science

chemistry

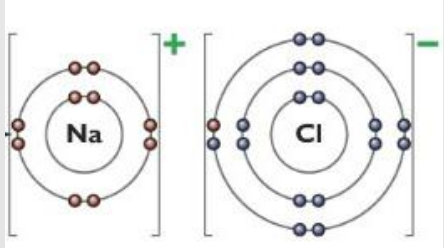
Topic 2– Structure & Bonding

1) What is an ion?

An ion is a charged atom. It becomes charged by gaining or losing electrons.

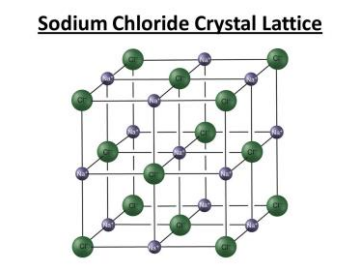
2) Ionic Bonding

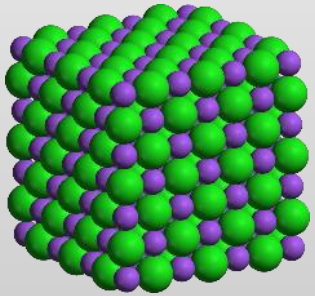
- An ionic bond is an electrostatic force of attraction between 2 or more oppositely charged ions.
- Forms between metals and non-metals.
- Metals give their electrons to the non-metals.



3) Ionic Structures

- **The ions attract each other to form an ionic lattice, which is a regular 3D arrangement of ions.**



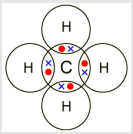


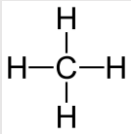
4) Properties of ionic substances:

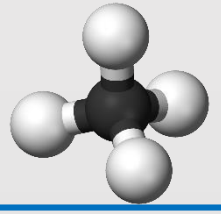
- High melting points - lots of energy needed to break all the strong electrostatic forces between ions.
- Do not conduct electricity when solid
- Conduct when liquid (molten or aqueous) as the ions are free to move and carry the charge.

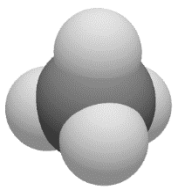
5) Using models to represent structures of compounds (HT)

- Simple particle model (s,l,g) assumes particle made up of solid spheres with no forces between them.
- Dot and cross model (shows where electrons come from but doesn't show 3D shape).
- Ball and stick model (limited as spheres are shown to be solid and inelastic and does not show true shape).
- Close packed model (difficult to see arrangement in 3D).



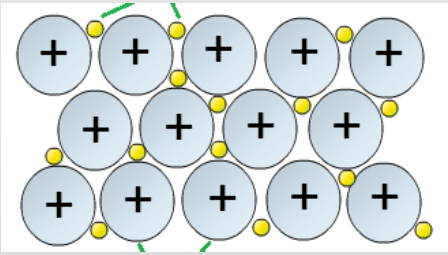






6) Metallic bonding

- Metals have a regular 3D arrangement of layers of positive ions held together by a sea of delocalised electrons.



7) Properties of metals:

- They are malleable & ductile because they the layers of metal ions can slide over each other and still be held together by the delocalised electrons.
- They are good conductors of electricity because the delocalised electrons are free to move through the whole structure and carry the charge.
- Good conductors of thermal energy because energy is transferred by the delocalised electrons.

8) Alloys

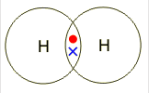
- Pure metals are too soft for many uses and so are **mixed** with other metals to make alloys which are harder.
- Different sized atoms in the alloy prevent layers from sliding.

9) What is a molecule?

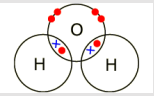
2 or more atoms joined together by covalent boning (sharing electrons)

10) Covalent Bonding

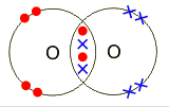
- A covalent bond is a pair of electrons shared between two atoms.
- Forms between two or more non-metals.
- Forms either simple molecules (see below) or giant molecules (see overleaf).
- A simple hydrogen molecule (H_2)



- A simple water molecule (H_2O)



- A simple oxygen molecule (O_2)



11) Properties of simple molecules such as the ones above:

- Low melting points and boiling points (usually gases) because they are held together by weak forces (intermolecular forces) of attraction so little energy is needed to break these forces.
- They do not conduct electricity because the molecules do not have a charge – there are no ions or free electrons to carry the charge.

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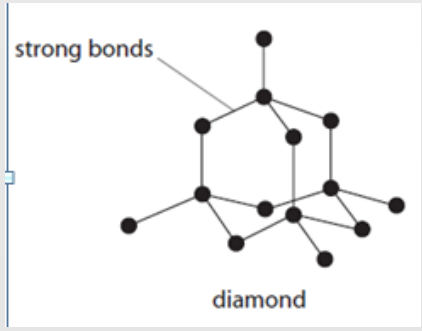
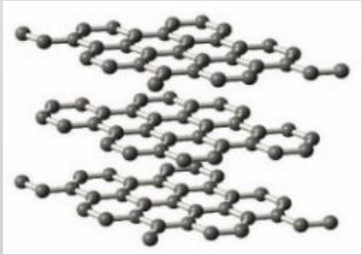
12) Giant Molecules (diamond and graphite)

- Diamond has a **giant covalent** structure with every carbon atom joined to 4 others by strong covalent bonds.
- Lots of energy is needed to break all these strong bonds which make diamonds very hard (useful as drill bits).



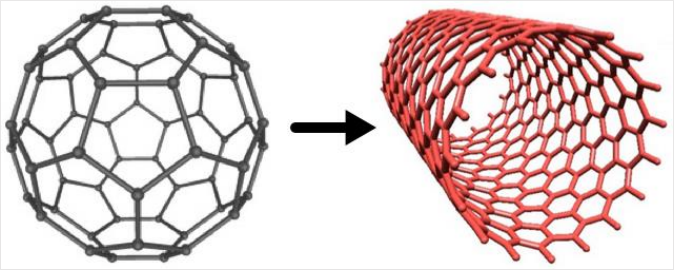
- Graphite has a **giant covalent** structure with every carbon joined to 3 others by strong covalent bonds.
- It forms layers of hexagons which can slide over each other as the layers are held together by weak forces of attraction (intermolecular forces).
- This makes graphite soft & slippery (pencils & lubricants)



- ***Graphite conducts electricity because these delocalised electrons are free to move through the structure and carry the charge.***

13) Graphene & fullerenes

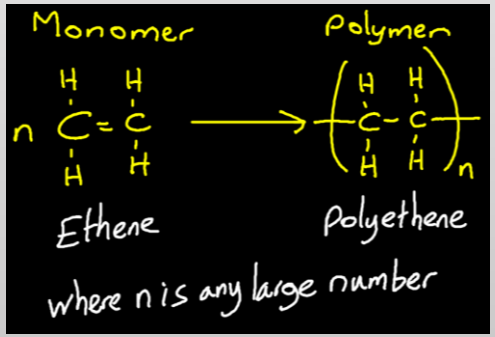
- Graphene is a single layer of graphite and only one atom thick (often called a 2D material).
- Carbon atoms can also form hollow balls or ring structures called fullerenes. Buckminsterfullerene (C_{60}) has a spherical shape like a football (used as a lubricant as it can roll).



- Carbon nanotubes can also form which are used in medicine.

14) Polymers

- Polymers have very large molecules. The atoms in the polymer molecules are linked to other atoms by strong covalent bonds.
- Intermolecular forces between polymer molecules are fairly strong and so these substances are solids at room temp.



15) Nanoparticles (SINGLE CHEMISTRY ONLY):

- Nanoscience refers to structures that are 1–100 nm in size (1nm = 1×10^{-9}m)
- Nanoparticles, are smaller than fine particles, which have diameters between 100 and 2500 nm (1×10^{-7} m and 2.5×10^{-6} m). Coarse particles (PM10) have diameters between 1×10^{-5} m and 2.5×10^{-6} m. Coarse particles are often referred to as dust.
- Nanoparticles have a high surface area to volume ratio so smaller quantities are needed to be effective compared to normal particle sizes.
- As the side of cube decreases by a factor of 10 the surface area to volume ratio increases by a factor of 10.
- Nanoparticles have many applications in medicine, in electronics, in cosmetics and sun creams, as deodorants, and as catalysts.

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Topic 3 – Quantitative Chemistry

1) Conservation of mass – In any chemical reaction no atoms are lost or made during a chemical reaction so the mass of the products equals the mass of the reactants.
If an element is heated in the air, it's mass may appear to go up as it is reacting with the oxygen in the air.
Metal carbonates, when heated, undergo thermal decomposition (breaking down using heat). The equation for the reaction is:
Calcium carbonate → Calcium oxide + Carbon dioxide.
$CaCO_{3\,(s)} \rightarrow CaO_{(s)} + CO_{2\,(g)}$
The mass might appear to go down but that is because one of the products made is a gas and this will escape into the environment.

2) Relative Atomic Mass (Ar)
The relative atomic mass of an element (Ar) is the average mass of the isotopes of that element compared to 1/12th the mass of carbon 12.

3) Relative Formula / Molecular Mass (Mr)
The relative formula mass (Mr) of a compound is the sum of the relative atomic masses of the atoms in the chemical formula.
(e.g. Mr of NaOH = 23 + 16 + 1 = 40)
The relative mass of any substance, in grams, is equivalent to 1mol of that substance.

4) Isotopes – Elements **with the same number of protons but different number of neutrons** (so they will have different relative masses).

5) Percentage by mass of an element
To calculate what percentage of a compound is made from a particular element:
e.g.
% of Mg in MgO
$\% \text{ Mg} = \frac{Mr\ of\ Mg}{Mr\ MgO}\ x\ 100 = \frac{24}{40} x100 = 60\%$

6) Avogadro and the mole
A mole is an amount of substance. 1mol of any substance will contain the same number of particles (6.02×10^{23}).
1 mol of H_2O will contain 6.02×10^{23} molecules.
This is known as Avogadro's number.

7) Calculating moles (HT only)

$$n = \frac{m}{Mr}$$

n = no. of moles (mols)

m = mass (grams)

Mr = Relative atomic mass

8) Reacting Masses

$$Mg + 2HCl \rightarrow MgCl_2 + H_2$$

If you have a 60g of Mg, what mass of HCl do you need to convert it to $MgCl_2$?

A_r : Mg =24 so mass of 1 mole of Mg = 24g
M_r : HCl (1 + 35.5) so mass of 1 mole of HCl = 36.5g

So 60g of Mg is 60/24 = 2.5 moles

Balanced symbol equation tells us that for every one mole of Mg, you need two moles of HCl to react with it.

So you need 2.5x2 = 5 moles of HCl

You will need 5 x 36.5g of HCl= 182.5g

9) Limiting reactants
The reactant that is used up first (the other one will be in excess). This limits the amount of product that is made and ensures that one of the reactants is completely used up. Work out the moles of each that you have & using the equation you can determine which one you have less moles of.

10) Concentration of solutions
Worked out by knowing the mass of solid dissolved and the volume of liquid it's dissolved in
$\text{Conc} = \frac{mass\ (g)}{volume\ (dm^3)}$ or $\text{conc} = \frac{mass\ (g)}{volume\ (cm^3)} x1000$

$1dm^3 = 1000cm^3$

11)Using moles to balance equations (HT only)
Convert the masses in grams to amounts in moles and convert the number of moles to simple whole number ratios.
e.g. 8.5g of sodium nitrate is heated in a test tube, it forms 6.9g of sodium nitrite and the remainder is oxygen. Balance the equation

	$NaNO_3$ →	$NaNO_2$ +	O_2
Work out mols	8.5/85 =0.1	6.9/69 = 0.1	1.6/32 = 0.05

then divide each by smallest

Equation is **2**$NaNO_3$ → **2**$NaNO_2$ + O_2

12) Uncertainty - determine whether the mean value falls within the range of uncertainty of the result
1. Calculate the mean
2. Calculate the range of the results
3. Estimate of uncertainty in mean would be half the range.

e.g.
Mean value is 46.5s
Range of results is 44s to 49s = 5s
So the time taken was 46.5s ±2.5s

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Topic 3 – Quantitative Chemistry (Single Chemistry)

13) Yield – The amount of a product in a reaction is known as the yield.

14) Percentage yield

This compares the actual mass you obtained in an experiment with the mass you should have got if all the reactants reacted together and there was no loss. This is know as the theoretical yield & is calculated from doing a moles calculation as before:

$$\% \text{ Yield} = \frac{actual\ mass\ of\ product}{\text{max}\ theoretical\ mass} \times 100$$

15) Factors affecting percentage yield

- Reaction may be reversible
- Unwanted / competing reactions
- Loss of product in handling / separation
- Reactants may not be pure.

16) Atom economy

This is a measure of how green /environmentally friendly the reaction is as it compares how much of your reactants are converted into useful products:

$$\text{Atom economy} = \frac{Mr\ of\ useful\ products}{\text{Mr of all the reactants}} \times 100$$

17) Calculating moles for solutions

$$n = \frac{CV}{1000} \quad \text{or} \quad n = CV$$

n= no. of moles (mols)

C = concentration (mol/dm^3)

V = Volume **(cm^3)**

n= no. of moles (mols)

C = concentration (mol/dm^3)

V = Volume **(dm^3)**

18) Volumes of gases

- 1mol of any gas has a volume of 24dm^3 (24litres or 24000cm^3)

So if you have 0.5mols of O_2 gas then it would have a volume of 12dm^3

19) Titrations

•The volumes of acid and alkali solutions that react with each other can be measured by titration using a suitable indicator.
•If the concentration of one of the reactants is known, the results of a titration can be used to find the concentration of the other reactant.
•You should be able to calculate the chemical quantities in titrations involving concentrations (in moles per dm^3) and masses (in grams per dm^3).
•Acid is usually placed in the burette and a known volume of known concentration of alkali is measured out using a pipette.
•The pipette is filled with a pipette filler and then emptied under gravity into a conical flask.
•2-3 drops of a suitable indicator (e.g. phenolphthalein) is added which will go pink in alkali and then colourless at end point.
•The flask is placed on a white tile so the colour change is clearer to see.
•The unknown acid is then added slowly from the burette to the alkali until the last drop of pink disappears. This is called the end-point.
•The reading on the burette is taken at eye level (to reduce any errors).
•The experiment is then repeated until you obtain **concordant** results (within ± 0.1cm^3 of each other)
•Looking at the results in the table below, you would use your two concordant results which would be from experiments 2 & 3. The average would be 21.33cm^3.

Attempt	1	2	3
Initial burette reading (cm^3)	0.00	0.00	0.00
Final burette reading (cm^3)	22.50	21.30	21.35
Titre value (cm^3)	22.50	21.30	21.35

20) Titration calculations

What is the concentration of hydrochloric acid if 21.33cm^3 of acid were needed to neutralise 25cm^3 of 0.5mol/dm^3 solution of sodium hydroxide?
e.g.

$$HCl + NaOH \rightarrow NaCl + H_2O$$

From the equation 1mol HCl needs 1mol of HCl

Using $n = \frac{CV}{1000}$ work out moles of NaOH

$$\text{Moles NaOH} = \frac{0.5\ x\ 25}{1000} = 0.0125$$

So from the equation the same amount of HCl would be needed so we have 0.0125mols of HCl.

Rearrange the equation to get the concentration:

$$\text{Conc} = \frac{1000\ x\ n}{V\ (acid)} = \frac{1000\ x\ 0.0125}{21.33} = 0.59\text{mol/dm}^3$$

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