respect to another.

Modified from [1].

**Saturating solute**

Solute that forms a saturated solution.

*Note:* Together with the term non-saturating solute, used to distinguish among solutes in ternary and higher-order systems.

**Saturation**

State of a saturated solution.

Modified from [1].

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Glossary of terms related to solubility

saturation vapor pressure
Pressure exerted by a pure substance (at a given temperature) in a system containing only the vapor and condensed phase (liquid or solid) of the substance.
From [1].

scaling equation
Dimensionless equation representing mutual solubilities in a number of chemically related liquid systems. Each mutual solubility is subtracted from the solubility at the critical solution point, and the temperature is divided by the critical solution temperature.

Schreinemakers’ method
See wet residue method.

Sechenov equation
Expression of the salt effect that relates the change in solubility of a nonelectrolyte (e.g., gas or organic liquid) to changing ionic strength of aqueous solutions,

\[ \lg\left(\frac{s_o}{s}\right) = K_s I_s \]

where \( s_o, s \) are the solubilities of nonelectrolyte in pure water and saline solution, respectively, \( K_s \) the Sechenov parameter (an empirical proportionality constant), and \( I_s \) the ionic strength of the saline solution.

Note 1: Positive values of the Sechenov parameter correspond to the commonly observed salting-out effect; negative values to the less common salting-in effect.

Note 2: Sechenov is the international and Setschenow the German transliteration of Сеченов.
From [5,28].

Sieverts’ law
Solubility of a diatomic gas in a molten metal is proportional to the square root of the partial pressure.

Note 1: The law follows from the solubility constant for equilibrium dissolution of an ideal diatomic gas \( X_2 \) in a metal \( M \) to form an ideal solution of dissociated atoms:

\[ X_2(g) \rightleftharpoons 2X(M) \]

for which, and for dilute solutions of the gas,

\[ K_{s,x} = \frac{a(X,M)^2}{a(X_2,g)} = x(X,M)^2 p^{1/2} p(X_2,g) \]

where \( x(X, M) \) is the solubility of the gas in the molten metal expressed as a mole fraction of H-atoms. The solubility constant \( K_{s,x} \) is called the Sieverts constant.

Note 2: Other forms of the solubility constant are common, especially in terms of amount concentrations or mass fractions (especially in metallurgical papers) instead of mole fractions.

See also Henry’s law constant, solubility constant.
Expanded from [29].

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smoothing equation
See fitting equation.

solid mixture
See mixture.

solid solution
See solution.

solubility, \( s \)
Analytical composition of a mixture or solution which is saturated with one of the components of the mixture or solution, expressed in terms of the proportion of the designated component in the designated mixture or solution.

Note 1: The definition refers to constituents \( B \); i.e., \( s_B \).

Note 2: Solubility may be expressed in any units corresponding to quantities that denote relative composition, such as mass, number or amount concentration, molality, mass fraction, mole fraction, mole ratio, etc.

Note 3: The mixture or solution may involve any physical state: solid, liquid, gas, vapor, supercritical fluid.

Note 4: The term “solubility” is also often used in a more general sense to refer to processes and phenomena related to dissolution.

Modified from [1,5].

solubility constant
Generalization of solubility product to include, for example, undissociated salt or uncharged cation-anion pair as solute, reaction between solution species and the cation or anion of the solid salt, and gas-solution equilibria.

Note 1: The solubility constant can also be considered in many cases as relating to a solubility equilibrium as described under solubility product plus one or more simultaneous homogeneous equilibria in solution.

Note 2: Examples and specialized notation for solubility constants are described in [30–34].

See also solubility product.

solubility parameter, \( \delta \)
Parameter used in predicting the solubility of non-electrolytes (including polymers) in a given solvent. For a substance \( B \),

\[
\delta_B = \left( \frac{\Delta_{\text{vap}} E_{m,B}}{V_{m,B}} \right)^{1/2}
\]

where \( \Delta_{\text{vap}} E_m \) is the molar energy of vaporization at zero pressure and \( V_m \) is the molar volume.

Note 1: For a substance of low molecular weight, the value of the solubility parameter can be estimated most reliably from the enthalpy of vaporization and the molar volume.
Note 2: The solubility of a substance B can be related to the square of the difference between the solubility parameters for supercooled liquid B and solvent at a given temperature, with appropriate allowances for entropy of mixing. Thus, a value can be estimated from the solubility of the solid in a series of solvents of known solubility parameter. For a polymer, it is usually taken to be the value of the solubility parameter of the solvent producing the solution with maximum intrinsic viscosity or maximum swelling of a network of the polymer. See [35] for the original definition, theory, and extensive examples.

Note 3: The SI units are Pa\(^{1/2}\) = J\(^{1/2}\) m\(^{-3/2}\), but units used frequently are (\(\mu\)Pa)\(^{1/2}\) = (J cm\(^{-3}\))\(^{1/2}\), where 1 (J cm\(^{-3}\))\(^{1/2}\) ≈ 2.045 (cal cm\(^{-3}\))\(^{1/2}\). The unit calorie is discouraged as obsolete.

Modified from [1,35].

**solubility product,** \(K_s\), \(K_s^o\)

Equilibrium constant for the dissolution process of an ionic solid or an addition compound that dissociates completely in solution; for example,

\[
M_{\nu_+} X_{\nu_-} \cdot r\text{H}_2\text{O}(s) \rightleftharpoons \nu_+ M^{2+} (\text{aq,sat}) + \nu_- X^{2-} (\text{aq,sat}) + r\text{H}_2\text{O} (l)
\]

where a hydrated ionic solid (mole ratio water/salt = \(r\)) has been shown as a common example of an addition compound one part of which is an ionic solid. The equilibrium constant is

\[
K_s = a^+ \nu_+ a^- \nu_- a_w^r
\]

or

\[
K_s = (\gamma^\pm m_o^\nu)^\nu (m^\nu)^\nu \exp(-rM_w \phi_m \Sigma \iota)
\]

where \(a^+, a^-,\) and \(a_w\) are the respective cation, anion and water activities, \(\nu = \nu_+ + \nu_-\) is the sum of the stoichiometric numbers and \(z^+; z^-\) the charge numbers of the cation and anion, \(m\) molality, \(\gamma^\pm\) the mean ionic activity coefficient (referenced to molality), \(m_o^\nu = 1 \text{ mol kg}^{-1}\) the standard molality, \(M_w\) the molar mass of water and \(\phi_m\) the osmotic coefficient (referenced to molality). The summation is over all species in solution.

Note 1: For the special case where there is no common cation or anion and no reaction with solution species and the cation or anion of the salt to form, e.g., an acid, base, or complex ion in the solution

\[
K_s = (\nu_+^\nu_+ m_o^\nu \nu_-^\nu_- \exp(-rM_w \phi_m))
\]

where \(\nu = (\nu_+^\nu_+ \nu_-^\nu_-)^{1/2}\) and \(m\) is the stoichiometric molality of the anhydrous salt, with \(m^\nu = \nu_+^\nu m, m^- = \nu_-^\nu m\).

Note 2: For the special case where a common anion, molality \(m_X^\nu\), with balancing cations of molality \(m_N^\nu\), exist in solution, but there is no reaction with solution species and the cation or anion of the salt

\[
K_s = (\gamma^\pm m_o^\nu)^\nu (\nu m^\nu + m_N^\nu) \exp[-rM_w (nm + m_N + m_X^\nu) \phi_m]
\]

Note 3: Solubility products have been written here as dimensionless quantities, a practice that is not always followed but avoids extra symbols to make the solubility constant dimensionless when its use as the argument of a logarithm is required.

See also solubility constant.

Modified from [1].

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**solubilization**
Process by which an agent increases the solubility or the rate of dissolution of a solid or liquid solute.

*Note 1:* The solute that is solubilized is called the solubilizate.

*Note 2:* Solubilization is sometimes used loosely to include processes caused by agents such as surfactants which hold a component in micelles or colloidal suspension rather than in solution.

Modified from entry “micellar solubilization” in [1].

**solute**
Minor component of a solution that is regarded as having been dissolved by the solvent.
Modified from [1].

**solute mole (or mass) fraction**
See Jänecke coordinates.

**solution**
Liquid or solid phase containing more than one substance, when, for convenience, one (or more) of the substances, called the solvent, is treated differently from the other substances, called solutes.

*Note 1:* When the sum of the mole fractions of the solutes is small compared with unity, the solution is called a dilute solution.

*Note 2:* A superscript $\infty$ attached to the symbol for a property of a solution denotes the property in the limit of infinite dilution.

See mixture.
From [2,13].

**solvent**
Major component of a solution that is regarded as having dissolved the solute.
Modified from [1].

**solvent mole (mass) fraction**, $x_{v,A}$ ($w_{v,A}$)
Mole or mass fraction of a solvent in a solution containing $s$ solute constituents ($i = 1, 2, \ldots, s$) and $p$ solvent constituents ($i = s + 1, s + 2, \ldots, p$). For a solvent A,

$$x_{v,A} = x_A / \sum_{i=s+1}^{p} x_i,$$

where $\sum_{i=1}^{s} x_i = \sum_{i=1}^{s} x_i + \sum_{i=s+1}^{p} x_i = 1$

and analogously for $w_{v,A}$, where the first and last summations are over all solvent constituents.
See also Jänecke coordinates.
From [5].

**solvent volume fraction**, $\varphi_v$
For a solvent A, volume fraction of A divided by the sum of the volume fractions of solvent in a solution containing both $s$ solutes ($i = 1, 2, \ldots, s$) and $p$ solvents ($i = s + 1, s + 2, \ldots, p$)

\[ \varphi_{\nu,A} = \frac{\varphi_A}{\sum_{i=1}^{p} \varphi_i} \]

where the summation is over the \( p \) solvent components only.

Note: Used to express solvent composition for a system containing a solute in a mixed solvent.

See also volume fraction.
From [5].

solvomolality
See aquamolality.

species mole (amount) fraction
See ionic mole (amount) fraction.

spinodal curve
Boundary of separation between metastable and unstable phases on a temperature-composition or pressure-composition phase diagram for a liquid or solid system of two or more components.
See spinodal decomposition.
From [10,11].

spinodal decomposition
Mechanism by which a system (solid or liquid) consisting of two or more components in a metastable phase transforms into two stable phases.

Note 1: On a temperature-composition phase diagram for a liquid or solid system of two or more components, systems with compositions lying within the areas between the spinodal and conodal curves can undergo spinodal decomposition.

Note 2: The mechanism is considered to involve a clustering reaction in which the mixture separates spontaneously into two phases, starting with small fluctuations and proceeding with a decrease in the Gibbs energy without a nucleation barrier.

See conodal curve, spinodal curve.
Modified from [1,12].

standard amount concentration, \( c^o \)
standard concentration
Chosen value of amount concentration.

Note: In principle, one may choose any value for the standard concentration, although the choice must be specified. The most common choice for standard concentration is \( c^o = 1 \text{ mol dm}^{-3} \), which is universally accepted.

Modified from [1,2].

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standard molality, \(m^o\), \(m^0\), \(b^o\), \(b^0\)

Chosen value of molality.

Note: In principle, one may choose any value for the standard molality, although the choice must be specified. The most common choice for standard molality is \(m^0 = 1 \text{ mol kg}^{-1}\), which is universally accepted.

Modified from [1,2].

standard pressure, \(p^o\), \(p^0\)

Chosen value of pressure.

Note: In principle, one may choose any value for the standard pressure, although the choice must be specified. In practice, the most common choice is \(p^0 = 0.1 \text{ MPa} = 100 \text{ kPa} (= 1 \text{ bar})\). The value for \(p^0 = 100 \text{ kPa}\) is the IUPAC recommendation since 1982, and is recommended for tabulating thermodynamic data. Prior to 1982, the standard pressure was usually taken to be \(p^0 = 101325 \text{ Pa} (= 1 \text{ atm}, \text{called the standard atmosphere})\). In any case, the value for \(p^0\) should be specified. The conversion of values corresponding to different \(p^0\) is described in [36–38]. The newer value of \(p^0\), 100 kPa, is sometimes called the standard state pressure.

Modified from [1,2].

standard state

State of a system chosen as standard for reference by convention. Three standard states are recognized:

For a gas phase, it is the (hypothetical) state of the pure substance in the gaseous phase at the standard pressure \(p = p^0\), assuming ideal behavior.

For a pure phase, or a mixture, or a solvent in the liquid or solid state, it is the state of the pure substance in the liquid or solid phase at the standard pressure \(p = p^0\).

For a solute in solution, it is the (hypothetical) state of solute at the standard molality \(m^0\), standard pressure \(p^0\), or standard concentration \(c^0\) and exhibiting infinitely dilute solution behavior.

Note: Either superscript, \(\phi\) or \(\circ\), is acceptable to designate standard state.

From [1].

supercritical fluid

State of a compound, mixture of fixed overall composition, or element above its critical pressure \((p_c)\) and critical temperature \((T_c)\).

From [1].

supersaturated solution

Solution that has a greater composition of a solute than one that is in equilibrium with undissolved solute at specified values of temperature and pressure.

See also saturated solution, unsaturated solution.

supersaturation

State of a supersaturated solution.

Modified from [1].

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syntectic reaction
Isothermal reversible reaction of two liquid phases \( l_1, l_2 \) which are transformed into a solid phase \( \alpha \) during the cooling of a system. For a binary system,
\[
l_1 + l_2 \rightleftharpoons \alpha
\]
where the forward arrow indicates the direction of cooling. The equilibrium reaction occurs along the syntectic line, characterized by the syntectic temperature. The syntectic composition and temperature, isobaric invariants of the system, define the syntectic point, which lies between the composition of the two liquid phases.

Note 1: From Greek: συν- = together and -τηκτος = fusible.

Note 2: Occasionally, a similar reaction takes place in the subsolidus region. Quite appropriately, it is called syntectoid reaction.

Note 3: Because the syntectic temperature is an isobaric invariant, the statement in [1,12] that it is the maximum temperature at which the syntectic reaction can occur is incorrect.

Example: system K + Zn [15].
Modified and corrected from [1,12,16].

synthetic method (in determination of solubility)
Class of experimental procedures for solubility determination in which a solution of known composition is prepared. Two major subclasses are recognized. In the isothermal method, one liquid component is titrated with a second liquid component at constant temperature until persistent turbidity is observed (i.e., to the cloud point). In the polythermal method, a mixture of known composition is heated above its solution temperature and monitored visually during cooling until turbidity (the cloud point) is observed.

system
Arbitrarily defined part of the universe, regardless of form or size.

Note: In the context of solubility phenomena, a system contains two or more components whose solubility is of interest.

Modified from [1].

ternary diagram
Gibbs’ triangular representation
Triangular plot (usually an equilateral or right-angled triangle) whose vertices represent the pure components of a ternary system at constant temperature and pressure. Any point within the triangle represents the composition in terms of the two independent mole (or mass) fractions. Lines parallel to the sides of the triangle represent constant proportions of the respective components (isopleths).

Note 1: In an isobaric diagram, an axis perpendicular to the plane of the triangle may be added to represent temperature.

Note 2: In addition, any triangle of any shape, including any triangle within the main triangle, can be used to obtain the composition of the system represented by a point within the triangle in terms of the compositions represented by its vertices.

From [39].

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ternary system
System containing three components.
See also unary system, binary system, ternary system.

thermodynamic component
See component.

tie-line
connodal conodal
Straight line connecting compositions of independent coexisting phases in equilibrium on a phase diagram.
From [1,12].

transition point
See peritonic reaction.

unary system
System containing one component.
See also binary system, ternary system, higher-order system.

unsaturated solution
Solution that has a lower proportion of a solute than one that is in equilibrium with undissolved solute at specified values of temperature and pressure.

unsaturation
undersaturation
State of an unsaturated solution.

upper critical solution temperature
See critical solution point.

Van der Waals’ equations
See Gibbs–Konovalov equations.

volume fraction, \( \phi \)
Volume of a constituent of a mixture divided by the sum of volumes of all constituents prior to mixing. For a substance B,
\[
\phi_B = x_B V_{m,B} / \sum_{j=1}^{C} x_j V_{m,j}^a
\]
where \( V_{m,j}^a \) is the molar volume of the pure constituent \( j \).
See also solvent volume fraction.
Extended from [1].
wet residue method
Schreinemakers’ method
initial complex method
Method for determining composition of a solid phase in a ternary system at constant temperature and pressure by analysis of the total mixture of solid and liquid in equilibrium.

Note 1: The method relies on the fact that the wet residue lies on the tie-line connecting the solid phase and the saturated liquid phase. Equally well, the initial mixture of solid and liquid used in the solubility experiment can be used, when it is referred to as the “initial complex” method. In either case, two or more tie-lines intersect at the composition of the solid phase.

Note 2: The method is generally more reliable than isolation and analysis of the solid phase, especially when the solid phase is a hydrate.

From [40,41].

3. REFERENCES


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4. QUANTITIES, SYMBOLS, AND UNITS USED IN THIS GLOSSARY

4.1 Quantities, symbols, and units

Entries in the table are consistent with terminology, symbols, and units given in [2,5].

Note that use of a + sign as a separator between formulas or names of components of mixtures is contrary to use in [12], where a hyphen is used. The use here accords with the most prevalent current use in research journals in thermodynamics.
<table>
<thead>
<tr>
<th>Name</th>
<th>Symbol</th>
<th>Definition</th>
<th>SI Unit</th>
</tr>
</thead>
<tbody>
<tr>
<td>absolute activity</td>
<td>( \gamma )</td>
<td>( \gamma_B = \exp(\mu_B/RT) )</td>
<td>1</td>
</tr>
<tr>
<td>absorption coefficient (in gas solubility)</td>
<td>( \beta^* )</td>
<td>( \beta_B^* = V B(T, p_A + p_B = p^o_B, n_B^1)/V(A, p^o, n_A) )</td>
<td>1</td>
</tr>
<tr>
<td>activity, relative activity</td>
<td>( a_x )</td>
<td>( RT \ln a_x B = \mu_B (T, p, x) \mu_B^o (T, p) )</td>
<td>1</td>
</tr>
<tr>
<td>activity coefficient, ref. Henry, amount conc. basis</td>
<td>( \gamma_c )</td>
<td>( RT \ln(\gamma_c c_B/c_B^o) = RT \ln a_B = \mu_B - \mu_B^o ) ( \mu_B^o = \lim_{c_B \to 0} [\mu_B - RT \ln(c_B/c_B^o)] )</td>
<td>1</td>
</tr>
<tr>
<td>activity coefficient, ref. Henry, molality basis</td>
<td>( \gamma_m )</td>
<td>( RT \ln(\gamma_m m_B) = RT \ln a_B = \mu_B - \mu_m(B) ) ( \mu_m(B) = \lim_{m_B \to 0} [\mu_B - RT \ln(m_B/m_o^B)] )</td>
<td>1</td>
</tr>
<tr>
<td>activity coefficient, ref. Henry, mole fraction basis</td>
<td>( \gamma_x )</td>
<td>( RT \ln(\gamma_x x_B) = RT \ln a_B = \mu_B - \mu_x(B) ) ( \mu_x(B) = \lim_{x_B \to 0} [\mu_B - RT \ln(x_B/x_o^B)] )</td>
<td>1</td>
</tr>
<tr>
<td>activity coefficient, ref. Raoult, mole fraction basis</td>
<td>( f )</td>
<td>( RT \ln(f_B x_B) = RT \ln a_B = \mu_B - \mu_B^* ) ( \mu_B^* : \text{standard chemical potential of pure B} )</td>
<td>1</td>
</tr>
<tr>
<td>activity coefficient, infinite dilution</td>
<td>( f_\infty )</td>
<td>( \ln f_\infty = \lim_{x_B \to 0} \left( \frac{\mu_B - \mu_B^*}{RT} - \ln x_B \right) )</td>
<td>1</td>
</tr>
<tr>
<td>amount concentration</td>
<td>( c, \text{[species]} )</td>
<td>( c_B = \text{[species B]} = n_B/V )</td>
<td>( \text{mol m}^{-3} )</td>
</tr>
<tr>
<td>amount of substance</td>
<td>( n )</td>
<td>(SI base unit)</td>
<td>( \text{mol} )</td>
</tr>
<tr>
<td>aquamolality</td>
<td>( m_\text{(C)} )</td>
<td>( m_\text{(C)}^B = m_B M/C )</td>
<td>( \text{mol m}^{-3} )</td>
</tr>
<tr>
<td>Bunsen coefficient, pure solvent basis</td>
<td>( \alpha^* )</td>
<td>( \alpha^*_B = V B(T^o, p^o, n_B^1)/V(A, p^o, n_A) )</td>
<td>1</td>
</tr>
<tr>
<td>Bunsen coefficient, solution basis</td>
<td>( \alpha )</td>
<td>( \alpha_B = V B(T^o, p^o, n_B^1)/V(A, p^o, n_A, n_B^1) )</td>
<td>1</td>
</tr>
<tr>
<td>chemical potential</td>
<td>( \mu )</td>
<td>( \mu_B = (\partial G/\partial n_B)_{T, p, n_i \neq n_B} )</td>
<td>( \text{J mol}^{-1} )</td>
</tr>
<tr>
<td>chlorinity</td>
<td>( w_{Cl} )</td>
<td>(see main entry)</td>
<td>1</td>
</tr>
<tr>
<td>compression (compressibility) factor</td>
<td>( Z )</td>
<td>( Z = pV_m/RT )</td>
<td>1</td>
</tr>
<tr>
<td>density, mass density</td>
<td>( \rho )</td>
<td>( \rho = \text{(total mass)}/V )</td>
<td>( \text{kg m}^{-3} )</td>
</tr>
<tr>
<td>energy of vaporization, molar</td>
<td>( \Delta_v E_m )</td>
<td>( \Delta_v E_m = E_m^g - E_m^l )</td>
<td>( \text{J mol}^{-1} )</td>
</tr>
<tr>
<td>enthalpy, molar</td>
<td>( H_m )</td>
<td>( H/n )</td>
<td>( \text{J mol}^{-1} )</td>
</tr>
<tr>
<td>enthalpy, partial molar</td>
<td>( H_B )</td>
<td>( H_B = (\partial H/\partial n_B)_{T, p, n_i \neq n_B} )</td>
<td>( \text{J mol}^{-1} )</td>
</tr>
</tbody>
</table>

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<table>
<thead>
<tr>
<th>Name</th>
<th>Symbol</th>
<th>Definition</th>
<th>SI Unit</th>
</tr>
</thead>
<tbody>
<tr>
<td>entropy, molar</td>
<td>$S_m$</td>
<td>$S/n$</td>
<td>J K⁻¹ mol⁻¹</td>
</tr>
<tr>
<td>entropy, partial molar</td>
<td>$S_B$</td>
<td>$S_B = (\partial S/\partial n_B)_{T,p,n_i \neq n_B}$</td>
<td>J K⁻¹ mol⁻¹</td>
</tr>
<tr>
<td>fugacity</td>
<td>$f, \bar{p}_B$</td>
<td>$f_B = \lambda_B \lim_{p \to 0} (p_B/\lambda_B)_T$</td>
<td>1</td>
</tr>
<tr>
<td>gas constant</td>
<td>$R$</td>
<td>$R = 8.314 472(15)$</td>
<td>J K⁻¹ mol⁻¹</td>
</tr>
<tr>
<td>Gibbs energy, molar</td>
<td>$G_m$</td>
<td>$G_m = H_m - TS_m$</td>
<td>J mol⁻¹</td>
</tr>
<tr>
<td>Henry’s law constant</td>
<td>$k_H$</td>
<td>$k_{H,B} = \lim_{x_B \to 0} (f_B/x_B) = (\partial f_B/\partial x_B)_{\lambda_B \to 0}$</td>
<td>Pa</td>
</tr>
<tr>
<td>ionic strength, molality basis</td>
<td>$I_m$</td>
<td>$I_m = \frac{1}{2} \sum z_i^2 m_i$</td>
<td>mol kg⁻¹</td>
</tr>
<tr>
<td>ionic strength, amount conc. basis</td>
<td>$I_c$</td>
<td>$I_c = \frac{1}{2} \sum z_i^2 c_i$</td>
<td>mol m⁻³</td>
</tr>
<tr>
<td>Kuenen coefficient, pure solvent basis</td>
<td>$S^*$</td>
<td>$S_B^* = V(T^<em>, p^</em>, n_B^<em>)/m^1(T, p^</em>, n_A)$</td>
<td>m³ kg⁻¹</td>
</tr>
<tr>
<td>Kuenen coefficient, solution basis</td>
<td>$S$</td>
<td>$S_B = V(T^<em>, p^</em>, n_B^<em>)/m^1(T, p^</em>, n_A, n_B^*)$</td>
<td>m³ kg⁻¹</td>
</tr>
<tr>
<td>mass</td>
<td>$m$</td>
<td>(SI base unit)</td>
<td>kg</td>
</tr>
<tr>
<td>mass concentration</td>
<td>$\gamma, \rho$</td>
<td>$\rho_B = m_B/V$</td>
<td>kg m⁻³</td>
</tr>
<tr>
<td>mass fraction</td>
<td>$w$</td>
<td>$w_B = m_B/\sum m_i$</td>
<td>1</td>
</tr>
<tr>
<td>mass ratio</td>
<td>$\zeta$</td>
<td>$\zeta_{A,B} = m_A/m_B$</td>
<td>1</td>
</tr>
<tr>
<td>molality</td>
<td>$m, b$</td>
<td>$m_B = n_B/M_A(n - \sum n_B)$</td>
<td>mol kg⁻¹</td>
</tr>
<tr>
<td>molar mass</td>
<td>$M$</td>
<td>$M_A = m_A/n_A$ (= mass/amount)</td>
<td>kg mol⁻¹</td>
</tr>
<tr>
<td>molar volume</td>
<td>$V_m$</td>
<td>$V_m = V/n$</td>
<td>m³ mol⁻¹</td>
</tr>
<tr>
<td>mole (amount) fraction</td>
<td>$x$</td>
<td>$x_B = n_B/\sum n_i$</td>
<td>1</td>
</tr>
<tr>
<td>mole ratio</td>
<td>$r$</td>
<td>$r_{12} = n_1/n_2$</td>
<td>1</td>
</tr>
<tr>
<td>number density of entities, number concentration</td>
<td>$C$</td>
<td>$C_B = N_B/V$</td>
<td>m⁻³</td>
</tr>
<tr>
<td>number of entities</td>
<td>$N$</td>
<td></td>
<td>1</td>
</tr>
<tr>
<td>number of phases</td>
<td>$C$</td>
<td></td>
<td>1</td>
</tr>
<tr>
<td>osmotic coefficient, mole fraction basis</td>
<td>$\varphi_x$</td>
<td>$\varphi_{x,A} = (\mu_A - \mu_A^*)/RT \ln x_A$</td>
<td>1</td>
</tr>
</tbody>
</table>

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<table>
<thead>
<tr>
<th>Name</th>
<th>Symbol</th>
<th>Definition</th>
<th>SI Unit</th>
</tr>
</thead>
<tbody>
<tr>
<td>osmotic coefficient, molality basis</td>
<td>( \phi_m )</td>
<td>( \phi_{m,A} = (\mu_A^* - \mu_A)/RTM_A \sum m_B )</td>
<td>1</td>
</tr>
<tr>
<td>Ostwald coefficient, pure solvent basis</td>
<td>( L^* )</td>
<td>( L_B^* = V^s(T, p, n_B^1)/V^l(T, p, n_A) )</td>
<td>1</td>
</tr>
<tr>
<td>Ostwald coefficient, solution basis</td>
<td>( L )</td>
<td>( L_B = V^s(T, p, n_B^1)/V^l(T, p, n_A, n_B^1) = c_B^1/c_B^s )</td>
<td>1</td>
</tr>
<tr>
<td>Poynting correction</td>
<td>( I_p(p) )</td>
<td>( I_p(p) = \int_{p_A}^{p_B} \frac{V^\infty_B}{RT} dp )</td>
<td>1</td>
</tr>
<tr>
<td>pressure</td>
<td>( p )</td>
<td>(normal force)/area</td>
<td>Pa</td>
</tr>
<tr>
<td>pressure, partial</td>
<td>( p_B )</td>
<td>( p_B = \gamma_B p )</td>
<td>Pa</td>
</tr>
<tr>
<td>salinity</td>
<td>( S )</td>
<td>(see entry)</td>
<td>1</td>
</tr>
<tr>
<td>solubility</td>
<td>( s )</td>
<td>composition of saturated solution or mixture (various)</td>
<td></td>
</tr>
<tr>
<td>solubility parameter</td>
<td>( \delta )</td>
<td>( \delta_B = (\Delta_{vap} E_{m,B}/V_{m,B})^{1/2} )</td>
<td>Pa(^{1/2})</td>
</tr>
<tr>
<td>solubility product</td>
<td>( K_s )</td>
<td>(see entry)</td>
<td>1</td>
</tr>
<tr>
<td>solute mole fraction; Jänecke mole fraction</td>
<td>( x_s )</td>
<td>( x_{s,B} = x_B\sum x_i )</td>
<td>1</td>
</tr>
<tr>
<td>solvent mole fraction</td>
<td>( x_v )</td>
<td>( x_{v,A} = x_A\sum x_i )</td>
<td>1</td>
</tr>
<tr>
<td>solute mole fraction</td>
<td>( x_s )</td>
<td>( x_{s,B} = x_B\sum x_i )</td>
<td>1</td>
</tr>
<tr>
<td>solvent volume fraction</td>
<td>( \phi_v )</td>
<td>( \phi_{v,A} = \phi_A\sum \phi_i )</td>
<td>1</td>
</tr>
<tr>
<td>stoichiometric number (of a salt)</td>
<td>( \nu )</td>
<td>( \nu = \nu_+ + \nu_- )</td>
<td>1</td>
</tr>
<tr>
<td>temperature (Celsius)</td>
<td>( \theta, t )</td>
<td>( \theta^\circ C = T/K - 273.15 )</td>
<td>( \circ C )</td>
</tr>
<tr>
<td>temperature, thermodynamic</td>
<td>( T )</td>
<td>(SI base unit)</td>
<td>K</td>
</tr>
<tr>
<td>volume</td>
<td>( V )</td>
<td></td>
<td>m(^3)</td>
</tr>
<tr>
<td>volume, partial molar</td>
<td>( V_B )</td>
<td>( V_B = (\partial V/\partial n_B)_{T,p,n_i\neq n_B} )</td>
<td>m(^3) mol(^{-1})</td>
</tr>
<tr>
<td>volume fraction</td>
<td>( \phi )</td>
<td>( \phi_B = x_B V^<em>_{m,B}\sum x_j V^</em>_{m,j} )</td>
<td>1</td>
</tr>
</tbody>
</table>
4.2 Subscripts and superscripts

4.2.1 Subscripts
A general constituent, usually solvent
B general constituent
c critical state
c concentration basis
i general constituent
m molar (divided by amount of substance)
m molality basis
r reference, reduced
sln solution
s saline solution
trs transition
vap vaporization
v volume
x mole fraction basis
+, – positive, negative charge

4.2.2 Superscripts

g vapor
l liquid
α, β, γ labels for phases
* pure substance
θ, o standard state
∞ infinite dilution
+, – positive, negative charge

4.2.3 Other symbols
Δ difference
Δβα change from phase α to phase β
∥ indicator of reciprocal salt system
⇌ equilibrium process
+ separator between components in mixture; e.g., Mg + Sn
% percent
‰ permil