WJEC Chemistry GCSE

1.1: The Nature of Substances and Chemical Reactions

Detailed notes

Welsh Specification

Atoms, elements and compounds

- Element - a substance made up of a single type of atom. These substances cannot be broken down into any simpler substance by chemical means.
- Chemical symbols represent an atom of an element
  e.g. Na represents an atom of sodium.
- Atom - the basic building block of all substances, the smallest particle of an element that can exist.
- Compound - a substance made from two or more elements that has formed from a chemical reaction, these reactions often involve an energy change.
  Compounds can only be separated into elements by chemical reactions and the compound will have completely unique properties to the elements which form it.
  Compounds can be represented by a chemical formula which uses chemical symbols to state the number and type of each atom present.
  e.g. H_2O represents a molecule with 2 atoms of hydrogen and 1 atom of oxygen.

Ionic compounds

Compounds can be formed from ions. An ion is an atom that has lost or gained at least one electron; since electrons have a negative charge, if an atom loses electrons it will have a positive charge and if an atom gains electrons then it will have a negative charge.

The charge of an ion can be worked out by looking at its group in the periodic table

- Group 1 metals form ions with a charge of 1+
  e.g. Li^+
- Group 2 metals form ions with a charge of 2+
  e.g. Mg^{2+}
- Group 6 elements form ions with a charge of 2-
  e.g. O^{2-}
- Group 7 elements form ions with a charge of 1-
  e.g. Cl^-

The overall charge of a compound is zero, so the number of positively and negatively charged ions must be balanced so there is no net charge. This is used to work out the chemical formula of the ionic compound.

Example: Magnesium chloride

Magnesium is in group 2 so forms a Mg^{2+} ion and chlorine is in group 7 so forms a Cl^- ion. The size of the charge on magnesium is double the size of the charge on a chloride ion, so for every 1 Mg^{2+} ion there must be 2 Cl^- ions to balance this charge and give an overall charge of zero. This gives the chemical formula $MgCl_2$.

Mixtures and separation techniques

What is a mixture?

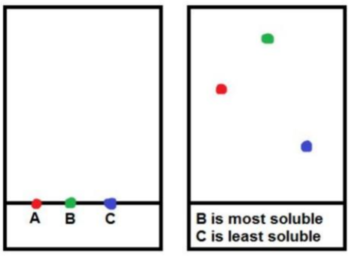
A mixture consists of 2 or more elements or compounds not chemically combined together. The chemical properties of each substance in the mixture are unchanged. Mixtures can be separated by several physical processes, which do not involve chemical reactions and no new substances are made.

These physical separation processes are:

- **Filtration**
- **Evaporation**
- **Chromatography**
- **Distillation**

Chromatography

- Used to separate mixtures and give information to help identify substances.
- Involves a stationary phase and a mobile phase.
- Separation depends on the distribution of substances between the phases - if a substance spends more time in the mobile phase, it'll move further.
- Rf value = distance moved by substance ÷ distance moved by solvent.
- Different compounds have different Rf values in different solvents, which can be used to help identify the compounds – and to distinguish pure from impure substances.
- Compounds in a mixture separate into different spots but a pure compound will produce a single spot in all solvents.

Paper chromatography 	• Analytical technique separating compounds by their relative speeds in a solvent as it spreads through paper. • The more soluble a substance is, the further up the paper it travels. • Separates different pigments in a coloured substance.
Pigment	Solid, coloured substance

Chemical reactions

What is a chemical reaction?

A chemical reaction is a process by which the atoms in the reactants are rearranged to form products. The number and type of each atom present in the reactants will also be present in the products; this is why chemical equations must be balanced.

Representing reactions

Chemical reactions can be represented by 2 types of equations:

- Word equations use the chemical names instead of formulas to show the reaction.
  E.g. methane + oxygen → carbon dioxide + water
- Chemical equations use the chemical formulas and show the ratio of molecules reacting
  E.g. $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

Observations

Some reactions may have no observations, as the products look the same as the reactants - for example both may be colourless solutions. However, some observations are evidence that a reaction has taken place:

- Effervescence - occurs when a gas is released
- Colour changes
- Temperature changes - exothermic reactions give off heat energy to the surroundings, causing an increase in temperature whereas endothermic reactions take in energy from the surroundings, causing a decrease in temperature.

Balancing equations

- In a chemical equation the number and type of atoms on the left-hand side of the equation (the reactants) must be the **same** as the number and type of atoms on the right-hand side of the equation (the products).
- When balancing equations you can only add numbers to the in front of the chemical formula, you can't change the chemical formula itself.
  E.g. The chemical formula for diatomic oxygen is O_2, this cannot be changed (not O_4 etc.), but it can be changed to $2O_2$.

EXAMPLE - magnesium + water → magnesium hydroxide + hydrogen

Step	Result
Write out the equation using chemical formulas, check to see if the number of atoms of each element are equal on both sides of the equation - they aren't.	$Mg + H_2O \rightarrow Mg(OH)_2 + H_2$
The magnesium and hydrogen atoms are balanced (1 x Mg and 2 x H on each side) but there is only 1 oxygen on the left side, but 2 on the right side, so a 2 is added in front of water. The resulting equation is now balanced.	$Mg + 2H_2O \rightarrow Mg(OH)_2 + H_2$

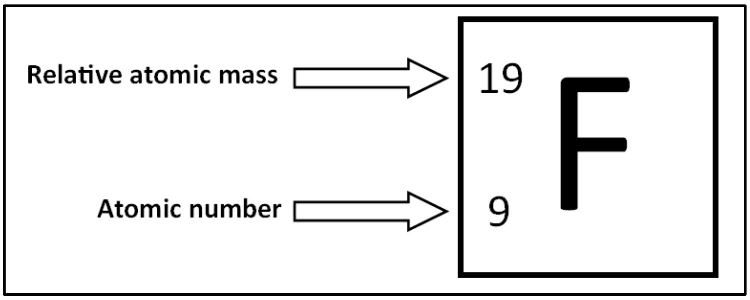
Quantitative chemistry

Mass of atoms

There are 3 subatomic particles in atoms: protons, neutrons and electrons, which will be discussed in more detail in section 1.2. The mass of an atom comes nearly all from protons and neutrons, which each have a mass of 1 relative to each other. The mass of atoms is described 'relatively' because their masses are extremely small - 1 hydrogen atom weighs $1.6735575 \times 10^{-27}$ kg!

- Mass number - the sum of the protons and neutrons in an atom
- Isotopes - atoms of the same element with different numbers of neutrons
- Relative atomic mass - an average value that takes account of the abundance of the isotopes of the element. This is the value given in the periodic table.
- Atomic number - the number of protons in that element.

This is how the relative atomic mass is shown in the periodic table:



Percentage Mass

The percentage mass of an element in a compound tells you what proportion of the mass of an entire compound is from a particular element.

It can be calculated using this formula:

$$\text{\% mass of atom 'X'} = \frac{\text{Mass of element X in formula}}{\text{Total mass of all atoms in formula}} \times 100$$

EXAMPLE - Fe_2O_3

Relative atomic mass of Fe = 55.8
Relative atomic mass of O = 16
Total mass of Fe = (2 x 55.8) = 111.6
Total mass of compound = (2 x 55.8) + (16 x 3) = 159.6

% mass of Fe = (111.6 ÷ 159.6) x 100 = 69.9%

Percentage Yield

$$\text{Percentage yield} = \frac{\text{Mass of product produced}}{\text{Theoretical maximum mass of product possible}} \times 100$$

- It is not always possible to obtain the calculated amount of a product for 3 reasons:
  - Reaction may not go to completion because it is reversible or the reaction is incomplete.
  - Some of the product may be lost when it is separated from the reaction mixture.
  - Some of the reactants may react in different ways to the expected reaction.
- Yield - mass of product obtained.
- To calculate the theoretical mass of a product from a given mass of reactant and the balanced equation for the reaction:
  - Calculate the number of moles of reactants by using mol = mass / molar mass
  - Use balancing numbers to find the number of moles of the desired product.

E.g. $2HCl + Mg \rightarrow MgCl_2$ if you have 2 mol of HCl, you would divide by 2 to get 1 mol. of $MgCl_2$

- ○ Calculate the theoretical maximum mass of a product by then using mass = mol x molar mass

Further quantitative chemistry - Higher tier only

Moles and the Avogradro constant

In chemistry the mass of atoms is so small that even a couple of grams of a substance contains over billions of atoms, this makes quantitative chemistry hard, so the 'mole' is used instead.

- **If you have 1 mole (abbreviated to mol) of atoms then it will weigh its relative atomic mass in grams.**
  - **1 mol of carbon would weigh 12g**
  - **1 mol of iodine would weigh 53 grams**
  - **2 mol of iodine would weigh 106 grams.**
- **This formula enables you to convert between moles, mass (in grams) and Mr of particles.**

$$\text{Number of moles} = \frac{\text{Mass of particle}}{\text{Relative atomic mass of particle}}$$

- **If you have 1 mole of a substance it actually means you have 6.02×10^{23} atoms /ions / molecules of that substance. This number is called the Avogadro constant. This constant enables you to convert between number of moles of a substance and the number of particles there is.**
- **This formula enables you to convert between moles and the number of particles present.**

$$\text{Number of moles} = \frac{\text{Number of particles}}{\text{Avogadro constant}}$$

Calculating the formula of a compound from reacting mass data

Moles, mass and Mr can be used to find the empirical formula of a compound from experimental data, which can then be used to find the chemical formula. An example will be used to illustrate the steps needed.

EXAMPLE: What is the chemical formula of a compound that contains only arsenic and oxygen and contains 75.74% of Arsenic by mass?

1. **Write out a table as shown and fill in with information from the question:**

Element	As	O
Mass	75.74	24.26
Mr	74.9	16
Mol		
Ratio		

2. The number of moles of each can be calculated using the formula above, mol = mass / Mr. The calculated values can be added to the table.

Element	As	O
Mass	75.74	24.26
Mr	74.9	16
Mol	1.011	1.52
Ratio		

3. The ratio is calculated by dividing each number of moles by the lowest number of moles present.
   So the ratio for arsenic is 1.011 / 1.011 = 1
   The ratio for oxygen is 1.52 / 1.011 = 1.50

Element	As	O
Mass	75.74	24.26
Mr	74.9	16
Mol	1.011	1.52
Ratio	1	1.5

4. You now have a ratio of As : O as 1:1.5, you can't have 1.5 atoms of oxygen in a compound, so the ratio is doubled so there are no decimals.
   This gives the empirical formula - As_2O_3, which in this case is also the chemical formula.

<u>Calculating the mass of reactants or products from a balanced chemical equation</u>

Experimental data can be used to find the mass of reactants or products from a balanced chemical equation. The method for this will again be illustrated with an example.

EXAMPLE:

The balanced chemical equation for the reaction is: $Mg(s) + H_2O(g)\ MgO(s) + H_2(g)$

The teacher used 3.50 g of magnesium. Use the equation to calculate the maximum mass of magnesium oxide produced.

1. Set up a table with the balanced chemical equation running along the top and mass, Mr and mol in the first column. Fill in the information given in the question and on the periodic table.

	Mg	H_2O	→	MgO	H_2
Mass	3.5	-		?	-
Mr	24	18		40	2
Mol					

2. Calculate the number of moles of Magnesium present by using the formula mol = mass / Mr.
   3.5 / 24 = 0.146

	Mg	H_2O	→	MgO	H_2
Mass	3.5	-		?	-
Mr	24	18		40	2
Mol	0.146				

3. In this equation the ratio of Mg : MgO is 1:1, so the number of moles of MgO will be equal to the number of moles of Mg. When the ratio is not 1:1 the number of moles must be multiplied accordingly - if the ratio was 1:3 then the number of moles must be tripled. The ratio is given by the balanced chemical equation.

	Mg	H_2O	→	MgO	H_2
Mass	3.5	-		?	-
Mr	24	18		40	2
Mol	0.146			0.146	

4. You now have enough information to calculate the mass of MgO, as mass = Mr x mol.

	Mg	H_2O	→	MgO	H_2
Mass	3.5	-		5.8	-
Mr	24	18		40	2

Mol	0.146			0.146	

Therefore the maximum possible mass of magnesium oxide produced is 5.8 grams to 2 significant figures.