Summer Examination Revision

These 3 subatomic particles are fundamental

Particle	Relative charge	Relative Mass	
Proton	+/	/	
Neutron	0	/	
Electron	-/	Nearly 0 (0.0005)	Neutron

Historical Development of the Atom

- **1. Democritus** Ancient Greek philosopher who suggested that matter cannot be cut into smaller pieces.
- 2. **Dalton**: Proposed the first **atomic theory**, suggesting that atoms are indivisible particles that combine in fixed ratios to form compounds.
- 3. **Crookes**: Investigated the properties of **cathode rays** using vacuum tubes, laying the groundwork for the discovery of the electron.
- 4. Stoney: Named the electron as the fundamental unit of electric charge.
- 5. **Thomson**: Discovered that cathode rays were composed of negatively charged particles, which he identified as **electrons**. He also measured the charge-to-mass ratio (**e/m**) of the electron.



- 6. Millikan: Measured the magnitude of the charge on the electron through the oil drop experiment.
- 7. **Rutherford**: Conducted gold foil **experiment**, leading to the discovery of the **nucleus** and proving that most of the atom is empty space. He also discovered **protons** in the nucleus.



- Most particles (alpha positively charged) went straight through = Empty space



- Some bounced back along original pathway = Hit dense concentrated positive nucleus
- Small amount deflected a @ large angles = Came close to nucleus
- 8. **Bohr**: Proposed a model where electrons move in fixed orbits around the nucleus with specific energy levels.
- 9. Chadwick: Discovered the neutron, completing the model of the atom.

Isotopes and Atomic Mass

Isotopes are atoms of the same element with the same **atomic number** but different **mass numbers** due to varying numbers of neutrons.

• Example: Carbon-12 and Carbon-14 are both isotopes of carbon, differing in neutron number.

Relative Atomic Mass (Ar):

- Defined as the weighted average of the mass numbers of all naturally occurring isotopes of an element, based on the **carbon-12** scale.
- Example: The relative atomic mass of chlorine is approximately 35.5 due to the presence of isotopes chlorine-35 and chlorine-37 in different natural abundances.



= 24.32g/mol

Mass Spectrometry

Mass spectrometry is a technique used to determine the **relative atomic mass** of elements by measuring the mass of their ions. The process involves several key stages:

- 1. Vaporisation: The sample is vaporised.
- 2. Ionisation: Atoms in the sample are converted into positive ions.
- 3. Acceleration: These ions are accelerated in an electric field.
- 4. **Separation**: The ions are separated based on their mass-to-charge ratio in a magnetic field.
- 5. **Detection**: The ions are detected, and their abundance is measured.

Electron

Bohr's Theory: Electrons occupy specific energy level**s** around the nucleus. He tested this theory using an instrument called:

Emission Line Spectrum:

- Electrons in ground state absorb energy.
- When they absorb energy, they become excited.
- They jump to a higher level (excited state)

- When electrons fall back down to lower energy levels, they emit energy in the form of a light (photon).

- Electrons release light of specific frequencies when falling between energy levels, producing line spectra.

- This supports the idea of quantised energy levels.

Another instrument that supports energy levels is:

Atomic Absorption Spectrometry (AAS): Measures the concentration of elements by detecting the light they absorb at specific wavelengths.



Both instruments are good for use in the chemical industry detecting unknown substances.



AAS: Used in water and blood analysis for metals like lead and mercury.

ELS: Used in astrophysics, identifying metals, flame test.

*Emission Line Spectrum releases energy in form of a light. Atomic absorption spectrometry absorbs energy in form of a light.

Balmer Series

This series is part of the hydrogen emission spectrum, where electrons fall to the n = 2 energy level in a hydrogen atom.

This results in the emission of visible light (bright lines at specific wavelengths).

It provides evidence for quantized energy levels in atoms — a key idea in atomic theory.

Limitations to Bohrs theory

- i. Bohrs theory worked to explain Hydrogen in emission line spectrum but did not account for atoms with more than one electron.
- ii. Did not consider electron wave motion.
- iii. Heisenberg's principle showed that finding an electron in a fixed position is a probability not a certainty.
- iv. Bohr did not consider sublevels with the splitting of certain lines in emission spectra.



Heisenberg's Uncertainty Principle

Heisenberg's uncertainty principle states that it is impossible to measure the position and velocity of an electron at the same time.

NB Wave Nature of the Electron:

Electrons exhibit **wave-particle duality**, meaning they have properties of both particles and waves.

Atomic Orbitals

An atomic orbital is a region of space where it is most likely to find an electron.

Properties of Atomic Orbitals

It is required that you know the shapes of the s and p orbitals.



Each orbital can hold 2 electrons. This follows Paulie's exclusion principle where he stated that no more than two electrons can occupy an orbital and when they do, they must have opposite spin.

Aufbau principle: electrons occupy the lowest energy level available when the electron is in the ground state.

The s orbitals are spherical there is only one s orbital for each sublevel.

The p orbitals have a dumbbell shape. There are three p orbitals for each sub level. They are called the px, py and the pz orbitals.



These p orbitals are at right angles to each other as shown in the diagram. It is important to note that any p sublevel can accommodate 6 electrons.

There are five d orbitals. It is not required that you know the names or shapes of these orbitals. It is important to note that the d sublevel can accommodate 10 electrons.

Energy Levels and Sub-levels:

The first 36 elements have electrons arranged in main energy levels (n = 1, 2, 3, etc.), which are further divided into sub-levels (s, p, d, f). The sub-levels are filled in a specific order of increasing energy.

- 1st Energy Level: Contains 1s (holds a maximum of 2 electrons).
- 2nd Energy Level: Contains 2s (2 electrons) and 2p (6 electrons).
- **3rd Energy Level**: Contains 3s (2 electrons), 3p (6 electrons), and part of 3d (but 3d is not filled until after 4s).
- **4th Energy Level**: Contains 4s (2 electrons) and part of 4p (starts being filled from element 31 onwards).

The order of filling is based on increasing energy, and follows this sequence:

 $1s \rightarrow 2s \rightarrow 2p \rightarrow 3s \rightarrow 3p \rightarrow 4s \rightarrow 3d \rightarrow 4p.$

Electronic Configuration of the First 36 Elements:

Here is how the electrons are arranged in the first 36 elements (with element numbers and names included for reference):

- 1. Hydrogen (H): 1s¹
- 2. **Oxygen (O)**: $1s^2 2s^2 2p^4$
- 3. Fluorine (F): $1s^2 2s^2 2p^5$
- 4. Ca (Ca)_____

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The Mole & Stoichiometry

Avogadro's Constant

- **Avogadro's constant:** 6 x 10²³, representing the number of particles (atoms, molecules, or ions) in one mole of any substance.
- Particles may be atoms or molecules
- The mole (mol) is the SI unit for the amount of substance.
- One mole contains exactly 6 x 10²³ Avogadro's number of particles.
- Relative molecular mass (Mr): The sum of the relative atomic masses of all atoms in a molecule compared to **one twelfth** of the mass of a carbon-12 atom.

Calculating the number of particles of a substance:

How many atoms are present in 2 mol of C

How many molecules are present in 2 mol of O_2

How many atoms are present in 2 mol of O_2





Empirical and Molecular Formulas

Empirical Formula: The simplest ratio of atoms in a compound. It shows the relative number of atoms of each element. For example, the empirical formula of hydrogen peroxide is HO.

Molecular Formula: This shows the actual number of atoms of each element in a molecule. For hydrogen peroxide, the molecular formula is H2O2, which indicates two hydrogen atoms and two oxygen atoms.

Calculating Empirical Formulas

- 1. If percentage is given, use as mass in grams.
- 2. Convert grams to moles by dividing by the atomic mass of each element.
- 3. Find the simplest ratio by dividing the number of moles of each element by the smallest number of moles calculated.
- 4. Adjust the ratio to whole numbers if needed.

Example 1: A compound contains 75g Carbon and 25g Hydrogen. What is the empirical formula?

This is your percentages	You work this out by dividing by the RAM					
	Carbon	Hydrogen				
Amount	75	25				
Moles 🖌	75/12 = 6.25	25/1 = 25				
Calculate smallest ratio 🔒	6.25 / 6.25	25/6.25				
Smallest ratio	1	4				
Empirical Formula	CH ₄					

You work this out by dividing by the smallest number of moles



Example 2: Determine the molecular formula of a compound whose composition by mass is carbon 82.66%, hydrogen 17.34% and whose relative molecular mass is 58.

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Stoichiometric calculations from balanced equations

Calculating moles using mass and Mr:

Example 1:

Magnesium reacts with oxygen to produce magnesium oxide according to the equation:

 $2Mg \ + \qquad O_2 \quad \rightarrow \ 2MgO$

If a student burns 9 g of magnesium in oxygen to ensure all of the magnesium reacts. Calculate the number of moles in magnesium oxide.







Calculating mass from moles and Mr:

Example 2:

Magnesium reacts with oxygen to produce magnesium oxide according to the equation:

$2Mg \ + \qquad O_2 \quad \rightarrow \ 2MgO$

If a student burns 9 g of magnesium in oxygen to ensure all of the magnesium reacts. What mass of magnesium oxide will be formed?



Calculating Mr from moles and mass:

A chemist heated 3.22 g of hydrated sodium sulphate crystals, $Na_2SO_4.xH_2O$ until no further loss in mass occurred. It was found that <u>1.42</u> g of anhydrous salt was obtained. Calculate the formula of the hydrated salt. (H = 1, O = 16, Na = 23, S = 32)

 $Na_2SO_4.xH_2O \quad + \quad Na_2SO_4 \quad \rightarrow \ xH_2O$



Calculating volume from moles and stp:

Carbon Dioxide is formed when carbon is burned in air according to the equation:

 $C \quad + \quad O_2 \quad \rightarrow \ CO_2$

What volume of carbon dioxide (at s.t.p) is formed when 20 g of carbon are burned in excess oxygen?

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Leaving Cert Exam-Style Question

When calcium carbonate (**CaCO**₃) reacts with hydrochloric acid (**HCl**), calcium chloride, water, and carbon dioxide gas are produced:

 $CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$

A student reacts **10.0 g of calcium carbonate (CaCO₃)** with excess hydrochloric acid.

(a) Calculate the number of moles of calcium carbonate used.

(b) Using the balanced equation, calculate the mass of calcium chloride (CaCl₂) formed.

(c) Calculate the volume of carbon dioxide gas produced at STP.

(d) Calculate the number of molecules of carbon dioxide produced.

Chemical Bonding

Ionic Bonding - The transfer of electrons

- Usually formed between metals and non-metals.
- Metals lose electrons to form positive ions, while non-metals gain electrons to form negative ions.

An ion is formed when atoms have either lost or gained electrons.

- A positive ion (cation) has lost one or more electrons.
- A negative ion (anion) has gained one or more electrons.

Dot Cross diagram NaCl

Characteristics of Ionic compounds

- 1. Conduct electricity when molten or in solution as their ions are free to move.
- 2. Have high melting and boiling points due to attraction between ions
- 3. Most ionic substances dissolve in water
- 4. Exist as solids at room temperature

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Covalent Bonding - Sharing of electrons

- Formed usually between non-metals.
- Atoms share electrons to complete their outer shells.

Dot Cross diagram of Cl₂

Characteristics of Covalent Substances

- 1. Low Melting and Boiling Points.
- 2. Intermolecular forces (e.g., van der Waals or dipole-dipole) between molecules are relatively weak.
- 3. Do not conduct electricity because they lack free-moving charged ions.
- 4. Generally insoluble in water (if non-polar) but soluble in non-polar solvents (e.g., oils, benzene). Polar covalent substances, however, may dissolve in water.

Electronegativity and Prediction of Bond Type

Electronegativity values help determine whether a bond is non-polar (pure) covalent, polar covalent, or ionic

• If electronegativity difference is zero, the bond is non-polar covalent (pure covalent), with electrons shared equally between the atoms.



- If electronegativity difference is between 0.4 and 1.7, the bond is polar covalent, where electrons are shared unequally, creating a dipole (e.g., HCl).
- If the electronegativity difference is greater than 1.7, the bond is ionic, where electrons are fully transferred from one atom to another (e.g., NaCl).

	Electronegativity values of the elements (Pauling scale)																
H		Electronegativity values of the elements (Pauling scale)											He				
2.1	R.												Na				
1.0	1.5												140				
Na	Ma													Ar			
0.9	1.2		1.5 1.8 2.1 2.5 3.0														
К	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
0.8	1.0	1.3	1.5	1.6	1.6	1.5	1.8	1.8	1.8	1.9	1.6	1.6	1.8	2.0	2.4	2.8	3.0
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	Т	Xe
0.8	1.0	1.2	1.4	1.6	1.8	1.9	2.2	2.2	2.2	1.9	1.7	1.7	1.8	1.9	2.1	2.5	2.6
Cs	Ba	La	Hf	Та	W	Re	Os	lr	Pt	Au	Hg	Ti	Pb	Bi	Po	At	Rn
0.7	0.9	1.1	1.3	1.5	1.7	1.9	2.2	2.2	2.2	2.4	1.9	1.8	1.8	1.9	2.0	2.2	2.4
Fr	Ra	Ac															
0.7	0.7	1.1															
Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu				
1.1	1.1	1.1	1.1	1.1	1.1	1.1	1.1	1.1	1.1	1.1	1.1	1.1	1.2				
Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr				
1.3	1.5	1.7	1.3	1.3	1.3	1.3	1.3	1.3	1.3	1.3	1.3	1.3					

Pauling Scale

Pauling scale can be used to identify type of bonding. (Located in log tables!)

• Na—Cl: Na = 0.9 Cl = 3.0 Electronegativity difference 3.0 – 0.9 = 2.1

Greater than 1.7 ∴ Ionic

• C—O: C = 2.5 O = 3.5 Electronegativity difference 3.5 – 2.5 = 1

Less than 1.7, Greater than 0.4 : Polar Covalent

• Cl—Cl Cl = 3.0 Cl = 3.0 Electronegativity difference 3.0 - 3.0 = 0

Equals 0 : Pure Covalent

Polarity

Polar Covalent Bonding

Two atoms share electrons unequally due to a difference in electronegativities.

This causes a partial positive charge on one atom and a partial negative charge on the other.



This separation of opposite charges in a molecule is called a dipole moment.



Pure (Non-Polar) Covalent Bonding

Type of bond where two atoms share electrons equally due to having identical or very similar electronegativities.

Resulting in no partial charges.

Example: Chlorine (Cl ₂)
Identical atoms, identical electronegativities.
Electrons do not favour either atom Shared evenly.

Polar and Non-Polar Molecules

Some molecules may have polar bonds but overall are a non-polar molecule.

Polar Molecules



Will be polar overall if one end is partially positive and one end is partially negative



Non-Polar Molecules

Will be non-polar overall if there's no distinct separation of charges.

If one end is not partially negative and other end is not partially positive



Polarity Experiment demonstration



Polarity and Solubility

Always think.... LIKE DISSOLVES LIKE.

Water is a polar molecule and is used as a main solvent.

∴ To dissolve in water, substances must also be **polar** or **ionic.** (Have charges)

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Example:

Sodium chloride (NaCl) – Ionic Water is polar and pulls apart the positive and negative ions, dissolving them.



Ammonia (NH3)- Polar Covalent.

These are polar covalent molecules, so they can form hydrogen bonds or interact well with water.



Iodine (I2) – Pure (Nonpolar) Covalent

As iodine is nonpolar i.e. no charges or partial charges, it will not dissolve in water.

 \therefore Nonpolar molecules will not dissolve in water or other polar molecules.

However, non-polar solvent cyclohexane will dissolve Iodine as both are non-polar.



To note:

Na⁺Cl[−] ^{δ+}H−Cl^{δ−} Cl−Cl ↑ Polar covalent bonding (partial charges) Non-polar covalent bonding (electronically symmetrical)

(full charges) (electronically symmetrical)

Bonding should be viewed as more of a spectrum as seen below:

We use tools provided to estimate the type of attraction between atoms.

*Remember, ionic always full chargers, polar covalent always partial charges, pure covalent no charges.

VSEPR (Valence Shell Electron Pair Repulsion) Theory

- Electron **pairs** around the central atom repel each other and position themselves as far apart as possible.
- Two types of electron pairs:
 - **Bond pairs** (involved in bonding).
 - Lone pairs (not involved in bonding).
- Lone pair has greater repulsion than bonded pair.
- The shape of the molecule is determined by:
 - 1. The number of **electron pairs** around the central atom.
 - 2. The type of **electron pairs (lone or bond pairs).**

<u>5 shapes you need to know</u>

<u>AX</u>2

Central atom A

Two atoms attached X₂

 Linear Bond angle 180° No lone pairs on central atom

Н — Ве — Н

0 = c = 0

2. Non-LinearBond angle 104.5°Lone pairs on central atom

н~°;

<u>AX3</u>

Central atom A

Three atoms attached X₃

- Trigonal Planar
 Bond angle 120°
 No lone pairs on central atom
- Pyramidal
 Bond angle 107°
 Lone pairs on central atom

<u>AX4</u>

Central atom A

Four atoms attached X₄

 Tetrahedral Bond angle 109.5° No lone pairs on central atom



*Notice how the lone pairs distort the shape. ie pushes the bonds further from central atom.

Van der Waals Forces

- Arise due to temporary dipoles.
- Weak forces easily separated.

Example: Hydrogen molecules (H2).

Two electrons are shared equally between two identical hydrogen atoms.

At any time, the shared electrons might move closer to one end of the molecule than the other.

Results in forming a temporary dipole in this molecule and induced (forced) dipole in neighbouring molecules. Van der Waals



Permanent Dipole-Dipole Interactions

- Attraction between positive and negative poles of molecules.
- Stronger than Van der Waals forces.
- \therefore higher boiling point than a non-polar molecule (Van der Waals.

Example: Hydrogen chloride (HCl).

Two electrons are shared unequally in a polar covalent bond.

Chlorine has a higher electronegativity than hydrogen.

 $\therefore\,$ has a greater pull on the electrons.

The chlorine molecule carries a partial negative charge, the hydrogen end carries a partial positive charge.

Attraction between these charges is called **dipole-dipole** attraction.

 $\begin{array}{c|c} \delta + & \longrightarrow & \delta - \\ H & \longrightarrow & CI \end{array} \begin{array}{c} \delta + & \longrightarrow & \delta - \\ H & \longrightarrow & CI \end{array}$

Hydrogen Bonding

- Special dipole-dipole interaction involving hydrogen bonded to fluorine, oxygen, or nitrogen.
- Strong due to the significant difference in electronegativity.

Example: Water (H2O).

- Oxygen is more electronegative than hydrogen.
- : greater share in the shared pair of electrons.
 - Oxygen is partially negative, and hydrogen is partially positive.



Effect of intermolecular forces on boiling points

- Strength of intermolecular forces influences boiling points.
- The stronger the attractive forces, the more energy required to break bonds, the higher the boiling point.
- Hydrogen bonding > Dipole-dipole interactions > Van der Waals forces.

*Remember: If asked to compare boiling point of two compounds always mention intermolecular forces!!

Examples:

$1.\,H_2\,vs.\,O_2$

- Boiling Point: $O_2 > H_2$
- Reason:

 O_2 molecules are heavier and have stronger van der Waals (dispersion) forces than lighter H_2 molecules.

 \blacktriangleright So, O₂ needs more energy (higher temperature) to boil.

2. C₂H₄ vs. CH₂O

- Boiling Point: $CH_2O > C_2H_4$
- Reason:
 - \circ CH₂O has a permanent dipole due to the polar C=O bond → stronger dipole-dipole forces.
 - C₂H₄ is nonpolar, only has weak van der Waals forces.
 ➤ So, CH₂O has a higher boiling point.

3. H₂O vs. H₂S

- Boiling Point: $H_2O > H_2S$
- Reason:
 - $_{\odot}$ $\,$ H_{2}O forms hydrogen bonds, which are very strong.
 - \circ H₂S is less polar and doesn't form hydrogen bonds.
 - ► Water boils at a much higher temperature.

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Sigma and Pi Bonding

Sigma (σ) Bonding

- Forms when two atomic orbitals overlap head-on.
- Sigma bonds are stronger than pi bonds due to greater overlap of orbitals.
- Found in all single covalent bonds.
- In double and triple bonds, one of the bonds is always a sigma bond.

Pi (π) Bonding

- Forms when two atomic orbitals overlap sideways, above and below the axis of the bond.
- Pi bonds are weaker than Sigma bonds because the sideways overlap is less extensive.
- Pi bonds occur in double and triple bonds. In a double bond, there is one sigma bond and one pi bond.
- In a triple bond, there is one sigma bond and two pi bonds.



Other important topics to consider:

Trends in Periodic Table

1. Definitions & Terms

Be sure you understand and can define:

- Atomic radius
- Ionisation energy
- Electronegativity
- Electron affinity (may come up in context but not as commonly)
- Shielding effect
- Effective nuclear charge

2. Trends Across a Period (Left to Right)

As you move across a period (e.g., from Na to Cl in Period 3):

- Atomic radius decreases: Increased nuclear charge pulls electrons closer.
- Ionisation energy increases: More energy is needed to remove an electron.
- Electronegativity increases: Atoms attract bonding electrons more strongly.
- Metallic character decreases: Elements become more non-metallic.

3. Trends Down a Group (Top to Bottom)

As you move down a group (e.g., Group 1 or Group 7):

- Atomic radius increases: More electron shells are added.
- Ionisation energy decreases: Outer electrons are further from the nucleus and more shielded.
- Electronegativity decreases: Atoms attract electrons less strongly.
- Reactivity changes:
 - Group 1 (alkali metals): Reactivity increases down the group.
 - Group 7 (halogens): Reactivity decreases down the group.

4. Specific Examples to Know

Be familiar with:

- Trends in atomic radius using Na, Mg, Al, Si, P, S, Cl.
- Trends in first ionisation energies across Period 2 and 3.
- Comparison of alkali metals (Group 1) and halogens (Group 7) reactivity.
- How noble gases differ (e.g., high ionisation energies, very low reactivity).

5. Explanations Required

You must be able to explain trends, not just state them. For example:

- Why ionisation energy increases across a period: "Due to increased nuclear charge and no additional shielding, electrons are held more tightly."
- Why atomic radius increases down a group: "Due to the addition of extra energy levels and increased shielding."
- 6. Graphs and Data Interpretation

You may be given:

- Graphs of atomic radius or ionisation energy vs. atomic number.
- Asked to interpret trends or identify anomalies (e.g., drop in ionisation energy between Be and B, or N and O).

7. Past Exam Questions

Practice with real Leaving Cert exam papers:

- Definitions
- Trends and explanations
- Short questions with graphs
- Full-length questions with structured parts

Radioactivity - Key Points

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1. Definitions to Know

Be able to define these precisely, as they often come up as short questions:

- Radioactivity: The spontaneous emission of radiation from the nucleus of an unstable atom.
- Alpha particle (a): A helium nucleus (2 protons, 2 neutrons) positively charged.
- Beta particle (β): A high-speed electron negatively charged.
- Gamma ray (γ): Electromagnetic radiation no charge, high penetration.
- Half-life: The time it takes for half of the nuclei in a sample of a radioactive isotope to decay.
- Isotope: Atoms of the same element with the same number of protons but different numbers of neutrons.
- Nuclear reaction: A reaction that alters the composition, structure, or energy of an atomic nucleus.

2. Types of Radiation

Know the properties and differences between alpha, beta, and gamma radiation:

Туре	Symbol	Charge	Penetration	Stopped by	Effect on Nucleus
Alpha	a or He ²⁺	+2	Low	Paper / skin	Mass -4, Atomic -2
Beta	β-	-1	Moderate	Thin aluminium	Mass same, Atomic +1
Gamma	Ŷ	0	High	Thick lead / concrete	No change

3. Detection of Radiation

- Geiger-Müller (GM) tube: Used to detect radiation.
- Understand how radiation is detected and measured in experiments.

4. Nuclear Reactions vs Chemical Reactions



You need to be able to distinguish between them:

- Nuclear reactions change the nucleus and can form new elements.
- Chemical reactions involve electrons and do not change the nucleus.

5. Writing Nuclear Equations

Be able to write and balance nuclear equations for:

- Alpha decay (e.g., Uranium-238 \rightarrow Thorium-234 + a)
- Beta decay (e.g., Carbon-14 \rightarrow Nitrogen-14 + β^{-})

Make sure:

• Mass numbers (top) and atomic numbers (bottom) are conserved.

6. Half-Life Calculations

- Be able to use a formula or table to find:
 - Remaining mass or number of atoms after n half-lives.
 - Time elapsed based on half-life and sample reduction.

7. Applications of Radioactivity

Know practical uses and risks:

- Carbon-14: Used in radiocarbon dating (age of once-living things).
- Cobalt-60: Cancer treatment (radiotherapy).
- Smoke detectors: Use of Americium-241 (alpha emitter).
- Food irradiation: Preservation using gamma rays.

8. Background Radiation

- Comes from natural sources (e.g., radon gas, cosmic rays).
- Human-made sources: medical procedures, nuclear power plants.

9. History & Discovery

- Henri Becquerel: Discovered radioactivity.
- Marie & Pierre Curie: Isolated polonium and radium.
- Know the historical timeline briefly (often in MCQs).

10. Safety Precautions

- Minimising exposure: shielding, distance, time.
- Importance of proper storage and disposal of radioactive materials.

11. Past Exam Focus Areas

- Definitions (especially half-life and types of radiation)
- Calculations involving half-life
- Writing balanced nuclear equations
- Applications and effects of radiation
- Short theory questions on detection and safety