Chapter 3: Composition of Substances and Solutions

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Molecular and Formula Mass

The **molecular mass** is the mass in atomic mass units (amu) of an individual molecule.

To calculate molecular mass, multiply the atomic mass of each element in a molecule by the number of atoms of that element and then total the masses

Molecular mass of H_2O:

2(atomic mass units of H) +(1)(atomic mass units of O)

2(1.008 amu) + (1)(16.00 amu) = 18.02 amu

Because the atomic masses on the periodic table are average atomic masses, the result of such a determination is an average molecular mass, sometimes referred to as the **molecular weight**.

Although an ionic compound does not have a molecular mass, we can use its empirical formula to determine its **formula mass** (the mass of a "formula unit"), sometimes called the **formula weight**.

The process is the same:
To calculate formula mass, multiply the atomic mass for each element in a formula unit by the number of atoms of that element and then total the masses

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Example Problem

Calculate the molecular mass or the formula mass, as appropriate, for each of the following:
(a) ethane, C_2H_6, (b) lithium hydroxide, (c) $CaCl_2$

(a) C_2H_6 : This says 2 carbon atoms and 6 hydrogen atoms. The molecular mass of ethane is:

$$(2)(12.01\text{ amu}) + (6)(1.008\text{ amu}) = 30.07\text{ amu}$$

(b) LiOH : This says 1 lithium atom and 1 oxygen atom and 1 hydrogen atom. The formula mass is:

$$(1)(6.94\text{ amu}) + (1)(16.00\text{ amu}) + (1)(1.008\text{ amu}) = 23.95\text{ amu}$$

(c) $CaCl_2$: This says 1 calcium atom and 2 chlorine atoms. The formula mass is:

$$(1)(40.078\text{ amu}) + (2)(35.453\text{ amu}) = 110.984\text{ amu}$$

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Practice Problem

Determine the Formula/Molecular weight for the following compounds:

Li_2NO_3	Formula mass = 75.89 amu
Potassium cyanide	Formula mass for KCN = 65.12 amu
Al_2O_3	Formula mass = 101.96 amu
Iron(III) sulfate	Formula mass for $Fe_2(SO_4)_3$ = 399.88 amu
Carbon dioxide	Molecular mass for CO_2 = 44.01 amu

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Percent Composition of Elements

A list of the percent by mass of each element in a compound is known as the compound's **percent composition by mass**.

$$\boxed{\text{mass percent of an element} = \frac{\text{n x atomic mass of element}}{\text{molecular or formula mass of compound}} \text{ x } 100\%}$$

where n is the number of atoms of the element in a molecule or formula unit of the compound

For a molecule of H_2O_2:

$$\%\mathrm{H} = \frac{(2)(1.008 \text{ amu H})}{34.02 \text{ amu } \mathrm{H_2O_2}} \text{ x } 100\%$$

$$\%\mathrm{H} = 5.926\ \%$$

$$\%\mathrm{O} = \frac{(2)(16.00 \text{ amu O})}{34.02 \text{ amu } \mathrm{H_2O_2}} \text{ x } 100\%$$

$$\%\mathrm{O} = 94.06\ \%$$

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Percent Composition of Elements

We could also have used the empirical formula of hydrogen peroxide (HO) for the calculation.

In this case, we could have used the **empirical formula mass**, the mass in amu of one empirical formula, in place of the molecular formula.

The empirical formula mass of H_2O_2 (the mass of HO) is 17.01 amu.

$$\%\mathrm{H} = \frac{(1)(1.008 \text{ amu H})}{17.01 \text{ amu HO}}\ x\ 100\%$$

$$\%\mathrm{H} = 5.926\ \%$$

$$\%\mathrm{O} = \frac{(1)(16.00 \text{ amu O})}{17.01 \text{ amu HO}}\ x\ 100\%$$

$$\%\mathrm{O} = 94.06\ \%$$

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Examples

What is the percent composition of O in the following compounds?

ClO_4^- $\quad \%O = \frac{4 * 16.00 \text{ amu O}}{99.45 \text{ amu ClO}_4^-} * 100\% = 64.35\%$

H_2SO_4 $\quad \%O = \frac{4 * 16.00 \text{ amu O}}{98.076 \text{ amu H}_2\text{SO}_4} * 100\% = 65.26\%$

SiO_2 $\quad \%O = \frac{2 * 16.00 \text{ amu O}}{60.09 \text{ amu SiO}_2} * 100\% = 53.25\%$

Calculate the percent composition of nitrogen in sodium azide (NaN_3). 64.64%

Lithium carbonate, Li_2CO_3, was the first "mood–stabilizing" drug approved by the FDA for the treatment of mania and manic-depressive illness, also known as bipolar disorder. Calculate the percent composition by mass of lithium carbonate. Li: 18.79%; C: 16.25%; O: 64.96%

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Determination of Empirical Formula and Molecular Formula from Percent Comp.

Using the concepts of the mole and molar mass, we can now use an experimentally determined percent composition to determine the empirical and/or molecular formula.

The empirical formula gives only the ratio of atoms in a molecule (many compounds can have the same empirical formula).

A compound's empirical formula can be determined from its percent composition. A compound's molecular formula is determined from the molar mass and empirical formula.

Given: Mass Percentages $\xrightarrow{\text{Use Molar Mass}}$ Moles $\rightarrow$ Mole Ratios $\rightarrow$ Empirical Formula $\xrightarrow{\text{Use Molar Mass}}$ Molecular Formula

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Empirical and Molecular Formulas

An orange compound was found to be 26.6% K, 35.4% Cr and 38.0% O. Determine its empirical formula.

1. Assume a 100.0 g sample. Percent becomes mass in grams.

   26.6% K becomes 26.6 g K; 35.4% Cr becomes 35.4 g Cr; 38.0% O becomes 38.0 g O

2. Divide each mass by its atomic mass. Gives the number of moles of each atom.

$$\left(\frac{26.6\text{ g K}}{}\right)\left(\frac{1\text{ mol K}}{39.10\text{ g K}}\right) = 0.6803\text{ mol K} \qquad \left(\frac{35.4\text{ g Cr}}{}\right)\left(\frac{1\text{ mol Cr}}{52.00\text{ g Cr}}\right) = 0.6808\text{ mol Cr}$$

$$\left(\frac{38.0\text{ g O}}{}\right)\left(\frac{1\text{ mol O}}{16.00\text{ g O}}\right) = 2.375\text{ mol O}$$

3. Divide each by the smallest number of moles. The smallest whole number ratio is the empirical formula.

$$\frac{0.6803}{0.6803} = 1 \qquad \frac{0.6808}{0.6803} = 1.001 \qquad \frac{2.375}{0.6803} = 3.491$$

These ratios tell us: $KCrO_{3.5}$. These are NOT all whole numbers:

4. If necessary, multiply to get whole number ratios:

   $(2)(KCrO_{3.5})$ = $K_2Cr_2O_7$

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Practice

Determine the empirical formula of a compound that is 30.45 percent nitrogen and 69.55 percent oxygen by mass. Given that the molar mass of the compound is approximately 92 g mol^{-1}, determine the molecular formula of the compound.

Assume 100 g of compound:
30.45 g N
69.55 g O

$$\left(\frac{30.45\text{ g N}}{}\right)\left(\frac{1\text{ mol N}}{14.01\text{ g N}}\right) = 2.173448\text{ mol N}$$

$$\left(\frac{69.55\text{ g O}}{}\right)\left(\frac{1\text{ mol O}}{16.00\text{ g O}}\right) = 4.346875\text{ mol O}$$

$$\frac{2.173448\text{ mol N}}{2.173448\text{ mol N}} = 1\text{ N:N} \qquad \frac{4.346875\text{ mol O}}{2.173448\text{ mol N}} = 2\text{ O : 1 N}$$

NO_2(Empirical): Molar Mass = 46.01 g

$$\frac{\text{Molar Mass Compound}}{\text{Molar Mass NO}_2} = \frac{\frac{92\text{ g}}{\text{mol}}}{\frac{46.01\text{ g}}{\text{mol}}} = 2 \longrightarrow (2)(\text{NO}_2) = \text{N}_2\text{O}_4$$

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Poll Question: Campusknot

(Ch. 6-Slide 11) How many carbon atoms are found in the empirical formula for a compound that contains 26.1% carbon, 4.3% hydrogen and 69.6 % oxygen?

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Concentration

Solution: is a homogeneous mixture of 2 or more materials.

Solvent : largest fraction (largest concentration)

Solute(s) : minor fraction(s) (smaller concentration)

Solution Video

Concentration is the relative amounts of solute and solvent; many ways to express.

Dilute : small amount of solute

Concentrated : a lot of solute

Concentration: Molarity = $\frac{\text{mol}}{\text{L}}$ = mol L^{-1}

Temp. dependent

$$M = \frac{\text{mole solute}}{\text{L}\ \textbf{solution}} \quad \textbf{or} \quad M = \frac{n_{\text{solute}}}{\text{L solution}}$$

[KCl] means the molarity of the KCl solution.

Concentration: Molality = $\frac{\text{mol}}{\text{kg}}$ = mol kg^{-1}

Temp. Independent

$$m = \frac{\text{mole solute}}{\text{kg}\ \textbf{Solvent}} \quad \textbf{or} \quad m = \frac{n_{\text{solute}}}{\text{kg}\ \textbf{Solvent}}$$

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Molarity

6.37 g of $Al(NO_3)_3$ are dissolved to make a 250. mL aqueous solution. Calculate (a) $[Al(NO_3)_3]$ (b) $[Al^{3+}]$ and $[NO_3^-]$.

$$\mathrm{Al(NO_3)_3 \text{ molar mass: }} \left(\frac{26.98\ \mathrm{g}}{\mathrm{mole}}\right) + \left[(3)\left(\frac{14.00\ \mathrm{g}}{\mathrm{mol}}\right)\right] + \left[(9)\left(\frac{16.00\ \mathrm{g}}{\mathrm{mol}}\right)\right] = \frac{213.0\ \mathrm{g}}{\mathrm{mol}}$$

$$\left(\frac{6.37\ \mathrm{g\ Al(NO_3)_3}}{}\right)\left(\frac{1\ \mathrm{mol\ Al(NO_3)_3}}{213.0\ \mathrm{g\ Al(NO_3)_3}}\right) = 2.991 \text{ x } 10^{-2}\mathrm{mol\ Al(NO_3)_3}$$

$$\mathrm{Al(NO_3)_3} = \frac{2.991\ x\ 10^{-2}\ \mathrm{mol}}{0.250\ \mathrm{L}} = 0.11964\ M$$

There are 3 NO_3^- and 1 Al^{3+}: Information given from the ionic compound $Al(NO_3)_3$

$$\left(\frac{0.11964\ \mathrm{mol\ Al(NO_3)_3}}{1\ \mathrm{L}}\right)\left(\frac{1\ \mathrm{mol\ Al^{3+}}}{1\ \mathrm{mol\ Al(NO_3)_3}}\right) = 0.11964\ M\ \mathrm{Al^{3+}}$$

$$\left(\frac{0.11964\ \mathrm{mol\ Al(NO_3)_3}}{1\ \mathrm{L}}\right)\left(\frac{3\ \mathrm{mol\ NO_3^-}}{1\ \mathrm{mol\ Al(NO_3)_3}}\right) = 0.35892\ M\mathrm{NO_3^-}$$

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Dilution

Dilution is the process of preparing a less concentrated solution from a more concentrated one.

Dilution Video

$$M_1V_1 = M_2V_2$$

moles of solute before dilution = moles of solute after dilution

A series of dilutions that may be used to prepare several increasingly dilute solutions is called ***Serial Dilution***.

1: Prepare a dilute solution from the stock

2: Dilute a portion of the prepared solution to make a more dilute solution

3: Repeat as needed

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Dilution

In an experiment, a student needs 1.00 L of a 0.400 *M* $KMnO_4$ solution. A stock solution of 1.00 *M* $KMnO_4$ is available. How much of the stock solution is needed?

Solution: Use the relationship that moles of solute before dilution = moles of solute after dilution.

$$(1.00\ M\ KMnO_4)(V_I) = (0.400\ M\ KMnO_4)(1.00\ L)$$

$$V_I = 0.400\ L \text{ or } 400\ mL$$

To make the solution:

1. Pipet 400 mL of stock solution into a 1.00 L volumetric flask.
2. Carefully dilute to the calibration mark.

Because most volumes measured in the laboratory are in milliliters rather than liters, it is worth pointing out that the equation can be written as

$$M_1 \times mL_1 = M_2 \times mL_2$$

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Practice

Commercial concentrated sulfuric acid is 17.8 *M*. If 75.0 mL of this acid is diluted to 1.00 L, what is the final concentration of the acid?

M_{conc} = 17.8 M V_{conc} = 75.0 mL

M_{dil} = ? V_{dil} = 1000. mL

$$M_{dil} = \frac{M_{con}V_{con}}{V_{dil}} = \frac{(17.8\ M)(75.0\ \cancel{mL})}{(1000\ \cancel{mL})} = 1.34\ M$$

Prepare a 0.5000 *M* solution of potassium permanganate in a 250.0 mL volumetric flask.

Mass of $KMnO_4$ required: to find mass we have to find moles. We can use the volume and concentration for this

$$\boxed{\text{Remember: } M = \frac{\text{mol}}{\text{L}}}$$

$$\left(\frac{0.2500\ L}{}\right)\left(\frac{0.5000\ mol\ KMnO_4}{L}\right) = 0.1250 \text{ mol } KMnO_4$$

$$\left(\frac{0.1250\ mol}{}\right)\left(\frac{158.03\ g\ KMnO_4}{1\ mol}\right) = 19.75\ g\ KMnO_4$$

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Practice

1. How many mL of 0.745 *M* solution do I use to make 750 mL of 0.680 *M* solution?

2. I have a standard solution $KMnO_4$ that is 0.500 M. How, using that standard solution, do I make 250 mL of a 0.218 *M* solution?

3. What volume of 0.62 *M* Na_2SO_4 solution should I use if I want to make 500 mL of 0.14 *M* solution?

4. What mass of 71% HNO_3 is contained in 0.500 L of solution knowing that the density of HNO_3 solution is 1.42 g/mL.

Could also ask how many moles of HNO_3?

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Concentration Units

$$\text{Mass Fraction} = \frac{\text{mass solute}}{\text{Total mass of solution}}$$

$$\text{Mass Percent} = \text{Mass Fraction} * 100\%$$

Example Problem:
4.6 g of NaCl is dissolved in 500 g of water. What is the mass percent of NaCl?

$$\textbf{Mass Fraction} = \frac{\mathbf{4.6\ g\ NaCl}}{\mathbf{(500\ g\ H_2O + 4.6\ g\ NaCl)}} = \mathbf{0.009116}$$

$$\textbf{Mass Percent} = \mathbf{(0.009116)\ (100)} = \mathbf{0.9116\%}$$

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Concentration: Mole fraction (no units)

$$\chi_{solute} = \frac{n_{solute}}{n_{solute} + n_{solvent}}$$

Large χ_{solute} = large amount of solute relative to solvent

$$\chi_{solvent} = \frac{n_{solvent}}{n_{solute} + n_{solvent}}$$

Large $\chi_{solvent}$ = large amount of solvent relative to solute

Concentration: Mole percent (mol %): mole fraction as a percentage

$$mol\ \%_{solute} = (\chi_{solute}) * (100\%)$$

$$mol\ \%_{solvent} = (\chi_{solvent}) * (100\%)$$

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Concentration: Parts by mass

$$\frac{\text{mass solute}}{\text{mass solution}} * \textbf{multiplication factor}$$

Percent by mass (%)	Mult. Factor = 100
Parts per million by mass (ppm)	Mult. Factor = 10^6
Parts per billion by mass (ppb)	Mult. Factor = 10^9

Concentration: Parts by volume

$$\frac{\text{volume solute}}{\text{volume solution}} * \textbf{multiplication factor}$$

Percent by volume (%)	Mult. Factor = 100
Parts per million by volume (ppm)	Mult. Factor = 10^6
Parts per billion by volume (ppb)	Mult. Factor = 10^9

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Converting Units

Sea water is 10,600 ppm Na^+. Calculate the mass fraction and molarity of sodium ions in sea water. The density of sea water is 1.03 g/mL.

$$10{,}600 \text{ ppm} = \frac{10{,}600 \text{ g Na}^+}{10^6 \text{g solution}}$$

$$\text{mass fraction} = \frac{10{,}600 \text{ g Na}^+}{10^6 \text{g solution}} = 0.0106$$

$$\text{Molarity} = \frac{\text{n}_{\text{solute}}}{\text{L}_{\text{solution}}}$$

$$\left(\frac{10{,}600 \text{ g Na}^+}{}\right)\left(\frac{1 \text{ mol Na}^+}{22.00 \text{ g Na}^+}\right) = 461.1 \text{ mol}$$

$$\left(\frac{10^6 \text{ g solution}}{}\right)\left(\frac{1 \text{ mL solution}}{1.03 \text{ g solution}}\right) = 9.709\ x\ 10^5 \text{ mL}$$

$$\text{Molarity} = \frac{461.1 \text{ mol}}{970.9 \text{ L}} = 0.475\ M$$

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Converting Units

A solution is prepared by dissolving 17.2 g ethylene glycol ($C_2H_6O_2$) in 0.500 kg of water (at 25°C). The final solution volume is 515 mL. Calculate the following concentrations: a.) *M* b.) *m* c.) % by mass d.) mol fraction solute e.) mol % solute

$$\left(\frac{17.2 \text{ g C}_2\text{H}_6\text{O}_2}{}\right)\left(\frac{1\ mol}{62.07 \text{ g}}\right) = 0.2771 \text{ mol C}_2\text{H}_6\text{O}_2$$

a.) $$\text{Molarity} = \frac{n_{\text{solute}}}{\text{L}_{\text{solution}}} = \frac{0.2771 \text{ mol}}{0.515 \text{ L}} = 0.538\ M$$

b.) $$\text{Molality} = \frac{n_{\text{solute}}}{\text{kg}_{\text{solvent}}} = \frac{0.2771 \text{ mol}}{0.500 \text{ kg}} = 0.554\ m$$

c.) $$\text{Mass \%} = \left(\frac{\text{mass}_{\text{solute}}}{\text{mass}_{\text{total}}}\right)(100) = \left(\frac{17.2 \text{ g}}{(17.2+500 \text{ g})}\right)(100) = 3.33\ \%$$

d.) $$\left(\frac{500 \text{ g H}_2\text{O}}{}\right)\left(\frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}}\right) = 27.75 \text{ mol} \qquad \text{X}_{\text{solute}} = \frac{0.2771 \text{ mol}}{(0.2771+27.75 \text{ mol})} = 9.89\ x10^{-3}$$

e.) $$\text{mol \% solute} = (9.89 \text{ x}10^{-3})(100) = 0.989\ \%$$

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More Practice

Calculate the molarity and molality of a 30.0% H_2O_2 aqueous solution (d = $\frac{1.11\,g}{mL}$).

Assume 100 g solution: Gives 30.0 g H_2O_2 and 70.0 g H_2O

$$\left(\frac{30.0\ g\ H_2O_2}{}\right)\left(\frac{1\ mol}{34.02\ g}\right) = 0.8818\ mol\ H_2O_2$$

$$\left(\frac{70.0\ g\ H_2O}{}\right)\left(\frac{1\ mol}{18.02\ g}\right) = 3.885\ mol\ H_2O$$

$$\left(\frac{100\ g\ solution}{}\right)\left(\frac{1\ mL}{1.11\ g}\right) = 90.09\ mL$$

$$m = \frac{n_{solute}}{kg_{solvent}} = \frac{0.8818\ mol}{0.070\ kg} = 12.60\ m$$

$$M = \frac{n_{solute}}{L_{solution}} = \frac{0.8818\ mol}{0.09009\ L} = 9.79\ M$$

Calculate $[CO_2]$ of a 0.400 m aq. solution (d = $\frac{1.025\,g}{mL}$).

$$\left(\frac{0.400\ mol\ CO_2}{1\ kg\ solvent}\right)\left(\frac{1\ kg}{1000\ g}\right)\left(\frac{1.025\ g}{1\ mL}\right)\left(\frac{1000\ mL}{1\ L}\right) = 0.41\ M$$

What's the mass % glucose ($C_6H_{12}O_6$ = $\frac{180.2\,g}{mol}$) in 2.50 M aqueous glucose solution (d = $\frac{1.149\,g}{mL}$)?

$$\left(\frac{1\ L}{}\right)\left(\frac{2.50\ mol}{1\ L}\right)\left(\frac{180.2\ g}{1\ mol}\right) = 450\ g\ C_6H_{12}O_6$$

$$\left(\frac{1\ L}{}\right)\left(\frac{1000\ mL}{1\ L}\right)\left(\frac{1.149\ g}{1\ mL}\right) = 1149\ g\ H_2O$$

$$\text{mass}\ \% = \frac{450\ g}{(450\ g + 1149 g)}(100) = 28.14\ \%$$

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More Practice

How do I make a 203 g solution that is 13.6% NaBr? Assume the solution has the density of water is 1.00 g/mL.

$$\frac{13.6\ \%}{100} = 0.136$$

$$(203\ g)(0.136) = 27.6\ g\ NaBr$$

$$203\ g - 27.608\ g = 175.39\ g\ H_2O$$

27.608 g NaBr:175.392 g H_2O

A 12-oz (355 mL) Pepsi contains 38.9 mg of caffeine (molar mass= 194.2 g/mol). Assume that the Pepsi, mainly water, has a density of 1.01 g/mL. For such a Pepsi, calculate the caffeine concentration in ppm.

$$\left(\frac{355\ mL}{}\right)\left(\frac{1.01\ g}{1\ mL}\right) = 358.55\ g$$

We have 38.9 mg/358.55 g solvent; we need to know how many mg/1000 g solvent per the ppm definition.

$$ppm = \frac{1\ mg}{1000\ g}$$

$$\left(\frac{1000\ g}{}\right)\left(\frac{38.9\ mg}{358.55\ g\ solvent}\right) = 108.49\ ppm$$

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More Practice

An aqueous solution is 6.75 % by mass NaCl and has a density of 1.02 g/mL. What is the a. molality, b. Molarity, and c. Mole percent of NaCl?

Assume 100 g of solution:
6.75 g NaCl
93.25 g H_2O

$$\left(\frac{6.75\text{ g NaCl}}{}\right)\left(\frac{1\text{ mol NaCl}}{58.44\text{ g NaCl}}\right) = 0.1155\text{ mol NaCl}$$

a. $m = \dfrac{n_{\text{solute}}}{\text{kg}_{\text{solvent}}} = \dfrac{0.1155\text{ mol NaCl}}{0.09325\text{ kg H}_2\text{O}} = 1.23\ m$

b. $M = \dfrac{n_{\text{solute}}}{L_{\text{solution}}}$

$$\left(\frac{100\text{ g solution}}{}\right)\left(\frac{1\text{ mL}}{1.02\text{ g}}\right)\left(\frac{1\text{ L}}{1000\text{ mL}}\right) = 0.0980\text{ L}$$

$$M = \frac{0.1155\text{ mol NaCl}}{0.0980\text{ L}} = 1.18\ M$$

c. $\text{mol \% solute} = \left(\dfrac{n_{\text{solute}}}{n_{\text{total}}}\right)(100)$

$$\left(\frac{93.25\text{ g H}_2\text{O}}{}\right)\left(\frac{1\text{ mol H}_2\text{O}}{18.02\text{ g H}_2\text{O}}\right) = 5.175\text{ mol H}_2\text{O}$$

$$\text{mol \% solute} = \left(\frac{0.1155\text{ mol NaCl}}{0.1155 + 5.18}\right)(100) = 2.18\ \%$$

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More Practice

In the lab, you are making 0.829 *m* manganese(II) sulfate solution using 350 grams of water. How many grams of $MnSO_4$ should you add?

$$\left(\frac{350\text{ g H}_2\text{O}}{}\right)\left(\frac{1\text{ kg H}_2\text{O}}{1000\text{ g H}_2\text{O}}\right)\left(\frac{0.829\text{ mol MgSO}_4}{1\text{ kg H}_2\text{O}}\right)\left(\frac{120.364\text{ g MgSO}_4}{1\text{ mol MgSO}_4}\right) = 34.924\text{ g MgSO}_4$$

Calculate the molality of a.) 2.50 M $CaCl_2$ solution (density of solution= 1.12 g/mL), b.) 48.2 % by mass KBr solution, and c.) 64.8 grams of ethylene glycol ($C_2H_6O_2$) mixed with water to make 3,500 grams of solution.

a. $\left(\frac{1\text{ L}}{}\right)\left(\frac{1000\text{ mL}}{1\text{ L}}\right)\left(\frac{1.12\text{ g}}{1\text{ mL}}\right) = 1120\text{ g}$

$$m = \frac{n_{\text{solute}}}{\text{kg}_{\text{solvent}}} = \frac{2.50\text{ mol CaCl}_2}{1.120\text{ kg}} = 2.23\ m$$

b. Assume a 100 g solution; 48.2 g KBr and 51.8 g H_2O

$$\left(\frac{48.2\text{ g KBr}}{}\right)\left(\frac{1\text{ mol KBr}}{119.002\text{ g KBr}}\right) = 0.405\text{ mol KBr}$$

$$m = \frac{n_{\text{solute}}}{\text{kg}_{\text{solvent}}} = \frac{0.405\text{ mol KBr}}{0.0518\text{ kg H}_2\text{O}} = 7.82\ m$$

c. $\left(\frac{64.8\text{ g C}_2\text{H}_6\text{O}_2}{}\right)\left(\frac{1\text{ mol C}_2\text{H}_6\text{O}_2}{62.08\text{ g C}_2\text{H}_6\text{O}_2}\right) = 1.0438\text{ mol C}_2\text{H}_6\text{O}_2$

$$m = \frac{n_{\text{solute}}}{\text{kg}_{\text{solvent}}} = \frac{1.0438\text{ mol C}_2\text{H}_6\text{O}_2}{3.435\text{ kg H}_2\text{O}} = 0.304\ m$$

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Practice Problems

A 1.000 L aq. solution contains 170.1 g of $C_6H_{12}O_6$ (180.0 g mol^{-1}) and has a density = $\frac{1.062\text{ g}}{\text{mL}}$. Calculate the (a) molar and (b) molal concentration.

a. $M = \frac{n_{\text{solute}}}{\text{L}_{\text{solution}}} \longrightarrow \left(\frac{170.1\text{ g}}{}\right)\left(\frac{1\text{ mol}}{180.0\text{ g}}\right) = 0.945\text{ mol}$ $\quad M = \frac{n_{\text{solute}}}{\text{L}_{\text{solution}}} = \frac{0.945\text{ mol}}{1\text{ L}} = 0.945\ M$

b. $m = \frac{n_{\text{solute}}}{\text{kg}_{\text{solvent}}}$

Need kg solvent, but only given information on solution density. These are linked:

$\text{kg}_{\text{solution}} = \text{kg}_{\text{solvent}} + \text{kg}_{\text{solute}}$

$$\left(\frac{1\text{ L}}{}\right)\left(\frac{1000\text{ mL}}{1\text{ L}}\right)\left(\frac{1.062\text{ g}}{1\text{ mL}}\right) = 1.062\ x10^3\text{g} = 1.062\text{ kg (total solution)}$$

$\text{kg}_{\text{solution}} = \text{kg}_{\text{solvent}} + \text{kg}_{\text{solute}}$

$1.062\text{ kg} - 0.1701\text{ kg} = 0.8919\text{ kg}$

$$m = \frac{n_{\text{solute}}}{\text{kg}_{\text{solvent}}} = \frac{0.945\text{ mol}}{0.8919\text{ kg}} = 1.06\ m$$

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