

CONCENTRATION

Lecture 7

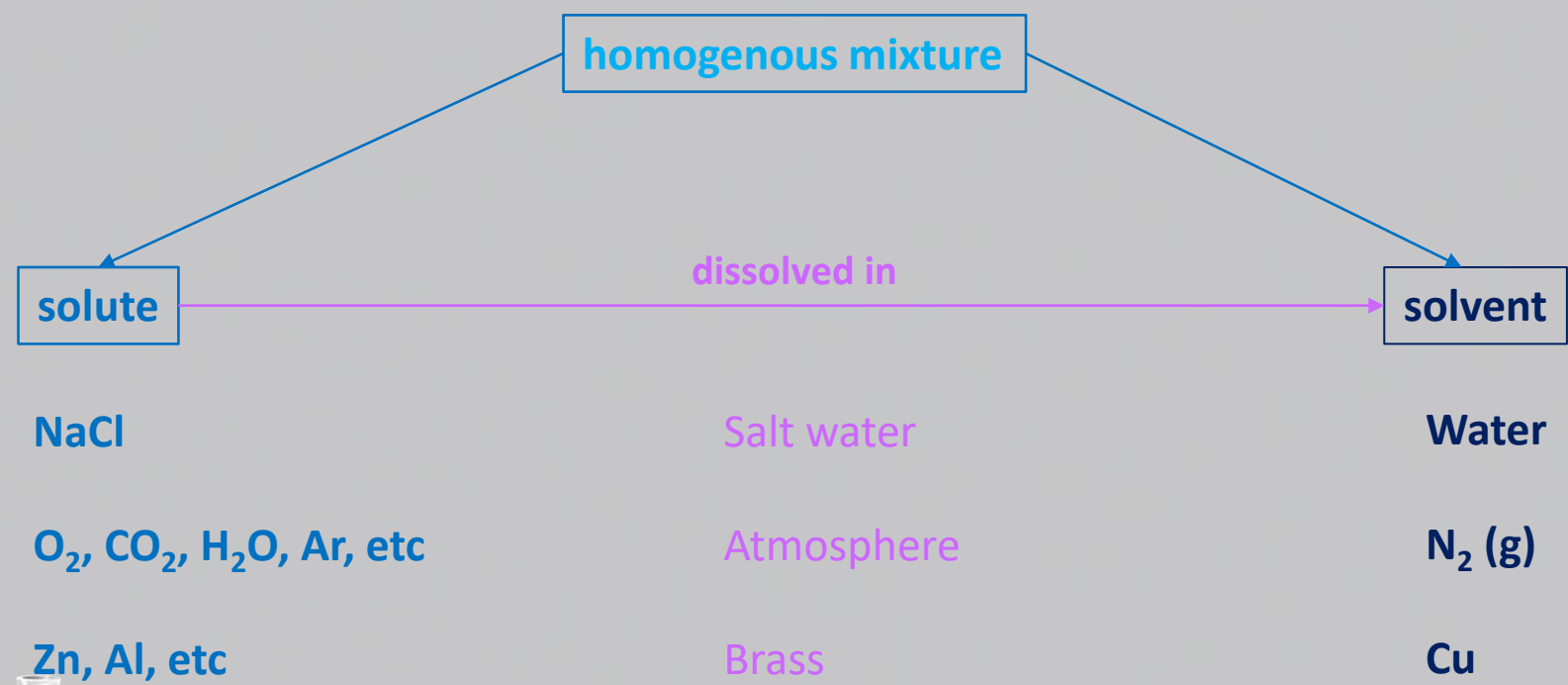


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SOLUTIONS

Solution - a homogenous mixture of two or more substances in relative amounts that can be varied continuously up to what is called the **limit of solubility**. The term solution is commonly applied to the liquid state of matter, but solutions of gases and solids are possible.



CONCENTRATIONS

The **concentration** of a solution is a measure of how 'crowded' the solute particles are. The more concentrated the solution, the more particles it contains in a given **volume**.

Percent of solute in solution

$$\frac{V}{V} = \frac{\text{volume of solute}}{\text{volume of solution}} \times 100\% \quad \text{n\% by volume}$$

Molarity

$$M = \frac{\text{moles of solute}}{\text{liters of solutions}} \quad \left(\frac{\text{mol}}{L}\right)$$



Relative amounts
Solute and solvent
Solute and solution

EXAMPLE

EXAMPLE 1: What is the molarity of a solution in which 68.28 g of CaBr_2 are dissolved in 0.91 litres of the solution?

$$M = \frac{n}{V}$$

$$g = n \times MM$$

$$n = \frac{68.28 \text{ g}}{200 \text{ g/mol}} = \underline{\underline{0.34 \text{ mol}}}$$

$$M = \frac{0.34}{0.91} = 0.375 \frac{\text{mol}}{\text{L}}$$

$$\frac{\text{mol}}{\text{L}} \approx M$$

$$\begin{aligned} MM &= 40.0 + \\ &2 \times 80.0 = \\ &= 200 \text{ g/mol} \end{aligned}$$



EXAMPLE

EXAMPLE 2: How many grams of glucose ($C_6H_{12}O_6$, MM = 180.0 g/mol) are contained in 0.600L of a 0.200M aqueous solution of glucose?



$$M = \frac{n}{V} \rightarrow n = M \cdot V$$

$\frac{\text{mol}}{L} \cdot L = \text{mol}$

$$n = 0.200 \frac{\text{mol}}{L} \cdot 0.600 L = 0.12 \text{ mol glucose}$$



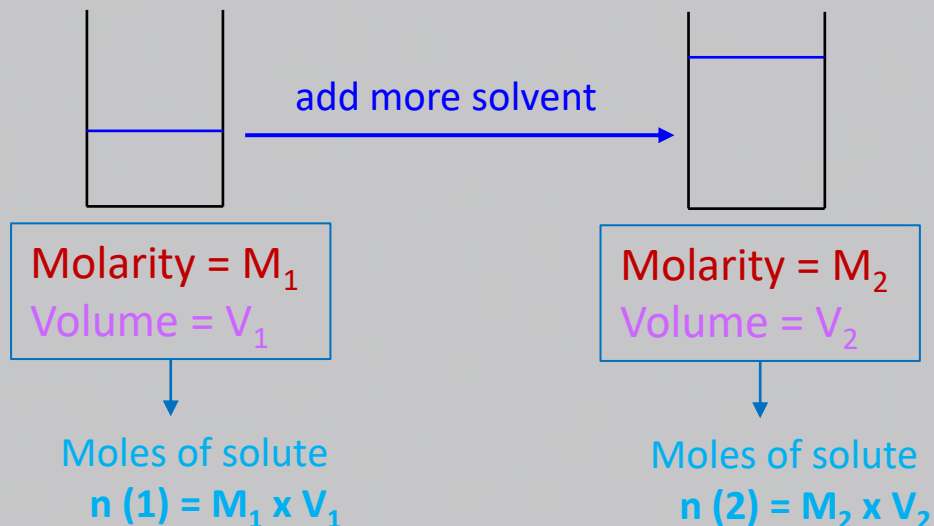
$$g = n \cdot MM = 0.12 \text{ mol} \cdot 180.0 \frac{g}{\text{mol}} = 21.6 \text{ g glucose}$$



DILUTION

process of decreasing the concentration of a solute in a solution, by mixing with more solvent without the addition of more solute

$$M = \frac{\text{moles of solute}}{\text{liters of solutions}} = \frac{n}{V}$$



SAME NUMBER OF MOLES OF SOLUTE !!!

$$n(1) = n(2)$$

$$M_1 \times V_1 = M_2 \times V_2$$



EXAMPLES

M_1

$$M = \frac{n}{V}$$

EXAMPLE 1: You have a bottle of 12.0 M HCl. How much of this will you have to use to make 1.25 L of 0.25 M HCl?

$$\underbrace{V_2} \quad \underbrace{M_2} \quad n_{HCl(1)} = n_{HCl(2)}$$

$$\underline{M_1} \underline{(V_1)} = \underline{M_2} \underline{V_2}$$

$$V_1 = \frac{M_2 \cdot V_2}{M_1} = \frac{0.25 \frac{\text{mol}}{\text{L}} \cdot 1.25 \text{ L}}{12.0 \frac{\text{mol}}{\text{L}}}$$

$$= 0.026 \text{ L}$$



EXAMPLES

solution 1

EXAMPLE 2: Calculate the volume of 2.050 M copper (II) nitrate that must be diluted with water to prepare 750.0 mL of 0.8553 M solution.

$$M = \frac{n}{V}$$

solution 2

$$n_1 = n_2$$

$$M_1 V_1 = M_2 V_2$$

$$V_1 = \frac{M_2 V_2}{M_1}$$

$$V_2 = 750.0 \text{ mL} = 0.75 \text{ L}$$

$$V_1 = \frac{0.8553 \frac{\text{mol}}{\text{L}} \times 0.75 \text{ L}}{2.050 \frac{\text{mol}}{\text{L}}}$$

$$V_1 = 0.313 \text{ L} = 313 \text{ mL}$$



CONCENTRATION

Relative amounts

Mass percent
(w/w)

$$\frac{\text{mass of solute}}{\text{mass of solution}} \times 100$$

$$\frac{\text{mass of solute}}{\text{mass of solute} + \text{mass of solvent}} \times 100$$

$$\frac{\text{mass of solvent}}{\text{mass of solution}} \times 100$$

$$\frac{\text{mass of solvent}}{\text{mass of solute} + \text{mass of solvent}} \times 100$$

law of conservation of mass

Volume percent
(v/v)

$$\frac{\text{volume of solute}}{\text{volume of solution}} \times 100$$

~~$$\frac{\text{volume of solute}}{\text{volume of solute} + \text{volume of solvent}} \times 100$$~~

but NO law of conservation of volume

percent is 'parts per hundred'

ppm is part per million

ppb is parts per billion

ppt is part per trillion

Common ppm: mg solute per kg solution
 mL solute per L solution
 mg solute per L solution



CONCENTRATION

Mole Fraction
(χ)

not % !!!

$$\frac{\text{moles of solute}}{\text{total moles}}$$

$$\frac{\text{moles of solute}}{\text{moles of solute} + \text{moles of solvent}}$$

$$\frac{\text{moles of solvent}}{\text{total moles}}$$

$$\frac{\text{moles of solvent}}{\text{moles of solute} + \text{moles of solvent}}$$

Molarity
(M)

$$\frac{\text{moles of solute}}{\text{L of solution}}$$

mol/L

Molality
(m)

$$\frac{\text{moles of solute}}{\text{KG of SOLVENT}}$$

mol/kg

TEMPERATURE INDEPENDENT!
mass does NOT depend on temp



EXAMPLE

EXAMPLE 1: 4.6 g ethanol, $\text{C}_2\text{H}_5\text{OH}$, is added to 90.0 g of water to make a solution that has a density of 0.95 g/mL. Calculate:

- a) mass percent of ethanol
- b) ppm (ethanol) by mass
- c) mole fraction of water
- d) molality]
- e) molarity

a) mass percent of ethanol

$$\frac{\text{mass of solute}}{\text{mass of solution}} \times 100 = \frac{4.6 \text{ g}}{4.6 \text{ g} + 90.0 \text{ g}} \times 100 = 4.9\%$$

b) ppm (ethanol) by mass

$$\frac{\text{mass of solute}}{\text{mass of solution}} \times 1\,000\,000 = \frac{4.6 \text{ g}}{4.6 \text{ g} + 90.0 \text{ g}} \times 1\,000\,000 = 48.626 \text{ ppm (49,000 ppm)}$$



EXAMPLE

c) mole fraction of water

$$\frac{\text{moles of solvent}}{\text{total moles}} = \frac{\text{moles of water}}{\text{moles of ethanol} + \text{moles of water}} = \frac{5.0 \text{ mol}}{0.1 \text{ mol} + 5.0 \text{ mol}} = 0.98$$

$$\text{moles water} = 90.0 \text{ g} \div 18 \text{ g/mol} = 5.0 \text{ mol}$$

$$\text{moles ethanol (C}_2\text{H}_5\text{OH)} = 4.6 \text{ g} \div 46 \text{ g/mol} = 0.1 \text{ mol}$$

d) molality

$$\frac{\text{moles of solvent}}{\text{KG of SOLVENT}} = \frac{0.1 \text{ mol}}{0.090 \text{ kg}} = 1.11 \text{ m}$$

e) molarity

$$\frac{\text{moles of solute}}{\text{L of solution}} = \frac{0.10 \text{ mol}}{0.0996 \text{ L}} = 1.00 \text{ M}$$

$$\text{density} = 0.95 \text{ g/mL} = \text{mass} / \text{volume}$$

$$0.95 \text{ g/mL} = (94.6 \text{ g} / \text{vol})$$

$$\text{Vol} = 94.6 \text{ mL} / 0.95 \text{ g/mL} = 99.6 \text{ mL}$$



EXAMPLE 2: A solution of potassium chloride that is 8.2% by mass has a density of 1.08 g/mL. Calculate the molarity of the solution

$$M = \frac{\text{moles of solute}}{L \text{ of solution}}$$

consider 100 g of solution !

$$\begin{aligned} &8.2 \text{ g KCl} \\ &/74.5 \text{ g/mol} \\ &= 0.11 \text{ mol KCl} \end{aligned}$$

$$\begin{aligned} 1.08 \text{ g/mL} &= 100 \text{ g / vol} \\ \text{Vol} &= 100 / 1.08 \\ &= 92.6 \text{ mL} \\ &= 0.0926 \text{ L solution} \end{aligned}$$

$$M = \frac{0.11 \text{ mol KCl}}{0.0926 \text{ L solution}} = 1.19 \text{ M}$$



LIMITING REACTANT

used up when the chemical reaction comes to completion

Example: Consider the reaction: $2 \text{K (s)} + 1 \text{H}_2\text{CO}_3 \text{ (aq)} \rightarrow 1 \text{K}_2\text{CO}_3 \text{ (aq)} + 1 \text{H}_2 \text{ (g)}$.
How many grams of H_2 will form if 7.0 g of K is taken for the reaction?

Step 1: Calculate number of moles of K

Step 2: Use the mole ratio to calculate corresponding moles of H_2

Step 3: Convert the moles of H_2 into grams of H_2

① $g = n \times \text{MM} \rightarrow n = g / \text{MM}$
 $\text{moles K} = 7.0 \text{ g} / 39.0 \text{ g/mol} = 0.18 \text{ mol K}$

② $\frac{\text{mol H}_2}{\text{mol K}} = \frac{1}{2} \rightarrow \text{mol H}_2 = \frac{0.18 \text{ mol K}}{2} = 0.09 \text{ mol H}_2$

③ $g = n \times \text{MM} \rightarrow g = 0.09 \text{ mol H}_2 \times 2.0 \text{ g/mol} =$
 $= \boxed{0.18 \text{ g H}_2}$



LIMITING REACTANT

Example: Consider the reaction: $2 \text{K (s)} + 1 \text{H}_2\text{CO}_3 \text{ (aq)} \rightarrow 1 \text{K}_2\text{CO}_3 \text{ (aq)} + 1 \text{H}_2 \text{ (g)}$.
How many grams of H_2 will form if 3.35 g of H_2CO_3 is taken for the reaction?

Step 1: Calculate number of moles of H_2CO_3 — $\text{MM} = 62.0 \text{ g/mol}$

Step 2: Use the mole ratio to calculate corresponding moles of H_2

Step 3: Convert the moles of H_2 into grams of H_2 — $\text{MM} = 2.0 \text{ g/mol}$

$$\textcircled{1} \quad n = g/\text{MM} = \frac{3.5 \text{ g}}{62.0 \text{ g/mol}} = 0.056 \text{ mol H}_2\text{CO}_3$$

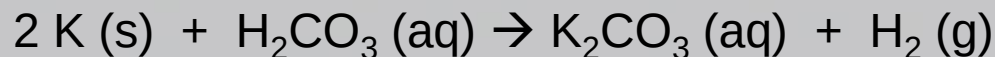
$$\textcircled{2} \quad \frac{\text{mol H}_2}{\text{mol H}_2\text{CO}_3} = \frac{1}{1} \Rightarrow \text{mol H}_2 = \text{mol H}_2\text{CO}_3$$

$$\text{mol H}_2 = 0.056 \text{ mol H}_2$$

$$\textcircled{3} \quad g = n \times \text{MM} \Rightarrow g = 0.056 \text{ mol} \times 2.0 \text{ g/mol} = 0.11 \text{ g H}_2$$



LIMITING REACTANT



7.0 g of K makes 0.18..... grams of H_2

3.35 g H_2CO_3 makes.....0.11.....grams of H_2

How many grams of H_2 will form if 7.0 g of K AND 3.35 g H_2CO_3 are both taken for the reaction?

0.11 g H_2

because H_2CO_3 is the limiting reactant



PERCENT YIELD

The **theoretical yield** is the maximum possible **mass** of a **product** that can be made in a chemical reaction.

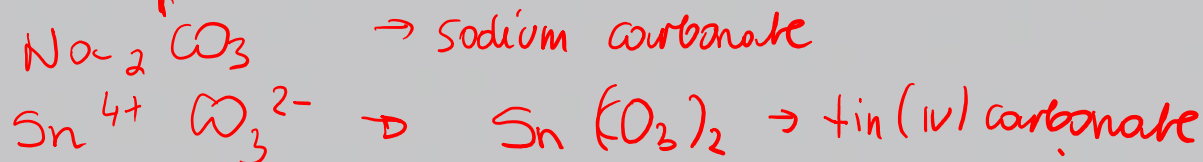
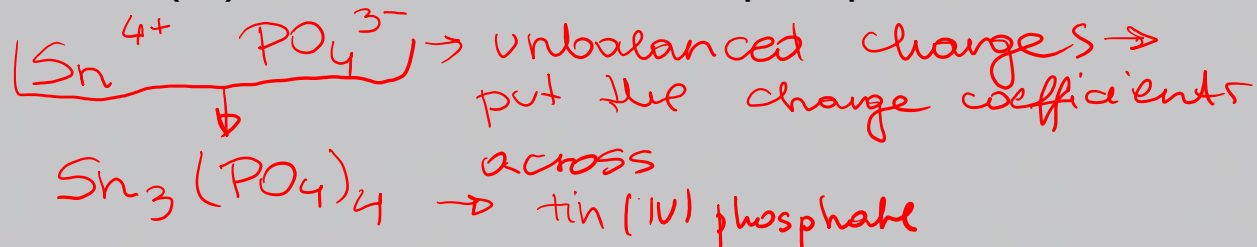
An **actual yield** is the mass of a product actually obtained from the reaction. It is usually less than the theoretical yield.

$$\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$



PERCENT YIELD

EXAMPLE: 1) Write a balanced equation for the reaction of tin (IV) phosphate with sodium carbonate to make tin (IV) carbonate and sodium phosphate.



PERCENT YIELD



2) If 36 grams of tin (IV) phosphate is mixed with an excess of sodium carbonate, how many grams of tin (IV) carbonate will form?

Molar masses:

$$\text{Sn}_3(\text{PO}_4)_4 = 3 \times 119 + 4 \times (31.0 + 4 \times 16.0) = 737 \text{ g/mol}$$

$$\text{Sn}(\text{CO}_3)_2 = 119 + 2(12.0 + 3 \times 16.0) = 239 \text{ g/mol}$$

$$a) \text{ moles } \text{Sn}_3(\text{PO}_4)_4 = g/\text{MM} = \frac{36 \text{ g}}{737 \text{ g/mol}} = 0.048 \text{ mol } \text{Sn}_3(\text{PO}_4)_4$$

$$\frac{\text{mol } \text{Sn}(\text{CO}_3)_2}{\text{mol } \text{Sn}_3(\text{PO}_4)_4} = \frac{3}{1} \rightarrow \text{moles } \text{Sn}(\text{CO}_3)_2 = 3 \times \text{mol } \text{Sn}_3(\text{PO}_4)_4$$

$$b) g = n \times \text{MM} \rightarrow 0.15 \text{ mol} \times 239 \text{ g/mol} =$$
$$= \boxed{35.9 \text{ g } \text{Sn}(\text{CO}_3)_2}$$

theoretical yield calculated from the
stoichiometry of reaction



PERCENT YIELD

↗ actual yield

3) If 29.8 grams of tin (IV) carbonate are actually formed when this reaction goes to completion, what is the percent yield?

$$\begin{aligned}\% \text{ yield} &= \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \\ &= \frac{29.8\text{g}}{35.0\text{g}} \cdot 100\% = \boxed{85.1\%}\end{aligned}$$

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PERCENT YIELD

$$MM_{Na_2CO_3} = 2 \times 23.0 + (12.0 + 3 \times 16.0) = 106 \text{ g/mol}$$
$$MM_{Na_3PO_4} = 3 \times 23.0 + (31 + 4 \times 16.0) = 164 \text{ g/mol}$$

4) If 7.3 grams of sodium carbonate are used in the reaction and this results in a 74.0% yield, how many grams of sodium phosphate will be formed?

1. Theoretical Yield (TY) — actual yield (AY)

a) moles of $Na_2CO_3 \rightarrow$ moles of $Na_3PO_4 \rightarrow$ mass Na_3PO_4

$$\frac{7.3 \text{ g } Na_2CO_3}{106 \text{ g/mol}} = 0.069 \text{ mol } Na_2CO_3$$

$$b) \frac{\text{mol } Na_3PO_4}{\text{mol } Na_2CO_3} = \frac{4}{6} \rightarrow \text{mol } Na_3PO_4 = \frac{4 \times 0.069}{6} = 0.046 \text{ mol } Na_3PO_4$$

$$c) g = n \times MM \rightarrow g_{Na_3PO_4} = 0.046 \text{ mol} \cdot 164 \text{ g/mol} = 7.53 \text{ g } Na_3PO_4 \rightarrow \text{TY}$$

$$2. \% \text{ yield} = \frac{AY}{TY} \cdot 100\% \rightarrow AY = \frac{\% \text{ yield} \cdot TY}{100\%}$$

$$AY = \frac{74.0 \cdot 7.53}{100} = 5.57 \text{ g}$$



ATOM ECONOMY

The **atom economy** of a reaction is a measure of the amount of starting materials that end up as useful products.
It is important for sustainable development.

$$\text{atom economy} = \frac{\text{mass of atoms in desired product}}{\text{mass total of reactant atoms}} \times 100\%$$

$$\text{atom economy} = \frac{\text{molar mass of desired product} \times \text{coefficient of desired product}}{\text{combined molar masses of reactants}} \times 100\%$$



Na₂SO₄ desired product

$$\text{atom economy} = \frac{(2 \times 23) + 32 + (4 \times 16)}{(2 \times 23) + (2 \times 1) + 32 + 94 \times 16)} \times 100\% = \frac{142}{144} \times 100\% = 98.6\%$$



ATOM ECONOMY



Na_2SO_4 desired product

$$\text{atom economy} = \frac{(2 \times 23) + 32 + (4 \times 16)}{(2 \times 23) + (2 \times 1) + 32 + 94 \times 16)} \times 100\% = \frac{142}{144} \times 100\%$$

= 98.6% GOOD

H_2 desired product

$$\text{atom economy} = \frac{(2 \times 1)}{(2 \times 23) + (2 \times 1) + 32 + 94 \times 16)} \times 100\% = \frac{2}{144} \times 100\%$$

= 1.4% BAD

