

These lessons in this download are *beta* versions. They are being posted in draft form to provide help on as many topics as possible during this semester. If you have need for help in this topic now, they will be useful, but not all topics are covered, and they will have a higher than usual rate of typos and errors.

If you return to this topic, download a newer version of this packet at www.ChemReview.Net.

Modules 19 and above have been re-numbered.
Module 26 on Mixtures has been added and is in this packet
The former Module 26 on Equilibrium is now Module 28
If you are looking for Equilibrium topics, check Module 28
At www.ChemReview.Net

Calculations In Chemistry



Module 26

Properties of Mixtures

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Module 26 — Properties of Mixtures

Timing: Start this module when you are asked to calculate the molality of a solution, or a boiling point elevation, or a freezing point depression.

Prerequisites: This module may be started at any point after Module 11 – Molarity.

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Lesson 26A: Measures of Mixtures

Timing: Complete this lesson if you are asked to calculate the molality of a solution.

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Types of Mixtures

Recall that by our definitions for discussing matter,

- An **element** is matter that contains only one kind of atom: Ne and O₂ are elements.
- A **compound** contains more than one kind of atom and its neutral units all have the same molecular formula: the same ratio between the atoms. NaCl and H₂SO₄ are compounds.
- A **substance** may be an element or a compound, but all of the neutral units of a substance must have the same molecular formula.
- A **mixture** contains more than one substance.
- A **solution** is a homogeneous mixture in which all of its substances are in the same phase. The mixture of gases in air, or the mixture of metals in a metal alloy such as bronze (copper and tin) or brass (copper and zinc), can be considered to be solutions.

However, the type of solution encountered more often in chemistry problems is a solution with two components: a large amount of a liquid substance termed the solvent, and a substance that dissolves in the solvent termed the solute. The solute can be a gas, liquid, or solid.

Mixtures can be divided into two types: those that are *uniform* and those that are not.

- Many rocks are non-uniform: they have visibly different components in different places, and the components do not mix at room temperature.
- For a non-reacting mixture of substances in which its particles can move about freely, if the mixture is stirred, the components of the mixture will distribute themselves randomly at the molecular level. At the visible level, the result of this random distribution is a *uniform concentration*: the number of chemical particles in any visible unit of volume will be the same, within experimental error, each time we measure.

In chemistry calculations, unless otherwise noted, you should assume solutions have a *uniform* concentration: that they are well mixed before and during their use.

Molarity, Mass Fraction, Mole Fraction, and Molality

1. In most calculations involving uniform mixtures, the measurement of the composition that is most useful is the concentration in moles per liter.

$$\text{Molarity (M)} = \frac{\text{moles of substance}}{\text{Liter of mixture}} \quad (\text{Equation 1})$$

Molarity is most often associated with the concentration of aqueous solutions, but it can also be used to measure components dissolved in non-water solvents such as *n*-octane (a liquid found in gasoline) or the concentration of a gas.

For some properties of uniform mixtures, other measures of the solution composition are needed to make predictions about the solution's behavior. Three of those measures are

2. **Mass fraction** = $\frac{\text{mass of a substance}}{\text{Total mass of mixture}}$ in any consistent mass units. (Eq. 2)

3. **Mole fraction (χ)** = $\frac{\text{moles of a substance}}{\text{Total moles of particles in mixture}}$ (Equation 3)

The symbol for mole fraction χ is from the ancient Greek: a lowercase letter *chi*.

The two fractions in equations 2 and 3 measure a part of a total, so the value of the fraction must always be 0.XX... : a number between zero and one. A fraction may also be expressed as a percent, which is equal to the fraction $\times 100\%$.

4. **Molality (*m*)** = $\frac{\text{moles of substance}}{\text{kilogram of solvent}}$ (Equation 4)

For very dilute solutions, the numeric value for the molality of a solution will be very close to the molarity, but when solutions are relatively concentrated, the values for molarity and molality will differ significantly.

Molarity, mole fraction, and mass fraction can be measured for any mixture, but molality is a measure that is applied only to solutions with a liquid solvent.

Molality calculations are unusual in that, though *kilograms* of a solvent are involved, you often will not use the molar mass (grams prompt) to solve. What is more often required in molality calculations is the density of the solvent: its mass to volume ratio. For water, the most common solvent, you will be expected to have memorized:

For liquid water: 1.00 g H ₂ O = 1 mL H ₂ O and/or 1.00 kg H ₂ O = 1 L H ₂ O
--

Dilute aqueous solutions are nearly all water, and the density of those solutions is the same as for pure liquid water, within the error in most calculations. Assume:

For dilute aqueous solutions: 1.00 g soln = 1 mL soln and 1.00 kg soln. = 1 L soln.

The definitions in Equations (1) to (4) must also be memorized. Each of these ratios is a different type of measure of concentration, but if no type of concentration is specified, you should assume that "concentration" means molarity: moles/liter.

In dealing with solutions, the unit moles per liter (molarity) is abbreviated with an upper case **M**, and moles per kilogram of solvent (molality) is abbreviated with a lower case

italicized m . However, because M has many uses in chemistry, and your handwriting may not easily distinguish between lower vs. upper case and regular vs. italicized letters, you may want to write out *molaR* and *molaL* during problem sets that use both.

Calculating Concentration In Non-Molar Units

The four properties above can be solved using conversions *or* equations, but using our equation method will help to separate these calculations into manageable pieces. For consistency in this module, we will solve for all four properties by equations.

Each of the four measures is an equation with a simple ratio. Our method to solve equations uses these same steps every time:

- Write the symbol or term and unit WANTED.
- Write the equation needed to solve. If you are not sure which equation to use, list the supplied DATA, assign symbols to each item, and recall which memorized equation relates those symbols.
- List each symbol or term in the equation in a DATA table. After the term or symbol, write the *unit* needed for consistent units or the units the equation requires.
- IN the DATA table, for each term, convert each supplied item of DATA to the unit needed for that term. Use conversions or equations. Getting the DATA ready in the DATA table simplifies the SOLVE step.
- Solve the equation for the WANTED symbol in symbols, *then* plug in the DATA.

For the four equations above:

- If an *amount* of a mixture is not specified, assume a 1.00 Liter or 100. gram sample – whichever simplifies the math of your first calculation. Since all 4 of these measures are ratios, a calculation can be based on any convenient-sized sample.
- *Label* each item of data with units, molecular formulas, and distinguishing words.

Apply those steps to this example.

- Q1.** Alcoholic beverages are a mixture of water and ethyl alcohol (formal name *ethanol*, formula C_2H_5OH , often abbreviated as EtOH). In the United States, the “proof” of an alcoholic beverage is the double the percent ethanol by volume: 100 proof means the mixture is 50% ethanol by volume and 50% water. The density of pure ethanol at 4°C is 0.790 g/mL. In 1.00 liters of 24.0 proof wine (an ethanol and water mixture with small amounts of other substances) at 4°C, calculate:
- a. The volume of alcohol and volume of water, in mL.
 - b. The mass fraction of the wine that is ethanol.
 - c. The mole fraction that is ethanol.
 - d. The molality of the ethanol solution.
 - e. The molarity of the ethanol.

Check your answers after each part. If you get stuck, look at a part of the answer below, then try again.

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Part a. By volume, 24.0 proof is 12.0% ethanol (half the proof), so 88.0% is water.

In 1.00 L, the 12.0% = 0.12 fraction of the 1.00 L volume that is ethanol =
= 0.120 L = **120. mL EtOH**. The remaining **880. mL** must be **water**.

(If needed, see Lesson 9A on fractions and percent.)

Part b. WANTED: Mass fraction that is EtOH = ? The equation is

$$\text{Mass fraction} = \frac{\text{mass of substance}}{\text{Total mass of mixture}} \quad \text{in any consistent mass units.}$$

DATA: List separately and solve for each term in the equation -- one at a time.

Mass fraction EtOH = ?

$$\text{Mass of ethanol} = 120. \text{ mL EtOH} \cdot \frac{0.790 \text{ g EtOH}}{1 \text{ mL EtOH}} = \mathbf{94.8 \text{ g EtOH}}$$

Total mass of mixture = mass of ethanol + mass of water = ?

Mass of water: At 4°C, water has a density of 1.00 g/mL, so 880. mL = **880. g H₂O**

$$\text{Total mass of mixture} = 94.8 \text{ g EtOH} + 880. \text{ g H}_2\text{O} = \mathbf{975 \text{ g total}}$$

SOLVE:

$$\text{Mass fraction EtOH} = \frac{\text{mass of EtOH}}{\text{Total mass soln.}} = \frac{94.8 \text{ g EtOH}}{975 \text{ g total}} = \mathbf{0.0972 \frac{\text{g EtOH}}{\text{g soln.}}}$$

Part c. WANTED: **Mole fraction EtOH = ?** The equation is

$$\text{Mole fraction } (\chi) = \frac{\text{moles of substance}}{\text{Total moles of particles in mixture}}$$

DATA: List separately and solve for each term in the equation -- one at a time.

Molar mass of C₂H₅OH = 46.0 g/mol

$$\text{Moles of ethanol} = 94.8 \text{ g EtOH} \cdot \frac{1 \text{ mol EtOH}}{46.0 \text{ g EtOH}} = \mathbf{2.06 \text{ mol EtOH}}$$

$$\text{Moles of H}_2\text{O} = 880. \text{ g H}_2\text{O} \cdot \frac{1 \text{ mol H}_2\text{O}}{18.0 \text{ g H}_2\text{O}} = \mathbf{48.9 \text{ mol H}_2\text{O}}$$

Total moles of particles in mixture = 2.06 + 48.9 = 50.96 = **51.0 moles total**

$$\text{Mole fraction } (\chi) = \frac{\text{moles of EtOH}}{\text{total moles}} = \frac{2.06 \text{ mol EtOH}}{51.0 \text{ moles total}} = \mathbf{0.0404 \frac{\text{mol EtOH}}{\text{mol total}}}$$

Part d. WANTED: **Molality EtOH = ?** The equation is

$$\text{Molality } (m) = \frac{\text{moles of solute}}{\text{kg of solvent}}$$

For the molality of *ethanol*, ethanol must be the solute, so water is the solvent.

DATA: **Moles of ethanol = 2.06 mol EtOH** from above.

$$\text{kg of solvent H}_2\text{O} = 880. \text{ g H}_2\text{O} \cdot \frac{1 \text{ kg}}{10^3 \text{ g}} = \mathbf{0.880 \text{ kg H}_2\text{O}}$$

SOLVE:

$$\text{Molality EtOH} = \frac{\text{moles EtOH}}{\text{kg of solvent H}_2\text{O}} = \frac{2.06 \text{ mol EtOH}}{0.880 \text{ kg H}_2\text{O}} = 2.34 \text{ m EtOH}$$

Part e. WANT: **Molarity EtOH = ?** The equation is

$$\text{Molarity (M)} = \frac{\text{moles of solute}}{\text{Liters of soln.}} = \frac{2.06 \text{ mol EtOH}}{1.00 \text{ L soln.}} = 2.06 \text{ M EtOH}$$

Practice: Memorize Equations (1) - (4), then do two problems below. Save one for your next practice session.

- How many grams of H_2SO_4 must be added to 0.500 L of water to form 4.00 *m* H_2SO_4 ?
- Find χ for the solution in Problem 1.
- Concentrated hydrochloric acid (a solution of the gas HCl and water) is typically sold as a 12.0 M HCl solution with a solution density of 1.19 g/mL. Calculate
 - The mass fraction that is HCl.
 - The mole fraction HCl
 - The HCl molality.
 - Which should be higher: the HCl molarity or molality value? Which was higher?

ANSWERS

1. WANT: ? g H_2SO_4 DATA: 98.1 g $\text{H}_2\text{SO}_4 = 1 \text{ mol H}_2\text{SO}_4$ (grams prompt)

The molality is known. The one equation we know that uses molality is

$$\text{Molality H}_2\text{SO}_4 = \frac{\text{moles H}_2\text{SO}_4}{\text{kg of solvent H}_2\text{O}}$$

That equation does not solve for grams H_2SO_4 , but it includes moles H_2SO_4 , and moles to grams we can do. List the terms in the equation, solve for the *term* missing in symbols, then substitute data.

$$\text{Molality H}_2\text{SO}_4 = 4.00 \text{ m H}_2\text{SO}_4 = 4.00 \text{ mol H}_2\text{SO}_4 / \text{kg H}_2\text{O}$$

$$\text{Moles H}_2\text{SO}_4 = ?$$

$$\text{kg of solvent H}_2\text{O} = 0.500 \text{ L} \cdot \frac{1.00 \text{ kg H}_2\text{O}}{1 \text{ L H}_2\text{O}} = 0.500 \text{ kg H}_2\text{O}$$

For molality calculations, it is helpful to memorize the equality: 1.00 kg liquid $\text{H}_2\text{O} = 1.00 \text{ L H}_2\text{O(l)}$

Solving the molality equation for the unknown symbol:

$$? = \text{mol H}_2\text{SO}_4 = (\text{Molality H}_2\text{SO}_4)(\text{kg H}_2\text{O}) = \frac{4.00 \text{ mol H}_2\text{SO}_4}{\text{kg H}_2\text{O}} \cdot 0.500 \text{ kg H}_2\text{O} = 2.00 \text{ mol H}_2\text{SO}_4$$

$$\text{WANTED} = ? \text{ g H}_2\text{SO}_4 = 2.00 \text{ mol H}_2\text{SO}_4 \cdot \frac{98.1 \text{ g H}_2\text{SO}_4}{1 \text{ mol H}_2\text{SO}_4} = 196 \text{ g H}_2\text{SO}_4$$

2. WANT: $\chi_{\text{H}_2\text{SO}_4}$ = mole fraction H_2SO_4 The equation is

$$\text{Mole fraction } (\chi) = \frac{\text{moles of substance}}{\text{Total moles of particles in mixture}}$$

Strategy: List DATA, then list and solve the 3 terms of the equation separately.

DATA: moles of substance = mol H_2SO_4 = **2.00 mol H_2SO_4** from Problem 1 .

$$\text{mol of H}_2\text{O} = 0.500 \text{ kg H}_2\text{O} \cdot \frac{10^3 \text{ g}}{1 \text{ kg}} \cdot \frac{1 \text{ mol H}_2\text{O}}{18.0 \text{ g H}_2\text{O}} = \mathbf{27.8 \text{ mol H}_2\text{O}}$$

Total moles in mixture = 2.00 + 27.8 = **29.8 moles total**

$$\text{SOLVE: } \chi_{\text{H}_2\text{SO}_4} = \frac{\text{mol H}_2\text{SO}_4}{\text{Total moles of particles in mixture}} = \frac{\mathbf{2.00 \text{ mol}}}{\mathbf{29.8 \text{ mol}}} = \mathbf{0.0671 \frac{\text{mol H}_2\text{SO}_4}{\text{mol mixture}}}$$

3a. WANTED: Mass fraction HCl = ? The equation is

$$\text{Mass fraction} = \frac{\text{mass of substance}}{\text{Total mass of mixture}} \text{ in any consistent mass units}$$

Strategy: List DATA, then list and solve the 3 terms of the equation separately.

DATA: 12.0 mol HCl = 1 L solution

1.19 g soln = 1 mL soln.

36.5 g HCl = 1 mol HCl (*mass of substance* in equation = grams prompt)

Strategy: Assume a sample size that simplifies the math. Since you want a mass measure, it is easier to assume a volume: assume exactly 1 liter of soln.

$$? \text{ Total g of mixture} = \mathbf{1 \text{ liter}} \cdot \frac{\mathbf{1 \text{ mL}}}{10^{-3} \text{ L}} \cdot \frac{\mathbf{1.19 \text{ g soln}}}{1 \text{ mL soln}} = \mathbf{1,190 \text{ g soln.}}$$

$$? \text{ g of HCl} = \mathbf{1 \text{ liter}} \cdot \frac{\mathbf{12.0 \text{ mol HCl}}}{1 \text{ L HCl}} \cdot \frac{\mathbf{36.5 \text{ g HCl}}}{1 \text{ mol HCl}} = \mathbf{438 \text{ g HCl}}$$

$$\text{SOLVE: Mass fraction HCl} = \frac{\text{mass of HCl}}{\text{Total mass soln.}} = \frac{\mathbf{438 \text{ g HCl}}}{\mathbf{1,190 \text{ g soln.}}} = \mathbf{0.368 \frac{\text{g HCl}}{\text{g soln.}}}$$

3b. WANT: χ_{HCl} = **Mole fraction** HCl = ? The equation is

$$\text{Mole fraction} = \frac{\text{mole of substance}}{\text{Total moles of particles in mixture}}$$

Strategy: Solve the 3 terms of the equation separately using conversions.

DATA: Use the DATA and answers from Part A. Assume the same exactly 1 liter of solution.

$$\mathbf{\text{Mol of substance HCl}} = \mathbf{1 \text{ liter}} \cdot \frac{\mathbf{12.0 \text{ mol HCl}}}{1 \text{ L HCl}} = \mathbf{12.0 \text{ mol HCl}}$$

Since the mixture total mass is 1,190 g, of which 438 g is HCl and the rest is water,

Grams water = 1,190 — 438 = 752 g H_2O

$$\text{mol H}_2\text{O} = 752 \text{ g H}_2\text{O} \cdot \frac{1 \text{ mol H}_2\text{O}}{18.0 \text{ g H}_2\text{O}} = \mathbf{41.8 \text{ mol H}_2\text{O}}$$

Total moles of particles in mixture = 12.0 mol HCl + 41.8 mol H₂O = **53.8 moles total**

SOLVE: Mole fraction HCl = $\frac{\text{mol HCl}}{\text{Total mol soln.}} = \frac{12.0 \text{ mol HCl}}{53.8 \text{ mol total}} = \mathbf{0.223 \frac{\text{mol HCl}}{\text{mol soln.}}}$

Check: In parts a and b, the fractions have a value between 0 and 1, as they must.

3c. WANT: Molality HCl = $m \text{ HCl} = ?$ The equation is: Molality (m) = $\frac{\text{moles of solute}}{\text{kg of solvent}}$

For a molality of HCl, HCl is the solute and H₂O the solvent.

Assume the same exactly 1 liter solution. Use the DATA and answers above.

Make a DATA table with the terms in the equation and conversions to the needed units.

DATA: **Moles of HCl = 12.0 mol HCl** from above.

kg of solvent H₂O = 752 g H₂O $\cdot \frac{1 \text{ kg}}{10^3 \text{ g}} = \mathbf{0.752 \text{ kg H}_2\text{O}}$

SOLVE: **Molality HCl** = $\frac{\text{moles HCl}}{\text{kg of solvent H}_2\text{O}} = \frac{12.0 \text{ mol HCl}}{0.752 \text{ kg H}_2\text{O}} = \mathbf{16.0 \text{ } m \text{ HCl}}$

3d. In concentrated solutions, the molaLity value is higher than the molaRity, and **16.0 m HCl < 12.0 M HCl**.

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Lesson 26B: Parts Per Million

Timing: Complete this lesson *only* if you are assigned calculations that include *ppm*.

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Low But Important Concentrations

For substances that are have high impact even in low concentration, such as toxic substances in the environment, concentrations are often expressed in **parts per million (ppm)**: the parts of a substance per million total parts. Parts per million can be a volume ratio, but unless otherwise specified, assume that ppm is a mass ratio.

Parts per million can be defined with an *equation* that includes a fraction:

Parts per million of a substance in a mixture = $\mathbf{\text{ppm} = \frac{\text{grams substance} \times 10^6}{\text{grams total}}}$

A measurement expressed as a ppm is similar to a percent:

- A percent is (part/total) times 100 ;
- A ppm is (mass part/mass total) times 1,000,000 (10⁶) .

A percent allows a fraction with a value between 0.01 and one to be expressed as number between 1 and 100. A ppm allows a fraction with a value between 0.000 001 and 0.000 1 to be expressed as a number between one and 100. Both are methods to express concentration without using small decimal equivalents or negative exponentials.

When % or PPM is WANTED

Conversion calculations are done using fractions. Our rule (from Lesson 9A) is:

If *percent* or *ppm* is WANTED, first solve for part/total (the *fraction*), then multiply by 100 to find percent or 1,000,000 to find ppm.

Using those steps, try this sample calculation.

- Q.** In earth's atmosphere, at room temperature and standard pressure, air has an average mass of 29.0 grams per 24.0 liters. The carbon dioxide in air has a current average concentration of 0.471 mg/liter. What is this concentration of CO₂ in parts per million?

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$$\text{WANTED: } \quad ? \text{ parts per million CO}_2 \text{ in air} = \frac{? \text{ grams CO}_2}{\text{gram air}} \times 10^6$$

The rule is: if % or ppm is WANTED, solve for the fraction first.

$$\begin{aligned} \text{DATA:} \quad 29.0 \text{ g air} &= 24.0 \text{ L air RT\&SP} && (\text{gas volumes need a P and t}) \\ 0.471 \text{ mg CO}_2 &= 1 \text{ L air RT\&SP} \end{aligned}$$

The labels distinguishing CO₂ and *air* are needed to arrange the conversions.

Solve for the fraction as a ratio unit, then convert to ppm.

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$$\text{SOLVE: } \frac{? \text{ g CO}_2}{\text{g air}} = \frac{0.471 \text{ mg CO}_2}{1 \text{ L air}} \cdot \frac{10^{-3} \text{ g}}{1 \text{ mg}} \cdot \frac{24.0 \text{ L air RT\&SP}}{29.0 \text{ g air}} = \frac{0.390 \times 10^{-3} \text{ g CO}_2}{\text{g air}}$$

Your conversions may be in a different order, but should be "right-side up" compared to those above. Now use the definition to change the fraction to ppm.

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$$\text{ppm} = \frac{\text{g substance}}{\text{g total}} \times 10^6 = \frac{0.390 \times 10^{-3} \text{ g CO}_2}{\text{g air}} \times 10^6 = 390. \text{ ppm CO}_2$$

When DATA Is Supplied As % or PPM

In the special case where percent or ppm is supplied as DATA, you can either

- Start by converting to the fraction using the definition equations, then use the fraction in conversions,

$$\text{ppm} \equiv \frac{\text{g substance}}{\text{g total}} \times 10^6 ; \quad \text{fraction} \equiv \frac{\text{g substance}}{\text{g total}} = (\text{ppm} \times 10^{-6}) \frac{\text{g substance}}{\text{g total}}$$

- *or* (this is easier) write the % or ppm in the DATA as a conversion in this format.

If "X % by mass" is supplied, write as DATA: **X g substance = 100 g total**

If "X ppm" is supplied, write as DATA: **X grams substance = 10⁶ grams total**

This “write the equality” method works because a

- percent is the number of parts per hundred parts total, and
- ppm is the number of grams per million grams total.

By using this equality format for % and ppm DATA, calculations can be solved without converting the percent or ppm to a decimal fraction first.

The same approach would apply to a *volume* % or ppm, using mL or L instead of grams or kilograms.

Apply that special rule for ppm DATA to this problem.

Q2. For dilute aqueous solutions, 1.0 ppm of a substance is how many mg/L? (Hint: use the density of a dilute aqueous soln.)

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WANT: ? $\frac{\text{mg substance}}{\text{L aq. soln.}}$

DATA: 1.0 g substance = 10^6 g total solution (using the ppm DATA equality)
 1.00 g soln. = 1 mL soln. (density of dilute aqueous solution)

SOLVE:

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$$\begin{aligned} ? \frac{\text{mg substance}}{\text{L aq. soln.}} &= \frac{1.0 \text{ g substance}}{10^6 \text{ g solution}} \cdot \frac{1 \text{ mg subst.}}{10^{-3} \text{ g subst.}} \cdot \frac{1.00 \text{ g soln.}}{1 \text{ mL soln.}} \cdot \frac{1 \text{ mL soln.}}{10^{-3} \text{ L soln.}} = \\ &= 1.0 \frac{\text{mg substance}}{\text{L if dilute aq. soln.}} \end{aligned}$$

Since our *given* quantity was 1.0 ppm, this answer provides a useful equality:

IF a dilute aqueous solution: # ppm = # mg substance/L

In dilute aqueous solutions, the *number* of ppm and the *number* of mg substance/L soln will be the same. In calculations involving ppm and aqueous solutions, writing the ppm as the mg/L in the DATA will often provide answers using fewer conversions.

For substances that are *not* in dilute aqueous solutions, such as mixtures of gases or highly concentrated aqueous solutions, the 1.00 g solution = 1 mL equality is *not* true, and # ppm \neq # mg/L. In the DATA for those cases, write: ppm# g substance = 10^6 g total .

To summarize:

$$\text{ppm} \equiv \frac{\text{g substance}}{\text{g total}} \times 10^6$$

When ppm is WANTED, find the fraction, then use this definition to change to ppm.

In DATA:

- List ppm# g substance = 10^6 g total
- Or, for a dilute aqueous solution, list ppm# mg substance = 1 L soln

and solve with conversions.

Practice: First learn the rules above, then do the problems below.

- In a trout stream where the water temperature is 5°C, dissolved oxygen has a concentration of 0.40 mmol/liter. What is this concentration of O₂
 - in parts per million?
 - in mg/L?
- If a dilute saline solution is composed of 550 ppm NaCl, find the [NaCl] in mmol/dL.

ANSWERS

1a. WANTED: $\frac{? \text{ parts per million O}_2 \text{ in water}}{\text{gram H}_2\text{O soln.}} = \frac{? \text{ grams O}_2}{\text{gram H}_2\text{O soln.}} \times 10^6$

The rule is: if % or ppm is WANTED, solve for the fraction first.

DATA: $0.40 \text{ mmol O}_2 = 1 \text{ L H}_2\text{O soln}$

Since *grams* H₂O and O₂ are wanted, you may need the molar masses of both (g prompts).

$$32.0 \text{ g O}_2 = 1 \text{ mol O}_2$$

$$18.0 \text{ g H}_2\text{O} = 1 \text{ mol H}_2\text{O} \quad (\text{For } m, \text{ you may need solvent density instead})$$

A small number of mol/L in water means a dilute aqueous solution. You can assume that

$$1.00 \text{ g dilute aqueous solution} \approx 1.00 \text{ g H}_2\text{O} = 1 \text{ mL solution} \text{ is the density.}$$

Use conversions to solve for the ratio unit *fraction*, then convert to ppm.

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$$\begin{aligned} \text{SOLVE: } \frac{? \text{ g O}_2}{\text{g H}_2\text{O}} &= \frac{0.40 \text{ mmol O}_2}{1 \text{ L H}_2\text{O}} \cdot \frac{10^{-3} \text{ mol O}_2}{1 \text{ mmol O}_2} \cdot \frac{32.0 \text{ g O}_2}{1 \text{ mol O}_2} \cdot \frac{10^{-3} \text{ L H}_2\text{O}}{1 \text{ mL H}_2\text{O}} \cdot \frac{1 \text{ mL dil. aq. soln.}}{1.00 \text{ g H}_2\text{O soln.}} \\ &= 13 \times 10^{-6} \text{ g O}_2 / \text{g H}_2\text{O soln.} \quad \text{Now convert from the fraction to ppm.} \end{aligned}$$

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$$\text{ppm} = \frac{\text{g substance}}{\text{g total}} \times 10^6 = \frac{13 \times 10^{-6} \text{ g O}_2}{\text{g H}_2\text{O soln.}} \times 10^6 = 13 \text{ ppm O}_2$$

1b. WANT: **mg O₂ / L** in a dilute aqueous solution

The rule for dilute aqueous solutions is: # ppm = # mg/L, so the answer should be 13 mg/L. But let's use the data above to be sure.

$$\text{SOLVE: } \frac{? \text{ mg O}_2}{\text{L H}_2\text{O}} = \frac{0.40 \text{ mmol O}_2}{1 \text{ L H}_2\text{O}} \cdot \frac{10^{-3} \text{ mol O}_2}{1 \text{ mmol O}_2} \cdot \frac{32.0 \text{ g O}_2}{1 \text{ mol O}_2} \cdot \frac{1 \text{ mg O}_2}{10^{-3} \text{ g O}_2} = 13 \text{ mg O}_2 / \text{L H}_2\text{O}$$

2. WANT: $\frac{? \text{ millimoles NaCl}}{\text{deciliter dilute aqueous solution}}$

DATA: In a dilute aqueous solution, when ppm DATA is supplied, you can use the ppm DATA rule:
 $550 \text{ g NaCl} = 10^6 \text{ g total}$
 But using the rule # ppm = # mg substance/L soln. will usually solve faster.
 $550 \text{ mg NaCl} = 1 \text{ L dilute aq. soln}$ Both ratios work to solve. We will use this one.
 $58.5 \text{ g NaCl} = 1 \text{ mol NaCl}$ (the mg NaCl in the DATA ratio is a grams prompt)

SOLVE: Want a ratio? Pick a ratio to start from.

* * * * *

$$\frac{? \text{ millimole NaCl}}{\text{deciliter dilute soln}} = \frac{550 \text{ mg NaCl}}{1 \text{ L soln}} \cdot \frac{1 \text{ mol NaCl}}{58.5 \text{ g NaCl}} \cdot \frac{10^{-1} \text{ L soln}}{1 \text{ dL soln.}} = \frac{0.94 \text{ mmol NaCl}}{\text{dL soln.}}$$

Since milli- is wanted, and m- is on top in the first conversion, and the prefixes are independent of the numbers and units, you can leave the m- in the conversions on top where it is WANTED.

Numbers, prefixes, and exponentials can all be grouped separately in the WANTED unit and in the conversions. Like units on the top and bottom cancel if multiplied. If a prefix is part of the WANTED unit, and the prefix appears in the conversions on top or bottom where it is WANTED, you may leave it un-cancelled.

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Lesson 26C: Colligative Properties And Gas Pressures

Timing: Complete this lesson *if* you are assigned calculations that include Raoult's law or Henry's law.

* * * * *

Vapor Pressure Review

Recall that all liquids (and some solids) have a significant tendency to release molecules into the gas or vacuum above them. This tendency is measured by the vapor pressure of the substance: the gas pressure exerted by the molecules of the gas above the liquid or solid at equilibrium (equilibrium meaning that the container is closed, the temperature is constant, and net changes in the system are not occurring).

The vapor pressure of a substance

- Depends on temperature: as temperature increases, the vapor pressure of a substance increases.
- Is characteristic: a given substance at a given temperature will always have the same vapor pressure.

A liquid boils when its vapor pressure is the same as the atmospheric pressure in the gas above it.

Mixtures of gases (if they do not react with each other) obey Dalton's law: the total pressure of a mixture of gases is equal to the sum of the pressures that the gases would exert alone. The pressure exerted by each gas in a mixture is its *partial pressure*: the pressure that the gas would exert in the given volume if it alone were present.

Dalton's Law Equations

For the partial pressure of a single gas in a mixture:

$$P_{\text{gas one}} = (\text{mole fraction gas one})P_{\text{total}} = (\text{volume fraction gas one})P_{\text{total}}$$

For a mixture of gases:

$$\begin{aligned} P_{\text{total}} &= P_{\text{gas one}} + P_{\text{gas two}} + \dots \quad (= \text{Partial } P \text{ gas one} + \text{Partial } P \text{ gas two} + \dots) \\ &= (\text{mole fraction gas one})P_{\text{total}} + (\text{mole fraction gas two})P_{\text{total}} + \dots \\ &= (\text{volume fraction gas one})P_{\text{total}} + (\text{volume fraction gas two})P_{\text{total}} + \dots \end{aligned}$$

Gas calculations frequently involve gases collected over water, where the gas mixture contains the gas formed plus water vapor. The vapor pressure of the water at a given temperature can be found in tables and subtracted from the total pressure to give the partial pressure of the gas formed.

Colligative Properties

Colligative properties are those that depend on the relative numbers of particles in a mixture, as measured by the molarity, molality, or mole fraction.

Among the colligative properties are:

- For a gas dissolved in a solvent at a given temperature, the partial pressure of a gas above the solution varies in proportion to the *mol/L* of the gas in the solution (and the reverse is true).
- If a non-volatile (non-gas-forming) solute is dissolved in a pure liquid solvent, at a given temperature, the vapor pressure of the solvent above the solution is proportional to the *mole fraction* of particles that are solvent in the solution.
- If a solute is dissolved in a pure liquid solvent, it lowers the freezing point of the solution, relative to the freezing point of the pure solvent, by an amount proportional to the molality (*m*) of the solute.
- If a non-volatile solute is dissolved in a pure liquid solvent, it elevates the boiling point of the solution, relative to the boiling point of the pure solvent, by an amount proportional to the molality (*m*) of the solute.
- The mol/L of particles dissolved in a solution can be used to predict the osmotic pressure of a solvent that can pass through a semi-permeable membrane.

Let's consider these properties one at a time.

Solubility of Gases in Solutions (Henry's Law)

The solubility of a gas in a liquid, in moles per liter, varies depending on the partial pressure of that gas above the liquid. The equation for this relationship termed

Henry's law: $\text{Concentration}_{\text{gas in solution}} = k_{\text{henry}} \cdot P_{\text{gas above}}$	(Equation 6)
--	--------------

where k_{henry} is the Henry's law constant, a value that stays constant at a given temperature for a given solvent and gas.

Henry's law has the form $y = (\text{constant}) x$; a direct proportion. This means that at constant temperature, for a mixture at equilibrium,

- as the pressure of a gas above a liquid doubles, the concentration of the gas dissolved in the liquid doubles;
- the ratio of the partial pressure to the mol/L that dissolves is constant.

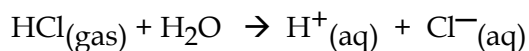
Henry's law can be used for any consistent units of concentration: any moles or mass units of gas per any units of solution volume.

Henry's law can also be written with the y and x terms reversed, as

$$P_{\text{gas above}} = k_{\text{henry}} \cdot \text{Concentration}_{\text{gas in solution}} .$$

Which form of the equation is being used can be identified by the units of k_{henry} and the logic of unit cancellation: Equation (6) above will have concentration units on top, the second form must have pressure units on top.

Henry's law does not work for gases that *react* with the solvent. For example, when HCl gas dissolves in water, it reacts ~100% to form ions.



When a covalent gas molecule reacts to form ions, the particles no longer behave as a dissolved gas in the solution, and Henry's law does not make accurate predictions about the behavior of the gas.

However, if a reaction of a gas with a solvent is slight, Henry's law results in reasonably accurate predictions. For example, carbon dioxide gas reacts with water, but only slightly:



For the case of carbon dioxide dissolving in water, Henry's law is relatively accurate.

Practice A: Memorize Henry's law and apply it to this problem.

1. At room temperature, if the solubility of CO_2 gas in water at 25°C is 1.45 g/L when the gas above the water is at 1.0 atm pressure, what is the solubility of CO_2 at 4.0 atm pressure?
2. In a trout stream at standard pressure, where the water temperature is 5°C , dissolved oxygen has a concentration of 0.40 mmol/liter . If 21% of the volume of air is O_2 , what is the Henry's law constant for O_2 at standard pressure and 5°C , in $\text{mol/kPa} \cdot \text{L}$?

Vapor Pressure Above A Solution (Raoult's Law)

If a relatively non-volatile (non-gas forming) substance is added to a volatile liquid at a constant temperature, the vapor pressure of the solution is lowered compared to the vapor pressure of the pure solvent. The equation that calculates the resulting vapor pressure is

$$\text{Raoult's law: } P_{\text{above soln.}} = \chi_{\text{volatile solvent particles}} \cdot P^{\circ}_{\text{solvent}} \quad (\text{Eq. 7})$$

where $P^{\circ}_{\text{solvent}}$ = the vapor pressure of the *pure* solvent liquid at that temperature.

Raoult's law is logical: it means that if only 80% of the particles in a solution are made of a volatile (gas forming) liquid, and the other particles do not form a gas, then the vapor pressure above the solution will be 0.80 times the vapor pressure of the pure liquid.

Practice B: Memorize Raoult's law, then apply it to this problem.

1. The vapor pressure of pure water at 80.°C is 47.2 kPa. The antifreeze *ethylene glycol* ($\text{C}_2\text{H}_6\text{O}_2$, molar mass = 62.0 g/mol) is relatively non-volatile compared to water (meaning it boils at a much higher temperature). If 1.00 kg of ethylene glycol is mixed with 2.00 kg of water, what will be the vapor pressure of the mixture at 80.°C?

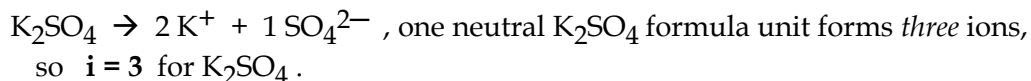
The van't Hoff Factor

Solvents are liquids, and liquids ionize only very slightly at normal laboratory temperatures. However, many *solutes* ionize when they dissolve in a solvent. Other solutes dissolve, but do not ionize. For example,

- sugars and alcohols dissolve in water because they are polar, but they are not ionic. They dissolve, but they do not ionize, in water.
- When a soluble ionic compound (such as NaCl) dissolves to form a dilute aqueous solution, the compound separates nearly 100% to form ions.

An aqueous solution of an soluble ionic compound is termed an **electrolyte** because it conducts electricity. One way to calculate the number of particles in an electrolyte is to use the **van't Hoff factor (i)**: defined as the number of separate particles that a single neutral formula unit of a soluble compound forms when it dissolves in water.

For example: for these solutes that dissolve and ionize ~100% in dilute aqueous solutions,



For substances that dissolve in a solvent but do not ionize, for each particle of a soluble solute added, one solute particle exists in the solution, so $\mathbf{i = 1}$ for soluble non-electrolytes such as sugars and alcohols.

In Raoult's law, we must count the moles of all of the separate particles in the solution: both solvent and solute. The rules we need are logical:

Moles of solute particles in solution = (i) (moles of *un-ionized* solute added)

For example, if 1.0 mol NaCl is added to water, since $i = 2$, **2.0** moles of solute particles are formed in the water.

Total moles of particles in a solution = moles of solvent + moles of solute *particles*
 = moles of solvent + (i) (moles of *un-ionized* solute added)

The i values can also be used to calculate the molarity and molality of solute particles.

Molarity_{solute particles in solution} = (i) Molarity_{*un-ionized* solute}

In 0.10 M NaCl, [solute particles] = [ions] = (**2**)(0.10 M) = 0.20 M solute particles

Molality_{solute particles in solution} = (i) Molality_{*un-ionized* solute}

In 0.30 *m* NaCl, $m_{\text{solute particles}}$ = m_{ions} = (**2**)(0.30 *m*) = 0.60 *m* solute particles

Practice C: First learn the rules for the van't Hoff factor, then try these.

- Predict the van't Hoff factor for each of these substances in dilute solutions.
 - NaBr
 - KNO₃
 - Na₃PO₄
 - O₂
- What will be the [solute particles] in these aqueous solutions?
 - 0.20 M KCl
 - 0.12 M K₃PO₄
- If [solute particles] = 1.0 M in Na₂SO₄ solution, [Na₂SO₄] = ?
- In 0.25 L of a 0.80 M NaCl solution, how many moles of ions are present?
- If the mole fraction of an aqueous KCl solution is 0.10 mol KCl/mole total, what is the mole fraction of ions in the solution?

Raoult's Law For Electrolyte Solutions

Raoult's law and other colligative properties depend on the fraction of *particles* that are solute and solvent. If a solute particle separates into ions when it dissolves, this separation affects the fraction of particles that are solvent. A solvent molecule is a particle, and each separate *ion* is a particle, in the measurements that decide colligative properties.

For Raoult's law and other colligative properties, the rule is

Each ion is a separate particle when calculating the moles, molarity, or molality of solute particles in the solution.

Keep the above in mind during the following example.

- Q.** Predict the vapor pressure (in kPa) of an aqueous solution that combines 17.0 moles of water with 1.00 moles of K₂SO₄ at 100.0 °C.

* * * * *

WANT: $P_{\text{above soln.}}$ in kPa To find the vapor pressure above a *solution* that contains a non-volatile solute, we use

$$\text{Raoult's law: } P_{\text{above soln.}} = \chi_{\text{solvent (volatile) particles}} \cdot P^{\circ}_{\text{solvent}}$$

Set up a data table that contains the equation's symbols.

DATA: $P_{\text{sbove soln.}} = ?$ kPa

$$\chi_{\text{particles that are solvent}} = ? = \text{mol solvent} / \text{Total moles of soln. particles}$$

Solve the $\chi_{\text{particles}}$ equation one term at a time.

Moles solvent particles = **17.0 mol water.**

Moles *solute* particles = **(3) 1.00 moles K_2SO_4 = 3.00 mol solute particles**

(In water, for every 1 mol K_2SO_4 added, 3 moles of ions form, so $i = 3$ for K_2SO_4 .)

Total moles particles = 17.0 mol water + 3.00 mol solute particles = **20.0 mol total**

$$\chi_{\text{particles that are solvent}} = ? = 17.0 \text{ mol solvent} / 20.0 \text{ moles total} =$$

$$= \mathbf{0.850 \text{ mol particles solvent/mol particles total}}$$

$$P^{\circ}_{\text{solvent}} = ? \quad \text{At } 100.0 \text{ }^{\circ}\text{C}, \text{ what is the vapor pressure of pure water in kPa?}$$

* * * * *

By definition, $100.0 \text{ }^{\circ}\text{C}$ = the boiling point of water at standard pressure (101 kPa).

By definition, a liquid boils when its vapor pressure = atmospheric pressure

So, at $100.0 \text{ }^{\circ}\text{C}$, the vapor pressure of pure water must be **101 kPa.**

If needed, adjust your work and complete the problem.

* * * * *

$$\text{SOLVE: } P_{\text{above soln.}} = \chi_{\text{solvent (volatile) particles}} \cdot P^{\circ}_{\text{solvent}}$$

$$= 0.850 \cdot 101 \text{ kPa} = \mathbf{85.8 \text{ kPa}} = \mathbf{\text{water vapor pressure above soln.}}$$

Actual van't Hoff Values

The van't Hoff factors above are *predicted* values based on an assumption of 100% ionization. For very dilute solutions, the predicted van't Hoff factor is generally accurate, but in more concentrated solutions, or for ionic compounds that have moderate covalent character, the *actual* value for the van't Hoff factor will be less than predicted.

One way to determine these actual van't Hoff factors is to the actual value of colligative properties experimentally.

For example, in NaCl, the predicted $i = 2$, in 0.001 M NaCl, the experimental $i = 1.97$; in more concentrated 0.10 M NaCl, the experimental $i = 1.87$.

The values indicate that these ions in water are most stable when they are separated and surrounded by water molecules, but that the Na^+ and Cl^- ions also remain attracted to each other. The ions behave as if they are separated *most* of the time, but not all the time.

In calculations, however, unless otherwise specified, you should use the whole-number i values predicted by the rules above.

Practice D

- In a solution that combines 12.0 moles of water with 1.0 mole of a soluble ionic solid, the vapor pressure of the solution at 25 °C is 17.8 mm Hg. If the vapor pressure of pure water at 25 °C is 23.8 mm Hg, what is the experimental van't Hoff factor for the ionic solid?

ANSWERS

Practice A

- Henry's law is a direct proportion. If the pressure of the gas above the liquid is quadrupled, the concentration of the gas in the solution will be quadrupled when an equilibrium is reached.

1.45 g/L x 4 = **5.80 g CO₂/L** dissolves at 4.0 atm pressure.

- WANT: k_{henry} in mol/kPa • L = ?

DATA: Henry's law: $\text{Concentration}_{\text{gas in solution}} = k_{\text{henry}} \cdot P_{\text{gas above}}$

Make a table with those symbols. Convert DATA to the WANTED units.

k_{henry} in mol/kPa = ? Convert the DATA to moles, liters, and kPa

Concentration_{gas in solution} in $\frac{\text{mol}}{\text{L}} = 0.40 \frac{\text{mmol}}{\text{L}} \cdot \frac{10^{-3} \text{ mol}}{1 \text{ mmol}} = 4.0 \times 10^{-4} \frac{\text{mol}}{\text{L}}$

$P_{\text{O}_2 \text{ gas above}}$ in kPa = ?

Since air is a mixture of gases, to find the partial pressure, use Dalton's law.

For a single gas in a mixture:

$$P_{\text{gas one}} = (\text{mole fraction gas one})P_{\text{total}} = (\text{volume fraction gas one})P_{\text{total}}$$

P_{total} in kPa = standard atmospheric pressure = 101 kPa

$P_{\text{O}_2} = \text{mol fraction O}_2 \cdot P_{\text{total}} = 0.21 \text{ volume fraction} \cdot 101 \text{ kPa} = \mathbf{21 \text{ kPa}}$

SOLVE the equation for the WANTED symbol in symbols first.

$$k_{\text{henry}} \text{ in mol/kPa} \cdot \text{L} = \frac{\text{Concentration}_{\text{gas in soln}}}{P_{\text{gas above}}} = 4.0 \times 10^{-4} \frac{\text{mol}}{\text{L}} \cdot \frac{1}{21 \text{ kPa}} = 1.9 \times 10^{-5} \frac{\text{mol}}{\text{kPa} \cdot \text{L}}$$

Practice B

- WANT: $P_{\text{above soln.}}$ in kPa

To find the vapor pressure of a gas above a solution with a non-volatile solute, use

Raoult's law: $P_{\text{above soln.}} = \chi_{\text{solvent (volatile) particles}} \cdot P^{\circ}_{\text{solvent}}$

DATA: $P_{\text{soln.}} = ?$ kPa

$$P^{\circ}_{\text{solvent}} = 47.2 \text{ kPa}$$

 $\chi_{\text{solvent particles}} = ? = \text{mol fraction of solvent} = \text{mol solvent} / \text{total moles of particles}$

$$\text{mol of H}_2\text{O solvent} = 2.00 \times 10^3 \text{ g H}_2\text{O} \cdot \frac{1 \text{ mol H}_2\text{O}}{18.0 \text{ g H}_2\text{O}} = \mathbf{111.1 \text{ mol H}_2\text{O}}$$

$$\text{mol of C}_2\text{H}_6\text{O}_2 = 1.00 \times 10^3 \text{ g C}_2\text{H}_6\text{O}_2 \cdot \frac{1 \text{ mol C}_2\text{H}_6\text{O}_2}{62.0 \text{ g C}_2\text{H}_6\text{O}_2} = \mathbf{16.13 \text{ mol C}_2\text{H}_6\text{O}_2}$$

$$\text{Moles total} = \mathbf{111.1 \text{ mol H}_2\text{O} + 16.13 \text{ mol C}_2\text{H}_6\text{O}_2 = 127.2 \text{ mol total}}$$

$$\chi_{\text{solvent}} = ? = \text{mol fraction of solvent} = 111.1 \text{ mol solvent} / 127.2 \text{ moles total} =$$

$$= \mathbf{0.8734 \text{ mol solvent/mol total}} \quad (\text{if you carry an extra } sf)$$

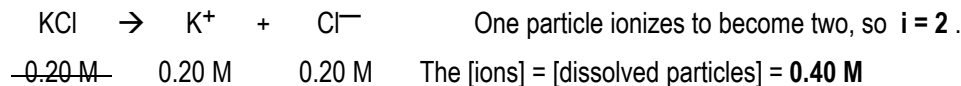
$$\text{SOLVE: } P_{\text{above soln.}} = \chi_{\text{solvent (volatile) particles}} \cdot P^{\circ}_{\text{solvent}}$$

$$= 0.8734 \cdot 47.2 \text{ kPa} = \mathbf{41.2 \text{ kPa} = \text{soln vapor pressure}}$$

Practice C1a. NaBr $i = 2$ b. KNO₃ $i = 2$ c. Na₃PO₄ $i = 4$ d. O₂ $i = 1$ NaBr, KNO₃, and Na₃PO₄ are all soluble and ionize in water (see Lesson 13A).O₂ is covalent and does not ionize. For substances that do not ionize, $i = 1$.

2a. [Particles] in 0.20 M KCl? KCl dissolves and separates ~100% in water.

To find the [ions] in an aqueous solution, write the REC steps.



$$\text{The [ions] = [dissolved particles] = } \mathbf{0.40 \text{ M}}$$

Since one particle becomes two ions, the [ions] is twice the [un-ionized compound].

$$[\text{solute particles}] = (i)[\text{un-ionized solute}] = (2)(0.20 \text{ M KCl}) = \mathbf{0.40 \text{ M solute particles}}$$

2b. $1 \text{ K}_3\text{PO}_4 \rightarrow 3 \text{ K}^+ + 1 \text{ PO}_4^{3-}$ Since one particle becomes four ions, $i = 4$.

$$[\text{solute particles}] = (i)[\text{un-ionized solute}] = (4)(0.12 \text{ M K}_3\text{PO}_4) = \mathbf{0.48 \text{ M solute particles}}$$

3. Since $\text{Na}_2\text{SO}_4 \rightarrow 2 \text{ Na}^+ + 1 \text{ SO}_4^{2-}$ Since one particle becomes three ions, $i = 3$.

$$? = [\text{un-ionized solute}] = [\text{solute particles}] / (i) = 1.00 \text{ M ions} / 3 = \mathbf{0.33 \text{ M Na}_2\text{SO}_4}$$

$$4. \quad ? \text{ mol ions} = 0.25 \text{ L soln} \cdot \frac{0.80 \text{ mol NaCl}}{1 \text{ L soln}} \cdot \frac{2 \text{ mol ions}}{1 \text{ mol NaCl}} = \mathbf{0.40 \text{ mol ions}}$$

5. The solution is 0.10 mol KCl/mole total, and the solution is aqueous. Since mole fraction is a ratio, it works for any size sample. If we have *one* mole of KCl, we must have how many moles of water?

* * * * *

We must have **9 mol water** -- to total 10 moles of solute and solvent for the 0.10 mole fraction. However, KCl ionizes. In that solution with one mol KCl, how many moles of ions are there? How many moles of water? What is the mole fraction that is ions?

* * * * *

One mole of KCl forms *two* total moles of ions in water.

Total moles of mixture = 2.0 mol ions + 9.0 mol solvent = 11.0 mol total

$$\chi_{\text{ions}} = \frac{\text{mol ions}}{\text{Total moles of mixture}} = \frac{2.0 \text{ mol ions}}{11.0 \text{ mol total}} = \mathbf{0.18 \frac{\text{mol ions}}{\text{mol total particles}}}$$

Practice D

1. WANT: i actual

To find the vapor pressure of a gas above a solution with a non-volatile solute, use

Raoult's law: $P_{\text{above soln.}} = \chi_{\text{solvent (volatile) particles}} \cdot P^{\circ}_{\text{solvent}}$

DATA: $P_{\text{above soln.}} = 17.8 \text{ mm Hg}$

$P^{\circ}_{\text{solvent}} = 23.8 \text{ mm Hg}$

$\chi_{\text{solvent particles}} = ? = \text{mol fraction of solvent} = \text{mol solvent} / \text{total moles of particles}$

$$\chi_{\text{solvent particles}} = \frac{P_{\text{above soln.}}}{P^{\circ}_{\text{solvent}}} = \frac{17.8 \text{ mm Hg}}{23.8 \text{ mm Hg}} = \mathbf{0.75}$$

mol of H₂O solvent = **12.0 mol H₂O**

Moles of *un-ionized solute* = 1.0 moles

$\chi_{\text{solvent particles}} = \mathbf{0.75} = 12.0 \text{ mol H}_2\text{O solvent} / Y \text{ moles total}$

$Y \text{ moles total} = 12.0 \text{ mol} / 0.75 = 16 \text{ moles total particles}$

$16 \text{ moles total particles} = 12.0 \text{ moles H}_2\text{O} + (i) 1.0 \text{ mole un-ionized solute}$

i must = **4**.

* * * * *

Lesson 26D: Colligative Properties Of Solutions**Boiling and Melting Points of Solutions**

Recall that a pure substance has a sharp and predictable melting point, and the melting temperature of a substance is the *same* as its freezing temperature.

At a given pressure, a pure substance has a sharp and predictable boiling point. Boiling points are much more sensitive to atmospheric pressure than melting points. A liquid boils when its vapor pressure equals the atmospheric pressure above it.

If a small amount of a soluble but non-volatile (non-gas-forming) substance is dissolved in a pure liquid, the range of temperature in which the liquid stays liquid expands at both ends. Adding soluble but non-volatile particles elevates the boiling point, compared to the boiling point of the pure liquid solvent. Adding soluble particles depresses the

melting/freezing point of the mixture, compared to the freezing point for the pure liquid solvent.

For example:

- If salt, alcohol, or ethylene glycol (sold as anti-freeze) is dissolved in pure water, the water will not freeze until a temperature substantially lower than 0 °C. These chemicals can therefore be used to prevent water from freezing in a car radiator and engine block. Freezing of radiator fluid would not be good because when water freezes, it expands. This expansion would crack open the coils in the expensive radiator and the cooling channels in the more expensive engine block. Ethylene glycol is preferred as an antifreeze because salt water corrodes metals, and most inexpensive alcohols will boil out of an alcohol/water mixture at typical engine operating temperatures.
- Ethylene glycol is relatively non-volatile compared to water: it boils at much higher temperature. As a result, this same anti-freeze that lowers the freezing point of water elevates water's boiling point. This means that in the summer, the water/antifreeze mixture in the engine block stays liquid until a higher temperature than water alone. If the engine gets hotter than normal, the mixture stays liquid until a higher temperature, and a liquid is far more efficient as a coolant than the steam that would form if the fluid were to boil. If your radiator fluid in the summer is water without ethylene glycol added, under the extra heat added due to an air conditioner, the cooling fluid tends to boil away at the engine operating temperature, the engine overheats, and expensive parts break.
- When making home-made ice cream, a container of cream and sugar is surrounded by ice. Salt is then added to the ice around the cream container. The ice/salt/water mixture adjusts to its freezing point, which is lower than for water alone, and is lower than the ice cream solution (water, sugar, and fat) freezing point. NaCl is both inexpensive and can lower the freezing temperature of salt/ice/water as low as -21°C. This very cold mixture freezes the ice cream mixture in the middle.

The *amount* of the elevation of the boiling temperature or the depression of a melting temperature, can be predicted by these equations:

$$\text{Elevation of boiling temp.} = \uparrow\Delta T_{\text{bp solvent}} = K_{\text{bp solvent}} \cdot m_{\text{solute particles}}$$

$$\text{or } \boxed{\text{Elevation BP} = \uparrow\Delta T_{\text{bp solvent}} = K_{\text{bp solvent}} \cdot i \cdot m_{\text{un-ionized solute}}} \quad (8)$$

$$\text{Depression of freezing/melting temp.} = \downarrow\Delta T_{\text{fp}} = K_{\text{fp solvent}} \cdot m_{\text{solute particles}}$$

$$\text{or } \boxed{\text{Depression FP/MP} = \downarrow\Delta T_{\text{fp solvent}} = K_{\text{fp solvent}} \cdot i \cdot m_{\text{un-ionized solute}}} \quad (9)$$

It may help to remember BP elevation and FP depression as “the two *Kim*” equations, in which

- K_{bp} = **molal boiling point constant** and K_{fp} = **molal freezing point constant**
- the constants **K** are values that are characteristic of the *solvent* and can be looked up in tables. The K value depends on the solvent, not the solute.

- The m is determined by the solute. The higher the molality of the solute particles in the solvent, the more the boiling point will elevate and freezing point will depress.
- i = the van't Hoff factor = the number of particles in the solution per *un-ionized* solute particle.
- $m_{\text{solute particles}} = i \cdot m_{\text{un-ionized solute}}$

In colligative properties, it is the number of *separate particles* of solute in the solution that must be measured to predict the value of the property. Each ion is a separate particle. To keep in mind the effect of ionization on colligative properties, it is a good idea to memorize the forms of the colligative property equations that include the symbol i .

Some examples of K values for boiling and melting/freezing are in the table at the right.

Solvent	BP at 1 atm (°C)	K_{bp} (°C·kg/mol)	MP/FP (°C)	K_{fp} (°C·kg/mol)
Water (H ₂ O)	100	0.51	0	1.86
Ethanol (C ₂ H ₅ OH)	78	1.22	−115	1.99
Benzene (C ₆ H ₆)	80.	2.53	6	5.12

Note that the K value for melting/freezing is a higher value. This is true in all cases. For a given solvent, the melting/freezing point will change more, for a given molality of solute, than the boiling point.

In experiments, K values for solvents can easily be measured, and if the FP depression is carefully measured for a known mass of an unknown solute, the molality of the solute can be found. From the moles in the molality, the g/mol (molar mass) can be calculated. The change in the FP is usually measured, rather than the change in BP, because the higher K_{fp} will have a larger change and therefore a lower uncertainty when measured experimentally.

The boiling point elevation equation predicts only the behavior of solutions where relatively non-volatile solutes (usually solids with relatively high boiling points) are added to the solvent, so that the only the solvent vaporizes significantly. In calculations for freezing point depression, the equation can usually be relied upon for predictions if the solute melting point is not close to that of the solvent.

To learn to use those equations, try this example.

- Q.** Using the answers for 24.0 proof wine in Q1 of the first lesson in this module, plus the table above, what will be the freezing temperature of 24.0 proof wine? (Alcohols do not ionize in water, and pure ethanol freezes below −100°C.).

* * * * *

Since the water is the larger component in the solution, treat water as the solvent and the ethanol as the solute. Adding a substance that dissolves in the water will lower the normal freezing point of pure water, which by definition is 0°C. The equation is

$$\text{Depression of FP/MP} = \downarrow \Delta T_{\text{fp solvent}} = K_{\text{fp solvent}} \cdot i \cdot m_{\text{un-ionized solute}}$$

DATA: Molality EtOH = 2.34 *m* EtOH = **2.34 mol/kg** from the first lesson.

$$K_{\text{fp solvent}} = K_{\text{fp water}} = 1.86 \text{ }^{\circ}\text{C} \cdot \text{kg/mol}$$

$i = 1$ because alcohols do not ionize in water.

SOLVE:

$$\downarrow \Delta T_{\text{fp}} = K_{\text{fp solvent}} \cdot i \cdot m_{\text{un-ionized solute}} = (1.86 \frac{\text{ }^{\circ}\text{C} \cdot \text{kg}}{\text{mol}})(1)(2.34 \frac{\text{mol}}{\text{kg}}) = 4.35 \text{ }^{\circ}\text{C}$$

The prediction is that the ethanol will *depress* water's normal freezing point by **4.35 °C**, from 0°C for pure water to **−4.35°C** for the wine. This means that in freezer, typically at −20 °C, wine will freeze, expand, and may crack open its bottle as it does so.

Non-Ideal Behavior

In colligative property calculations, an assumption is “ideal behavior:” that no reaction takes place between the solvent and solute, and that solute particles do not attract each other. In alcohols mixed with water, there is a “hydrogen bond” weak attraction between the —OH group in water and the —OH group present in all alcohols. This hydrogen-bond attraction is one form of “non-ideal” behavior that limits the accuracy of the prediction using the colligative property equations.

In addition, the equations for boiling point elevation and freezing point depression are approximations derived from more complex equations, and those approximations assume both ideal and dilute solutions. As solutions become more concentrated, the predictions of the equations are less accurate.

Practice A: Memorize Equations (8) and (9), then complete the problems below. Use data as needed from the table in the lesson.

- 68.4 grams of an unknown sugar is dissolved in 500. mL of water. At standard pressure, the boiling point of the solution is found to be 100.20 °C. In water, sugars dissolve but do not ionize. Find
 - The molality of the sugar solution.
 - The estimated molar mass of the sugar.
- If 11.7 grams of NaCl is dissolved in 250. mL of water, what will be the freezing point of the solution?

Osmotic Pressure

Osmosis is the movement of solvent through a type of **semi-permeable membrane** that allows solvent molecules to pass through, but not solute particles. For the membrane, solvent can flow in both directions, but the *net* flow is in the direction that tends to *equalize the concentration* of the particles on both sides of the membrane.

If a container contains a membrane that separates the pure solvent from a solution of the solvent, and the container is open to the atmosphere on both sides, the movement of molecules from the pure solvent into the solution by osmosis can cause the solvent level on the solution side of the membrane to be higher than the pure solvent side. At equilibrium,

when no further net change is occurring, the pressure difference due to the higher fluid on the solution side is equal to what is termed the **osmotic pressure** of the solution.

The osmotic pressure of the solution can also be defined as the minimum pressure that is needed to stop the osmosis.

If two solutions with the same solvent but different concentrations are separated by a semi-permeable membrane, the solution with the higher concentration of solute is termed **hypertonic** and the less concentrated is **hypotonic**. If the two solutions have the same osmotic pressure, they are **isotonic**.

The osmotic pressure of a solution can be predicted by the equation

$$\text{Osmotic pressure} = \boxed{\Pi = (i) M_{\text{un-ionized solute}} R T} = M_{\text{particles}} R T \quad (\text{Eq. 10})$$

In this equation:

Π (capital pi) = osmotic pressure in any pressure units

i = the van't Hoff factor = the number of ions that form in the solution per un-ionized solute particle.

(For molecules that do not ionize, $i = 1$; for Na_2SO_4 , $i = 3$, etc.)

$M_{\text{un-ionized solute}}$ = the *molarity* of the un-ionized solute

$$\{ M_{\text{particles}} = (i)(M_{\text{un-ionized solute}}) \}$$

R = the gas constant in *liters* and the pressure unit used in the problem.

T = absolute temperature in kelvins

As a reminder that the number of separate particles in the solute must be measured, it is best to memorize the form of the equation that includes (i). Apply the equation to this example.

Q1. Predict the osmotic pressure in atm. of a 0.200 M K_3PO_4 solution at 25 °C .
($R = 0.0821 \text{ L} \cdot \text{atm}/\text{mol} \cdot \text{K}$).

* * * * *

WANT: Osmotic P = $\Pi = ?$ Write the equation: $\Pi = (i) M_{\text{un-ionized solute}} R T$

As always, use a DATA table to break the problem into pieces. In the table, convert the supplied data to the units needed for the equation and its constants.

DATA: $i = 4$ In water, $1 \text{ K}_3\text{PO}_4 \rightarrow 3 \text{ K}^+ + 1 \text{ PO}_4^{3-}$

$$M_{\text{un-ionized solute}} = 0.200 \text{ M K}_3\text{PO}_4$$

$$R = 0.0821 \text{ L} \cdot \text{atm}/\text{mol} \cdot \text{K}$$

$$T \text{ in K} = 25 \text{ }^\circ\text{C} + 273 = 298 \text{ K}$$

SOLVE:

$$\Pi = (i) M_{\text{un-ionized solute}} R T = (4) \frac{0.200 \text{ mol}}{\text{L}} \cdot 0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \cdot 298 \text{ K} = \mathbf{19.6 \text{ atm}}$$

Measurements of osmotic pressure are often used to find the molar mass of large molecules (polymers and large biomolecules). Knowing the mass of dissolved particles, osmotic pressure, temperature, and solution volume, the mol/L and then the grams/mole can be calculated. Try this example.

- Q2.** 26.0 mg of a large biomolecule is dissolved in water to form 2.50 mL soln. at 10 °C . The osmotic pressure of the solution is found to be 0.750 kPa. Assuming that the biomolecule does not ionize, find its molar mass ($R = 8.31 \text{ kPa} \cdot \text{L/mol} \cdot \text{K}$).

* * * * *

WANT: Molar mass = g/mol We know the g of sample. Can we find its moles?

DATA: The osmotic pressure equation is $\Pi = (i) M_{un-ionized \text{ solute}} R T$

$$\Pi = 0.750 \text{ kPa}$$

$i = 1$ for compounds that do not ionize

$M_{un-ionized \text{ solute}} = ? =$ Molarity of biomolecule since it does not ionize.

$$R = 8.31 \text{ kPa} \cdot \text{L/mol} \cdot \text{K}$$

$$T \text{ in K} = 10 \text{ }^\circ\text{C} + 273 = 283 \text{ K}$$

$$26.0 \text{ mg molecule} = 2.50 \text{ mL soln.} \quad (2 \text{ measures of same solution})$$

$$\text{SOLVE: } M_{\text{solute}} = \frac{\Pi}{i R T} = \frac{0.750 \text{ kPa}}{(1) 283 \text{ K}} \cdot \frac{\text{mol} \cdot \text{K}}{8.31 \text{ kPa} \cdot \text{L}} = 3.19 \times 10^{-4} \frac{\text{mol}}{\text{L}}$$

Can the molar mass be solved from here?

* * * * *

$$\text{Molar Mass} = ? \frac{\text{g biomolecule}}{\text{mol}} = \frac{26.0 \text{ mg}}{2.50 \text{ mL}} \cdot \frac{1 \text{ L}}{3.19 \times 10^{-4} \text{ mol}} = 3.26 \times 10^4 \frac{\text{g}}{\text{mol}}$$

Recall that prefixes are abbreviations of exponentials. Exponentials and prefixes can be grouped separately from numbers and units, and like prefixes can cancel.

The answer is a large molar mass, but that is consistent with a large biomolecule.

Osmotic Pressure and Kinetic Molecular Theory

The equation for osmotic pressure is essentially the same as the ideal gas law.

$$PV = nRT \text{ for gases; } P(\text{osmotic}) = (\text{moles of particles}/V \text{ in liters})RT = (i)MRT$$

This is because the ideal gas law and osmotic pressure equations can both be derived from the Kinetic Molecular Theory equations: the physics that governs the motion of particles at a constant temperature with elastic collisions. In a gas, the particles move relatively far between collisions, and in a solution, the distance between collisions is relatively short, but the math describing the collisions is essentially the same.

The KMT equations (Module 19) can be used to predict many properties of gases, liquids, and solids. However, as with ideal gases, the molecules of real substances do not always collide without significant interaction, and this is one factor that can cause variations between predicted (ideal) and real molecular behavior.

Practice B: Memorize the numbered equations in this module, then complete the problems below.

- List the 5 colligative properties covered in this module. After each, write the equation that predicts the value for the property.
- Which colligative properties are calculated based on
 - Molality?
 - Molarity?
- To mix physiological saline solution, 9.07 g of NaCl is dissolved in water to make one liter of solution. Predict the osmotic pressure in kPa of this solution at the mammalian normal temperature of 37 °C ($R = 8.31 \text{ kPa} \cdot \text{L/mol} \cdot \text{K}$).

For More Practice

- Find these topics in your regular chemistry textbook.
- For each topic, find the examples embedded in the textbook reading on the topic.
- Cover the example *answer* with a sticky note or cover sheet.
- Work the problem in your problem notebook, then check your answer.

Problems can be solved in many ways, and the text may solve problems differently, but the answer should be the same as the answer found using the methods here. If your answer differs, the answer worked out in the text will help you to find your error.

For more practice, work “end of the textbook chapter” problems that have answers you can check.

ANSWERS

Practice A

1a. WANT: ? $m = ?$ mol sugar/kg solvent

DATA: For BP elevation, the equation is “kim:” $\uparrow \Delta T_{\text{bp solvent}} = K_{\text{bp solvent}} \cdot i \cdot m_{\text{un-ionized solute}}$

$\uparrow \Delta T_{\text{bp solvent}} = +0.20 \text{ }^\circ\text{C}$, since pure water at std. P boils at exactly 100°C by definition.

$K_{\text{bp solvent}} = K_{\text{bp water}} = 0.51 \text{ }^\circ\text{C kg/mol}$ from table above.

$i = 1$ since sugar does not ionize

$m_{\text{un-ionized solute}} = ?$

SOLVE: $m_{\text{un-ionized solute}} = ? = \frac{\Delta T_{\text{bp solvent}}}{(i) K_{\text{bp solvent}}} = \frac{0.20 \text{ }^\circ\text{C}}{(1) 0.51 \text{ }^\circ\text{C kg}} = 0.39 \text{ mol sugar/kg water}$

1b. WANT: ? MM sugar = ? g sugar/mol sugar

DATA: g sugar = 68.4 **g sugar**

Since you know the grams of sugar in the sample, you only need the moles.

Molality solution = 0.39 **mol sugar/kg water**

Kg water = 500. mL water = 500. g water = 0.500 kg water

Since $m = \text{mol solute/kg solvent}$

$$\text{Mol sugar} = (\text{molality sugar})(\text{kg H}_2\text{O}) = (0.39 \frac{\text{mol sugar}}{\text{kg H}_2\text{O}})(0.500 \text{ kg H}_2\text{O})$$

$$= 0.195 \text{ mol sugar} \quad (\text{carrying an extra } sf \text{ until the end})$$

$$\text{SOLVE: } ? \text{ MM sugar} = ? \frac{\text{g sugar}}{\text{mol sugar}} = \frac{68.4 \text{ g sugar}}{0.195 \text{ mol sugar}} = \mathbf{350 \frac{\text{g sugar}}{\text{mol sugar}}}$$

2. WANT: Final freezing point = ? Find $\downarrow\Delta T_{\text{fp solvent}} = ?$

DATA: For freezing point depression, $\downarrow\Delta T_{\text{fp solvent}} = K_{\text{fp solvent}} \cdot i \cdot m_{\text{un-ionized solute}}$

In the DATA table, compute a value for each symbol.

$K_{\text{fp solvent}} = K_{\text{fp water}} = 1.86 \text{ }^\circ\text{C} \cdot \text{kg/mol}$ from the table above.

$i = 2$ for NaCl. In water, $1 \text{ NaCl} \rightarrow 1 \text{ Na}^+ + 1 \text{ Cl}^-$ One particle ionizes to form two.

$$\text{Molality NaCl un-ionized} = \frac{\text{mol NaCl}}{\text{kg of solvent H}_2\text{O}} = ?$$

$$\text{Mol NaCl} = 11.7 \text{ g NaCl} \cdot \frac{1 \text{ mol NaCl}}{58.5 \text{ g NaCl}} = \mathbf{0.200 \text{ mol NaCl}}$$

kg of solvent H₂O = 250 mL H₂O = 250 g H₂O = **0.250 kg H₂O solvent**

$$\text{Molality NaCl un-ionized solute} = \frac{\text{mol NaCl}}{\text{kg H}_2\text{O}} = \frac{0.200 \text{ mol NaCl}}{0.250 \text{ kg H}_2\text{O}} = \mathbf{0.800 \text{ m NaCl}}$$

SOLVE:

$$\downarrow\Delta T_{\text{fp solvent}} = K_{\text{fp solvent}} \cdot i \cdot m_{\text{un-ionized solute}} = (1.86 \text{ }^\circ\text{C} \cdot \frac{\text{kg}}{\text{mol}})(2)(0.800 \frac{\text{mol}}{\text{kg}}) = \downarrow\mathbf{2.98 \text{ }^\circ\text{C}}$$

Since pure water freezes at **0 °C**, and adding a solute lowers the solvent freezing point, the freezing temperature of this salt water mixture = **−2.98 °C**

Practice B

1. See summary at the end of this module.

2a. Henry's law relating gas solubility in a solvent to pressure holds within experimental error for *any* measure of concentration, including both molarity and molality.

Molality (m) is used to calculate freezing point depression and melting point elevation in a solution, compared to the properties of the solvent if pure.

2b. Molarity is used to predict osmotic pressure. Molarity can be used with Henry's law.

3. WANT: osmotic pressure = Π = ? The equation is $\Pi = (i) M_{un-ionized\ solute} RT$
- DATA: 9.07 g NaCl = 1 L solution
 58.5 g NaCl = 1 mol NaCl (grams prompt)
- Make the DATA table to match the terms in the equation:
- $i = 2$ for NaCl In water, 1 NaCl \rightarrow 1 Na⁺ + 1 Cl⁻
- $M_{un-ionized\ solute} = ?$
- $$\text{Molarity NaCl} = M_{\text{NaCl}} = \frac{\text{mol NaCl}}{\text{L soln}} = \frac{9.07\text{ g NaCl}}{1\text{ L}} \cdot \frac{1\text{ mol NaCl}}{58.5\text{ g NaCl}} = 0.155\text{ M NaCl}$$
- $R = 8.31\text{ kPa} \cdot \text{L/mol} \cdot \text{K}$
- $T\text{ in K} = ^\circ\text{C} + 273 = 37^\circ\text{C} + 273 = 310.\text{ K}$
- SOLVE: $\Pi = (i) M_{un-ionized\ solute} RT = (2) 0.155 \frac{\text{mol}}{\text{L}} \cdot 8.31 \frac{\text{L} \cdot \text{kPa}}{\text{mol} \cdot \text{K}} \cdot 310.\text{ K} = 799\text{ kPa}$

* * * * *

SUMMARY: Properties of Mixtures

You may want to flashcard these equations.

1. Five measures of the composition of a mixture of two substances are:

- Molarity (M)** = $\frac{\text{moles of substance}}{\text{Liter of mixture}}$
- Mass fraction** = $\frac{\text{mass of a substance}}{\text{Total mass of mixture}}$ in any consistent mass units
- Mole fraction (χ)** = $\frac{\text{moles of a substance}}{\text{Total moles in mixture}}$

For fractions that measure a part of a total, the value of the fraction must be a number between zero and one.

- Molality (m)** = $\frac{\text{moles of substance}}{\text{kilogram of solvent}}$

In molality calculations, for the density of *dilute* aqueous solutions, use

$$1.00\text{ g H}_2\text{O} = 1\text{ mL H}_2\text{O} \quad \text{and/or} \quad 1.00\text{ kg H}_2\text{O} = 1\text{ L H}_2\text{O}$$

- Parts per million** = $\text{ppm} \equiv \frac{\text{g substance}}{\text{g total}} \times 10^6$

2. **Percent and PPM**

- If *percent* or *ppm* is WANTED, solve for the part/total fraction, then convert to percent or ppm.
- In the special case where percent or ppm is supplied as DATA, either
 - start by converting to the fraction using the definition equations, then use the fraction in conversions, or

- write the percent or ppm in the DATA as a conversion in this format.
If “X % by mass” is supplied, write as DATA: X g substance = 100 g total
If “X ppm” is supplied, write as DATA: X g substance = 10⁶ g total
For dilute aq. soln, if ppm is supplied, write: # ppm = # mg substance/L

3. **van't Hoff factor (i)** = the number of separate particles that a substance forms when it dissolves in water.

Examples: Predict **i = 3** for K₂SO₄; **i = 2** for NaCl; **i = 1** for non-electrolytes.
Moles of solute particles dissolved = (**i**) (moles un-ionized solute)

4. For colligative properties, the rule is

Each ion is a separate particle when counting the moles, molarity, or molality of solute particles in the solution.

5. Five Colligative Properties

All colligative properties assume that in a mixture, the particles will not significantly react. If the particles *ionize* but do not otherwise react in a solvent, colligative properties can be calculated based on the van't Hoff factor for the number of *ions* in the solution.

- a. Henry's law: $\text{Concentration}_{\text{gas in solution}} = k_{\text{henry}} \cdot P_{\text{gas above}}$

where the Henry's law constant = k_{henry} is constant at a given temperature for a given gas and solvent combination. Concentration may be measured in any consistent units.

- b. Partial pressure of solvent vapor above a solution with a non-volatile solute:

Raoult's law: $P_{\text{above soln.}} = \chi_{\text{solvent (volatile) particles}} \cdot P^{\circ}_{\text{solvent}}$

where $P^{\circ}_{\text{solvent}}$ = the vapor pressure of the *pure* solvent at that temperature.

- c. Elevation of boiling pt. = $\uparrow \Delta T_{\text{bp solvent}} = K_{\text{bp solvent}} \cdot i \cdot m_{\text{un-ionized solute}}$

- d. Depression of FP/MP = $\downarrow \Delta T_{\text{fp}} = K_{\text{fp solvent}} \cdot i \cdot m_{\text{un-ionized solute}}$

where the K values for elevation and depression are a characteristic property of the solvent. This means that for each solvent, the elevation and depression will vary depending on the molality of the solute particles dissolved in the solution, but not what those particles are.

BP elevation and FP depression are the “two kim ” equations.

- e. Osmotic pressure = $\Pi = (i) M_{\text{un-ionized solute}} R T = M_{\text{all dissolved particles}} R T$

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