

Structure of Atom

Here is the **Master Study Plan** we can follow to master this chapter:

Phase 1: The History & Subatomic Particles

- Discovery of Electrons, Protons, and Neutrons.
- Early models: Dalton, J.J. Thomson, and Rutherford (and their limitations).
- Isotopes, Isobars, and Isoelectronic species (very common 1-mark questions).

Phase 2: Bohr's Model & The Hydrogen Spectrum

- Bohr's model of the atom and energy quantization.
- Calculating the energy, radius, and velocity of an electron in a specific orbit.
- The Hydrogen Spectrum (Lyman, Balmer, Paschen, Brackett, Pfund series) and Rydberg's formula.

Phase 3: The Quantum Revolution

- Dual Nature of Matter (De Broglie's equation).
- Heisenberg's Uncertainty Principle (Concept + Numericals).
- The difference between an *orbit* (Bohr) and an *orbital* (Quantum mechanics).

Phase 4: Quantum Numbers & Shapes of Orbitals

- Principal (n), Azimuthal (l), Magnetic (m_l), and Spin (m_s) quantum numbers.
- Shapes of s , p , and d orbitals.
- Calculating Radial and Angular Nodes.

Phase 5: Electronic Configuration (The Heart of the Chapter)

- **Aufbau Principle, Pauli's Exclusion Principle, and Hund's Rule of Maximum Multiplicity.**
- Writing configurations for neutral atoms and ions.
- Exceptions: Why are Chromium (Cr) and Copper (Cu) special? (Hint: Stability of half-filled and fully-filled orbitals).

How we will work together:

1. **Concept First:** For each phase, I will explain the concept to you in the simplest, most relatable way possible.
2. **Doubt Clearing:** You ask me as many "silly" questions as you want until it clicks.
3. **Question Bank Attack:** We will pull the exact questions from your attached PDFs for that specific topic and solve them together. I'll teach you *how to think* during the exam to avoid tricks.

Phase 1: The History, Subatomic Particles & The "Iso" Family.

1. The 3 Subatomic Particles

Think of an atom like a tiny solar system.

4. **Protons (p^+)**: Positively charged. Found in the center (nucleus). Heavy. (Discovered by Goldstein, named by Rutherford).
5. **Neutrons (n^0)**: No charge (neutral). Also found in the nucleus. Heavy (almost the same mass as a proton). (Discovered by James Chadwick).
6. **Electrons (e^-)**: Negatively charged. They revolve around the nucleus. They are extremely light (almost zero mass compared to protons). (Discovered by J.J. Thomson).

(*Protons + Neutrons together are called **Nucleons** because they live in the nucleus and make up almost all the mass of the atom*).

2. The Early Models (and why they failed)

Teachers love asking "*What were the limitations of...*" in exams. Here is the cheat sheet:

7. **Dalton's Theory**: He said atoms are indivisible (cannot be broken). **Failed** because electrons, protons, and neutrons were discovered later!
8. **J.J. Thomson's Model**: The "Watermelon" or "Plum Pudding" model. He said the atom is a positive sphere with negative electrons embedded inside like seeds. **Failed** because it couldn't explain Rutherford's experiment.
9. **Rutherford's Model**: He fired positive Alpha particles at a gold foil. Most went straight through (meaning the atom is mostly empty space). A few bounced right back (meaning there is a heavy, dense, positive center called the **nucleus**).
 - o **The Big Failure (Limitation)**: Rutherford said electrons revolve around the nucleus like planets around the sun. But according to classical physics, a moving charged particle must constantly lose energy. If this happens, the electron should spiral into the nucleus and the atom would collapse! Therefore, **Rutherford could not explain the STABILITY of the atom**. (Mark this, it's a very common 2-mark question).

3. The "Iso" Family (★ Super Important for 1-mark MCQs)

First, two golden rules:

10. **Atomic Number (Z)** = Number of Protons. (In a neutral atom, Protons = Electrons).
This is the ID card of an element.
11. **Mass Number (A)** = Protons + Neutrons.

Now, meet the family:

1. **Isotopes:** Same element, so they have the **same Atomic Number (Z)**, but **different Mass Number (A)** (because they have different numbers of neutrons).
Example: Carbon-12 and Carbon-14.
 2. **Isobars:** Different elements, so they have **different Atomic Numbers (Z)**, but they happen to have the **same Mass Number (A)**. *Example: Argon (Mass 40) and Calcium (Mass 40).*
 3. **Isoelectronic Species (VVIP):** These are atoms or ions that have the **exact same number of electrons**.
 - *Trick to calculate:* If it has a positive charge (+), subtract that number from the atomic number. If it has a negative charge (-), add that number.
 - *Example:* Na (Sodium) has 11 electrons. Na^+ has lost 1 electron. So Na^+ has 10 electrons. Neon (Ne) also has 10 electrons. Therefore, Na^+ and Ne are Isoelectronic!
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Your Turn! (Questions from your PDF)

Grab a rough notebook. Try to solve these three questions based on what we just discussed. Take your time and reply with your answers and your thought process.

Question 1 (From PDF 1, Page 1, MCQ 10): Iso-electronic species are: (a) F^- , O^{2-} (b) F^- , O (c) F^- , O^+ (d) F^- , O^{2+} (*Hint: Atomic number of F is 9, O is 8. Calculate the total electrons for each!*)

Question 2 (From PDF 1, Page 6, MCQ 2): They have same mass number, different atomic number. These are isobars. In which of the following pairs, the ions are isoelectronic? (a) Na^+ , Mg^{2+} (b) Al^{3+} , O^- (c) Na^+ , O^{2-} (d) Both (a) and (c) are correct. (*Note: I tweaked the options slightly to test you better based on the PDF's answers!*) (*Hint: Atomic numbers - Na: 11, Mg: 12, Al: 13, O: 8*)

Question 3 (Theoretical check): Imagine a 2-mark question in your exam: "Why was Rutherford's model of the atom rejected?" How would you explain it in your own words based on what you read above?

Question 1 (Isoelectronic Species)

Answer: (a) F^- , O^{2-}

Explanation: Isoelectronic species are those that have the same total number of electrons.

- **Atomic numbers:** Fluorine (F) = 9, Oxygen (O) = 8.
- F^- : Neutral F has 9 electrons. Adding 1 (due to the -1 charge) = **10 electrons**.
- O^{2-} : Neutral O has 8 electrons. Adding 2 (due to the -2 charge) = **10 electrons**.
- For other options: O has 8, O^+ has 7, and O^{2+} has 6 electrons. None of these match F^- 's count of 10.

Question 2 (Pair of Isoelectronic Ions)

Answer: (d) Both (a) and (c) are correct.

Explanation: Let's calculate the electrons for each ion (Atomic numbers: Na=11, Mg=12, Al=13, O=8):

- **Option (a):** Na^+ ($11 - 1 = 10$) and Mg^{2+} ($12 - 2 = 10$). (**Isoelectronic**)
 - **Option (b):** Al^{3+} ($13 - 3 = 10$) and O^- ($8 + 1 = 9$). (Not isoelectronic)
 - **Option (c):** Na^+ ($11 - 1 = 10$) and O^{2-} ($8 + 2 = 10$). (**Isoelectronic**) Since both pairs in (a) and (c) have 10 electrons each, option (d) is the right choice.
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Question 3 (Theoretical Check: Rutherford's Model)

Exam Answer (2 Marks): Rutherford's model was rejected primarily because it could not explain the **stability of the atom**.

1. **Electromagnetic Theory:** According to Maxwell's theory, any charged particle (like an electron) moving in a circular path undergoes acceleration and must continuously radiate energy.
2. **Atomic Collapse:** If an electron loses energy, its orbit would shrink, causing it to spiral inward and eventually crash into the nucleus. This would make the atom unstable, which contradicts the fact that atoms are stable in reality. (*Bonus: It also failed to explain the discrete line spectra of atoms.*)

You have a very logical brain, which is exactly what Chemistry requires. You are absolutely on track for full marks!

Let's keep this momentum going and move to the next level.

Phase 2: Bohr's Model & The Hydrogen Spectrum

Rutherford left us with a collapsing atom. Niels Bohr stepped in to save the day in 1913 by introducing a magical word: **Quantization**.

1. Bohr's Model (The Fix)

Bohr said: "*Electrons don't just orbit anywhere. They orbit in specific, fixed circular paths called **Stationary Orbits** (or Shells: K, L, M, N...).*"

- **The Golden Rule:** As long as an electron stays in its designated orbit, it **does not lose or gain energy**. (This fixed Rutherford's collapsing atom problem!).
- **Jumping Electrons:** An electron only absorbs energy when it jumps to a *higher* orbit (excited state) and emits energy when it falls back down to a *lower* orbit (ground state).
- **Angular Momentum is Quantized:** An electron can only revolve in orbits where its angular momentum (mvr) is a multiple of $\frac{h}{2\pi}$. Formula: $mvr = \frac{nh}{2\pi}$ (where $n = 1, 2, 3\dots$).

2. The 3 Magic Formulas for Hydrogen (and H-like ions like He^+ , Li^{2+})

You don't need to derive these, just know how to use them:

- **Radius (r):** $r_n = 0.529 \times \frac{n^2}{Z} \text{ \AA}$
- **Energy (E):** $E_n = -2.18 \times 10^{-18} \times \frac{Z^2}{n^2} \text{ Joules}$ (Notice the negative sign! It means the electron is bound to the nucleus. The closer to the nucleus, the more negative/lower the energy).

3. The Hydrogen Line Spectrum (The "Staircase" Analogy)

Imagine energy levels as steps on a staircase. When you pass electricity through Hydrogen gas, its electron jumps to higher steps. But it is unstable there, so it falls back down, releasing the extra energy as **light**.

Depending on which step the electron *lands* on, we get different "series" of light:

1. **Lyman Series:** Falls to $n_1 = 1$. (Huge drop → High energy → **Ultraviolet region**)
2. **Balmer Series:** Falls to $n_1 = 2$. (Medium drop → **Visible region** - We can actually see these lines!)

3. **Paschen Series:** Falls to $n_1 = 3$. (**Infrared region**)
 4. **Brackett Series:** Falls to $n_1 = 4$. (**Infrared region**)
 5. **Pfund Series:** Falls to $n_1 = 5$. (**Infrared region**)
- **Super Trick:** If an exam asks, "What is the maximum number of emission lines when an electron falls from the n^{th} shell to the ground state?" Use the formula:
$$\frac{n(n - 1)}{2}$$
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Your Turn! (Questions from your PDF)

Let's test this! Grab your notebook again.

Question 1 (From PDF 2, Page 5, MCQ 1 - Conceptual): Bohr model can explain:

- (a) the solar spectrum
- (b) the spectrum of hydrogen molecule (H_2)
- (c) spectrum of any atom or ion containing one electron only
- (d) the spectrum of hydrogen atom only

(Hint: Look closely at the heading for the 3 magic formulas above!)

Question 2 (From PDF 2, Page 10, Section E - Quick Math): What is the maximum number of emission lines when the excited electron of a hydrogen atom in $n = 6$ drops to the ground state? *(Hint: Use the super trick formula!)*

Question 3 (From PDF 2, Page 8, Section C - 3 Marks Theory): Why do atoms emit a **line spectrum** instead of a continuous spectrum, based on the concept of quantization? *(Hint: Think about the "staircase" analogy. Can an electron stand between two steps? If it can only jump between fixed steps, what does that mean for the energy it releases?)*

Question 1

Answer: (c) spectrum of any atom or ion containing one electron only

My logic: I remember Bohr is basically the "one-electron guy." His rules only work for things like H , He^+ , or Li^{2+} . If there's more than one electron, the math gets too hard for his model, so it's not just for Hydrogen, but for anything that *acts* like Hydrogen by having only one electron.

Question 2

Answer: 15

$$\frac{n(n - 1)}{2}$$

My math: I used that "super trick" formula from the notes:

- Here, $n = 6$
 $\frac{6 \times (6 - 1)}{2}$
 - So, $\frac{6 \times 5}{2} = \frac{30}{2}$
 - **Answer = 15 lines.** (I just plugged it in and it worked!)
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Question 3

My explanation: Think of it like a staircase. When you're walking down stairs, you can stand on the 1st step or the 2nd step, but you can't just float in the air between them.

It's the same for electrons in an atom. They have "fixed" energy levels (this is what quantization means). When an electron jumps from a high step to a lower step, it loses a very specific amount of energy. Because that energy is always the same specific amount, it shows up as a single, sharp "line" of color.

If the electron could slide down a ramp instead of jumping down stairs, it would release all kinds of energy and we'd see a messy rainbow (continuous spectrum). But since it's stuck on "steps," we only get specific lines.

Examiners *love* it when students use clear analogies combined with key terms like "fixed energy levels" and "discrete lines."

You have successfully mastered the classical models! Now, we enter the modern era.

Phase 3: The Quantum Revolution (Wave-Particle Duality & Uncertainty)

Bohr's model was great for Hydrogen, but it completely failed for multi-electron atoms. Why? Because Bohr treated the electron purely as a *particle* moving on a fixed track (orbit). Two scientists came along and completely destroyed this idea, birthing **Quantum Mechanics**.

1. Louis de Broglie: The Dual Nature of Matter (1924)

De Broglie said: "*Wait a minute. If light (a wave) can act like a particle (photon), then matter (a particle, like an electron) must also act like a wave!*"

- **The Rule:** EVERY moving object has a wave associated with it.
- **The Formula:** $\lambda = \frac{h}{mv}$ (where λ is wavelength, h is Planck's constant, m is mass, v is velocity, and mv is momentum p).
- **The Catch:** Notice that mass (m) is in the denominator. This means if an object is heavy (like a cricket ball or a car), its wavelength (λ) becomes so incredibly small that we can't even measure it. But for a tiny electron, the wave character is massive and cannot be ignored!

2. Werner Heisenberg: The Uncertainty Principle (1927)

Heisenberg delivered the final blow to Bohr's "fixed circular orbits." He said: "*It is impossible to determine simultaneously, the exact position and exact momentum (or velocity) of a microscopic particle like an electron.*"

- **The Formula:** $\Delta x \cdot \Delta p \geq \frac{h}{4\pi}$ (where Δx is uncertainty in position, and $\Delta p = m\Delta v$ is uncertainty in momentum).
- **Why does this happen?** To "see" an electron, you have to bounce light (a photon) off it. But because the electron is so tiny, hitting it with a photon instantly changes its speed and direction. By the time you know *where* it is, it's already speeding away at an unknown velocity!

3. The Big Shift: Orbit vs. Orbital

Because of Heisenberg, we realized electrons don't travel in fixed paths. We can only talk about the *probability* of finding an electron.

- **Orbit (Bohr):** A well-defined, 2D circular path around the nucleus. (Like a train on tracks. *This concept is now outdated/wrong*).

- **Orbital (Quantum Mechanics):** A 3D region of space around the nucleus where the **probability of finding an electron is maximum** (usually >90%). (Like a fuzzy cloud or a swarm of bees around a hive).
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Your Turn! (Questions from your PDF)

Let's test this new Quantum mindset.

Question 1 (From PDF 1, Page 1, MCQ 1): Which of the following is responsible to rule out the existence of definite paths or trajectories of electrons?

- (a) Pauli's exclusion principle.
- (b) Heisenberg's uncertainty principle.
- (c) Hund's rule of maximum multiplicity.
- (d) Aufbau principal.

Question 2 (From PDF 2, Page 18, Section B - 2 Marks Conceptual): We don't see a car

moving as a wave on the road, why? (*Hint: Use De Broglie's formula $\lambda = \frac{h}{mv}$. What happens to λ when m is very large?*)

Question 3 (Based on PDF 1, Page 4 - 3 Marks Theory): Write 3 points of difference between an **orbit** and an **orbital**. (*Hint: Think about 2D vs 3D, certainty vs probability, and Bohr vs Quantum mechanics!*)

Question 1

Answer: (b) Heisenberg's uncertainty principle.

My reasoning: I remember this one because "uncertainty" means we aren't sure. Heisenberg said we can't know the exact position and the exact speed of an electron at the same time. If we can't even point to where it is exactly, there's no way it can have a "definite path" or a fixed track like a train. It's all just a blur.

Question 2

My explanation: The formula is $\lambda = \frac{h}{mv}$.

- In this, **m** is the mass and λ is the wavelength.
- Since the car is super heavy (it has a huge mass), the "m" in the bottom of the formula makes the wavelength (λ) come out to be extremely, extremely small.

- Because the wave is so tiny, we can't see it or feel it, and the car just looks like it's moving in a straight line like a normal object.
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Question 3

My 3 points for Orbit vs. Orbital:

1. **2D vs 3D:** An **Orbit** is just a flat, circular path (like a 2D drawing of planets), but an **Orbital** is a 3D space or a "cloud" where the electron hangs out.
2. **Being Sure:** In an **Orbit**, we say the electron is *definitely* moving on that specific line. In an **Orbital**, we only say there is a high *chance* (about 90% probability) of finding the electron there.
3. **The Rules:** **Orbit** comes from Bohr's old model (which was simpler), while **Orbital** is part of the new Quantum Mechanics way of thinking.

Phase 4: Quantum Numbers & Shapes of Orbitals (The "Address" of an Electron)

Since we can't pinpoint an electron perfectly (thanks, Heisenberg!), we use **Quantum Numbers** to give the *probability address* of where an electron most likely lives.

Imagine you want to send a letter to an electron. You need its City, Street, House Number, and Roommates.

1. Principal Quantum Number (n) → The "City" (Shell)

- Tells you the **Shell** (K, L, M, N...) and the **Size/Energy** of the orbital.
- Values: $n = 1, 2, 3, 4\dots$

2. Azimuthal / Angular Momentum Quantum Number (l) → The "Street" (Subshell)

- Tells you the **Subshell** and the **Shape** of the orbital.
- **The Rule:** For a given n , l can be anything from 0 **up to** $(n - 1)$.
- **The Cheat Code (MEMORIZE THIS):**
 - If $l = 0 \rightarrow \text{s}$ subshell (Shape: Spherical)
 - If $l = 1 \rightarrow \text{p}$ subshell (Shape: Dumb-bell)
 - If $l = 2 \rightarrow \text{d}$ subshell (Shape: Double dumb-bell)
 - If $l = 3 \rightarrow \text{f}$ subshell (Shape: Complex) (*Example: If $n = 2$, l can only be 0 or 1. So the 2nd shell only has 2s and 2p subshells. There is no 2d!*)

3. Magnetic Quantum Number (m or m_l) → The "House" (Orbital)

- Tells you the **Orientation** (how the orbital is tilted in 3D space).
- **The Rule:** For a given l , m goes from $-l$ **to** $+l$ (including zero).
- *Example:* If we are in the **p** subshell ($l = 1$), m can be $-1, 0, +1$. This means there are exactly **3 houses** (orbitals) in the p subshell (p_x, p_y, p_z).

4. Spin Quantum Number (s or m_s) → The "Roommate" (Spin)

- Tells you the **Spin** of the electron.
- An orbital (house) can only hold a maximum of **2 electrons**. If they live together, they must spin in opposite directions so they don't repel each other too much.
- Values: $+1/2$ (spin up \uparrow) or $-1/2$ (spin down \downarrow).

★ Bonus Concept: NODES (Places where electrons NEVER go)

A node is a region where the probability of finding an electron is exactly **ZERO**.

- **Angular Nodes** = l (Depends entirely on the shape!)
- **Radial Nodes** = $n - l - 1$ (Depends on size and shape)
- **Total Nodes** = $n - 1$

Your Turn! (Questions from your PDF)

Let's test if you can read the "addresses" of these electrons.

Question 1 (From PDF 1, Page 1, MCQ 8): Quantum numbers $n = 2, l = 1$ represent:

- (a) 1s orbital
- (b) 2s orbital
- (c) 2p orbital
- (d) 3d orbital

Question 2 (From PDF 1, Page 3, Section C - 3 Marks Theory): An atomic orbital has $n = 3$. What are the possible values of l and m_l ? (*Hint: Remember the rules: l goes from 0 to $n - 1$. Then, for EVERY value of l you find, list the m_l values from $-l$ to $+l$.*)

Question 3 (From PDF 1, Page 2, Section B - 2 Marks Math): Calculate the total number of **angular nodes** and **radial nodes** present in a **3p** orbital. (*Hint: First, identify n and l for "3p". Then plug them into the nodes formulas!*)

Question 1

Answer: (c) 2p orbital

My logic:

- The first number $n = 2$ just stays as **2**.
 - Then I used the "cheat code" for l :
 - If $l = 0$, it's **s**
 - If $l = 1$, it's **p**
 - So, $n = 2$ and $l = 1$ together make **2p**.
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Question 2

My answer: For $n = 3$, we just follow the rules step-by-step:

1. **Values for l :** It goes from 0 up to $(n - 1)$. Since n is 3, l can be **0, 1, and 2**.
2. **Values for m_l :** Now I have to list the m values for each l I just found:
 - If $l = 0$, then $m_l = 0$
 - If $l = 1$, then $m_l = -1, 0, +1$
 - If $l = 2$, then $m_l = -2, -1, 0, +1, +2$

(Basically, m is just the negative to positive version of whatever l is!)

Question 3

My math: First, for a **3p** orbital:

- $n = 3$
- $l = 1$ (because it's a "p" orbital)

Now I just plug them into the two formulas:

1. **Angular nodes** $= l = 1$
2. **Radial nodes** $= n - l - 1$
 - $= 3 - 1 - 1$
 - $= 1$

So, there is **1 angular node** and **1 radial node**. Total is 2, but the question asked for them separately!

Phase 5: Electronic Configuration (The Rules of the House)

Now that we know the "addresses" (orbitals), how do we actually put the electrons inside them? We have to follow three strict rules.

1. Aufbau Principle (The Energy Rule)

- **The Rule:** Electrons are lazy. They will always fill the **lowest energy orbital first** before moving to higher energy ones.
- **The $(n+l)$ Trick (SUPER IMPORTANT):** How do we know which orbital has lower energy? Just add $(n+l)$!
 - The orbital with the lower $(n+l)$ value gets filled first.
 - *Tie-breaker:* If two orbitals have the *same* $(n+l)$ value, the one with the lower n (smaller city) gets filled first.
 - *Example:* 4s vs 3d. For 4s: $n = 4, l = 0 \rightarrow (4+0) = 4$. For 3d: $n = 3, l = 2 \rightarrow (3+2) = 5$. Since $4 < 5$, **4s is filled BEFORE 3d**.

2. Pauli's Exclusion Principle (The Roommate Rule)

- **The Rule:** An orbital can hold a maximum of **2 electrons**, and they **MUST** have **opposite spins** (one \uparrow , one \downarrow).
- **Scientific Definition:** No two electrons in an atom can have the exact same set of all four quantum numbers (n, l, m_l, m_s). If they live in the same house, their spin *must* be different.

3. Hund's Rule of Maximum Multiplicity (The Empty Bus Seat Rule)

- **The Rule:** Imagine getting on a bus with empty double seats. Strangers will always sit in empty seats first before they start sitting next to someone else.
- Electrons do the same! In degenerate orbitals (orbitals with the same energy, like the three houses of the p subshell: p_x, p_y, p_z), **every orbital gets ONE electron (all with parallel spins \uparrow) before any pairing ($\uparrow\downarrow$) begins.**

★ THE EXCEPTIONS: Chromium (Cr) and Copper (Cu)

Teachers *obsess* over this in exams.

- Normally, Chromium ($Z = 24$) should be: $4s^2 3d^4$
- But it is actually: $4s^1 3d^5$
- **WHY?** Because a **half-filled (d^5)** or **fully-filled (d^{10})** subshell is extremely stable.
- *Two reasons for this extra stability:*
 1. **Symmetrical distribution** of electrons (nature loves symmetry).
 2. **Exchange energy:** Electrons in degenerate orbitals love to swap positions.
More swaps = more energy released = more stability!

Your Final Test! (Questions from your PDF)

Let's finish strong!

Question 1 (From PDF 1, Page 1, MCQ 3): For which of the following sets of quantum numbers, an electron will have the **highest energy**? (a) $n = 3, l = 2, m_l = +1, m_s = +1/2$ (b) $n = 4, l = 2, m_l = -1, m_s = +1/2$ (c) $n = 4, l = 1, m_l = 0, m_s = -1/2$ (d) $n = 5, l = 0, m_l = 0, m_s = +1/2$

(Hint: Use the $(n + l)$ rule for each option! Ignore m_l and m_s , they don't affect energy here.)

Question 2 (From PDF 2, Page 2, Section B - 2 Marks Theory): Which one of the following electronic configurations is correct for Chromium? a) $4s^2 3d^4$ OR b) $4s^1 3d^5$? Justify your answer. (Hint: Use the exact keywords from the exceptions section above!)

Question 3 (From PDF 2, Page 2, Section B - 2 Marks Theory): Out of 3d and 4s orbitals, which is filled first? Give the reason. (Hint: Show the math using the trick you just learned!)

Question 1

Answer: (b) $n = 4, l = 2, m_l = -1, m_s = +1/2$

My math: I just added n and l for each one like you said:

- (a) $3 + 2 = 5$
- (b) $4 + 2 = 6$
- (c) $4 + 1 = 5$
- (d) $5 + 0 = 5$

Since **6** is the biggest result, option (b) has the highest energy. It doesn't matter what the other numbers (m_l and m_s) are, they don't change the "score."

Question 2

Answer: (b) $4s^1 3d^5$

My justification: Chromium is one of those "weird" exceptions we have to memorize. It doesn't follow the normal rule because a **half-filled** d-subshell ($3d^5$) is extra **stable**. To get that stability, one electron jumps from the $4s$ over to the $3d$. It's all about being more stable and symmetrical!

Question 3

Answer: $4s$ is filled first.

My reasoning (using the trick): I compared the $(n + l)$ values for both:

- **For $4s$:** $n = 4$ and $l = 0$ (because it's an 's' orbital). So, $4 + 0 = 4$.
- **For $3d$:** $n = 3$ and $l = 2$ (because it's a 'd' orbital). So, $3 + 2 = 5$.

Since **4 is lower than 5**, the $4s$ orbital has lower energy. The rule is that electrons always go into the lower energy spot first, so $4s$ gets filled before $3d$.

Look at what you just accomplished!

Look at what you just did:

1. You mastered the history and subatomic particles (Isotopes/Isobars/Isoelectronic).
2. You understood Bohr's model, the Hydrogen spectrum, and the math behind the lines.
3. You grasped high-level Quantum Mechanics (De Broglie and Heisenberg) and the difference between Orbitals and Orbitals.
4. You learned how to extract the "address" of an electron using Quantum Numbers and calculate nodes.
5. You successfully applied the Aufbau, Pauli, and Hund's rules to write configurations and explain exceptions like Chromium.

You are highly logical, you pick up patterns quickly, and you know exactly how to apply formulas to new questions. This is the exact mindset needed to score 100/100 in Chemistry.

What's Next?

You are fully equipped to tackle your "Structure of Atom" exam. Here is what I recommend you do on your own time to lock this in permanently:

1. **Write out the configurations** of elements 1 to 30 (especially Copper, Cu, which is the other exception: $4s^13d^{10}$).
2. **Skim your NCERT textbook** once—you will be amazed at how easily you understand the text now.
3. **Keep this chat as your personal cheat sheet** to revise before the exam.

You did an amazing job today. Be proud of yourself!