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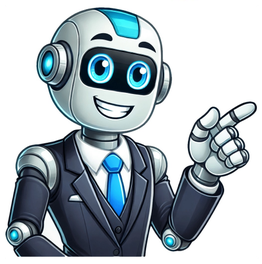

Concentration ConversionsThis article details the most common methods to quantify the concentration of a component in a mixture, and details how to convert between these concentrations. Concentration is commonly expressed in moles per volume, mass per volume, mass percentage, volume percentage, mole percentage and parts per million. Using density and/or molecular weight properties it is possible to convert between any of the above measures.Definitions$m$:mass of componentMW:Molecular weightppm:Parts per million$ppm(w)$:Parts per million, weight basis$ppm(v)$:Parts per million, volume basis$ppm(mol)$:Parts per million, mole basis$v$:Volume fraction$w$:Mass fraction$x$:Mol fractionρ:Density or mass per volume concentrationBelow are tabulated the concentration conversions between many measures of concentration. In all cases we are describing the concentration of substance 1 in total substance, and this is used in subscripts in the formulae below.Click to jump to the section for converting from:Mol FractionConvert ToTypical UnitsConvert From Mol/Mol (x)mass/mass*-$$ \displaystyle w = x \times \frac{MW_{1}}{MW_{tot}} $$volume/volume*-$$ \displaystyle v = x \times \frac{MW_{1}}{MW_{tot}} \times \frac{\rho_{tot}}{\rho_{1}} $$molar concentrationmol/L$$ \displaystyle c = x \times \frac{\rho_{tot}}{MW_{tot}} $$mass/volg/L$$ \displaystyle \frac{m_{1}}{v_{tot}} = x \times \frac{MW_{1}}{MW_{tot}} \times \rho_{tot} $$parts per million(ppm mol)mol/mol$$ \displaystyle ppm(mol) = x \times 1,000,000 $$parts per million(ppm weight)mg/kg$$ \displaystyle ppm(w) = x \times \frac{MW_{1}}{MW_{tot}} \times 1,000,000 $$parts per million(ppm volume)L/L$$ \displaystyle ppm(v) = x \times \frac{MW_{1}}{MW_{tot}} \times \frac{\rho_{tot}}{\rho_{1}} \times 1,000,000 $$* multiply by 100 for mass, mole or volume %.Mass 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percent1. A solution is prepared by adding 3 g of a substance A to 17 g of water. Calculate the mass per cent of the solute.Solution$Weight \; percent= \frac {mass \; of \; the \; solute}{mass \; of \; the \; solution} \times 100$$=\frac {3}{3+17} \times 100$=15%2. How many grams of 5.0% by weight NaCL solutions are required to 6.4 g NaCL?SolutionA 5% by weight NaCL solutions contains 5 g of NaCL in 100 g of solutionSo 1 g of NaCL will be contained in = 100/5 =20 g of solutionTherefore 6.4 g of NaCL will be present in = 20 X 6.4 = 128 g of solutionAlternate Method$Weight \; percent= \frac {mass \; of \; the \; solute}{mass \; of \; the \; solution} \times 100$$5= \frac {6.4}{x} \times 100$or x = 128 gHow to convert Weight percent to mole fractionNow if weight percent of the solution is known, we can calculate the mole fraction of the solution or the solvent easily as explained with below exampleExampleCalculate the mole fraction of HCl and H2O in a solution of HCL acid in water, containing 20% HCl by weight.SolutionThe solution contains 20 grams of HCL acid and 80 grams of water. Also,Molar mass of HCl is 36.5 grams/mole.Molar mass of water is 18 grams/mole.Moles of HCl = 20/36.5 = .547 moles of HClMoles of water = 80/18 = 4.44 moles of H2O. Share copy and redistribute the material in any medium or format for any purpose, even commercially. Adapt remix, transform, and build upon the material for any purpose, even commercially. The licensor cannot revoke these freedoms as long as you follow the license terms. Attribution You must give appropriate credit , provide a link to the license, and indicate if changes were made . You may do so in any reasonable manner, but not in any way that suggests the licensor endorses you or your use. ShareAlike If you remix, transform, or build upon the material, you must distribute your contributions under the same license as the original. No additional restrictions You may not apply legal terms or technological measures that legally restrict others from doing anything the license permits. You do not have to comply with the license for elements of the material in the public domain or where your use is permitted by an applicable exception or limitation . No warranties are given. The license may not give you all of the permissions necessary for your intended use. For example, other rights such as publicity, privacy, or moral rights may limit how you use the material. Skills to Develop Calculate formula masses for covalent and ionic compounds Define the amount unit mole and the related quantity Avogadros number Explain the relation between mass, moles, and numbers of atoms or molecules, and perform calculations deriving these quantities from one another Compute the percent composition of a compoundWe can argue that modern chemical science began when scientists started exploring the quantitative as well as the qualitative aspects of chemistry. For example, Daltons atomic theory was an attempt to explain the results of measurements that allowed him to calculate the relative masses of elements combined in various compounds. Understanding the relationship between the masses of atoms and the chemical formulas of compounds allows us to quantitatively describe the composition of substances. Formula Mass In an earlier chapter, we described the development of the atomic mass unit, the concept of average atomic masses, and the use of chemical formulas to represent the elemental makeup of substances. These ideas can be extended to calculate the formula mass of a substance by summing the average atomic masses of all the atoms represented in the substances formula. For covalent substances, the formula represents the numbers and types of atoms composing a single molecule of the substance; therefore, the formula mass may be correctly referred to as a molecular mass. Consider chloroform (CHCl3), a covalent compound once used as a surgical anesthetic and now primarily used in the production of tetrafluoroethylene, the building block for the anti-stick polymer, Teflon. The molecular formula of chloroform indicates that a single molecule contains one carbon atom, one hydrogen atom, and three chlorine atoms. The average molecular mass of a chloroform molecule is therefore equal to the sum of the average atomic masses of these atoms. Figure \(\PageIndex{1}\) outlines the calculations used to derive the molecular mass of chloroform, which is 119.37 amu. Figure \(\PageIndex{1}\): The average mass of a chloroform molecule, CHCl3, is 119.37 amu, which is the sum of the average atomic masses of each of its constituent atoms. The model shows the molecular structure of chloroform. Likewise, the molecular mass of an aspirin molecule, C9H8O4, is the sum of the atomic masses of nine carbon atoms, eight hydrogen atoms, and four oxygen atoms, which amounts to 180.15 amu (Figure \(\PageIndex{2}\)). Figure \(\PageIndex{2}\): The average mass of an aspirin molecule is 180.15 amu. The model shows the molecular structure of aspirin, C9H8O4. Example \(\PageIndex{1}\): Computing Molecular Mass for a Covalent Compound Ibuprofen, C13H18O2, is a covalent compound and the active ingredient in several popular nonprescription pain medications, such as Advil and Motrin. What is the molecular mass (amu) for this compound? Solution Molecules of this compound are comprised of 13 carbon atoms, 18 hydrogen atoms, and 2 oxygen atoms. Following the approach described above, the average molecular mass for this compound is therefore: Exercise \(\PageIndex{1}\) Acetaminophen, C8H9NO2, is a covalent compound and the active ingredient in several popular nonprescription pain medications, such as Tylenol. What is the molecular mass (amu) for this compound? Answer 151.16 amu Ionic compounds are composed of discrete cations and anions combined in ratios to yield electrically neutral bulk matter. The formula mass for an ionic compound is calculated in the same way as the formula mass for covalent compounds: by summing the average atomic masses of all the atoms in the compounds formula. Keep in mind, however, that the formula for an ionic compound does not represent the composition of a discrete molecule, so it may not correctly be referred to as the molecular mass. As an example, consider sodium chloride, NaCl, the chemical name for common table salt. Sodium chloride is an ionic compound composed of sodium cations, Na+, and chloride anions, Cl, combined in a 1:1 ratio. The formula mass for this compound is computed as 58.44 amu (Figure \(\PageIndex{3}\)). Figure \(\PageIndex{3}\): Table salt, NaCl, contains an array of sodium and chloride ions combined in a 1:1 ratio. Its formula mass is 58.44 amu. Note that the average masses of neutral sodium and chlorine atoms were used in this computation, rather than the masses for sodium cations and chlorine anions. This approach is perfectly acceptable when computing the formula mass of an ionic compound. Even though a sodium cation has a slightly smaller mass than a sodium atom (since it is missing an electron), this difference will be offset by the fact that a chloride anion is slightly more massive than a chloride atom (due to the extra electron). Moreover, the mass of an electron is negligibly small with respect to the mass of a typical atom. Even when calculating the mass of an isolated ion, the missing or additional electrons can generally be ignored, since their contribution to the overall mass is negligible, reflected only in the nonsignificant digits that will be lost when the computed mass is properly rounded. The few exceptions to this guideline are very light ions derived from elements with precisely known atomic masses. Example \(\PageIndex{2}\): Computing Formula Mass for an Ionic Compound Aluminum sulfate, Al2(SO4)3, is an ionic compound that is used in the manufacture of paper and in various water purification processes. What is the formula mass (amu) of this compound? Solution The formula for this compound indicates it contains Al3+ and SO42 ions combined in a 2:3 ratio. For purposes of computing a formula mass, it is helpful to rewrite the formula in the simpler format, Al2S3O12. Following the approach outlined above, the formula mass for this compound is calculated as follows: Exercise \(\PageIndex{2}\) Calcium phosphate, Ca3(PO4)2, is an ionic compound and a common anti-caking agent added to food products. What is the formula mass (amu) of calcium phosphate? Answer 310.18 amu In units 1 and2, we learned about measuring quanity of atoms using the mole. As a reminder, the mole is an amount unit similar to familiar units like pair, dozen, gross, etc.The number of entities composing a mole has been experimentally determined to be \(6.02214179 \times 10^{23}\), a fundamental constant named Avogadros number (NA) or the Avogadro constant in honor of Italian scientist Amedeo Avogadro. This constant is properly reported with an explicit unit of per mole, a conveniently rounded version being \(6.022 \times 10^{23}/\ce{mol}\). The relationships between formula mass, the mole, and Avogadros number can be applied to compute various quantities that describe the composition of substances and compounds. In unit 2, we used Avogadros number to convert between mass of an element, moles of an element,and number of atoms in the sample. Now, we can useAvogadros number and formula mass (also called molecular weight)to convent between mass of a sample, moles, and number of molecules.For example, if we know the mass and chemical composition of a substance, we can determine the number of moles and calculate number of atoms or molecules in the sample. Likewise, if we know the number of moles of a substance, we can derive the number of atoms or molecules and calculate the substances mass.sExample \(\PageIndex{7}\): Deriving Moles from Grams for a Compound Our bodies synthesize protein from amino acids. One of these amino acids is glycine, which has the molecular formula C2H5O2N. How many moles of glycine molecules are contained in 28.35 g of glycine? Solution We can derive the number of moles of a compound from its mass following the same procedure we used for an element in Example \(\PageIndex{6}\): The molar mass of glycine is required for this calculation, and it is computed in the same fashion as its molecular mass. One mole of glycine, C2H5O2N, contains 2 moles of carbon, 5 moles of hydrogen, 2 moles of oxygen, and 1 mole of nitrogen: The provided mass of glycine (~28 g) is a bit more than one-third the molar mass (~75 g/mol), so we would expect the computed result to be a bit greater than one-third of a mole (~0.33 mol). Dividing the compounds mass by its molar mass yields: \[\mathrm{28.35\:\cancel{g}\:glycine\left(\dfrac{mol\: glycine}{75.07\:\cancel{g}}\right)=0.378\:mol\: glycine} onumber\] This result is consistent with our rough estimate. Exercise \(\PageIndex{7}\) How many moles of sucrose, \(C_{12}H_{22}O_{11}\), are in a 25-g sample of sucrose? Answer 0.073 mol Example \(\PageIndex{8}\): Deriving Grams from Moles for a Compound Vitamin C is a covalent compound with the molecular formula C6H8O6. The recommended daily dietary allowance of vitamin C for children aged 48 years is 1.42 104 mol. What is the mass of this allowance in grams? Solution As for elements, the mass of a compound can be derived from its molar amount as shown: The molar mass for this compound is computed to be 176.124 g/mol. The given number of moles is a very small fraction of a mole (~104 or one-ten thousandth); therefore, we would expect the corresponding mass to be about one-ten thousandth of the molar mass (~0.02 g). Performing the calculation, we get: \[\mathrm{1.42\times10^{-4}\:\cancel{mol}\:vitamin\: C\left(\dfrac{176.124\:g}{\cancel{mol}\:vitamin\: C}\right)=0.0250\:g\: vitamin\: C} onumber\] This is consistent with the anticipated result.Exercise \(\PageIndex{8}\) What is the mass of 0.443 mol of hydrazine, \(N_2H_4\)? Answer 14.2 g Example \(\PageIndex{9}\): Deriving the Number of Molecules from the Compound Mass A packet of an artificial sweetener contains 40.0 mg of saccharin (C7H5NO3S), which has the structural formula: Given that saccharin has a molar mass of 183.18 g/mol, how many saccharin molecules are in a 40.0-mg (0.0400-g) sample of saccharin? How many carbon atoms are in the same sample? Solution The number of molecules in a given mass of compound is computed by first deriving the number of moles, as demonstrated in Example \(\PageIndex{8}\), and then multiplying by Avogadros number: Using the provided mass and molar mass for saccharin yields: \[\mathrm{0.0400\:\cancel{g}\:\ce{C7H5NO3S}\left(\dfrac{\cancel{mol}\:\ce{C7H5NO3S}}{183.18\:\cancel{g}\:\ce{C7H5NO3S}}\right)\left(\dfrac{6.022\times10^{23}\:\ce{C7H5NO3S}\:molecules}{1\:\cancel{mol}\:\ce{C7H5NO3S}}\right)}\\=\mathrm{1.31\times10^{20}\:\ce{C7H5NO3S}\:molecules}\] The compounds formula shows that each molecule contains seven carbon atoms, and so the number of C atoms in the provided sample is: \[\mathrm{1.31\times10^{20}\:\ce{C7H5NO3S}\: molecules\left(\dfrac{7\:C\: atoms}{1\:\ce{C7H5NO3S}\: molecule}\right)=9.20\times10^{21}\:C\: atoms} onumber\]Exercise \(\PageIndex{9}\) How many \(C_4H_{10}\) molecules are contained in 9.213 g of this compound? How many hydrogen atoms? Answer \(9.545 \times 10^{22}\; \text{molecules}\; C_4H_{10}\) \(9.545 \times 10^{23 }\;\text{atoms}\; H\) Video \(\PageIndex{3}\): A preview of some of the uses we will have for moles in upcoming units In the previous section, we discussed the relationship between the bulk mass of a substance and the number of atoms or molecules it contains (moles). Given the chemical formula of the substance, we were able to determine the amount of the substance (moles) from its mass, and vice versa. But what if the chemical formula of a substance is unknown? In this section, we will explore how to apply these very same principles in order to derive the chemical formulas of unknown substances from experimental mass measurements. The elemental makeup of a compound defines its chemical identity, and chemical formulas are the most succinct way of representing this elemental makeup. When a compounds formula is unknown, measuring the mass of each of its constituent elements is often the first step in the process of determining the formula experimentally. The results of these measurements permit the calculation of the compounds percent composition, defined as the percentage by mass of each element in the compound. For example, consider a gaseous compound composed solely of carbon and hydrogen. The percent composition of this compound could be represented as follows: \[\mathrm{\%H=\dfrac{mass\: H}{mass\: compound}\times100\%}\] \[\mathrm{\%C=\dfrac{mass\: C}{mass\: compound}\times100\%}\] If analysis of a 10.0-g sample of this gas showed it to contain 2.5 g H and 7.5 g C, the percent composition would be calculated to be 25% H and 75% C: \[\mathrm{\%H=\dfrac{2.5\:g\: H}{10.0\:g\: compound}\times100\%=25\%}\] \[\mathrm{\%C=\dfrac{7.5\:g\: C}{10.0\:g\: compound}\times100\%=75\%}\] Example \(\PageIndex{10}\): Calculation of Percent Composition Analysis of a 12.04-g sample of a liquid compound composed of carbon, hydrogen, and nitrogen showed it to contain 7.34 g C, 1.85 g H, and 2.85 g N. What is the percent composition of this compound? Solution To calculate percent composition, we divide the experimentally derived mass of each element by the overall mass of the compound, and then convert to a percentage: \[\mathrm{\%C=\dfrac{7.34\:g\: C}{12.04\:g\: compound}\times100\%=61.0\%} onumber\] \[\mathrm{\%H=\dfrac{1.85\:g\: H}{12.04\:g\: compound}\times100\%=15.4\%} onumber\] \[\mathrm{\%N=\dfrac{2.85\:g\: N}{12.04\:g\: compound}\times100\%=23.7\%} onumber\] The analysis results indicate that the compound is 61.0% C, 15.4% H, and 23.7% N by mass.Exercise \(\PageIndex{10}\) A 24.81-g sample of a gaseous compound containing only carbon, oxygen, and chlorine is determined to contain 3.01 g C, 4.00 g O, and 17.81 g Cl. What is this compounds percent composition? Answer 12.1% C, 16.1% O, 71.8% Cl Video \(\PageIndex{4}\): A video overview of how to calculate percent composition of a compound based on its chemical formula. Percent composition is also useful for evaluating the relative abundance of a given element in different compounds of known formulas. As one example, consider the common nitrogen-containing fertilizers ammonia (NH3), ammonium nitrate (NH4NO3), and urea (CH4N2O). The element nitrogen is the active ingredient for agricultural purposes, so the mass percentage of nitrogen in the compound is a practical and economic concern for consumers choosing among these fertilizers. For these sorts of applications, the percent composition of a compound is easily derived from its formula mass and the atomic masses of its constituent elements. A molecule of NH3 contains one N atom weighing 14.01 amu and three H atoms weighing a total of (3 1.008 amu) = 3.024 amu. The formula mass of ammonia is therefore (14.01 amu + 3.024 amu) = 17.03 amu, and its percent composition is: \[\mathrm{\%N=\dfrac{14.01\:amu\: N}{17.03\:amu\:NH_3}\times100\%=82.27\%}\] \[\mathrm{\%H=\dfrac{3.024\:amu\: N}{17.03\:amu\:NH_3}\times100\%=17.76\%}\] This same approach may be taken considering a pair of molecules, a dozen molecules, or a mole of molecules, etc. The latter amount is most convenient and would simply involve the use of molar masses instead of atomic and formula masses, as demonstrated Example \(\PageIndex{2}\). As long as we know the chemical formula of the substance in question, we can easily derive percent composition from the formula mass or molar mass. Example \(\PageIndex{11}\): Determining Percent Composition from a Molecular Formula Aspirin is a compound with the molecular formula C9H8O4. What is its percent composition? Solution To calculate the percent composition, we need to know the masses of C, H, and O in a known mass of C9H8O4. It is convenient to consider 1 mol of C9H8O4 and use its molar mass (180.159 g/mole, determined from the chemical formula) to calculate the percentages of each of its elements: \[\begin{align*}\%\ce C&=\mathrm{\dfrac{9\:mol\: C\times molar\: mass\: C}{molar\: mass\:\ce{C9H18O4}}}\times100=\dfrac{9\times12.01\:g/mol} onumber{180.159\:g/mol}\times100} onumber\\\%\ce H&=\mathrm{\dfrac{108.09\:g/mol}{180.159\:g/mol}\times100} onumber\\\%\ce C&=\mathrm{60.00\,\%\,C} onumber\end{align*}\] \[\begin{align*}\%\ce H&=\mathrm{\dfrac{8\:mol\: H\times molar\: mass\: H}{molar\: mass\:\ce{C9H18O4}}}\times100=\dfrac{8\times1.008\:g/mol} onumber {180.159\:g/mol}\times100=\dfrac{8.064\:g/mol}{180.159\:g/mol}\times100} onumber\\\%\ce H&=4.476\,\%\,\ce Honumber\end{align*}\] \[\begin{align*}\%\ce O&=\mathrm{\dfrac{4\:mol\: O\times molar\: mass\: O}{molar\: mass\: \ce{C9H18O4}}}\times100=\dfrac{4\times16.00\:g/mol} onumber{180.159\:g/mol}\times100=\dfrac{64.00\:g/mol}{180.159\:g/mol}\times100} onumber \\\%\ce O&=35.52\% onumber\end{align*}\] Note that these percentages sum to equal 100.00% when appropriately rounded.Exercise \(\PageIndex{11}\) To three significant digits, what is the mass percentage of iron in the compound \(Fe_2O_3\)? Answer 69.9% Fe The formula mass of a substance is the sum of the average atomic masses of each atom represented in the chemical formula and is expressed in atomic mass units. The formula mass of a covalent compound is also called the molecular mass. A convenient amount unit for expressing very large numbers of atoms or molecules is the mole. Experimental measurements have determined the number of entities composing 1 mole of substance to be 6.022 1023, a quantity called Avogadros number. The mass in grams of 1 mole of substance is its molar mass. Due to the use of the same reference substance in defining the atomic mass unit and the mole, the formula mass (amu) and molar mass (g/mol) for any substance are numerically equivalent (for example, one H2O molecule weighs approximately 18 amu and 1 mole of H2O molecules weighs approximately 18 g). Footnotes 1 Omiatek, Donna M., Amanda J. Bressler, Ann-Sofie Cans, Anne M. Andrews, Michael L. Heien, and Andrew G. Ewing. The Real Catecholamine Content of Secretory Vesicles in the CNS Revealed by Electrochemical Cytometry. Scientific Report 3 (2013): 1447, accessed January 14, 2015, doi:10.1038/srep01447. Read more about the redefinition of SI units including the kilogram here (Laura Howe, CE&N, Nov. 16, 2018). Avogadros number (NA) experimentally determined value of the number of entities comprising 1 mole of substance, equal to 6.022 1023 mol1 formula mass sum of the average masses for all atoms represented in a chemical formula; for covalent compounds, this is also the molecular mass mole amount of substance containing the same number of atoms, molecules, ions, or other entities as the number of atoms in exactly 12 grams of 12C molar mass mass in grams of 1 mole of a substance percent composition percentage by mass of the various elements in a compound Contributors Paul Flowers (University of North Carolina - Pembroke),Klaus Theopold (University of Delaware) andRichard Langley (Stephen F. Austin State University) with contributing authors.Textbook content produced by OpenStax College is licensed under a Creative Commons Attribution License 4.0 license. Download for free at 9.110). Adelaide Clark, Oregon Institute of Technology Crash Course Chemistry: Crash Course is a division of Complexly and videos are free to stream for educational purposes. TED-Eds commitment to creating lessons worth sharing is an extension of TEDs mission of spreading great ideas. Within TED-Eds growing library of TED-Ed animations, you will find carefully curated educational videos, many of which represent collaborations between talented educators and animators nominated through the TED-Ed website. Fuse School, Open Educational Resource free of charge, under a Creative Commons License: Attribution-NonCommercial CC BY-NC (View License Deed: 5) Determine the molarity of the solution: 17.9710 mol / 1.000 L = 17.97 M (to four sig figs) 6) Determine the molality of the solution: 73.44 g = 0.07344 kg 17.9710 mol / 0.07344 kg = 244.7 m Scorpio reacts. Solution assuming a certain mass of the solution is present: 1) Assume 100.0 g of the solution is present. Determine the mass of each solution component: H2SO4 ---> (100.0 g) (0.9600) = 96.00 gH2O ---> (100.0 g) (0.0400) = 4.00 g 2) Determine the mole fraction: H2SO4 ---> 96.00 g / 98.078 g/mol = 0.978813 molH2O ---> 4.00 g / 18.015 g/mol = 0.222037 mol 0.978813 mol + 0.222037 mol = 1.20085 mol H2SO4 ---> 0.978813 mol / 1.20085 mol = 0.8151H2O ---> 0.222037 mol / 1.20085 mol = 0.1849 or, the H2O can be obtained by subtraction: 1 0.8151 = 0.1849 3) Determine the molality: 4.00 g = 0.00400 kg 0.978813 mol / 0.00400 kg = 244.7 m 4) Determine the molarity: 100.0 g / 1.836 g/mL = 54.46623 mL = 0.05446623 L 0.978813 mol / 0.05446623 L = 17.97 M 5) By the way, you could consider this solution to be some water (the solute) dissolved in some sulfuric acid (the solvent). Calculate the molality of the water: 1762.56 g = 1.76256 kg 4.0766 mol / 1.76256 kg = 2.313 m Example #2: Given a density of 1.769 g/mL, and a H2SO4 mole fraction of 0.5000, find the molality, molarity, and mass percent. Solution: 1) We will use a mole fraction of 0.5000 to mean 0.5000 mole is present in a total of 1.0000 mole of solution. Determine the mass of each mole fraction: H2SO4 ---> (0.5000 mol) (98.078 g/mol) = 49.039 gH2O ---> (0.5000 mol) (18.015 g/mol) = 9.0075 g Comment: A mole fraction of 0.50 could mean 1.0 mol of one component in 2.0 total moles. Using 0.5 and 1 is the simplest meaning of a mole fraction of 0.5. We could have used any paring of numbers that gives a mole fraction of 0.5. The final answers would be the same, but the numbers in the calculations would be different. 2) Determine the mass percent of each component: 49.039 g + 9.0075 g = 58.0465 g H2SO4 ---> 49.039 g / 58.0465 g = 84.48%H2O ---> 9.0075 g / 58.0465 g = 15.52% Often the last mass percent is obtained by subtraction: H2O ---> 100.00 84.48 = 15.52 3) Determine the molality: 9.0075 g = 0.0090075 kg 0.5000 mol / 0.0090075 kg = 55.51 m 4) Determine the molarity: 58.0465 g / 1.769 g/mL = 32.81 mL = 0.03281 L 0.5000 mol / 0.03281 L = 15.24 M Example #3: Given a density of 1.059 g/mL and a H2SO4 molarity of 1.000 M, find the molality, mole fraction, and mass percent. Solution: 1) Assume 1.0000 L of the solution is present. Determine the mass of the solution: (1.059 g/mL) (1000.0 mL) = 1059 g 2) Determine the mass percent of each component: H2SO4 ---> 98.078 g (remember, it's a 1 M solution)H2O ---> 1059 g 98.078 g = 960.922 g H2SO4 ---> 100 90.74 = 9.26%H2O ---> 960.922 g / 1059 g = 90.74% Note that I calculated the larger value by division and the smaller value by subtraction. 3) Determine the mole fraction: H2SO4 ---> 98.078 g / 98.078 g/mol = 1.000 mol (or just remember it's a 1 M solution)H2O ---> 960.922 g / 18.015 g/mol = 53.349 mol 53.349 mol + 1.000 mol = 54.349 mol H2SO4 ---> 1 0.9816 = 0.0184H2O ---> 53.349 mol / 54.349 mol = 0.9816 4) Determine the molality: 960.922 g = 0.960922 kg 1.000 mol / 0.960922 kg = 1.041 m Example #4: Given a density 1.122 g/mL and a H2SO4 molality of 4.500 m, find the molarity, mole fraction and mass percent. Solution: 1) The given molality means 4.500 mol dissolved in 1.000 kg of water. Determine the mass of each component: H2SO4 ---> (4.500 mol) (98.078 g/mol) = 441.351 gH2O ---> 1.000 kg = 1000. g 2) Determine mass percentages: 1000. g + 441.351 g = 1441.351 g (total mass of the solution) H2O ---> 1000. g / 1441.351 g = 69.38%H2SO4 ---> 100 69.38 = 30.62% 3) Determine mole fraction: H2SO4 ---> 4.500 molH2O ---> 1000. g / 18.015 g/mol = 55.509 mol 55.509 mol + 4.500 mol = 60.009 mol H2SO4 ---> 1 0.9250 = 0.0750H2O ---> 55.509 mol / 60.009 mol = 0.9250 4) Determine the molarity: 1441.351 g / 1.122 g/mL = 1296.179 mL = 1.296179 L 4.500 mol / 1.296179 L = 3.472 M Example #5: Calculate the percent cadmium(II) bromide by mass in an aqueous solution with a solute mole fraction of 0.114 and a density of 1.047 g/mL. Solution: 1) Determine mole fraction of the water: 1 0.114 = 0.886 2) Assume a solution with 0.114 mol of CdBr2 and 0.886 mol of water is present. Determine the mass of each solution component: CdBr2 ---> (0.114 mol) (272.218 g/mol) = 31.033 g H2O ---> (0.886 mol) (18.015 g/mol) = 15.889 g 3) The mass percent of CdBr2 is: [31.033 g / (31.033 g + 15.889 g)] * 100 = 66.1% The solution density is not required. Notice that the density is given in each of the above examples. This provides a necessary bridge between the volume-based concentration unit of molarity and the other concentration units (molality, mole fraction and mass percent) none of which uses volume in its definition. If you were only given two of these: (1) mass percent or (2) molality or (3) mole fraction (and no density), you could not get the molarity (or the density). However, suppose you are given one of these: (1) mass percent or (2) molality or (3) mole fraction and then also given molarity instead of the density. Could you get to the density and the other units? In other words, can we swap density and molarity in the given part of the problem? Let's see . . . . Example #6: Reagent grade nitric acid is 70.40% HNO3 (63.0119 g/mol) by mass and its molarity is 16.00 M. Calculate the density, molality and mole fraction of nitric acid in the solution. Solution: 1) 16.0 molar means this: 16.0 moles of HNO31.000 L (which equals 1000. mL) of solution. 2) The key point is that the 16.00 moles of HNO3 is 70.40% of the entire mass of the 1000. mL of solution. (63.0119 g/mol) (16.0 mol) = 1008.19 g 1008.19 g / 0.7040 = 1432 g (this is the total mass of the solution) 3) The density is: 1432 g / 1000. mL = 1.432 g/mL 4) Molality: 1432 g 1008.19 g = 423.81 g = 0.42381 kg 16.00 mol / 0.42381 kg = 37.75 m 5) Mole fraction 423.81 g / 18.015 g/mol = 23.5254 mol (of water) 16.00 mol + 23.5254 mol = 39.5254 mol (total moles) 23.5254 mol / 39.5254 mol = 0.5952 (mole fraction of water) 1 0.5952 = 0.4048 (mole fraction of the nitric acid) Here's the question again: . . . can we swap density and molarity in the given part of the problem? And the answer is a very firm YES. Let's do another . . . . Example #7: Reagent grade nitric acid (HNO3, MW = 63.0119 g/mol) has a molarity of 16.00 M and its molality is 37.75 m. Calculate the density, mass percent and mole fraction of nitric acid in the solution. Solution: 1) Use the molality: 37.75 mol (of HNO3)1.000 kg of solvent (water) 2) Compute the mass of the above solution: (37.75 mol) (63.0119 g/mol) = 2378.7 g 1000 g + 2378.7 g = 3

to division of both numerator and denominator by the sum of molar amounts of components. This property has consequences for representations of phase diagrams using, for instance, ternary plots.Mixing binary mixtures with a common component gives a ternary mixture with certain mixing ratios between the three components. These mixing ratios from the ternary and the corresponding mole fractions of the ternary mixture x1(123), x2(123), x3(123) can be expressed as a function of several mixing ratios involved, the mixing ratios between the components of the binary mixtures and the mixing ratio of the binary mixtures to form the ternary one. x 1 (123) = n (12) x 1 (12) + n 13 x 1 (13) n (12) + n (13) {\displaystyle x_{1(123)}={\frac {n_{(12)}x_{1(12)}+n_{13}x_{1(13)}}{n_{(12)}+n_{(13)}}}} Multiplying mole fraction by 100 gives the mole percentage, also referred as amount/amount percent [abbreviated as (n/n)% or mol%].The conversion to and from mass concentration i is given by: x i = i M M i i = x i M i M {\displaystyle {\begin{aligned}x_{i}&={\frac {\rho _{i}}{\rho }}{\frac {\bar {M}}{M_{i}}}\\[3pt]\Leftrightarrow \rho _{i}&=x_{i}\rho {\frac {M_{i}}{\bar {M}}}\end{aligned}}} where M is the average molar mass of the mixture.The conversion to molar concentration ci is given by: c i = x i c = x i M = x i j x j M j {\displaystyle {\begin{aligned}c_{i}&=x_{i}c\\[3pt]&={\frac {x_{i}\rho }{\bar {M}}}={\frac {x_{i}\rho }{\sum _{j}x_{j}M_{j}}}\end{aligned}}} where M is the average molar mass of the solution, c is the total molar concentration and is the density of the solution.The mole fraction can be calculated from the masses mi and molar masses Mi of the components: x i = m i M i j m j M j {\displaystyle x_{i}={\frac {\frac {m_{i}}{M_{i}}}{\sum _{j}{\frac {m_{j}}{M_{j}}}}}} In a spatially non-uniform mixture, the mole fraction gradient triggers the phenomenon of diffusion.^ a b c IUPAC, Compendium of Chemical Terminology, 5th ed. (the "Gold Book") (2025). Online version: (2006) "amount fraction". doi:10.1351/goldbook.A00296^ Zumdahl, Steven S. (2008). Chemistry (8thed.). Cengage Learning. p.201. ISBN978-0-547-12532-9.^ Rickard, James N.; Spencer, George M.; Bodner, Lyman H. (2010). Chemistry: Structure and Dynamics (5thed.). Hoboken, N.J.: Wiley. p.357. ISBN978-0-470-58711-9.^ a b "ISO 80000-9:2019 Quantities and units Part 9: Physical chemistry and molecular physics". ISO. 2013-08-20. Retrieved 2023-08-29.^ "SI Brochure". BIPM. Retrieved 2023-08-29.^ a b Thompson, A.; Taylor, B. N. (2 July 2009). "The NIST Guide for the use of the International System of Units". National Institute of Standards and Technology. Retrieved 5 July 2014.Retrieved from " fraction is the number of moles of a component divided by the total number of moles. Mole percent is this value multiplied by 100%.In chemistry, the mole fraction is a unit of concentration that is the number of moles of a component divided by the total number of moles of a solution or mixture. The mole fraction is a dimensionless number. The sum of all of the mole fractions equals 1. The symbol for mole fraction is the capital letter X or the lowercase Greek letter chi (). The terms amount fraction or amount-of-substance fraction mean the same as mole fraction.The formula for mole fraction is the moles of one component divided by the total number of moles:XA = moles A / total molesFor example, in a mixture consisting of 0.25 moles of component A and 0.40 moles of component B, you can find the mole fractions of A and B.XA = moles A / total moles = 0.25 / (0.25 + 0.40) = 0.38 (rounded)XB = moles B / total moles = 0.40 / (0.25 + 0.40) = 0.62 (rounded)Remember, adding up the mole fractions equals 1.XA + XB = 10.38 + 0.62 = 1If the mixture consists of more than two components, the same rules apply.A related term is mole percent. Mole percent or mole percentage is the mole fraction multiplied by 100%.mol% = XA x 100%The sum of all of the mole percents of a mixture equals 100%The mole fraction offers advantages over some of the other units of concentration.Unlike molarity, mole fraction is not temperature dependent.Preparing a solution using mole fraction is easy because you simply weigh the masses of the components and then combine them.There is no confusion over which component is the solvent and which is the solute. The unit is symmetric in this respect because the roles of solute and solvent are reversible, depending on the mole fraction.In a mixture of ideal gases or most real gases, the mole fraction is the same as the ratio of partial pressure of a gas to the total pressure of the mixture. In other words, mole fraction follows Daltons law of partial pressure.For example, find the mole fraction of carbon tetrachloride in a mixture consisting of 1 mole benzene, 2 moles carbon tetrachloride, and 7 moles acetone. XA = moles A / total molesXCCl4 = 2 / (1 + 2 + 7) = 2/10 = 0.2Find the mole fraction of formaldehyde (CH2O) when you dissolve 25.7 grams of CH2O in 3.25 moles of carbon tetrachloride (CCl4).Here, the amount of CCl4 is already in moles, but you cant find mole fraction until you convert grams of CH2O into moles, too. Look up the atomic masses of carbon, hydrogen, and oxygen on the periodic table and use the chemical formula for formaldehyde to calculate the number of moles.1 mole CH2O = 12.01 g + 21.01 g + 16.00 g = 30.03 gUse this relationship and find the number of moles of CH2O.moles CH2O = 25.7 g x (1 mol/30.03 g) = 0.856 molNow, solve for mole fraction. XA = moles A / total molesXA = 0.856 moles CH2O / (0.856 moles CH2O + 3.25 moles CCl4) = 0.208Molality (m) is the moles of solute per kilogram of solvent. Using these units, you can calculate mole fraction if you know molality. For example, find the mole fraction of table sugar or sucrose (C6H12O6) in a 1.62 m solution of sucrose in water.Given the definition of molality, you know the following:1.2 m sucrose = 1.62 moles sucrose / 1 kg waterNext, find how many moles there are of water. Use the atomic masses from the periodic table and find that the molar mass of water is 18.0 (21.01 + 16.00).1 kg = 1000 g = 1 mol / 18.0 g = 55.5 moles H2OKnowing the moles of sucrose and the moles of water, find the mole fraction of sucrose.XA = moles A / total molesXsucrose = moles sucrose / total moles = 1.62 / (1.62 + 55.5) = 0.0284With small numbers like this, its often better expressing mole fraction as mole percent. The solution is 2.84% sugar in water.IUPAC (1997). Amount fraction. Compendium of Chemical Terminology (the Gold Book) (2nd ed.). Blackwell Scientific Publications. ISBN 0-9678550-9-8. doi:10.1351/goldbook.A00296Rickard, James N.; Spencer, George M.; Bodner, Lyman H. (2010). Chemistry: Structure and Dynamics (5th ed.). Hoboken, N.J.: Wiley. ISBN 978-0-470-58711-9.Thompson, A.; Taylor, B. N. (2009). Special Publication 811. The NIST Guide for the use of the International System of Units. National Institute of Standards and Technology.Zumdahl, Steven S. (2008). Chemistry (8th ed.). Cengage Learning. ISBN 978-0-547-12532-9.Related Posts The calculator below can be used to convert between common concentration units. The mole weight of the solvent and solute and the density of the solution must be filled in: Density of water solutions of: organic acids, organic substances as sugars and alcohols, inorganic chlorides, inorganic sodium salts, inorganic potassium salts and some other inorganic substances . See also Solutions, molarity and dilution, Mixtures, Solutions and Suspensions and Solubility product constants. The fomulas used by the calculater is given in the table below:For full table - rotate the screen! Molality, Molarity and Mole fractions Desired unit - as function of Wt% Molality Molarity Mole fraction g solute/l solution Wt% of solute 100*E*B 1000 + E*B F*B 10*C 100*G*B G*B + (1 - G)*A H 10*C Molality, mol/kg solvent 1000*D B*(100 - D) 1000*F 1000*C - F*B 1000*G A - G*A 1000*H B*(1000*C - H) Molarity, mol/liter solution 10*C*D B 1000*C*E 1000 + B*E 1000*C*G G*B + (1 - G)*A H B Mole fraction, molsolute /moltotal D/B D/B + (100 - D)/A A*E A*E + 1000 A*F F*(A - B) + 1000*C A*H H*(A - B) + 1000*C*B Gram of solute/liter solution 10*D*C 1000*C*E*B 1000 + E*B F*B 1000*C*G*B G*B + (1-G)*A where A = mole weight of solvent, g/mol B = mole weight of solute, g/mol C = density of solution*, g/ml D = wt% of solute E = molality, mol/kg solvent F = molarity, mol/liter solution G = mole fraction, molsolute /moltotal H = gram of solute/liter solution* Density is not needed for conversons between wt%, mole fraction and molality or between molarity and grams of solute per liter solution. Density of low viscous liquids can be measured by exact weighing of a well-defined volume of the liquid. Then, the density is calculated by using the formula density = mass (g) / volume (ml)Density of water solutions of: organic acids, sugar and alchohols, inorganic chlorides, inorganic potassium salts, inorganic sodium salts, other common inorganic substanses Shrinkage and estimation of density of a liquid-liquid solution How to measure density of petroleum products Descriptive Terms for SolubilityDescriptive Terms for SolubilityDescriptive TermParts of solvent required for 1 part of soluteVery soluble< 1Freely soluble1 < 10Soluble10 < 30Sparingly soluble30 < 100Slightly soluble100 < 1000Very slightly soluble1000 < 10000Pracically insoluble (Insoluble)> 10000 An overview of common test methods and typical ranges of variation of petroleum quality parameters. What, why and how do the different test? Changes in density of aqueous solutions with changes in concentration at 20C. Density of inorganic chlorides in water is plotted as function of wt%, mol/kg water and mol/l solution. Changes in density of aqueous solutions with changes in concentration at 20C. Density of potassium salts in water is plotted as function of wt%, mol/kg water and mol/l solution. Changes in density of aqueous solutions with changes in concentration at 20C. Density of inorganic sodium salts in water is plotted as function of wt%, mol/kg water and mol/l solution. Changes in density of aqueous solutions with changes in concentration at 20C. Density of acetic acid, citric acid, formic acid, D-lactic acid, oxalic acid and trichloroacetic acid in water is plotted as function of wt%, mol/kg water and mol/l solution. Online density converter with commonly used units. Changes in density of aqueous solutions with changes in concentration at 20C. Density of some sugars, alcohols and other organic substances in water is plotted as function of wt%, mol/kg water and mol/l solution. Changes in density of aqueous solutions with changes in concentration at 20C. Density of inorganic substances in water is plotted as function of wt%, mol/kg water and mol/l solution. It is possible to estimate the density of a liquid-liquid solution from the density of the solute and the solvent. However, due to shrinkage, the estimate will be a bit too low. Mixtures vs. solutions vs. suspensions. An overview of common test methods and specifications of petroleum fuels. What, why and how do the different test? The equilibrium constant, Ksp, for aqueous solutions of ionic compounds at 25C. Definitions and examples of how to calculate wt%, molarity and how to prepare dilutions. Common converting units for Acceleration, Area, Density, Energy, Energy per unit mass, Force, Heat flow rate, Heat flux, Heat generation per unit volume and many more. Convert between units with the unit factor or factor-label methodMass fraction (chemistry). In chemistry, the mass fraction of a substance within a mixture is the ratio (alternatively denoted) of the mass of that substance to the total mass of the mixture. Expressed as a formula, the mass fraction is:Mass fraction can also be expressed, with a denominator of 100, as percentage by mass (in commercial contexts often called percentage by weight, abbreviated wt%; see mass versus weight). It is one way of expressing the composition of a mixture in a dimensionless size; mole fraction (percentage by moles, mol%) and volume fraction (percentage by volume, vol%) are others.Video advice: Mass and Mole Fractions (Conversions)An overview on how to change molar fractions to mass fractions and vice-versa. Percent concentration does not refer to this quantity. This improper name persists, especially in elementary textbooks. In biology, the unit % is sometimes (incorrectly) used to denote mass concentration, also called mass/volume percentage. A solution with 1g of solute dissolved in a final volume of 100mL of solution would be labeled as 1% or 1%m/v (mass/volume). This is incorrect because the unit % can only be used for dimensionless quantities. Instead, the concentration should simply be given in units of g/mL. Percent solution or percentage solution are thus terms best reserved for mass percent solutions (m/m, m%, or mass solute/mass total solution after mixing), or volume percent solutions (v/v, v%, or volume solute per volume of total solution after mixing). The very ambiguous terms percent solution and percentage solutions with no other qualifiers continue to occasionally be encountered.Difference Between Mole Fraction and Weight PercentThe key difference between mole fraction and weight percent is that mole fraction gives the composition of a substance in a mixture by means of moles while.Weight percent or mass percent may be the number of the mass of the component inside a mixture. It is called the mass fraction too. Further, this term describes the ratio between your mass from the preferred component and also the total mass from the mixture multiplied by 100 to obtain a percentage value. The formula with this determination is really as follows:What is Weight Percent?Mole fraction and weight percentage are two different ways of expressing the concentration of a component in a mixture. The key difference between mole fraction and weight percent is that the mole fraction gives the composition of a substance in a mixture by means of moles, while the weight percent gives the composition by means of mass. Furthermore, the value of mole fraction is given as a fraction or in one or two decimal places while weight percent is given as a percentage value.Molarity of a Percent by Weight Solution You can determine the concentration of solute in a solution as a weight-to-weight percentage, a weight-to-volume percentage or volume-to-volume percentage. In this context, weight is synonymous with mass, so a mass percentage means the relative weight of the solute to the weight of the solution, and you can express it as percent by weight. However, its also common to relate weight to volume and express the result as percent weight to volume. Either way, as long as you know the chemical formulas for the solute and solvent (which is usually water), the weight percentage allows you to calculate how many moles of solute are present in the solution. From this, its possible to determine the molarity of the solution, which is the number of moles of solute per liter of solution. TL;DR (Too Long; Didnt Read) If you know the percent by weight of a solution, you can find the weight of solute. Divide this by its molecular weight to find the number of moles and divide by the volume of solution to find the molarity.How to Calculate Mass Percent Composition These are worked example problems showing how to calculate mass percent composition. Examples include sodium bicarbonate, water, and carbon dioxide.The units of mass are typically grams. Mass percent is also known as percent by weight or w/w%. The molar mass is the sum of the masses of all the atoms in one mole of the compound. The sum of all the mass percentages should add up to 100%. Watch for rounding errors in the last significant figure to make sure all the percentages add up.mole fraction and mass percent Kristen A.Mole Fraction Calculator The mole fraction calculator provides you with the solution to your chemical questions using two different methods!Our mole fraction calculator is a handy tool that computes the mole fraction of your solutions using two different methods. It doesnt matter if you only know the total number of moles in a solution, or your task requires you to work with solvents and solutes were here for you to help!Textbook-specific videos for college studentsOur videos prepare you to succeed in your college classes. Let us help you simplify your studying. If you are having trouble with Chemistry, Organic, Physics, Calculus, or Statistics, we got your back! Our videos will help you understand concepts, solve your homework, and do great on your exams.Step One. Determine the moles of acetone within the solution. Were given a molality of just one. 00 m solution of acetone in ethanolAssume that theres 1 kg of solvent ethanol:1. 00molkg=molesacetone1kg1. 00molkg1kg=molesacetone1kg1kgmolesacetone=1. 00molacetoneStep 2. Calculate the mass of acetone (in g) from the solutionWe can calculate the mass of acetone using molar mass. The molar mass of acetone (CH3COCH3) is:3 C x 12. 01 g/mol C = 36. 03 g/mol6 H x 1. 008 g/mol H = 6. 048 g/mol1 O x 16. 00 g/mol O = 16. 00 g/molMolar mass = 58. 078 g/mol CH3OHCH3Calculate the mass of acetone in kg:massacetone=1. 00molacetone58. 078gacetone1molacetonemassacetone=58. 078gacetoneStep 3. Calculate thevolume of ethanol(in L) contained in the answer. We are able to calculate the amount by utilizing mass and density:Mass (density) volumeACETONEGiven values:Mass = 69. 694 g acetoneDensity of acetone = . 788 g/cm3The volume within the density given is within cm3, we will need to convert it towards the needed unit in L. Conversion units: 1 mL = 1 cm3 1 mL = 10-3 LCalculate the level of acetone:Put units you have to block out on opposite places (if its within the numerator then place the conversion factor with similar unit within the denominator)Vacetone=58.Video advice: Converting between Mass and Mole Fractions13.3: Units of ConcentrationA general chemistryLibretextsTextmaporganized around the textbookChemistry: Principles, Patterns, and Applicationsby Bruce A. Averill.B To calculate the mole fraction of acetic acidity within the solution, we have to know the amount of moles of both acetic acidity and water. The amount of moles of acetic acidity is . 0629 mol, as calculated partly (a). We all know that 100. g of vinegar contains 3. 78 g of acetic acidity therefore, the solution also includes (100. g 3. 78 g) = 96. 2 g water. Weve.Contributors and AttributionsThere are several different ways to quantitatively describe the concentration of a solution. For example, molarity was introduced in Chapter 4 as a useful way to describe solution concentrations for reactions that are carried out in solution. Mole fractions, introduced in Chapter 10, are used not only to describe gas concentrations but also to determine the vapor pressures of mixtures of similar liquids. Example 4 reviews the methods for calculating the molarity and mole fraction of a solution when the masses of its components are known.Calculations involving molality, molarity, density, mass percent, mole fraction Problems 1 10.Comment: A mole fraction of 0. 50 could mean 1. 0 mol of one component in 2. 0 total moles. Using 0. 5 and 1 is the simplest meaning of a mole fraction of 0. 5. We could have used any paring of numbers that gives a mole fraction of 0. 5. The final answers would be the same, but the numbers in the calculations would be different.Mole Fraction and Mole Percent In chemistry, the mole fraction, xi, is defined as the amount of moles of a constituent, ni, divided by the total amount of moles of all constituents in a mixture, ntot:First, we assume a total mass of 100. 0 g, although any mass could be assumed. This means that we have 50. 0 g of urea and 50. 0 g of cinnamic acid. We can then calculate the moles present by dividing each by its molecular weight. We have 0. 833 moles urea and 0. 388 moles cinnamic acid, so we have 1. 22 moles total.Learning ObjectiveKey PointsTermsYou can convert the percent by weight of solute in a solution to molarity, which is the number of moles per liter.Molarity of a Percent by Weight Solution You can determine the concentration of solute in a solution as a weight-to-weight percentage, a weight-to-volume percentage or volume-to-volume percentage. In this context, weight is synonymous with mass, so a mass percentage means the relative weight of the solute to the weight of the solution, and you can express it as percent by weight. However, its also common to relate weight to volume and express the result as percent weight to volume. Either way, as long as you know the chemical formulas for the solute and solvent (which is usually water), the weight percentage allows you to calculate how many moles of solute are present in the solution. From this, its possible to determine the molarity of the solution, which is the number of moles of solute per liter of solution. TL;DR (Too Long; Didnt Read) If you know the percent by weight of a solution, you can find the weight of solute. Divide this by its molecular weight to find the number of moles and divide by the volume of solution to find the molarity.Mass Fraction: Definition, Examples, Problems, Calculation ~ ChemistryGodMass Fraction, Weight Fraction, Percentage by Mass, Percentage by Weight, Weight fraction, Definition, Examples, Problems, Calculation, Solution, Molarity, Mass Concentration, Mole Fraction.Example 1: WineExample 2: Sodium Chloride SolutionMass PercentageExample 3: SteelExample 4: Manganese (II) SulphateMixing RatioMole FractionVolume FractionMass ConcentrationMolarity or Molar ConcentrationPercentage CompositionProblem 1: AirSolved ProblemsThe mass fraction is a property of a mixture. A mixture is the physical combination of two or more different substances. The types of mixtures are solutions, suspensions, and colloids. Figure 1: Types of Mixtures: Solutions (salt and water), Suspensions (soil and water), and Colloids (milk) (from left to right)(Image Sources: Wikimedia, Flickr, and Public Domain)In chemistry, most mixtures are solutions, homogeneous mixture of a solute and a solvent. But the below definition of the mass fraction is also applicable to the other types. 1 Definition and Formula2 Relationship with Other Quantities3 Solved Problems4 Practice Problems5 Associated ArticlesDefinition and FormulaThe mass fraction of a substance in a mixture is the ratio of the mass of the substance to the total mass of the mixture. It is also known as mass per cent or percentage by mass when expressing in percentage. Since the mass fraction is a ratio of mass to mass, it is a dimensionless quantity. For a mixture,Here, wi is the mass fraction of the ith component, mi is the mass of the ith component, and mT is the total mass of the mixture.Video advice: Mass Percent & Mole FractionMolarity serves as the bridge between moles to liters and can also incorporate the concept of mass percent and mole fraction. Mass Percent is represented by the mass of the solute per mass of solution, which is the combined mass of the solute and solvent. [FAQ] SolutionFind the weight of Solute and solvent from weight percent or mass percent.Find the molecules mass of the solute and solvent.Calculate the moles of solute and solvent using weight and molecular mass.Apply the mole fraction formula to obtain the mole fraction. The key difference between mole fraction and weight percent is that the mole fraction gives the composition of a substance in a mixture by means of moles, while the weight percent gives the composition by means of mass. Calculate the mole percent of one of the components by dividing its number of moles by the total number of moles of all substances and multiplying the result by 100. Mole percent is the percentage of the total moles that is of a particular component. Mole percent is equal to the mole fraction for the component multiplied by 100: mol % a = a 100. In a mixture, the mass fraction is the amount of mass of one substance, divided by the mass of the total mixture. The sum of all the mass fractions is equal to 1. Mass fraction can also be expressed, with a denominator of 100, as percentage by mass.

How to convert mole fraction to mass percent. How to find mole fraction from percent by mass. How to find mole fraction from density and mass percent. Mole fraction to mass percentage. How to calculate mole fraction given percent mass. How is the mole fraction related to the mass percentage of a component in a solution. Mole fraction to mass percent calculator. How to convert mole fraction to mass percentage. How to go from mole fraction to mass percent. How to get mass percent from mole fraction.