

**The Islamic University of Gaza
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Civil Engineering Department**

**Environmental Engineering
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Environmental Measurements

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Mass Concentration Units

Chemical concentration is one of the most important determinants in almost all aspects of chemical fate, transport, and treatment in both natural and engineered systems

Important prefixes to know include pico (10^{-12} , abbreviated as p), nano (10^{-9} , abbreviated as n), micro (10^{-6} , abbreviated as μ), milli (10^{-3} , abbreviated as m), and kilo (10^{+3} , abbreviated as k). Other important units are the tonne (which is also called the metric ton by some in the United States), which equals 1,000 kg (or 2,204 lb), and the common ton, which equals 2,000 lb. In addition, 1 teragram (Tg) = 10^{12} g = 1 million metric tons.

Mass Concentration Units

Table / 2.1

Common Units of Concentration Used in Environmental Measurements

Representation	Example	Typical Units
Mass chemical/total mass	mg/kg in soil	mg/kg, ppm _m
Mass chemical/total volume	mg/L in water or air	mg/L, µg/m ³
Volume chemical/total volume	volume fraction in air	ppm _v
Moles chemical/total volume	moles/L in water	M

MASS/MASS UNITS

$$\text{ppm}_m = \text{g of } i \text{ in } 10^6 \text{ g total}$$

example/ 2.1 Concentration in Soil

A 1 kg sample of soil is analyzed for the chemical solvent trichloroethylene (TCE). The analysis indicates that the sample contains 5.0 mg of TCE. What is the TCE concentration in ppm_m and ppb_m ?

solution

$$\begin{aligned} [\text{TCE}] &= \frac{5.0 \text{ mg TCE}}{1.0 \text{ kg soil}} = \frac{0.005 \text{ g TCE}}{10^3 \text{ g soil}} \\ &= \frac{5 \times 10^{-6} \text{ g TCE}}{\text{g soil}} \times 10^6 = 5 \text{ ppm}_m = 5,000 \text{ ppb}_m \end{aligned}$$

Note that in soil and sediments, mg/kg equals ppm_m , and $\mu\text{g/kg}$ equals ppb_m .

MASS/VOLUME UNITS: mg/L AND $\mu\text{g}/\text{m}^3$

In the atmosphere, it is common to use concentration units of mass per volume of air, such as mg/m^3 and $\mu\text{g}/\text{m}^3$. In water, mass/volume concentration units of mg/L and $\mu\text{g}/\text{L}$ are common. In most aqueous systems, ppm_m is equivalent to mg/L . This is because the density of pure water is approximately $1,000 \text{ g}/\text{L}$

MASS/VOLUME UNITS: mg/L AND $\mu\text{g}/\text{m}^3$

example/ 2.2 Concentration in Water

One liter of water is analyzed and found to contain 5.0 mg of TCE. What is the TCE concentration in mg/L and ppm_m ?

solution

$$[\text{TCE}] = \frac{5.0 \text{ mg TCE}}{1.0 \text{ L H}_2\text{O}} = \frac{5.0 \text{ mg}}{\text{L}}$$

To convert to ppm_m , a mass/mass unit, it is necessary to convert the volume of water to mass of water. To do this, divide by the density of water, which is approximately 1,000 g/L:

$$\begin{aligned} \text{TCE} &= \frac{5.0 \text{ mg TCE}}{1.0 \text{ L H}_2\text{O}} \times \frac{1.0 \text{ L H}_2\text{O}}{1,000 \text{ g H}_2\text{O}} \\ &= \frac{5.0 \text{ mg TCE}}{1,000 \text{ g total}} = \frac{5.0 \times 10^{-6} \text{ g TCE}}{\text{g total}} \times \frac{10^6 \text{ ppm}_\text{m}}{\text{mass fraction}} \\ &= 5.0 \text{ ppm}_\text{m} \end{aligned}$$

In most dilute aqueous systems, mg/L is equivalent to ppm_m .

In this example, the TCE concentration is well above the allowable U.S. drinking water standard for TCE, 5 $\mu\text{g}/\text{L}$ (or 5 ppb), which was set to protect human health. Five ppb is a small value. Think of it this way: Earth's population exceeds 6 billion people, meaning that 30 individuals in one of your classes constitute a human concentration of approximately 5 ppb!

Volume/Volume and Mole/Mole Units

$$\text{ppm}_v = \frac{V_i}{V_{\text{total}}} \times 10^6$$

example/ 2.3 Concentration in Air

What is the carbon monoxide (CO) concentration expressed in $\mu\text{g}/\text{m}^3$ of a 10 L gas mixture that contains 10^{-6} mole of CO?

solution

In this case, the measured quantities are presented in units of moles of the chemical per total volume. To convert to mass of the chemical per total volume, convert the moles of chemical to mass of chemical by multiplying moles by CO's molecular weight. The molecular weight of CO (28 g/mole) is equal to 12 (atomic weight of C) plus 16 (atomic weight of O).

$$\begin{aligned} [\text{CO}] &= \frac{1.0 \times 10^{-6} \text{ mole CO}}{10 \text{ L total}} \times \frac{28 \text{ g CO}}{\text{mole CO}} \\ &= \frac{28 \times 10^{-6} \text{ g CO}}{10 \text{ L total}} \times \frac{10^6 \mu\text{g}}{\text{g}} \times \frac{10^3 \text{ L}}{\text{m}^3} = \frac{2,800 \mu\text{g}}{\text{m}^3} \end{aligned}$$

Partial-Pressure Units

In the atmosphere, concentrations of chemicals in the gas and particulate phases may be determined separately. A substance will exist in the gas phase if the atmospheric temperature is above the substance's boiling (or sublimation) point or if its concentration is below the saturated vapor pressure of the chemical at a specified temperature

Partial-Pressure Units

Table / 2.3

Composition of the Atmosphere

Compound	Concentration (% volume or moles)	Concentration (ppm _v)
Nitrogen (N ₂)	78.1	781,000
Oxygen (O ₂)	20.9	209,000
Argon (Ar)	0.93	9,300
Carbon dioxide (CO ₂)	0.039	391
Neon (Ne)	0.0018	18
Helium (He)	0.0005	5
Methane (CH ₄)	0.00018	1.813
Krypton (Kr)	0.00011	1.1
Hydrogen (H ₂)	0.00005	0.50
Nitrous oxide (N ₂ O)	0.000032	0.324
Ozone (O ₃)	0.000002	0.020

Partial-Pressure Units

Remember, the ideal gas law states that, at a given temperature and volume, pressure is directly proportional to the number of moles of gas present; therefore, pressure fractions are identical to mole fractions (and volume fractions). For this reason, partial pressure can be calculated as the product of the mole or volume fraction and the total pressure. For example:

$$\begin{aligned} P_i &= [\text{volume fraction}_i \text{ or mole fraction}_i \times P_{\text{total}}] \\ &= [(\text{ppm}_v)_i \times 10^{-6} \times P_{\text{total}}] \end{aligned}$$

$$\text{ppm}_v = \frac{P_i}{P_{\text{total}}} \times 10^6$$

Mole/Volume Units

Units of moles per liter (molarity, M) are used to report concentrations of compounds dissolved in water. Molarity is defined as the number of moles of compound per liter of solution. Concentrations expressed in these units are read as molar.

Mole/Volume Units

example/ 2.6 Concentration as Partial Pressure

The concentration of gas-phase PCBs in the air above Lake Superior was measured to be 450 picograms per cubic meter (pg/m³). What is the partial pressure (in atm) of PCBs? Assume the temperature is 0°C, the atmospheric pressure is 1 atm, and the average molecular weight of PCBs is 325.

solution

The partial pressure is defined as the mole or volume fraction times the total gas pressure. First, find the number of moles of PCBs in a liter of air. Then use the ideal gas law (Equation 2.7) to calculate that 1 mole of gas at 0°C and 1 atm occupies 22.4 L. Substitute this value into the first expression to determine the mole fraction of PCBs:

$$450 \frac{\text{pg}}{\text{m}^3 \text{ air}} \times \frac{\text{mole}}{325 \text{ g}} \times 10^{-12} \frac{\text{g}}{\text{pg}} \times 10^{-3} \frac{\text{m}^3}{\text{L}} = 1.38 \times 10^{-15} \frac{\text{mole PCB}}{\text{L air}}$$

$$1.38 \times 10^{-15} \frac{\text{mole PCB}}{\text{L air}} \times \frac{22.4 \text{ L}}{\text{mole air}} = 3.1 \times 10^{-14} \frac{\text{mole PCB}}{\text{mole air}}$$

Multiplying the mole fraction by the total pressure (1 atm) (see Equation 2.10) yields the PCB partial pressure of 3.1×10^{-14} atm.

Mole/Volume Units

example/ 2.7 Concentration as Partial Pressure Corrected for Moisture

What would be the partial pressure (in atm) of carbon dioxide (CO₂) when the barometer reads 29.0 inches of Hg, the relative humidity is 80 percent, and the temperature is 70°F? Use Table 2.3 to obtain the concentration of CO₂ in dry air.

solution

The partial-pressure concentration units in Table 2.3 are for dry air, so the partial pressure must first be corrected for the moisture present in the air. In dry air, the CO₂ concentration is 391 ppm_v. The partial pressure will be this volume fraction times the total pressure of dry air. The total pressure of dry air is the total atmospheric pressure (29.0 in. Hg) minus the contribution of water vapor. The vapor pressure of water at 70°F is 0.36 lb/in². Thus, the total pressure of dry air is

$$\begin{aligned} P_{\text{total}} - P_{\text{water}} &= 29.0 \text{ in. Hg} - \left[0.36 \frac{\text{lb}}{\text{in.}^2} \times \frac{29.9 \text{ in. Hg}}{14.7 \text{ lb/in.}^2} \times 0.8 \right] \\ &= 28.4 \text{ in. Hg} \end{aligned}$$

The partial pressure of CO₂ would be:

$$\text{vol. fraction} \times P_{\text{total}} = 391 \text{ ppm}_v \times \frac{10^{-6} \text{ vol. fraction}}{\text{ppm}_v} \times \left[28.4 \text{ in. Hg} \times \frac{1 \text{ atm}}{29.9 \text{ in. Hg}} \right] = 3.7 \times 10^{-4} \text{ atm}$$

Mole/Volume Units

example/ 2.8 Concentration as Molarity

The concentration of TCE is 5 ppm. Convert this to units of molarity. The molecular weight of TCE is 131.5 g/mole.

solution

Remember, in water, ppm_m is equivalent to mg/L, so the concentration of TCE is 5.0 mg/L. Conversion to molarity units requires only the molecular weight:

$$\begin{aligned}\frac{5.0 \text{ mg TCE}}{\text{L}} \times \frac{1 \text{ g}}{10^3 \text{ mg}} \times \frac{1 \text{ mole}}{131.5 \text{ g}} &= \frac{3.8 \times 10^{-5} \text{ moles}}{\text{L}} \\ &= 3.8 \times 10^{-5} \text{ M}\end{aligned}$$

Often, concentrations below 1 M are expressed in units of millimoles per liter, or millimolar (1 mM = 10⁻³ moles/L), or in micromoles per liter, or micromolar (1 μM = 10⁻⁶ moles/L). Thus, the concentration of TCE could be expressed as 0.038 mM or 38 μM.

Mole/Volume Units

example/ 2.9 Concentration as Molarity

The concentration of alachlor, a common herbicide, in the Mississippi River was found to range from 0.04 to 0.1 $\mu\text{g/L}$. What is the concentration range in nmoles/L? The molecular formula for alachlor is $\text{C}_{14}\text{H}_{20}\text{O}_2\text{NCl}$, and its molecular weight is 270.

solution

The lowest concentration range in nmoles/L can be found as follows:

$$\frac{0.04 \mu\text{g}}{\text{L}} \times \frac{\text{mole}}{270 \text{ g}} \times \frac{10^{-6} \text{ g}}{\mu\text{g}} \times \frac{10^9 \text{ nmole}}{\text{mole}} = \frac{0.15 \text{ nmole}}{\text{L}}$$

Similarly, the upper limit (0.1 $\mu\text{g/L}$) can be calculated as 0.37 nmoles/L.

Other Types of Units

NORMALITY

Normality (equivalents/L) typically is used in defining the chemistry of water, especially in instances where acid–base and oxidation–reduction reactions are taking place. Normality is also used frequently in the laboratory during the analytical measurement of water constituents.

NORMALITY

example/ 2.10 Calculations of Equivalent Weight

What are the equivalent weights of HCl, H₂SO₄, NaOH, CaCO₃, and aqueous CO₂?

solution

To find the equivalent weight of each compound, divide the molecular weight by the number of equivalents:

$$\text{eqv wt of HCl} = \frac{(1 + 35.5)\text{g/mole}}{1 \text{ eqv/mole}} = \frac{36.5 \text{ g}}{\text{eqv}}$$

$$\text{eqv wt of H}_2\text{SO}_4 = \frac{(2 \times 1) + 32 + (4 \times 16)\text{g/mole}}{2 \text{ eqv/mole}} = \frac{49 \text{ g}}{\text{eqv}}$$

$$\text{eqv wt of NaOH} = \frac{(23 + 16 + 1)\text{g/mole}}{1 \text{ eqv/mole}} = \frac{40 \text{ g}}{\text{eqv}}$$

$$\text{eqv wt of CaCO}_3 = \frac{40 + 12 + (3 \times 16)\text{g/mole}}{2 \text{ eqv/mole}} = \frac{50 \text{ g}}{\text{eqv}}$$

Determining the equivalent weight of aqueous CO₂ requires additional information. Aqueous carbon dioxide is not an acid until it hydrates in water and forms carbonic acid (CO₂ + H₂O → H₂CO₃). So aqueous CO₂ really has 2 eqv/mole. Thus, one can see that the equivalent weight of aqueous carbon dioxide is

$$\frac{12 + (2 \times 16)\text{g/mole}}{2 \text{ eqv/mole}} = \frac{22 \text{ g}}{\text{eqv}}$$

NORMALITY

example/ 2.11 Calculation of Normality

What is the normality (N) of 1 M solutions of HCl and H₂SO₄?

solution

$$1 \text{ M HCl} = \frac{1 \text{ mole HCl}}{\text{L}} \times \frac{1 \text{ eqv}}{\text{mole}} = \frac{1 \text{ eqv}}{\text{L}} = 1 \text{ N}$$

$$1 \text{ MH}_2\text{SO}_4 = \frac{1 \text{ mole H}_2\text{SO}_4}{\text{L}} \times \frac{2 \text{ eqv}}{\text{mole}} = \frac{2 \text{ eqv}}{\text{L}} = 2 \text{ N}$$

Note that on an equivalent basis, a 1 M solution of sulfuric acid is twice as strong as a 1 M solution of HCl.

NORMALITY

example/ 2.12 Use of Equivalents in Determining the Accuracy of a Water Analysis

Prof. Mihelcic was in the city of Dunedin in New Zealand to view yellow-eyed penguins and albatrosses. The label on a bottle of New Zealand mineral water he purchased there states that a chemical analysis of the mineral water resulted in the following cations and anions being identified with corresponding concentrations (in mg/L):

$$[\text{Ca}^{2+}] = 2.9 \quad [\text{Mg}^{2+}] = 2.0 \quad [\text{Na}^{+}] = 11.5 \quad [\text{K}^{+}] = 3.3$$

$$[\text{SO}_4^{2-}] = 4.7 \quad [\text{F}^{-}] = 0.09 \quad [\text{Cl}^{-}] = 7.7$$

Is the analysis correct?

solution

First, convert all concentrations of major ions to an equivalent basis. To do this, multiply the concentration in mg/L by a unit conversion (g/1,000 mg) and then divide by the equivalent weight of each substance (g/eqv). Then sum the concentrations of all cations and anions on an equivalent basis. A solution with less than 5 percent error generally is considered acceptable.

NORMALITY

Cations

$$[\text{Ca}^{2+}] = \frac{1.45 \times 10^{-4} \text{ eqv}}{\text{L}}$$

$$[\text{Mg}^{2+}] = \frac{1.67 \times 10^{-4} \text{ eqv}}{\text{L}}$$

$$[\text{Na}^{+}] = \frac{5 \times 10^{-4} \text{ eqv}}{\text{L}}$$

$$[\text{K}^{+}] = \frac{8.5 \times 10^{-5} \text{ eqv}}{\text{L}}$$

Anions

$$[\text{SO}_4^{2-}] = \frac{9.75 \times 10^{-5} \text{ eqv}}{\text{L}}$$

$$[\text{F}^{-}] = \frac{4.73 \times 10^{-6} \text{ eqv}}{\text{L}}$$

$$[\text{Cl}^{-}] = \frac{2.17 \times 10^{-4} \text{ eqv}}{\text{L}}$$

The total amount of cations equals $9.87 \times 10^{-4} \text{ eqv/L}$, and the total amount of anions equals $3.2 \times 10^{-4} \text{ eqv/L}$.

The analysis is not within 5 percent. The analysis resulted in more than three times more cations than anions on an equivalent basis. Therefore, either of two conclusions is possible: (1) One or more of the reported concentrations are incorrect, assuming all major cations and anions are accounted for. (2) One or more important anions were not accounted for by the chemical analysis. (Bicarbonate, HCO_3^{-} , would be a good guess for the missing anion, as it is a common anion in most natural waters.)