Example: You need 1.75 moles of iron. What mass of iron do you need to weigh out on the balance?

Fe: 55.85 g Fe = mol Fe

$$1.75\ \cancel{\text{mol Fe}} \times \frac{55.85\ \text{g Fe}}{\cancel{\text{mol Fe}}} = \boxed{97.7\ \text{g Fe}}$$

WHAT ABOUT COMPOUNDS? FORMULA WEIGHT

Example: 25.0 g of WATER contain how many MOLES of water molecules?

H_2O :   H : 2 x 1.008 = 2.016

O : 1 x 16.00 = 16.00

18.016 FORMULA WEIGHT of water

FORMULA WEIGHT is the mass of one mole of either an element OR a compound.

18.016 g H_2O = mol H_2O

$$25.0\,g\,H_2O \times \frac{mol\,H_2O}{18.016\,g\,H_2O} = \boxed{1.39\,mol\,H_2O}$$

Formula weight goes by several names:

- For atoms, it's the same thing as ATOMIC WEIGHT
- For molecules, it's called MOLECULAR WEIGHT
- Also called "MOLAR MASS"

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

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 $BaCl_2$

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

or, 7.60×10^2

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: $1 \times 137.3 = 137.3$ ← Mass of barium in 1 mol barium chloride

Cl: $2 \times 35.45 = 70.90$ ← Mass of chlorine in 1 mol barium chloride

$208.2\ g\ BaCl_2 = mol\ BaCl_2$

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

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

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

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?

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

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

11800 mL

This volume is quite large for lab-scale work

What if we used 6.00 M HCl?

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

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

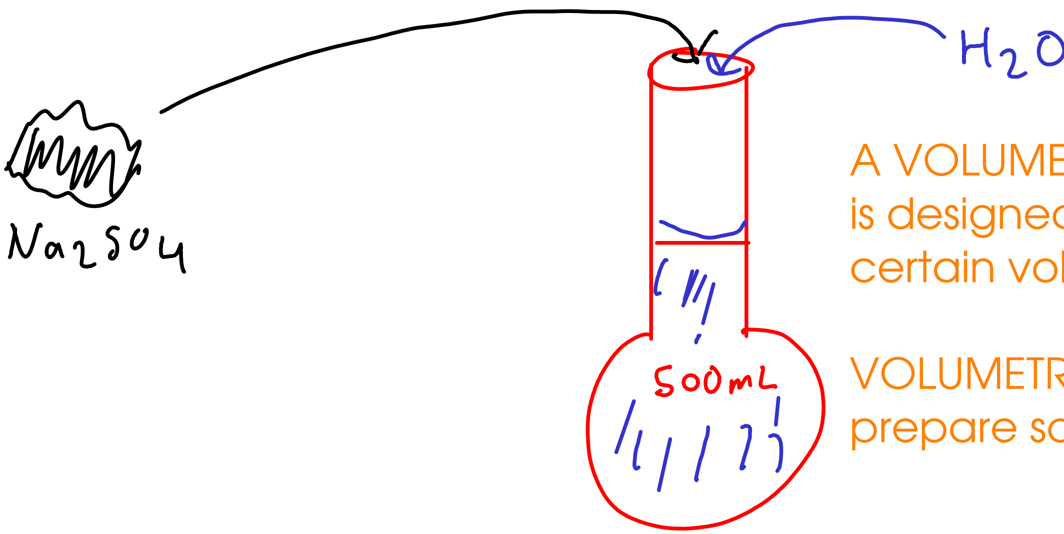
110. mL

This volume is more reasonable for lab-scale work

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 know we need 500. mL of 0.500 M sodium sulfate. Calculate how many moles of sodium sulfate would be in the soluution, then change that number to grams.

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

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

Weigh out 35.5 grams sodium sulfate, put into 500 mL volumetric flask, and dilute to the mark with water.

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.

$$M \times V$$

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

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

$$M_1 V_1 = M_2 V_2$$

before diution

after dilution

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

$$M_1V_1 = M_2V_2$$ ... 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?

$$M_1V_1 = M_2V_2$$

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

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

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

$$V_1 = 99.9\ mL$$

Take 99.9 mL of 0.500 M sodium sulfate and add enough water to get 150. mL total.