BioMedical Admissions Test (BMAT)

Section 2: Chemistry

Topic C4: Quantitative Chemistry

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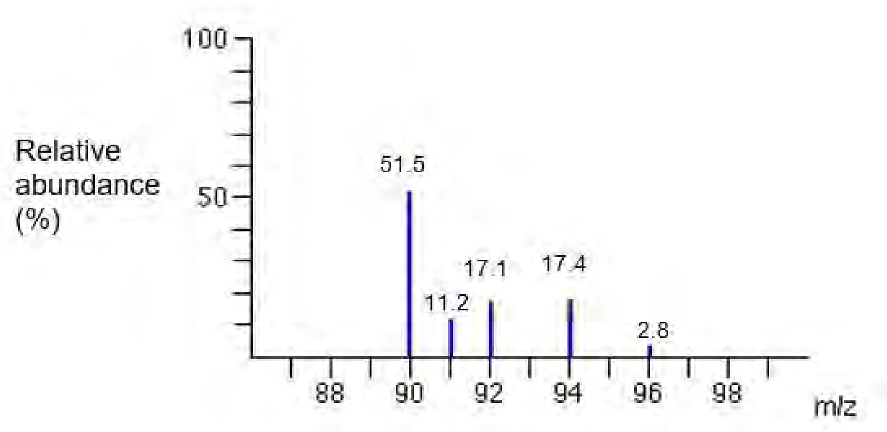
Relative Atomic Mass:

- The **relative atomic mass (A_r)** of an element is the **same as the mass number** of the element.
    - The mass number denotes the number of **protons + neutrons** in the element.
- Because there can be **isotopes of elements all with different masses**, the **A_r is the weighted average of the different masses of all the isotopes of that element**.
    - Isotopes are atoms of the same element that all have the same atomic number but have different mass numbers and therefore a different number of neutrons to each other.

- **Mass spectrometry** can be used to find relative atomic masses. **Ions** are **accelerated** through a **magnetic field**, which allows the mass to be detected. A mass spectrum showing the masses of the different isotopes of an element can be used to determine the **relative atomic mass**, which is the **weighted average mass of all the isotopes.**
    - The positive ion that is accelerated through the magnetic field is formed when electrons collide with the atom and in doing so remove 1 electron from the atom.

Worked example:

For example finding the A_r of zirconium from its mass spectrum



For each peak, read the **% relative isotopic abundance from the y-axis and the relative isotopic mass from the x-axis.**
Multiply these two values together, this will give you the **total mass for each isotope**:

51.5 x 90 = 4635
11.2 x 91 = 1019.2
17.1x92 = 1573.2
17.4 x 94 = 1635.6
2.8 x96 =268.8

The **total mass** of all the different isotopes would therefore be:

4635 + 1019.2 + 1573.2 + 1635.6 +268.8 =9131.8

Because the relative isotopic abundance was given in %, divide the answer by 100

9131.8 ÷ 100 = 91.3 (3sf)

Therefore, the relative atomic mass of zirconium is 91.3

Important: if the relative abundance is NOT given in %, instead of dividing by 100 in the final step, divide by the total relative abundance. This is because it may not be out of 100.

Mass Spectrometry

Mass spectrometry can also be used to work out the **relative formula mass M_r** of a molecular compound in the same way. However the molecules will be broken up into **fragments** when they are bombarded with electrons in the mass spectrometer, causing a fragmentation pattern in the mass spectrum.

- The **fragmentation pattern is used to identify molecules and their structure**.
- A molecular ion M^+ is formed when electrons collide with the molecule and remove 1 electron. The molecular ion will give the peak in the mass spectrum that is the furthest to the right; this will be **the peak with the highest mass**. The **mass of M^+ is the M_r of the molecule.**

Percentage Composition by Mass

- Using the **relative atomic mass** of an element and the **relative formula mass** of the compound, the **percentage mass of an element in a compound can be found**
- The % mass of an element in a compound refers to the **% of a certain element that is found in the compound**. *e.g.* in CO, the % mass of C is 50% and the % mass of O is 50%

$$\%\ mass\ of\ an\ element = \frac{Ar \times no.of\ atoms\ of\ that\ element}{Mr\ of\ whole\ compound} \times 100$$

Worked example

Find the percentage mass of magnesium in magnesium chloride, $MgCl_2$

A_r of Mg $= 25$ A_r of Cl $= 35$

M_r of $MgCl_2$ $25 + 35 + 35 = 95$

$$\% \, mass \, of \, an \, element = \frac{Ar \times no.of \, atoms \, of \, that \, element}{Mr \, of \, whole \, compound} \times 100$$

$$\Rightarrow \frac{25 \times 1}{95} \times 100 = 26.3\% \, (3.s.f)$$

$$\Rightarrow \% \, mass \, of \, Mg \, in \, MgCl_2 = 26.3\%$$

The Mole

The mole is a measure of the **amount of substance.**

To find the mass of 1 mole of a substance, work out the relative formula mass and attach the units "grams".

$$Number \, of \, moles = \frac{mass \, (g)}{Mr \, of \, element/compound}$$

Worked Example 1

For example, fluorine gas, F_2 has a M_r of $38 \, (19 \times 2)$. So 1 mole of F_2 weighs 38 g.

Worked Example 2

For example, 1 mol of $CaCO_3$ weights 100 g. How much does 0.2 mol weigh?

Using the formula above, rearrange so that:

Mass (g) = No. of moles x Mr of compound

0.2 mol weighs 0.2 x 20 g = 20 g.

Worked Example 3

Example Question: How many moles is 54 g of water, H_2O?

$$Number\ of\ moles\ =\ \frac{mass\ (g)}{Mr\ of\ element/compound}$$

The A_r of H $= 1$. A_r of O $= 16$. Therefore M_r of H_2O $= 1 + 1 + 16 = 18$

$$\Rightarrow\ \frac{54}{(18)} = 3\ mol$$

Empirical Formulas

The empirical formula shows the **simplest ratio of the various atoms** in a molecule or compound.

Worked example

Example Question: A compound contains 4.6g of Sodium, 2.8g of Nitrogen and 9.6 of Oxygen. Find the empirical formula of the compound.

Step 1: Determine the masses of each element

Na =4.6g **N =2.8g** **O=9.6g**

Step 2: Determine the A_r values for each element

Na=23 **N=14** **O=16**

Step 3: Determine the number of moles of atoms of each element by using moles = mass ÷ mr

Moles Na = 4.6 ÷ 23 = 0.2

Moles N = 2.8 ÷ 14 = 0.2

Moles O = 9.6 ÷ 16=0.6

Step 4: Determine the ratio of moles and from this find the simplest (empirical) fomrmula

	Na	**N**	**O**
Ratio of moles	0.2	0.2	0.6
Simplest formula	1	1	3

⇒ Empirical formula is $NaNO_3$

The Molar Volume of a Gas

1 mol of any gas at room temperature and pressure (RTP) occupies a fixed volume of $24dm^3$.

- Room temperature = 25˚C. Room pressure = 1 atmosphere
- The molar volume will vary with changing temperature and pressure.

Worked Example

Example question: Calculate the volume of 0.01g of hydrogen at RTP

The question states the conditions are room temperature and pressure (RTP). At RTP, 1 mole of gas= $24dm^3$.

$$Number\ of\ moles\ =\ \frac{mass\ (g)}{Mr\ of\ element/compound}$$

Determine the mass (g) of H_2 that is 1 mol:

1 x M_r of H_2 = mass (g) ⇒ 1x2 = 2g

So 1 mol of H_2 = 2g. 1 mol occupies $24dm^3$ ⇒ 2g of H_2 occupies $24dm^3$

2g = $24dm^3$
0.01g = ?

0.01g of H_2 occupies $\frac{0.01}{2} \times 24dm^3$
= $0.12dm^3$

Solutions

Another key type of calculation in chemistry is surrounding solutions.

- Concentration is a measure of the amount of the solute in a volume of solution.
  - It can be given in mol dm^{-3} – this is the number of moles of a solute per decimetre cubed of solution.
  - An alternative unit is g dm^{-3} – which is the mass of a solute per decimetre cubed of solution.
  - For example 73 of HCl in $1dm^3$ of solution would have the concentration of either 73g dm^{-3} or 2 mol dm^{-3} given HCl has a molar mass of 36.5.

- To convert cm^3 to dm^3, divide by 1000.
- To convert from grams to moles as above, you need to know the molar mass.
    - This is the sum of the mass number of each element in the molecule.
    - For example: HCl is made from H – mass number 1 and Cl – mass number 35.5 so has a molar mass of 36.5g mol^{-1}

- From the concentration and volume of a solution, we can calculate the mass of the solute present.
    - Volume of solution (dm^3) x Concentration (mol dm^3) = amount of solute (mol)
- Some solutions can be described as saturated.
    - This is a solution in which no more solute can dissolve in at the same temperature. You don't need to know why solubility varies with temperature.
    - It is often measured in g of solute per 100g of solvent.
    - The curves drawn can be used for other calculations such as how much precipitate will form if a warmed saturated solution is cooled.

Titration Calculations

Titration calculations are often used to calculate the amount of a base required to neutralise an acid and vice versa.

- First make sure you know the reaction ratio.
    - For example, is it a diprotic acid such as H_2SO_4.
    - HCl has a reaction ratio of 1:1 with NaOH but H_2SO_4 has a 1:2 ratio.
- Then calculate the number of moles of the known substance using the equation n=CV.
    - You will be given the concentration and volume for either the base or the acid.
    - For example 500cm^3 of 0.200 mol dm^3 NaOH will have (0.5 x 0.2) mol of NaOH which is 0.0100 mol.
- Take the reaction ratio and work out how many moles of acid will be needed to neutralise this.
    - For example, for NaOH + HCl → NaCl + H_2O, there is a 1:1 ratio.
    - Therefore there will be 0.0100 mol of HCl.
- You will know either the concentration or the volume of the acid. Use the two values you know to find the missing value.
    - For example, if 1000cm^3 of acid was required, 0.0100 = c x 1
    - Thus the concentration would be 0.0100 mol dm^{-3}.

Worked example:
What volume of 0.200 mol dm^{-3} H_2SO_4 would be required to neutralise 50ml of 0.100 mol dm^{-3} NaOH?

A) 50ml
B) 25ml
C) 12.5ml
D) 6.25ml
E) Can't tell

1. Find the reaction ratio. H_2SO_4 + **2**NaOH → Na_2SO_4 + H_2O. So, it is a 1:2 ratio
2. Here there is a short cut. The concentration of the base is half as strong as the acid.
3. This means that overall the acid is four times stronger than the base.
4. This means one quarter of the volume of the base will be required to neutralise it.
5. 50/4 = 12.5ml – **Answer C.**

Conservation of Mass

In chemical reactions, **mass is conserved**. This means **no atoms are created and no atoms are destroyed**, reactants are simply rearranged to form products. (This is simpler to the conservation of energy principle in physics).

- The conservation mass means that there will always be the same number of each type of atom on both the left and right side of the equation
- The conservation of mass principle can be used to determine reacting masses in a reaction. For example in the reaction of 20g of Na reacting with 35g of Cl, we know that the only product formed in this reaction is NaCl. Following the conservation of mass principle, this means 20g of Na and 35g of Cl will react to form 55g of NaCl.

Worked Example

Example question: What mass of carbon dioxide is produced when 18g of carbon is burnt in air?

Step 1: Determine the balanced equation for the reaction described

carbon + oxygen → carbon dioxide

$$C + O_2 \rightarrow CO_2$$

Step 2: Determine the relative formula masses of the elements in question:

C= 12 CO_2 = 12+16+16 = 44

This tells us that 12g of C react to give 44g of CO_2

Step 3: Determine the masses of each substance

$C \rightarrow CO_2$
$12 \rightarrow 44$
$18 \rightarrow ?$

To get from 18g of carbon to 12g of carbon, divide 18 by 12 (18 ÷ 12 = 1.5)

Therefore to get from 44g of CO_2 the amount in ?, multiply 44 by 1.5

= 66g of CO_2 formed from 18g of carbon burning in oxygen

However, the calculation above assumes that when the reaction occurs that there are **no other products that are "waste products" being formed or that no reactants or products are lost.** In reality, a **100% yield of the desired product is never the case**. Waste products can be formed and atoms of reactants and products can be lost along the way, e.g. when transferring a solution from one beaker to another, there will be some drops of the solution still left in the original beaker.

Percentage yield calculations can be done to compare the actual yield that is obtained and the theoretical predicted yield.

Percentage Yields

Percentage yield = actual yield (grams) ÷ predicted yield (grams) x 100

- **Low yields are undesirable.** It can mean less profit but also wasted chemicals, which is **not sustainable for development**
    - Sustainable development concerns not using resources faster they can be replaced.

Why are actual yields less than 100%?

- The reaction taking place is **reversible.**
    - Because the reaction is going both forwards (to make the products) and backwards (to remake the reactants), 100% of the product will never be formed

at any one time. Some of the products will be reacting together to reform the original reactants and so the actual yield of the products will be lower.

- **Loss of substance when transferring liquid**s
    - 100% of a liquid in 1 container will not be transferred to another container. Some liquid may get spilled or splashed out of the container as droplets, and there will be some liquid left in the original container.
- **Evaporation** of liquids
    - Liquids evaporate greatly when being heated, but there is also evaporation occurring when they are not being heated. This means some of the liquid is lost in the gaseous state
- **Incomplete reactions**
    - Some reactions will remain incomplete; not all the reactants will react together to form products. This could be due to impurities in the starting reactants which will not react to give the desired product.
- **Filtration**
    - When filtering liquids to remove solids, some liquid will also be lost along with the solid.
    - When filtering out a liquid to keep the solid, some of the solid will be lost with the liquid.
- **Unexpected reactions** occur
    - There may be side reactions occuring between the reactants meaning some of the reactants are converted to side products which are undesired. There is now less reactant left to make the desired product.