INORGANIC CHEMISTRY

Key GCSE ideas

Revise the following GCSE topics:

- ***Fundamental Ideas***
- ***Introduction to Bonding and Structure and Bonding***
- ***The Periodic Table***
- ***Quantitative Chemistry***
- ***Reaction Rates***
- ***Equilibria***
- ***Energetics***

Introduction

Inorganic Chemistry is about all the matter and the chemical reactions that take place between atoms. This makes Chemistry a vast subject. In the Inorganic Chemistry, you will study the foundation concepts which form the building blocks for all the other modules in your A-level course. You will also study aspects of the Physical Chemistry.

You will learn how to use the periodic table and how to calculate the amount of substance using different types of available data. You will study the role of electrons in chemical bonding and details of the structure and bonding which is essential for all further topics.

You will also study enthalpy changes and their determination from experimental data and data tables. Reaction rates and equilibria will focus on how changing reaction conditions affect the speed of reaction and the position of equilibrium.

Fundamental Ideas

1. Key definitions:
   Element - a substance made up of only one type of atom

   Compound - a substance made up of two or more different atoms chemically bonded

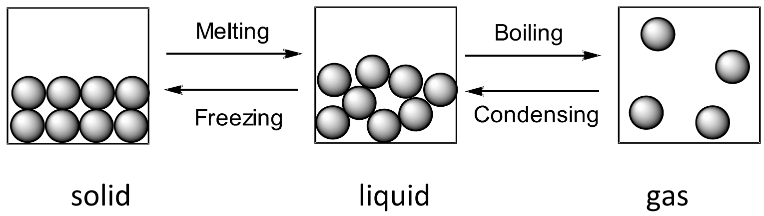
   Mixture - a substance made up of two or more different chemicals (elements or compounds) <u>not</u> chemically bonded

   Atom- the smallest part of an element that can exist on its own

2. Mixtures can be separated using physical methods of separation. Examples:

Method	Example of mixture	Details
Filtration	Sand and water	The solid (sand) stays on the filter paper but water drains through
Crystallisation	Copper sulphate solution	Evaporate the water by heating, leaving the solid behind
Distillation	Ethanol and water	Separates by boiling point. Evaporate and then condense.
Chromatography	Soluble dyes	Soak water up paper. Separates by solubility.

3. There are three states of matter:



4. Summary of atomic structure:

Subatomic Particle	Relative Mass	Relative Charge	Position in Atom
Proton	1	+1	Nucleus
Neutron	1	0	Nucleus
Electron	0	-1	Electron Shell

5. There were many models of atom. Below are diagrams of the "Plum Pudding Model" and the Bohr model of an atom.

The key differences between these models include:

a. Bohr model has a nucleus

b. Bohr model has electron shells

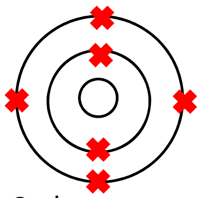
6. Key facts about the current model of atom:

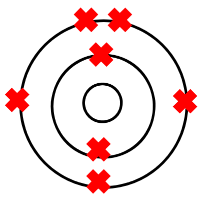
   The number of protons in an atom is equal to the proton/atomic number. In an atom, the number of electrons is equal to the proton/atomic number. To work out how many neutrons are in an atom, we subtract the atomic number from the atomic mass.

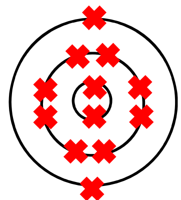
   Atoms have no overall charge because protons have a positive charge, electrons have a negative charge and the number of protons and electrons in an atom is the same.

   Electrons in an atom are found on the electron shells. The first electron shell holds a maximum of two electrons, the second holds a maximum of eight electrons and the third holds a maximum of eight.

   Examples of electronic structures of atoms:


Carbon atom


Nitrogen atom


Magnesium atom

Structure and Bonding

When two non-metal atoms react they form a **covalent bond**. This type of bond can be described as a shared pair of electrons. The resulting structure is called a molecule.

e.g. hydrogen, H_2

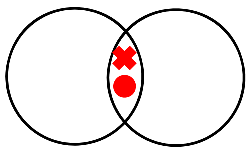
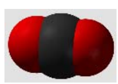


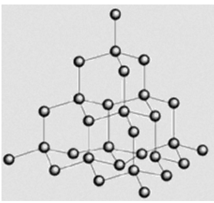
Diagram of molecules:

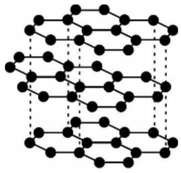


Properties of simple covalent molecules:

1. Conductivity – There is no overall charge on a molecule, so there is no conduction of electricity.
2. Melting and boiling point - Low melting and boiling points as there are only weak intermolecular forces between the molecules. Not much energy is needed to separate the molecules. Larger molecules have stronger intermolecular forces.
3. Hardness – Molecular substances tend to be relatively soft

Some covalent substances form giant covalent lattices, rather than molecules. They include diamond (C), graphite (C) and silica (SiO_2).



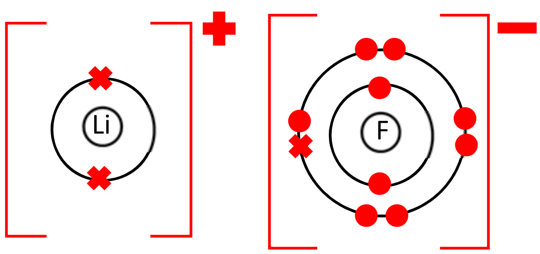


Properties of giant covalent lattices:

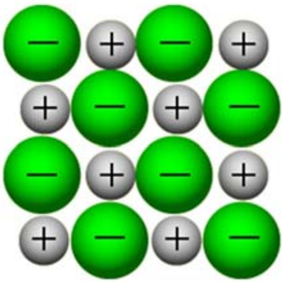
1. Conductivity – Diamond and silicon dioxide have no charged particles so do not conduct electricity. Graphite has only 3 covalent bonds per carbon atom. Its fourth electron is a delocalised electron so can carry a charge as it can move through the structure. Graphite conducts electricity.
2. Melting and boiling point - High melting and boiling point as atoms are held together by many strong covalent bonds. Lots of energy is needed to break these bonds.
3. Hardness – Diamond is hard because each atom has 4 strong covalent bonds holding it in place, so atoms do not easily move. Graphite is soft because it is arranged in layers with only weak forces between them. Layers can easily slide.

When a metal and non-metal react they form an **ionic bond**. This type of bond can be described as electrostatic attraction between oppositely charged ions. The resulting structure is called a giant ionic lattice.

e.g. lithium fluoride, LiF



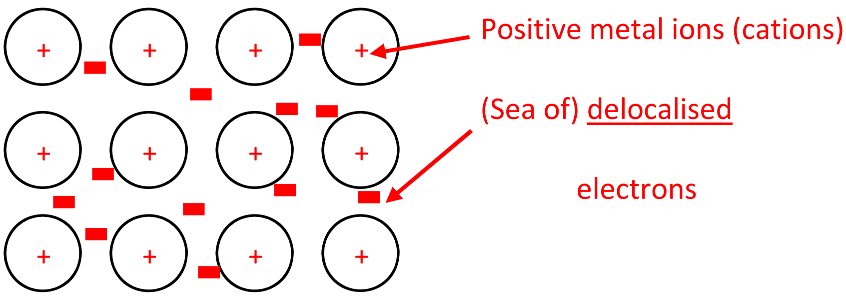
The diagram of an example of the giant ionic structure:



Properties of ionic compounds:

1. Conductivity - Ions are charged. Ions cannot flow in a solid, so cannot carry a charge and conduct electricity. Ions can flow in a liquid so can carry a charge and conduct electricity. Ions can flow when dissolved in water so can carry a charge and conduct electricity.
2. Melting and boiling point - High melting and boiling point as ions have strong electrostatic attractions holding them together. Large amounts of energy are required to separate the ions.
3. Hardness – Ionic solids are brittle – they shatter easily.

When atoms of metal bond together, they do so with the **metallic bonding**. The nature of this type of bonging can be described as a strong electrostatic attraction between positive metal ions and the negative sea of delocalised electrons:



Properties of metallic substances:

1. Conductivity – Delocalised electrons are charged, and can move through the structure so metals conduct electricity.
2. Melting and boiling point - High melting point as the attraction between the ions and delocalised electrons is strong. Lots of energy is needed to break the attraction.
3. Hardness – Pure metals are relatively soft as they are arranged in neat layers that can slide over each other. Alloys have different sized ions so the layers are distorted and do not slide over each other easily. This makes them harder.

Periodic Table

The elements in the periodic table are arranged in order of increasing ***atomic*** number. Elements with similar properties are in columns, known as ***groups***. Elements in the same group have the same number of outer shell electrons.

Only outer shell electrons are involved in bonding, so it is the number of these electrons that control the reactions of an element. This is why elements in the same group all react in a similar way.

Elements show trends across periods.

Definition of periodicity - A repeating pattern of properties at regular intervals.

The period number indicates the number of electron shells in an atom.

E.g. The element with 3 electron shells and 7 electrons in the outer shell is Chlorine, Cl.

Number of electron shells	Number of outer shell electrons	Element
3	2	Mg
4	1	K
5	2	Sr
3	3	Al

Group	Name
1	Alkali metals
2	Alkali Earth Metals
7	Halogens
8	Noble Gases

Development of the Periodic Table

Initially the periodic table was arranged in order of increasing atomic weights. The early periodic tables were incomplete and elements were not placed in appropriate groups.

Mendeleev left gaps for undiscovered elements, and, also changed the order so elements which reacted in a similar way were in the same group.

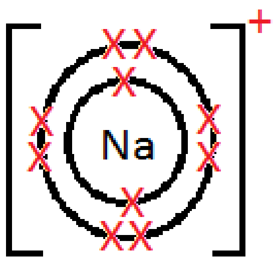
Metals and Non-Metals

Below is part of the periodic table, draw the ‘stair case’ that separates the metals from the non-metals.

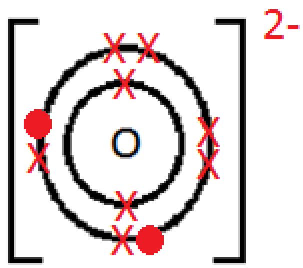
					He
B	C	N	O	F	Ne
Al	Si	P	S	Cl	Ar
Ga	Ge	As	Se	Br	Kr
In	Sn	Sb	Te	I	Xe
Tl	Pb	Bi	Po	At	Rn

Metals form ***positive*** ions, whereas non-metals form ***negative*** ions. The exception to this is hydrogen, which is a non-metal and forms a positive ion.

Example: the diagram of a sodium **<u>ion</u>**:

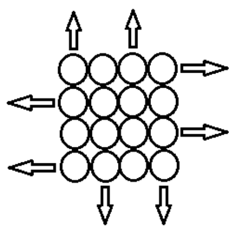


Example: the diagram of an oxide ion:



Metals have metallic bonding. They have the following properties:

- High melting point: There is a strong electrostatic attraction positive metal ions and delocalised electrons.
- Good electrical conductor: delocalised electrons can carry a charge.
- Malleable and ductile: metal atoms are arranged in neat layers which can slide and move past each other.

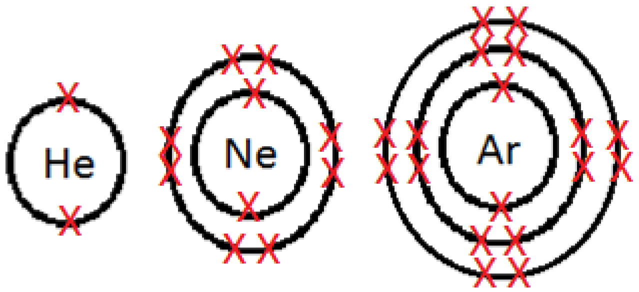


Non-metals have covalent bonding. They usually have the following properties:

- Low melting point: Not a lot of energy is needed to overcome weak intermolecular forces between simple covalent molecules, hence why most covalent substances are gases at room temperature and pressure.
- Poor electrical conductor: molecules have no overall charge.

Group 0

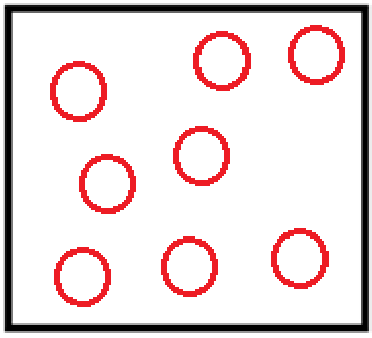
All noble gases have a full outer shell of electrons, so they are said to be inert (very unreactive). This means that they do not easily form molecules.



Uses of noble gases include:

- Filling food packaging
- Filling lightbulbs
- Atmospheres for carrying out reactions.

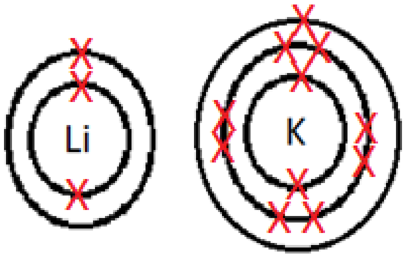
Due to Group 0 elements having a full outer shell of electrons, they exist as monatomic gases.



Going down group 0, the boiling point increases. This is because as the atoms increase in size, the number of electrons increases. This causes the forces between the ***atoms*** to become ***stronger*** and more energy is needed to overcome them.

Group 1

All alkali metals have one outer shell electron. This means that they all react in a ***similar*** way.



Example: reaction with water

Metal	Observation when reacted with water
Lithium	Fizzing
Sodium	Strong fizzing, metal melts
Potassium	Vigorous fizzing, lilac flame

When an alkali metal reacts with water, it produces a metal hydroxide and hydrogen gas.

Alkali metal + water → metal hydroxide + hydrogen

lithium + water → lithium hydroxide + hydrogen

$2Na + 2H_2O \rightarrow 2NaOH + H_2$

$2K + 2H_2O \rightarrow 2KOH + H_2$

Alkali metals will also react with chlorine and oxygen.

Metal + chlorine → metal chloride

Sodium + chlorine → sodium chloride

$2K + Cl_2 \rightarrow 2KCl$

Metal + oxygen → metal oxide

Lithium + oxygen → Lithium oxide

$4Na + O_2 \rightarrow 2Na_2O$

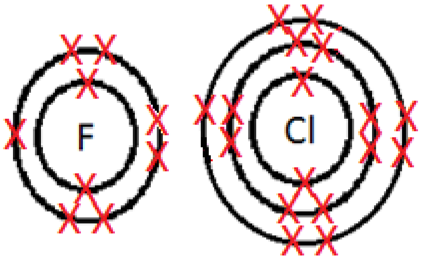
As group 1 metals increase in atomic number, (descend group):

- Density ***increases***
- Strength ***decreases***
- Melting point ***decreases***
- Reactivity ***increases***

Alkali metals react by losing one electron. Reactivity ***increases*** down the group because:

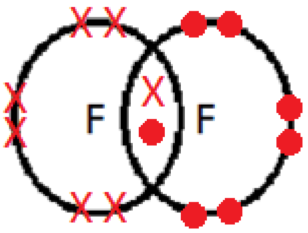
- Increased atomic radius
- Increased electron shielding – more FULL electron shells means that the outer electron is more shielded from the nucleus.
- Decreased attraction to the nucleus – the electron is ***more/less*** easily lost.

Group 7



All halogens have seven electrons in their outer shell. They exist as diatomic molecules consisting of two atoms.

This is the dot and cross diagram for a fluorine molecule:



As the atomic number of the halogen increases, the molecular mass ***increases***, this means there are greater intermolecular forces, which require ***more*** energy to overcome, hence ***higher*** melting points and boiling points:

Halogen	Melting pt/^{o}C	Boiling pt. / ^{o}C	State at room temperature
F_2	-220	-188	Gas
Cl_2	-101	-35	Gas
Br_2	-7	59	Liquid
I_2	114	184	Solid

Halogens react with metals to form halide ions, e.g. F^-, Cl^-, I^- by ***gaining*** an electron to complete their outer shell. In general, halogens react with metals to form metal halides with ***ionic*** bonding.

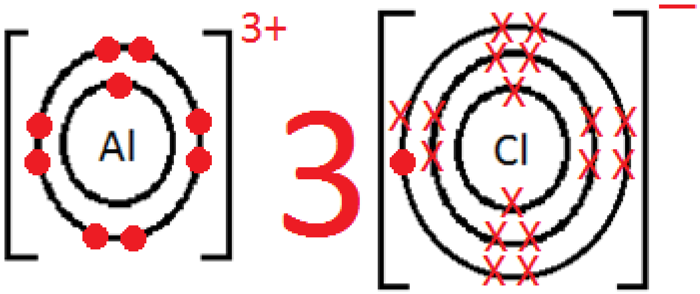
For example:

Aluminium + iodine → aluminium iodide

iron + bromine → iron bromide

$Mg + Cl_2 \rightarrow MgCl_2$

This is the dot and cross diagram for aluminium chloride:



Halogens react with non-metals, e.g. hydrogen to form molecules. The bonding is ***covalent.***

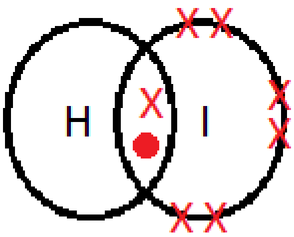
For example:

Hydrogen + bromine → hydrogen bromide

$H_2 + Cl_2 \rightarrow 2HCl$

$H_2 + I_2 \rightarrow 2HI$

This is the dot and cross diagram for hydrogen bromide:



Halogens undergo displacement reactions. In these reactions, the <u>more</u> reactive halogen displaces the less reactive halogen.

E.g. sodium bromide + chlorine → sodium chloride + bromine

$2NaI + Br_2 \rightarrow 2NaBr + I_2$

Halogens react by ***gaining*** one electron to form halide ions, Cl^-, Br^-, I^-.

Reactivity decreases down the group because:

- Increased atomic radius
- Increased electron shielding – more full electron shells means that the electron is more shielded and less attracted.
- Less attraction to the nucleus – the electron to be gained is less attracted.

Quantitative Chemistry

Key definitions:

- Reactant - A substance that takes part in and undergoes change in a reaction.
- Product – The substance made in a chemical reaction.
- State symbol – Shows the physical state of a substance in a reaction.

The law of conservation of mass states that no atoms are lost or made during a chemical reaction, so the mass of the products equals the mass of reactants.

A reaction may appear to have gained mass because one of the reactants is a gas. A reaction may also appear to have lost mass because one of the products is a gas.

We use the law of conservation of mass to balance equations:

$Mg + 2HCl \rightarrow MgCl_2 + H_2$

$2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$

$C_3H_8 + 5O_2 \rightarrow 4H_2O + 3CO_2$

When two elements combine, a compound has the ending of 'ide'.

When two or more elements combine with oxygen, a compound ends with 'ate'.

When two identical elements combine, the name of the substance is identical.

The charge on elemental ions can be predicted from the periodic table.

Group	1	2	3	5	6	7
Charge	1+	2+	3+	3-	2-	1-

The charge on a transition metal is denoted by Roman numerals in brackets when naming a compound.

Ionic compounds have no overall charge. When the charges on the ions are the same and opposite, they cancel out. When the charges are different, the charges are swapped.

This is how we work out formulae of compounds using charges on ions:

$Mg^{2+} + O^{2-} \rightarrow MgO$

$Al^{3+} + Cl^{-} \rightarrow AlCl_3$

$Li^{+} + S^{2-} \rightarrow Li_2S$

Some compounds contain group ions that have their own formula:

Hydroxide: OH^-	Carbonate: CO_3^{2-}	Ammonium: NH_4^+
Sulfate: SO_4^{2-}	Nitrate: NO_3^-	

More definitions:

- Isotope: Atoms of the same element with a different number of neutrons.
- Relative atomic mass (A_r): The mass of an atom expressed in atomic mass units.
- Relative formula mass (M_r): The sum of all relative atomic masses of the atoms present in the formula.

The relative atomic mass (A_r) of an element is the same as its mass number in the periodic table. Examples:

The relative atomic mass of Na is: 23

The relative formula mass of NaCl is: 58.5

The relative formula mass of $CaCO_3$ is: 100

To calculate the **relative atomic mass** of chlorine given the percentage abundance of each isotope (75% of ^{35}Cl and 25% of ^{37}Cl)

A_r of Cl = $\left(\frac{75}{100} \times 35\right) + \left(\frac{25}{100} \times 37\right) = 35.5$

This is the formula for calculating the **percentage composition** by mass of an element in a compound:

$$\%\ of\ element\ A = \frac{A_r\ of\ A\ \times Number\ of\ atoms\ of\ A}{M_r} \times 100$$

Example: calculating the percentage by mass of Mg, S and O in $MgSO_4$.

% Mg = ((24*1)/120) * 100 = 20%

% S = ((32*1)/120) * 100 = 26.7%

% O = ((16*4)/120) * 100 = 53.3%

The sum of all the percentages should be 100%.

Chemical amounts are measured in **moles**. The unit of this is the mol. The number of atoms, molecules or ions in one mole of a substance is equal to the Avogadro's constant (6.02×10^{23} per mole).

The mass of one mole of a substance in grams is equal to its relative mass.

The moles equation:

Moles (mol) = Mass (g) ÷ Relative mass (g/mol)

We can rearrange it to find mass:

Mass (g) = Moles (mol) X Relative mass (g/mol)

Balanced equations show the molar ratio of reactants to products.

E.g. $2Mg + O_2 \rightarrow 2MgO$

This means that 2 moles of magnesium reacts with one mole of oxygen gas to produce 2 moles of magnesium oxide. This allows us to work out the mass of magnesium and oxygen are needed to react and how much product will form.

Example: if 36g of Mg reacts in air, what mass of MgO will be produced?

Moles of Mg = 36 ÷ 24 = 1.5 mol

Since the molar ratio is 2:2, this means the number of moles of MgO produced will also be 1.5 mol.

Mass of MgO = 1.5 x 40 = 60 g

The balancing numbers in a symbol equation can be calculated from the masses of reactants and products. The masses can be converted into moles, and then these can be converted into whole number ratios.

Example: 28 g of iron reacts with exactly 8 g of oxygen molecules. The reaction produces 36 g of iron oxide (FeO). What is the simplest whole number ratio of iron that reacts with oxygen?

Moles of iron that react: 28/56 = 0.5 mol

Moles of oxygen that react: 8/32 = 0.25 mol

Moles of iron oxide produced: 36/72 = 0.5 mol

Ratio of Fe : O_2 : FeO = 0.5:0/25:0.5

Whole number ratio: 2:1:2

In a chemical reaction it is common to use an **excess** of one of the reactants to ensure that all of the other reactant is used. This is normally done when one reactant is more expensive than another. It is called the **limiting react**ant because is limits the amounts of product formed.

The definition of the limiting reactant – The reactant that is completely used up in a reaction.

Uncertainties can occur in experiments. This involves calculating the mean and the range. When calculating a mean, it is important to exclude any anomalous results. The mean should also match the number of significant figures given the question. To calculate the range, the largest value is subtracted from the smallest value.

The equation for uncertainty is: range/2

If a mean is 25.1 cm^3 is found, and a range of 0.4 is calculated. Then the uncertainty for the measurement would be written as 25.1 ±0.2.

More definitions:

- Solute: A substance that dissolves in a liquid.
- Solvent: A liquid used to dissolve a solute.
- Solution: A mixture of a soluble solute which has been dissolved in a solvent.
- Saturated: A solution that has the maximum amount of solute dissolved in it.

Many chemical reactions take place in solution. The **concentration of a solution** can be measured in **mass per unit volume**. It has the units of g/dm^3. 1 dm^3 is equal to 1000 cm^3 which is equal to 1 litre. To increase the concentration of a solution you can either add more solute whilst keeping the volume the same, or reduce the volume whilst keeping the mass of solute the same.

The equation for calculating concentration:

Concentration (g/dm^3) = Mass (g) ÷ Volume (dm^3)

For example, calculate the concentration of 10 g of NaCl in 500 cm3 of water:

Concentration = 10 g ÷ 0.5 dm^3

= 20 g/dm^3

Further Quantitative Chemistry

Percentage yield and atom economy

The equation for percentage yield:

$$percentage\ yield\ = \frac{mass\ of\ product\ actually\ made}{maximum\ theoretical\ mass\ of\ product} \times 100$$

E.g. Calculate the percentage yield of a reaction between 8.00g of iron (III) oxide and excess magnesium if only 3.64g of iron was actually produced.

Theoretical mass of iron produced = 5.6g
Percentage Yield = 65.0%

The reasons as to why a reaction might not obtain 100% yield:

- The reaction may not go to completion as it is reversible
- Some of the product may be lost during separation
- Some of the reactants may react in ways different to the expected reaction

Atom economy - A measure of the amount of starting materials that end up as useful products.
The equation for atom economy:

$$Atom\ economy = \frac{relative\ formula\ mass\ of\ desired\ product\ from\ equation}{sum\ of\ relative\ formula\ masses\ of\ all\ reactants\ from\ equation} \times 100$$

E.g. Calculate the atom economy for forming hydrogen in the following reaction.

$$CH_4 + H_2O \rightarrow CO + 3H_2$$

Atom Economy = 17.6%

The reasons as to why high atom economy is important for sustainable development:

- Low atom economy produces waste which can be expensive to remove/dispose
- Raw materials are expensive and high atom economy reactions use fewer raw materials

Volumes of gases

At room temperature and pressure one mole of a gas occupies a volume of 24 dm^3 or 24,000 cm^3.

To calculate the volume of a gas or the moles of a gas the following equation is used:

$$number\ of\ moles = \frac{volume\ of\ gas\ (dm^3)}{24\ (dm^3)}$$

E.g. Calculate the volume taken up by 0.125mol N_2 gas.

$v(N_2) = 3\ dm^3$

E.g. Calculate the number of moles of 16dm^3 O_2 gas.

$n(O_2) = 0.67mol$

To link the volume to mass of gas the following equation is used:

$$volume\ of\ gas\ (dm^3) = \frac{mass\ of\ gas\ (g)}{Mr\ of\ gas} \times 24$$

E.g. Calculate the mass of 30 dm^3 of He at room temperature and pressure.

Mass He = 5g

Equations can be used to calculate volumes of gases:

E.g. Calculate the volume of carbon dioxide released when 18g of pentane completely combusts.

$C_5H_{12} + 8O_2 \rightarrow 5CO_2 + 6H_2O$

$n(C_5H_{12}) = 0.25$

$n(CO_2) = 1.25$

$v(CO_2) = 30dm^3$

More about concentration

Concentration of a solution can be measured in mol/dm^3. The following equation can be used to calculate the concentration of a solution:

$$concentration\ (mol/dm^3) = \frac{moles\ of\ solute\ (mol) \times 1000}{volume\ of\ solution\ (cm^3)}$$

E.g. Calculate the concentration of 6.25×10^{-3} moles of a 25cm^3 NaCl solution.

$c(NaCl) = 0.25\ mol/dm^3$

Concentration of a solution can be calculated if mass of the solute and volume of the solvent is known. The above equation and the equation linking moles to mass is also used.

E.g. Calculate the concentration of a 250cm^3 solution containing 2.9g KF.

c(KF) = 0.2 mol/dm^{-3}

To convert between the units of concentration (g/dm^3 ←→ mol/dm^{-3}) the following equation is used:

$$concentration\ (mol/dm^3) = \frac{concentration\ (g/dm^3)}{Mr}$$

E.g. Calculate the concentration of g/dm^3 of 0.5 mol/dm^3 NaOH.

c(NaOH) = 20g/dm^3

Titrations

Titrations are used to determine the concentration of either an acids or a base.
A volumetric pipette is used to measure out an aliquot.

Definition of aliquot: A sample of a total amount of liquid taken for chemical analysis.

A burette is designed to measure various volumes of a solution. The volumes from a burette are always recorded to 2 decimal places. An indicator is used to show the end point of the reaction. This is the point at which the acid and alkali have reacted completely. Three examples of indicators suitable for titrations include: phenolphthalein, methyl orange, litmus

Example of a detailed method for the process of carrying out a titration:

1. Fill burette with acid (e.g. HCl), record starting titre to 2 decimal places, ending in 0 or 5.
2. Use a volumetric pipette to measure out a specific aliquot of alkali (e.g. 25cm^3 NaOH)
3. Add the alkali to the conical flask, add (phenolphthalein) indicator, and place on a white tile.
    - The phenolphthalein will turn **<u>pink</u>**
4. Slowly add the acid dropwise from the burette whilst swirling the conical flask.
5. Stop adding acid when the indicator when the indicator changes colour (e.g. colourless) to show the endpoint.
6. Record the final titre on the burette to 2 decimal places, ending in 0 or 5.
7. Subtract the final titre from the starting titre to calculate the volume added.
8. Repeat titration until concordance is reached (same volume +/- 0.1 cm^3)

The equation $n = \frac{c \times v}{1000}$ is used in titrations to calculate the unknown concentration of an acid/base.