CHEMISTRY REVISION NOTES

FOR AQA GCSE (9-1)

SIMPLE, CLEAR & MEMORABLE

PAPER 1

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USING THIS BOOK

This is **Higher Tier** only material – this means you will only need to revise this if you are sitting the higher tier Biology paper.

This is **Chemistry (separate science)** only material – this means you will only need to revise this if you are sitting the triple award separate science Biology paper (**8462**).

This is **Higher Tier and Chemistry (separate science)** only material – this means you will only need to revise this if you are sitting the higher tier Biology paper (**8462**).

THIS IS A SPECIFICATION CHAPTER

1.1 THIS IS A SPECIFICATION TOPIC

1.1.1 This is a specification subtopic

THE PERIODIC TABLE

← This is an identical copy of the periodic table provided in the AQA GCSE Chemistry exams

Key

relative atomic mass
atomic symbol
name
atomic (proton) number

1	2											3	4	5	6	7	0
							1 **H** hydrogen 1										4 **He** helium 2
7 **Li** lithium 3	9 **Be** beryllium 4											11 **B** boron 5	12 **C** carbon 6	14 **N** nitrogen 7	16 **O** oxygen 8	19 **F** fluorine 9	20 **Ne** neon 10
23 **Na** sodium 11	24 **Mg** magnesium 12											27 **Al** aluminium 13	28 **Si** silicon 14	31 **P** phosphorus 15	32 **S** sulfur 16	35.5 **Cl** chlorine 17	40 **Ar** argon 18
39 **K** potassium 19	40 **Ca** calcium 20	45 **Sc** scandium 21	48 **Ti** titanium 22	51 **V** vanadium 23	52 **Cr** chromium 24	55 **Mn** manganese 25	56 **Fe** iron 26	59 **Co** cobalt 27	59 **Ni** nickel 28	63.5 **Cu** copper 29	65 **Zn** zinc 30	70 **Ga** gallium 31	73 **Ge** germanium 32	75 **As** arsenic 33	79 **Se** selenium 34	80 **Br** bromine 35	84 **Kr** krypton 36
85 **Rb** rubidium 37	88 **Sr** strontium 38	89 **Y** yttrium 39	91 **Zr** zirconium 40	93 **Nb** niobium 41	96 **Mo** molybdenum 42	[98] **Tc** technetium 43	101 **Ru** ruthenium 44	103 **Rh** rhodium 45	106 **Pd** palladium 46	108 **Ag** silver 47	112 **Cd** cadmium 48	115 **In** indium 49	119 **Sn** tin 50	122 **Sb** antimony 51	128 **Te** tellurium 52	127 **I** iodine 53	131 **Xe** xenon 54
133 **Cs** caesium 55	137 **Ba** barium 56	139 **La*** lanthanum 57	178 **Hf** hafnium 72	181 **Ta** tantalum 73	184 **W** tungsten 74	186 **Re** rhenium 75	190 **Os** osmium 76	192 **Ir** iridium 77	195 **Pt** platinum 78	197 **Au** gold 79	201 **Hg** mercury 80	204 **Tl** thallium 81	207 **Pb** lead 82	209 **Bi** bismuth 83	[209] **Po** polonium 84	[210] **At** astatine 85	[222] **Rn** radon 86
[223] **Fr** francium 87	[226] **Ra** radium 88	[227] **Ac*** actinium 89	[261] **Rf** rutherfordium 104	[262] **Db** dubnium 105	[266] **Sg** seaborgium 106	[264] **Bh** bohrium 107	[277] **Hs** hassium 108	[268] **Mt** meitnerium 109	[271] **Ds** darmstadtium 110	[272] **Rg** roentgenium 111	[285] **Cn** copernicium 112	[286] **Nh** nihonium 113	[289] **Fl** flerovium 114	[289] **Mc** moscovium 115	[293] **Lv** livermorium 116	[294] **Ts** tennessine 117	[294] **Og** oganesson 118

* The Lanthanides (atomic numbers 58 – 71) and the Actinides (atomic numbers 90 – 103) have been omitted.

Relative atomic masses for **Cu** and **Cl** have not been rounded to the nearest whole number.

1 ATOMIC STRUCTURE AND THE PERIODIC TABLE

Equations are in **bold.**

1.1 A SIMPLE MODEL OF THE ATOM, SYMBOLS, RELATIVE ATOMIC MASS, ELECTRONIC CHARGE AND ISOTOPES

1.1.1 Atoms, elements and compounds

- All substances are made of atoms.
- An atom is the smallest part of an element that can exist.
- Atoms of each element are represented by a chemical symbol.
- There are about 100 elements, all shown in the periodic table.

- Compounds:
    - are two or more elements chemically combined in fixed proportions
    - have combined properties of their atoms
    - can only be separated by chemical reactions
    - are represented by formulae using the symbols of their atoms

- Chemical reactions:
    - are represented by word or symbol equations
    - involve the formation of one or more new substances
    - often involve a detectable energy change

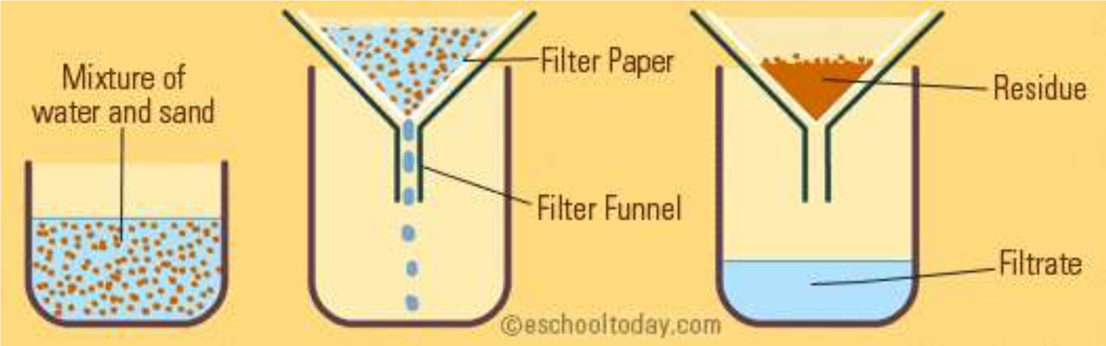
- The Law of Conservation of mass says that the total mass of the products of a reaction is equal to the total mass of the reactants, so we must balance symbol equations to make this applicable.

- WORD EQUATION: sodium + water → sodium hydroxide + hydrogen
- SYMBOL EQUATION: $Na + H_2O \rightarrow NaOH + H_2$
- BALANCED: $2Na + 2H_2O \rightarrow 2NaOH + H_2$

1.1.2 Mixtures

- Mixtures:
  - are two or more elements or compounds not chemically combined together
  - contain molecules with unchanged chemical properties
  - can be separated by physical processes (no reaction)

- Physical separation techniques:
- Filtration:

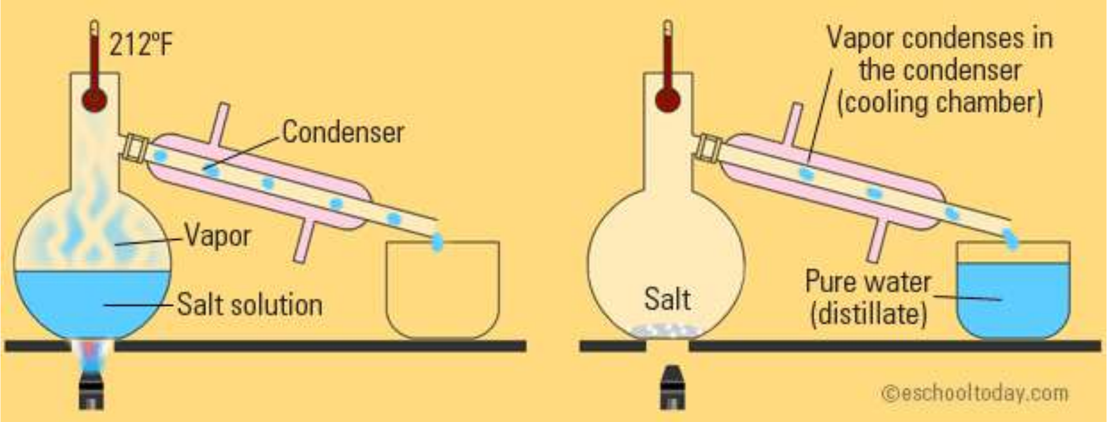


  - used to separate insoluble substances from a solvent
  - involves the use of filter paper in a filter funnel
  - large particles cannot go through the paper and become the residue

- Crystallisation:

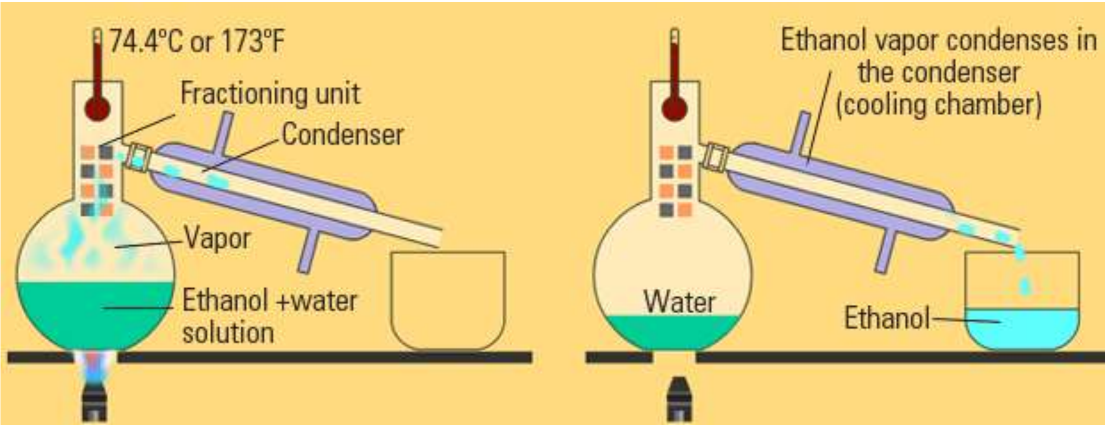
  - used to separate soluble substances from a solution
  - involves the application of heat to a solution
  - solutes with a higher boiling point than the liquid will not evaporate

- Simple distillation:



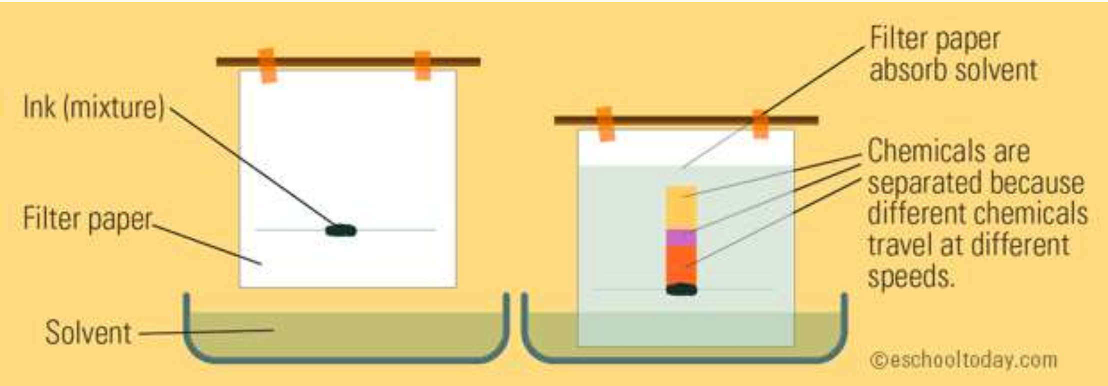
  - used to separate a solvent from a solution
  - involves the evaporation and condensation of a solvent
  - solutes with a higher boiling point than the liquid will not evaporate and will not enter the condenser

- Fractional distillation:



  - used to separate miscible liquids with different boiling points
  - involves the use of glass beads in the fractionating column
  - liquids with a higher boiling point will not evaporate and will condense in the column

- Chromatography:



  - used to separate and identify mixtures in a solution
  - involves the use of a drop of the solution on filter paper in a solvent
  - substances with different solubilities travel at different speeds

1.1.3 The development of the model of the atom

- Before the discovery of the electron, atoms were thought to be **tiny spheres** that could not be divided.
- The **discovery of the electron led to the plum pudding model** of the atom in which the atom is a ball of positive charge with embedded negative electrons.
- The **alpha particle scattering experiment led to the nuclear model** in which the **mass of the atom was concentrated at the centre** (charged nucleus) because some alpha particles were repelled back to their source and did not penetrate the gold foil.
- **Niels Bohr** adapted the nuclear model by suggesting that **electrons orbit the nucleus** at specific distances; theoretical calculations agreed with this.
- Further experiments led to the idea that the **nucleus contained smaller particles** (protons) each having the same amount of positive charge.
- **James Chadwick** provided evidence to show the **existence of uncharged particles** (neutrons) in the nucleus.

Plum pudding model	Nuclear model
sphere of positive charge	protons in nucleus
equal mass across sphere	mass concentrated at nucleus
electrons embedded	electrons orbit nucleus

1.1.4 Relative electrical charges of subatomic particles

Particle	Relative charge
Proton	+1
Neutron	0
Electron	-1

- atomic number = proton number = electron number

- Atoms have no overall electrical charge.
- Atoms of a different elements have different numbers of protons.

1.1.5 Size and mass of atoms

- Atomic mass is concentrated in the nucleus.

Particle	Relative mass
Proton	1
Neutron	1
Electron	[neglible]

- atomic mass = atomic number + neutrons

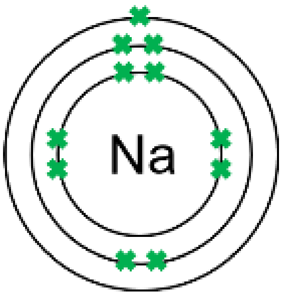
- Isotopes:
  - are atoms of the same element with a different number of neutrons
  - have the same atomic number but a different mass number
  - may be less stable or radioactive
  - have different physical properties
  - have the same chemical properties
  - have the same electronic structure

1.1.6 Relative atomic mass

- The relative atomic mass of an element is an average value that considers the abundance of the isotopes of an element.

1.1.7 Electronic structure

- Electrons of an atom occupy the innermost available shells.
- The lowest (innermost) energy level can have a maximum of 2 electrons; other shells can have up to 8 electrons.
- E.g. sodium's electronic structure is 2,8,1 and is drawn as follows:



There is a nucleus containing 11 protons and 12 neutrons.

There are:
- 2 electrons in the lowest energy level
- 8 electrons in the second energy level
- 1 electron in the highest energy level

1.2 THE PERIODIC TABLE

1.2.1 The periodic table

- The elements in the periodic table are arranged in order of atomic number.
- Elements with similar properties are in groups (columns).
- Elements in the same group have the same number of electrons in their outer shell (which gives them similar chemical properties).

1.2.2 Development of the periodic table

- Before the discovery of subatomic particles, the elements were arranged in order of atomic weight and some elements were placed inappropriately as a result of this strict protocol.
- Dmitri Mendeleev left gaps for undiscovered elements and changed the order of some elements to keep certain patterns in properties.
- Elements were later discovered and filled the gaps.
- Knowledge of isotopes explained reasons for some disorders in the table.

1.2.3 Metals and non-metals

- Atoms gain or lose electrons (whichever is easier) to become more stable, so:

- Metals are elements that react to form positive ions by losing electrons from their outermost shell.
- Non-metals are elements that react to form negative ions by gaining electrons from their outermost shell.

- Metals are found to the left and bottom of the periodic table.
- Non-metals are found to the right and top of the periodic table.

Properties	Metals	Non-metals
Physical	high melting points good electrical conductor good heat conductor shiny high density malleable and ductile	low melting points poor electrical conductor poor heat conductor dull low density brittle
Chemical	form basic oxides	form acidic oxides

1.2.4 Group 0

- Noble gases
- As they have full outer shells, they are unreactive.
- Boiling points increase going down the group.
- Reactivity increases going down the group.

1.2.5 Group 1

- Alkali metals
- They have one electron in their outer shell.
- Reactivity increases going down the group because the outermost electron is easier to lose when the distance to the nucleus is longer and the electrostatic force of attraction to the nucleus is weaker.

1.2.6 Group 7

- Halogens
- They have seven electrons in their outer shell.
- Halogens exist as diatomic molecules.

- magnesium + chlorine → magnesium chloride
- sodium + bromine → sodium bromide
- potassium + iodine → potassium iodide

- Boiling points increase going down the group.
- Reactivity decreases going down the group because it is harder for an electron to attract to the nucleus as it is further away.

- A more reactive halogen can displace a less reactive halogen from an aqueous solution of its salt.

- magnesium iodide + sodium → sodium iodide + magnesium
- calcium chloride + potassium → potassium chloride + calcium

1.3 PROPERTIES OF TRANSITION METALS

1.3.1 Comparison with Group 1 elements

- The transition elements:
    - good electrical conductors (iron, Fe)
    - good heat conductors (zinc, Zn)
    - high densities (gold, Au)
    - high melting points (except mercury, Hg)
    - hard (nickel, Ni)
    - shiny (chromium, Cr)
    - less reactive (copper, Cu)

1.3.2 Typical properties

- Many transition elements form coloured ions and are useful as catalysts:
    - copper(II) sulfate is blue
    - nickel(II) carbonate is pale green
    - chromium(III) oxide is dark green
    - manganese(II) chloride is pale pink

2 BONDING, STRUCTURE, AND THE PROPERTIES OF MATTER

2.1 CHEMICAL BONDS, IONIC, COVALENT AND METALLIC

2.1.1 Chemical bonds

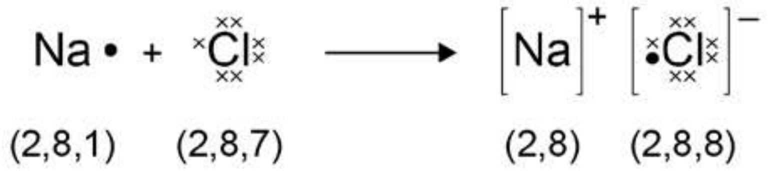
- There are three types of chemical bonds:
    - ionic (oppositely charged ions – metals and non-metals)
    - covalent (shared electrons – non-metals)
    - metallic (shared delocalised electrons - metals)

2.1.2 Ionic bonding

- When metals and non-metals react electrons in the outer shell of the metal atom are transferred to the non-metal.
- Metals lose electrons to become positively charged ions.
- Non-metals gain electrons to become negatively charged ions.
- Ions have full outer shells.

- Electron transfer can be represented by a dot and cross diagram, e.g.:

- $Na + Cl \rightarrow Na^{+} + Cl^{-}$

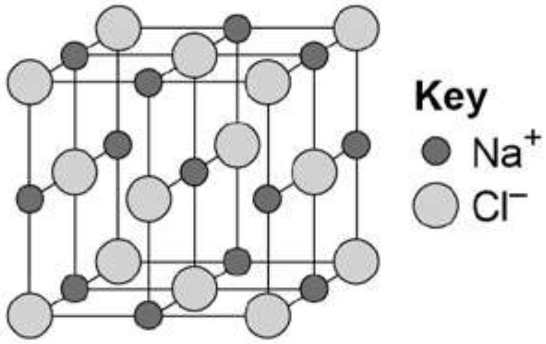


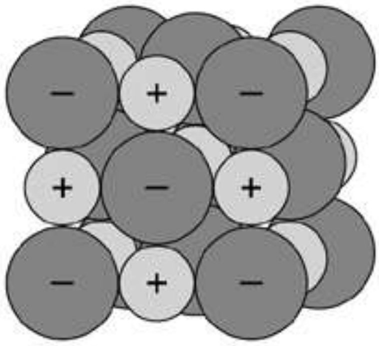
- Ionic charges are related to groups as follows:
    - Group 1: 1+
    - Group 2: 2+
    - Group 3: 3+
    - Group 5: 3-
    - Group 6: 2-
    - Group 7: 3-

2.1.3 Ionic compounds

- Ionic bonding is when ionic compounds are held together by strong electrostatic forces of attraction between oppositely charged ions.

- E.g. this is a giant ionic lattice for the structure of sodium chloride:





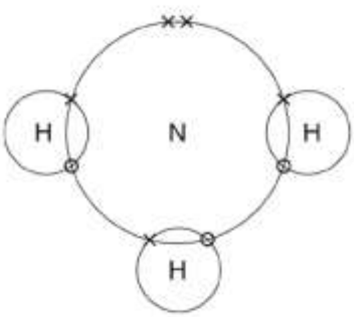
- Limitations of the particle models:

Limitation	Dot and cross	Ball and stick	2D diagram	3D diagram
may not show relative size	✓	✓	✓	✓
may not show real shape	✓		✓	
does not show electrostatic forces	✓	✓		✓
shows ions as solid spheres		✓	✓	✓
does not show ions touching		✓	✓	

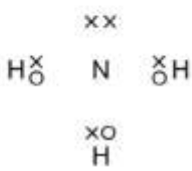
2.1.4 Covalent bonding

- Covalent bonds are stronger bonds where atoms share pairs of electrons.
- Covalently bonded atoms may consist as simple molecules like O_2 or Cl_2.
- Covalent bonds can be represented in the following forms:

For ammonia (NH_3)



and/or



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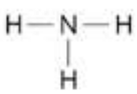

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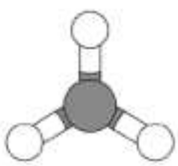
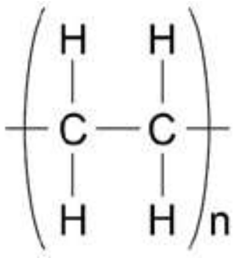
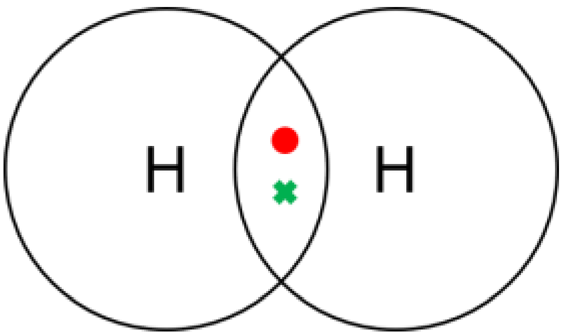
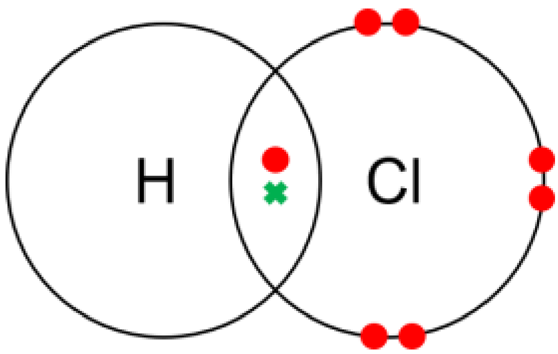


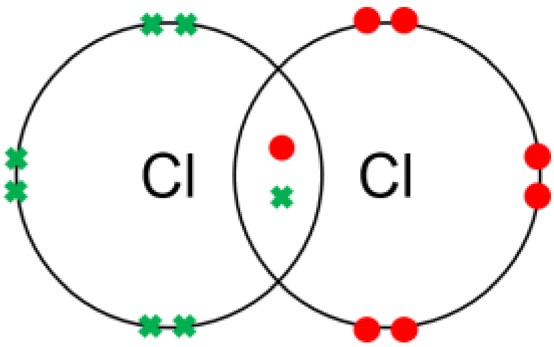
- A polymer is a substance made of long molecules of repeating units.
- E.g. poly(ethene) is made of small ethene molecules (C_2H_4) reacting together to form a long chain.
- Poly(ethene) can be represented by the diagram below, where n is a large number:

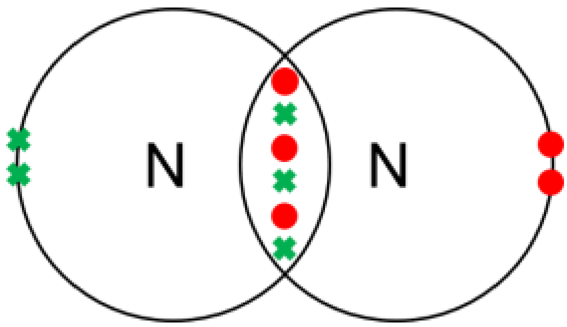


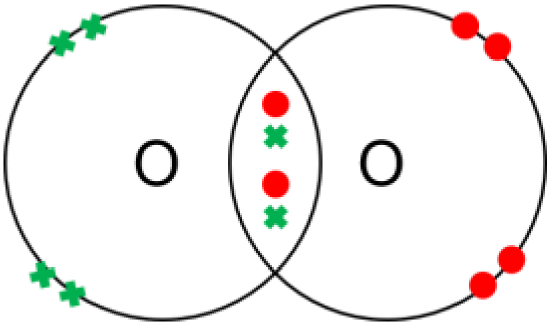
- Here are some more covalent bonding diagrams:

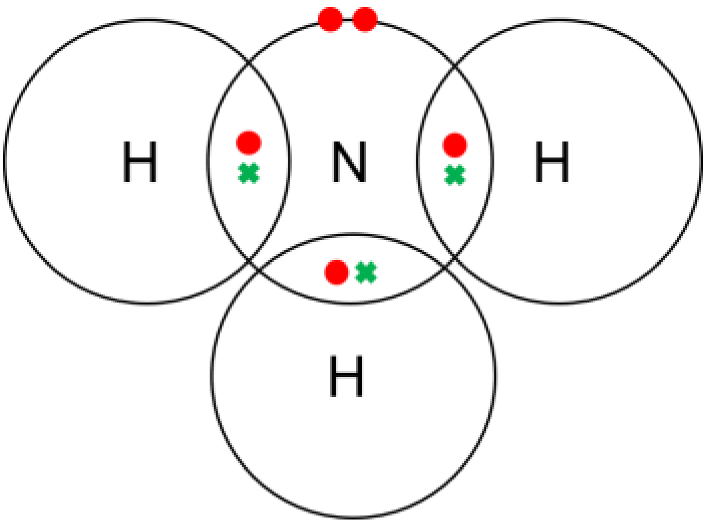


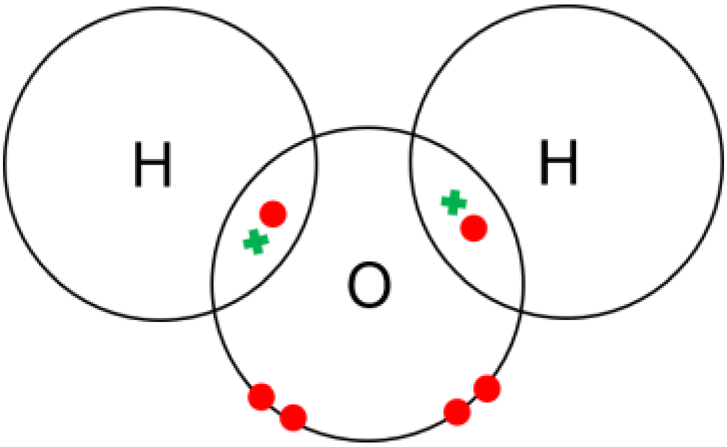


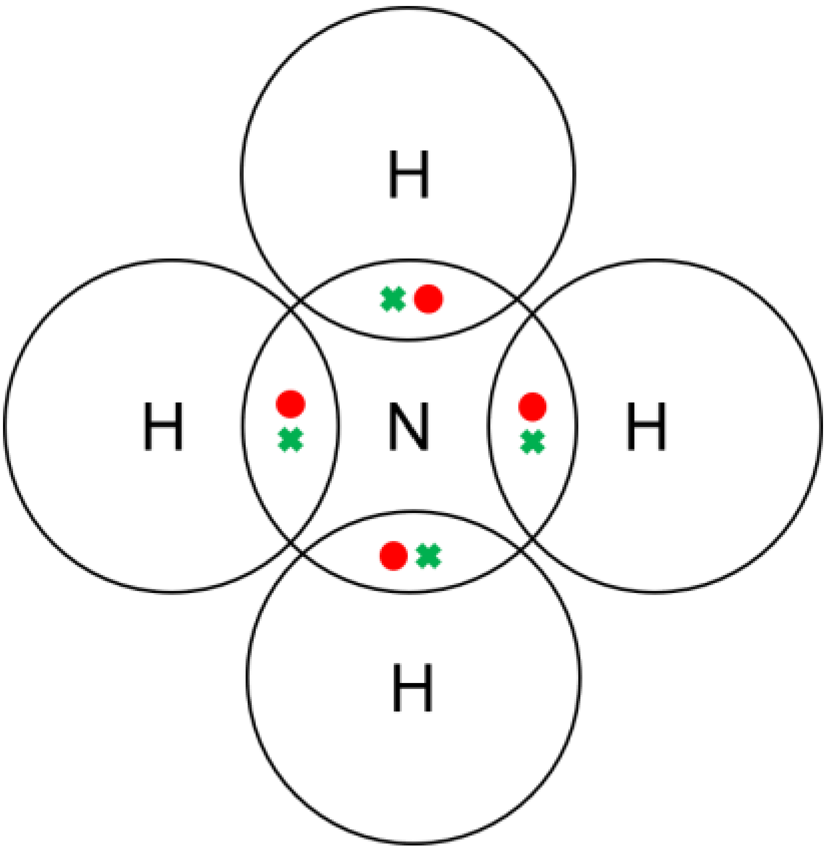






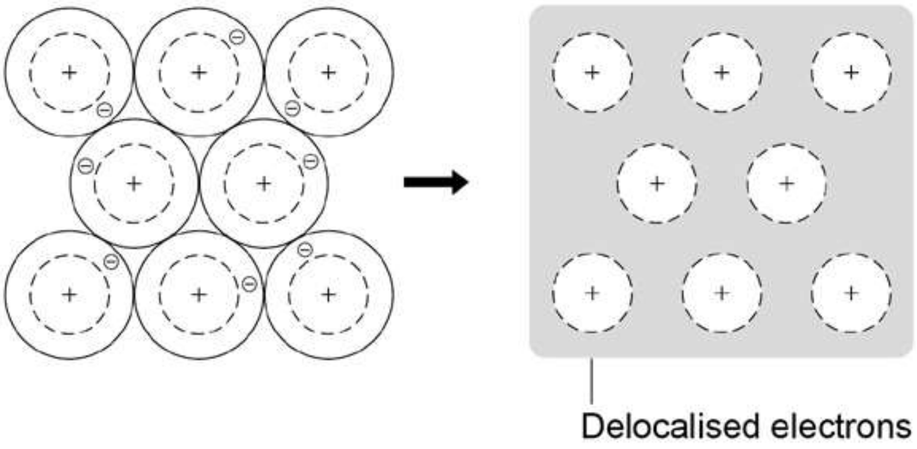

H
N
H
H


H
H
O


H
H
N
H
H

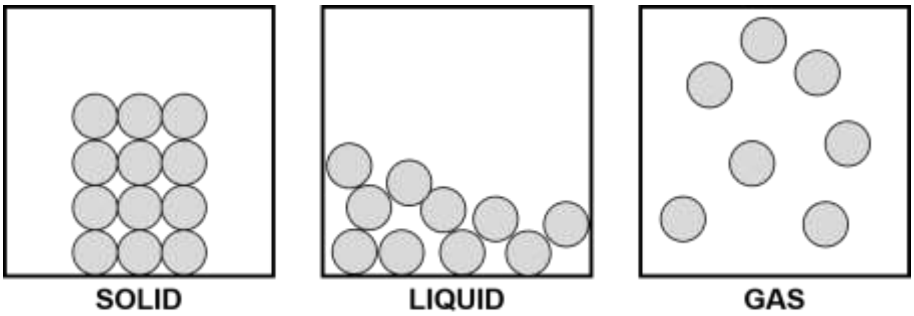
2.1.5 Metallic bonding

- Metals consist of giant structures of atoms arranged in a regular pattern.
- Electrons in the outer shell of atoms are delocalised and so are free to move through the whole structure.
- The shared 'delocalised sea of electrons' produces strong metallic bonds.
- Metallic bonding can be represented as follows:



2.2 HOW BONDING AND STRUCTURE ARE RELATED TO THE PROPERTIES OF SUBSTANCES

2.2.1 The three states of matter



- Melting → solid to liquid
- Evaporation → liquid to gas
- Condensation → gas to liquid
- Freezing → liquid to solid
- Sublimation → gas to solid (direct)

- The stronger the intermolecular forces of a substance, the higher its melting and boiling point.
- Atoms themselves do not have the bulk properties of materials.

2.2.2 State symbols

- In chemical equations the following state symbols are used:
- (s) – solid
- (l) – liquid
- (g) – gas
- (aq) – aqueous solution

- Lithium (s) + water (l) → lithium hydroxide (aq) + hydrogen (g)

2.2.3 Properties of ionic compounds

- Ionic compounds have:
    - regular structures (giant ionic lattices)
    - strong electrostatic forces of attraction in all directions between oppositely charged ions
    - high melting/boiling points (large amounts of energy needed to overcome strong electrostatic forces)

- Ionic compounds **do not** conduct electricity when **solid** (ions can only vibrate, not move)
- Ionic compounds **do** conduct electricity when **molten** or **dissolved** (ions can move)

2.2.4 Properties of small molecules

- Substances consisting of small molecules have low melting and boiling points.
- They have weak intermolecular forces.
- These **do not** conduct electricity (molecules have no overall charge).

2.2.5 Polymers

- Polymers have very large molecules.
- Atoms of molecules are linked by strong covalent bonds.
- They have strong intermolecular forces.

2.2.6 Giant covalent structures

- Substances consisting of giant covalent structures have very high melting points.
- Atoms of these structures are linked by strong covalent bonds.
- Diamond, graphite and silicon dioxide (silica) are examples of giant covalent structures.

2.2.7 Properties of metals and alloys

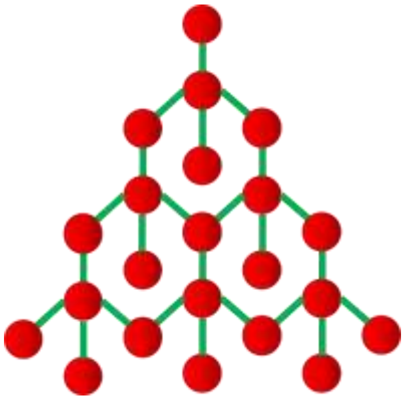
- Metals have giant structures of atoms with strong metallic bonding.
- In pure metals, atoms are arranged in layers (so are malleable and ductile).
- Pure metals are mixed with others to make harder alloys in which layers are interrupted.

2.2.8 Metals as conductors

- Metals **do** conduct electricity because the delocalised electrons can move and carry electrical charge.
- Metals **do** conduct thermal energy because energy is transferred by the delocalised electrons.

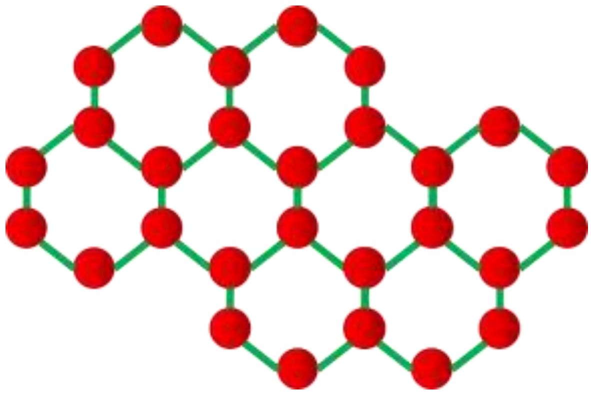
2.3 STRUCTURE AND BONDING OF CARBON

2.3.1 Diamond



- In diamond, each carbon atom forms four covalent bonds with other carbon atoms in a giant covalent structure.
- Diamond is very hard, has a very high melting point and **does not** conduct electricity.

2.3.2 Graphite



- In graphite, each carbon atom forms three covalent bonds with other carbon atoms in a giant covalent structure forming disconnected layers of hexagonal rings.
- Graphite **does** conduct electricity (has one delocalised electron per atom).

2.3.3 Graphene and fullerenes

- Graphene is a single layer of graphite having useful properties in electronics and composites.
- Graphene is made of atoms of carbon each having three covalent bonds with other carbon atoms, along with one delocalised electron per atom.

- Fullerenes are molecules of carbon atoms with hollow shapes.
- They are made with mainly hexagonal (and pentagonal/heptagonal) rings of carbon atoms.
- The first fullerene to be discovered was Buckminsterfullerene (C_{60}).

- Carbon nanotubes are cylindrical fullerenes with very high length to diameter ratios, useful in nanotechnology, electronics and materials.

2.4 BULK AND SURFACE PROPERTIES OF MATTER INCLUDING NANOPARTICLES

2.4.1 Sizes of particles and their properties

- Nanoscience refers to structures less than 100nm in size.
- Nanoparticles are smaller than fine particles (100 – 2500nm).
- Nanoparticles are smaller than coarse particles (2500nm – 10000nm)

- Nanoparticles have a high surface area to volume ratio so are different from their same materials in bulk (which have a lower SA:V).

2.4.2 Uses of nanoparticles

- Nanoparticles are used:
  - in medicine
  - in electronics
  - in cosmetics
  - in sun cream
  - as deodorants
  - as catalysts
  - (scientists are researching applications for nanoparticulate materials)

Advantages	Disadvantages
can kill bacteria	undiscovered side effects
faster catalysts due to high SA:V	might be toxic for cells
can block UV light (invisibly)	can be airborne and inhaled

3 QUANTITATIVE CHEMISTRY

3.1 CHEMICAL MEASUREMENTS, CONSERVATION OF MASS AND THE QUANTITATIVE INTERPRETATION OF CHEMICAL EQUATIONS

3.1.1 Conservation of mass and balanced chemical equations

- The law of conservation of mass states that no atoms are lost or made during a chemical reaction so:
  mass of reactants = mass of products

- Symbol equations have to be balanced so that both sides have the same number of each atom.

- E.g. $Al + O_2 \rightarrow Al_2O_3$
  becomes $\mathbf{4}Al + \mathbf{3}O_2 \rightarrow \mathbf{2}Al_2O_3$
- E.g. $Mg + HCl \rightarrow MgCl_2 + H_2$
- becomes $Mg + \mathbf{2}HCl \rightarrow MgCl_2 + H_2$

3.1.2 Relative formula mass

- The relative formula mass (M_r) of a compound is the sum of the relative atomic masses of the atoms in the formula.

- E.g. in KNO_3 (potassium nitrate):
  $M_r = 1 \times K + 1 \times N + 3 \times O$
  $M_r = 1 \times 39 + 1 \times 14 + 3 \times 16$
  $M_r = 39 + 14 + 48$
  $M_r = 101$

- In a balanced chemical equation, the sum of the relative formula masses of the reactants equals that of the products.

3.1.3 Mass changes when a reactant or product is a gas

- A non-enclosed system may involve a change in mass due to a reactant or product being a gas.
- If a reactant is a gas, the product mass will be greater than the reactant mass.
- If a product is a gas, the product mass will be less than the reactant mass.

3.1.4 Chemical measurements

- Uncertainty is given in the form: $mean \pm \frac{range}{2}$
- E.g. for the measurements: 3 5 8 8 9
- Mean = 6.6
- Range = 6
- Uncertainty = 6.6 ± 3

3.2 USE OF AMOUNT OF SUBSTANCE IN RELATION TO MASSES OF PURE SUBSTANCES

3.2.1 Moles

- A mole is a chemical unit of measurement, with the symbol mol.
- The mass of one mole of a substance is numerically equal to its relative formula mass.

- The number of atoms, molecules or ions in a mole of a given substance is the Avogrado Constant – 6.02×10^{23} per mole.
- E.g. there are 6.02×10^{23} atoms in a mole of carbon
- E.g. there are also 6.02×10^{23} molecules in a mole of carbon dioxide

number of moles = $\frac{\textbf{mass (g)}}{\textbf{molar mass}}$

- E.g. in 24g of carbon: number of moles = $\frac{24}{12}$ = 2 mol

3.2.2 Amounts of substances in equations

- Chemical equations can be interpreted in terms of moles:

- E.g. $Mg + 2HCl \rightarrow MgCl_2 + H_2$
- E.g. every **one** mole of magnesium reacts with **two** moles of hydrochloric acid to produce **one** mole of magnesium chloride and **one** mole of hydrogen gas

- <u>Calculating masses of reactants and products:</u>

Symbols:	Mg	+ 2HCl	→	$MgCl_2$	+ H_2
Masses per mole:	24	+ 36.5	→	95	+ 2
Masses:	24g	+ 73g	→	95g	+ 2g

- <u>Calculating masses in relation to a given reactant or product:</u>
  Q) If you have a solution containing 84g of carbon, what mass of oxygen gas is needed to react all the carbon to produce carbon monoxide?
  Symbols: $2C + O_2 \rightarrow 2CO$ (carbon monoxide)
  Ratio: 2 : 1 : 2

- Every **two** moles of carbon react with **one** mole of oxygen gas to produce **two** moles of carbon monoxide.
- The number of moles of carbon = 84 ÷ 12 = 7 mol.
- The required number of moles of oxygen gas is 7 ÷ 2 (according to the ratio), so 3.5 mol.
- Mass of oxygen = moles x molar mass = 3.5 x 32 = 112g.

3.2.3 Using moles to balance equations

- When given masses of reactants and/or products:
  - convert masses to moles
  - divide number of moles by smallest number to get whole numbers
  - re-multiply ratios if there are still decimals
  - insert ratios into symbol equation

- E.g. 4g of $KHCO_3$ decomposes to form 2.76g of K_2CO_3 and 0.36g of H_2O.
- Symbols: $KHCO_3$ + K_2CO_3 → H_2O
- Moles: 0.04 mol + 0.02mol → 0.02 mol
- Ratio: 2 : 1 : 1
- Balanced: $2KHCO_3$ + K_2CO_3 → H_2O

3.2.4 Limiting reactants

- In a chemical reaction, it is common to use an excess of one or more reactants to ensure all the other reactant is used.
- The limiting reactant is the reactant that is completely used up and limits the amount of product(s).

- E.g. in: $N_2 + 3H_2 \rightarrow 2NH_3$
- Ratio: 1 : 3 : 2
- If we have: 2 mol 5 mol

- There will not be enough H_2 molecules so hydrogen is the limiting reactant.
- As there are 2 moles of nitrogen and not 3 times that of hydrogen, only 1.67 moles of nitrogen will react with 5 moles of hydrogen to produce 2.5 moles of ammonia.
- The excess reactant will thereafter be 0.33 moles of hydrogen.

3.2.5 Concentration of solutions

- Concentration is the amount of solute per unit of volume of solvent.

concentration (g/dm^3) = $\dfrac{\textbf{mass (g)}}{\textbf{volume (dm}^3\textbf{)}}$

1 dm = 10 cm
1 dm^3 = 1000cm^3 = 1 litre

concentration (mol/dm^3) = $\dfrac{\textbf{number of moles}}{\textbf{volume (dm}^3\textbf{)}}$

- As concentration increases:
  - the amount of solute increases **and/or**
  - the amount of solvent decreases

3.3 YIELD AND ATOM ECONOMY OF CHEMICAL REACTIONS

3.3.1 Percentage yield

- Even though no atoms are gained or lost in a chemical reaction, it is not always possible to obtain the calculated amount of a product because:
  - reaction may be reversible (will not go to completion)
  - some product may be lost when it is separated from the reaction mixture
  - some reactants may react in unexpected ways

- Yield is the amount of product obtained.
- Percentage yield compares the actual yield with the maximum theoretical yield.

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

- To calculate the theoretical mass of a product from a given mass of reactant and a balanced equation, you need to calculate the number of moles and multiple it by the M_r.

3.3.2 Atom economy

- The atom economy (atom utilisation) is a measure of the amount of starting materials that end up as useful products.
- Atom economy is important for:
  - sustainable development
  - economic reasons

$$\textbf{atom economy} = \frac{\textbf{M}_r \textbf{ of desired product}}{\textbf{sum of M}_r \textbf{ of all reactants}} \times 100$$

- When choosing a particular reaction pathway, the following must be considered:
  - atom economy
  - yield
  - rate of reaction
  - equilibrium position
  - usefulness of by-products

3.4 USING CONCENTRATIONS OF SOLUTIONS IN MOL/DM^3

- The concentration of a solution can be measured in mol/dm^3.

- The amount in moles of solute or the mass in grams of solute in a given volume of solution can be calculated from its concentration in mol/dm^3.
- If the volumes of two solutions that react completely are known and the concentration of one solution is known, the concentration of the other solution can be calculated.

3.5 USE OF AMOUNT OF SUBSTANCE IN RELATION TO VOLUMES OF GASES

- Equal amounts in moles of gases occupy the same volume under the same conditions of temperature and pressure.

- One mole of any gas at room temperature and pressure (20 ˚C and 1 atmosphere pressure) is $24dm^3$.

- E.g.: $Ca_{(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2\,(aq)} + H_{2\,(g)}$
  Q) 0.54g of calcium reacts with hydrochloric acid to form calcium chloride and hydrogen gas. Calculate the volume of hydrogen gas produced.
  A) moles (Ca) = 0.54 / 40 = 0.0135 mol
  moles (H_2) = 0.0135 mol (1:1 ratio)
  volume (H_2) = 0.0135 mol x 24 dm^3 = 0.324 dm^3 = 324 cm^3

4 CHEMICAL CHANGES

4.1 REACTIVITY OF METALS

4.1.1 Metal oxides

- **metal + oxygen → metal oxide**
- Metals react with oxygen to produce metal oxides.

- Oxidation reaction: gain of oxygen
- Reduction: substance loss of oxygen

4.1.2 The reactivity series

- The reactivity of a metal is its tendency to form positive ions.

- Potassium	K
- Sodium	Na
- Lithium	Li
- Calcium	Ca
- Magnesium	Mg
- Aluminium	Al
- **Carbon**	**C**
- Zinc	Zn
- Iron	Fe
- Tin	Sn
- Lead	Pb
- **Hydrogen**	**H**
- Copper	Cu
- Silver	Ag
- Gold	Au
- Platinum	Pt

- The non-metals carbon and hydrogen are used to extract metals from their ores by displacement.

4.1.3 Extraction of metals and reduction

- Metals (except the unreactive ones) are found in their compounds.

- Metals less reactive than carbon can be extracted from their oxides by reduction with carbon, e.g.:

- lead oxide + carbon → lead + carbon dioxide
- $2PbO_{(s)} + C_{(s)} \rightarrow 2Pb_{(s)} + CO_{2(s)}$

- zinc oxide + carbon → zinc + carbon dioxide

- $2ZnO_{(s)} + C_{(s)} \rightarrow 2Zn_{(s)} + CO_{2(s)}$

4.1.4 Oxidation and reduction in terms of electrons

- Oxidation: loss of electrons
- Reduction: gain of electrons

Ionic equations
- Ionic equations show only the atoms and ions that change in a reaction, e.g.:

- $Mg_{(s)} + CuSO_{4(aq)} \rightarrow MgSO_{4(aq)} + Cu_{(s)}$

- remove state symbols and unchanged ions…

- $Mg + Cu \rightarrow Mg + Cu$ add the charges…

- $Mg + Cu^{2+} \rightarrow Mg^{2+} + Cu$ add the state symbols…

- $Mg_{(s)} + Cu^{2+}_{(aq)} \rightarrow Mg^{2+}_{(aq)} + Cu_{(s)}$

Half equations
- Half equations show what happens to each reactant in a reaction, e.g.:

- $Fe_{(s)} \rightarrow Fe^{2+}_{(aq)} + 2e^-$ (**oxidation** of iron – loses electrons)
- $Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-$ (**oxidation** of zinc – loses electrons)

- $Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$ (**reduction** of copper ion – gains electrons)
- $K^+_{(aq)} + e^- \rightarrow K_{(s)}$ (**reduction** of potassium ion – gains electrons)

4.2 REACTIONS OF ACIDS WITH METALS

4.2.1 Reactions of acids with metals

- **acid + metal → salt + hydrogen**
- **acid + base → salt + water**
- **acid + carbonate → salt + water + carbon dioxide**

- Metal: pure metal, e.g. sodium, magnesium, iron
- Base: insoluble metal oxide or hydroxide, e.g. copper oxide
- Alkali: soluble metal oxide or hydroxide, e.g. sodium hydroxide
- Carbonate: metal carbonate, e.g. calcium carbonate

- Hydrochloric acid is HCl and makes Cl^- ions
- Sulfuric acid is H_2SO_4 and makes SO_4^{2-} ions.
- Nitric acid is HNO_3 and makes NO_3^- ions.

- magnesium + hydrochloric acid → magnesium chloride + hydrogen
- magnesium + sulfuric acid → magnesium sulfate + hydrogen

- $Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$
- $Zn_{(s)} + 2HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$
- $Fe_{(s)} + 2HCl_{(aq)} \rightarrow FeCl_{2(aq)} + H_{2(g)}$

- $Mg_{(s)} + H_2SO_{4(aq)} \rightarrow MgSO_{4(aq)} + H_{2(g)}$
- $Zn_{(s)} + H_2SO_{4(aq)} \rightarrow ZnSO_{4(aq)} + H_{2(g)}$
- $Fe_{(s)} + H_2SO_{4(aq)} \rightarrow FeSO_{4(aq)} + H_{2(g)}$

- Reactions of metals with acids are redox reactions as the metal is oxidised and the hydrogen is reduced.
- This redox reaction happens because the metal is more reactive than hydrogen so displaces it from its compound (the acid) to form a salt.

4.2.2 Neutralisation of acids and salt production

- Acids are neutralised by alkalis (soluble) and bases (insoluble).

- **acid + carbonate → salt + water + carbon dioxide**

- hydrochloric acid + calcium carbonate → calcium chloride + water + carbon dioxide
- sulfuric acid + sodium carbonate → sodium sulfate + water + carbon dioxide

- $2HCl_{(aq)} + Na_2CO_{3(s)} \rightarrow 2NaCl_{(aq)} + H_2O_{(l)} + CO_{2(g)}$
- $H_2SO_{4(aq)} + CaCO_{3(s)} \rightarrow CaSO_{4(aq)} + H_2O_{(l)} + CO_{2(g)}$

- The formula of a salt depends on its ions, e.g.:
    - sodium chloride is made of Na^+ and Cl^-, so it is $NaCl$
    - sodium sulfate is made of Na^+ and SO_4^{2-}, so it is Na_2SO_4
    - magnesium nitrate is made of Mg^{2+} and NO_3^-, so it is $Mg(NO_3)_2$

4.2.3 Soluble salts

- How to make pure, dry samples of soluble salts:
    - heat acid to increase rate of reaction
    - add solid (metal/metal oxide/metal hydroxide/metal carbonate) to acid until no more reacts to ensure all acid has reacted
    - filter excess solid using filter paper and funnel
    - pour salt solution into evaporating basin
    - heat solution with Bunsen burner until crystals begin to form
    - leave the rest to evaporate at room temperature (for about 1 day)
    - dry crystals using paper/napkin

4.2.4 The pH scale and neutralisation

- Acids produce hydrogen (H^+) ions in aqueous solution.
- Alkalis produce hydroxide (OH^-) ions in aqueous solutions.

- The pH scale (0-14) is a measure of the acidity or alkalinity of a solution:

pH scale														
← more acidic ←							neutral	→ more alkaline →						
0	1	2	3	4	5	6	7	8	9	10	11	12	13	14

- pH can be measured using:
    - universal indicator
    - pH probe

  detect concentration of H^+ ions

- Ionic equation for neutralisation:
    - $H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$

4.2.5 Titrations

- Titrations are a way to measure the volumes of acid and alkali solutions that react together using universal indicator.

- To carry out a titration for neutralisation:
  - add a measured volume of alkaline solution to a conical flask and place it on a white tile
  - add a few drops of phenolphthalein indicator
  - fill the burette with acid and record starting volume
  - slowly add the acid to the solution until the indicator changes colour
  - record the volume of acid added to reach the end-point
  - repeat titre adding alkaline solution dropwise as you near the previous result to obtain a more accurate reading
  - now you can repeat the experiment without universal indicator to obtain a pure neutralised sample of the initial alkaline solution

Q) 25 cm^3 of sodium hydroxide solution is neutralised by 23.4 cm^3 of 0.0998 mol/dm^3 hydrochloric acid solution. Work out the concentration of the sodium hydroxide solution.

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

moles (HCl) = ($23.4cm^3/1000$) dm^3 x 0.998 mol = 0.0233532 mol

moles (NaOH) = 0.0233532 mol (1:1 ratio)

concentration = 0.0233532 mol / (25 $cm^3/1000$) dm^3 = 0.934 mol/dm^3

4.2.6 Strong and weak acids

- Strong acid: completely ionised in an aqueous solution.
- E.g. hydrochloric acid, nitric acid, sulfuric acid.

- Weak acid: only partially ionised in an aqueous solution.
- E.g. ethanoic acid, citric acid, carbonic acid.

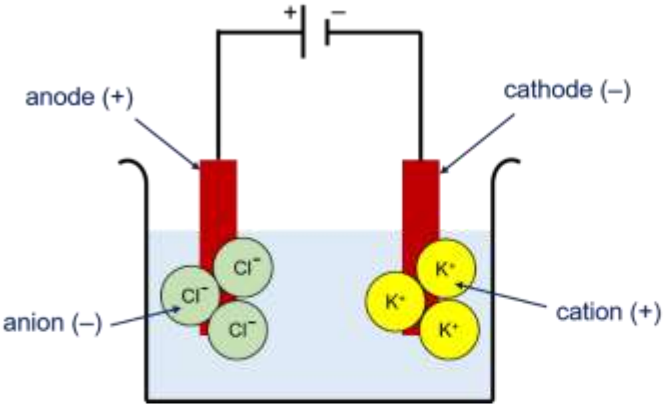
- Dilute: an acid with a low number of acid particles per unit of volume.
- Concentrated: an acid with a high number of acid particles per unit of volume.

- For a given concentration of aqueous solutions, the stronger the acid, the lower the pH.
- An acid of pH 0-3 is stronger than that of pH 4-6.

- As pH increases by one unit, hydrogen ion concentration increases by one order of magnitude.
- E.g. pH 0 at 1 mol/dm^3
- increase pH by 1 (/10): pH 1 at 0.1 mol/dm^3
- increase pH by 3 (/1000): pH 4 at 0.0001 mol/dm^3

4.3 ELECTROLYSIS

4.3.1 The process of electrolysis



- An electrolyte is a molten or dissolved ionic compound that is able to conduct electricity because its ions can move.

- When an electric current is passed through the electrolytes, the ions move to the corresponding electrodes.
- Positively charged ions (cations) move to the negative electrode (cathode).
- Negatively charged ions (anions) move to the positive electrode (anode).
- Electrons are transferred from the anions to the cations to produce neutral elements at the electrodes.
- In this example, solid potassium forms on the cathode and chlorine gas is produced at the anode.

- $2Cl^-_{(aq)} + 2e^- \rightarrow Cl_{2\,(g)}$ (half equation for chlorine ion)
- $K^+_{(aq)} - e^- \rightarrow K_{(s)}$ (half equation for potassium ion)

4.3.2 Electrolysis of molten ionic compounds

- When a simple ionic compound (e.g. lead bromide) is electrolysed in the molten state using inert electrodes:
  - the metal (lead) is produced at the cathode
  - the non-metal (bromine) is produced at the anode

4.3.3 Using electrolysis to extract metals

- Electrolysis is used to extract metals that are more reactive than carbon.
- It is an expensive process because it requires large amounts of energy to melt the compounds and produce an electric current.

- Electrolysis of aluminium:
  - aluminium oxide → aluminium + oxygen
  - $2Al_2O_{3(l)} \rightarrow 4Al_{(l)} + 3O_{2(g)}$

  - cryolite is mixed with aluminium oxide to lower the melting point
  - the anode is made of carbon
  - Al^{3+} ions are attracted to the cathode
  - O^{2-} ions are attracted to the anode
  - at the anode, the O_2 react with the carbon anode to produce CO_2 gas

4.3.4 Electrolysis of aqueous solutions

- sodium chloride solution → hydrogen + chlorine + sodium hydroxide solution
- $2NaCl_{(aq)} + 2H_2O_{(l)} \rightarrow H_{2(g)} + Cl_{2(g)} + 2NaOH_{(aq)}$

- In electrolysis of brine (sodium chloride solution):
  - H^+ and OH^- ions are produced by electrolysing the water

  - H^+ ions are attracted to the cathode, where they are reduced (gain one electron) to make diatomic molecules of H_2 gas
  - $2H^+_{(aq)} + 2e^- \rightarrow H_{2(g)}$

  - Cl ions are attracted to the anode, where they are oxidised (lose one electron) to make diatomic molecules of Cl_2 gas
  - $2Cl^-_{(aq)} \rightarrow Cl_{2(g)} + 2e^-$

  - Na^+ ions are more reactive than H^+ ions, and Cl^- ions are more reactive than OH^- ions so they remain in solution as aqueous ions
  - In the solution, the Na^+ and OH^- ions react to form sodium hydroxide solution ($NaOH_{(aq)}$)

- At the anode either oxygen or a halogen gas is produced.
- At the cathode hydrogen is produced if the metal is more reactive than hydrogen.

4.3.5 Representation of reactions at electrodes as half equations

- At the anode, positively charged ions gain electrons by reduction.
- At the cathode, negatively charged ions lose electrons by oxidation.

- At the anode: $2H^+_{(aq)} + 2e^- \rightarrow H_{2(g)}$
- At the cathode: $2Cl^-_{(aq)} \rightarrow Cl_{2(g)} + 2e^-$ **or** $2Cl^-_{(aq)} - 2e^- \rightarrow Cl_{2(g)}$

5 ENERGY CHANGES

5.1 EXOTHERMIC AND ENDOTHERMIC REACTIONS

5.1.1 Energy transfer during exothermic and endothermic reactions

- Exothermic reactions transfer energy to the surroundings so the temperature of the surroundings increases.
- Endothermic reactions take in energy from the surroundings so the temperature of the surroundings decreases.

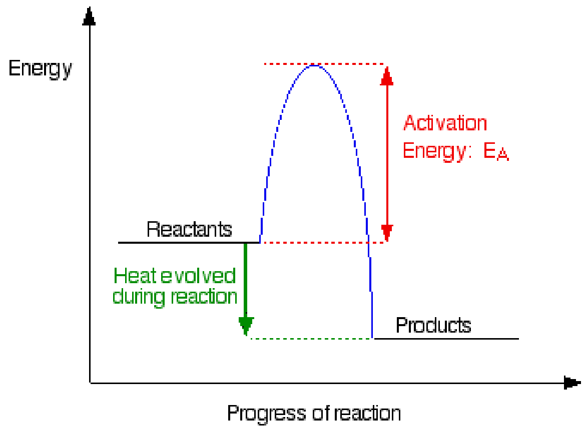
- Exothermic reactions include:
  - combustion
  - oxidation
  - neutralisation

- Endothermic reactions include:
  - thermal decomposition
  - citric acid and sodium hydrogencarbonate
  - photosynthesis
  - sports injury packs

5.1.2 Reaction profiles

- Chemical reactions can only occur when reacting particles collide with each other and with sufficient energy.
- The activation energy, E_a, is the minimum amount of energy particles must have to react.

- Reaction profiles show the relative energies of reactants and products, the activation energy and the enthalpy change (ΔH) of a reaction.



5.1.3 The energy change of reactions

- Breaking bonds **takes in** energy.
- Making bonds **releases** energy.

- Bond energies are measured in kJ/mol and you will be supplied with the appropriate bond energies in your exam.

- <u>Bond energy calculation example:</u>
- Hydrogen + chlorine → hydrogen chloride
- $H_2 + Cl_2 \rightarrow 2HCl$

- *Bond breaking:* 1 x (H—H) + 1 x (Cl—Cl)
- substitute energies: 436 + 243
- calculate: 679 kJ/mol

- *Bond making:* 2 x (H—Cl)
- substitute energies: 2 x 432
- calculate: 864 kJ/mol

- Energy change = bond breaking – bond making
- $\Delta H = 679 - 864$
- $\Delta H = -185$ kJ/mol ← As this is negative, the reaction has lost energy so it is exothermic (releases 185kJ/mol)

5.2 CHEMICAL CELLS AND FUEL CELLS

5.2.1 Cells and batteries

- Cells contain chemicals which react to produce electricity

- A cell is made by connecting two different metals in contact with an electrolyte.
- Batteries consist of two or more cells connected together in series to provide a greater voltage.

- Voltage produced by a cell is dependent upon:
  - **type of electrode:** the greater the difference in reactivity, the greater the voltage
  - **type of electrolyte:** the greater the concentration, the more ions it carries, and the greater the voltage

- In non-rechargeable cells/batteries:
  - chemical reactions stop when one reactant is used up
  - the battery is disposed of

- In rechargeable cells/batteries:
  - an external current reverses the chemical reactions
  - the battery can be used again

Advantages	Disadvantages
portable power supply	non-rechargeable cells disposed of
reusable rechargeable batteries	lots of metal used to make them
reduce pollution (electric cars)	non-renewable energy

5.2.2 Fuel Cells

- Fuel cells are supplied by an external source of fuel (e.g. hydrogen) and oxygen or air.
- The fuel is electrochemically oxidised within the fuel cell to produce a potential difference.
- The overall reaction in a hydrogen fuel cell involves the oxidation of hydrogen to produce water.
- Hydrogen fuel cells offer a potential alternative to rechargeable cells and batteries.

- $2H_{2\,(g)} + O_{2\,(g)} \rightarrow 2H_2O_{\,(l)}$

- Half equations:
  - $2H_2 + 4OH^- \rightarrow 4H_2O + 4e^-$
  - $O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$

Hydrogen fuel cells	Rechargeable batteries
+ no CO2 emissions	+ reusable + cheaper
- hydrogen gas is explosive - needs constant supply	- capacity degrades over time - uses fossil-fuel consuming electricity

CHEMISTRY PAPER 1

1	ATOMIC STRUCTURE AND THE PERIODIC TABLE
2	BONDING, STRUCTURE, AND THE PROPERTIES OF MATTER
3	QUANTITATIVE CHEMISTRY
4	CHEMICAL CHANGES
5	ENERGY CHANGES