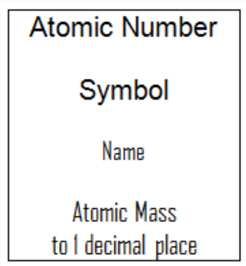
1.2 – Formulas

1.2.1 - Define the terms relative atomic mass (A_r) and relative molecular mass (M_r)

Relative Atomic Mass (A_r) - The weighted mean of the masses of the naturally occurring isotopes, on a scale in which the mass of an atom of the carbon-12 isotope is 12 units exactly.

Relative Molecular Mass (M_r) - The sum of the relative atomic masses of the constituent elements, as given in the molecular formula.

Relative Formula Mass - The same as the Relative Molecular Mass, but it only applies to ionic compounds.



1.2.2 - Calculate the mass of one mole of a species form its formula

The mass of one mole of any substance is equivalent in to its relative atomic (or molecular, or formula) mass, measured in grams.

The mass of one mole of any substance is called the molar mass (M), where the molar mass is equal to the relative atomic mass in units of grams per mole (g mol-1)

For example, if we wanted to find the mass of one mole of HNO_3, we would:

$$
\begin{aligned}
M_r(HNO_3) &= A_r(H) + A_r(N) + 3A_r(O) \\
&= 1.01 + 14.01 + 3 \times 16.00 \\
&= 63.02
\end{aligned}
$$

$\therefore$ one mole of HNO_3 = 63.02 g

1.2.3 - Solve problems involving the relationship between the amount of substance in moles, mass and molar mass

$$n = \frac{m}{M}$$

n = number of moles
m = mass (g)
M = molar mass ($g\ mol^{-1}$)

For example, if we wanted to find the number of moles there are in 127.1g of copper atoms, we do:

$$127.1\,\cancel{g\ Cu} \times \frac{1\ mol\ Cu}{63.55\,\cancel{g\ Cu}} = 2.000\,mol\ Cu$$

To find the mass of 100 atoms of copper, we do:

$$100\,\cancel{atoms\ Cu} \times \frac{1\,\cancel{mol}}{6.02\times10^{23}\,\cancel{atoms}} \times \frac{63.55\,g\ Cu}{1\,\cancel{mol}} = 1.06\times10^{-20}\,g\ Cu$$

1.2.4 - Distinguish between the terms empirical formula and molecular formula

Empirical Formula - The simplest whole number ratio of atoms of different elements in the compound

Molecular Formula - The actual number of atoms of different elements covalently bonded in a molecule

1.2.5 - Determine the empirical formula from the percentage composition or from other experimental data

There a few simple steps to determining the empirical formula:

1. Write the elements present in the compound as a ratio
2. Write the percentage compostion below each element
3. Divide the percentage by the relative atomic mass of the element, giving a molar ratio
4. Divide each ratio by the smallest one to get a whole-number ratio
5. Express the ratios in the empirical formula

1. H : O
2. 11.2% : 88.8%
3. $\div 1.01$ $\div 16.00$
   11.1 : 5.55
4. $\div 5.55$ $\div 5.55$
   = 2 = 1
5. H_2O

To determine the percentage composition by mass of a compound, given its formula, divide the relative atomic mass of each element present by the relative atomic mass of the compound and express the result as a percentage.

$\boxed{H_2O}$

$$\% H_2 = \frac{H_2}{H_2O} = \frac{2.02}{18.02} = 11.2\%$$

$$\% O = \frac{O}{H_2O} = \frac{16.00}{18.02} = 88.8\%$$

100.0%

Remember to add all your percentages together at the end to identify any mistakes.

1.2.6 - Determine the molecular formula when given both the empirical formula and experimental data

To determine the **molecular formula**, we must first know the **empirical formula** and the molar mass of the compound. The molar mass of the empirical formula will be in direct proportion to the molar mass of the molecular formula.

For example, if we knew that the empirical formula of a hydrocarbon was CH and its molar mass was 26.04g mol-1, to determine its molecular formula, we would:

$$M(CH) = 13.02\,g\,mol^{-1} \qquad M(\text{compound}) = 26.04\,g\,mol^{-1}$$

$$\frac{M(\text{compound})}{M(\text{empirical formula})} = \frac{26.04}{13.02} = 2$$

So, the molecular formula is 2 times the empirical formula. Therefore, it is: C_2H_2.