Stoichiometry I

Refers to calculations using atoms, chemical formulas, and chemical equations.

Atoms form compounds and undergo chemical rxns in specific ratios.

For example, if we inspect the following chemical formulas:

Na_2SO_4

2 atoms Na

1 atom S

4 atoms O

MgO

1 atom Mg

1 atom O

CH_4

1 atom C

4 atoms H

Now inspect the following chemical equation:

$$C_{(s)} + O_{2(g)} \longrightarrow CO_{2(g)}$$

one atom C

2 atoms O

1 molecule CO_2

s = solid
l = liquid
g = gas
aq = aqueous (in water)

In all cases what the chemist is concerned with is the number of atoms, ratio of atoms, involved in the chemical formula or chemical equation.

But counting atoms individually is impossible because they (atoms) are to tiny to see and count. So we need a method that makes it convenient to count atoms.

We are going to use MASS to COUNT atoms. There are two parts to this: First we must learn about the MOLE.

Before we talk about the mole, I want to remind you of one of the ways we count objects everyday → the Dozen.

DOZEN

Everybody knows that 1 dozen means 12. we could be talking about 12 cars or 12 nails or 12 kids or 12 eggs. whatever the case, we understand that one dozen is 12 of whatever - its a number independent of mass.

We know 12 eggs weighs different than 12 cars - BUT the number is the same.

However, 12 sixteen-penny nails will have the same average mass as any other average sample of 12 sixteen-penny nails - so mass can be used to count nails of the same type.

In chemistry we need a unit bigger than the Dozen because atoms are tiny and there are lots of them.

<u>The Mole</u>: This is a quantity that refers to Avogadro's Number (N_A) which says that <u>1 mole contains 6.02×10^{23} particles.</u>

In Math terms: $N_A = \dfrac{6.02 \times 10^{23} \text{ particles}}{\text{mole}}$

<u>words</u>: 1 mole is 6.02×10^{23} particles per mole.

<u>particles can refer to</u>: atoms, units, molecules, formula units, apples, oranges, widgets, bolts, nails, people, whatever..

Now for the second part - <u>mass</u>

Let's turn to the periodic table:

6
C
12

we now will focus on the lower number: <u>Atomic Mass</u>

Atomic Mass is in <u>atomic mass units</u> or <u>amu</u>.

$1 \text{ amu} = 1.66 \times 10^{-24} \text{ g}$

Now since C is 12.01 amu, one carbon atom is $\left(\frac{12.01\,\cancel{amu}}{atom}\right)\left(\frac{1.66\times10^{-24}\,g}{\cancel{amu}}\right) =$

$$= \frac{1.99\times10^{-23}\,g}{atom}$$

Now remember that <u>1 mole</u> of <u>carbon</u> has <u>6.02×10^{23} atoms C</u>.

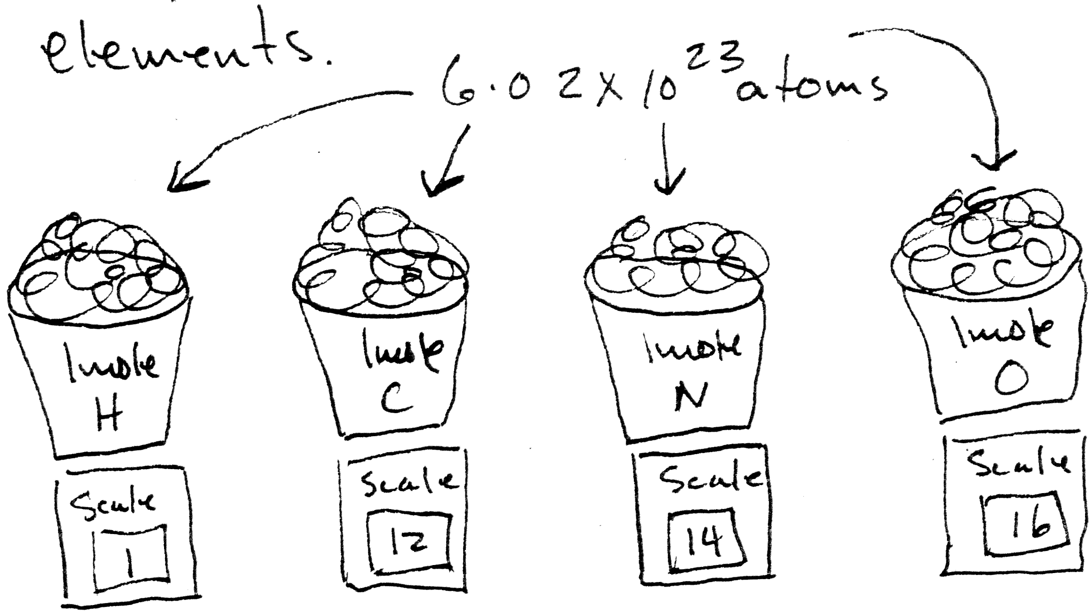
So the <u>mass of 1 mole</u> is:

$$\frac{1.99\times10^{-23}\,g}{\cancel{atom}} \times \frac{6.02\times10^{23}\,\cancel{atoms}\,C}{mole} =$$

$$= \frac{11.996\,g\,C}{mole} \text{ or } \frac{12\,g}{mole} \Rightarrow \text{ so if}$$

we use a scale and weigh out 12g of carbon (diamond or graphite) we would have exactly 6.02×10^{23} atoms C.

Comparison of C with other elements.



The moles of H, C, N, O have different masses because they each have different numbers of protons and neutrons.

Bottom line: we can use a scale to mass a given element and then using atomic mass (Molar Mass) and N_A we can know the number of atoms.

(A) How many atoms are in 10g carbon?

$$10\cancel{g}\ \cancel{Carbon} \times \frac{1\,mole}{12\cancel{g}\cancel{C}} = 0.83 \text{ mole } C$$

$$\frac{0.83\ \cancel{mole}\ C}{1} \times \frac{6.02 \times 10^{23} \text{ atoms } C}{\cancel{mole}} =$$

$$= 5 \times 10^{23} \text{ atoms } C.$$

(B) We can also work backwards from atoms to mass.

$$5.6 \times 10^{24}\ \cancel{atoms}\ N \times \frac{1\,mole}{6.02 \times 10^{23}\ \cancel{atoms}} = 9.30 \text{ mole } N$$

$$\cancel{9.6} \quad 9.30 \text{ mole } N \times \frac{14.01\,g\,N}{mole} = 130.3\,g\,N$$

Molar Mass of Compounds and Calculations using molar mass.

1. Molar Mass of Compounds

Sum the molar mass of all elements in the compound accounting for how many times a particular element appears in the compound.

Example

a) $$\text{Mm of } H_2O = \underset{\substack{\text{number}\\ \text{of}\\ H}}{2}\left(\overset{\substack{\text{molar mass}\\ H}}{\frac{1g}{mole}}\right) + \underset{\substack{\text{number}\\ \text{of}\\ O}}{1}\left(\overset{\substack{\text{molar mass}\\ O}}{\frac{16g}{mole}}\right) = \boxed{\frac{18g}{mole}}$$

b) $$\text{Mm of } CO_2 = \underset{\#C}{1}\left(\overset{\text{Mm C}}{\frac{12g}{mole}}\right) + \underset{\#O}{2}\left(\overset{\text{Mm O}}{\frac{16g}{mole}}\right) = \boxed{\frac{44g}{mole}}$$

② Calculations using molar mass

a. Finding formula units or molecules from mass:

Given 10g Na_2CO_3 find formula units:

$$10\,g\ \cancel{Na_2CO_3} \times \frac{1\ \cancel{mole}}{106\,g\ \cancel{Na_2CO_3}} \times \frac{6.022 \times 10^{23}\ \text{f. units}}{\cancel{mole}} =$$

$$= 5.68 \times 10^{22}\ \text{formula units,}$$

b. Find moles from mass:

$$21\,g\ NH_3 \times \frac{1\ \text{mole}}{17\,g\ NH_3} = 1.24\ \text{mole}\ NH_3$$

c. Find molecules from moles:

$$0.25\ \cancel{mole\ HC_2H_3O_2} \times \frac{6.022 \times 10^{23}\ \text{molecules}}{\cancel{mole}} =$$

$$= 1.5 \times 10^{23}\ \text{molecules}\ HC_2H_3O_2$$

d. Find mass from molecules or formula units:

$$2.4 \times 10^{22} \text{ molecules } H_2O \times \frac{1 \text{ mole}}{6.022 \times 10^{23} \text{ molecules}} \times \frac{18 g}{\text{mole}} =$$

$$= 0.72 g\ H_2O$$

3. Calculation of % Mass Composition

a. % Mass Composition: The percentage by mass of an element in a given compound.

$$\% \text{ element} = \frac{\left(\text{Molar mass}_{\text{atom}}\right)\left(\overset{\substack{\text{\# of times} \\ \text{it appears} \\ \downarrow}}{x}\right)}{\text{Molar Mass}_{\text{compound}}} \times 100$$

b. Example % mass composition

Calculate % mass of each element for: C_2H_6O

a. Find Mm compound = 46g/mole

b. Mm of each element accounting for numbers:

$C = 2(12) = 24$

$H = 6(1) = 6$

$O = 1(16) = 16$

c. % Mass is :

$$\%C = \frac{24}{46} \times 100 = 52.2$$

$$\%H = \frac{6}{46} \times 100 = 13$$

$$\%O = \frac{16}{46} \times 100 = 34.8$$

(4) Using % Composition to find Empirical Formula (EF).

Empirical Formula : The EF is the simplest chemical formula of a compound containing the lowest whole number ratio of elements.

Example

The compound glucose, $C_6H_{12}O_6$, has an $EF = CH_2O$. Each of the subscripts in $C_6H_{12}O_6$ is divisible by 6.

Before we proceed we must learn the rules for making calculations.

[redacted]

Rules for Making Chemistry Calculations

1. For calculations using molar or atomic masses, use all the decimals provided on the periodic table. No rounding allowed.
2. For any other calculation, use however many decimals provided. No rounding allowed.
3. As the calculation proceeds on the calculator, use ALL the numbers shown on the screen. No rounding allowed. If you must delete the screen between calculations, then you are required to write down all the numbers and then re-enter them.
4. After the whole string of calculations has been completed the answer will include THREE decimal places—whatever three numbers are on the right side of the decimal point. NO ROUNDING.
5. The final answer is circled.

(EF) Empirical Formula: A chemical formula of a compound in the lowest whole number ratio possible.

EF ionic compounds: The EF is the full chemical formula for an ionic compound. EF = molecular formula. eg. Na_2SO_4 = EF and molecular formula (MF).

EF of molecular compounds: The EF may not equal the molecular formula. Molecular formula = MF

Example EF

when 0.3546g sample of vanadium metal is heated in air, it reacts with O_2 to achieve a final mass of 0.6330g. Calculate the EF of this vanadium oxide.

① Mass O in compound is $0.6330 - 0.3546 =$

$$= 0.2784\,g\ O$$

② Next Find moles V, O:

$$\text{moles V} = 0.3546\,g \times \frac{1\,\text{mole}}{50.94\,g} = 0.006961$$

$$= 0.006961\ \text{moles V}$$

$$\text{moles O} = 0.2784\,g\ O \times \frac{1\,\text{mole}}{16.00\,g} = 0.01740\ \text{moles O}$$

③ Divide both values by smallest mole amount.

$$V = \frac{0.006961}{0.006961} = 1$$

$$O = \frac{0.01740}{0.006961} = 2.5$$

(4) Find a whole number to multiply through 1 and 2.5 to give only whole numbers:

$$V = 1 \times 2 = 2$$

$$O = 2.5 \times 2 = 5$$

EF is $\boxed{V_2O_5}$

Determining Empirical Formula from % Composition

Caffeine 49.48% carbon
5.15% hydrogen
28.87% nitrogen
16.49% oxygen

Molar Mass $\frac{194.2g}{mole}$

Assume 100g so:

$$49.48g\ C \times \frac{1mole}{12.01g} = 4.11990 mole\ C$$

$$5.15g\ H \times \frac{1mole}{1.008g} = 5.10912 mole\ H$$

$$28.87g\ N \times \frac{1mole}{14.01g} = 2.06067\ mole\ N$$

$$16.49g\ O \times \frac{1mole}{16g} = 1.03063\ mole\ O$$

Now divide by smallest mole value:

$$C \rightarrow \frac{4.11990}{1.03063} = 3.99746 \quad 4$$

$$H \rightarrow \frac{5.10912}{1.03063} = 4.95727 \quad 5$$

$$N \rightarrow \frac{2.06067}{1.03063} = 1.99942 \quad 2$$

$$O \rightarrow \quad = 1 \quad 1$$

$4(12.01) + 5(1.008) + 2(14.01) + 16 =$
$48.04 \quad 5.04 \quad 28.02$

$EF = C_4H_5N_2O \qquad EFMM = 97.1$

$MM = \frac{194.2g}{mole}$

Divide $\frac{MM}{EF_{MM}} = \frac{194.2}{97.1} = \underline{2}$

Now $2(C_4H_5N_2O) = C_8H_{10}N_4O_2$

Chem 7 Additional Notes on using mass percent.

$$\text{Mass percent} = \underbrace{\left[\frac{\text{total mass element}}{\text{mass compound}}\right]}_{\text{decimal }\%} \times 100$$

① Find mass % of C in CO_2:

$$\%C = \frac{12\ g/mole}{44\ g/mole} = \underbrace{\left[0.2727\right]}_{\text{decimal }\%}$$

$$\left[0.2727 \times 100\right] = 27.27\%$$

② Find the mass of carbon in 10g CO_2

$$\text{mass carbon} = 10g\ CO_2 \times 0.2727\left(\frac{g_C}{g_{CO_2}}\right) =$$

$$= 2.72g\ C \text{ in } 10g\ CO_2$$

(3) How many grams of CO_2 contain 2.72g carbon?

Two ways to solve this problem:

a) $2.72g\,\cancel{C} \times \frac{1\,mole\,C}{12g\,\cancel{C}} = 0.2266\,mole\,C$

$0.2266\,\cancel{mole\,C} \times \frac{1\,mole\,CO_2}{1\,\cancel{mole\,C}} = 0.2266\,mole\,CO_2$

$0.2266\,mole\,CO_2 \times \frac{44g\,CO_2}{mole} = 9.97g\,CO_2 = \tilde{=} 10g\,CO_2$

b) $\boxed{\text{Decimal } \% = \left(\frac{g\,c}{g\,compound}\right)}\left(g\,compound\right) = \text{grams carbon}$

$\underset{Dec.\,\%}{(0.2727)}\left(\text{grams compound}\right) = \text{grams carbon}$

So

$\rightarrow 0.2727\,(g\,compound) = 2.72$

$\text{grams compound} = \frac{2.72}{0.2727} = \tilde{}\,10g$ Compoun

Review % mass and Empirical Formula

(A) % mass or % composition

1. In words: % mass is the fraction of mass represented by an element within its compound.

2. Formula:

molar mass element; # of times element appears in formula

$$\% \text{ mass} = \frac{(Mm)(n)}{Mm \text{ compound}} \times 100$$

molar mass compound

3. Example

Calc % mass C in $CHCl_3$.

$$\% = \frac{(12 \text{ g/mole})}{119.35 \text{ g/mole}} \times 100 = 10.05\%$$

4. what 10.05% C means. This means that in any given mass of $CHCl_3$, 10% of that mass is represented by carbon.

5. The general formula:

$$\% \text{ mass} = \frac{\text{g element}}{\text{g compound}} \times 100$$

Ⓑ Empirical Formula (EF)

A chemical formula that represents the lowest whole number ratio of the elements within that compound.

1. Ionic compounds: molecular formula = EF

2. Molecular Compounds: molecular formula ≠ EF

Examples

a) $MgSO_4 = MgSO_4$

molecular ↑ ; EF ↓

b) $C_6H_{12}O_6 \neq CH_2O$

molecular ↑ ; EF ↓

Ⓒ Finding the EF given a formula:

Divide the subscripts in the chemical formula by a common divisional so as to reduce the formula to the <u>lowest whole number ratio</u>. $C_6H_{12}O_6$ was divided by 6.

(c) continued

example

$$C_8H_{12}O_6 \rightarrow C_{\frac{8}{2}}H_{\frac{12}{2}}O_{\frac{6}{2}} = C_4H_6O_3$$

↑ molecular

EF

(d) Finding EF when given % mass of the elements in the compound.

1. Given %'s of each element, convert to grams.
2. Convert grams to moles
3. Divide by smallest value - this usually gives the correct whole #'s.

Example

A compound is 59.9% Ti and 40.1% O.

Don't round

$$59.9\%\,Ti \rightarrow 59.9g\,Ti \times \frac{1\,mole}{47.9g} = \frac{1.25052192\,moles\,Ti}{1.2505\ldots} = 1$$

$$40.1\%\,O \rightarrow 40.1g\,O \times \frac{1\,mole}{16g} = \frac{2.50625\,moles\,O}{1.2505\ldots} = 2$$

$\rightarrow TiO_2$

(e) Finding EF when given mass of each element: Convert directly to moles.

Example

A compound contains 2.514g Ca and 1.004g O.

$$2.514g\ Ca \times \frac{1\,mole}{40.08g} = \frac{0.0627245509\,mole\ Ca}{0.06272\ldots} = 1$$

$$1.004g\ O \times \frac{1\,mole}{16g} = \frac{0.06275\,mole\ O}{0.06272\ldots} = 1$$

So CaO

(f) Finding EF when given moles:

Example

A compound contains 0.4285506 mole Ba, 0.42843779 mole S, and 1.714375 mole O.

$$\frac{0.4285506\,mole\ Ba}{0.4284\ldots} = 1$$

$$\frac{0.42843779\,mole\ S}{0.4284\ldots} = 1$$

$$\frac{1.714375\,mole\ O}{0.4284\ldots} = 4$$

$BaSO_4$

(c) Doing calculations using % mass.

Basic Formula: (1) $\% \text{ mass} = \dfrac{g \text{ element}}{g \text{ compound}} \times 100$

(2) $\text{decimal } \% = \dfrac{g \text{ element}}{g \text{ compound}}$

Example Problem

Find the mass of Ti in 10g TiO_2

Step 1

$$\text{Dec. } \% \text{ Ti} = \frac{47.87}{79.87} = 0.5993$$

Step 2

g's Ti in 10g TiO_2:

$$\underset{\substack{\uparrow \\ \text{Dec. } \%}}{(0.5993)}\underset{\substack{\uparrow \\ \text{mass compound}}}{(10g\ TiO_2)} = 5.993g \text{ Ti in } 10g\ TiO_2$$