Solution Concentrations

% Concentration by Mass

grams of solute per 100 grams of solution

$$\text{Mass \%} = \frac{\text{grams solute}}{\text{100 grams solution}} \times 100$$

$$\text{Mass \%} = \frac{\text{grams solute}}{\text{grams solution}} \times 100$$

$$\text{Mass \%} = \frac{\text{grams solute}}{(\text{g solute} + \text{g solvent})} \times 100$$

NOTICE:
g solution = g solute + g solvent

Example

5.5% (by mass) dextrose:
5.5 g dextrose dissolved in 100 g solution
5.5 g dextrose dissolved in 94.5 g water

May also be expressed as 5.5 % (m:m)

Calculate the mass % of a solution made by dissolving 3.8 g $CaBr_2$ in 58.0 g H_2O.

$$\text{Mass \%} = \frac{\text{grams solute}}{(\text{g solute} + \text{g solvent})} \times 100$$

$$\text{Mass \%} = \frac{3.8\ \text{g}}{(3.8\ \text{g} + 58.0\ \text{g})} \times 100$$

Mass % = 3.8 g / 61.8g x 100 = 6.14887 ➔ 6.1

How many grams of sucrose are contained in 235 grams of a 4.82% (by mass) aqueous sucrose solution?

$$\text{Mass \%} = \frac{\text{grams solute}}{\text{grams solution}} \times 100$$

$$\frac{(\text{Mass \%})(\text{grams solution})}{100} = \text{grams solute}$$

(4.82) (235 g) / 100 = 11.327 ➔ 11.3 g

How much of a 13.5% (by mass) NaCl solution is needed to obtain 47.0 grams NaCl?

$$\text{Mass \%} = \frac{\text{grams solute}}{\text{grams solution}} \times 100$$

$$\text{grams solution} = \frac{(100)(\text{grams solute})}{(\text{Mass \%})}$$

$$\text{grams solution} = \frac{(100)(47.0\text{ g})}{(13.5)}$$

grams solution = 348.148 g ➔ 348 g

Two Forms of the % by Mass Problem

Solute = NaBr
Solvent = Water
Solution = NaBr + Water

Determine % by Mass for a solution:

1. Prepared by dissolving 22.4 g of NaBr in 287 g of water:
% NaBr (by mass) = 22.4 g / (22.4 g + 287 g) x 100 = 7.24 %

2. Prepared by dissolving 22.4 g of NaBr in water to make 287 g of solution:
% NaBr (by mass) = 22.4 g / (287 g) x 100 = 7.80 %

Find the % concentration of a solution prepared by dissolving 2.20 g $BaCl_2$ in 57.9 g of water.

$$\text{Mass \%} = \frac{\text{grams solute}}{(\text{g solute} + \text{g solvent})} \times 100$$

$$= \frac{2.20\text{ g}}{(2.20\text{ g} + 57.9\text{ g})} \times 100$$

$$= 3.66$$

How many grams of sodium sulfate are in 505 g of a 15.0% solution? How many grams of water?

$$\textbf{Mass \%} = \frac{\textbf{grams solute}}{\textbf{grams solution}} \times \textbf{100}$$

$$\frac{(\text{grams solution})(\text{Mass \%})}{100} = \text{grams solute}$$

$$\text{grams } Na_2SO_4 = \frac{(505\text{ g})(15.0)}{100}$$

grams = 75.75 ➔ 75.8 g

grams H_20 = 505 g – 75.8 g

grams H_20 = 429.2 g

Weight (mass): Volume

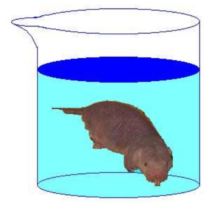
Weighing solvents often cumbersome
So, another form of practical measurement
Weigh solute
Dissolve in solvent
Bring (accurately) to desired volume
Express as % (w:v … weight:volume)

Example

5.5 g of solute brought to 100 ml solution
5.5 % (w:v)

Molarity

Primary means of calculating solution concentrations



A 1.0 molar solution

1 molar solution = molar mass dissolved in 1 L of solution

$$\mathbf{M} = \frac{\text{moles solute}}{\text{liters solution}}$$

Preparing Molar Solutions

Weigh solute
Dissolve in small amount of solvent
Bring (accurately) to desired volume using a volumetric flask

Molar Solutions defined as

$$\frac{\text{Moles of Solute}}{\text{Volume of Solution}}$$

A 1.50 M aqueous solution of HCl contains 1.50 moles of HCl dissolved in enough water to make 1.00 liter of solution. How many *grams* of HCl would be in 1.0 liter of this solution?

Given: 1.50 moles HCl
Wanted: g HCl

Grams requested, need molar mass for HCl (36.46)

$$1.50 \text{ moles HCl} \times \frac{36.46 \text{ g}}{1 \text{ mole}} = 54.69 \text{ g} \rightarrow 54.7 \text{ g}$$

Calculate the molarity of a solution prepared by dissolving 23.9 grams of KBr in 400.0 mL (0.4000 L) of solution.

Given: 23.9 g KBr in 400 mL
Wanted: Molarity (M/L)

Grams requested, need molar mass for KBr (119.01)

$$23.9 \text{ g} \times \frac{1 \text{ mole}}{119.01 \text{ g}} \times \frac{1}{0.4000 \text{ L}} = 0.502059 \text{ M} \rightarrow 0.502 \text{ M}$$

How many grams of KBr must be added to water to prepare 250.0 mL of a 0.188 M KBr solution

Given: 0.188 M/L KBr
Wanted: g KBr

Grams requested, need molar mass for KBr (119.01)

$\frac{0.188 \text{ Moles}}{1 \text{ L}} \times 250.0 \text{ ml} \times \frac{1 \text{ L}}{1000 \text{ ml}} \times \frac{119.01 \text{ g}}{1 \text{ Mole}} = 5.59347 \text{ g} \rightarrow 5.59 \text{ g}$

Check:

$5.59 \text{ g} \times \frac{1 \text{ mole}}{119.01 \text{ g}} \times \frac{1}{0.250 \text{ L}} = 0.187883 \text{ moles/L} \rightarrow 0.188 \text{ M}$

How many mL of a 0.475 M KBr solution can be prepared from 9.51 g KBr?

Given: 9.51 g KBr
Wanted: mL of 0.475 M/L solution

Grams requested, need molar mass for KBr (119.01)

$9.51 \text{ g} \times \frac{1 \text{ mole}}{119.01 \text{ g}} \times \frac{1 \text{ L}}{0.475 \text{ M}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 168.230 \rightarrow 168 \text{ mL}$

How many moles of sodium sulfate (Na_2SO_4) are present in 250 mL of a 0.150 M solution of sodium sulfate?

Given: 250 mL; 0.150 M solution
Wanted: moles sodium sulfate

All calculations in moles; no need for molar mass

$\frac{0.150 \text{ M}}{1 \text{ L}} \times 0.250 \text{ L} = 0.0375 \text{ moles}$

How would you prepare 2.50 L of a 0.360 M solution of sulfuric acid (H2SO4) starting with 18.0 M sulfuric acid

Given: Dilution of 18.0 M H2SO4
Needed: 2.50 L of 0.360 M solution

Hint: moles in final solution ➔same as moles added

$0.360 \text{ M} \times 2.50 \text{ L} = 18.0 \text{ M} \times X \text{ Liters}$

$X = \frac{0.360 \text{ M} \times 2.50 \text{ L}}{18.0 \text{ M}}$

$X = 0.0500 \text{ L} \rightarrow 50.00 \text{ mL}$

So, Dilute 5.00 mL 18 M H2SO4 to 2.50 L of solution

KI is the additive in "iodized" table salt. Calculate the molarity of a solution prepared by dissolving 2.41 g of KI in water and diluting to 50.0 mL.

Given: 2.41 g KI (molar mass = 166.01)
Wanted: molarity of 50.0 mL solution

Determine Moles:

$2.41\text{ g} \times \frac{1 \text{ mole}}{166.01\text{ g}} = 1.452 \times 10^{-2}$ moles

Determine Molarity (moles/L)

$\frac{1.452 \times 10^{-2}\text{ moles}}{50.0\text{ mL}} \times \frac{1000\text{ mL}}{1\text{ L}} = 0.2903 \rightarrow 0.290\text{ M}$

Potassium hydroxide is used in making liquid soap, as well as many other things. How many grams would you use to prepare 2.50 L of 1.40 M KOH?

Given: 2.50 L of 1.40 M KOH
Wanted: grams KOH (molar mass = 56.11)

Moles present in solution

$2.50\text{ L} \times \frac{1.40\text{ M}}{1\text{ L}} = 3.5$ moles

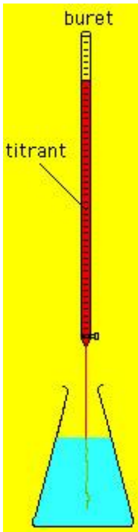
Gram equivalent

$3.5\text{ moles} \times \frac{56.11\text{ g}}{1\text{ mole}} = 196.4\text{ g} \rightarrow 196\text{ g}$

Solution Stoichiometry

How many mL of a 0.155 M $CaCl_2$ solution are required to react with Na_2SO_4 to form 15.8 g $CaSO_4$? **$Na_2SO_4 + CaCl_2 \rightarrow 2\,NaCl + CaSO_4(s)$**

Given: 15.8 g $CaSO_4$
Wanted: mL 0.155 M $CaCl_2$ solution



Grams requested, need molar mass for $CaSO_4$ (136.14)

Start with # moles of given (known) substance:

$15.8\text{ g } CaSO_4 \times \frac{1 \text{ mole}}{136.14\text{ g}} = 0.1161\text{ moles} \rightarrow 0.116\text{ moles } CaSO_4$

Use per expression from reaction coefficients → moles wanted

$0.116\text{ moles } CaSO_4 \times \frac{1\text{ mole } CaCl_2}{1\text{ mole } CaSO_4} = 0.116\text{ moles } CaCl_2$

Convert moles wanted to equivalent solution concentration

$0.116 \text{ moles } CaCl_2 \times \frac{1 \quad L}{0.155 \text{ M}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 748 \text{ mL } CaCl_2$

How many mL of a 0.155 M $CaCl_2$ solution will react with
47.7 mL of a 0.248 M Na_2SO_4 solution? $Na_2SO_4 + CaCl_2 \rightarrow 2\,NaCl + CaSO_4(s)$

Given: 47 mL of 0.248 Na_2SO_4
Wanted: mL 0.155 M $CaCl_2$ solution

$\frac{0.248 \text{ moles } Na_2SO_4}{1000 \text{ mL}} \times 47.7 \text{ mL} \times \frac{1 \text{ mole } CaCl_2}{1 \text{ mole } Na_2SO_4} \times \frac{1000 \quad \text{mL}}{0.155 \text{ mole}} = 76.3 \text{ mL}$

How many grams of AgCl can be precipitated by adding excess NaCl to 65.0 mL of 0.757 M $AgNO_3$? $AgNO_{3(aq)} + NaCl_{(aq)} \rightarrow AgCl_{(s)} + NaNO_{3(aq)}$

$\frac{0.757 \text{ moles } AgNO_3}{1000 \text{ mL}} \times 65.0 \text{ mL} \times \frac{1 \text{ mole AgCl}}{1 \text{ mole } AgNO_3} \times \frac{143.2 \text{ g}}{\text{mole}} = 7.05 \text{ g AgCl}$

How many mL of 0.084 M $AgNO_3$ solution would be needed to react with excess NaCl solution to produce 0.64 g of solid AgCl

What mass of barium fluoride can be precipitated from 25.0 mL of 0.465 M NaF by adding excess barium nitrate solution?

$\frac{0.465 \text{ moles NaF}}{L} \times \frac{1 \quad L}{1000 \text{ mL}} \times 25.0 \text{ mL} \times \frac{1 \text{ mole } BaF_2}{2 \text{ moles NaF}} \times \frac{175.34 \text{ g}}{1 \text{ mole}} = 1.02 \text{ g}$

Assignment

Continue taking Unit 9 Practice Test
The Practice Quiz is very similar to the Unit Exam
Success on Unit exam is directly related to practice exam experience