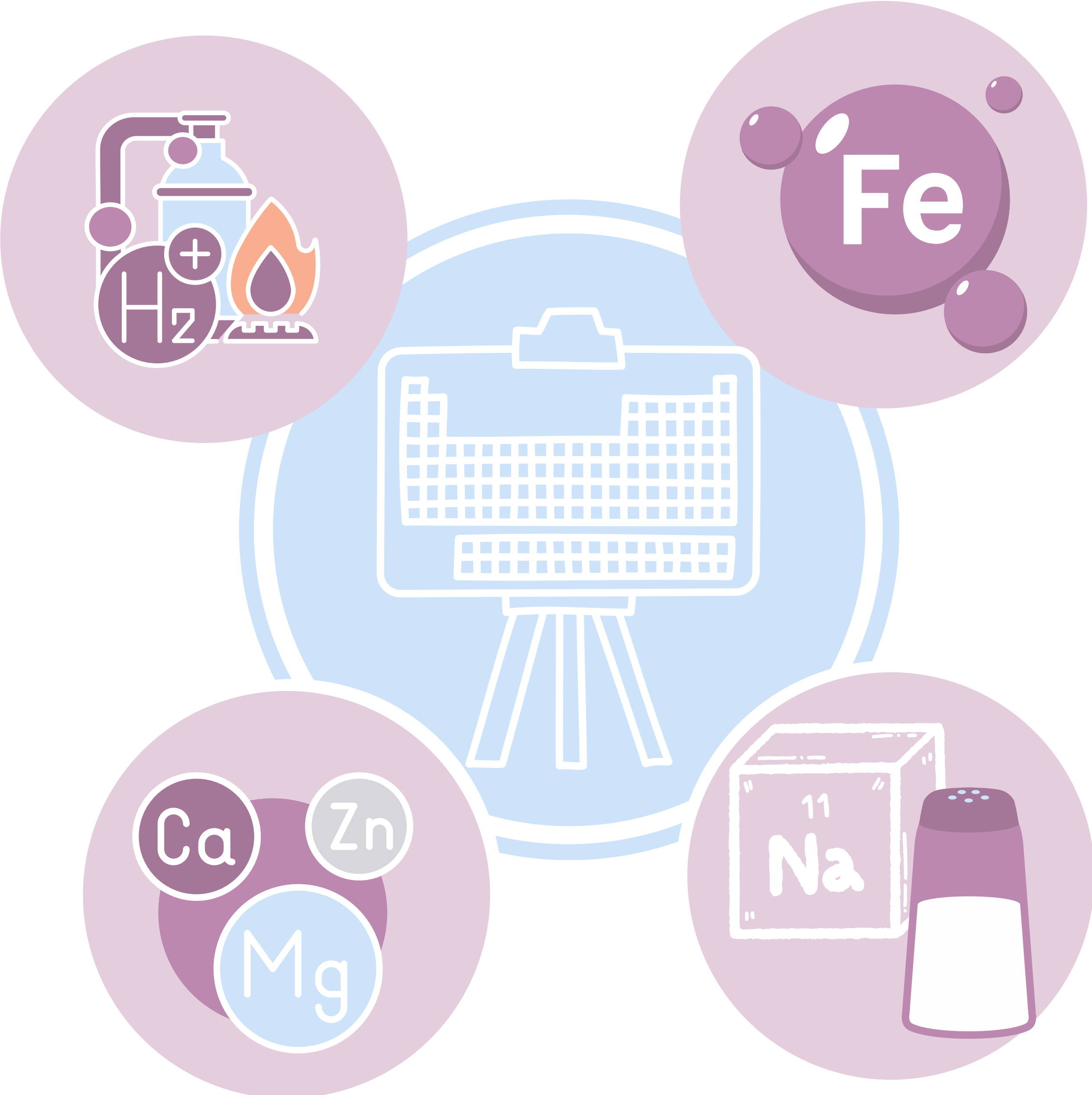
PERIODICITY IN ELEMENTS



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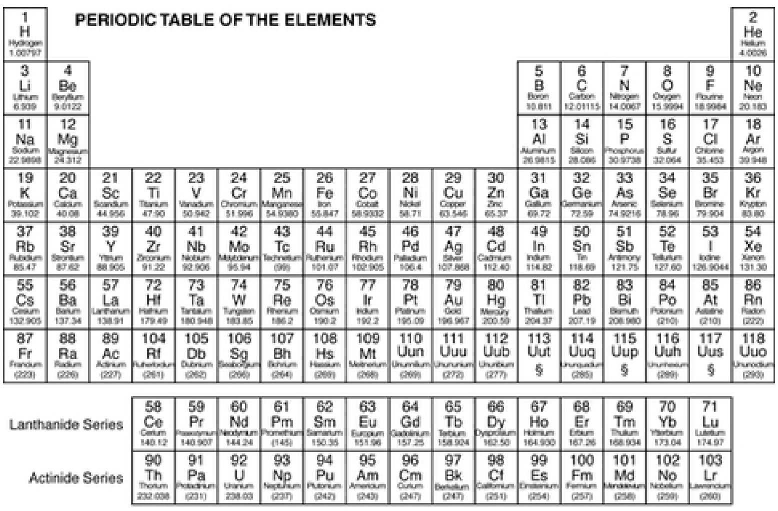
NOTES

NOTES PERIODICITY IN ELEMENTS

INTRODUCTION TO THE PERIODIC TABLE

THINGS TO REMEMBER:

- **Periodic table:** Elements arranged in order of increasing proton number.
- **Group number:** Number of elections in the outermost shell valence electrons.
- **Period number:** Number of shells of electrons.
- **Valence electrons:** Electrons in the outermost shell.
- **Metals:** Elements with 1-5 valence electrons except boron.
- **Non-Metals:** Elements with 4-7 valence electrons.
- **Inert gases:** Elements with full outermost shells.
- **Atomic number (Z):** Number of protons in the nucleus of an atom, which determines the element.
- **Mass number (A):** Sum of protons and neutrons in the nucleus.
- **Isotopes:** Atoms of the same element with the same atomic number but different mass numbers.
- **Ionization energy:** The energy required to remove an electron from an atom in its gaseous state.
- **Electronegativity:** A measure of an atom's ability to attract and bond with electrons.



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EVOLUTION OF THE PERIODIC TABLE

- The early history of ideas leading up to the periodic classification of elements is fascinating, but will not be treated here in detail.
- Those who made memorable contributions in this field are
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1. DOBERENER'S TRADS:

- In 1829, a German chemist Dobereiner arranged then known elements in groups. called triads, as each contained three elements with similar properties.
- According to this law of triads:

"The atomic mass of a middle element is the average of atomic masses of the other two elements of a triad."

- For example,

TRAID	ATOMIC MASS
Li	5
Na	23
K	39
Conclusion	Atomic mass of Na = 7+39/2 = 23

Defects:

- Dobereiner's idea of classification failed because only a few elements could be covered under triads.

2. NEWLAND'S OF OCTAVES:

- In 1864, an English chemist, John Newland classified 62 elements, known at that time, in increasing order of their atomic masses.
- According to his law of octaves.

"If the elements are arranged in the increasing order of their atomic masses, every 8th element had some properties in common with the first one."

Li(7)	Na(23)	K(39)
Be(9)	Mg(24)	Ca(40)
Be(11)	Al(27)	Ti(48)
C(12)	Si(28)	Cr(52)
N(14)	P(31)	Mn(55)
O(16)	S(32)	Fe(56)
F(19)	Cl(35.5)	Ni and Co

Defects:

- It was found that the law of octaves was applicable only up to calcium, as after calcium every 8 elements did not possess properties similar to that of the first.

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- It was assumed by Newland that only 62 elements existed in nature and no more elements would be discovered in the future. But later on, several new elements were discovered, whose properties did not fit into the law of octaves.
- The above sequence of elements was disturbed by the discovery of noble gases.

3. LOTHAR MEYER:

- Lothar Meyer conducted a series of experiments in the late 1860s and early 1870s that led him to develop his own periodic table of the elements.
- His experiments were based on the idea that the physical and chemical properties of the elements are related to their atomic weights.
- One of Meyer's most important experiments involved plotting the atomic volumes of the elements against their atomic weights.
- He found that the resulting graph showed a repeating pattern, with elements with similar properties having similar atomic volumes. This pattern is now known as the Lothar Meyer curve.
- Mever's experiments provided strong evidence for the existence of a periodic relationship between the elements. However, his periodic table was not as widely accepted as that of Dmitri Mendeleev, who had also developed a periodic table around the same time.
- Mendeleev's table was more successful because it made more accurate predictions about the properties of undiscovered elements.
- Here is a more detailed explanation of some of Meyer's experiments:
- **Atomic volume:** Meyer calculated the atomic volumes of the elements by dividing their molar masses by their densities. For example, the alkali metals (Group 1 of the periodic table) have the lowest atomic volumes, while the noble gases (Group 18 of the periodic table) have the highest atomic volumes.
- **Chemical properties:** Meyer also studied the chemical properties of the elements. He found that elements with similar atomic weights also had similar chemical properties. For example, he found that the alkali metals (Croup 1 of the periodic table) all have one valence electron, and they all form similar jonic compounds. He also found that the halogens (Group 17 of the periodic table) all have seven valence electrons, and they all form similar ionic compounds. Meyer's experiments were important because they provided strong evidence for the existence of a periodic relationship between the elements. His work helped to pave the way for the development of the modern periodic table, which is one of the most important tools in chemistry.

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4. MENDELEEV'S PERIODIC TABLE:

- In 1871, Russian chemist, Dimitri Mendeleev, give a more useful and comprehensive scheme for the classification of elements. He presented the first regular periodic table in which elements of similar chemical properties were arranged in 8 vertical columns called groups. The horizontal rows of the table were called periods.

Mendeleev's periodic law:
"If the elements are arranged in ascending order of their atomic masses, their chemical properties repeat in periodic manner."

Significance of Mendeleev's periodic table:

- Mendeleev's periodic table had a profound impact on the development of chemistry. It provided a framework for understanding the properties of the elements and their relationships to each other. The periodic table also allowed chemists to predict the properties of new elements that had not yet been discovered. Some of the specific benefits of Mendeleev's periodic table include:

1. It helped chemists to understand the relationships between the elements and their properties.

2. It allowed chemists to predict the properties of new elements that had not yet been discovered.

3. It helped chemists to develop new theories about the structure and behavior of atoms.

It provided a foundation for the development of new technologies, such as fertilizers, plastics, and pharmaceuticals.

Defects of Mendeleev's periodic table:
Although Mendeleev's periodic table was a major breakthrough, it did have some defects. Some of the specific defects of Mendeleev's periodic table include:

1. It did not explain why the elements had repeating patterns of properties.

2. It did not include all of the elements that are now known to exist

3. It placed some elements in the wrong order based on their atomic weights.

4. It left some gaps for elements that had not yet been discovered, but which Mendeleev predicted would exist.

The periodic table has been updated and refined over the years, but it remains one of the most important tools in chemistry today.

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5. MOSELEY PERIODIC LAW:

- Moseley periodic law states that:
  "The physical and chemical properties of elements are periodic function of their atomic numbers."

Periodicity:

- The repetition of properties after regular intervals is called periodicity.

- Moseley's periodic law also states that the square root of the frequency of the characteristic

- X- ray emitted by an element is proportional to its atornic number. This means that the higher the atomic number of an element, the higher the energy of its characteristic X-rays.

- Moseley discovered this law in 1913 while studying the X-ray spectra of different elements. He found that the X-ray spectrum of each element was unique, and that the frequencies of the X-rays emitted by an element could be used to determine its atomic number,

- Moseley's law was a major breakthrough in the development of the periodic table.

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Before Moseley, the periodic table was arranged by atomic mass. However, Moseley's law showed that atomic number is a more fundamental property of an element, and that the periodic table should be arranged by atomic number

Moseley's periodic law was a significant advancement in understanding the periodic table. By examining the X-ray spectra of elements, Moseley discovered that each element emits unique X-rays, which vary predictably with atomic number. This relationship provided a more accurate method for organizing elements, as opposed to

Mendeleev's method based on atomic mass. Moseley's work led to the reordering of several elements and the prediction of new ones, refining the periodic table's structure. His law highlighted the importance of atomic number in determining element properties and periodicity, reinforcing the concept that element properties recur at regular intervals.

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MODERN PERIODIC TABLE

- The modern periodic law states:

"If elements are arranged in ascending order of their atomic numbers, their chemical properties repeat in a periodic manner."

- This law was introduced by Moseley from his X-rays experiment in 1913.
- On the basis of this periodic law he arranged the element in the modern periodic table.

SALIENT FEATURES OF MODERN PERIODIC TABLE:

Groups:

Vertical columns of elements are called groups.

Properties:

- There are 8 groups in periodic table, and 18 according to the advance classification.
- They are numbered by Roman numerals and VIII.
- Each group is divided into two
- sub group A and B.
  - Sub group A contain normal and representative elements.
  - Sub group B contain less typical elements or transition elements.

Periods:

- Horizontal rows of elements are called periods.
- They are represented by Arabic numbers 1 to 7.

1st period:

- 1st period contains only 2 elements Hand He It is the shortest period of periodic table.
- It include only S block elements.

2nd and 3rd period:

- Each period contain 8 elements.
- These are called short periods
- All the elements belong to sub group A and are representative elements. In these periods every eighth element resembles in properties with first element.
- It includes S block and P block elements.

4th and 5th periods:

- Each period contains 18 elements.
- These are called Long periods.
- First 2 elements are 5 block, next 10 are D-block and next six are P-block.
- In these periods properties repeat after 18 elements e.g K is similar to Rb.

6th periods:

- This period contains 32 elements.
- It is a very long period. It contains 8 normal, 10 transition and 14 inner transition elements(lanthanides). Lanthanide is placed at the bottom of periodic table.

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7th period:

- It is last period of periodic table.
- It is incomplete and contain 28 elements
- It contains 2 normal, 10 transition 14 inner transition elements (actinides).
- All the elements of this period are reactive.

Blocks in periodic table:

The periodic table is divided into four blocks S-block P-block, D-block, and F-block. These blocks are determined by the valence electron orbital. Valence electrons are the electrons in the outermost shell of an atom, and they are the electrons that participate in chemical bonding.

s-block:

The s-block elements are located in groups IA and IIA of the periodic table, They have one or two valence electrons in the s orbital. S-block elements are generally soft, reactive metals.

p-block:

The p-block elements are located in groups IIIA-VIIIA of the periodic table. They have three to eight valence electrons in the porbital p-block elements include metals, nonmetals, and metalloids.

d-block:

The d-block elements are located in the middle of the periodic table, from groups 3-12. They have one to ten valence electrons in the d orbital, d-block elements are all metals, and they are known as transition metals.

f-block:

The f-block elements are located below the main body of the periodic table. They have one to fourteen valence electrons in the f orbital. F-block elements are all metals, and they are known as inner transition metals.

The blocks of the periodic table can be used to predict the chemical properties of elements. For example, S-block elements are generally more reactive than P-block elements, and D-block elements are generally more reactive than F-block elements.

The blocks can also be used to explain the physical properties of elements. For example, S-block and P-block elements are generally more ductile and malleable than D-block and F-block elements.

SOME FAMILIES IN PERIODIC TABLE:

Here are some families in the periodic table with their mnemonics:

Alkali metals: Group IA/1

Mnemonic:

LiNa Ki Ruby Cse Friendship hai

Elements:

Lithium (Li), sodium (Na), potassium (K), rubidium (Rb), caesium (Cs), and francium (Fr).

Alkaline earth metals: Group IIA/2

Mnemonic:

Be Mg Ca Sr Ba Ra

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Elements:
Beryllium (Be), magnesium (Mg), Calcium (Ca), strontium (Sr), barium (Ba), and radium (Ra).

Boron group: Group IIIA/13
Mnemonic:
BAGIT
Elements:
Boron (B), aluminium (Al), gallium (Gal, indium (In), and thallium (Tl)

Carbon group: Group 14 / IV A
Mnemonic:
Chemistry Sir Gives Sanki Problems
Elements:
Carbon (C), silicon (Si), germanium (Ge), Lin (Sn), and lead (Pb)

Nitrogen group: Group 15/VA
Mnemonic:
Nahi Pasand Aise Sab Elements Bhai
Elements:
Nitrogen (N), phosphorus (P). arsenic (As), antimony (Sb), and bismuth (Bi)

Oxygen group: Group 16/VIA
Mnemonic:
Oh! Style Se Tel Polish
Elements:
Oxygen (O), sulfur (S), selenium (Se), tellurium (Te), and polonium (Po)

Halogens: Group 17/VIIA
Mnemonic:
Fir Call kar Bahar Aayi Aunty
Elements:
Fluorine (F), chlorine (Cl), bromine (Br), iodine (1), and astatine (At)

METALS, NON-METALS, AND METALLOIDS IN THE PERIODIC TABLE:

The periodic table is a chart of all the known chemical elements, arranged in order of increasing atomic number. The elements are also grouped by their chemical properties, with metals on the left, non-metals on the right, and metalloids in between.

Metals:
Metals are generally shiny, good conductors of heat and electricity, malleable and ductile (meaning they can be hammered into thin sheets and drawn into wires). Metals also tend to form positive ions when they react with other elements. Some common metals include Iron, Aluminum, Copper, and Gold.

Non-metals:
Non-metals are generally dull, poor conductors of heat and electricity, and brittle. They also tend to form negative ions when they react with other elements. Some common non-metals include Oxygen, Nitrogen, Carbon, and Sulfur.

Metalloids:
Metalloids have properties that are intermediate between those of metals and non-metals. They can be good conductors of electricity, but they are not as shiny or malleable as metals. Metalloids can also form both positive and negative ions when they react with other elements. Some common metalloids include Silicon, Cermanium, and Arsenic.

NOTES

PERIODIC TRENDS

PROPERTY	DEFINITION	TREND DOWN THE GROUP	RESPONSIBLE FACTORS	TREND ALONG PERIOD FROM LEFT TO RIGHT
Atomic radius	Half of the distance between the centres of two adjacent atoms of any element.	Increases	Nuclear charge and number of shells	Decreases
Ionic radius	Distance between the centre of an ion and the outer boundary of its electron cloud.	For similar charged ions increases	Nuclear charge and number of shells	Decreases for Iso- electronic +ve and -ve
Ionization energy	Minimum quantity of energy, which is required to remove an electron from the outermost shell of its isolated, gaseous atom in its ground state.	Decreases	Nuclear charge, atomic size and shielding effect	Increases

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Electron affinity	Energy released or observed when an electron is added to a gaseous atom to form a negative ion.	Decreases	Size of atom, nuclear charge and vacancies in valence shell	Increases
Electronegativity	The tendency of an atom to attract a shared pair electron toward itself	Decreases	I.E and E.A of bonded atoms the structure of the atom and the number and kind of atoms with which it may combine	Increases
Metallic character	Tendency to lose electron and form cation and basic oxides	Increases	Atomic size and nuclear charge	Decreases
Electrical Conductance	Ability to conduct electricity	(Increases for IA and IIA) no regular trend for transition metals.	Lose electrons in outer shell	Decreases
Melting and boiling points	Specific temperature at which an element changes to liquid or gaseous state	In case of IA and IIA decreases and for VIIA increases	Number of valence electrons	I-A to IV-A increases while V to VIII decreases

NOTES

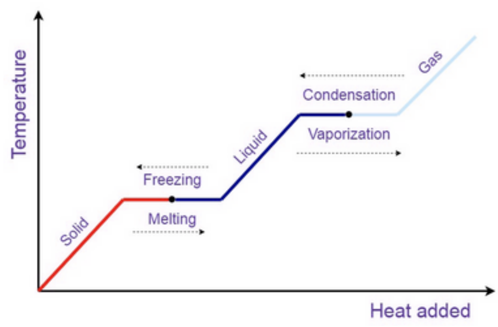
Ionic Radius

The ionic radius is the measure of an atom's ion in a crystal lattice. It is typically half the distance between two ions that are barely .

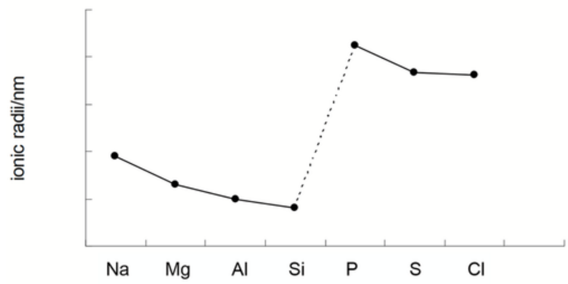
melting point

The melting point is the temperature at which a solid turns into a liquid. It is a measure of the strength of the bonds between atoms or ions in a solid. .

Ionic radius



Melting point



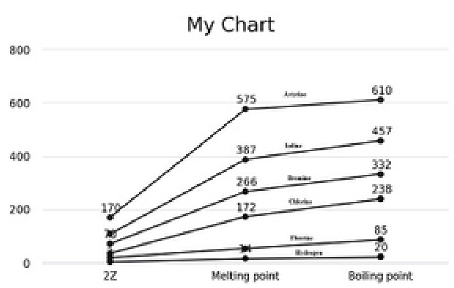
boiling point

The boiling point is the temperature at which a liquid changes to a gas (boils). It is a measure of the strength of the intermolecular forces present in the substance.

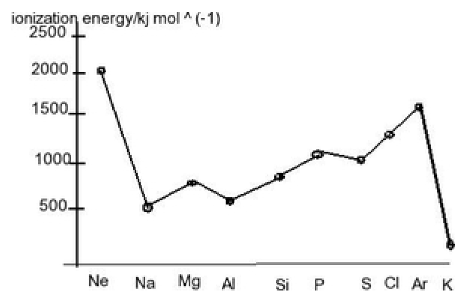
ionization energy

Ionization energy is the energy required to remove an electron from a gaseous atom or ion. The first ionization energy is the energy needed to remove the first electron, the second ionization energy is for the removal of the second electron, and so on.

Boiling point



ionization energy



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POSITION OF HYDROGEN

Position of an element in periodic table is determined by electronic configuration and properties. Hydrogen resembles the elements of group, I-A, IV-A and VII-A in the same respects. Properties of Hydrogen do not completely match anyone of the above mention groups. That's why position of electron is still undecided.

HYDROGEN AND I-A GROUP (ALKALL METALS):

Similarities:

- Hydrogen and alkali metals have one electron in their valence "s" subshell (ns)).
- Valance shell electronic configuration of hydrogen and alkali metals is same.
- Both hydrogen and alkali metals by using their valance electron form unipositive ion.
- HH+ + 10- MM++ le- Both hydrogen and alkali metals have strong tendency to combine with the electronegative elements such as halogens.
- Both form ionic compounds which dissociate in water(ionize in water).
- Both hydrogen and alkali metals are good reducing agents.

Dissimilarities:

HYDROGEN	I-A GROUP
It is a non metal.	The elements of this group are metals.
Hydrogen is a gas at room temperature.	Alkali metals are solid at room temperature
Hydrogen does not lose electrons as easily.	Due to low I.E., they lose electron easily as most of the alkali metals.
It gains electrons.	They do not gain electrons.
Hydrogen exists in molecular form i.e. H2 in open temperature.	They do not exist in the molecular form.
Hydrogen is unstable in water and form hydronium ion.	Sodium and potassium are stable in water.

HYDROGEN AND VII-A GROUP (HALOGENS):

Similarities:

- Both Hydrogen and halogen gain one electron to complete their valence shell, forming negative ions.

  H + 1e ----> H-

  X + 1e ----> X-
- Both Hydrogen and halogens are non metal in true sense.
- Both exist as molecular form in open temperature i.e H_2, F_2.

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- Hydrogen is the gas like that of most halogen i.e fluorine, chlorine.
- Both hydrogen and hydrogen form stable ionic compound with alkali metals

H_2 + 2Na ----> 2NaH

Cl_2 + 2Na ---->2NaCl

Dissimilarities:

HYDROGEN	VII-A (HALOGENS)
It has one electron in its valence shell.	Halogens have 7 electrons in their valance shell.
It losses its valance electrons to form hydrogen ion.	They cannot lose electrons.
It is s block element.	They belong to p block.
It combines with oxygen to form stable oxides(H2O).	They cannot formstable oxides e.g (ClO2,Cl2O7).
It is a gas at room temperature.	Bromine is a liquid while iodine is solid at room temperature.

HYDROGEN AND IV-A GROUP (CARBON FAMILY):

Similarities:

- Hydrogen has half-filled outermost shell like the elements of the IV-A group.
- Both Hydrogen and elements of the IV-A group combine with other elements for covalent bond.
- Both Hydrogen and elements of the IV-A group combine with other elements for covalent bond.
- Like Carbon, Hydrogen also show remarkable reducing property.
- Ionization energy and electron affinity of hydrogen and carbon are comparable.
- Electronegativity of Hydrogen and Carbon family are almost similar. [H=2.1,C=2.5,Si=2.4].

HYDROGEN	IVA GROUP ELEMENTS
Hydrogen belongs to s-Block.	Elements of this group belongs to p-block.
It is a gas at room temperature.	IV-A group elements are solid at room temperature.
It needs one electron to complete its valence shell.	Elements of this group needs 4 electron to complete their valance shell.
It does not form long chain compounds.	Carbon and silicon from long chain compound when they combine with their atoms.
Due to one electron in the valence shell, only one element at a time.	Carbon can simultaneously from bonds with more hydrogen can combine with than one element.

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CONCLUSION:

- Some of the properties of Hydrogen are similar to those of elements of I-A,IV-A and VII-A groups. But Hydrogen is a unique element whose property do not match exactly with any of the group in the periodic table. However due to partial resemblance in properties with alkali metals and monovalent nature Hydrogen is placed at the top of elements in group I-A.

TEST YOURSELF

Q1 HYDROGEN IS PLACED IN GROUP 1 OF THE PERIODIC TABLE BUT ALSO SHOWS SIMILARITIES WITH ELEMENTS OF WHICH OTHER GROUP?

A) Group 2
B) Group 14
C) Group 16
D) Group 17

Q2 WHICH OF THE FOLLOWING ELEMENTS HAS THE HIGHEST BOILING POINT?

A. Sodium (Na)
B. Magnesium (Mg)
C. Aluminum (Al)
D. Silicon (Si)

NOTES

ELECTRONIC CONFIGURATION

In order to understand the distribution of electron in an atom which should know the following facts;

1. Any degenerate orbital like in orbital s, p, d and f can have at most two electrons.
2. The maximum number of electrons that can be accommodated in a shell is given by 2n^2 formula. where n is principal quantum number and it cannot have zero value. Moreover, following rules have been adopted to distribute the electrons in subshell or orbital.
   1. n+l rule
   2. Aufbau principle
   3. Pauli's exclusion principle
   4. Hund's rule

1. n+l RULE:

This rule is also called the Wiswesser rule. The (n + l) rule is a mnemonic used to remember the order in which atomic orbitals are filled during the construction of the ground state electron configuration of the elements The rule states that "Orbitals with lower (n + l) values are filled first."

Where

n is the principal quantum number, which represents the energy level of the orbital.

l is the azimuthal quantum number, which represents the shape of the orbital.

Mnemonics:

Here are some mnemonics for remembering the (n+1) rule:

a. "No low levels left behind."

This mnemonic reminds us that orbitals with lower (n-1) values are filled first.

b. "Newer lower levels first."

This mnemonic reminds us that orbitals with lower (n+1) values are filled first, even if they are in a higher energy level.

c. "Never leave low levels."

This mnemonic is a simpler version of the first two mnemonics.

Examples: Here are some examples of how the (n+1) rule is used to predict the order in which atomic orbitals are filled.

- 1s: n = 1, l = 0; (n + l) = 1; filled first.
- 2s: n = 2, l = 0; (n + l) = 2; filled second.
- 2p: n = 2,1=1; (n + l) = 3; filled third.
- 3s: n = 3,1=0; (n + l) = 3; filled fourth.
- 3p: n = 3,1 = 1; (n + l) = 4; filled fifth.

Note that the (n + l) rule is not a perfect predictor of orbital filling order. In some cases other factors, such as electron-electron repulsion can override the (n-1) rule. However, the (n + 1) rule is a good general rule of thumb for predicting orbital filling order.

NOTES

1s,2s,2p,3s,3p, 4s, 3d, 4p,5s, 4d, 5p,6s,4f,5d,6p, 7s and so on. To learn these, mnemonics is given below;

School, school, public school, public school, division public school, division public school, federal division public school, federal division public school.

ORBITAL	n	l	n+l
1s	1	0	1
2s	2	0	2
2p	2	1	3
3s	3	0	3
3p	3	1	4
3d	3	2	5
4s	4	0	4
4p	4	1	5
4d	4	2	6
4f	4	3	7
5s	5	0	5
5p	5	1	6
5d	5	2	7
5f	5	3	8
6s	6	0	6
6p	6	1	7
6d	6	2	8
7s	7	0	7

NOTES

2. AUFBAU PRINCIPLE:

The Aufbau principle, also known as the Aufbau rule, states that electrons fill subshells of the lowest available energy, then they fill subshells of higher energy. For example, the is subshell is filled before the 2s subshell is occupied. in this way, the electrons of an atom or ion form the most stable electron configuration possible.

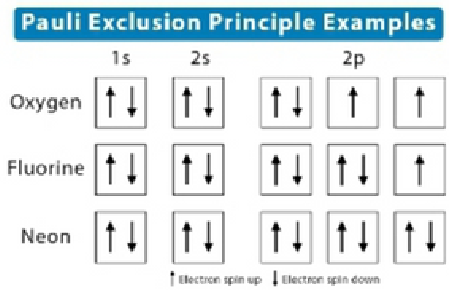
Exceptions to the Aufbau principle: There are a few exceptions to the Aufbau principle. These exceptions occur when electron-electron repulsion overrides the tendency of electrons to fill lower-energy orbitals first. One exception is the case of Chromium. The Aufbau principle would predict that the valence shell electronic configuration of chromium is 3d4 452. However, the actual valence shell electronic configuration is 3d5 451. This is because the 3d orbitals are closer to the nucleus than the 4s orbitals, but the 3d orbitals are also more negatively charged. The negative charge of the 3d electrons repels each other, making it more stable for the 3d orbitals to be only half filled.

Another exception is the case of Copper. The Aufbau principle would predict that the valence shell electronic configuration of copper is 3d9 4s2. However, the actual valence shell electronic configuration is 3010 451. This is because the 3d orbitals are more stable when they are filled with 10 electrons.

3. PAULI'S EXCLUSION PRINCIPLE:

Pauli's exclusion principle states that:

"No two electrons in an atom can have the same set of four quantum numbers."



The four quantum numbers that uniquely identify an electron in an atom are:

Principal quantum number (n):
This quantum number represents the energy level of the electron.

Azimuthal quantum number (l):
This quantum number represents the shape of the electron's orbital.

Magnetic quantum number (m):
This quantum number represents the orientation of the electron's orbital under the influence of magnetic field.

Spin quantum number(s):
This quantum number represents the spin of the electron.

Hence the Pauli exclusion principle can be summarized as follows:

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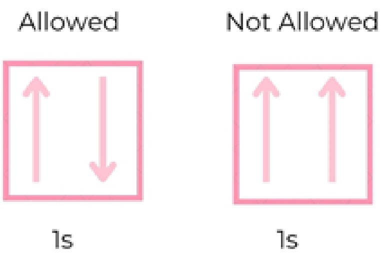
"No two electrons in an atom can have the same set of four quantum numbers."

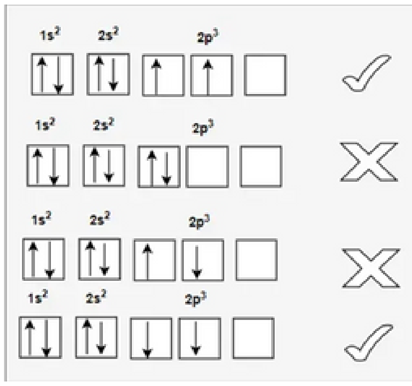
This means that each electron in an atom must have its own unique set of quantum numbers.

The Pauli exclusion principle has a number of important consequences. For example, it explains why the electron shells of atoms are filled in a specific order. It also explains why atoms cannot collapse on top of themselves, since the Paull exclusion principle prevents the electrons from occupying the lowest energy orbital.



5. Hund's rule:

Hund's rule is a rule of quantum mechanics that states that, "If degenerate orbitals are available and more than one electrons are to be placed in them they should be placed in separate orbital with the same speed rather than putting them in the same orbital with opposite spins."



TEST YOURSELF

Q3 WHICH OF THE FOLLOWING STATEMENTS BEST DESCRIBES HUND'S RULE?

A) Electrons fill orbitals starting with the lowest energy level first.
B) No two electrons in an atom can have the same set of four quantum numbers.
C) Every orbital in a subshell is singly occupied with one electron before any one orbital is doubly occupied, and all electrons in singly occupied orbitals have the same spin.
D) The total energy of the electrons in an atom is minimized by filling orbitals in order of increasing energy.

ANSWER KEY

Q.1
Answer: D
Explanation: Hydrogen is placed in Group 1 of the periodic table because it has one electron in its outermost shell, similar to the alkali metals (e.g., lithium, sodium, potassium). However, hydrogen also shows similarities with Group 17 elements (the halogens, e.g., fluorine, chlorine) in its ability to gain an electron to form a negative ion

Q.2
Answer: D
Explanation: As you move across Period 3 from left to right, the boiling points generally increase. Silicon, being a metalloid with a giant covalent structure, has a much higher boiling point compared to the metals (Na, Mg, Al) in the same period due to the strong covalent bonds between silicon atoms.

Q.3
Answer: C
Explanation: Hund's rule states that every orbital in a given subshell is singly occupied with one electron before any orbital is doubly occupied. This minimizes electron-electron repulsion within an atom. Additionally, all electrons in singly occupied orbitals must have the same spin to maximize the total spin of the electrons. This is essential for understanding the electron configuration of atoms, particularly when dealing with elements in the same group of the periodic table. This principle helps in determining the ground-state electronic configuration of atoms and is crucial for predicting the chemical behavior of elements.