SAMPLE CONTENT

FOCUS CHEMISTRY

BASED ON THE LATEST TEXTBOOK OF MAHARASHTRA STATE BOARD

Build Powerful Concepts

- Quick Review
- Important Formulae
- Complete Coverage of Textual and Intext Questions
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STD XI

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Focus

CHEMISTRY

Std. XI Sci.

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CONTENTS

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Sample Content

4 Structure of Atom

Quick Review

➢ **Properties of subatomic particles:**

Name	Symbol	Absolute charge/C	Relative charge	Symbol for charge	Mass/kg	Mass/u	Approximate mass/u
Electron	e^-	-1.6022×10^{-19}	–1	–e	9.10938×10^{-31}	0.00054	0
Proton	p	$+1.6022 \times 10^{-19}$	+1	+e	1.6726×10^{-27}	1.00727	1 u
Neutron	n	0	0	—	1.67493×10^{-27}	1.00867	1 u

➢ **Isotopes, isobars and isotones:**

Atomic species	Similarities	Differences	Examples
Isotopes	Isotopes have same: • atomic number (Z) • number of protons • number of electrons • electronic configuration • position in the periodic table • chemical properties (similar)	Isotopes have different: • mass number (A) • number of neutrons • physical properties	$^{1}_{1}H$, $^{2}_{1}H$, $^{3}_{1}H$
Isobars	Isobars have same: • mass number (A) • number of nucleons	Isobars have different: • atomic number (Z) • number of protons, electrons and neutrons • electronic configuration • position in the periodic table • chemical properties	$^{40}_{18}Ar$, $^{40}_{19}K$, $^{40}_{20}Ca$
Isotones	Isotones have same: number of neutrons	Isotones have different: • Atomic number and mass number • protons and electrons • electronic configuration • position in the periodic table	• $^{30}_{14}Si$, $^{31}_{15}P$, $^{32}_{16}S$ • $^{39}_{19}K$, $^{40}_{20}Ca$

Sample Content

- **Evolution of quantum theory:**

Classical theory
Matter is composed of particles and radiation has wave nature.

↓ ***Einstein*** | ***Planck***

Energy is Quantized.

↓ ***Bohr***

Line spectra (Bohr's H atom)

↓ ***De Broglie***

Matter wave

↓ ***Heisenberg***

Uncertainty principle

↓ ***Schrodinger***

Quantum theory
Radiation and matter both have particle and wave nature.

- **Series of emission spectral lines for hydrogen:**

Series	n_1	n_2	Spectral region
Lyman	1	2,3….	Ultraviolet
Balmer	2	3,4,….	Visible
Paschen	3	4,5,….	Infrared
Bracket	4	5,6,….	Infrared
Pfund	5	6,7,….	Infrared

- **Quantum numbers:**

Quantum numbers

Principal quantum number (n)	**Azimuthal (or subsidiary quantum number (l))**	**Magnetic orbital quantum number (m_l)**	**Spin quantum number (m_s)**
'n' is a positive integer and has values 1, 2, 3, 4 and so on.	Can have the positive values between 0 and (n – 1).	Can have values between $-l$ and $+l$ through 0.	Can have values +1/2 for clockwise spin and –1/2 for anticlockwise spin.

Important Formulae

1. **Number of neutrons (N) = A – Z**
   where,
   A = Mass number
   Z = Atomic number

2. **Frequency (ν)** $= \dfrac{c}{\lambda}$
   where,
   ν = Frequency of electromagnetic radiation
   c = Speed of light $= 3 \times 10^{8}\ \text{m s}^{-1}$
   λ = Wavelength of electromagnetic radiation

3. **Wave number** $(\overline{v}) = \dfrac{1}{\lambda}$

4. **Energy of quantum of radiation (E)** $= h\nu = h\dfrac{c}{\lambda} = hc\overline{v}$
   where, h = Planck's constant $= 6.626 \times 10^{-34}$ J s

Sample Content

5. **Wave number of series of hydrogen spectrum:**
$\bar{v} = R\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$ cm^{-1}
where,
$\bar{v}$ = wave number
R = Rydberg constant = 109677 cm^{-1}
n_1 = lower energy level
n_2 = higher energy level

6. **Bohr's frequency rule:**
$\nu = \frac{\Delta E}{h} = \frac{E_2 - E_1}{h}$
where,
E_1 = Energy of lower energy state
E_2 = Energy of higher energy state

7. **Energy of an electron (E_n) in the n^{th} orbit of hydrogen atom:**
$E_n = -2.18 \times 10^{-18}\left(\frac{Z^2}{n^2}\right)$ J
where, Z = Atomic number

8. **Radius of Bohr's orbit (r)** = $52.9 \times \frac{n^2}{Z}$ pm

9. **de Broglie equation:** $\lambda = \frac{h}{mv} = \frac{h}{p}$
p = mv = momentum of particle

10. **Heisenberg's uncertainty principle:**
$\Delta x\, \Delta p_x \geq \frac{h}{4\pi}$
where,
Δx = uncertainty in position,
Δp_x = uncertainty in momentum
$\Delta x \times \Delta(mv_x) \geq \frac{h}{4\pi}$ $(\because p = mv_x)$
$\therefore$ $\Delta x \times m \times \Delta v_x \geq \frac{h}{4\pi}$

11. **Maximum number of electrons in each principal shell** = $2n^2$ electrons

12. **Maximum number of electrons in each subshell** = $2(2l + 1)$ electrons

4.1 Subatomic particles

Q.1. Can you recall? *(Textbook page no. 35)*

i. What is the smallest unit of matter?
Ans: The smallest unit of matter is atom.

ii. What is the difference between molecules of an element and those of a compound?
Ans: The molecules of an element are made of atoms of same element while the molecules of a compound are made of atoms of different elements.

iii. Does an atom have any internal structure or is it indivisible?
Ans: Yes, an atom has internal structure. Different subatomic particles such as protons, electrons and neutrons constitute an atom. So, it is divisible.

iv. Which particle was identified by J. J. Thomson in the cathode ray tube experiment?
Ans: Electron was identified by J.J. Thomson in the cathode ray tube experiment.

Sample Content

v. **Which part of an atom was discovered by Ernest Rutherford from the experiment of scattering of α-particles by gold foil?**

Ans: Nucleus of an atom was discovered by Ernest Rutherford from the experiment of scattering of α-particles by gold foil.

4.2 Atomic number and atomic mass number

***Q.2. If an element 'X' has mass number 11 and it has 6 neutrons, then write its representation.**

Ans: The representation of the given element is ${}^{11}_{7}X$.

Reading between the lines

Element is represented as ${}^{A}_{Z}X$.
Now, $A = Z + N$
$\therefore$ $Z = A - N = 11 - 6 = 7$
Hence, element is represented as ${}^{11}_{7}X$.

***Q.3. Match the following:**

	Element		No. of neutron
i.	${}^{40}_{18}Ar$	a.	7
ii.	${}^{14}_{6}C$	b.	21
iii.	${}^{40}_{19}K$	c.	8
iv.	${}^{14}_{7}N$	d.	22

Ans: i – d, ii – c, iii – b, iv – a

4.3 Isotopes, isobars and isotones

***Q.4. Complete the following information about the isotopes in the chart given below:**

Substance	Mass number	Number of		
		Protons	Neutrons	Electrons
Carbon-14	----------	----------	----------	----------
Lead-208	----------	----------	----------	----------
Chlorine-35	----------	----------	----------	----------
Uranium-238	----------	----------	----------	----------
Oxygen-18	----------	----------	----------	----------
Radium-223	----------	----------	----------	----------

(Hint: Refer to the Periodic Table if required)

Ans:

Substance	Mass number	Number of		
		Protons	Neutrons	Electrons
Carbon-14	14	6	8	6
Lead-208	208	82	126	82
Chlorine-35	35	17	18	17
Uranium-238	238	92	146	92
Oxygen-18	18	8	10	8
Radium-223	223	88	135	88

Sample Content

***Q.5. Differentiate between isotopes and isobars.**

Ans:

No.	Isotopes	Isobars
i.	Isotopes are atoms of same element.	Isobars are atoms of different elements.
ii.	They have same atomic number but different atomic mass number.	They have same atomic mass number but different atomic numbers.
iii.	They have same number of protons but different number of neutrons.	They have different number of protons and neutrons.
iv.	They have same number of electrons.	They have different number of electrons.
v.	They occupy same position in the modern periodic table.	They occupy different positions in the modern periodic table.
vi.	They have similar chemical properties.	They have different chemical properties.
e.g.	${}^{12}_{6}C$ and ${}^{14}_{6}C$	${}^{14}_{6}C$ and ${}^{14}_{7}N$

***Q.6. Define isotones.**

Ans: ***Isotones*** *are defined as the atoms of different elements having same number of neutrons in their nuclei.*

e.g. ${}^{11}_{5}B$ and ${}^{12}_{6}C$ having 6 neutrons each are isotones.

4.4 Drawbacks of Rutherford atomic model

***Q.7. Write the drawbacks of Rutherford's model of an atom.**

Ans: Drawbacks of Rutherford's model of an atom:

i. Rutherford's model of an atom resembles the solar system with the nucleus playing the role of the massive sun and the electrons are lighter planets. Thus, according to this model, electrons having negative charge revolve in various orbits around the nucleus. However, the electrons revolving about the nucleus in fixed orbits pose a problem. Such orbital motion is an accelerated motion accompanied by a continuous change in the velocity of electron as noticed from the continuously changing direction. According to Maxwell's theory of electromagnetic radiation, accelerated charged particles would emit electromagnetic radiation. Hence, an electron revolving around the nucleus should continuously emit radiation and lose equivalent energy. As a result, the orbit would shrink continuously and the electron would come closer to the nucleus by following a spiral path. It would ultimately fall into the nucleus. Thus, Rutherford's model has an intrinsic instability of atom. However, real atoms are stable.

ii. Rutherford's model of an atom does not describe the distribution of electrons around the nucleus and their energies.

4.5 Developments leading to the Bohr's atomic model

***Q.8. Name the element that shows simplest emission spectrum.**

Ans: The element that shows simplest emission spectrum is hydrogen.

4.6 Bohr's model for hydrogen atom

***Q.9. Write postulates of Bohr's theory of hydrogen atom.**

Ans: Postulates of Bohr's theory of hydrogen atom:

i. The electron in the hydrogen atom can move around the nucleus in one of the many possible circular paths of fixed radius and energy. These paths are called orbits, stationary states or allowed energy states. These orbits are arranged concentrically around the nucleus in an increasing order of energy.

ii. The energy of an electron in the orbit does not change with time. However, the electron will move from a lower stationary state to a higher stationary state if and when the required amount of energy is absorbed by the electron. Energy is emitted when electron moves from a higher stationary state to a lower stationary state. The energy change does not take place in a continuous manner.

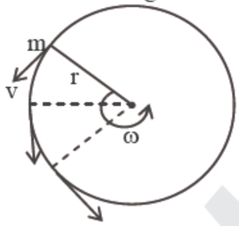
iii. The frequency of radiation absorbed or emitted when transition occurs between two stationary states that differ in energy by ΔE is given by the following expression:

$$\nu = \frac{\Delta E}{h} = \frac{E_2 - E_1}{h} \quad .....(1)$$

Where E_1 and E_2 are the energies of the lower and higher allowed energy states respectively. This expression is commonly known as **Bohr's frequency rule**.

iv. The angular momentum of an electron in a given stationary state can be expressed as $\mathbf{mvr = n \times h/2\pi}$
where, n = 1, 2, 3
Thus, an electron can move only in those orbits for which its angular momentum is integral multiple of $h/2\pi$. Thus, only certain fixed orbits are allowed.

Note:
Angular momentum is a product of moment of inertia (I) and angular velocity (ω).
Angular momentum = $I \times \omega$
However, $I = m r^2$ and $\omega = v/r$
$\therefore$ Angular momentum = $mr^2 \times v/r = mvr$



Q.10. Just Think *(Textbook page no. 41)*
What does the negative sign of electron energy convey?

Ans: Negative sign for the energy of an electron in any orbit in a hydrogen atom indicates that the energy of the electron in the atom is lower than the energy of a free electron at rest. A free electron at rest is an electron that is infinitely far away from the nucleus and is assigned the energy value of zero.
As the electron gets close to the nucleus, value of 'n' decreases and E_n becomes large in absolute value and more negative. The negative sign corresponds to attractive forces between electron and nucleus.

***Q.11. Mention the demerits of Bohr's atomic model.**

Ans: Demerits of Bohr's atomic model:

i. Bohr's atomic model (theory) failed to account for finer details of the atomic spectrum of hydrogen as observed in sophisticated spectroscopic experiments.
ii. Bohr's atomic model (theory) was unable to explain the spectrum of atoms other than hydrogen.
iii. Bohr's atomic model (theory) could not explain the splitting of spectral lines in the presence of a magnetic field (Zeeman effect) or electric field (Stark effect).
iv. Bohr's atomic model (theory) failed to explain the ability of atoms to form molecules by chemical bonds.

***Q.12. State Heisenberg uncertainty principle.**

Ans: Heisenberg uncertainty principle states that *"It is impossible to determine simultaneously, the exact position and exact momentum (or velocity) of an electron"*.

4.7 Quantum mechanical model of atom

***Q.13. Give the names of quantum numbers.**

Ans: The four quantum numbers are: principal quantum number (n), azimuthal or subsidiary quantum number (l), magnetic quantum number (m_l) and electron spin quantum number (m_s).

***Q.14. Write a note on principal quantum number.**

Ans: Principal quantum number (n):

i. Principal quantum number indicates the principal shell or main energy level to which the electron belongs.
ii. It is denoted by 'n' and is a positive integer with values 1, 2, 3, 4, 5, 6, ….
iii. A set of atomic orbitals with given value of 'n' constitutes a single shell. These shells are also represented by the letters K, L, M, N, etc.
iv. With increase of 'n', the number of allowed orbitals in that shell increases and is given by n^2.
v. The allowed orbitals in the first four shells are given below:

Principal quantum number (n)	Shell symbol	Allowed number of orbitals (n^2)	Size of shell
1	K	1	↓ increases
2	L	4	
3	M	9	
4	N	16	

vi. As the value of 'n' increases, the distance of the shell from the nucleus increases and the size of the shell increases. Its energy also goes on increasing.

Sample Content

***Q.15. Explain in brief the significance of azimuthal quantum number.**

Ans: Azimuthal quantum number (*l*):

i. Azimuthal quantum number is also known as subsidiary quantum number and is represented by letter *l*.

ii. It represents the subshell to which the electron belongs. It also defines the shape of the orbital that is occupied by the electron.

iii. Its value depends upon the value of principal quantum number 'n'. It can have only positive values between 0 and $(n - 1)$.

iv. Atomic orbitals with the same value of 'n' but different values of '*l*' constitute a subshell belonging to the shell for the given 'n'. The azimuthal quantum number gives the number of subshells in a principal shell. The subshells having *l* equal to 0, 1, 2,3 … are represented by symbols s, p, d, f, … respectively.

Principal shell	Value of n	Permissible value of *l*	Possible subshell	Number of subshells in shell
K	1	0	s	1
L	2	0, 1	s, p	2
M	3	0, 1, 2	s, p, d	3
N	4	0, 1, 2, 3	s, p, d, f	4

***Q.16. Write orbital notations for electrons in orbitals with the following quantum numbers.**

i. $n = 2, l = 1$ ii. $n = 4, l = 2$ iii. $n = 3, l = 2$

Ans:

i. 2p ii. 4d iii. 3d

***Q.17. If n = 3, what are the quantum number *l* and m_l?**

Ans: For a given n, $l = 0$ to $(n - 1)$ and for given l, $m_l = -l$...., 0, ... $+ l$.

Therefore, the possible values of l and m_l for $n = 3$ are:

Value of n	Value of *l*	Values of m_l
3	0	$m_l = 0$
	1	$m_l = -1$ $m_l = 0$ $m_l = +1$
	2	$m_l = -2$ $m_l = -1$ $m_l = 0$ $m_l = +1$ $m_l = +2$

***Q.18. Using concept of quantum numbers, calculate the maximum numbers of electrons present in the 'M' shell. Give their distribution in shells, subshells and orbitals.**

Ans:

i. Each main shell contains a maximum of $2n^2$ electrons.
For 'M' shell, $n = 3$.
Therefore, the maximum numbers of electrons present in the 'M' shell $= 2 \times (3)^2 = 18$.

ii. The distribution of these electrons in shells, subshells and orbitals can be given as follows:

Value of n	Values of *l*	Values of m_l	Values of m_s
3	**0**	**0**	±½
	1	**–1**	±½
		0	±½
		+1	±½
	2	**–2**	±½
		–1	±½
		0	±½
		+1	±½
		+2	±½

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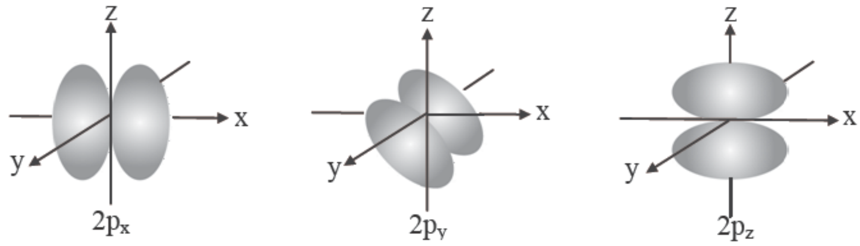
***Q.19. Draw shapes of 2s and 2p orbitals.**

Ans:

2s
Shape of 2s orbital



Shapes of 2p orbitals

***Q.20. State the order of filling atomic orbitals following Aufbau principle.**

Ans: Aufbau principle:

i. Aufbau principle gives the sequence in which various orbitals are filled with electrons.

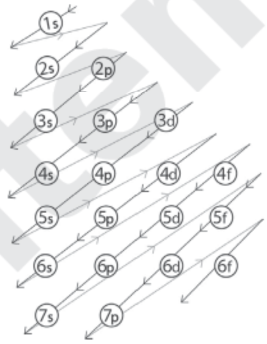
ii. In the ground state of an atom, the orbitals are filled with electrons based on increasing order of energies of orbitals, Pauli's exclusion principle and Hund's rule of maximum multiplicity.

iii. Increasing order of energies of orbitals:

a. Orbitals are filled in order of increasing value of $(n + l)$

b. In cases where the two orbitals have same value of $(n + l)$, the orbital with lower value of n is filled first.

iv. The increasing order of energy of different orbitals in a multi-electron atom is:
1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s < 4f < 5d < 6p < 7s and so on.



Increasing order of orbital energy

Note:

i. The energy of an electron in a hydrogen atom is determined solely by the principal quantum number. This is because the only interaction in these species is attraction between the electron and the nucleus. Thus, the energy of the orbitals increases as follows :
1s < 2s = 2p < 3s = 3p = 3d < 4s = 4p = 4d = 4f

ii. The orbitals having same energies are called **degenerate orbitals.**

iii. The energy of an electron in a multi-electron atom, depends on both its principal quantum number (shell) and azimuthal quantum number (sub-shell).

iv. The main reason for having different energies of the sub-shells is the mutual repulsion among the electrons in a multi-electron atom.

v. In a multi-electron atom, electrons ocupy different orbitals. The lowest total electronic energy corresponds to the most stable, that is, the **ground state of an atom**.

***Q.21. State and explain Pauli's exclusion principle.**

Ans: Pauli's exclusion principle:

i. **Statement:** *"No two electrons in an atom can have the same set of four quantum numbers".* **OR**
"Only two electrons can occupy the same orbital and they must have opposite spins."

ii. The capacity of an orbital to accommodate electrons is decided by Pauli's exclusion principle.

iii. According to this principle, for an electron belonging to the same orbital, the spin quantum number must be different since the other three quantum numbers are the same.

iv. The spin quantum number can have two values: $+\frac{1}{2}$ and $-\frac{1}{2}$.

v. Example, consider helium (He) atom with electronic configuration $1s^2$.
For the two electrons in 1s orbital, the four quantum numbers are as follows:

Electron number	Quantum number				Set of values of quantum numbers
	n	l	m	s	
1^{st} Electron	1	0	0	$+\frac{1}{2}$	$(1, 0, 0, +\frac{1}{2})$
2^{nd} Electron	1	0	0	$-\frac{1}{2}$	$(1, 0, 0, -\frac{1}{2})$

Sample Content

Thus, in an atom, any two electrons can have the same three quantum numbers, but the fourth quantum number must be different.

vi. This leads to the conclusion that an orbital can accommodate maximum of two electrons and if it has two electrons, they must have opposite spin.

***Q.22. State Hund's rule of maximum multiplicity with suitable example.**

Ans: Hund's rule of maximum multiplicity:

i. **Statement:** *"Pairing of electrons in the orbitals belonging to the same subshell does not occur unless each orbital belonging to that subshell has got one electron each."*

ii. Example, according to Hund's rule, each of the three-degenerate p-orbitals must get one electron of parallel spin before any one of them receives the second electron of opposite spin. Therefore, the configuration of four electrons occupying p-orbitals is represented as

↑↓	↑	↑	and not as	↑↓	↑↓	

iii. As a result of Hund's rule, the atom with fully filled and half-filled set of degenerate orbitals has extra stability.

***Q.23. Make the pairs:**

	'A'		'B'
i.	Neutrons	a.	Six electrons
ii.	p-orbitals	b.	-1.6×10^{-19} C
iii.	Charge on electron	c.	Ultraviolet region
iv.	Lyman series	d.	Chadwick

Ans: i – d, ii – a, iii – b, iv – c

***Q.24. Define electronic configuration.**

Ans: ***Electronic configuration*** *of an atom is defined as the distribution of its electrons in orbitals.*

***Q.25. An atom of an element contains 29 electrons and 35 neutrons. Deduce:**

i. The number of protons

ii. The electronic configuration of that element

Ans:

i. In an atom, number of protons is equal to number of electrons.
The given atom contains 29 electrons.

∴ Number of protons = 29

ii. The electronic configuration of an atom of an element containing 29 electrons is:
$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^{10}$

Note: Given element is copper (Cu) with Z = 29.

***Q.26. Explain the anomalous behaviour of copper and chromium.**

Ans:

i. Copper:

a. Copper (Cu) has atomic number 29.
b. Its expected electronic configuration is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^9$.
c. The 3d orbital is neither half-filled nor fully filled. Hence, it has less stability.
d. Due to interelectronic repulsion forces, one 4s electron enters into 3d orbital. This makes 3d orbital completely filled and 4s orbital half-filled which gives extra stability and the electronic configuration of Cu becomes, $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^{10}$.

ii. Chromium:

a. Chromium (Cr) has atomic number 24.
b. Its expected electronic configuration is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^4$.
c. The 3d orbital is less stable as it is not half-filled.
d. Due to inter electronic repulsion forces, one 4s electron enters into 3d orbital. This makes 4s and 3d orbitals half-filled which gives extra stability and the electronic configuration of Cr becomes,
$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^5$.

Sample Content

***Q.27. The electronic configuration of oxygen is written as $1s^2\ 2s^2\ 2p_x^2\ 2p_y^1\ 2p_z^1$ and not as $1s^2\ 2s^2\ 2p_x^2\ 2p_y^2\ 2p_z^0$. Explain.**

Ans:

i. According to Hund's rule of maximum multiplicity "*Pairing of electrons in the orbitals belonging to the same subshell does not occur unless each orbital belonging to that subshell has got one electron each.*"

ii. Oxygen has 8 electrons. The first two electrons will pair up in the 1s orbital, the next two electrons will pair up in the 2s orbital and this leaves 4 electrons, which must be placed in the 2p orbitals.

iii. Each of the three degenerate p-orbitals must get one electron of parallel spin before any one of them receives the second electron of opposite spin. Therefore, two p orbitals have one electron each and one p-orbital will have two electrons.

Thus, the electronic configuration of oxygen is written as $1s^2\ 2s^2\ 2p_x^2\ 2p_y^1\ 2p_z^1$ and not as $1s^2\ 2s^2\ 2p_x^2\ 2p_y^2\ 2p_z^0$.

***Q.28. Write condensed orbital notation of electronic configuration of the following elements:**

i. Lithium (Z = 3) **ii. Carbon (Z = 6)**
iii. Oxygen (Z = 8) **iv. Silicon (Z = 14)**
v. Chlorine (Z = 17) **vi. Calcium (Z = 20)**

Ans:

No.	Element	Condensed orbital notation
i.	Lithium (Z = 3)	[He] $2s^1$
ii.	Carbon (Z = 6)	[He] $2s^2\ 2p^2$
iii.	Oxygen (Z = 8)	[He] $2s^2\ 2p^4$
iv.	Silicon (Z = 14)	[Ne] $3s^2\ 3p^2$
v.	Chlorine (Z = 17)	[Ne] $3s^2\ 3p^5$
vi.	Calcium (Z = 20)	[Ar] $4s^2$

***Q.29. Write electronic configurations of Fe, Fe^{2+}, Fe^{3+}.**

Ans:

Species	Orbital notation
Fe	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^6$ OR [Ar] $4s^2\ 3d^6$
Fe^{2+}	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^6$ OR [Ar] $3d^6$
Fe^{3+}	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5$ OR [Ar] $3d^5$

***Q.30. Indicate the number of unpaired electrons in:**

i. Si (Z = 14) **ii. Cr (Z = 24)**

Ans:

i. **Si (Z = 14):** $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2$

Orbital diagram:

1s	2s	2p			3s	3p		
↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑	

Number of unpaired electrons = 2

ii. **Cr (Z = 24):** $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^5$

Orbital diagram:

1s	2s	2p			3s	3p			4s	3d				
↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑	↑	↑	↑	↑

Number of unpaired electrons = 6

***Q.31. Define isoelectronic species.**

Ans: ***Isoelectronic species*** *are defined as atoms and ions having the same number of electrons.*

e.g. Ar, Ca^{2+} and K^+ containing 18 electrons each.

Sample Content

***Q.32. Identify from the following the isoelectronic species:**
Ne, O^{2-}, Na^{+} OR Ar, Cl^{2-}, K^{+}

Ans: Atoms and ions having the same number of electrons are isoelectronic.

Species	No. of electrons
Ne	10
O^{2-}	$8 + 2 = 10$
Na^{+}	$11 - 1 = 10$
Ar	18
Cl^{2-}	$17 + 2 = 19$
K^{+}	$19 - 1 = 18$

Hence, Ne, O^{2-}, Na^{+} are isoelectronic species.

Reading between the lines

For atoms: No. of electrons = Atomic number (Z)
For monoatomic cation: No. of electrons = Z – [charge on cation]
For monoatomic anion: No. of electrons = Z + [charge on anion]

Multiple Choice Questions

*1. The energy difference between the shells goes on _______ when moved away from the nucleus.
(A) increasing (B) decreasing
(C) equalizing (D) static

*2. The value of Planck's constant is _______.
(A) 6.626×10^{-34} J s (B) 6.023×10^{-24} J s
(C) 1.667×10^{-28} J s (D) 6.626×10^{-28} J s

*3. Principal quantum number describes _______.
(A) shape of orbital
(B) size of the orbital
(C) spin of electron
(D) orientation of the orbital electron cloud

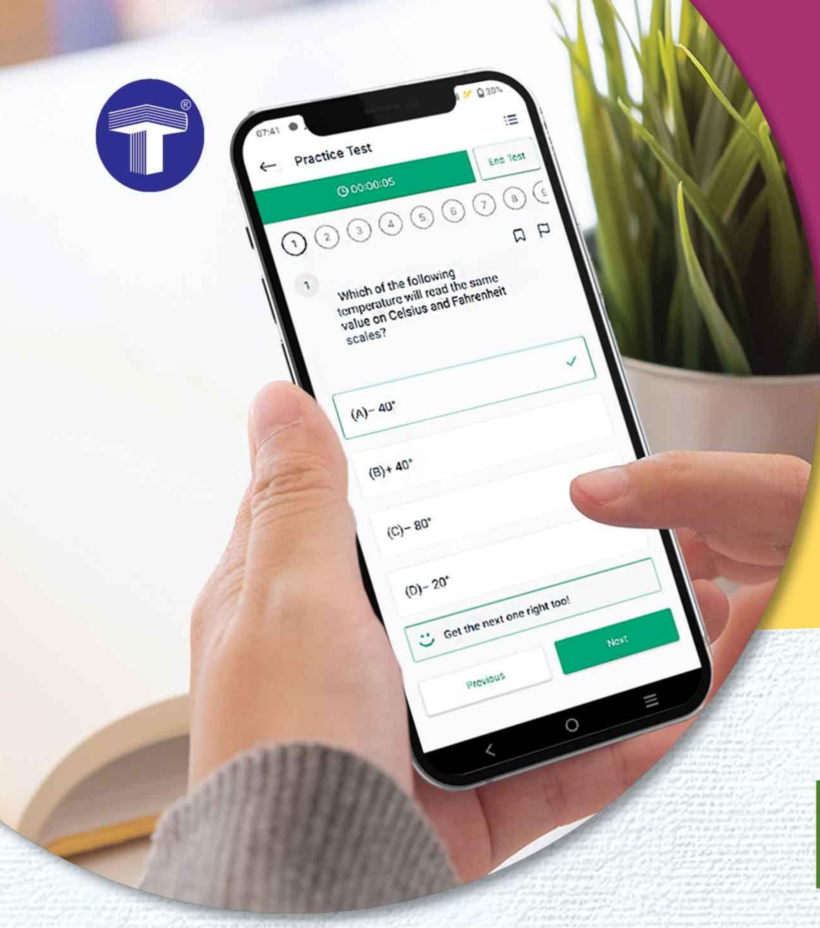
*4. p-Orbitals are _______ in shape.
(A) spherical (B) dumbbell
(C) double dumbbell (D) diagonal

*5. "No two electrons in the same atom can have identical set of four quantum numbers". This statement is known as _______.
(A) Pauli's exclusion principle
(B) Hund's rule
(C) Aufbau rule
(D) Heisenberg uncertainty principle

Answers to Multiple Choice Questions

1. (B) 2. (A) 3. (B) 4. (B)
5. (A)

Sample Content



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