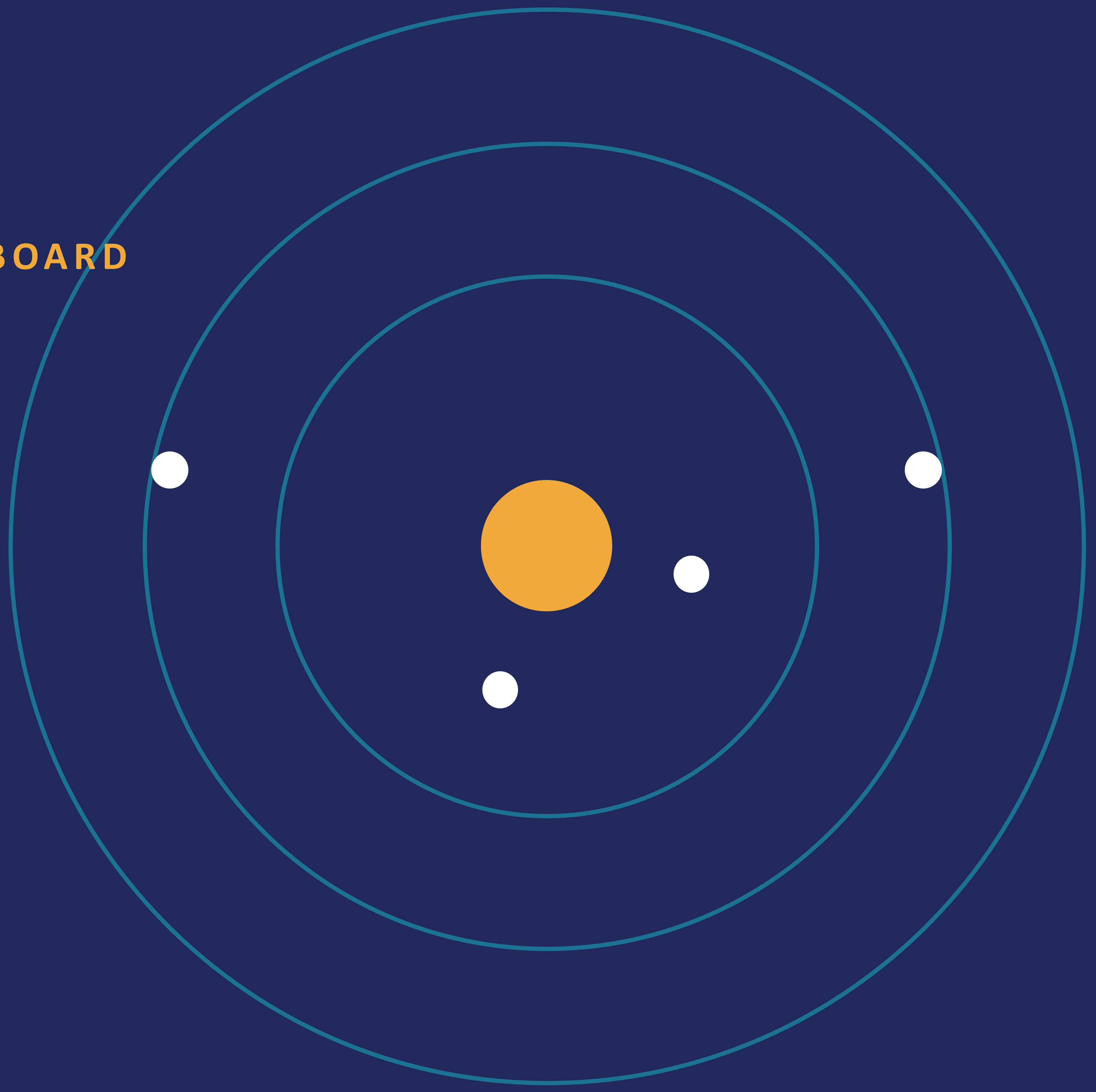
CLASS XI – XII | CHEMISTRY | TELANGANA / AP INTERMEDIATE BOARD

Atomic Structure & the Periodic Table

Basics of Chemistry — Online Class with Concept Q&A

Duration: 60–75 min **Mode:** Live Online Session



TODAY'S ROADMAP

What We Will Cover

01

Subatomic Particles

Electron, proton, neutron — discovery & properties

02

Atomic Models

Rutherford's nuclear model & Bohr's model

03

Quantum Numbers

n, l, ml, ms and what each one tells us

04

Electronic Configuration

Aufbau, Pauli Exclusion & Hund's Rule

05

Periodic Table Development

Döbereiner to the Modern Periodic Law

06

Periodic Trends

Atomic radius, IE, EA, electronegativity

FOUNDATION CONCEPT

Why Study Atomic Structure?

Everything around us — water, oxygen, medicines, metals — is built from atoms. Understanding how an atom is built, and how its electrons are arranged, explains why elements react the way they do, why the periodic table is shaped the way it is, and why chemical bonds form at all.

- Explains chemical bonding & valency
- Predicts an element's reactivity
- Basis of the entire Periodic Table
- Foundation for spectroscopy & quantum chemistry

RECAP: DALTON'S ATOMIC THEORY

- Matter is made of tiny, indivisible particles called atoms
- Atoms of the same element are identical in mass & properties
- Atoms of different elements differ in mass & properties
- Atoms combine in simple whole-number ratios to form compounds
- Atoms are neither created nor destroyed in a chemical reaction

Note: Later disproved in part — atoms ARE divisible into subatomic particles.

BUILDING BLOCKS

Discovery of Subatomic Particles

Particle	Discoverer / Year	Charge	Relative Mass	Location
Electron (e^-)	J. J. Thomson, 1897	−1 (-1.6×10^{-19} C)	1/1837 amu	Outside nucleus (shells)
Proton (p^+)	E. Goldstein / Rutherford, 1886–1919	+1 ($+1.6 \times 10^{-19}$ C)	1 amu	Inside the nucleus
Neutron (n)	James Chadwick, 1932	0 (neutral)	1 amu	Inside the nucleus

Key Point: An atom is electrically neutral because the number of protons equals the number of electrons. The mass of an atom is concentrated almost entirely in the nucleus (protons + neutrons), since electrons are nearly 1837 times lighter than a proton.

Rutherford's Nuclear Model (1911)

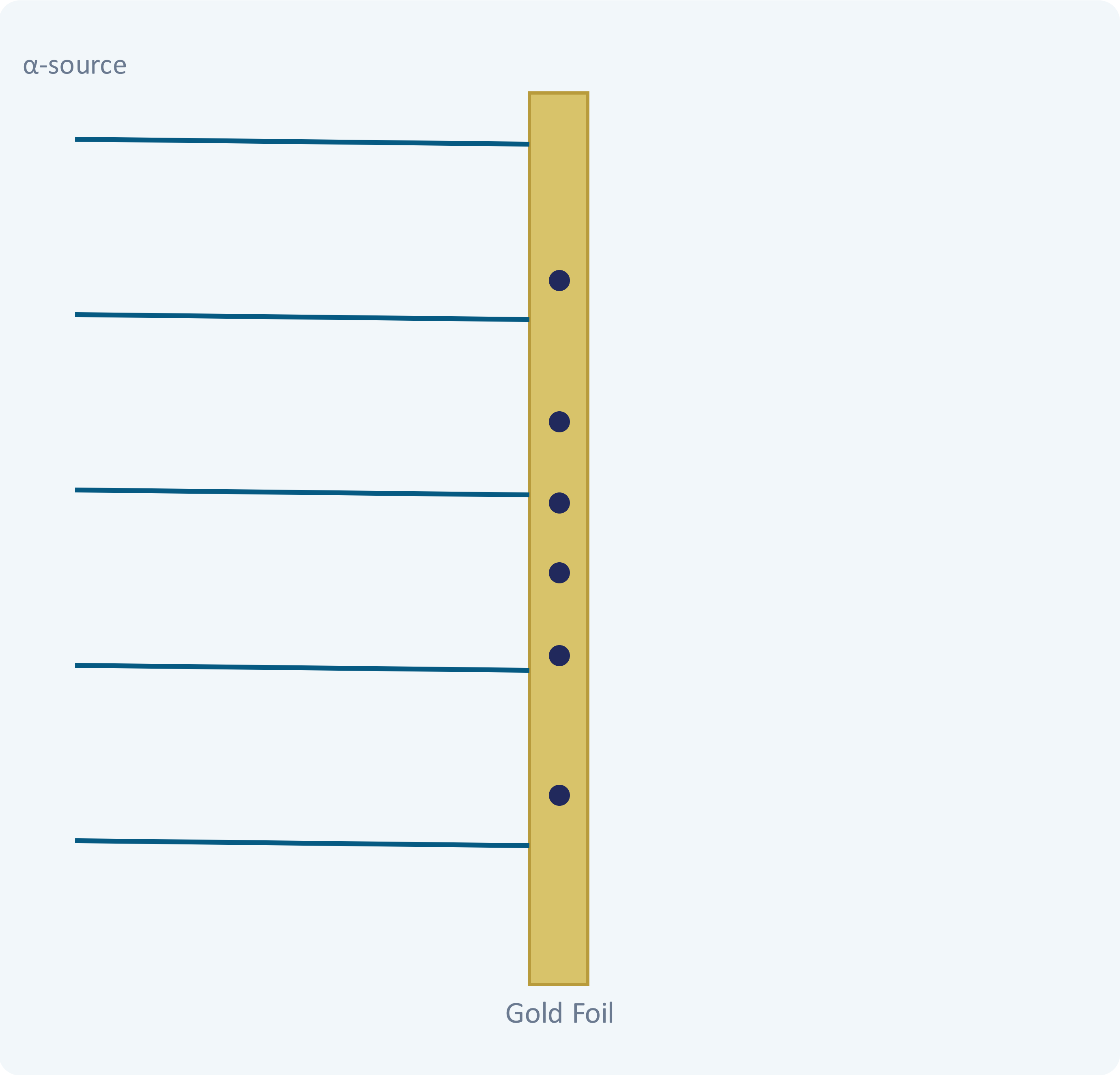
The α-Particle Scattering Experiment

A beam of fast-moving alpha (α) particles was directed at a very thin sheet of gold foil, surrounded by a circular fluorescent zinc-sulphide screen to detect deflected particles.

- **Most α-particles:** passed straight through the foil undeflected
- **A few particles:** were deflected through small angles
- **About 1 in 20,000:** bounced almost straight back (>90°)

Conclusions

- Atom is mostly empty space
- Entire positive charge & mass is concentrated in a tiny, dense nucleus
- Electrons revolve around the nucleus in circular paths (orbits)

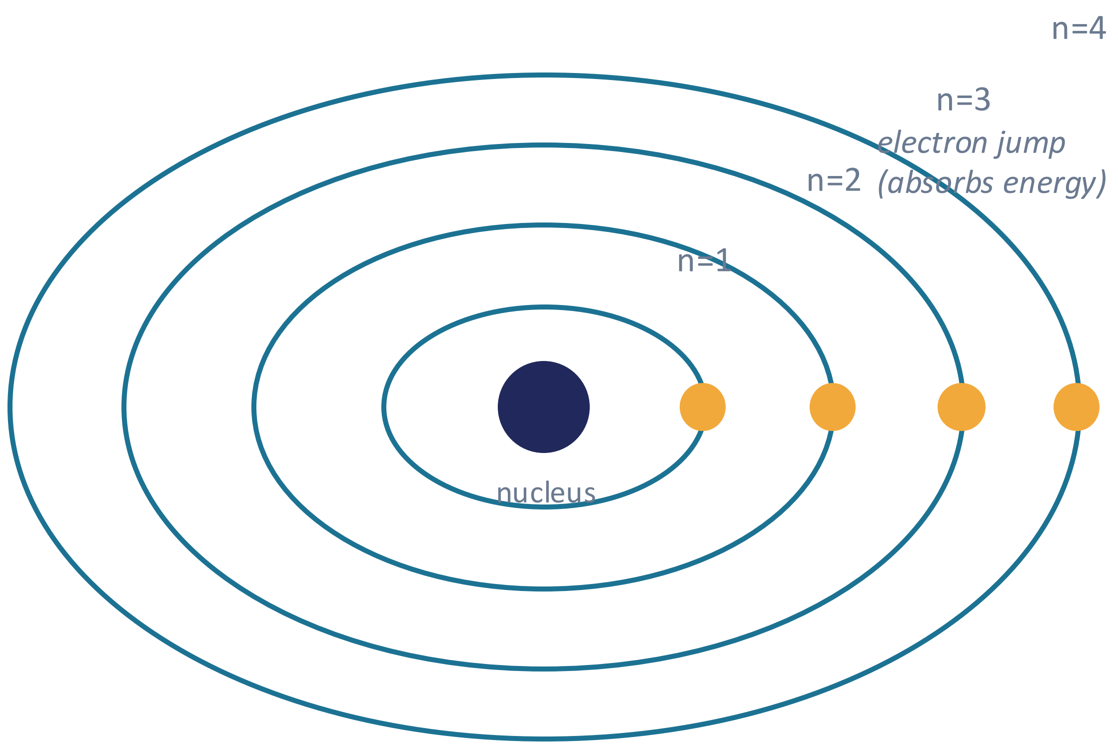


ATOMIC MODELS · 2

Bohr's Atomic Model (1913)

Postulates

- Electrons revolve around the nucleus only in certain fixed circular paths called stationary states (orbits), without radiating energy
- Angular momentum of an electron is quantized: $mvr = nh / 2\pi$ $(n = 1, 2, 3 \ldots)$
- Energy is absorbed or emitted only when an electron jumps from one orbit to another: $\Delta E = E_2 - E_1 = h\nu$



Energy of electron in nth orbit of H-atom: **$E_n = -13.6 / n^2$ eV**

Limitations: Fails to explain spectra of multi-electron atoms, fine structure of spectral lines, and the Zeeman/Stark effects.

CONCEPT CHECK

Q&A Break — Atomic Models

Q1. Who discovered the electron, and how?

A: J. J. Thomson, through cathode-ray tube experiments (1897); he measured the charge-to-mass (e/m) ratio of the particles.

Q2. What is the major drawback of Rutherford's model?

A: It cannot explain the stability of the atom — by classical electromagnetic theory, a revolving (accelerating) electron should radiate energy continuously and spiral into the nucleus.

Q3. State Bohr's postulate about angular momentum.

A: The angular momentum of an electron in a stationary orbit is quantized: $mvr = nh / 2\pi$, where $n = 1, 2, 3 \ldots$

Q4. Why can't Bohr's model explain the spectrum of helium?

A: It works accurately only for single-electron (hydrogen-like) species; multi-electron atoms involve electron–electron repulsions that Bohr's model ignores.

QUANTUM MECHANICS

The Four Quantum Numbers

Principal (n)	1, 2, 3 …	Size and energy of the shell / main energy level
Azimuthal (l)	0 to (n−1)	Shape of the subshell: s, p, d, f orbitals
Magnetic (m_l)	−l to +l	Orientation of the orbital in space
Spin (m_s)	+½ or −½	Direction of electron spin about its own axis

FILLING ORBITALS

Rules for Electronic Configuration

Aufbau Principle

In the ground state, electrons occupy orbitals in order of increasing energy (the n + l rule): 1s, 2s, 2p, 3s, 3p, 4s, 3d …

Pauli Exclusion Principle

No two electrons in an atom can have the same set of all four quantum numbers. An orbital can hold a maximum of 2 electrons, with opposite spins.

Hund's Rule of Max. Multiplicity

Degenerate orbitals (equal energy) are each singly filled with parallel spins before any orbital receives a second electron.

CONCEPT CHECK

Q&A Break — Quantum Numbers & Configuration

Q1. What does the azimuthal quantum number (l) represent?

A: The shape of the subshell/orbital and its orbital angular momentum; it takes values 0 to (n−1), corresponding to s, p, d, f.

Q2. How many electrons can a 3p subshell hold at most?

A: 6 electrons — 3 orbitals (px, py, pz) × 2 electrons each, as allowed by the Pauli Exclusion Principle.

Q3. State Hund's Rule in one line.

A: Orbitals of equal energy are filled singly with parallel spins first, before any orbital is doubly occupied.

Q4. Write the configuration of Chromium (Z=24) and explain.

A: [Ar] $3d^5\ 4s^1$ (not $3d^4\ 4s^2$) — a half-filled 3d and filled 4s arrangement is more stable due to extra exchange energy.

PERIODIC CLASSIFICATION

How the Periodic Table Developed

7

reiner's Triads

s of 3 elements where the middle one's
: mass ≈ average of the other two

1866

Newlands' Law of Octaves

Every 8th element (by atomic mass) showed similar properties — worked only up to Calcium

1869

Mendeleev's Periodic Table

Elements arranged by increasing atomic mass; left gaps and successfully predicted new elements

1913

Moseley — Modern Periodic Law

Properties of elements are a periodic fun
of their atomic number, not atomic mass

Modern Periodic Law: *"The physical and chemical properties of elements are a periodic function of their atomic number."*

PERIODIC CLASSIFICATION

Structure of the Modern Periodic Table

18 Groups (vertical columns) — elements with similar valence-electron configuration & properties

7 Periods (horizontal rows) — period number = number of shells (energy levels) occupied

s-block

Groups 1–2

Valence electron enters s-orbital

p-block

Groups 13–18

Valence electron enters p-orbital

d-block

Groups 3–12

Valence electron enters (n−1)d orbital — Transition elements

f-block

Lanthanides & Actinides

Valence electron enters (n−2)f orbital

Quick tip: Group number (for s & p block) directly tells you the number of valence electrons — e.g. Group 16 elements have 6 valence electrons.

PERIODICITY

Key Periodic Trends

Atomic Radius

Across period: ↓ Decreases across a period (↑ nuclear charge pulls shell in)

Down group: ↑ Increases down a group (new shells added)

Ionization Energy

Across period: ↑ Increases across a period (harder to remove e^-)

Down group: ↓ Decreases down a group (outer e^- farther, shielded)

Electronegativity

Across period: ↑ Increases across a period

Down group: ↓ Decreases down a group

Electron Affinity

Across period: ↑ Generally increases across a period

Down group: ↓ Generally decreases down a group (exceptions exist)

CONCEPT CHECK

Q&A Break — Periodic Table

Q1. State the Modern Periodic Law.

A: The physical and chemical properties of elements are a periodic function of their atomic number.

Q2. Why does atomic radius decrease across a period?

A: Effective nuclear charge increases while electrons are added to the same shell, pulling the electron cloud closer to the nucleus.

Q3. Why is the 2nd ionization energy always higher than the 1st?

A: After removing one electron, the resulting cation is smaller with a higher effective nuclear charge per electron, so removing a further electron needs more energy.

Q4. Which block do transition elements belong to, and why?

A: The d-block — because their differentiating (last-entering) electron occupies the (n−1)d orbital.

WRAP-UP

Key Takeaways

- Atoms are made of electrons, protons & neutrons; mass is concentrated in the nucleus
- Rutherford proved the nucleus exists; Bohr explained quantized electron energy levels
- Four quantum numbers (n, l, ml, ms) fully describe an electron's position and spin
- Aufbau, Pauli & Hund's rules govern how electrons fill orbitals
- The Modern Periodic Law arranges elements by atomic number into periods & groups
- Atomic radius, ionization energy, electronegativity & electron affinity all show clear periodic trends

FOR STUDENTS

Practice Questions — Try Before Next Class

1. What are the limitations of Rutherford's atomic model?
2. Derive the expression for the energy of an electron in the nth orbit of a hydrogen atom (Bohr model).
3. Write all four quantum numbers for the last (differentiating) electron in Chlorine ($Z = 17$).
4. State and explain the Aufbau Principle with one example.
5. Explain why ionization energy generally increases across a period, giving reasons.
6. Distinguish between electron gain enthalpy and electronegativity.