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

Fe: 55.85 (atomic weight)

55.85 g Fe = mol Fe

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

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

18.016 g H_2O = mol H_2O

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

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

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?

First, we need to find the chemical formula of barium chloride:

Ba^{2+} Cl^-

Cl^-

$BaCl_2$

Second, calculate formula weight of barium chloride.

$BaCl_2$ – Ba: 1 x 137.3

Cl: 2 x 35.45

208.2

208.2 g $BaCl_2$ = mol $BaCl_2$

Finally, calculate the required mass by using the formula weight as a conversion factor:

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

$(76\bar{0} \text{ g})$

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!

$208.2\ g\ BaCl_2 = \text{mol}\ BaCl_2$

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

These percentages should sum to 100% (within roundoff error)

$$\%Cl = \frac{70.90\ g\ Cl_2}{208.2\ g\ BaCl_2} \times 100 = 34.05\%\ Cl$$

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

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

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

*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 mol HCl = L

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

(11800 mL)

What if we used 6.00 M HCl?

6.00 mol HCl = L

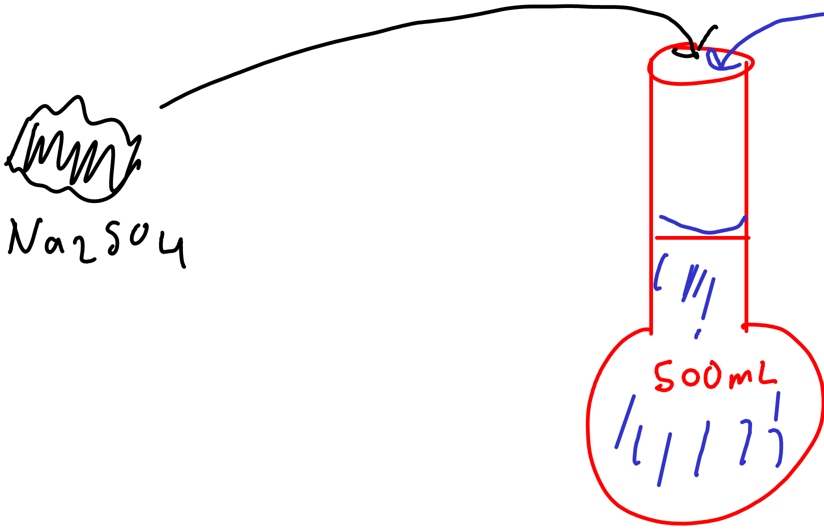
$$0.657 \text{ mol HCl} \times \frac{\text{L}}{6.00 \text{ mol HCl}} = \boxed{0.110 \text{ L}}$$

($11\bar{0}$ mL)

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 can calculate the MOLES of sodium sulfate required from the solution's volume (500 mL) and the solution's molarity (0.500 M). Then, we can calculate the MASS of sodium sulfate needed from the moles and formula weight.

$0.500 \text{ mol } Na_2SO_4 = L \qquad mL = 10^{-3} L$

$$500.\,mL \times \frac{10^{-3} L}{mL} \times \frac{0.500 \text{ mol } Na_2SO_4}{L} = 0.250 \text{ mol } Na_2SO_4$$

$142.05 \text{ g } Na_2SO_4 = \text{mol } Na_2SO_4$

$$0.250 \text{ mol } Na_2SO_4 \times \frac{142.05 \text{ g } Na_2SO_4}{\text{mol } Na_2SO_4} = \boxed{35.5 \text{ g } Na_2SO_4}$$

... to make 500 mL of 0.500 M 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)

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

   "stock 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!