

Concentration Units

How many ways can you express how concentrated a solution is, and how do you convert between them?

●●○ Intermediate

Solutions

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Before you start: it helps to understand [Molarity and Solution Stoichiometry](#) and [The Mole Concept](#).

What is it?

A **concentration** tells you how much solute is in a given amount of solution or solvent. Different jobs use different units:

Unit	Definition	Typical use
Molarity , c (mol/L)	$\frac{\text{mol solute}}{\text{L solution}}$	Reactions and titrations in the lab
Mass percent , % (m/m)	$\frac{\text{mass solute}}{\text{mass solution}} \times 100\%$	Labels, concentrated reagents, alloys
Parts per million , ppm	$\frac{\text{mass solute}}{\text{mass solution}} \times 10^6$	Traces: pollutants, fluoride in water
Mole fraction , x	$\frac{\text{mol of component}}{\text{total mol}}$	Gas mixtures, vapour pressure
Molality , m (mol/kg)	$\frac{\text{mol solute}}{\text{kg solvent}}$	Colligative properties

💡 KEY IDEA

Watch the **denominator**. Molarity divides by the volume of **solution**; molality divides by the mass of **solvent**; mass percent and ppm divide by the mass of **solution**. Converting between them is a matter of finding the missing amount, often with the **density** of the solution.

Why does it matter?

- **Safety and health.** Drinking-water limits are given in ppm or ppb (for example, fluoride is added at about 0.7 ppm); a 3 % hydrogen peroxide solution is safe on skin, but 30 % is corrosive.
- **Temperature independence.** Molarity changes slightly as a solution expands when warmed; molality, mass percent and mole fraction do not, because they use masses.

- **Reading labels.** Concentrated acids are sold as mass percent with a density; to use them in a reaction you need their molarity.

How does it work?

1. Mass percent and ppm

Both compare the mass of solute with the mass of the **whole solution** (solute + solvent). For very small amounts, ppm (and ppb, $\times 10^9$) avoid tiny percentages. In dilute water solutions, 1 L has a mass of about 1 kg, so **1 ppm $\approx$ 1 mg/L**.

2. Mole fraction

Convert every component to moles and divide each by the total. Mole fractions have no unit and always add up to 1.

3. Molality

Divide the moles of solute by the mass of **solvent in kilograms**. Because it doesn't depend on volume, molality is used for boiling-point and freezing-point calculations.

4. Converting with density

To go from a “per volume” unit to a “per mass” unit, take a convenient amount of solution (for example 1 L) and use the density to find its mass. Then split that mass into solute and solvent.

THINK OF IT LIKE THIS

Concentration units are like describing a fruit salad. You can say “one in five pieces is a strawberry” (mole fraction), “20 % of the weight is strawberry” (mass percent) or “three strawberries per bowl” (molarity). All describe the same salad; you just need to know how big the pieces and the bowl are to convert.

MORE PRECISELY

Mass percent is sometimes written % (w/w); volume percent, % (v/v), is used for liquid mixtures such as alcoholic drinks, and % (w/v) (grams per 100 mL) in medicine. For very dilute aqueous solutions, molarity and molality are almost equal, because 1 L of solution contains almost exactly 1 kg of water. The approximation 1 ppm $\approx$ 1 mg/L also holds only when the density is close to 1.00 g/mL.

Visualise it

What each unit divides by

Teal: solute (top) · amber: what it is divided by (bottom)

Molarity	$\frac{\text{mol solute}}{\text{L of solution}}$
Molality	$\frac{\text{mol solute}}{\text{kg of solvent only}}$
Mass % ($\times 100\%$)	$\frac{\text{mass solute}}{\text{mass of solution}}$
ppm ($\times 10^6$)	$\frac{\text{mass solute}}{\text{mass of solution}}$
Mole fraction	$\frac{\text{mol of one component}}{\text{total mol (all parts)}}$

The units differ in what they divide by: solution or solvent, volume or mass, moles or grams.

Worked example

WORKED EXAMPLE: Mass percent

Question: 5.00 g of NaCl is dissolved in 95.0 g of water. What is the mass percent of NaCl?

$$\begin{aligned}\% \text{ NaCl} &= \frac{5.00 \text{ g}}{5.00 \text{ g} + 95.0 \text{ g}} \times 100\% \\ &= 5.00\%\end{aligned}$$

 WORKED EXAMPLE: Parts per million

Question: A 2.50 kg sample of water contains 3.5 mg of lead. Express the lead concentration in ppm.

Step 1 Convert to the same unit: 3.5 mg = 0.0035 g and 2.50 kg = 2500 g.

Step 2 Concentration:

$$\begin{aligned}\text{ppm} &= \frac{0.0035 \text{ g}}{2500 \text{ g}} \times 10^6 \\ &= 1.4 \text{ ppm}\end{aligned}$$

 WORKED EXAMPLE: Mole fraction

Question: 23.0 g of ethanol is mixed with 54.0 g of water. Find the mole fraction of ethanol.

Step 1 Moles: ethanol $\frac{23.0 \text{ g}}{46.07 \text{ g/mol}} = 0.4992 \text{ mol}$; water $\frac{54.0 \text{ g}}{18.02 \text{ g/mol}} = 2.997 \text{ mol}$.

Step 2 Mole fraction:

$$\begin{aligned}x_{\text{ethanol}} &= \frac{0.4992 \text{ mol}}{0.4992 \text{ mol} + 2.997 \text{ mol}} \\ &= 0.143\end{aligned}$$

 WORKED EXAMPLE: From molarity to molality

Question: A 2.00 mol/L NaCl solution has a density of 1.08 g/mL. What is its molality?

Step 1 Take 1.000 L of solution. Its mass is $1000 \text{ mL} \times 1.08 \text{ g/mL} = 1080 \text{ g}$.

Step 2 Mass of NaCl: $2.00 \text{ mol} \times 58.44 \text{ g/mol} = 116.9 \text{ g}$.

Step 3 Mass of water: $1080 \text{ g} - 116.9 \text{ g} = 963.1 \text{ g} = 0.9631 \text{ kg}$.

Step 4 Molality:

$$\begin{aligned}m &= \frac{2.00 \text{ mol}}{0.9631 \text{ kg}} \\ &= 2.08 \text{ mol/kg}\end{aligned}$$

 WORKED EXAMPLE: From mass percent to molarity

Question: Concentrated hydrochloric acid is 37.0 % HCl by mass, with a density of 1.19 g/mL. What is its molarity?

Step 1 Take 1.000 L: mass of solution = 1000 mL $\times$ 1.19 g/mL = 1190 g.

Step 2 Mass of HCl = 0.370 $\times$ 1190 g = 440.3 g.

Step 3 Moles and molarity:

$$\begin{aligned}n &= \frac{440.3 \text{ g}}{36.46 \text{ g/mol}} \\&= 12.08 \text{ mol} \\c &= \frac{12.08 \text{ mol}}{1.000 \text{ L}} \\&= 12.1 \text{ mol/L}\end{aligned}$$

Common mistake

 COMMON MISTAKE: Dividing by the solvent instead of the solution

Mass percent uses the mass of the **solution**: 5.00 g of NaCl in 95.0 g of water is 5.00 %, not $\frac{5.00}{95.0} \times 100\% = 5.26\%$.

 COMMON MISTAKE: Confusing molarity and molality

Molarity (c or M) is per **litre of solution**; molality (m) is per **kilogram of solvent**. They are close only for dilute water solutions.

 COMMON MISTAKE: Mixing units in ppm

Solute and solution must be in the **same** mass unit before multiplying by 10^6 : 3.5 mg in 2.50 kg means 0.0035 g in 2500 g.

 NOTATION NOTE

- Molarity: mol/L, mol L⁻¹ or M. Molality: mol/kg or *m* (italic *m*, not to be confused with *m* for mass or metre).
- 1 ppm = 1 mg/kg; 1 ppb = 1 µg/kg.

Remember this

REMEMBER THIS

- Molarity = $\text{mol solute} \div \text{L solution}$. Molality = $\text{mol solute} \div \text{kg solvent}$.
- Mass percent = $\text{mass solute} \div \text{mass solution} \times 100 \%$; ppm uses $\times 10^6$.
- Mole fraction = $\text{mol of component} \div \text{total mol}$; fractions add to 1.
- Convert between volume-based and mass-based units with the density, starting from 1 L of solution.

References

1. Brown, T. L.; LeMay, H. E., Jr.; Bursten, B. E.; Murphy, C. J.; Woodward, P. M.; Stoltzfus, M. W. Chemistry: The Central Science, 15th ed.; Pearson, 2022.

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