

Solution Concentrations



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% Concentration by Mass

grams of solute per 100 grams of solution

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



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

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



NOTICE:
g solution = g solute + g solvent

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% Concentration by Mass

Example

5.5% (by mass) dextrose:

5.5 g dextrose dissolved in 100 g solution

$$\% = 5.5 \text{ g} / 100\text{g} \rightarrow 5.5$$

5.5 g dextrose dissolved in 94.5 g water

$$\% = 5.5 \text{ g} / (5.5 + 94.5) \text{ g} \rightarrow 5.5$$

May be expressed as:

5.5 % (m:m)

5.5 % (w:w)



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Problem: % Concentration by Mass

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

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

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

$$\text{Mass \%} = 3.8 \text{ g} / 61.8\text{g} \times 100 = 6.14887 \rightarrow 6.1$$



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Problem: % Concentration by Mass

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

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

$$4.82 = \frac{\text{grams solute}}{235 \text{ g}} \times 100$$

$$\text{Solute} = (4.82) (235 \text{ g}) / 100 = 11.327 \rightarrow 11.3 \text{ g}$$



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Problem: % Concentration by Mass

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

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

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

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

$$\text{grams solution} = 348.148 \rightarrow 348 \text{ g}$$



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% by Mass Problem Has Two Forms

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:

$$\% \text{ NaBr (by mass)} = 22.4 \text{ g} / (22.4 \text{ g} + 287 \text{ g}) \times 100 = 7.24 \%$$



2. Prepared by dissolving 22.4 g of NaBr in water to make 287 g of solution:

$$\% \text{ NaBr (by mass)} = 22.4 \text{ g} / (287 \text{ g}) \times 100 = 7.80 \%$$

Pay attention to terms: Solute, Solvent, and Solution



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Problem: % Concentration by Mass

How many grams of sodium sulfate are in 505.0 g of a 15.0% solution? How many grams of water?

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

$$15.0 = \frac{\text{grams Na}_2\text{SO}_4}{505.0 \text{ g}} \times 100$$

$$\text{grams Na}_2\text{SO}_4 = \frac{(505 \text{ g})(15.0)}{100}$$

$$\text{grams Na}_2\text{SO}_4 = 75.75 \rightarrow 75.8 \text{ g}$$

$$\text{grams H}_2\text{O} = 505.0 \text{ g} - 75.8 \text{ g}$$

$$\text{grams H}_2\text{O} = 429.2 \text{ g}$$



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Weight (Mass): Volume

Weighing solvents often cumbersome

Practical measurement:

Weigh solute (grams)

Dissolve in solvent

Bring (accurately) to desired volume

Express as % (w:v ... weight: volume)

Express as % (m:v ... mass: volume)



Example

5.5 g of solute brought to 100 ml solution

5.5 % (w:v) or 5.5 % (m:v)

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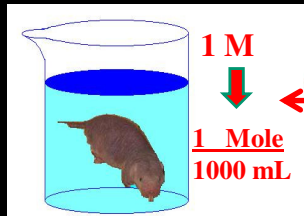
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Molarity

Primary means of calculating solution concentrations

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

$$M = \frac{\text{moles solute}}{\text{liters solution}} = \text{moles / L} \quad (\text{mol / L or mol / 1000 mL})$$



Think



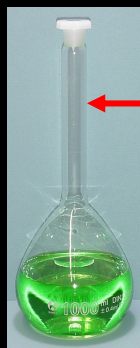
$$\# \text{ moles solute} = \frac{\# \text{ moles}}{1000 \text{ mL}} \times \text{mL}$$

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Molar Solutions

Weigh solute (grams for molar equivalent)
Dissolve in small amount solvent
Bring (accurately) to desired volume
Use Volumetric Flask (Beyond CEM 101)



Only 1 mark (desired volume)

Molar Solutions:

$$\frac{\text{Moles of Solute}}{\text{Volume of Solution (L)}}$$

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Problem: Grams Solute From Molarity

How many grams of HCl would be in 1.00 liter of a 1.50 M solution?

Given: 1.50 moles HCl per L; 1.00 L of solution

Wanted: g HCl

Grams requested, need molar mass for HCl (36.46 g / mole)

$$\frac{1.50 \text{ moles HCl}}{1 \text{ L}} \times 1.00 \text{ L} \times \frac{36.46 \text{ g}}{1 \text{ mole HCl}} = 54.69 \rightarrow 54.7 \text{ g HCl}$$



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Problem: Molarity From Grams Solute and Solution Volume

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 400.0 mL

Wanted: Molarity (M/L)

Grams requested, need molar mass for KBr (119.01 g/mole)

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



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Problem: Grams Solute Given Solution Volume and Molarity

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 moles/liter (M) KBr

Wanted: g KBr

Grams requested, need molar mass for KBr (119.01 g/mole)

$$\frac{0.188 \text{ moles}}{1 \text{ L}} \times \frac{1 \text{ L}}{1000 \text{ ml}} \times 250.0 \text{ ml} \times \frac{119.01 \text{ g}}{\text{mole}} = 5.59347 \rightarrow 5.59 \text{ g KBr}$$

Check:

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



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Problem: Volume From Grams Solute and Solution Molarity

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 (moles/L) solution

Grams requested, need molar mass for KBr (119.01 g/mole)

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



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Problem: Moles From Volume and Solution Molarity

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

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

Check:

$$0.0375 \text{ moles} / 0.250 \text{ L} = 0.150 \text{ M}$$



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Problem: Molarity From Grams Solute and Solution Volume

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 g/mole)

Wanted: molarity of 50.0 mL solution



$$2.41 \text{ g} \times \frac{1 \text{ mole}}{166.01 \text{ g}} \times \frac{1}{50.0 \text{ mL}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 0.290 \text{ moles/L}$$



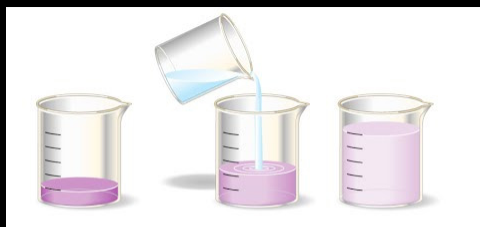
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Dilutions



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Dilutions: General



Molarity = $\frac{\text{Moles}}{\text{Liter}}$ → Moles = Molarity x Liter

Since dilution does not change # moles present:

Volume₁ x Molarity₁ = # moles = Volume₂ x Molarity₂

$$V_1M_1 = V_2M_2$$



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Dilutions: Problem

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 diluted

$$0.360 \text{ M} \times 2.50 \text{ L} = 18.0 \text{ M} \times \text{Needed Volume}$$

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

$$\text{Needed Volume} = 0.0500 \text{ L} \rightarrow 50.0 \text{ mL}$$

So, Dilute 50.0 mL 18 M H₂SO₄ to 2.50 L of solution



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Dilutions: Problem



How many mL of solvent must be added to 345mL of a 14.5 M solution of sodium nitrate to dilute the solution to 6.75 M?

Given: Dilution of 345 mL 14.5 M NaNO_3

Needed: mL of 6.75 M NaNO_3 solution

Hint: moles in final solution \rightarrow same as moles diluted

$$(345 \text{ mL}) (14.5 \text{ M}) = V (6.75 \text{ M})$$

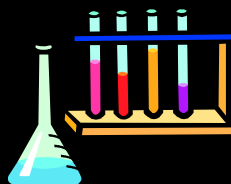
$$V = 741.111 \text{ mL} \rightarrow 741 \text{ mL}$$



741 is the final volume of 6.75 M Solution
So, $741 \text{ mL} - 345 \text{ mL} = 396 \text{ mL H}_2\text{O}$ added

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Normal Solutions



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Equivalent Mass (Weight)

Used for Acids and Bases

For Acids:

One Equivalent = amount that furnishes 1 mole of H^+

For Bases:

One Equivalent = amount that furnishes 1 mole of OH^-

$$\text{Acid Equivalent Mass} = \frac{\text{Molar Mass}}{\text{\# of protons}}$$



$$\text{Base Equivalent Mass} = \frac{\text{Molar Mass}}{\text{\# of hydroxides}}$$

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Normal Solutions

1 Normal Solution = $\frac{\text{Equivalent mass}}{1 \text{ L of solution}}$



Calculate Normality of 154.0 g H_2SO_4 in 500 mL H_2O
Molar Mass $H_2SO_4 = 98.0$

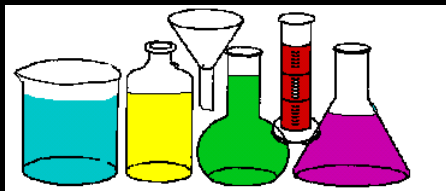
$$154.0 \text{ g} \times \frac{1 \text{ mole}}{98.0 \text{ g}} \times \frac{2 \text{ equivalents}}{1 \text{ mole}} \times \frac{1}{0.500 \text{ L}} = 6.286 \text{ N}$$



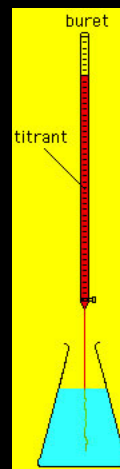
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Solution Stoichiometry



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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 ? $\text{Na}_2\text{SO}_4 + \text{CaCl}_2 \rightarrow 2 \text{NaCl} + \text{CaSO}_4(\text{s})$

Given: 15.8 g CaSO_4 (Molar Mass 136.14 g / mole)

Wanted: mL 0.155 M CaCl_2 solution

Start with # moles of given (known) substance:

$$15.8 \text{ g CaSO}_4 \times \frac{1 \text{ mole}}{136.14 \text{ g}}$$

Use per expression from reaction coefficients \rightarrow moles wanted

$$\times \frac{1 \text{ mole CaCl}_2}{1 \text{ mole CaSO}_4}$$

Convert moles wanted to equivalent solution concentration

$$\times \frac{1 \text{ L}}{0.155 \text{ mole}} \times \frac{1000 \text{ mL}}{1 \text{ L}}$$

As Linear String: (no need to isolate intermediate value)

$$15.8 \text{ g CaSO}_4 \times \frac{1 \text{ mole}}{136.14 \text{ g}} \times \frac{1 \text{ mole CaCl}_2}{1 \text{ mole CaSO}_4} \times \frac{1000 \text{ mL}}{0.155 \text{ mole CaCl}_2} = 749 \text{ mL}$$



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How many mL of a 0.155 M CaCl₂ solution will react with 47.7 mL of a 0.248 M Na₂SO₄ solution? $\text{Na}_2\text{SO}_4 + \text{CaCl}_2 \rightarrow 2 \text{NaCl} + \text{CaSO}_4(\text{s})$

Given: 47.7 mL of 0.248 M Na₂SO₄

Wanted: mL 0.155 M CaCl₂ solution



$$\frac{0.248 \text{ moles Na}_2\text{SO}_4}{1000 \text{ mL}} \times 47.7 \text{ mL} \times \frac{1 \text{ mole CaCl}_2}{1 \text{ mole Na}_2\text{SO}_4} \times \frac{1000 \text{ mL}}{0.155 \text{ mole}} = 76.3 \text{ mL}$$

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How many grams of AgCl can be precipitated by adding excess NaCl to 65.0 mL of 0.757 M AgNO₃? $\text{AgNO}_3(\text{aq}) + \text{NaCl}(\text{aq}) \rightarrow \text{AgCl}(\text{s}) + \text{NaNO}_3(\text{aq})$

Given: 65.0 mL of 0.757 M AgNO₃

Wanted: grams AgCl (molar mass = 143.32 g/mole)



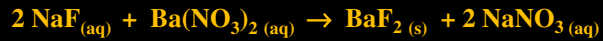
$$\frac{0.757 \text{ moles AgNO}_3}{1000 \text{ mL}} \times 65.0 \text{ mL} \times \frac{1 \text{ mole AgCl}}{1 \text{ mole AgNO}_3} \times \frac{143.32 \text{ g}}{\text{mole}} = 7.05 \text{ g AgCl}$$

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



Given: 25.0 mL of 0.465 M NaF

Wanted: # grams BaF_2 (molar mass = 175.34 g/mole)



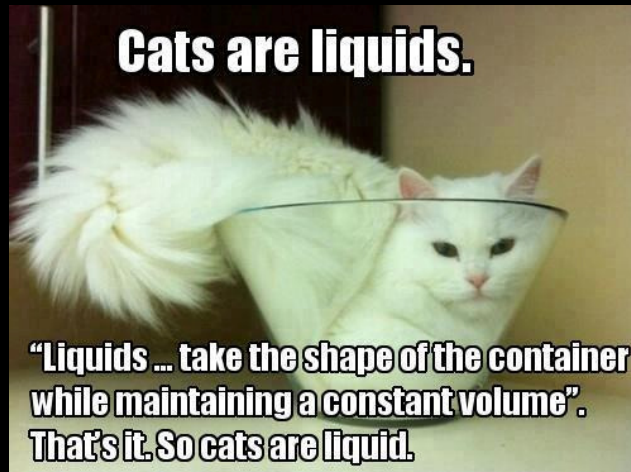
$$\frac{0.465 \text{ moles NaF}}{\text{L}} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times 25.0 \text{ mL} \times \frac{1 \text{ mole BaF}_2}{2 \text{ moles NaF}} \times \frac{175.34 \text{ g}}{1 \text{ mole}} = 1.02 \text{ g}$$

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Cats are liquids.



“Liquids ... take the shape of the container while maintaining a constant volume”. That’s it. So cats are liquid.

Q: What do you call a tooth in a glass of water?

A: One molar solution.

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