A First Year blog on Calculations in Chemistry: Part 2, 2nd March 2019

(This is a **revision** blog. It is merely a *summary* of some of the things that you need to know. *Everything in this blog can be found in Chapters 38-41 of the First Year, Foundation Chemistry Book*.)

Last week I told you that reaction equations such as

$$2H_2 + O_2 \longrightarrow 2H_2O \qquad \text{(cf. footnote[1])}$$

do two things i.e.

a) they state what the reactants and the products are, and
b) they specify the **proportions** in which the reactants react to produce the product (or products).

The equation is thus saying that if you react

2 MOLECULES of Hydrogen + 1 MOLECULE of Oxygen you will get 2 MOLECULES of Water

and this means that it is also saying that if you react

2 MOLES of Hydrogen molecules + 1 MOLE of Oxygen molecules you get 2 MOLES of Water molecules.

I also explained to you that the number of units of small entities involved (i.e. atoms/molecules/ions) **in any and in EVERY** chemical reaction is absolutely massive, and this would make the mathematics involved for even simple calculations totally and utterly unwieldy. For example, there are

334,000,000,000,000,000,000,000

or a more accurate answer would be $3.342,796,149 \times 10^{23}$ molecules of water in just one 10 ml / 10 cm^3 teaspoonful of water, and this would make simple calculations for the amount of energy involved in a reaction / the concentrations involved / the masses of the entities involved / etc extraordinarily unwieldy. However, at a stroke, the whole problem of trying to work with very large numbers of units becomes resolved by choosing "**one mole**" as the counting unit in Chemistry.

One "mole" consists of exactly the same number of Carbon atoms as there are in 12 grams of Carbon-12 (this number being very roughly 6.0×10^{23}).

Nobody knows *exactly* how many units there are in one mole (but it is roughly $6.0\text{x}10^{23}$), but whatever that number may be **it is exactly as many atoms of Carbon-12 as there are in 12 grams of Carbon-12**.

That is the counting unit in Chemistry.

We cannot easily manipulate numbers such as

$$\mathbf{6.022,140,760 \times 10^{23}}$$

or even **600,000,000,000,000,000,000,000** $(= 6.0 \times 10^{23})$

but it is easy to manipulate the number "1" (as in "one mole").

Last week I told you about the "mole", and this week I am going to tell you about RAM/RFM/and RMM.

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[1] Strictly speaking the equation should be written as a reversible reaction, but let us not worry about that for the moment.

A) Relative Atomic Mass (RAM or A_r)

Since the mass of 1 atom of Carbon-12 in effect consists of 12 units i.e. **6 protons + 6 neutrons**[2] , and *one-twelfth* of 12 equals "1" – it is then the convention in Chemistry to express the Relative Atomic Mass of an element as **the ratio of the mass of one atom of the element divided by ONE-TWELFTH of the mass of one atom of Carbon.** The RELATIVE ATOMIC MASS (RAM or A_r) of an element is thus

$$\textbf{RAM or } A_r = \frac{\textbf{Mass of 1 ATOM of the element in grams (g)}}{\textbf{One-twelfth of the mass of 1 ATOM of Carbon-12 in grams (g)}}$$

The "grams" will cancel each other out, and RAM or A_r therefore has NO UNITS![3]

However, the RAM of an element can also just as easily be expressed in terms of **1 mole** of its atoms i.e.

$$\textbf{RAM or } A_r = \frac{\textbf{Mass of 1 mole of ATOMS of the element in grams (g)}}{\textbf{One-twelfth the mass of 1 mole of Carbon-12 ATOMS in grams (i.e. 1g)}}$$

NB One mole of Carbon-12 atoms has a mass (for 'A' Level purposes) of 12 grams (or 12g), therefore **one-twelfth of the mass of one mole of Carbon-12 atoms = 1g**. When you take into account the electrons and the different isotopes of Carbon, the mass of 1 mole of isotopic Carbon will be more than 1g.

The RAM of an element has NO UNITS (it is unitless/dimensionless)

Relative Mass is merely a ratio! Relative Mass has **NO UNITS** because the units in both the Numerator and in the Denominator (e.g. grams) cancel each other out!

I will say it again, Relative Mass (i.e. Relative Atomic Mass/Relative Molecular Mass/Relative Formula Mass is merely a ratio and

RELATIVE MASS HAS NO UNITS!

The concept of " g mol^{-1} "

RAM/RMM/RFM all have NO UNITS, but 1 atom has a mass, and 1 mole of atoms also has a mass – and you will thus often see RAM expressed in the following manner

"RAM of Ne = 20.2 g mol^{-1}" or "RAM of Ne = 20.2 grams per mole"
and what this means is that **"1 mole of atoms of Neon has a mass of 20.2g"**,
and the RAM of any element can therefore be written in the following manner

RAM of Ne = 20.2 g mol^{-1} viz. "1 mole of Neon atoms has a mass of 20.2g"
RAM of He = 4.0 g mol^{-1} viz. "1 mole of Helium atoms has a mass of 4.0g"
RAM of Kr = 83.8 g mol^{-1} viz. "1 mole of Krypton atoms has a mass of 83.8g"
.... and so on.

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[2] The mass of 6 electrons being small by comparison.
[3] The number at the top left hand corner of the symbol for an element in a Periodic Table is that of the Atomic Mass of the element and it has the same numeric value as **RAM** or **A_r** – thus for example, the RAM of Hydrogen is "1", and the RAM of Carbon-12 is "12", and the RAM of Nitrogen-14 is "14", and so on.

We have as yet not gone into how Isotopes affect Relative Atomic Mass (RAM), but even so I should like you now to learn the following definition by heart (and then I will tell you about Isotopes and RAM).

The Relative Atomic Mass of an element is the arithmetically weighted average mass of one atom of the commonly occurring isotopes of the given element relative to 1/12th of the mass of one atom of Carbon-12.

Learn this definition off by heart and if they ask it in the First Year exams (they will not ask it in the Second Year exams), you've straight away got one mark – and it will have taken you just 30 seconds to get that mark. OK, now let us see why some elements have RAMs that are not whole numbers.

Isotopes/Atomic Mass/Relative Atomic Mass

You would not be doing 'A' Level subjects such as Chemistry/Physics/Maths/Biology/etc if you were not intellectually reasonably bright (or perhaps even *very* bright) – therefore I am sure that it has not escaped your notice that some of the Atomic Mass numbers in the Periodic Table look wrong because they are not whole numbers (i.e. they are not *integers*), and instead they are decimal numbers – and you will have said to yourself *"you cannot have a part of an atom (after all, it is either an atom or it is not an atom!) therefore how on earth can you have a decimal number for a RAM"*!

In fact, these decimal numbers **are** correct because on Earth many elements exist in different forms of isotopes. [You will remember that isotopes of an element contain the same number of protons but differing numbers of neutrons, therefore the atoms of isotopes of the same element will have differing masses.] If therefore an element on earth has two or more isotopic forms, then this would normally give rise to non-integer Atomic Masses (and thus to non-whole number RAMs). Let us use Chlorine to see why this is the case.

On earth, there are two isotopes of Chlorine, i.e. Chlorine-35 (Cl-35) and Chlorine-37 (Cl-37), and if they existed in equal quantities, then the average RAM of 1 atom of Chlorine would be (35+37)÷2 = 36; and if there were twice as much Cl-35 as there were Cl-37 then the average RAM of 1 atom of Chlorine would be [(2x35)+(1x37)] ÷ 3 = 35.7.

As it happens, 75.53% of all Chlorine atoms are Cl-35, and 24.47% are Cl-37, therefore the weighted average of the naturally occurring atoms of Chlorine is [(75.53 x 35) + (24.47 x 37)] ÷ 100 = 35.4894 or approximately **35.5**.

Many elements (e.g. H/O/N/C/etc) under normal circumstances have only one type of atom therefore their Atomic Masses are whole numbers (i.e. they are "integers"). However, those elements that have two or more naturally occurring isotopes will normally have RAMs that are not integers (because their RAMs are the **weighted average** of the naturally occurring isotopes of the element).

NB **It is not possible to have a part of an atom**, therefore (with the numbers that we use at 'A' Level) it follows that any element that has a RAM that is not a whole number **MUST** have more than one isotope on earth (and the RAM in the Periodic Table will be the weighted average of the RAMs of its isotopes).

Relative Abundance and Weighted Averages

If on earth 90.9% of all Neon atoms have an Atomic Mass of 20, and 8.8% of all Neon atoms have an Atomic Mass of 22, and 0.26% have an Atomic Mass of 21, then the percentages "90.9%", "8.8%" and "0.26%" represent the "**Relative Abundances**" of Neon on earth. An average that takes into account the relative abundance of the things that are being averaged is called a ***weighted*** average.

The *"simple"* average of 35 and 37 is 36 [i.e. (35 + 37) ÷ 2 = 36], but if the Relative abundance of the Cl-35 isotope is 75.53%, and the Relative Abundance of the Cl-37 isotope is 24.47%, then (as we saw on page 3) the ***weighted*** average RAM of the two forms is **35.5** (to one decimal point).[4]

OK, now that we know how isotopes enter into the calculation, we can write RAM or A_r as

$$\textbf{RAM or } A_r = \frac{\text{Mass of 1 isotopic weighted average atom of the element in g}}{\text{One-twelfth of the mass of one C-12 atom in g}}$$

or

$$\textbf{RAM or } A_r = \frac{\text{Mass of 1 MOLE of the isotopic wtd. Av. Atoms of the element in g}}{\text{One-twelfth of the mass of 1 MOLE of C-12 atoms in g (i.e. 1g)}}$$

You will almost certainly be given a RAM calculation to do in your First Year exams, therefore learn the *technique* for calculating weighted averages and use that technique to do the question in the exam. For example, go back to the Chlorine example on page 3 and learn the technique/the methodology for calculating a RAM where there are *two* isotopes involved. I shall now show you how to do the calculation for the weighted average RAM for an element that has *three* isotopes[5]. OK, let us do the calculation for Neon which has three stable isotopes on earth i.e. Ne-20, Ne-21, and Ne-22.

ISOTOPE	RELATIVE ABUNDANCE (%)
Neon-20	90.9
Neon-21	0.26
Neon-22	8.80

$$\begin{aligned}\text{RAM or } A_r &= \frac{[90.9 \times 20] + [0.26 \times 21] + [8.80 \times 22]}{100} \\ &= \frac{1818 + 5.46 + 193.6}{100} \\ &= 20.17\end{aligned}$$

and, if you refer to a data book, you will find that the RAM of Neon (when calculated to 5 significant figures) is given as 20.179 – but for 'A' Level purposes, 3 significant figures is appropriate, so you will find the RAM of Neon on the data sheet in the exam is given as 20.2.

We have talked fairly extensively about the Mass of an object in relation to the mass of Carbon-12, and we can now develop that concept to allow us to talk about the absolute Mass of an object.

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[4] The Relative abundances that I used were to 4 significant figures therefore I should really have given my answer to 4 significant figures i.e. 35.49. If you want to know about "significant figures" (and you will get 1 extra mark in your Practical Exam/Coursework if you give your answers to the correct number of significant figures), then please refer to *"Maths For Chemistry"* in the Appendix to this Chapter.

[5] The examiners HAVE in the past set a RAM question for an element with three isotopes, and in fact there is absolutely no reason whatsoever why they should not ask you to do one for an element with even four isotopes!

B) Atomic Mass

Chemists have designed Periodic Tables so that they show a list of all the elements known to man[6], and in a Periodic Table, in each box there is the Chemical symbol of an element, and there are two numbers. Ignoring decimal numbers caused by isotopes, the smaller of these two numbers states the **Proton Number** (or the **Atomic Number**) of the element (i.e. how many protons there are in one atom of that element), and in earlier Chapters we learnt that the larger of the two numbers, i.e. the **Atomic Mass** states the number of (Protons + Neutrons) in an atom of that element.

If the yardstick for the measurement of the mass of atoms is roughly speaking *the number of times that the mass of that atom is larger than one-twelfth the mass of 1 atom of Carbon*, then the **Atomic Mass** of an element is also numerically the **Relative Atomic Mass** of the element!

We have therefore derived a new relationship[7] viz. that the numerical value of the Relative Atomic Mass of an element is the same as the numerical value of the Atomic Mass of that element![8]

You will remember that two numbers (i.e. Atomic Mass and Atomic Number) are written

Atomic Mass[9] = (No of protons + No of neutrons) —> A

X <— symbol for the element

Atomic Number or Proton Number —> Z

We have now seen that **Relative Atomic Mass** (RAM) is *numerically* the same as **Atomic Mass**, and the Mass of 1 mole of any element can thus be calculated by multiplying its RAM by 1g

i.e. Mass = RAM x 1g

In other words, if you are asked to calculate the MASS of 1 mole of any element, then all that you need to do is to look up its Atomic Mass in a Periodic Table and multiply it by 1g and that will be the numerical value of the Mass of 1 mole of that element!

The Unit for Mass : The mass of 1 mole of the atoms of an element can however be expressed in whichever units the observer may so desire e.g. in pounds/in kilos/in ounces/in tonnes/or whatever; ***but, it is conventional to express Atomic Mass in grams.***

NB There are therefore no calculations involved in an Atomic Mass. You just look up the relevant number for an element in your Periodic Table and that is its Atomic mass in grams!

It is difficult to pitch explanations between being too detailed and being not detailed enough. What I have tried to do is not to get you involved in stuff that is First Year Degree Level but to give you explanations that are more informative than GCSE stuff.

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[6] I am sure that you remember that an "element" is a substance that is composed of only ONE type of atoms (i.e. atoms that all contain the same number of protons).
[7] ... but only for 'A' Level purposes, and NOT for when you get to University!
[8] **Don't you just LOVE the feeling of intellectual satisfaction that comes when you use your brain to derive a new truth from stuff that you already know!** Whenever I do it, I find the experience is just sheer magic!
[9] And also Relative Atomic Mass!

OK, let us now see whether we can do the calculation for the number of molecules of H_2O that there are in one 10 ml / 10 cm^3 teaspoonful of Water.

By definition,	1 litre of Water has a mass of	1000g, and
since the RAM for H= 1 and O = 16	1 mole of Water has a mass of	18.0g [(2x1)+(1x16)]
or if you use the more accurate figure	1 mole of Water has a mass of	18.01528g
therefore	X moles of Water will have a mass of	1000g

Where $X = \frac{1 \text{ mol} \times 1000\text{g}}{18.01528\text{g}} = 55.50843506 \text{ mol}$

But, there are	$6.022{,}140{,}760 \times 10^{23}$ units in every mole of anything/of everything
therefore there are	$6.022{,}140{,}760 \times 10^{23}$ x 55.50843506 molecules in 1 litre/1 dm^3 of Water
or there are	$3.342796093 \times 10^{25}$ molecules of Water in 1 litre /1 dm^3 / 1000 cm^3 of Water.

Therefore, there are $3.342796093 \times 10^{23}$ molecules of Water in 0.01 litre or 10 cm^3 of Water.

and $3.342796093 \times 10^{23}$ = 334,279,609,300,000,000,000,000 .

As you can now see very clearly, to do calculations in Chemistry you could not use absolute numbers. The calculations would be just too unwieldy.

You have to use MOLES, and
you have to use RAMs / RMMs / RFMs
and you must be numerate. You do not need the complicated bits of 'A' Level Maths such as Integrals and Derivatives – but you must be able to handle numbers competently and confidently.