Example: How many grams of barium chloride do we need to weigh out to get 3.65 moles of barium chloride?

Find the FORMULA of barium chloride:

Ba^{2+} Cl^-

Cl^-

$BaCl_2$

Find the FORMULA WEIGHT

Ba : 1 x 137.3

Cl : 2 x 35.45

208.2 g $BaCl_2$ = mol $BaCl_2$

$$3.65 \text{ mol } \cancel{BaCl_2} \times \frac{208.2 \text{ g } BaCl_2}{\text{mol } \cancel{BaCl_2}} = \boxed{760. \text{ g } BaCl_2}$$

PERCENTAGE COMPOSITION

- sometimes called "percent composition" or "percent composition by mass"

- the percentage of each element in a compound, expressed in terms of mass

Example: Find the percentage composition of barium chloride.

$BaCl_2$: Ba: $1 \times 137.3 = \boxed{137.3}$

Cl: $2 \times 35.45 = \boxed{70.90}$

These numbers are the masses of each element in a mole of the compound!

$208.2\ g\ BaCl_2 = mol\ BaCl_2$

$$Ba: \frac{137.3\ g\ Ba}{208.2\ g\ BaCl_2} \times 100\% = 65.95\%\ Ba$$

$$Cl: \frac{70.90\ g\ Cl}{208.2\ g\ BaCl_2} \times 100\% = 34.05\%\ Cl$$

As a check, these should sum to 100% within roundoff error!

So far, we have

- looked at how to determine the composition by mass of a compound from a formula
- converted from MASS to MOLES (related to the number of atoms/molecules)
- converted from MOLES to MASS

Are we missing anything?

- What about SOLUTIONS, where the desired chemical is not PURE, but found DISSOLVED IN WATER?
- How do we deal with finding the moles of a desired chemical when it's in solution?

MOLAR CONCENTRATION *

- unit: MOLARITY (M): moles of dissolved substance per LITER of solution

dissolved substance

$$M = \text{molarity} = \frac{\text{moles of SOLUTE}}{\text{L SOLUTION}}$$

6.0 M HCl solution: $\frac{6.0 \text{ mol HCl}}{\text{L}}$

If you have 0.250 L (250 mL) of 6.0 M HCl, how many moles of HCl do you have?

6.0 mol HCl = L

$$0.250 \cancel{L} \times \frac{6.0 \text{ mol HCl}}{\cancel{L}} = \boxed{1.5 \text{ mol HCl}}$$

*See SECTIONS 4.7 - 4.10 for more information about MOLARITY and solution calculations (p 154 - 162 - 9th edition) (p 156-164 - 10th edition)

If you need 0.657 moles of hydrochloric acid, how many liters of 0.0555 M HCl do you need to measure out?

$$\frac{0.0555 \text{ mol HCl}}{L}$$

$$0.657 \text{ mol HCl} \times \frac{L}{0.0555 \text{ mol HCl}} = \boxed{11.8 L}$$

11800 mL

In most lab situations, this solution wouldn't be practical if we needed 0.657 mol

What if we used 6.00 M HCl?

$$\frac{6.00 \text{ mol HCl}}{L}$$

$$0.657 \text{ mol HCl} \times \frac{L}{6.00 \text{ mol HCl}} = \boxed{0.110 L}$$

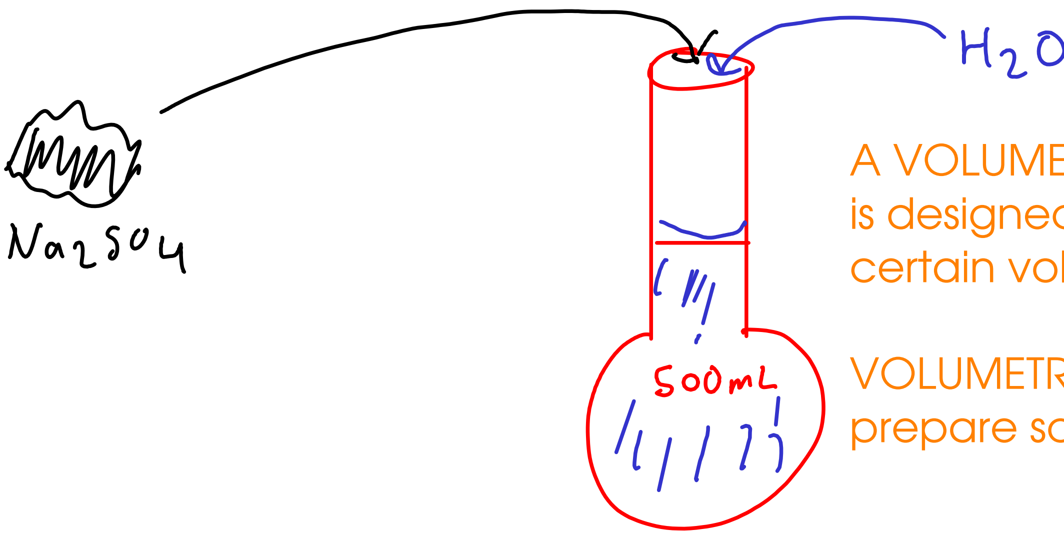
110. mL

This is a more lab-scale volume than the first one.

Example: How would we prepare 500. mL of 0.500 M sodium sulfate in water?

Na_2SO_4 : 142.05 g/mol

Dissolve the appropriate amount of sodium sulfate into enough water to make 500. mL of solution.



A VOLUMETRIC FLASK is a flask that is designed to precisely contain a certain volume of liquid.

VOLUMETRIC FLASKS are used to prepare solutions.

volumetric flask

We will calculate the moles of sodium sulfate that are present in 500. mL of solution. Then, we will convert the moles of sodium sulfate to mass.

$0.500 \text{ mol } Na_2SO_4 = L$ | $mL = 10^{-3} L$ | $142.05 \text{ g } Na_2SO_4 = \text{mol } Na_2SO_4$

$$500.\,\cancel{mL} \times \frac{10^{-3}\,\cancel{L}}{\cancel{mL}} \times \frac{0.500\,\cancel{\text{mol } Na_2SO_4}}{\cancel{L}} \times \frac{142.05 \text{ g } Na_2SO_4}{\cancel{\text{mol } Na_2SO_4}} = \boxed{35.5 \text{ g } Na_2SO_4}$$

To prepare the solution, put 35.5 g sodium sulfate into a 500 mL volumetric flask, and add enough distilled water to get 500 mL of solution.

More on MOLARITY

To prepare a solution of a given molarity, you generally have two options:

1. Weigh out the appropriate amount of solute, then dilute to the desired volume with solvent (usually water)

"stock solution"

2. Take a previously prepared solution of known concentration and DILUTE it with solvent to form a new solution

- Use DILUTION EQUATION

The dilution equation is easy to derive with simple algebra.

$$M \times V$$

$$\frac{mol}{L} \times L = \text{moles solute}$$

... but when you dilute a solution, the number of moles of solute REMAINS CONSTANT. (After all, you're adding only SOLVENT)

$$M_1V_1 = M_2V_2$$

Since the number of moles of solute stays the same, this equality must be true!

before diution

after dilution

$$M_1V_1 = M_2V_2$$ ... the "DILUTION EQUATION"

M_1 = molarity of concentrated solution

V_1 = volume of concentrated solution

M_2 = molarity of dilute solution

V_2 = volume of dilute solution (total volume, not volume of added solvent!)

The volumes don't HAVE to be in liters, as long as you use the same volume UNIT for both volumes!

Example: Take the 0.500 M sodium sulfate we discussed in the previous example and dilute it to make 150. mL of 0.333 M solution. How many mL of the original solution will we need to dilute?

$$M_1V_1 = M_2V_2$$

$M_1 = 0.500\,M$ $\quad M_2 = 0.333\,M$

$V_1 = ?$ $\quad V_2 = 150.\,mL$

$$(0.500\,M) \times V_1 = (0.333\,M) \times (150.\,mL)$$

$$V_1 = \boxed{99.9\,mL \text{ of } 0.500\,M\ Na_2SO_4}$$

To prepare the new solution, measure out 99.9 mL of the 0.500 M sodium sulfate, then add water until the total volume equals 150. mL.

CHEMICAL CALCULATIONS CONTINUED: REACTIONS

- Chemical reactions proceed on an ATOMIC basis, NOT a mass basis!

- To calculate with chemical reactions (i.e. use chemical equations), we need everything in terms of ATOMS ... which means MOLES of atoms

$$\underline{2}\,Al(s) + \underline{3}\,Br_2(l) \longrightarrow \underline{2}\,AlBr_3(s)$$

coefficients are in terms of atoms and molecules!

2 atoms Al = 3 molecules Br_2 = 2 formula units $AlBr_3$

2 mol Al = 3 mol Br_2 = 2 mol $AlBr_3$

- To do chemical calculations, we need to:
  - Relate the amount of substance we know (mass or volume) to a number of moles
  - Relate the moles of one substance to the moles of another using the equation
  - Convert the moles of the new substance to mass or volume as desired