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?

Translate "ammonium carbonate" to a chemical formula!

NH_4^+ CO_3^{2-}

NH_4^+

$(NH_4)_2CO_3$

$N: 2 \times 14.01$

$H: 8 \times 1.008$

$C: 1 \times 12.01$

$O: 3 \times 16.00$

96.094

Once we know the FORMULA, we can find the FORMULA WEIGHT!

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

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

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 : $3 \times 16.00 = 48.00$

$80.052 \text{ g } NH_4NO_3 = 1 \text{ mol } NH_4NO_3$

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

$$\%N = \frac{28.02 \text{ g N}}{80.052 \text{ g total}} \times 100\% = 35.0\% \text{ N}$$

$$\%H = \frac{4.032 \text{ g H}}{80.052 \text{ g total}} \times 100\% = 5.0\% \text{ H}$$

$$\%O = \frac{48.00 \text{ g O}}{80.052 \text{ g total}} \times 100\% = 60.0\% \text{ O}$$

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)

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

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

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

(11,800 mL)

This is too large of a volume for lab-scale work, so we should pick a hydrochloric acid solution that is more concetrated than this one!

What if we used 6.00 M HCl?

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

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

($11\bar{0}$ mL)

This is a more lab-friendly quantity. We can measure this amount easily with a 250 mL cylinder (or just use a 100 mL cylinder twice!)

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!

$$\frac{2 \text{ atoms } Al = 3 \text{ molecules } Br_2 = 2 \text{ formula units } AlBr_3}{2 \text{ mol } Al = 3 \text{ mol } Br_2 = 2 \text{ mol } AlBr_3}$$

- 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

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

* Given that we have 25.0 g of liquid bromine, how many grams of aluminum would we need to react away all of the bromine? How many grams of aluminum bromide would be produced?

1. Convert grams of bromine to moles: Need formula weight Br_2: $2 \times 79.90 = 159.80$

$159.80\,g\,Br_2 = 1\,mol\,Br_2$

$$25.0\,g\,Br_2 \times \frac{1\,mol\,Br_2}{159.80\,g\,Br_2} = 0.15645\,mol\,Br_2$$

2. Use the chemical equation to relate moles of bromine to moles of aluminum

$2\,mol\,Al = 3\,mol\,Br_2$

$$0.15645\,mol\,Br_2 \times \frac{2\,mol\,Al}{3\,mol\,Br_2} = 0.10430\,mol\,Al$$

3. Convert moles aluminum to mass: Need formula weight Al: 26.98

$26.98\,g\,Al = 1\,mol\,Al$

$$0.10430\,mol\,Al \times \frac{26.98\,g\,Al}{1\,mol\,Al} = \boxed{2.81\,g\,Al}$$

You can combine all three steps on one line if you like!

$$25.0\,g\,Br_2 \times \frac{1\,mol\,Br_2}{159.80\,g\,Br_2} \times \frac{2\,mol\,Al}{3\,mol\,Br_2} \times \frac{26.98\,g\,Al}{1\,mol\,Al} = \boxed{2.81\,g\,Al}$$

1 2 3

You can solve the second part of the question using CONSERVATION OF MASS - since there's only a single product and you already know the mass of all reactants.

$$\begin{array}{r} 25.0\,g\,Br_2 \\ +\ \ 2.81\,g\,Al \\ \hline 27.8\,g\,AlBr_3 \end{array}$$

But ...

...what would you have done to calculate the mass of aluminum bromide IF you had NOT been asked to calculate the mass of aluminum FIRST?

$$25.0\,g\,Br_2 \times \frac{1\,mol\,Br_2}{159.80\,g\,Br_2} \times \frac{2\,mol\,AlBr_3}{3\,mol\,Br_2} \times \frac{266.694\,g\,AlBr_3}{1\,mol\,AlBr_3} = 27.8\,g\,AlBr_3$$

1 convert mass bromine to moles

2 convert moles bromine to moles aluminum bromide

3 convert moles aluminum bromide to mass

Example:

How many milliliters of 6.00M hydrochloric acid is needed to completely react with 25.0 g of sodium carbonate?

$$2HCl(aq) + Na_2CO_3(s) \longrightarrow H_2O(l) + CO_2(g) + 2NaCl(aq)$$

1 - Convert 25.0 g of sodium carbonate to moles. Use formula weight of sodium carbonate
2 - Convert moles of sodium carbonate to moles hydrochloric acid. Use chemical equation
3 - Convert moles hydrochloric acid to volume. Use molarity (6.0M) of solution.

1 Na_2CO_3: Na: 2×22.99

C: 1×12.01

O: 3×16.00

$105.99 \text{ g } Na_2CO_3 = \text{mol } Na_2CO_3$

Formula weight of sodium carbonate

$$25.0 \text{ g } Na_2CO_3 \times \frac{\text{mol } Na_2CO_3}{105.99 \text{ g } Na_2CO_3} = 0.235871 \text{ mol } Na_2CO_3$$

2 $2 \text{ mol } HCl = \text{mol } Na_2CO_3$

$$0.235871 \text{ mol } Na_2CO_3 \times \frac{2 \text{ mol } HCl}{\text{mol } Na_2CO_3} = 0.471743 \text{ mol } HCl$$

Example:

How many milliliters of 6.00M hydrochloric acid is needed to completely react with 25.0 g of sodium carbonate?

$$2HCl(aq) + Na_2CO_3(s) \longrightarrow H_2O(l) + CO_2(g) + 2NaCl(aq)$$

~~1 - Convert 25.0 g of sodium carbonate to moles. Use formula weight of sodium carbonate~~

~~2 - Convert moles of sodium carbonate to moles hydrochloric acid. Use chemical equation~~

3 - Convert moles hydrochloric acid to volume. Use molarity (6.0M) of solution.

3 $6.00 \text{ mol HCl} = L$ $\quad mL = 10^{-3} L$

M = moles solute (HCl) per liter

$$0.471743 \text{ mol HCl} \times \frac{L}{6.00 \text{ mol HCl}} \times \underbrace{\frac{mL}{10^{-3} L}} = \boxed{78.6 \text{ mL of } 6.00 \text{ M HCl}}$$

This step converts volume in L to volume in mL

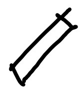
On one line...

$$25.0 g\, Na_2CO_3 \times \underbrace{\frac{mol\, Na_2CO_3}{105.99\, g\, Na_2CO_3}}_{1} \times \underbrace{\frac{2\, mol\, HCl}{mol\, Na_2CO_3}}_{2} \times \underbrace{\frac{L}{6.00\, mol\, HCl} \times \frac{mL}{10^{-3} L}}_{3} = 78.6\, mL$$

CONCEPT OF LIMITING REACTANT

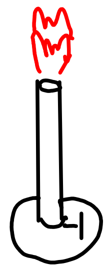
- When does a chemical reaction STOP?

$$2Mg(s) + O_2(g) \xrightarrow{\Delta} 2MgO(s)$$



Magnesium strip

+



Flame, oxygen from air

Magnesium oxide powder

- When does this reaction stop? When burned in open air, this reaction stops when all the MAGNESIUM STRIP is gone. We say that the magnesium is LIMITING.

- This reaction is controlled by the amount of available magnesium

- At the end of a chemical reaction, the LIMITING REACTANT will be completely consumed, but there may be amount of OTHER reactants remaining. We do chemical calculations in part to minimize these "leftovers".

These are often called "excess" reactants, or reactants present "in excess"

LIMITING REACTANT CALCULATIONS

- To find the limiting reactant, calculate how much product would be produced from ALL given reactants. Whichever produces the SMALLEST amouht of product is the limiting reactant, and the smallest anount of product is the actual amount of product produced.

Example:

56.08 12.01 64.10 <- Formula weights

$$CaO(s) + 3\,C(s) \xrightarrow{\Delta} CaC_2(s) + CO(g)$$

If you start with 100. g of each reactant, how much calcium carbide would be produced?

CaO: 56.08 g CaO = mol CaO | 1 mol CaO = 1 mol CaC_2 | 64.10 g CaC_2 = mol CaC_2

$$100.\text{ g CaO} \times \frac{\text{mol CaO}}{56.08\text{ g CaO}} \times \frac{1\text{ mol } CaC_2}{1\text{ mol CaO}} \times \frac{64.10\text{ g } CaC_2}{\text{mol } CaC_2} = \boxed{114\text{ g } CaC_2}$$

C: 12.01 g C = mol C | 3 mol C = 1 mol CaC_2 | 64.10 g CaC_2 = mol CaC_2

$$100.\text{ g C} \times \frac{\text{mol C}}{12.01\text{ g C}} \times \frac{1\text{ mol } CaC_2}{3\text{ mol C}} \times \frac{64.10\text{ g } CaC_2}{\text{mol } CaC_2} = 178\text{ g } CaC_2$$

114 g of calcium carbide should be produced. Calcium oxide runs out when the reaction has produced 114 g of calcium carbide, so no further product can be produced at this point.

We would say that calcium oxide is "limiting", while carbon is present "in excess".

PERCENT YIELD

- Chemical reactions do not always go to completion! Things may happen that prevent the conversion of reactants to the desired/expected product!

① SIDE REACTIONS:

$$C + O_2 \longrightarrow CO_2$$

This reaction occurs when there is a large amount of oxygen available

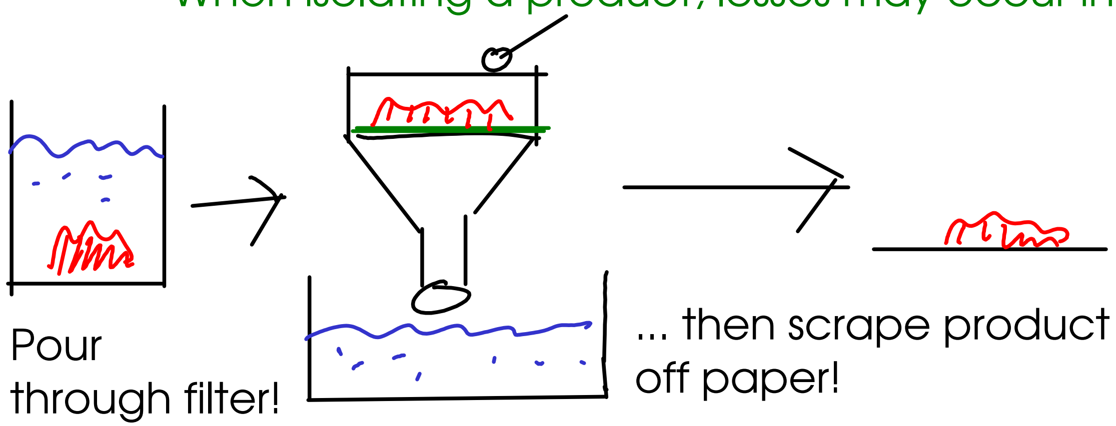
$$2C + O_2 \longrightarrow 2CO$$

... while this reaction is more favorable in low-oxygen environments!

... so in a low-oxygen environment, you may produce less carbon dioxide than expected!

② TRANSFER AND OTHER LOSSES

- When isolating a product, losses may occur in the process. Example: filtering



During each step of this process, some amount of product loss will occur!

③ EQUILIBRIUM

- Reactions may reach an equilbrium between prodcuts and reactants. We'll talk more about this in CHM 111. The net results is that the reaction will appear to stop before all reactants have been consumed!

- All of these factors cause a chemical reaction to produce LESS product than calculated. For many reactions, this difference isn't significant. But for others, we need to report the PERCENT YIELD.

Determined EXPERIMENTALLY!

$$\text{PERCENT YIELD} = \frac{\text{ACTUAL YIELD}}{\text{THEORETICAL YIELD}} \times 100\%$$

Calculated based on the limiting reactant. (The chemical calculations you've done up to now have been theoretical yields!)

... the percent yield of a reaction can never be greater than 100% due to conservation of mass! If you determine that a percent yield is greater than 100%, then you've made a mistake somewhere - either in a calculation or in the experiment itself!