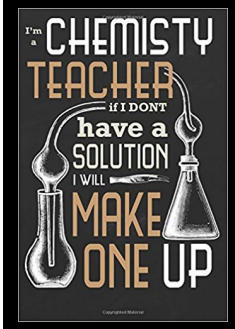



# Solution Concentrations



LPT

## % Concentration by Mass

**grams of solute per 100 grams of solution**

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

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

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

**NOTICE:**
**g solution = g solute + g solvent**



LPT

## % Concentration by Mass

**Example**

**5.5% (by mass) dextrose:**

**5.5 g dextrose dissolved in 100 g solution**

**% = 5.5 g / 100g ➔ 5.5**

**5.5 g dextrose dissolved in 94.5 g water**

**% = 5.5 g / (5.5 + 94.5) g ➔ 5.5**

**May be expressed as:**
**5.5 % (m:m)**
**5.5 % (w:w)**



LPT

## Problem: % Concentration by Mass

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

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

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

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



LPT

## Problem: % Concentration by Mass

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

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

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

Solute = (4.82) (235 g) / 100 = 11.327 → 11.3 g



LPT

## Problem: % Concentration by Mass

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

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

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

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

grams solution = 348.148 → 348 g



LPT

## % 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:

% 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 %

**Pay attention to terms: Solute, Solvent, and Solution**



LPT

## 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?

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

$$15.0 = \frac{\text{grams } Na_2SO_4}{505.0\ g} \times 100$$

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

grams $Na_2SO_4$ = 75.75 ➔ 75.8 g

grams $H_20$ = 505.0 g – 75.8 g

grams $H_20$ = 429.2 g



LPT

## 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)



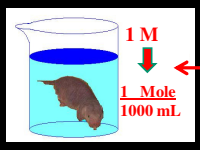
LPT

## Molarity

Primary means of calculating solution concentrations

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

$M = \frac{\text{moles solute}}{\text{liters solution}}$ = moles / L (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:

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



LPT

## 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)

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

Let the units drive the solution



LPT

**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)

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

Let the units drive the solution



LPT

**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)

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

Check:

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

Let the units drive the solution



LPT

**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)

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

Let the units drive the solution



LPT

## 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

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

Check:

0.0375 moles / 0.250 L = 0.150 M

Let the units drive the solution



LPT

## 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

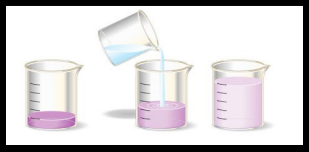
$$2.41\ g \times \frac{1\ mole}{166.01\ g} \times \frac{1}{50.0\ mL} \times \frac{1000\ mL}{1\ L} = 0.290\ moles/L$$

Let the units drive the solution



LPT

## Dilutions





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## Dilutions: General

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

Since dilution does not change # moles present:

$Volume_1 \times Molarity_1 = \#\ moles = Volume_2 \times Molarity_2$

$V_1M_1 = V_2M_2$



LPT

## Dilutions: Problem

How would you prepare 2.50 L of a 0.360 M solution of sulfuric acid ($H_2SO_4$) starting with 18.0 M sulfuric acid?

Given: Dilution of 18.0 M $H_2SO_4$

Needed: 2.50 L of 0.360 M solution

Hint: moles in final solution → same as moles diluted

0.360 M x 2.50 L = 18.0 M x Needed Volume

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

Needed Volume = 0.0500 L → 50.0 mL

So, Dilute 50.0 mL 18 M $H_2SO_4$ to 2.50 L of solution



LPT

## 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 → same as moles diluted

(345 mL) (14.5 M) = V (6.75 M)

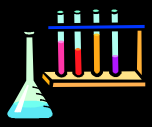
V = 741.111 mL → 741 mL

741 is the final volume of 6.75 M Solution

So, 741 mL – 345 mL = 396 mL $H_2O$ added



LPT



## Normal Solutions



LPT

## 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 $OH^-$

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

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



LPT

## Normal Solutions

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

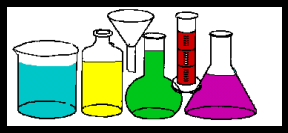
Calculate Normality of 154.0 g $H_2SO_4$ in 500 mL $H_2O$
Molar Mass $H_2SO_4$ = 98.0

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

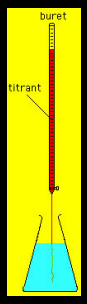
Let the units drive the solution



LPT



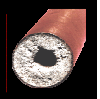
# Solution Stoichiometry






LPT

**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 $CaSO_4$ (Molar Mass 136.14 g / mole)
Wanted: mL 0.155 M $CaCl_2$ solution

Start with # moles of given (known) substance:

$$15.8\ g\ CaSO_4 \times \frac{1\ mole}{136.14\ g}$$

Use per expression from reaction coefficients ➔ moles wanted

$$\times \frac{1\ mole\ CaCl_2}{1\ mole\ CaSO_4}$$

Convert moles wanted to equivalent solution concentration

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

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

$$15.8\ g\ CaSO_4 \times \frac{1\ mole}{136.14\ g} \times \frac{1\ mole\ CaCl_2}{1\ mole\ CaSO_4} \times \frac{1000\ mL}{0.155\ mole\ CaCl_2} = 749\ mL$$



LPT

**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: 47.7 mL of 0.248 M $Na_2SO_4$
Wanted: mL 0.155 M $CaCl_2$ solution

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

Let the units drive the solution



LPT

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

Given: 65.0 mL of 0.757 M $AgNO_3$
Wanted: grams AgCl (molar mass = 143.32 g/mole)

$$\frac{0.757 \text{ moles } AgNO_3}{1000 \text{ mL}} \times 65.0 \text{ mL} \times \frac{1 \text{ mole AgCl}}{1 \text{ mole } AgNO_3} \times \frac{143.32 \text{ g}}{\text{mole}} = 7.05 \text{ g AgCl}$$

Let the units drive the solution



LPT

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

$2\,NaF_{(aq)} + Ba(NO_3)_{2\,(aq)} \rightarrow BaF_{2\,(s)} + 2\,NaNO_{3\,(aq)}$

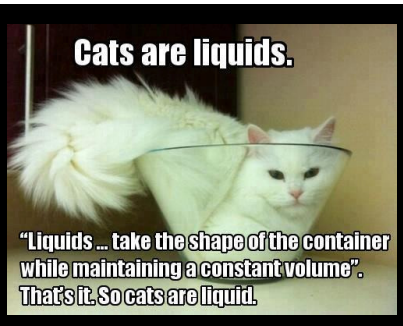
Given: 25.0 mL of 0.465 M NaF
Wanted: # grams $BaF_2$ (molar mass = 175.34 g/mole)

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

Let the units drive the solution



LPT




**Q: What do you call a tooth in a glass of water?**
**A: One molar solution.**



LPT