RELATING MASS AND MOLES

- Use DIMENSIONAL ANALYSIS (a.k.a "drag and drop")

- Need CONVERSION FACTORS - where do they come from?

- We use ATOMIC WEIGHT as a conversion factor.

Mg : 24.31 | 24.31 g Mg = 1 mol Mg

Atomic mass

"mol" is the abbreviation for "mole"

Example: How many moles of atoms are there in 250. g of magnesium metal?

24.31 g Mg = mol Mg

$$250.\ g\ \cancel{Mg} \times \frac{mol\ Mg}{24.31\ g\ \cancel{Mg}} = \boxed{10.3\ mol\ Mg}$$

ATOMIC WEIGHT is a MEASURED number - in other words, it has significant figures. Usually we can find atomic weights with more significant figures if necessary.

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

Fe: 55.85

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

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

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 ammonium carbonate do we need to weigh out to get 3.65 moles of ammonium carbonate?

First, find out the correct FORMULA of ammonium carbonate...

$NH_4^+ \quad CO_3^{2-}$

NH_4^+

$(NH_4)_2CO_3$

Then, find the FORMULA WEIGHT

N: 2×14.01

H: 8×1.008

C: 1×12.01

O: 3×16.00

96.094 g $(NH_4)_2CO_3$ = mol $(NH_4)_2CO_3$

Finally, do the calculation...

$$3.65 \text{ mol } (NH_4)_2CO_3 \times \frac{96.094 \text{ g } (NH_4)_2CO_3}{\text{mol } (NH_4)_2CO_3} = \boxed{351 \text{ g } (NH_4)_2CO_3}$$

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 ammonium nitrate.

NH_4NO_3: N: 2 x 14.01 = 28.02

H: 4 x 1.008 = 4.032

O: 3 x 16.00 = 48.00

80.052 g NH_4NO_3 = mol NH_4NO_3

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

$$\%N = \frac{28.02 \text{ g N}}{80.052 \text{ g total}} \times 100\% = 35.00\%\ N$$

$$\%H = \frac{4.032 \text{ g H}}{80.052 \text{ g total}} \times 100\% = 5.04\%\ H$$

$$\%O = \frac{48.00 \text{ g O}}{80.052 \text{ g total}} \times 100\% = 59.96\%\ O$$

Check ... all these percentages 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: $\dfrac{6.0 \text{ mol HCl}}{\text{L}}$

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?

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

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

11800 mL

This volume is too large for lab-scale work. We shoiuld use a more DILUTE solution!

What if we used 6.00 M HCl?

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

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

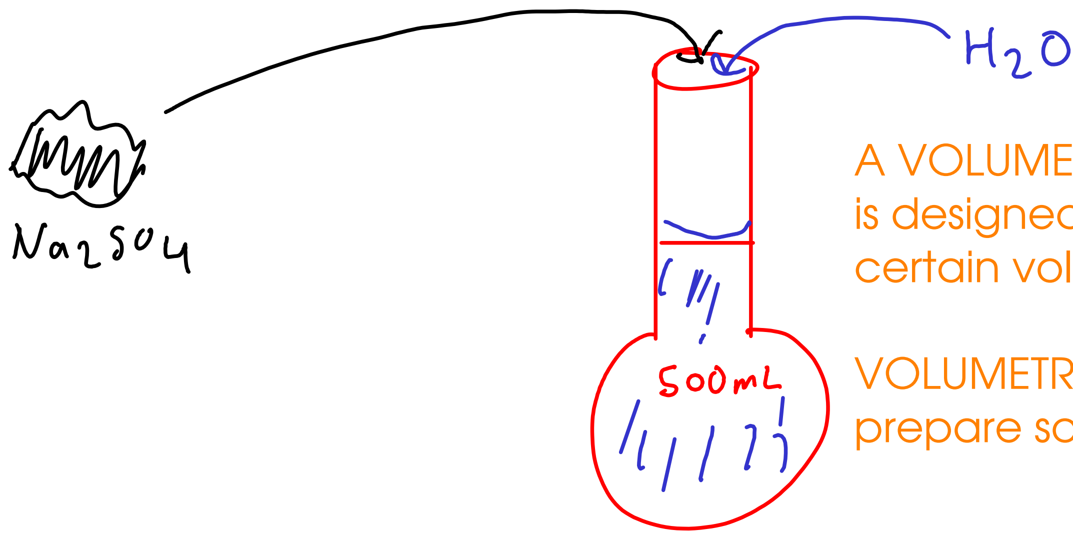
110. mL

This volume is easy to measure with our normal glassware!

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 that we need 500 mL of solution. We ALSO know that the solution must be 0.500 M sodium sulfate. First, find the MOLES sodium sulfate to put in the flask, then convert moles to MASS sodium sulfate.

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

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

To make the solution, weigh out 35.5 grams sodium sulfate, put into a 500 mL volumetric flask, and add water to the mark.