CHEMICAL CALCULATIONS - RELATING MASS AND ATOMS

$$\_\ Na_2CO_3 + \_\ 2HCl \rightarrow \_\ H_2O + \_\ CO_2 + \_\ 2NaCl$$

Chemical equations are written and balanced in terms of ATOMS and MOLECULES

- While chemical equations are written in terms of ATOMS and MOLECULES, that's NOT how we often measure substances in lab!

- measurements are usually MASS (and sometimes VOLUME), NOT number of atoms or molecules!

Na_2CO_3 solid   HCl solution

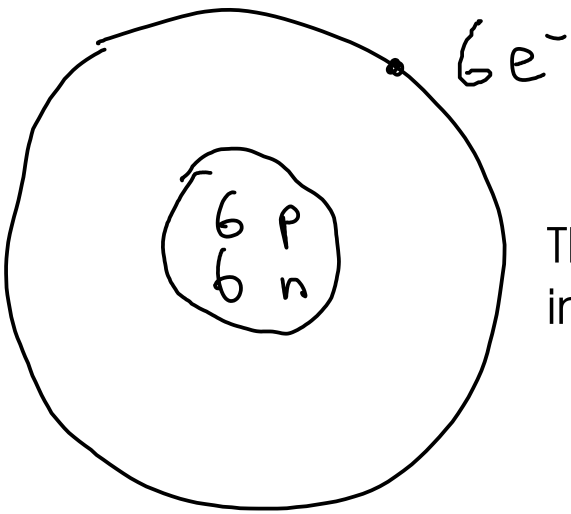
... so how do we relate atoms and molecules with things we routinely measure in lab - like grams and milliliters?

THE MOLE CONCEPT

- A "mole" of atoms is 6.022×10^{23} atoms

Why so big? Because atoms are so small!

- Why - in the metric dominated world of science - do we use such a strange number for quantity of atoms?



carbon-12

The mole is also defined as the number of carbon-12 atoms in exactly 12 g of carbon-12

THE MOLE CONCEPT

- Why define the mole based on an experimentally-measured number?

- The atomic weight of an element (if you put the number in front of the unit GRAMS) is equal to the mass of ONE MOLE of atoms of that element!

Carbon (C): Atomic mass 12.01 a~~m~~u → $\boxed{12.01\text{ g}}$

the mass of ONE MOLE of naturally-occurring carbon atoms

Magnesium (Mg): 24.31 g = the mass of ONE MOLE OF MAGNESIUM ATOMS

- So, using the MOLE, we can directly relate a mass and a certain number of atoms!

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

"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

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

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

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

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

WHAT ABOUT COMPOUNDS? FORMULA WEIGHT

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

(H_2O)

H_2O

$$H: 2 \times 1.008 = 2.016$$
$$O: 1 \times 16.00 = 16.00$$
$$18.016$$

FORMULA WEIGHT of water

$18.016\,g\ H_2O = mol\ H_2O$

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

First, let's find the FORMULA of the compound. We know the name, so ...

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

$\downarrow$

$(NH_4)_2CO_3$

N: 2 x 14.01
H: 8 x 1.008
C: 1 x 12.01
O: 3 x 16.00

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

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

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_4^+ NO_3^-

$\underline{NH_4NO_3}$:

$N: 2 \times 14.01 = 28.02$

$H: 4 \times 1.008 = 4.032$

$O: 3 \times 16.00 = 48.00$

$80.052\ g\ NH_4NO_3 = mol\ NH_4NO_3$

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

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

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

These percentages SHOULD sium to 100%, but you may see a small amount of "roundoff error" in the total.