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

$$55.85\,g\,Fe = mol\,Fe$$

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

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

Find the formula of ammonium carbonate!

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

NH_4^+

$(NH_4)_2CO_3$

Find formula weight!

N : 2 x 14.01

H : 8 x 1.008

C : 1 x 12.01

O : 3 x 16.00

96.094

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

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

$$= \boxed{351 \text{ g } (NH_4)_2CO_3}$$

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: $2 \times 14.01 = 28.02$

H: $4 \times 1.008 = 4.032$

O: $3 \times 16.00 = 48.00$

$80.052\ g\ NH_4NO_3 = 1\ mol\ NH_4NO_3$

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

$\%N: \dfrac{28.02\ g\ N}{80.052\ g\ total} \times 100\% = 35.0\%\ N$

$\%H: \dfrac{4.032\ g\ H}{80.052\ g\ total} \times 100\% = 5.0\%\ H$

$\%O: \dfrac{48.00\ g\ O}{80.052\ g\ total} \times 100\% = 60.0\%\ O$

Check: All the percentages should sum to 100% - taking roundoff error into account!

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

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

dissolved substance

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?

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

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

11800 mL

This is too large a volume for lab-scale work. To get a more reasonable volume, we should use a more concentrated solution ... like the 6.00 M HCl described below!

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

110 mL is a reasonable volume for lab-scale work. We can measure this out easily in a 250 mL cylinder.