


**Glossary of Terms Related to Solubility**

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INTERNATIONAL UNION OF PURE AND APPLIED CHEMISTRY

ANALYTICAL CHEMISTRY DIVISION\*

SUBCOMMITTEE ON SOLUBILITY AND EQUILIBRIUM DATA\*\*

## GLOSSARY OF TERMS RELATED TO SOLUBILITY

### (IUPAC RECOMMENDATIONS 200-)

*Prepared for publication by*

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# Glossary of Terms Related to Solubility

## (IUPAC Recommendations 200-)

*Abstract:*

This glossary defines 151 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.

### 1. INTRODUCTION

Disciplines concerned with solubility and related phenomena extend well beyond the traditional branches of chemistry to a wide range of bio-medical, 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 [1] and quantities, units, and symbols [2] at the time of its preparation (2006).

This glossary defines 151 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* at 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* (Gold Book), particularly the online version [1]. The authors also recommend the IUPAC publication *Quantities, Units and Symbols in Physical Chemistry* (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 object 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. Where synonymous terms are in common usage the recommended term ("main term") is followed on the second line of the glossary entry 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)".

### 2. GLOSSARY OF TERMS

1  
2  
3  
4  
5  
6 **1. absorption coefficient,  $\beta^*$**  (in gas solubility)

7 Volume  $V^g$  of an amount  $n_B^1$  of a dissolved gas at a given *standard temperature*, usually  $T^\circ =$   
8 273.15 K, and total *standard pressure*  $p^\circ$  divided by the volume of the pure dissolving liquid of  
9 volume  $V^l$  that contains an amount  $n_A$  of *solvent* at the same temperature  $T$  and pressure  $p$ .  
10

11 Note 1: There is only one absorption coefficient, as compared to the *Bunsen*, *Ostwald* and  
12 *Kuenen coefficients*, because the volume of gas in the definition contains both solute and solvent.  
13 The mathematical definition is:  
14

15 absorption coefficient, pure solvent reference

$$\beta_B^* = V^g(T, p_A + p_B = p^\circ, n_B^1) / V^l(T, p^\circ, n_A)$$

16 where  $p_A, p_B$  are the partial pressures of solvent and gas.  
17

18 Note 2: For an ideal gas, the absorption coefficient and *Bunsen coefficient* are related by:  
19

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

20 since  $V^g$  is inversely proportional to pressure.  
21

22 Note 3: The relations between the *molality*  $m_B(p^\circ)$  or *amount fraction*  $x_B(p^\circ)$  of dissolved gas  
23 and the absorption coefficients are:  
24

$$\frac{1}{x_B} = 1 + \frac{1}{m_B(p^\circ) M_A} = 1 + \frac{RT^\circ Z_B^\circ (1 + p_A / p^\circ)}{p^\circ V_{m,A} \beta_B^*}$$

25 where  $V_A, V_{m,A}$  are the respective partial molar volume and molar volume of the solvent and  $Z_B$   
26 is the compression factor of the gas.  
27

28 Note 4: The absorption coefficient and the related quantities for expression of gas solubility;  
29 Bunsen coefficient, Kuenen coefficient, and Ostwald coefficient appear frequently in the older  
30 literature of gas solubility determination. However, the modern practice, recommended here, is  
31 to express gas solubility as *molality*, *amount fraction* or *amount ratio*.  
32

33 **2. activity coefficient,  $f, \gamma_m, \gamma_c$**

34 Dimensionless correction factor that multiplies the quantity used to express the composition of  
35 the substance (usually *amount fraction*, *molality* or *amount concentration*) to produce the  
36 (relative) activity of the substance.  
37

38 (a) Referenced to *Raoult's law*, and for a substance B in a liquid or solid *mixture* containing  
39 amount fractions  $x_B, x_C, \dots$  of the substances B, C,  $\dots$ : a dimensionless quantity  $f_B$  defined in  
40 terms of the chemical potential  $\mu_B$  of B in the mixture by:  
41

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

where  $x$  denotes the set of amount fractions  $x_B, x_C, \dots$  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*, and for a *solute* B in a *solution* (especially a dilute liquid solution) containing molalities  $m_B, m_C, \dots$ , because  $m_B$  denotes the molality of solute B and  $m_A$  denotes the mass of solvent A) of *solutes* B, C, ... 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:

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

$$\mu_{m,B}^\ominus = \lim_{m_B \rightarrow 0} [\mu_B - RT \ln(m_B / m^\ominus)]$$

where  $m^\ominus = 1 \text{ mol kg}^{-1}$  is the standard molality and  $\mu_B^\ominus$  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.

(c) Referenced to *Henry's law*, and for a *solute* B in a *solution* (especially a dilute liquid solution) containing amount concentrations  $c_B, c_C, \dots$  of solutes B, C, ... 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:

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

$$\mu_{c,B}^\ominus = \lim_{c_B \rightarrow 0} [\mu_B - RT \ln(c_B / c^\ominus)]$$

where  $c^\ominus = 1 \text{ mol dm}^{-3}$  is the standard amount concentration and  $\mu_{c,B}^\ominus$  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].

### 3. activity coefficient at infinite dilution, $f^\infty$

For a substance B, *activity coefficient*  $f_B$  extrapolated to *infinite dilution*:

$$\ln f_B^\infty = \lim_{x_B \rightarrow 0} \left( \frac{\mu_B - \mu_B^*}{RT} - \ln x_B \right)$$

Note 1: Useful for dilute *mixtures* as an alternative to the standard chemical potential on a *molality* basis, and as such a fundamental thermodynamic property.

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:

$$\ln f_B^\infty = \frac{\mu_{m,B}^\ominus - \mu_B^*}{RT} - \ln(M_A m_B^\ominus)$$

where  $M_A$  is the molar mass of the *solvent*.

See also *activity coefficient*

**4. amount,  $n$** 

amount of substance

Base quantity in the system of quantities upon which SI is based. It is the number of elementary entities divided by the Avogadro constant.

Note 1: Since amount is proportional to the number of entities, the proportionality constant being the reciprocal Avogadro constant and the same for all substances, it has to be treated almost identically with the number of entities. Thus the counted elementary entities must always be specified.

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 the simplest formula unit [e.g. amount of sodium chloride usually means  $n(\text{NaCl})$ ] and for metals 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. The 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.

[1]

**5. amount concentration,  $c$** 

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

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

Note 1: For constituent B it is often denoted [B].

Note 2: The common unit is mole per cubic decimeter ( $\text{mol dm}^{-3}$ ) or mole per liter ( $\text{mol L}^{-1}$ ) sometimes denoted by (small capital) M.

Modified from [1].

**6. amount fraction,  $x$** 

mole fraction

*Amount* of substance of a *constituent* divided by the total amount of all constituents in a *mixture*.

Modified from [1].



**7. amount ratio,  $r$** 

mole ratio

*Amount* of one *constituent* divided by the amount of a second constituent in the same *system*. For constituents 1 and 2,

$$r_{12} = n_1 / n_2$$

[3]

**8. analytical method**

Class of experimental procedures for *solubility* determination in which a *saturated solution* is prepared and then analyzed to determine composition.

See also *synthetic method*.

**9. binary system**

*System* containing two *components*.

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

**10. binodal curve**

See *coexistence curve*.

**11. Bunsen coefficient,  $\alpha$** 

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

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

$$\begin{aligned} \text{Bunsen coefficient, solution reference} & \quad \alpha_B = V^g(T^\circ, p^\circ, n_B^l) / V^l(T, p^\circ, n_A, n_B^l) \\ \text{Bunsen coefficient, pure solvent reference} & \quad \alpha_B^* = V^g(T^\circ, p^\circ, n_B^l) / V^l(T, p^\circ, n_A) \end{aligned}$$

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

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

$$\frac{1}{x_B^\circ} = 1 + \frac{1}{m_B(p^\circ) M_A} = 1 + \frac{RT^\circ Z_B^\circ(T^\circ)}{p^\circ V_A \alpha_B}$$

$$\frac{1}{x_B^\circ} = 1 + \frac{1}{m_B(p^\circ)M_A} = 1 + \frac{RT^\circ Z_B^\circ(T^\circ)}{p^\circ V_{m,A} \alpha_B^*}$$

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*, *amount fraction* or *amount ratio*.

## 12. catatectic reaction

See *metatectic reaction*

## 13. chlorinity, *Cl*

Mass of chlorine (in all forms, usually chloride) divided by mass of saline *solution*.

Note 1: Used in describing composition of saline solutions, e.g., seawater, brackish waters and brines.

Note 2. Usual unit for this mass fraction is *permil*.

See also *salinity*.

[4]

## 14. Clarke-Glew equation

Semi-empirical equation describing the temperature dependence of the standard Gibbs energy of solution:

$$\Delta_{\text{sln}}G^\circ(T, p^\circ) / RT = A_0 + A_1(T_r / T) + A_2 \ln(T / T_r) + A_3(T / T_r) + A_4(T / T_r)^2 + \dots$$

where the  $A_i$  are constants that can be related to thermodynamic quantities and  $T_r$  is a reference temperature.

See also *Clarke-Glew-Weiss equation*

[5, 6]

## 15. Clarke-Glew-Weiss equation

*Clarke-Glew equation* with  $T_r = 100$  K.

[5, 7]

## 16. cloud point



1  
2  
3 *Critical solution point*, particularly when used as an end point of a turbidometric titration to  
4 determine *solubility*.  
5  
6

7 See also *synthetic method*  
8  
9

### 10 **17. coexistence curve**

11 binodal curve

12 conodal curve

13 Boundary of stable *phase* separation (limits of solubility) in a liquid or solid *system* of two or  
14 more *components*.  
15  
16

17 Note: In a *binary system*, a plot of experimental variables such as temperature or pressure, or of  
18 theoretical variables such as Gibbs energy, against *amount fraction* displays the coexistence  
19 curve. In *ternary* or *higher-order systems*, coexistence curves are displayed on a plot of  
20 composition at constant temperature or pressure, usually on a *ternary diagram*.  
21  
22

23 See also *spinodal curve*, *conjugate phases*.  
24

25 From [8]  
26  
27

### 28 **18. common ion effect**

29 Decrease in solubility of a salt when a second *non-saturating* salt with one ion in common with  
30 the salt is added to its *saturated solution*.  
31  
32

33 Note: Restricted in practice to salts with sufficiently low *solubility* and at sufficiently *low ionic*  
34 *strengths* that the *activity coefficient* of the salt is close to unity. At higher ionic strengths, or for  
35 salts with high solubility, the activity coefficient usually decreases with addition of non-  
36 saturating salt, resulting in a higher solubility.  
37  
38  
39

### 40 **19. component**

41 thermodynamic component

42 independent component

43 *Constituent* of a *mixture* the *amount* or concentration of which can be varied independently. The  
44 number of components in a given *system* is the minimum number of independent species  
45 necessary to define the composition of all *phases* of that system. Relations among the  
46 constituents, such as charge balance or chemical *equilibria*, must be taken into account in  
47 determining the number of components. The term component is also often used in the more  
48 general sense defined here under constituent.  
49

50 Modified from [1].  
51  
52  
53

### 54 **20. compression factor, $Z$**

55 compressibility factor

56 Product of pressure and molar volume divided by the gas constant and thermodynamic  
57  
58  
59  
60

1  
2  
3 temperature. For an ideal gas it is equal to 1.

4 [1]  
5  
6  
7

## 8 **21. conjugate phases**

9 Two *phases* of variable composition in mutual thermodynamic *equilibrium*.  
10

11 See also *coexistence curve*.

12 From [8]  
13

## 14 **22. conodal**

15 connodal

16 See *tie line*

17 From [8]  
18  
19  
20

## 21 **23. conodal (connodal) curve**

22 See *coexistence curve*.  
23  
24  
25

## 26 **24. constituent**

27 Chemical species present in a *system*.  
28

29 See also *component*  
30

31 Modified from [1].  
32  
33  
34

## 35 **25. critical index $\beta$**

36 Index in the basic equation describing a *coexistence curve*:

$$37 \beta = \lim_{T \rightarrow T_c} \left( \frac{\partial \ln(x'' - x')}{\partial \ln|T - T_c|} \right)$$

38 where  $x''$ ,  $x'$  are the compositions of one *component* in the respective concentrated and dilute  
39 phases and  $T_c$  is the *critical solution temperature*.  
40

41 From [8]  
42  
43  
44  
45  
46

## 47 **26. critical point**

48 Temperature and pressure at which intensive properties of liquid and vapor (density, heat  
49 capacity, etc.) become equal.  
50

51 Note: The highest temperature (*critical temperature*) and pressure (*critical pressure*) at which  
52 both a gaseous and a liquid phase of a given compound can exist.  
53

54 [1]  
55  
56  
57

## 58 **27. critical pressure, $p_c$**

1  
2  
3 In *binary, ternary* and *higher-order liquid-liquid systems*, the point on a *phase diagram* of  
4 pressure against composition (at constant temperature) at which the liquid and vapor composition  
5 become equal.  
6  
7

## 28. critical solution point

cloud point

consolute point

plait point (ternary and higher-order systems)

10  
11  
12  
13 In a *binary system*, point with coordinates critical solution temperature (CST) or critical solution  
14 pressure (CSP) and critical composition on a temperature-composition or pressure-composition  
15 *phase diagram* at which the distinction between coexistent *phases* disappears.  
16

17  
18 In *ternary* and *higher-order systems*, composition below or above which, on a ternary or higher  
19 composition *phase diagram* at constant temperature and pressure, the distinction between  
20 coexistent phases disappears.  
21

22 Note 1: Disappearance of the distinction between phases corresponds to disappearance of a  
23 *miscibility gap*; see *mutual solubility*.  
24

25  
26 Note 2: In solid-solid, solid-liquid and liquid-liquid systems both upper and lower critical  
27 solution temperatures (UCST, LCST) or upper and lower critical solution pressures (UCSP,  
28 LCSP) can occur. In some systems, both can be observed.  
29

30 Modified from [1].  
31

## 29. critical temperature, $T_c$

32  
33 In *binary, ternary* and *higher-order liquid-liquid* or *solid-solid systems*, point on a *phase*  
34 *diagram* of temperature against composition (at constant pressure) at which the liquid and vapor  
35 composition become equal.  
36  
37

38 Note: The highest temperature at which vapor-liquid *equilibrium* can exist.  
39

40 Modified from [1].  
41  
42

## 30. crystallization

43  
44 Formation of a crystalline solid *phase* from either (a) a solid, liquid or gaseous *mixture* or  
45 *solution*, or (b) a pure liquid or gas.  
46  
47

48 Note: Crystallization usually occurs under laboratory conditions by altering the temperature or  
49 pressure of a *system*, or by evaporation of a *solvent*.  
50

51 [1]  
52  
53

## 31. data

54  
55 Experimental results, often numerical  
56  
57  
58  
59  
60

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, represented by smooth lines.

### 32. dissolution

Process of mixing of two or more *phases* with the formation of one new homogeneous phase (i.e., the *solution*)

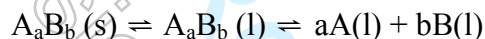
Modified from [1].

### 33. dystectic reaction

congruent melting

indifferent melting [9]

Isobaric, reversible melting or *dissolution* with either complete or partial dissociation on heating of a solid compound,  $A_aB_b$ , formed by *components* A and B:



where the forward arrow indicates 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.

[9]

Note 2: Derived from Greek, meaning “difficult (or highest) melting”.

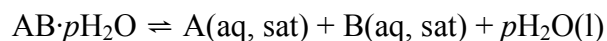
Examples: systems Mg-Sn [10],  $H_2SO_4 - H_2O$

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

### 34. dystonic reaction

congruent dissolution

Reversible *dissolution* in an isothermal, isobaric *system* of three or more *components* characterized by congruent dissolution of and *saturation* with a stoichiometric compound consisting of two or more of these components. The *equilibrium* process is, for example:



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: Derived from Greek, meaning “difficult (or highest) tension (or vapor pressure)”.

Examples: systems  $\text{Na}_2\text{SO}_4\text{-ZnSO}_4\text{-H}_2\text{O}$ ,  $\text{Na}_2\text{SO}_4\text{-CdSO}_4\text{-H}_2\text{O}$ ,  $\text{Na}_2\text{SO}_4\text{-CuSO}_4\text{-H}_2\text{O}$

### 35. equilibrium

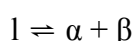
State of a *system* in which the macroscopic properties of the system become uniform, independent of time. If the temperature is uniform throughout the system, a state of thermal equilibrium has been reached; if the pressure is uniform, a state of mechanical equilibrium has been reached; and if the chemical 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 also 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. For these variables, the condition for stable equilibrium is that a variation in the total Gibbs energy at equilibrium caused by variations in any of the variables is zero,  $\delta G(T, p, \mu_i) = 0$ ; 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 amount 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 amount 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.

### 36. 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:



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 composition 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.

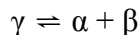
Note 2: The eutectic line and composition (hence point) are isobaric invariants of the system, and represent the composition and the minimum melting temperature along the two intersecting melting curves.

Note 3: Derived from Greek, meaning “easy (or lowest) melting”.

Example: system Ag-Cu [10] and many salt-water systems

### 37. 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*:



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 univariants of the system, define the eutectoid point.

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

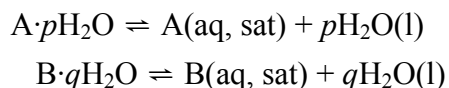
Example: system Fe-C  
[10]

### 38. eutonic reaction

isothermally invariant reaction

(isothermally invariant) reaction of double saturation

Reversible isoplethic *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:



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 non-volatile 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: The term is derived from Greek meaning “easy (or lowest) tension (or 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 *peritonic reaction*

### 39. 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.

### 40. fugacity

Of a substance B,  $f_B$  or  $\tilde{p}_B$ , in a gaseous *mixture* is defined by  $f_B = \lambda_B \lim_{p \rightarrow 0} (p_B / \lambda_B)_T$ , where  $p_B$  is the *partial pressure* of B and  $\lambda_B$  its absolute activity.

SI unit: Pa.  
[1], [2]

### 41. fugacity coefficient, $\varphi$

*fugacity* divided by the *partial pressure* of a gaseous *constituent*.

Modified from [1]

### 42. 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:

$$\sum_{i=1}^c x_i (d\mu_i + S_i dT - V_i dp) = 0$$

where  $S_i$ ,  $V_i$ ,  $x_i$  are the respective partial molar entropy, partial molar volume and amount fraction of component  $i$ .

Note 3: There is a Gibbs-Duhem equation for each phase in a *system* exhibiting multi-phase 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.

[9]

### 43. Gibbs energy of transfer

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

### 44. 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,  $\beta$ :

$$-\frac{(1-x_B^\beta)\Delta_\alpha^\beta H_A + x_B^\beta \Delta_\alpha^\beta H_B}{T} dT + [(1-x_B^\beta)\Delta_\alpha^\beta V_A + x_B^\beta \Delta_\alpha^\beta V_B] dp + (x_B^\alpha - x_B^\beta) \left( \frac{\partial^2 G_m^\alpha}{\partial x_B^{\alpha 2}} \right)_{T,p} dx_B^\alpha = 0$$

$$-\frac{(1-x_B^\alpha)\Delta_\alpha^\beta H_A + x_B^\alpha \Delta_\alpha^\beta H_B}{T} dT + [(1-x_B^\alpha)\Delta_\alpha^\beta V_A + x_B^\alpha \Delta_\alpha^\beta V_B] dp + (x_B^\alpha - x_B^\beta) \left( \frac{\partial^2 G_m^\beta}{\partial x_B^{\beta 2}} \right)_{T,p} dx_B^\beta = 0$$

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

Note 1:  $\left( \frac{\partial^2 G_m}{\partial x_B^2} \right)_{T,p} > 0$  (condition for diffusional stability). This quantity may also be expressed

in terms of the derivatives of the chemical potentials, using

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

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 these equations are called van der Waals' equations.

Note 4: Sometimes the spelling Konovalow is found.

[9]

#### 45. Henry's law

*Partial pressure (fugacity) of a solute (B) in a solution is directly proportional to the relative activity,  $a_B$ , of the solute:*

$$p_B = a_B/a_B^\infty$$

where  $a_B^\infty$  is the activity at *infinite dilution*, i.e., for pure *solvent*.

Note 1: For the solvent (A) the relationship is called Raoult's law, and the proportionally factor is the fugacity of the pure solvent,  $p_A = p_A^* a_A$ .

Note 2. Henry's law is a limiting law, but can sometimes be used for converting solubility *data* from the experimental pressure to standard partial pressure, provided the *amount(mole) fraction* of the gas in the liquid is small and that the difference in pressures is small.

See also *Henry's law constant*.

Modified from [1].

#### 46. Henry's law constant, $k_H$

Henry coefficient, Henry's law coefficient, Henry constant

For a gas B,

$$k_H = \lim_{x_B \rightarrow 0} (f_B / x_B) = (\partial f_B / \partial x_B)_{x_B \rightarrow 0}$$

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 *amount fractions* has units Pa, but is sometimes expressed in terms of molalities or *amount concentrations*, with corresponding units Pa kg mol<sup>-1</sup>, Pa m<sup>3</sup> mol<sup>-1</sup>, 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.

See also *Henry's law*.

From [2, 5].

#### 47. higher-order system

multicomponent system

*System containing more than three components.*

**48. incongruent reaction**

See *peritectic reaction*

**49. independent component**

See *component*

**50. 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.

**51. infinite miscibility**

Property of a *system* of forming a single *phase* at all relative proportions of its *components*.

See also *mutual solubility*

**52. initial complex method**

See *wet residue method*

**53. ionic amount (mole) fraction,  $x_+$ ,  $x_-$**   
species amount fraction, species mole fraction

(a) For an ionized salt B in *solution*:

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

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:

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

Note 1: These are generalizations of terms defined in [11], and are used in formulating fitting equations for *solubility* of salts, in defining *activity coefficients* on the *amount fraction* scale, and in discussing salt effects on solubilities of gases.

Note 2: Note that:

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

[5]

#### 54. ionic strength

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

(a) on molality basis:  $I_m = \frac{1}{2} \sum z_i^2 m_i$

(b) on amount concentration basis:  $I_c = \frac{1}{2} \sum z_i^2 c_i$

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

From [2].

#### 55. 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*.

[12]

#### 56. isobar

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

#### 57. isopleth

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

#### 58. isotherm

Line joining points of equal temperature on a *phase diagram*.

#### 59. isothermal method

See: *synthetic method*

#### 60. Jänecke coordinates

solute amount (mass) fraction.

Coordinates used mainly in representation of multicomponent *phase diagrams* that distinguish *solutes* and a single *solvent*. The *amount* or *mass fractions* of the  $s$  solutes are chosen as the primary composition variables and the amount 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 amount fraction of solvent component A is replaced by a special case of the *solvent amount fraction*  $x_{v,A} = x_A / \sum_{i=s+1}^p x_i$ , 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 amount (mass) fraction*.

[5, 13]

### 61. Kuenen coefficient, $S$

Volume  $V^g$  of an amount  $n_B^1$  of a dissolved gas at a given *standard temperature*  $T^\circ$  (usually 273.15 K) and given *standard pressure*  $p^\circ$  (usually 1 atm) divided by the mass  $m^1$  of the dissolving liquid containing an amount  $n_A$  of *solvent* at temperature  $T$  and the given pressure  $p^\circ$ .

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

$$\begin{aligned} \text{Kuenen coefficient, solution reference} \quad S_B &= V^g(T^\circ, p^\circ, n_B^1) / m^1(T, p^\circ, n_A, n_B^1) \\ \text{Kuenen coefficient, pure solvent reference} \quad S_B^* &= V^g(T^\circ, p^\circ, n_B^1) / m^1(T, p^\circ, n_A) \end{aligned}$$

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

$$\begin{aligned} \frac{1}{x_B^\circ} &= 1 + \frac{1}{m_B(p^\circ) M_A} = 1 + \frac{RT^\circ Z_B^\circ(T^\circ)}{p^\circ M_A S_B} \\ \frac{1}{x_B^\circ} &= 1 + \frac{1}{m_B(p^\circ) M_A} = 1 + \frac{RT^\circ Z_B^\circ(T^\circ)}{p^\circ M_A S_B^*} \end{aligned}$$

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*, *amount fraction* or *amount ratio*.

### 62. lower critical solution temperature

See *critical solution point*



**63. mass concentration,  $\gamma, \rho$** 

mass density

Mass of a *constituent* divided by the volume of a *mixture*.

From [2].

**64. mass density**

Mass of a pure substance, *mixture* or *solution* divided by its volume.

From [2].

**65. mass fraction,  $w$** 

Mass of a particular *constituent* divided by the sum of the masses of all *system* constituents.

$$w_B = m_B / \sum_i m_i$$

Note: *Solubility* may be expressed in terms of mass fraction as mg(1) / kg solution. However, it is recommended that the term “parts per million, ppm” be avoided since “parts” may be measured on either a mass or volume basis. If it is used, the physical quantity to which it refers must be specified.

From [2].

**66. mass percent**

For a *system constituent* 100 times the *mass fraction* of that constituent.

**67. mass ratio**

Mass of one *constituent* divided by the mass of a second constituent in the same *system*.

**68. mass solubility,  $C_w$** 

Obsolete term for *molality* of a dissolved gas.

**69. melt**

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

**70. metastable state**

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.

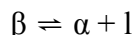
Note 2: Although a driving force for the transition of a metastable state to a stable state exists ( $\Delta_{tr}G < 0$  at constant  $T$  and  $p$ ) the transition is retarded during the time of observation.

Modified from [1].

### 71. metatectic reaction

catatectic reaction [14]

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*:



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: Derived from Greek, meaning “behind, or second-order, after melting”.

Example: system Fe-Zr  
[10]

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

### 72. miscibility

Ability of two liquids to undergo spontaneous mixing to form a homogeneous *mixture*.

### 73. miscibility gap

See *mutual solubility*.

### 74. mixture

Gaseous, liquid or solid *phase* containing more than one substance, when all the substances are treated in the same way.

See also *solution*.  
[2, 15]

### 75. molality, $m$ , $b$

*Amount* of a *solute* divided by the mass of the *solvent*.

Modified from [1]

### 76. mole fraction, $x$

See *amount fraction*

**77. mole percent**

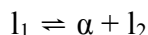
For a *system constituent* 100 times the *amount fraction* of that constituent.

**78. mole ratio,  $r$** 

See *amount ratio*

**79. monotectic reaction**

Isothermal reversible reaction of a solid and a liquid *phase* to form a new solid phase during cooling of a *system*. In a *binary system*:



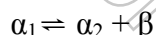
where the forward arrow indicates the direction of cooling. The *equilibrium* transformation occurs along the monotectic line, characterized by the monotectic temperature. The monotectic composition and temperature, isobaric invariants of the system, define the monotectic point, which lies between the compositions of liquid  $l_2$  and solid phase  $\alpha$ .

Note: Derived from Greek, meaning “single melting”.

Examples: systems Pb-Zn [10], methanol-cyclohexane

**80. 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*:



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

Example: system Al-Zn

[10]

Modified from [1].

**81. multicomponent system**

See *higher-order system*

**82. 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*.

**83. 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*

**84. number concentration,  $C$** 

Number of entities of a constituent in a *mixture* divided by the volume of the mixture.

[1]

See also *amount*.

**85. osmotic coefficient  $\varphi_x, \varphi_m$** 

Factor to correct for non-ideal behavior of the *solvent* in a *solution*.

(a) *Amount fraction* basis. For solvent A,

$$\varphi_{x,A} = (\mu_A - \mu_A^*)/RT \ln x_A$$

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  $\sum m_B$ ,

$$\varphi_{m,A} = (\mu_A^* - \mu_A)/RTM_A \sum m_B$$

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  $\nu$  and molality  $m_B$ ,

$$\varphi_{m,A} = (\mu_A^* - \mu_A)/\nu RTM_A m_B$$

From [2].

**86. Ostwald coefficient,  $L$** 

Volume  $V^g$  of an amount  $n_B^1$  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  $L_B = V^g(T, p, n_B^1) / V^l(T, p, n_A, n_B^1) = c_B^l / c_B^g$

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*  $m_B(p^\ominus)$  or *amount fraction*  $x_B(p^\ominus)$  of dissolved gas and the Ostwald coefficients are:

$$\frac{1}{x_B} = 1 + \frac{1}{m_B(p)M_A} = 1 + \frac{RTZ_B}{V_A p_B L_B}$$

$$\frac{1}{x_B} = 1 + \frac{1}{m_B(p)M_A} = 1 + \frac{RTZ_B}{pV_{m,A}L_B^*}$$

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 [16]

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

### 87. Ostwald ripening

Growth of larger crystals from those of smaller size which have a lower *solubility* than the larger ones.

[1]

### 88. partial miscibility

See *mutual solubility*.

### 89. partial pressure

For a *mixture* of gases the contribution by a gaseous *constituent* B to the total pressure.

Partial pressure  $p_B = y_B p$ , where  $y_B$  is the *amount fraction* of constituent B and  $p$  is the total pressure.

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

### 90. percent, %

One part in a hundred parts.

Example: The *amount fraction*  $x = 2.8 \times 10^{-2} = 2.8 \%$

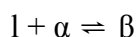
[1]

### 91. peritectic reaction

incongruent reaction.

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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, ... ( $n - 1$ , where  $n$  is the number of *components*) new solid *phases*. For example, in a binary system



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  $l$  and  $\alpha$ .

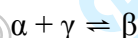
Note: Derived from Greek, peri- = around and -tektos = fusible.

Example: alloy system Cu-Zn [10], salt-water system  $\text{Na}_2\text{SO}_4\text{-H}_2\text{O}$

## 92. peritectoid reaction

metatectoid 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:



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 acceptable as synonym for peritectoid.

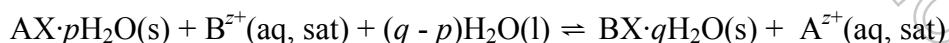
Example: systems Al-Cu [10], hexacosane ( $n\text{-C}_{26}\text{H}_{54}$ ) – octacosane ( $n\text{-C}_{28}\text{H}_{58}$ )

Modified from [1].

## 93. peritonic reaction

transition (reaction)

Isothermal (or isobaric), isoplethic reaction, especially in *ternary systems*, where a pure compound and its binary compound, or two binary compounds containing one ionic *component* (of an ionic compound) or one neutral *component* (of a binary compound) in common, or a binary and a ternary compound with one or more components in common, exist in *equilibrium* with saturated liquid. The equilibrium process is, for example:



Note: Named from a combination of eutonic and peritectic.



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4  
5 Example: system KCl-MgCl<sub>2</sub>-H<sub>2</sub>O  
6  
7

8 **94. permil, ‰**

9 per mille

10 permille

11 promille

12 One part in a thousand parts or  $1000 \times$  (mass fraction) or  $\text{g kg}^{-1}$ .

13 See also *salinity*  
14  
15

16  
17  
18 **95. phase**

19 *System* or portion of a system which is uniform in chemical composition and physical state.  
20

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

23 Modified from [1].  
24  
25

26  
27 **96. phase diagram**

28 Graphical representation of chemical *equilibrium* by points, lines, and surfaces in unary, *binary*,  
29 *ternary*, and *higher-order systems* containing two or more *phases*.  
30  
31

32  
33 **97. phase rule**

34 Gibbs' phase rule

35 Relation connecting number of possible stable *phases P* in an *equilibrium system* with the  
36 number of *components, C*, and the number of degrees of freedom, *F*, i.e., the number of variables  
37 that can be assigned free values:  
38

$$F = C + 2 - P$$

39  
40  
41 Note 1: The phase rule as stated holds under the conditions: (a) negligible surface contributions  
42 (unless the curvature is constant); (b) uniform normal pressure is the only external force; (c)  
43 Interphase surfaces are deformable, heat-conducting and permeable to all components.  
44  
45

46 Note 2: Sometimes the Phase Rule is written as  $\nu = (C - r) + 2 - P$  where *r* is the number of  
47 chemical reactions which can take place in the system. This form requires that *C* be the number  
48 of species assumed for the system, rather than the number of components.  
49  
50

51 Note 3: For the phase rule in the presence of surface phases see [17].

52 Modified from [1, 17]  
53

54 **98. plait point**

55 See *critical solution point*.  
56  
57  
58  
59  
60

**99. polythermal method**

See *synthetic method*

**100. precipitation**

The sedimentation 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.

[1]

**101. 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.

**102. Raoult's Law**

See *Henry's law*

**103. reciprocal salt system**

*System* containing  $n$  ionic species and a single solvent, Because of electrical neutrality, the system has  $n$  components.

Note 1: Notation for reciprocal salt systems specifies all ionic species. For example:



denotes a four-component (quaternary) reciprocal salt system.

Note 2: In this example, the solution may become saturated with respect to any one of KCl, NaCl, KNO<sub>3</sub>, NaNO<sub>3</sub> or their hydrates.

Note 3: Ternary systems containing a common ion can be considered as a limiting case of this class of system, where there is only one cation or anion. For example:

**104. 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.

1  
2  
3 (2) In ternary liquid systems with *coexisting phases*, where the *critical solution point* is not a  
4 maximum on a *ternary diagram*, under specific conditions it is possible to pass from a  
5 homogeneous mixture in which the composition of a given component is greater than that of the  
6 critical mixture to a heterogeneous mixture and back to a homogeneous mixture by altering the  
7 relative amounts of the two other components.  
8  
9 [18]  
10

### 11 12 13 **105. salinity, $S$**

- 14 (a) Conventional salinity,  $S_T$ : mass of total salt (in all forms) divided by mass of saline *solution*.  
15 (b) Relative to *chlorinity*:  $S = 1.80655 Cl$  in seawater and brackish waters.  
16

17 Note 1: Used in describing composition of saline solutions, e.g., seawater, brackish waters and  
18 brines.  
19

20 Note 2: Usual unit for this mass fraction is *permil*.  
21

22 See also *chlorinity*.  
23

24 [6]  
25  
26

### 27 **106. salt effect (in solubility)**

28 Change in *solubility* of a *solute* in aqueous solution on addition of a salt that does not possess a  
29 *common ion* with the original *solute*.  
30

31 Note: If the solubility increases on addition of a salt, the addition is said to cause salting-in; if the  
32 opposite, it causes salting-out.  
33

34 See also *Sechenov equation*  
35  
36  
37

### 38 **107. saturated solution**

39 *Solution* which has the same concentration of a *solute* as one that is in *equilibrium* with  
40 undissolved solute at specified values of temperature and pressure.  
41

42 Note: *Ternary* and *higher-order systems* can be saturated with respect to one *component* while  
43 being unsaturated with respect to another.  
44

45 Modified from [1].  
46  
47

### 48 **108. saturating solute**

49 *Solute* which forms a *saturated solution*.  
50  
51

52 Note: Together with the term *non-saturating solute* used to distinguish among solutes in *ternary*  
53 and *higher-order systems*  
54  
55  
56  
57  
58  
59  
60

**109. saturation**

State of a *saturated solution*.

Modified from [1].

**110. saturation vapor pressure**

The 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.

[1]

**111. 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*.

**112. Schreinemakers' method**

See *wet residue method*

**113. Sechenov equation**

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

$$\lg(s_0 / s) = K_s I_s$$

where  $s_0$ ,  $s$  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: Positive values of the Sechenov parameter correspond to the commonly observed salting-out effect; negative values to the less common salting-in effect.

[5]

**114. Sieverts' law**

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

**115. smoothing equation**

See *fitting equation*

**116. solid mixture**

See *mixture*

**117. solid solution**

See *solution*

**118. solubility  $s$** 

The 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*, *amount (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*.

**119. solubility constant**

Generalization of *solubility product* to include the undissociated salt or uncharged cation-anion pair as *solute* and, as well, reaction between *solution* species and the cation or anion of the solid salt.

Note 1: The solubility constant can also be considered 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 [19-23]. See also *solubility product*.

**120. solubility parameter,  $\delta$** 

Parameter used in predicting the *solubility* of non-electrolytes (including polymers) in a given *solvent*. For a substance B:

$$\delta_B = (\Delta_{\text{vap}}E_{\text{m,B}}/V_{\text{m,B}})^{1/2}$$

where  $\Delta_{\text{vap}}E_{\text{m}}$  is the molar energy of vaporization at zero pressure and  $V_{\text{m}}$  is the molar volume [24].

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 [24] for the original definition, theory, and extensive examples.

Note 3: The SI units are  $\text{Pa}^{1/2} = \text{J}^{1/2} \text{m}^{-3/2}$ , but units used frequently are  $(\mu\text{Pa})^{1/2} = (\text{J cm}^{-3})^{1/2}$  or  $(\text{cal cm}^{-3})^{1/2}$ , where  $1 (\text{J cm}^{-3})^{1/2} \approx 2.045 (\text{cal cm}^{-3})^{1/2}$ .

Modified from [1, 24].

### 121. solubility product, $K_s$

*Equilibrium* constant for the *dissolution* process of an ionic solid or an addition compound that dissociate completely in *solution*; for example:



where a hydrated ionic solid (mole ratio water/salt =  $r$ ) has been shown as a common occurrence of a combined ionic solid and addition compound. The equilibrium constant is::

$$K_s = a_{+}^{\nu_+} a_{-}^{\nu_-} a_w^r$$

or

$$K_s = (\gamma_{\pm} / m^0)^{\nu} (m_+)^{\nu_+} (m_-)^{\nu_-} \exp(-rM_w \sum_i m_i \varphi_m)$$

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),  $m^0 = 1 \text{ mol kg}^{-1}$  the standard molality,  $M_w$  the molar mass of water and  $\varphi_m$  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:

$$K_s = (\nu_{\pm} \gamma_{\pm} m / m^0)^{\nu} \exp(-r\nu M_w m \varphi_m)$$

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



Note 2: For the special case where common anions, molality  $m_X$ , with balancing cations of molality  $m_N$ , exist in solution, but there is no reaction with solution species and the cation or anion of the salt:

$$K_s = (\gamma_{\pm}/m^{\circ})^{\nu} m_{+}^{\nu_{+}} (m_{-} + m_X)^{\nu_{-}} \exp[-rM_w(\nu m + m_N + m_X)\phi_m]$$

See also *solubility constant*.

Modified from [1]

### 122. solubilization

Process or agent that increases *solubility* or the rate of *dissolution* of a *solute*.

Note: Sometimes used loosely to include agents such as surfactants which hold a *component* in micelles or colloidal suspension rather than in *solution*.

### 123. solute

Minor component of a *solution* which is regarded as having been dissolved by the *solvent*.

Modified from [1].

### 124. solute amount (mass) fraction

See *Jänecke coordinates*.

### 125. 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 *amount 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*.

[2, 15]

### 126. solvent

Major *component* of a *solution* which is regarded as having dissolved the *solute*.

Modified from [1].

### 127. solvent amount (mass) fraction

Amount or mass fraction of a solvent in a solution containing  $s$  solute constituents and  $p - s$  solvent constituents. For a solvent A:

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

and the summation is over all solvent substances,  $p - s$  in number.

See also *Jänecke coordinates*.

[5]

### 128. 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  $p$  solvents and  $s$  solutes:

$$\varphi_{v,B} = \varphi_B / \sum_{j=s+1}^p \varphi_j$$

where the summation is over the  $p - s$  solvent components.

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

See also *volume fraction*.

[5]

### 129. species amount (mole) fraction

See *ionic amount (mole) fraction*

### 130. 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*.  
From [8]

### 131. 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  $c^\ominus = 1 \text{ mol dm}^{-3}$ , which is universally accepted.

Modified from [1] and [2].

### 132. standard molality, $m^\ominus$ , $m^\circ$ , $b^\ominus$ , $b^\circ$

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  $m^\ominus = 1 \text{ mol kg}^{-1}$ , which is universally accepted.

Modified from [1] and [2].

### 133. standard pressure, $p^\ominus$ , $p^\circ$

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  $p^\ominus = 0.1 \text{ MPa} = 100 \text{ kPa} (= 1 \text{ bar})$ . The value for  $p^\ominus = 100 \text{ kPa}$ , is the IUPAC recommendation since 1982, and is recommended for tabulating thermodynamic data. Prior to 1982 the standard pressure was usually taken to be  $p^\ominus = 101\,325 \text{ Pa} (= 1 \text{ atm})$ , called the *standard atmosphere*. In any case, the value for  $p^\ominus$  should be specified. The conversion of values corresponding to different  $p^\ominus$  is described in [25-27]. The newer value of  $p^\ominus$ , 100 kPa is sometimes called the standard state pressure.

Modified from [1] and [2].

### 134. 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*  $p = p^\ominus$ , 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  $p = p^\ominus$ .

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

[1]

### 135. supercritical fluid

State of a compound, *mixture* or element above its *critical pressure* ( $p_c$ ) and *critical temperature* ( $T_c$ ).

[1]

### 136. supersaturated solution

1  
2  
3 *Solution* which has a greater concentration of a *solute* than one that is in *equilibrium* with  
4 undissolved solute at specified values of temperature and pressure.  
5  
6

7 See also *saturated solution*, *unsaturated solution*.  
8  
9

### 10 **137. supersaturation**

11 State of a *supersaturated solution*.

12 Modified from [1].  
13  
14

### 15 **138. syntectic reaction**

16 Isothermal reversible reaction of two liquid *phases*  $l_1$ ,  $l_2$  which are transformed into a solid phase  
17  $\alpha$  during the cooling of a *system*. For a binary system:  
18  
19



21 where the forward arrow indicates the direction of cooling. The equilibrium reaction occurs  
22 along the syntectic line, characterized by the syntectic temperature. The syntectic composition  
23 and temperature, isobaric invariants of the system, define the syntectic point, which lies between  
24 the composition of the two liquid phases.  
25  
26

27 Note 1: Derived from Greek, meaning “melting together”.  
28  
29

30 Example: system K-Zn

31 [10]  
32

33 Note 2: Occasionally a similar reaction takes place in the subsolidus region. Quite appropriately  
34 it is called syntectoid reaction.  
35  
36  
37

### 38 **139. synthetic method**

39 Class of experimental procedures for *solubility* determination in which a *solution* of known  
40 composition is prepared. Two major subclasses are recognized. In the isothermal method one  
41 liquid *component* is titrated with a second liquid component at constant temperature until  
42 persistent turbidity is observed (i.e., to the *cloud point*). In the polythermal method a *mixture* of  
43 known composition is heated above its solution temperature and monitored visually during  
44 cooling until turbidity (the cloud point) is observed.  
45  
46  
47

### 48 **140. system**

49 Arbitrarily defined part of the universe, regardless of form or size.  
50  
51

52 Note: In the context of *solubility* phenomena a system contains two or more *components* whose  
53 solubility is of interest.  
54

55 Modified from [1].  
56  
57  
58  
59  
60

**141. 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 amount (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 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.

[28]

**142. ternary system**

System containing three *components*.

**143. thermodynamic component**

See *component*

**144. tie-line**

connodal, conodal

Straight line connecting *phases* in *equilibrium* on a *phase diagram*.

[24]

**145. transition point**

See peritonic point.

**146. unsaturated solution**

*Solution* which has a lower concentration of a *solute* than one that is in *equilibrium* with undissolved solute at specified values of temperature and pressure.

**147. unsaturation**

State of an *unsaturated solution*.

**148. upper critical solution temperature**

See *critical solution point*

**149. Van der Waals' equations**

See *Gibbs-Konovalov equations*

**150. volume fraction,  $\varphi$** 

Volume of a *constituent* of a *mixture* divided by the sum of volumes of all constituents prior to mixing. For a substance B:

$$\varphi_{\text{B}} = x_{\text{B}} V_{\text{m,B}}^* / \sum_{j=1}^{\text{C}} x_j V_{\text{m,j}}^*$$

where  $V_{\text{m,j}}^*$  is the molar volume of the pure constituent  $j$ .

See also *solvent volume fraction*.

Extended from [1].

**151. 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.  
[29, 30]

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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, 4].

Name	Symbol	Definition	SI Unit	Notes
absorption coefficient (in gas solubility)	$\beta^*$	$\beta_B^* = V^g(T, p_A + p_B = p^\circ, n_B^1) / V^l(T, p^\circ, n_A)$	1	
activity coefficient, amount fraction basis	$f$	$RT \ln(x_B f_B) = \mu_B(T, p, x) - \mu_B^*(T, p)$	1	
activity coefficient, molality basis	$\gamma_m$	$RT \ln(\gamma_{m,B} m_B / m^\circ) = \mu_B - \mu_B^\circ = \lim_{m_B \rightarrow 0} [\mu_B - RT \ln(m_B / m^\circ)]$	1	
activity coefficient at infinite dilution	$f_B^\infty$	$\ln f_B^\infty = \lim_{x_B \rightarrow 0} \left( \frac{\mu_B - \mu_B^*}{RT} - \ln x_B \right)$	1	
amount concentration	$c$ , [species], M	$c_B = [\text{species B}] = n_B / V$	mol m <sup>-3</sup>	
amount (mole) fraction	$x$	$x_B = n_B / \sum_{i=1}^C n_i$	1	
amount of substance	$n$	(SI base unit)	mol	
amount ratio	$r$	$r_{12} = n_1 / n_2$	1	
Bunsen coefficient, pure solvent basis	$\alpha^*$	$\alpha_B^* = V^g(T^\circ, p^\circ, n_B^1) / V^l(T, p^\circ, n_A)$	1	
Bunsen coefficient, solution basis	$\alpha$	$\alpha_B = V^g(T^\circ, p^\circ, n_B^1) / V^l(T, p^\circ, n_A, n_B^1)$	1	
chemical potential	$\mu$	$\mu_B = (\partial G / \partial n_B)_{T, p, n_{i \neq B}}$	J mol <sup>-1</sup>	
compression factor, compressibility factor	$Z$	$Z_B = p V_{m,B} / RT$	1	
(molar) energy of vaporization	$\Delta_{\text{vap}} E_m$	$\Delta_{\text{vap}} E_m = E_m^g - E_m^l$	J mol <sup>-1</sup>	
(molar) enthalpy	$H_m$	$H/n$	J mol <sup>-1</sup>	
(molar) entropy	$S_m$	$S/n$	J K <sup>-1</sup> mol <sup>-1</sup>	
Gibbs energy	$G$	$G = H - TS$	J	
Kuenen coefficient, pure solvent basis	$S^*$	$S_B^* = V^g(T^\circ, p^\circ, n_B^1) / m^l(T, p^\circ, n_A)$	m <sup>3</sup> kg <sup>-1</sup>	
Kuenen coefficient, solution basis	$S$	$S_B = V^g(T^\circ, p^\circ, n_B^1) / m^l(T, p^\circ, n_A, n_B^1)$	m <sup>3</sup> kg <sup>-1</sup>	
mass	$m$	(SI base unit)	kg	
mass concentration	$\gamma, \rho$	$\rho_B = m_B / V$	kg m <sup>-3</sup>	

mass fraction	$w$	$w_B = m_B / \sum_{i=1}^C m_i$	1	
molality	$m, b$	$m_B = n_B / M_A(n - \sum_B n_B)$		
molar mass	$M$	$M_A = m_A / n_A$ (= mass / amount)	$\text{kg mol}^{-1}$	
number density of entities, number concentration	$C$	$C_B = N_B / V$	$\text{m}^{-3}$	
osmotic coefficient, amount fraction basis	$\varphi_x$	$\varphi_{x,A} = (\mu_A - \mu_A^*) / RT \ln x_A$	1	
osmotic coefficient, molality basis	$\varphi_m$	$\varphi_{m,A} = (\mu_A^* - \mu_A) / RT M_A \sum m_B$	1	
Ostwald coefficient, pure solvent basis	$L^*$	$L_B^* = V^g(T, p, n_B) / V^l(T, p, n_A)$	1	
Ostwald coefficient, solution basis	$L$	$L_B = V^g(T, p, n_B) / V^l(T, p, n_A, n_B) = c_B^g / c_B^l$	1	
partial molar enthalpy	$H_B$	$H_B = (\partial H / \partial n_B)_{T, p, n_i \neq n_B}$	$\text{J mol}^{-1}$	
partial molar entropy	$S_B$	$S_B = (\partial S / \partial n_B)_{T, p, n_i \neq n_B}$	$\text{J K}^{-1} \text{mol}^{-1}$	
partial molar volume	$V_B$	$V_B = (\partial V / \partial n_B)_{T, p, n_i \neq n_B}$	$\text{m}^3 \text{mol}^{-1}$	
pressure	$p$	(normal force) / area	Pa	
solubility	$s$	composition of saturated solution or mixture	(various)	
solubility parameter	$\delta$	$\delta_B = (\Delta_{\text{vap}} E_{m,B} / V_{m,B})^{1/2}$	$\text{Pa}^{1/2}$	
solute amount fraction; Jänecke amount fraction	$x_s$	$x_{s,B} = x_B / \sum_{i=1}^s x_i$	1	
solvent amount fraction	$x_v$	$x_{v,A} = x_A / \sum_{i=s+1}^p x_i$	1	
stoichiometric number (of a salt)	$\nu$	$\nu = \nu_+ + \nu_-$	1	
temperature (Celsius)	$\theta, t$	$\theta / ^\circ\text{C} = T / \text{K} - 273.15$	$^\circ\text{C}$	
temperature, thermodynamic	$T$	(SI base unit)	K	
volume	$V$		$\text{m}^3$	
volume fraction	$\varphi$	$\varphi_B = x_B V_{m,B}^* / \sum_{j=1}^C x_j V_{m,j}^*$	1	

## 4.2 Subscripts and superscripts

#### 4.2.1 Subscripts

A	general constituent, usually solvent
B	general constituent
m	molar (divided by amount of substance)
vap	vaporization
1, 2	general constituent

#### 4.2.2 Superscripts

g	vapor
l	liquid
*	pure substance
$\ominus, 0$	standard state
$\infty$	infinite dilution

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