TOPIC / THEME: ATOMIC AND ELECTRONIC STRUCTURE OF THE ATOM

Learning Outcomes (Competencies)

By the end of the topic, learners should be able to:

1. Describe the arrangement of electrons in an atom based on Aufbau's principle, Pauli's exclusion principle, and Hund's rule.
2. Write electronic configurations of elements using energy levels and sublevels.
3. Relate the electronic configuration of elements to their positions in the Periodic Table.
4. Apply the rules of electron arrangement to predict chemical properties of elements.

Introduction

The study of atomic and electronic structure helps us understand the fundamental nature of matter. Atoms are known to consist of fundamental subatomic particles (protons, neutrons, and electrons), and the arrangement of electrons in the atoms determines the chemical behaviour of elements.

Particle	Symbol	Relative Mass	Relative Charge	Location
Proton	p^+	1	+1	Nucleus
Neutron	N	1	0	Nucleus
Electron	e^-	~1/1840	-1	Electron cloud

- Protons and neutrons form the positively charged nucleus.
- Electrons are found outside the nucleus occupying regions of space along energy levels (sometimes called the quantum shells). The number of energy levels shown in the distribution of electrons depends on the number of electrons a particular atom contains. The number of energy levels increases as the number of electrons atoms contain increase. Energy levels are numbered as 1, 2, 3, etc as they move away from the nucleus. The 1st energy level is closest to the nucleus, followed by the 2nd energy level, then the 3rd energy level, the 4th etc.

Electron Orbitals

An **orbital** *is a region or volume of space around the nucleus within which there is a high probability of finding an electron.*

Electrons can occupy four different types of orbitals labelled as s, p, d and f.

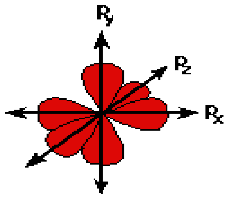
(a) The s-orbital: The s-orbital is **a spherically-shaped region that describes where an electron can be found**, within a certain degree of probability.

The s-orbital is occupied by the first electron in any given energy level**.**

The electrons in the s-orbital are closer to the nucleus than those of other types of orbitals in any given energy level in order to slightly reduce the energy of electrons in the orbital. The nearer to the nucleus the electrons get, the lower their energy.

The s-orbital consists of one orbital that accommodates a maximum of two electrons. Hence, at the first energy level, the only orbital available to electrons is the s-orbital.

(b) The p orbitals: The p-orbitals are **a dump bell-shaped region that describes where an electron can be found**, within a certain degree of probability.



At the second energy level, the orbitals available to electrons are the 2s orbital and ***2p orbitals.***

All energy levels except for the 1st energy level have p-orbitals. They consist of a set of three p-orbitals which equivalent or similar in energy pointing mutually at right angles to each other.

Thus, there are a total of four orbitals altogether in the 2nd energy level.

These are arbitrarily given the symbols p_x, p_y and p_z simply for convenience. The x, y or z direction changes constantly as the atom tumbles in space.

A p-orbital is like 2 identical balloons tied together at the centre.

The orbitals that have similar or equivalent energies and are within the same sub energy level are termed as ***degenerate orbitals***.

(c) The d orbitals: The d-orbitals are **a complex-shaped region that describes where an electron can be found**, within a certain degree of probability. They consist of a set of five orbitals with equivalent or similar energies that accommodates a maximum of ten electrons.

All energy levels except for the 1st and 2nd energy level have d-orbitals. i.e. At the third level, there is a set of five equivalent ***d orbitals*** on addition to those in the 3s and 3p orbitals. Thus, there are a total of nine orbitals altogether in the 3rd energy level.

(d) The f orbitals: The f-orbitals are **a complex-shaped region that describes where an electron can be found**, within a certain degree of probability.

They consist of a set of seven orbitals with equivalent or similar energies that accommodates a maximum of fourteen electrons. At the fourth level, on addition to the 4s and 4p and 4d orbitals there is a set of seven equivalent ***f orbitals***. Thus, there are 16 orbitals altogether in the 4th energy level.

The s, p, d and f orbitals are then available at all other higher energy levels.

Quantum Number

This is a number used when describing the energy levels available to atoms. These include:

(a) ***Principal quantum number, n***. This is a number that represents the main *energy level* (*quantum shell*) of an atom that contains electrons and determines the energy of an electron.

It is designated by a whole number 1, 2, 3, 4, etc.

The total number of electrons in any of the first four given energy levels is given by $2n^2$ where n is the principal quantum number.

Thus, the maximum number of electrons in each energy level is as shown below:

Energy Level	1st	2nd	3rd	4th
Maximum number of electrons	2	8	18	32

(b) ***Subsidiary (****angular / secondary / azimuthal****) quantum number, l.*** This is a number used to represent a *sub energy level or sub shell* that is a division of the main energy level and it mainly contains degenerate orbitals.

The sub-energy levels are designated as s, p, d or f.

The subsidiary quantum number mainly occurs for principal quantum numbers greater than one (beyond the 1st energy level i.e. from the 2nd energy level). **l = n-1.**

L	*Sub energy level*	*Number of orbitals*	*Total number of electrons*
1-1= 0	s	1	2
2-1 = 1	p	3	6
3-1 = 2	d	5	10
4-1 = 3	f	7	14

In summary, the 1st energy level has no sub energy level, the 2nd energy level has two sub energy levels, s & p. The 3rd energy level has three sub energy levels s, p & d, while the 4th energy level has four sub energy levels s, p, d & f.

(c) ***Magnetic quantum number, m.*** This is a number that represents *orbitals within a sub energy level* all of which have similar energy values. Thus, the *s*-sub energy level has one orbital, the *p*-sub energy level has three orbitals, etc.
(d) ***Spin quantum number, s.*** This is a number that represents *energy for each of the two electrons within an orbital that have opposite spins*.

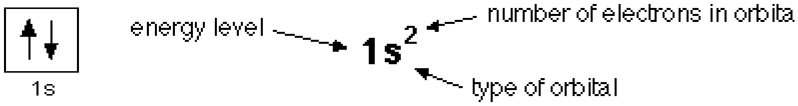
ELECTRONIC CONFIGURATION OF ATOMS AND IONS

Definition: The electronic configuration of an atom is one that describes how electrons are distributed (arranged) in energy levels (shells) and sublevels (orbitals) around the nucleus of an atom.
There are **three common ways** electronic configuration is represented:
(a) Energy Level (Shell) Notation: This shows how electrons are distributed among the **main energy levels (shells)** around the nucleus. The shells are labeled **K, L, M, N…** or **1, 2, 3, 4…**
Each shell has a maximum number of electrons; 1st shell = 2 electrons, 2nd shell = 8 electrons, 3rd shell = 18 electrons (often simplified to 8 for lighter elements in basic ordinary level chemistry), etc.
Example: Sodium, Na with 11 electrons. Electronic configuration is: **2:8:1** or **2, 8, 1**.
This means that 2 electrons in the first shell, 8 in the second shell and 1 in the third shell. This method is common in **basic chemistry and early science education at ordinary level**.
(b) Subshell (Orbital) Notation / Spectroscopic Notation: This shows the distribution of electrons in **sub energy levels (orbitals)** such as **s, p, d, and f**. The order follows the **Aufbau principle**.
Format is; Number of energy level + Orbital type + Number of electrons.
Example: Sodium, Na with 11 electrons. Electronic configuration is: **$1s^2$ $2s^2$ $2p^6$ $3s^1$.**
This means that, 2 electrons are in 1s, 2 electrons are in 2s orbital, 6 electrons in 2p orbitals and 1 electron in the 3s orbital. This method is widely used in **advanced chemistry and physics at Advanced Level.**
(c) Orbital Diagram (Box Diagram): this shows electrons being represented by **arrows (↑ or ↓)** inside boxes representing orbitals. The order follows **Pauli's Exclusion Principle, Hund's Rule** and **Aufbau Principle**.
Example: Sodium, Na with 11 electrons.
Electronic configuration is: **1s[↑↓]** **2s[↑↓]** **2p[↑↓][↑↓][↑↓]** **3s[↑]**
This method clearly shows **electron spins and pairing**.

✅ **Summary:**

Representation	**Description**	**Example**
Energy level notation	Electrons per energy level	2:8:1
Subshell notation	Electrons in s, p, d, f orbitals	**$1s^2$ $2s^2$ $2p^6$ $3s^1$**
Orbital diagram	Boxes and arrows showing electron spin	**1s[↑↓] 2s[↑↓] 2p[↑↓][↑↓][↑↓] 3s[↑]**

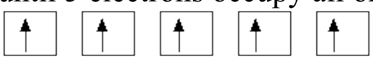
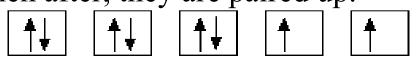
The arrangement of electrons follows three fundamental principles or rules:
(a) Pauli's Exclusion Principle.
It states that, "electrons are arranged in a free atom in such a way that an orbital can accommodate a maximum of two electrons which must spin in opposite directions." **or** "electrons are arranged in a free atom such that no two electrons can have the same values for quantum numbers." (↑↓)
Electron arrangement is represented using boxes (orbitals) and arrows (electrons) but it is commonly represented using the principal quantum number, symbol of the orbital and the number of electrons.



(b) Hund's Rule (Rule of maximum multiplicity or "Bus Seat Rule").

It states that, "electrons are arranged in a free atom where there are orbitals of equal or similar energy in such a way that electrons are filled in all the orbitals singly with parallel spins before pairing them up with opposite spins." This arrangement of electrons occurs in order to minimise the repulsions between electrons, keeping them as far away from each other as possible hence making the atom more stable.

e.g. arrangement of electrons in the d-orbitals. There are five d-orbitals, the electrons will inhabit them singly first until 5 electrons occupy all orbitals. Then after, they are paired up.

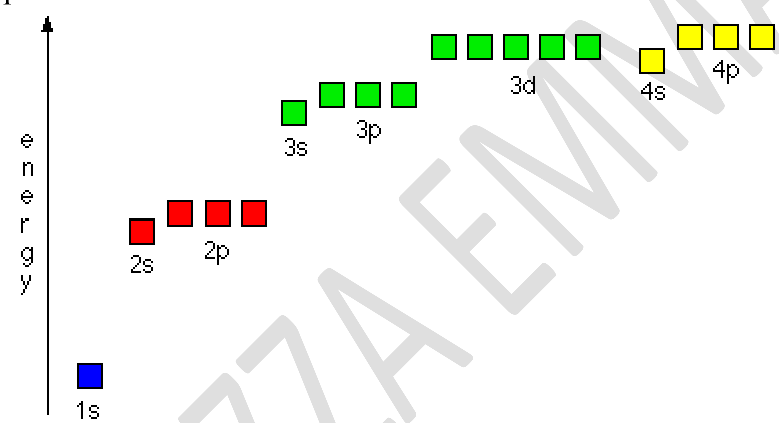
d^5 is  d^8 is 

(c) Aufbau Principle ("Building-Up Principle").

Aufbau is a German word meaning *building up* or *construction*. The principle states that, "electrons are arranged in a free atom in such a way that they fill lower energy levels / orbitals or sub energy levels first before they fill higher energy levels / orbitals".

This is because lower energy levels are closer to the nucleus and hence are more stable.

The diagram below summarises the level of energies of the orbitals up to the 4p level as illustrated using the Aufbau Principle.



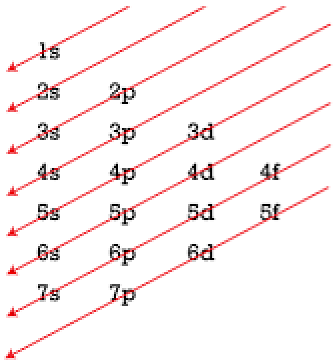
i.e. The first energy level is lowest in energy, followed by the second energy level, then the third, etc.

On the other hand, the s-orbital always has the lowest energy than the p-orbitals at the same energy level, so the s-orbital always fills with electrons before the corresponding p-orbitals.

However, it is observed that the 3d-orbitals are at a slightly higher energy level than the 4s orbital hence the 4s orbital which is filled first, followed by all the 3d orbitals and then the 4p orbitals etc. This phenomenon is also realized in other higher sub energy levels.

Thus, from the Aufbau principle, the order in which electrons are arranged in an atom of an element in the respective sub-energy levels is arrived at is shown by the diagonal rule of the Aufbau diagram below:

The diagonal rule
for electron
filling order.



The order of filling follows the energy level sequence:
1s → 2s → 2p → 3s → 3p → 4s → 3d → 4p → 5s → 4d → 5p → 6s → 4f → 5d → 6p → 7s → 5f → 6d → 7p

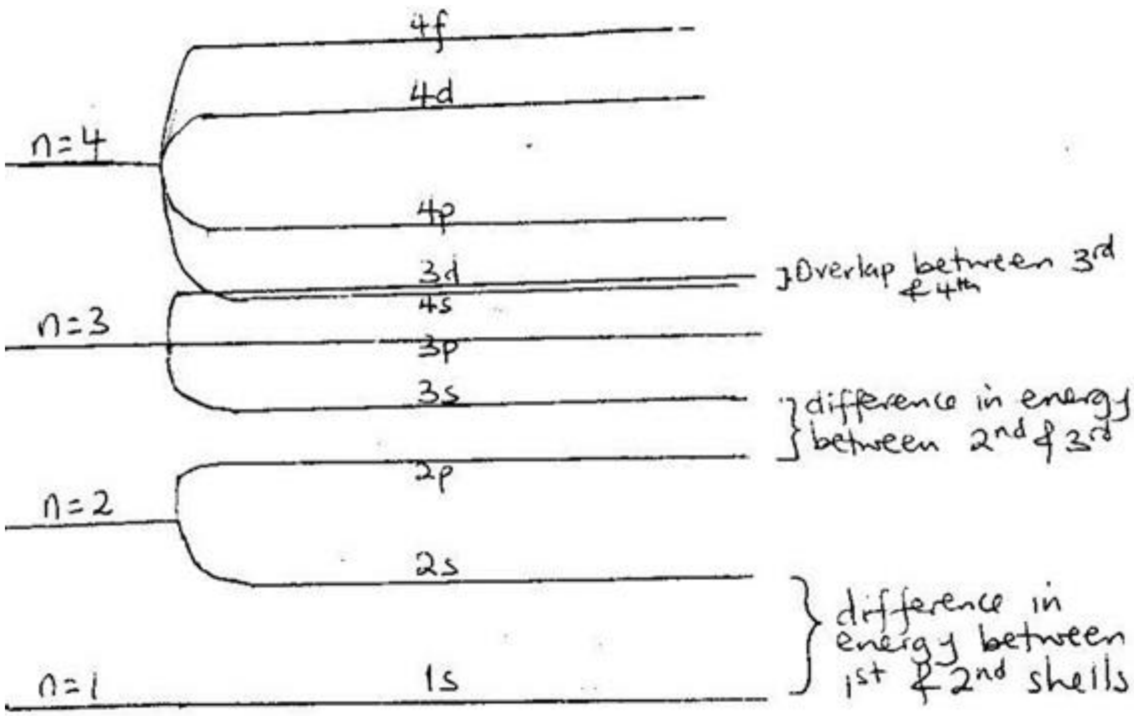
Writing Electronic Configuration of Atoms

There are two common notations for writing electronic configurations:
(a) Expanded Notation: This shows the complete or full arrangement of electrons in sublevels.
Example: Carbon (C) → $1s^2\ 2s^2\ 2p^2$
(b) Noble Gas Notation (Shorthand): This uses the noble gas core (symbols) for simplicity.
Example: Chlorine (Cl) → $[Ne]\ 3s^2\ 3p^5$

Item:
The S.5 chemistry class at a secondary school is preparing for an upcoming practical and theory assessment examination on **atomic structure**. Their teacher wants to ensure that the students understand how electrons are arranged in atoms and how the arrangement follows the basic rules of filling electron shells and subshells.
During the lesson, the teacher reminds the class that the first twenty elements of the periodic table are commonly used to illustrate how electrons occupy energy levels according to the **Aufbau principle**, the **Pauli exclusion principle**, and **Hund's rule**.
To help the students connect theory with real chemical elements, the teacher displays a periodic table in the laboratory and asks the class to imagine they are building atoms one by one, starting from hydrogen and moving across the table to calcium. For each element, the students must determine how the electrons fill the available orbitals in the order **1s, 2s, 2p, 3s, 3p, 4s**, ensuring that the number of electrons matches the atomic number of the element.
The teacher then assigns the students an exercise. They must carefully write the **electronic configuration** for every element from atomic number 1 to 20, showing how the electrons are distributed in the orbitals. This task will help them understand periodic trends, chemical reactivity, and how the structure of atoms influences the behavior of elements in chemical reactions.
Task: As a chemistry student, based on your understanding of electron arrangement in atoms, **write the electronic configuration of all the first twenty elements of the periodic table (from hydrogen to calcium).**

(Leave a page as space for the task)

Important Points to Note:
(a) As electrons are distributed in these atoms, the energy levels get further away from the nucleus and the energy difference between the energy levels decreases successively until they eventually become closer in energy or they merge. This leads to an overlap between orbitals of different energy levels.
Thus, the difference in energy between the second and third energy level is less than that between the first and second energy level. As the 3d sub-energy level is reached, there's an overlap between the third and fourth energy level as shown in the diagram below:



Where n is the quantum number.

(b) After distribution of electrons in the atomic orbitals, the electrons in the 3d orbitals repel those in the 4s orbital. And as such, the 4s orbital shifts to a position at a higher energy level in order to minimize the repulsion. Consequently, the 4s orbital is located as the outermost orbital.

e.g. the filling of electrons in scandium using the Aufbau diagram is $1s^22s^22p^63s^23p^64s^23d^1$

After rearrangement, the 4s orbital located as the outermost orbital i.e. $1s^22s^22p^63s^23p^63d^14s^2$

(c) Rearrangement of electrons in both the 3d and 4s orbitals for some atoms of elements (chromium and copper) occurs since there is a possibility of forming more stable electronic structures. An electronic structure is energetically (thermodynamically) stable when the orbitals are half filled and / or completely (fully) filled.

(i) Chromium (24) has an electronic structure of $1s^22s^22p^63s^23p^63d^54s^1$ and not $1s^22s^22p^63s^23p^63d^44s^2$ as might be expected. This is because the electronic structure $1s^22s^22p^63s^23p^63d^54s^1$ has half-filled 3d and 4s orbitals which have lower energy and therefore more energetically stable. The 4s orbital being occupied by one electron creates equal distribution of charge around the atom.

i.e. [Ar] $3d^5$ | ↑ | ↑ | ↑ | ↑ | ↑ | $4s^1$ | ↑ |

where [Ar] is the electronic configuration of Argon, $1s^22s^22p^63s^23p^6$

(ii) Copper (29) has an electronic structure of $1s^22s^22p^63s^23p^63d^{10}4s^1$ rather than $1s^22s^22p^63s^23p^63d^94s^2$. This is because the electronic structure $1s^22s^22p^63s^23p^63d^{10}4s^1$ has a fully filled 3d and a half filled 4s orbitals which are energetically stable. There is uniform distribution of charge in the sub energy levels making the electronic structure more stable than $1s^22s^22p^63s^23p^63d^94s^2$.

i.e [Ar] $3d^{10}$ | ↑↓ | ↑↓ | ↑↓ | ↑↓ | ↑↓ | $4s^1$ | ↑ | is more stable than, [Ar] $3d^9$ | ↑↓ | ↑↓ | ↑↓ | ↑↓ | ↑ | $4s^2$ | ↑↓ |

Item:

At **Kisubi Secondary School**, the Senior five learners are carrying out a practical activity in the chemistry laboratory. Their teacher, **Mr. Okello**, has asked them to investigate the properties of **transition elements** found in the periodic table. During the lesson, the students observe samples and diagrams of elements in the **first transition series**, ranging from **Scandium (Sc) to Zinc (Zn)**.

Mr. Okello explains that understanding the **electronic configuration** of these elements helps scientists know how they form compounds, conduct electricity, and are used in industries such as making tools, coins, and electrical materials. He then divides the class into groups and gives each group a task.

One group is asked to study how electrons are arranged in the atoms of the elements from **Scandium to Zinc**. The learners are reminded that electrons fill orbitals according to the **Aufbau principle**, where electrons occupy lower energy levels before filling higher ones, and that in the first transition series electrons begin to fill the **3d subshell** after the **4s orbital**.

Task: As a chemistry student, use your knowledge of atomic structure to **write down the electronic configuration of the elements from Scandium (Sc) to Zinc (Zn).**

Response:

The electronic configuration of elements from atomic number 21 to 30 are written as follows.

Scandium, Sc (21)	$1s^22s^22p^63s^23p^6\mathbf{3d^14s^2}$
Titanium, Ti (22)	$1s^22s^22p^63s^23p^6\mathbf{3d^24s^2}$
Vanadium, V (23)	$1s^22s^22p^63s^23p^6\mathbf{3d^34s^2}$
Chromium, Cr (24)	$1s^22s^22p^63s^23p^6\mathbf{3d^54s^1}$
Manganese, Mn (25)	$1s^22s^22p^63s^23p^6\mathbf{3d^54s^2}$
Iron, Fe (26)	$1s^22s^22p^63s^23p^6\mathbf{3d^64s^2}$
Cobalt, Co (27)	$1s^22s^22p^63s^23p^6\mathbf{3d^74s^2}$
Nickel, Ni (28)	$1s^22s^22p^63s^23p^6\mathbf{3d^84s^2}$
Copper, Cu (29)	$1s^22s^22p^63s^23p^6\mathbf{3d^{10}4s^1}$
Zinc, Zn (30)	$1s^22s^22p^63s^23p^6\mathbf{3d^{10}4s^2}$

Writing Electronic Configuration of Ions

The electronic configuration of an ion is obtained from the electronic configuration of the atom of the element from which it is derived.

(a) Cations are formed when atoms lose electrons. This implies that the electronic configuration a cation is written by removing a specified number of electrons from the outermost orbital(s) of the electronic configuration of the atom from which it is derived.

e.g. the electronic configurations of Fe^{2+} and Fe^{3+} are $1s^22s^22p^63s^23p^63d^6$ and $1s^22s^22p^63s^23p^63d^5$ respectively. This means that two electrons are removed from the outermost 4s orbital of the atom of iron $1s^22s^22p^63s^23p^63d^64s^2$ to form iron(II) ions $1s^22s^22p^63s^23p^63d^6$. Similarly, two electrons are removed from the outermost 4s orbital and one electron is removed from the 3d orbitals of the atom of iron $1s^22s^22p^63s^23p^63d^64s^2$ to form iron(III) ions $1s^22s^22p^63s^23p^63d^5$.

(b) Anions are formed by gaining of electrons by atoms. This implies that the electronic configuration of an anion is written by addition of electrons to the electronic configuration of the atom from which it is derived.

e.g. the electronic configuration of the phosphide ion P^{3-} is $1s^22s^22p^63s^23p^6$. This means that three electrons are added to the outermost orbitals of the phosphorous atom, P $1s^22s^22p^63s^23p^3$ to obtain the electronic structure of the phosphide ion, P^{3-}.

Using the idea above the electronic structure of these metal ions becomes;

Element			Electronic structure of atom	Common ion	Electronic structure of ion
Scandium	21	Sc	$1s^22s^22p^63s^23p^63d^14s^2$	Sc^{3+}	$1s^22s^22p^63s^23p^6$
Titanium	22	Ti	$1s^22s^22p^63s^23p^63d^24s^2$	Ti^{4+}	$1s^22s^22p^63s^23p^6$
Vanadium	23	V	$1s^22s^22p^63s^23p^63d^34s^2$	V^{3+}	$1s^22s^22p^63s^23p^63d^2$
Chromium	24	Cr	$1s^22s^22p^63s^23p^63d^54s^1$	Cr^{3+}	$1s^22s^22p^63s^23p^63d^3$

Manganese	25	Mn	$1s^22s^22p^63s^23p^63d^54s^2$	Mn^{2+}	$1s^22s^22p^63s^23p^63d^5$
Iron	26	Fe	$1s^22s^22p^63s^23p^63d^64s^2$	Fe^{2+}	$1s^22s^22p^63s^23p^63d^6$
				Fe^{3+}	$1s^22s^22p^63s^23p^63d^5$
Cobalt	27	Co	$1s^22s^22p^63s^23p^63d^74s^2$	Co^{2+}	$1s^22s^22p^63s^23p^63d^7$
Nickel	28	Ni	$1s^22s^22p^63s^23p^63d^84s^2$	Ni^{2+}	$1s^22s^22p^63s^23p^63d^8$
Copper	29	Cu	$1s^22s^22p^63s^23p^63d^{10}4s^1$	Cu^{+}	$1s^22s^22p^63s^23p^63d^{10}$
				Cu^{2+}	$1s^22s^22p^63s^23p^63d^9$
Zinc	30	Zn	$1s^22s^22p^63s^23p^63d^{10}4s^2$	Zn^{2+}	$1s^22s^22p^63s^23p^63d^{10}$

Item:

Students at **Kampala Secondary School** visited a nearby **materials research laboratory** at Makerere University where scientists study metals and non-metals used in making electronics, fertilizers, and industrial materials. During the visit, the laboratory technologist explained that many substances used in industry exist in the form of **ions**.

He told the students that understanding the **electronic configuration of ions** helps scientists predict their **reactivity, bonding, and applications** in areas such as metal extraction, semiconductor production, and fertilizer manufacture.

The technologist then gave the students a worksheet containing different ions formed when atoms either **lose electrons to form cations** or **gain electrons to form anions**.

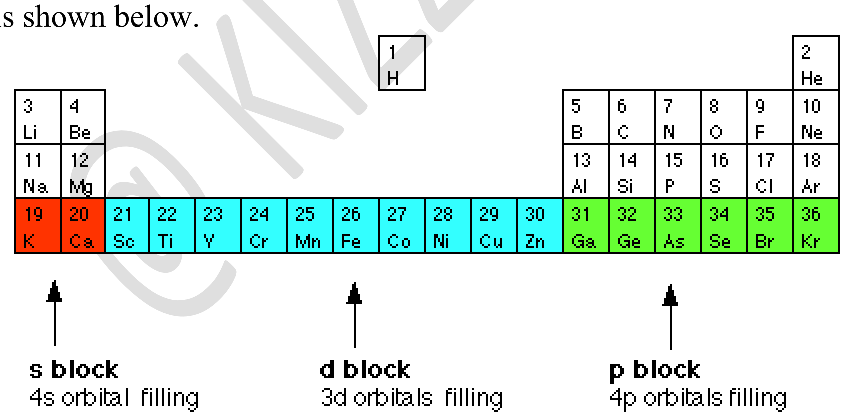
Task: As a chemistry learner, you are required to help the scientists determine the electronic arrangements of the ions listed below.

(a) Chromium(II), Cr^{2+}.
(b) Carbide ion, C^{4-}.
(c) Silicide ion, Si^{4-}.
(d) Manganese(IV), Mn^{4+}.
(e) Manganese(VII), Mn^{7+}.
(f) Nitride ion, N^{3-}.
(g) Aluminium ion, Al^{3+}.
(h) Nickel(III), Ni^{3+}.
(i) Sulphide ion, S^{2-}.
(j) Vanadium(II), V^{2+}.

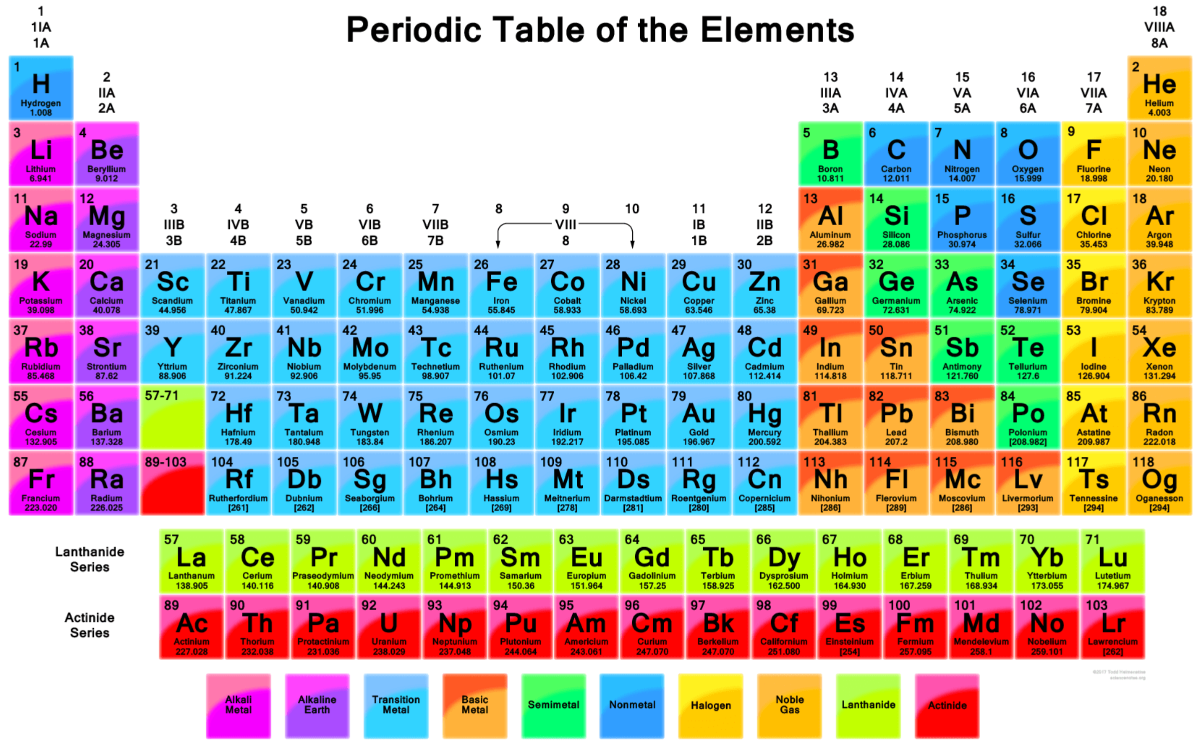
Note that: Electrons in the outer most quantum shell determine chemical behavior of elements. These electrons are involved in bonding hence referred to as **valence electrons**.

Chemical behavior of elements that have the same electron arrangement in the outermost quantum shell are related hence resemble each other. They are therefore placed in the same families referred to as groups. Groups are vertical columns in a periodic table.

The Periodic Table has got therefore 4 blocks each representing elements with the outermost electron being located in that orbital. These blocks include; the s-block, p-block, d-block and f-block. Part of the Periodic Table is shown below.



The **f block** constitutes of Lanthanides (in Pereiod 6 Lathanium being the 1st member) and Actinides (in Period 7 Actinium being the first member)



Sample Items About Electronic Structure:

1. (a) Aufbau's Principle in Electronics.

A new electronics technician is assembling a circuit board and needs to place the components in the correct order for the circuit to function.

Task: How does Aufbau's principle relate to this situation?

Response: Aufbau's principle states that electrons fill the lowest energy levels first before moving to higher ones. Similarly, in circuit board assembly, the technician must place the fundamental components (like resistors and capacitors) in the correct sequence to ensure the circuit operates efficiently. If critical base components are missing or placed incorrectly, the system may fail just as an atom cannot function properly if electrons skip lower energy orbitals and jump directly to higher ones.

(b) Aufbau's Principle in a Multi-Level Parking Garage.

A new driver enters a multi-level parking garage with multiple floors of parking spots. Most of the lower-level spots are filled, while upper levels still have space.

Task: How does this reflect Aufbau's Principle?

Response: Aufbau's Principle states that electrons fill lower energy orbitals first before occupying higher energy orbitals. Similarly, in a parking garage, drivers tend to park in the lowest available spots before moving to higher floors. Just as electrons seek the most stable (lowest energy) arrangement, drivers prefer closer, more accessible spots before going to higher, less convenient ones.

(c) Aufbau's Principle – Building a House.

Imagine you are constructing a house. The foundation must be laid first before building the walls and roof.

Task: How does this process relate to Aufbau's Principle in electron configuration?

Response: Aufbau's Principle states that electrons fill lower energy orbitals first before occupying higher ones. Similarly, when building a house, the foundation must be completed before constructing the walls and roof. If we try to build the roof before the foundation, the house would collapse—just as an atom's stability depends on electrons filling lower energy levels first before moving to higher ones.

Competency Developed: *Logical thinking and application of scientific concepts to real-life situations.*

(d) Aufbau's Principle – Filling Water Bottles.

If you are filling a water bottle, you must first fill the bottom part before reaching the top.

Task: How does this relate to Aufbau's Principle?

Response: Just as water naturally fills the bottom of a bottle before rising to the top, Aufbau's Principle states that electrons fill lower energy levels first before moving to higher ones. If you try to fill the top of the bottle first, the water will spill, just like an electron configuration cannot skip lower orbitals and remain stable.

Competency Developed: *Observation of natural processes and their relation to scientific laws.*

2. (a) Pauli's Exclusion Principle in Dormitory Room Sharing.

Two students in a boarding school are assigned to a cubicle in a small dormitory room, but the school has a rule that no two students can have the exact same ID number.

Task: How does this situation reflect Pauli's Exclusion Principle?

Response: Pauli's Exclusion Principle states that "electrons are arranged in a free atom in such a way that an orbital can accommodate a maximum of two electrons which must spin in opposite directions." i.e. no two electrons in an atom can have the same set of quantum numbers (they must differ in at least one value, such as spin). Similarly, the university enforces unique ID numbers to distinguish students, just as nature ensures that no two electrons in the same atomic orbital have identical properties. This principle helps maintain order in both the atom and the dormitory!

(b) Pauli's Exclusion Principle in Barcode Scanning.

A store has a policy where no two products can have the same barcode because each must be uniquely identified.

Task: How does this relate to Pauli's Exclusion Principle?

Response: Pauli's Exclusion Principle states that "electrons are arranged in a free atom in such a way that an orbital can accommodate a maximum of two electrons which must spin in opposite directions." i.e. no two electrons in an atom can have the same set of quantum numbers. Similarly, in a store, every product must have a unique barcode to differentiate it from others. This ensures proper organization; just as unique quantum states allow for orderly electron configurations in atoms.

(c) Pauli's Exclusion Principle – Identity Cards in School.

In a school, every student has a unique admission number to differentiate them.

Task: Why is it necessary for each student to have a different number, and how does this relate to Pauli's Exclusion Principle?

Response: Pauli's Exclusion Principle states that "electrons are arranged in a free atom in such a way that an orbital can accommodate a maximum of two electrons which must spin in opposite directions." i.e. no two electrons in an orbital of an atom can have the same set of spin quantum number. Similarly, in a school, every student is assigned a unique admission number to avoid confusion and ensure proper record-keeping. This principle ensures that electrons in an atom are uniquely identified, just like students in a school.

Competency Developed: *Understanding of uniqueness in scientific and social systems.*

(d) Pauli's Exclusion Principle – Traffic Lights and Road Rules.

Traffic lights ensure that vehicles from different directions do not move at the same time in an intersection.

Task: Why is this rule important, and how does it relate to Pauli's Exclusion Principle?

Response: Pauli's Exclusion Principle states that no two electrons in an atom can have the same set of quantum numbers in the same orbital. Similarly, traffic lights prevent two vehicles from occupying the same space at the same time to avoid accidents. Just as electrons follow strict quantum rules, road users must follow traffic laws for safety and order.

Competency Developed: *Understanding the importance of rules in both science and society.*

3. (a) Hund's Rule in Bus Seating.

A group of school students enters a nearly empty bus and begins taking seats. Most of them sit alone in separate seats rather than pairing up.

Task: How does this relate to Hund's Rule?

Response: Hund's Rule states that when electrons fill degenerate orbitals (orbitals of the same energy), they occupy separate orbitals before pairing up. This is similar to how the students prefer sitting alone in separate seats but in a regular fashion before having to share a seat with someone else. Electrons, like the students, avoid unnecessary repulsion that make the atom unstable and only pair up when no empty orbitals (or seats) are left.

(b) Hund's Rule – Passengers Boarding a Kamunye Taxi (Public Transport).

A group of passengers is boarding a matatu (public minibus). There are empty rows of seats available, but most people choose to sit alone before sitting next to someone else.

Task: How does this relate to Hund's Rule?

Response: Hund's Rule states that electrons will occupy empty orbitals of the same energy level before pairing up. Similarly, passengers prefer to sit in separate seats first before sharing with others. This minimizes discomfort and maximizes space, just as electrons try to minimize repulsion by filling separate orbitals first before pairing.

Competency Developed: *Application of scientific principles to social behavior and daily life situations.*

(c) Hund's Rule – Distributing Homework Assignments in a Group.

A teacher assigns different roles to a group of students. Each student takes one role before getting a second.

Task: How does this relate to Hund's Rule?

Response: Hund's Rule states that electrons will occupy separate orbitals before pairing up. Similarly, when distributing homework, each student takes one role before anyone gets a second one. This ensures a balanced workload and prevents one student from being overloaded, just as electrons avoid unnecessary repulsion by filling empty orbitals first.

Competency Developed: *Fair resource distribution and teamwork.*

Using Electronic Configuration to Determine the Reactivity of Elements

1. Introduction:

Reactivity of elements depends on how easily they lose, gain, or share electrons. Electronic configuration provides insight into an element's stability and its tendency to undergo chemical reactions.

2. Understanding Electronic Configuration:

Electronic configuration describes how electrons are distributed in atomic orbitals. It follows the **Aufbau Principle**, **Pauli's Exclusion Principle**, and **Hund's Rule** to fill orbitals in order of increasing energy.

e.g. **Sodium (Na) – 11 electrons**: Configuration: $1s^2\ 2s^2\ 2p^6\ 3s^1$

Chlorine (Cl) – 17 electrons: Configuration: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$

3. Relationship Between Electronic Configuration and Reactivity:

(a) Reactivity of Metals in **Groups I, II and III.** Metals lose electrons to form **positive ions (cations)**.

e.g. Typical metals of elements in **Groups I and II** are highly reactive because they have:
(i) few valence electrons. Movement of electrons involves energy. This means that when few electrons are lost, less energy is involved and hence the electrons are lost more easily and the metals become more reactive.
(ii) larger atomic radius (size). Increased atomic size leads to a longer distance between the positively charged protons in the nucleus and the negatively charged outermost electrons. The nuclear attraction for the outermost electrons reduces as the electrons are lost more easily and the metals become more reactive.
Trend: Reactivity increases **down the group** (due to increased atomic size). e.g. Alkali Metals (Group I)
Lithium (Li): $1s^2\ 2s^1 \rightarrow$ loses 1 electron $\rightarrow Li^+$
Potassium (K): $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1 \rightarrow$ loses 1 electron $\rightarrow K^+$
Potassium is more reactive than lithium because its outermost electron is farther away from the nucleus, reducing the nuclear attraction.
Trend: Reactivity decreases **across the Period** (due to increased number of valence electrons and decreased atomic size). e.g. Period 3 metallic elements (Na, Mg and Al).
Sodium (Na): $1s^2\ 2s^2\ 2p^6\ 3s^1 \rightarrow$ loses 1 electron $\rightarrow Na^+$
Magnesium (Mg): $1s^2\ 2s^2\ 2p^6\ 3s^2 \rightarrow$ loses 2 electrons $\rightarrow Mg^{2+}$
(b) Reactivity of Non-metals in **Groups IV, V, VI and VII.**
Non-metals gain electrons to form **negative ions (anions)**.
e.g Typical nonmetals of Groups VI and VII (Chalcogens and Halogens) are highly reactive because they:
(i) need only a few electrons to complete their valence shell. Movement of electrons involves energy. This means that when few electrons are gained, less energy is involved and hence the electrons are gained more easily and the non-metals become more reactive.
(ii) have smaller atomic radius (size). Decreased atomic size leads to a shorter distance between the positively charged protons in the nucleus and the negatively charged electrons being gained. The nuclear attraction for the gained electrons increases as the electrons are gained more easily and the non-metals become more reactive.
Trend: Reactivity decreases **down the group** (due to increased atomic size and making electron gain harder).
e.g. Halogens (Group VII). **Fluorine (F):** $1s^2\ 2s^2\ 2p^5 \rightarrow$ gains 1 electron $\rightarrow F^-$
Chlorine (Cl): $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5 \rightarrow$ gains 1 electron $\rightarrow Cl^-$
Fluorine is more reactive than chlorine because its gained electron experiences a stronger nuclear attraction.
Trend: Reactivity increases **across the Period** (due to increased number of valence electrons hence fewer electrons being gained and decreased atomic size). e.g. Period 3 non-metallic elements.
Sulphur (S): $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4 \rightarrow$ gains 2 electrons $\rightarrow S^{2-}$
Chlorine (Cl): $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5 \rightarrow$ gains 1 electron $\rightarrow Cl^-$
(c) Noble Gases (Group VIII). Complete valence shells (e.g., **Ne: $1s^2\ 2s^2\ 2p^6$**) make noble gases **chemically inert**. They rarely react because they have **stable electron configurations**.

4. Conclusion:

Elements react to achieve a **stable electron configuration** (often a noble gas configuration).
- Metals lose electrons (forming cations), and their reactivity increases **down a group** but decreases **across the period**.
- Non-metals gain electrons (forming anions), and their reactivity decreases **down a group** but increases **across the period**.
- Noble gases are inert due to their full outer shells.