How to calculate the moles of an element

Get help with your researchJoin ResearchGate to ask questions, get input, and advance your work. Process Technology Consulting Process Technology Consulting Process Technology Consulting Philipps University of Marburg University of Trinidad and Tobago (UTT) Independent Consultant Related Pages Molar Volume Stoichiometry Lessons Molecular Mass Writing A Balanced Chemical Equation Chemistry Lessons The following diagram shows the conversion between Mole and Mass. Scroll down the page for more examples and solutions. Mole-Mass Equation mass = number of moles × molar mass where mass is in grams and the molar mass is in grams per mole. Moles to Mass Calculation We can use the above equation to find the mass of a substance when we are given the number of moles of the substance. Example: Calculate the mass of (a) 2 moles and (b) 0.25 moles of iron. (Relative atomic mass: Fe = 56) Solution: a) mass of 2 moles of iron = number of moles × molar mass = 2 × 56 = 112 g b) mass of 0.25 mole of iron = number of moles × molar mass = 0.25 × 56 = 14 g Example: Calculate the mass of (a) 3 moles and (b) 0.2 moles of carbon dioxide gas, CO2. (Relative atomic mass: C = 12; O = 16) Solution: a) mass of 1 mole of CO2 = (1 × 12) + (2 × 16) = 44 g mass of 3 moles of CO2 = 3 × 44 = 132g b) mass of 0.2 mole of CO2 = 0.2 × 44 = 8.8 g Examples of moles to mass calculation Example: If an experiment calls for 0.200 mol acetic acid (HC2H3O2), how many grams of glacial acetic acid do we need? Formula: m = nM Show Video Lesson Example: If an experiment calls for 0.500 mol CaCO3, how many grams of pure calcium carbonate do we need? Show Video Lesson Mass to Moles Calculation If we are given the mass of a substance and we are asked to find the number of moles of the substance, we can rewrite the above equation as Example: Calculate the number of moles of aluminum present in (a) 108 g and (b) 13.5 g of the element. (Relative atomic mass: Al = 27) Solution: a) b) Example: Calculate the number of moles of magnesium oxide, MgO in (a) 80 g and (b) 10 g of the compound. (Relative atomic mass: O = 16, Mg = 24) Solution: a) Mass of 1 mole of MgO = (1 × 24) + (1 × 16) = 40 g b) Examples of mass to mole calculation How many moles of acetic acid (HC2H3O2) are present in a 5.00 g sample of pure acetic acid? Show Video Lesson How to use formula mass to convert grams to moles and moles to grams? Examples: How many moles of NAOH are represented by 80.0 grams of NAOH? How many grams will 3.5 moles of NAOH weigh? Show Video Lesson Try the free Mathway calculator and problem solver below to practice various math topics. Try the given examples, or type in your own problem and check your answer with the step-by-step explanations. We welcome your feedback, comments and questions about this site or page. Please submit your feedback or enquiries via our Feedback page. Percent Compostition Why do we care about % composition? When an elemental analysis is performed we do not get the molecular formula. The result of an elemental analysis is always reported as a percent mass of each of the elements for which the analysis was performed. An elemental analysis of a white crystalline compound believed to be C6H12O6, analyzed for C, H and O gave the following results. C = 40.00 % , H = 6.71 %, O = 53.28 % by wt. Is the the compound actually C6H12O6? The problem is really just a conversion from % by wt. to moles. Why a conversion to moles? Because that is what a formula is, a formula relates moles of one element to moles of another element. Start the problem with the assumption that you have 100.0 g of sample. The percent composition tells you how much of each element is present: Why use 100 g? Because it is convenient. The formula is determined by the mole ratio of the elements, not the mass ratios. So, we need to convert the number of grams to number of moles. Finally, the mole ratio is determined. Divide each molar amount by the smallest molar amount. The ratio is 1 C for 2 H for 1 O. According to our calculations we have CH2O or formaldehyde (which is a gas)! Are we wrong...NO! We have determined the ratios in which the elements are combined to form the compound, but not the actual amount of the elements present. To determine the actual amount of each element present we need to know the Molar Mass (often refered to as the molecular weight or MW). The molar mass must be determined using another experiment (A mass spectrometer could determine the molar mass. There are other less exspensive ways to determine molar mass, and we will discuss some of them next semester.) The molar mass was determined to be 180.16 g/mol. To determine the molecular formula we will use the empirical formula CH2O which we just found. The molecule must be made up of some number of empirical formula units; that way the ratio of the elements remains the same: molecular formula = (a number) x (empirical formula) If we knew any two of these numbers then we could find the third, but we only know one. If the molecule is made of some number of empirical formula units the the molar mass must be some multiple of the empirical mass. molar mass = (a number) x (empirical mass) Since, empirical mass CH2O = 30.25 g/mole 180.16 g/mol = (a number) x 30.25 g/mol (a number) = 5.956 or 6 molecular formula = (6) x (CH2O) So, we did have C6H12O6 Essentially, we found how many empirical formula units are in the whole. A bottle of improperly labeled "iron oxide" is found in a laboratory. Elemental analysis reveals that the sample is iron and oxygen, and the sample is 69.94 % iron by mass. What is the formula of the material, and what should the bottle be labeled; that is, what is the name of the compound? To find the formula we need to know the % mass of each of the elements. Since, there are only 2 different elements in iron oxide we know... %Fe + %O = 100 % 100 % - %Fe = %O So, 100 % - 69.94 % = %O %O = 30.06 % Now that we have % mass of both elements let's find the formula. Once again assume a100 g sample. Reduce to lowest ratio by dividing each number by the smallest number... Formulae do not (usually) contain fractions. So, we must find the lowest common whole number multiple. Multiplying 1.5 by 2 gives a whole number, 3. So, multiply each number by 2 1.5 x 2 = 3 mol O and 1 Fe x 2 = 2 mol Fe So, the formula is Fe2O3 The name is iron (III) oxide To find a formula we do not need to have % composition. If we are told the mass of each element present in a compound we can find the formula. The mass of the elements can be converted to moles of the elements. The mole ratio reveals the empirical formula. For example, A compound containing only Mn and Cl contains 1.9228 g Mn and 2.4817 g Cl. Determine the empirical formula. We simply need to convert the masses of the elements to moles. The empirical formula is determined by the mole ratio of Mn to Cl. The formula is MnCl2 Combustion analysis is used to determine the empirical formula of hydrocarbons (a compound containing only hydrogen and carbon). Combustion of a hydrocarbon with a molar mass of 78.11 g/mol produced 2.6406 g CO2 and 0.5400 g H2O. Determine the molecular formula. Since we do not know the formula we cannot balance the equation. The important thing to note is that all the carbon in the CO2 must come from the hydrocarbon. All the hydrogen in the water must also come from the hydrocarbon! If we determine the number of moles of C in the CO2 we will have the number of moles of carbon in the hydrocarbon. If we determine the number of moles of H in the H2O we have the number of moles of H in the hydrocarbon. The empirical formula is CH So, the molecular formula is made up of some number of empirical formula units. Use the molar mass and the empirical formula mass to determine that number. The molecule contains 6 empirical formula units. The molecular formula is C6H6 back to GenChem Home Page The relative formula mass of a compound is calculated by adding together the relative atomic mass values for all the atoms in its formula. Moles are units used to measure substance amount. The identity of a substance is defined not only by the types of atoms or ions it contains, but by the quantity of each type of atom or ion. For example, water, H2O, and hydrogen peroxide, H2O2, are alike in that their respective molecules are composed of hydrogen and oxygen atoms. However, because a hydrogen peroxide molecule contains two oxygen atoms, as opposed to the water molecule, which has only one, the two substances exhibit very different properties. Today, we possess sophisticated instruments that allow the direct measurement of these defining microscopic traits; however, the same traits were originally derived from the measurement of macroscopic properties (the masses and volumes of bulk quantities of matter) using relatively simple tools (balances and volumetric glassware). This experimental approach required the introduction of a new unit for amount of substances, the mole, which remains indispensable in modern chemical science. The mole is an amount unit similar to familiar units like pair, dozen, gross, etc. It provides a specific measure of the number of atoms or molecules in a bulk sample of matter. A mole is defined as the amount of substance containing the same number of discrete entities (such as atoms, molecules, and ions) as the number of atoms in a sample of pure 12C weighing exactly 12 g. One Latin connotation for the word "mole" is "large mass" or "bulk," which is consistent with its use as the name for this unit. The mole provides a link between an easily measured macroscopic property, bulk mass, and an extremely important fundamental property, number of atoms, molecules, and so forth. The number of entities composing a mole has been experimentally determined to be \(6.02214179 \times 10^{23}\), a fundamental constant named Avogadro's 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}\). Video \(\PageIndex{1}\): What is Avogadro's Number? Consistent with its definition as an amount unit, 1 mole of any element contains the same number of atoms as 1 mole of any other element. The masses of 1 mole of different elements, however, are different, since the masses of the individual atoms are drastically different. The molar mass of an element (or compound) is the mass in grams of 1 mole of that substance, a property expressed in units of grams per mole (g/mol) (Figure \(\PageIndex{1}\)). Figure \(\PageIndex{1}\): Each sample contains \(6.022 \times 10^{23}\) atoms —1.00 mol of atoms. From left to right (top row): 65.4 g zinc, 12.0 g carbon, 24.3 g magnesium, and 63.5 g copper. From left to right (bottom row): 32.1 g sulfur, 28.1 g silicon, 207 g lead, and 118.7 g tin. (credit: modification of work by Mark Ott). Because the definitions of both the mole and the atomic mass unit are based on the same reference substance, 12C, the molar mass of any substance is numerically equivalent to its atomic or formula weight in amu. Per the amu definition, a single 12C atom weighs 12 amu (its atomic mass is 12 amu). The former definition of the mole was that a mole was 12 g of 12C contains 1 mole of 12C atoms (its molar mass is 12 g/mol). This relationship holds for all elements, since their atomic masses are measured relative to that of the amu-reference substance, 12C. Extending this principle, the molar mass of a compound in grams is likewise numerically equivalent to its formula mass in amu. On May 20, 2019 the definition was permanently changed to Avogadro's number: a mole is \(6.02214179 \times 10^{23}\) of any object, from atoms to apples.1 Table \(\PageIndex{2}\): Mass of one mole of elements Element Average Atomic Mass (amu) Molar Mass (g/mol) Atoms/Mole C 12.01 12.01 \(6.022 \times 10^{23}\) H 1.008 1.008 \(6.022 \times 10^{23}\) O 16.00 16.00 \(6.022 \times 10^{23}\) Na 22.99 22.99 \(6.022 \times 10^{23}\) Cl 33.45 35.45 \(6.022 \times 10^{23}\) While atomic mass and molar mass are numerically equivalent, keep in mind that they are vastly different in terms of scale, as represented by the vast difference in the magnitudes of their respective units (amu versus g). To appreciate the enormity of the mole, consider a small drop of water after a rainfall. Although this represents just a tiny fraction of 1 mole of water (~18 g), it contains more water molecules than can be clearly imagined. If the molecules were distributed equally among the roughly seven billion people on earth, each person would receive more than 100 billion molecules. Video \(\PageIndex{2}\): The mole is used in chemistry to represent \(6.022 \times 10^{23}\) of something, but it can be difficult to conceptualize such a large number. Watch this video and then complete the "Think" questions that follow. Explore more about the mole by reviewing the information under "Dig Deeper." The relationships between formula mass, the mole, and Avogadro's number can be applied to compute various quantities that describe the composition of substances and compounds. 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 substance's mass. Example \(\PageIndex{1}\): Deriving Moles from Grams for an Element According to nutritional guidelines from the US Department of Agriculture, the estimated average requirement for dietary potassium is 4.7 g. What is the estimated average requirement of potassium in moles? Solution The mass of K is provided, and the corresponding amount of K in moles is requested. Referring to the periodic table, the atomic mass of K is 39.10 amu, and so its molar mass is 39.10 g/mol. The given mass of K (4.7 g) is a bit more than one-tenth the molar mass (39.10 g), so a reasonable "ballpark" estimate of the number of moles would be slightly greater than 0.1 mol. The molar amount of a substance may be calculated by dividing its mass (g) by its molar mass (g/mol): The factor-label method supports this mathematical approach since the unit "g" cancels and the answer has units of "mol:" \[\mathrm{4.7\; \cancel{g} K \left (\dfrac{mol\; K}{39.10\;\cancel{g}}\right)=0.12\;mol\; K} onumber\] The calculated magnitude (0.12 mol K) is consistent with our ballpark expectation, since it is a bit greater than 0.1 mol. Exercise \(\PageIndex{1}\): Beryllium Beryllium is a light metal used to fabricate transparent X-ray windows for medical imaging instruments. How many moles of Be are in a thin-foil window weighing 3.24 g? Answer 0.360 mol Example \(\PageIndex{2}\): Deriving Grams from Moles for an Element A liter of air contains \(9.2 \times 10^{−4}\) mol argon. What is the mass of Ar in a liter of air? Solution The molar amount of Ar is provided and must be used to derive the corresponding mass in grams. Since the amount of Ar is less than 1 mole, the mass will be less than the mass of 1 mole of Ar, approximately 40 g. The molar amount in question is approximately one-one thousandth (~10−3) of a mole, and so the corresponding mass should be roughly one-one thousandth of the molar mass (~0.04 g): In this case, logic dictates (and the factor-label method supports) multiplying the provided amount (mol) by the molar mass (g/mol): \[\mathrm{9.2 \times10^{-4}\; \cancel{mol} \; Ar \left(\dfrac{39.95\;g}{\cancel{mol}\;Ar} \right)=0.037\;g\; Ar} onumber\] The result is in agreement with our expectations, around 0.04 g Ar. Exercise \(\PageIndex{2}\) What is the mass of 2.561 mol of gold? Answer 504.4 g Example \(\PageIndex{3}\): Deriving Number of Atoms from Mass for an Element Copper is commonly used to fabricate electrical wire (Figure \(\PageIndex{6}\)). How many copper atoms are in 5.00 g of copper wire? Figure \(\PageIndex{6}\): Copper wire is composed of many, many atoms of Cu. (credit: Emilian Robert Vicol) Solution The number of Cu atoms in the wire may be conveniently derived from its mass by a two-step computation: first calculating the molar amount of Cu, and then using Avogadro's number (NA) to convert this molar amount to number of Cu atoms: Considering that the provided sample mass (5.00 g) is a little less than one-tenth the mass of 1 mole of Cu (~64 g), a reasonable estimate for the number of atoms in the sample would be on the order of one-tenth NA, or approximately 1022 Cu atoms. Carrying out the two-step computation yields: \[\mathrm{5.00\:\cancel{g}\:Cu\left(\dfrac{\cancel{mol}\:Cu}{63.55\:\cancel{g}}\right)\left(\dfrac{6.022\times10^{23}\:atoms}{\cancel{mol}}\right)=4.74\times10^{22}\:atoms\: of\: copper}\] The factor-label method yields the desired cancellation of units, and the computed result is on the order of 1022 as expected. Exercise \(\PageIndex{3}\) A prospector panning for gold in a river collects 15.00 g of pure gold. How many Au atoms are in this quantity of gold? Answer \(4.586 \times 10^{22}\; Au\) atoms Example \(\PageIndex{4}\): 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 compound's 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 \(\PageIndex{4}\) How many moles of sucrose, \(C_{12}H_{22}O_{11}\), are in a 25-g sample of sucrose? Answer 0.073 mol Example \(\PageIndex{5}\): 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 4–8 years is 1.42 × 10−4 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 (~10−4 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 \(\PageIndex{5}\) What is the mass of 0.443 mol of hydrazine, \(N_2H_4\)? Answer 14.2 g Example \(\PageIndex{6}\): 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 Avogadro's 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 compound's 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 \(\PageIndex{6}\) 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 \(\PageIndex{5}\): A preview of some of the uses we will have for moles in upcoming units 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 Contributors Paul Flowers (University of North Carolina - Pembroke), Klaus Theopold (University of Delaware) and Richard 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 Fuse School, Open Educational Resource free of charge, under a Creative Commons License: Attribution-NonCommercial CC BY-NC (View License Deed: TED-Ed's commitment to creating lessons worth sharing is an extension of TED's mission of spreading great ideas. 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