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Chemistry Stoichiometry Notes

Before starting to learn about stoichiometry, think about this little analogy.

Let's say I have two cubes, one being 100 grams of pure Carbon, and another being 100 grams of Aluminium.

Now, are there the same number of atoms within both of the cubes?

No, because Carbon, being a lighter element, will require more atoms to maintain a specific weight than a metal such as Aluminium.

This is where **moles** come in. Essentially, moles are an intermediary unit to describe the relationship between molar mass (which is basically a measurement of relative atomic mass between elements) and mass (in grams).

Now, if you have a periodic table, you might notice some numbers under elements. This is the **relative atomic mass** of an element (this number is based on the average natural abundance of this element which is analysed using mass spectrometry (you don't need this)).

We could measure everything in grams, but then all our calculations would use *really* small numbers, that would be overly complicated to use for calculations. (e.g. Carbon-12 weighs 1.9926×10^{-23}g)

Instead, we use **relative atomic mass** to compare the mass of an element with the mass of **one twelfth the mass of a Carbon-12 atom**.

Thus, for example, we would describe Chlorine's *relative atomic mass* as 35.45, as the average atom of Chlorine is 35.45 times the mass of one twelfth the mass of a Carbon-12 atom.

Similarly, we have another unit called **molar mass** (M). It is defined as:

> The mass of one mole of an element's atoms.

But what is a mole?

As said before, a mole is a unit we use to connect number of atoms to grams.

Essentially, a mole of **any element** has a mass that is equal to its **relative atomic mass**, in grams.

For example, if I had one mole of Carbon, I would have exactly 12.01 grams of Carbon. (its relative atomic mass is 12.01)

> Note: Similarly, One mole of Carbon-12 is 12 grams.

That's cool and all, but surely there has to be some universal constant to represent this.

Well, **relative atomic mass** is based on **one twelfth the mass of a Carbon-12 atom**. So naturally, we need some equivalent in describing a mole.

This is where **Avogadro's Number** comes in.

Avogadro's Number is the exact quantity of Carbon-12 atoms to have a mass of 12 grams.

In general,

> Avogadro's number represents the number of an element's atoms required to have a mass of that element's relative atomic mass.

Having this quantity of an element is equivalent to a **mole** of that element.

Therefore, given that Molar Mass is the **mass of one mole of its atoms**, and a **mole** is related to **relative atomic mass**, unsurprisingly **Molar Mass** and **Relative atomic mass** are **exactly the same!**

Except that molar mass is measured in grams per mole, whereas Relative atomic mass does not have a unit of measurement, as its *relative*.

Thus, if I had 22.99 grams of Sodium, which is numerically equivalent to the Molar mass of Sodium (and Sodium's relative atomic mass), I would also have 1 mole of Sodium, which has Avogadro's number of atoms.

Now, here's a quick question:

What's the molar mass of oxygen *gas*?

Well, we know that oxygen gas contains two atoms of oxygen. We also know that molar mass is exactly the same as Relative atomic mass.

The Relative Atomic Mass of Oxygen is 16.00, but this is for a single atom. We're looking for the relative atomic mass of 2 atoms, so we just **times by 2**.

Therefore, the Molar mass of Oxygen is 32.00 $g\,mol^{-1}$. *Remember to write $g\,mol^{-1}$!*

Here's a harder question:

What's the molar mass of Sulphuric acid?

First, let's identify the following:

1. What elements are we looking for
2. What ratio are these elements present in

Sulphuric acid's formula is H_2SO_4.

Thus, there are:

- 2 Hydrogen atoms
- 1 Sulfur atom
- 4 Oxygen atoms.

Now, we can write each of the element's Molar mass (i.e. its relative atomic mass (i use relative atomic mass cause its alot easier)) and like before, multiply it by how many atoms there are of the element.

$2 \times 1.008 + 1 \times 32.07 + 4 \times 16.00$

= 98.086 $g\,mol^{-1}$.

Now that's done, you're ready for one of the most easiest formulas in the world.

> **Info**
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> $n = \frac{m}{M}$

where:

- n = number of moles, measured in moles (mol)
- m = mass, measured in grams (g)
- M = molar mass, measured in grams per mole ($g\,mol^{-1}$)

This formula will covert grams to moles, and moles to grams, so you can have the best of both worlds (i.e. a unit of measurement used everywhere and a unit of measurement that can be converted to number of atoms easily)

Now, example question:

What is the number of atoms in 50 grams of Silver Ethanoate ($CH_3COOHAg$)(btw Ms Pilling said for ethanoates you put the cation at the end. Something to do with intermolecular bonding)?

First, there are:

- 2 Carbon atoms (12.01)
- 4 Hydrogen atoms (1.008)
- 2 Oxygen atoms (16.00)
- 1 Silver atom (107.868)

Therefore, the molar mass of Silver Ethanoate is:

$2 \times 12.01 + 4 \times 1.008 + 2 \times 16.00 + 1 \times 107.868$

Which is 167.92 $g\,mol^{-1}$.

But we're not done yet! Now we can use our amazing formula:

$n = \frac{m}{M}, n = \frac{50}{167.92}$

Thus, the number of moles in 50g of Silver Ethanoate is 0.29776 **mol**.

Therefore, since a mole represents 6.022×10^{23} atoms,

Then numbe of atoms in 50g of Silver Ethanoate is:

$0.29776 \times 6.022 \times 10^{23}$

Which is: **1.79312×10^{23} atoms!

Ok, now you understand basic stoichiometry, and moles to grams conversions. But what about moles inside of moles?

Let's say I had 2 moles of Sulphuric acid (H_2SO_4). How many moles of Hydrogen atoms are contained inside?

Well, the answer is surprisingly very simple. We look at the **subscript** of the element we're looking for. In this case, Hydrogen's subscript is 2, so we simply **multiply** the moles of the **whole molecule** by the subscript.

Thus, the number of Hydrogen moles is **4**.

Similarly the number of Sulphate ion (SO_4^{-2}) moles is **2**. (There is only one molecule of Sulphate) (Imagine H_2SO_4 as $H_2(SO_4)_1$)

Also, the number of Oxygen moles is 8. (Oxygen has a subscript of 4.)

Ok, time to throw you in the deep end.

I have 5 kilograms of Nitric Acid ($HNO_{3(aq)}$) and I mixed it with an excess amount of Sodium Hydroxide ($NaOH_{aq}$).

Find all the products of this reaction, and find how many moles of Sodium Nitrate are produced (hint hint sodium nitrate is one of the products)

First, we write the equation:

$$HNO_{3(aq)} + NaOH_s \rightarrow NaNO_{3(aq)} + H_2O_{(l)}$$

Note that the ratio between all the reactants and products is: **1 : 1 : 1 :1**

Now here's the thing about moles within chemical formulas:

Let's say I had 2 substances in my chemical reaction. They could be **any** of the substances, reactants or products.

Let's call them aA, and bB, where the small letters are the coefficients and the big elements are the substances.

(In our case, we'll choose Nitric Acid and Sodium Nitrate.)

Now, here's another amazingly simple formula:

For two **substances** in an equation (aA, bB),

> **Info**
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> $n(A) = \frac{a}{b} \times n(B)$

Thus,

$n(HNO_{3(aq)}) = \frac{1}{1} \times n(NaNO_{39(aq)})$

$\therefore n(HNO_{3(aq)}) = n(NaNO_{3(aq)})$

Ok, now we need to actually calculate the number of moles of Nitric Acid:

$n = \frac{m}{m}, n = \frac{5000}{1 \times 1.008 + 1 \times 14.01 + 3 \times 16.00}$

$\therefore n(HNO_{3(aq)}) = 79.342$ mol.

$\therefore n(NaNO_{3(aq)}) =$ 79.232 mol.

Great! You understand the basics of stoichiometry now.

Here are a few practice questions:

1. For a science experiment, I burn 4 moles of ethanol ($C_2H_6O_{(l)}$) in an open flame. How much moles of oxygen is consumed, assuming complete combustion (yea there are other types of combustion .-.)?
2. Pure oxygen is prepared for laboratory use through the decomposition of hydrogen peroxide (sidenote, its catalytic) in the following formula: $2H_2O_{2(aq)} \rightarrow O_{2(g)} + 2H_2O_{(l)}$ If I needed 20 grams of oxygen gas for an experiment, how many moles of hydrogen peroxide would I need?
3. Given the formula: $Na_2CO_3.10H_20_{(aq)} + 2HCl_{(aq)} \rightarrow CO_{2(g)} + 11H_2O_{(l)} + 2Na$

How many **grams** of Sodium Carbonate-10-water produces 60 grams of Sodium Chloride?

Answers: (don't look if you actually want to do it yourself, otherwise I guess you can see how I would do it)

1. First, write the chemical equation:

$$2C_2H_6OH_{(l)} + 6O_{2(g)} \rightarrow 4CO_{2(g)} + 6H_2O_{(l)}$$

Note the 2 : 6 : 4 : 6 ratio.

I'm comparing ethanol and oxygen. Therefore, I'm looking at a 2 to 6 ratio.

Therefore, $n(O_{2(g)}) = \frac{6}{2} \times n(C_2H_6OH_{(l)})$

= $\frac{6}{2} \times 4$ = 12 mol.

2. First, we're in grams so I'm going to convert into moles (in general you should always convert into moles, at least you'll get a mark for that)

$n = \frac{m}{m}, n(O_{2(g)}) = \frac{20}{2 \times 16} = 0.625$ mol.

Therefore,

$n(H_2O_{2(aq)}) = \frac{2}{1} \times n(O_{2(g)})$

= 1.25 mol.

3. Convert grams to moles.

$n = \frac{m}{M}, n(NaCl_{(aq)}) = \frac{60}{22.99+35.45} = 1.0267$ mol.

Therefore,

$n(Na_2CO_3.10H_20_{(aq)}) = \frac{1}{2} \times n(NaCl_{aq}) = 0.51335$ mol.

However, we're looking for **grams**. Thus, we have to reconvert from moles to grams.

$n = \frac{m}{M}, m = nM, m(Na_2CO_3.10H_20_{(aq)}) = 0.51335 \times (2 \times 22.9$
(preferably its not on one giant line)

= 146.89 **g**!

Great! Now you have a relatively good understanding of basic stoichiometry.

Now, time for **gasses.**

Remember that laboratory question about oxygen? Perhaps you noticed something wrong about measuring oxygen in grams. You would be correct, as in general, we do not have oxygen in solid form, which is what we use grams for.

Instead, since oxygen is a gas, it occupy **volume**.

Instead of grams, we use **litres**.

Sounds logical, right? All we need to do is change the grams to litres, change some numbers around and everything will be fine, right?

No.

Gasses are influenced by both **temperature** and **pressure**.

If I increase the temperature of a gas, following the kinetic theory of gas, then on average gas particles will have more kinetic energy, making them spread out and thus want to occupy more space (if they don't get this space, pressure increases but you don't need to know that).

Similarly, increasing pressure will reduce the volume of a gas. Think about it as you're crushing gas particles in some sort of compressor. Adding pressure will push the particles closer.

We measure pressure in pascals (Pa).

For example, average atmospheric pressure at sea level is described as 101.325 kPa, which is equivalent to one standard atmosphere (atm) of pressure.

We measure temperature in kelvin (K) (actually we usually use degrees Celcius).

What makes kelvin useful is the concept of absolute zero.

Here's a side-note about absolute zero you can read if you're really interested:

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The Kinetic theory of gasses states that all particles usually have some form of motion. This motion accounts for the 'heat' of the particles, e.g. in high temperatures particles would move faster than if they were at lower temperatures, assume the particles remain the same element.

The idea of absolute zero is that, theoretically, even though substances can be gasses, there exists a state where all particles of matter become motionless. They would be much like a solid (at that temperature they pretty much are), but there would be no lattices or bonds between separate particles.

This also means that ideal gasses at absolute zero have no volume. Emphasis on ideal, as the concept of an ideal gas revolves around the idea that their particles have negligible volume, and are not affected by attraction/repulsion.

Thus, ideal gasses at absolute zero have no kinetic energy, and no volume.

Note: Temperature(kelvin) = Temperature(degrees celcius) + 273.15

To measure some molar amount of a gas, given that volume is influenced by both temperature and pressure, we have to use some constant of both temperature and pressure.

IUPAC (big science people) recommends the **standard temperature and pressure** (STP).

This describes gasses at 273.15 K or $0^\circ C$, and at a pressure of 100 kPa.

Given a gas at STP,

One mole → 22.71 L

This applies **regardless** of the type of gas, except it only applies for gasses at STP.

Thus, we have a new formula:

> **Info**
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> $$n = \frac{V_{stp}}{22.71}$$

where n = number of moles (mol)
V = Volume of gas at STP (L)

Congratulations! Now you can convert for grams to moles, moles to grams, litres to moles, and moles to litres.

By the way, this isn't even proper stoichiometry.

You're a quarter of the way there. Good job tho :)