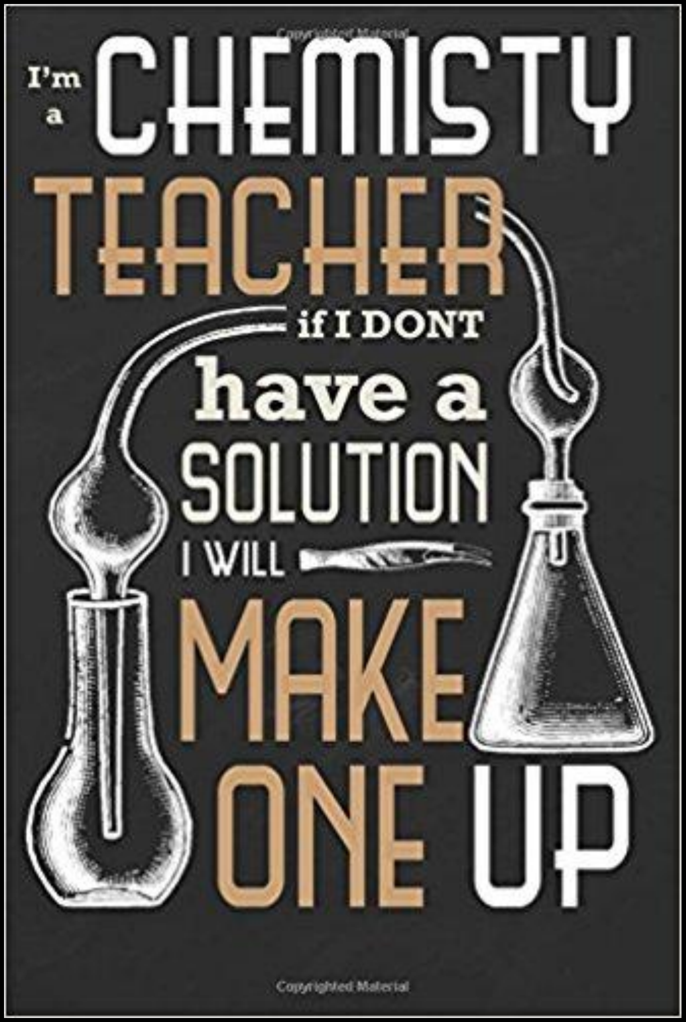


Solution Concentrations

LPT

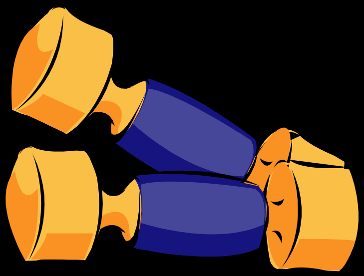
% Concentration by Mass

grams of solute per 100 grams of solution

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

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

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



NOTICE:

g solution = g solute + g solvent

% 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.

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

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

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

LPT

Problem: % Concentration by Mass

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

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

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

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

LPT

Problem: % Concentration by Mass

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

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

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

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

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

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

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?

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

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

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

$$\text{grams } Na_2SO_4 = 75.75 \rightarrow 75.8 \text{ g}$$

$$\text{grams } H_20 = 505.0 \text{ g} - 75.8 \text{ g}$$

$$\text{grams } H_20 = 429.2 \text{ g}$$

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

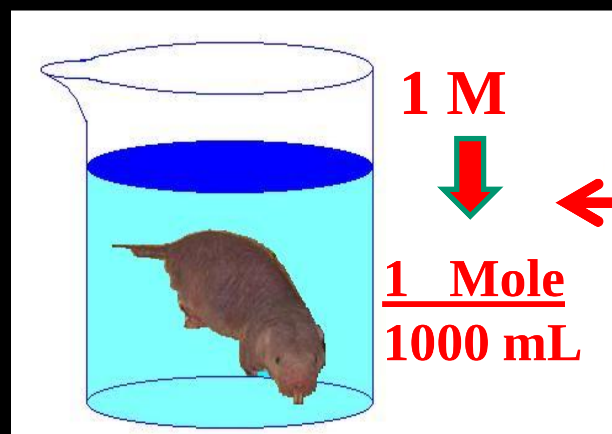
5.5 % (w:v) or 5.5 % (m:v)

Molarity

Primary means of calculating solution concentrations

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

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

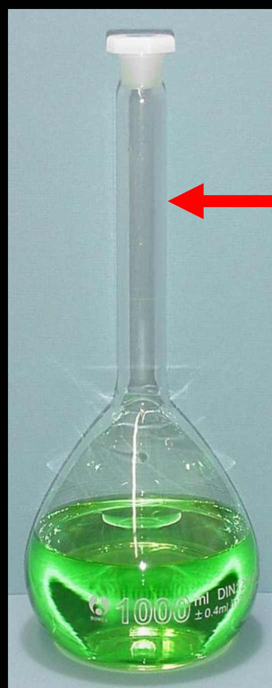


Think

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

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)

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

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)

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

Check:

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

Let the units drive the the solution

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)

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

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

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

Check:

0.0375 moles / 0.250 L = 0.150 M

Let the units drive the solution

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)

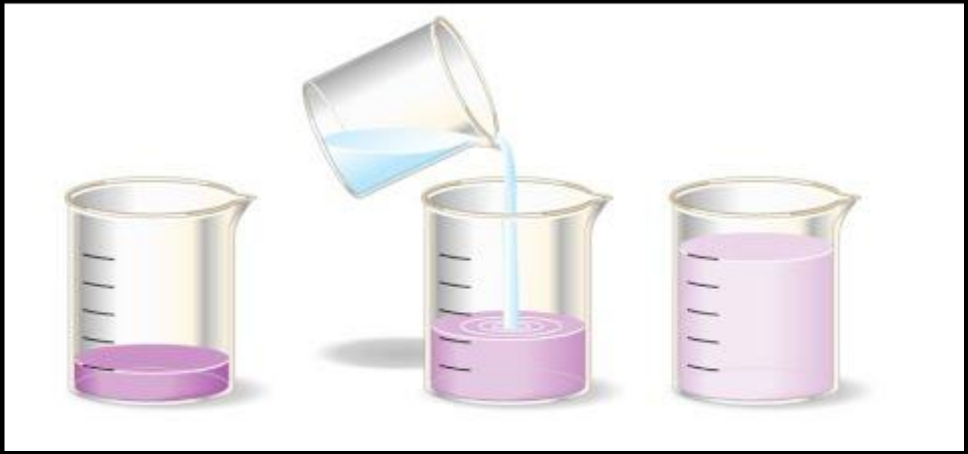
Wanted: molarity of 50.0 mL solution

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

Let the units drive the solution

LPT

Dilutions



LPT

Dilutions: General

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

Since dilution does not change # moles present:

$\text{Volume}_1 \times \text{Molarity}_1 = \# \text{ moles} = \text{Volume}_2 \times \text{Molarity}_2$

$V_1M_1 = V_2M_2$

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

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

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?

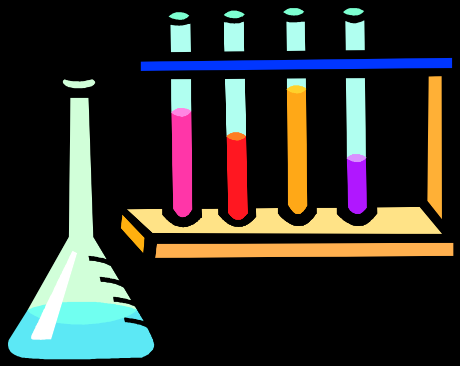
Given: Dilution of 345 mL 14.5 M $NaNO_3$
Needed: mL of 6.75 M $NaNO_3$ solution

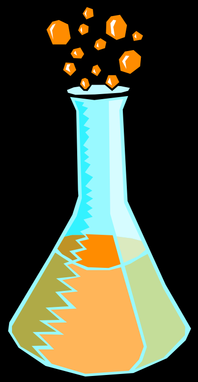
Hint: moles in final solution ➔ same as moles diluted

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

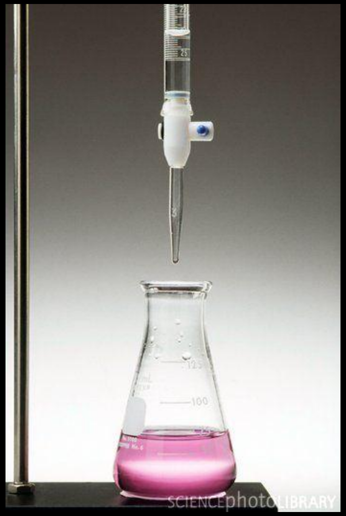
V = 741.111 mL ➔ 741 mL

741 is the final volume of 6.75 M Solution
So, 741 mL – 345 mL = 396 mL H_2O added





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

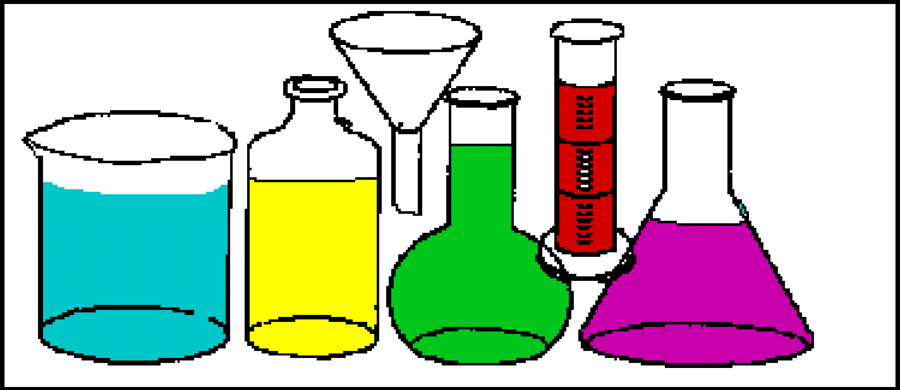
$$1\ \text{Normal Solution} = \frac{\text{Equivalent mass}}{1\ \text{L of solution}}$$

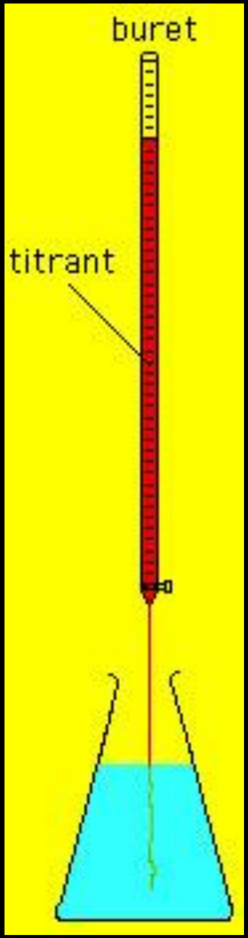
Calculate Normality of 154.0 g H_2SO_4 in 500 mL H_2O

Molar Mass H_2SO_4 = 98.0

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

Let the units drive the solution





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:

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

Use per expression from reaction coefficients ➔ 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)

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

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

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

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?

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

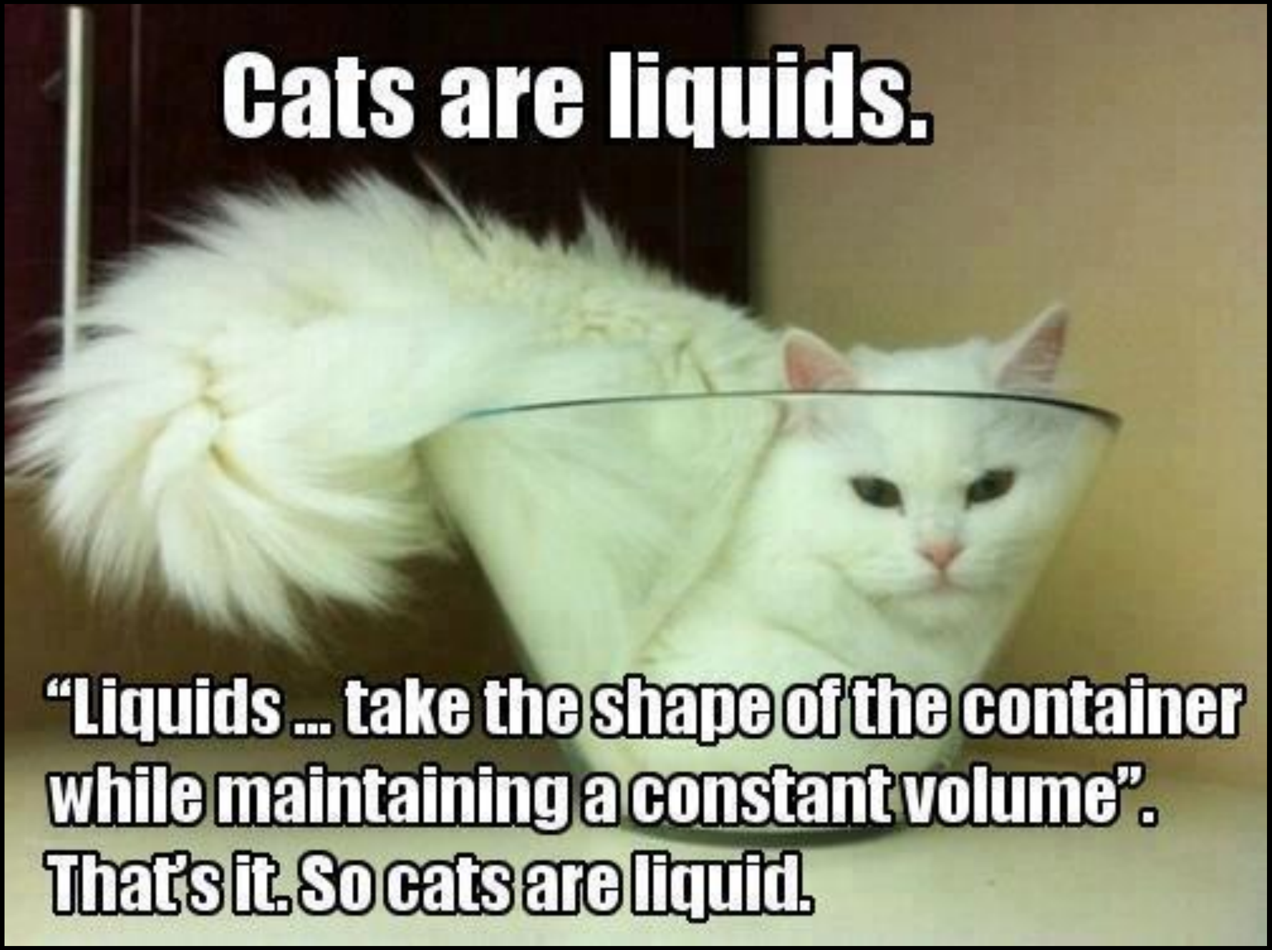
Given: 25.0 mL of 0.465 M NaF

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