

CHEMISTRY

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PERIODIC TABLE OF THE ELEMENTS

The periodic table shows elements arranged by atomic number. It includes color-coded regions: s-block (pink), p-block (light blue), d-block (light green), and f-block (light orange). Sub-periodic tables show the orbital configurations for each block: s-block (1 orbital), p-block (3 orbitals), d-block (5 orbitals), and f-block (7 orbitals).



ATOMIC STRUCTURE

Atomic Number, Z , is of protons in the nucleus. For neutral atom, $Z = +$ of electrons. For ions of charge, $+Q$, $Z = +Q$ (cations) or $Z = -Q$ (anions). Atomic mass, A , is the sum of protons and neutrons in the nucleus. Isotopes atoms with the same Z , different A . The atomic mass is a weighted average of isotopes. Use atomic mass, A , to determine a number of isotopes. In practice, work with the atomic mass. Number of Atomic Weight is a weighted average of isotopes for a given element.

ATOMIC QUANTUM NUMBERS AND ORBITALS

Quantum Numbers:
 1. Principal quantum number (n): $n = 1, 2, 3, \dots$
 2. Azimuthal quantum number (l): $l = 0, 1, 2, \dots, n-1$
 3. Magnetic quantum number (m_l): $m_l = -l, -(l-1), \dots, 0, \dots, (l-1), l$
 4. Spin quantum number (m_s): $m_s = +\frac{1}{2}, -\frac{1}{2}$

Orbitals:
 s-orbital: $l = 0$, 1 type
 p-orbital: $l = 1$, 3 types
 d-orbital: $l = 2$, 5 types
 f-orbital: $l = 3$, 7 types

MANY-ELECTRON ATOMS

Orbitals and quantum numbers are used to describe all many-electron atoms and ions. A many electron atom has more electrons than single electrons and orbital properties. The main point is that orbitals have different energies. The filling of levels is guided by the Aufbau Principle.

n	l	m_l	m_s	Orbital
1	0	0	$\pm\frac{1}{2}$	1s orbital
2	0	0	$\pm\frac{1}{2}$	2s orbital
2	1	0	$\pm\frac{1}{2}$	2p _z orbital
2	1	± 1	$\pm\frac{1}{2}$	2p _x orbital
2	1	± 1	$\pm\frac{1}{2}$	2p _y orbital
3	0	0	$\pm\frac{1}{2}$	3s orbital
3	1	0	$\pm\frac{1}{2}$	3p _z orbital
3	1	± 1	$\pm\frac{1}{2}$	3p _x orbital
3	1	± 1	$\pm\frac{1}{2}$	3p _y orbital
3	2	0	$\pm\frac{1}{2}$	3d _{z²} orbital
3	2	± 1	$\pm\frac{1}{2}$	3d _{xy} orbital
3	2	± 2	$\pm\frac{1}{2}$	3d _{x²-y²} orbital
3	2	± 2	$\pm\frac{1}{2}$	3d _{xz} orbital
3	2	± 2	$\pm\frac{1}{2}$	3d _{yz} orbital

Atomic Configuration: Shows orbital occupancy. Pauli Exclusion Principle: Each orbital has a unique set of quantum numbers. An orbital may hold up to two electrons, one with spin up, $m_s = +\frac{1}{2}$, and one with spin down, $m_s = -\frac{1}{2}$. Hund's Rule: Electrons fill orbitals of equal energy as a many-electron atom. The number of electrons that fill orbitals is the number of electrons. Electrons are added to orbitals in order of increasing energy. The order of filling of orbitals is given by the Aufbau Principle. Hund's Rule: Electrons fill orbitals of equal energy as a many-electron atom. The number of electrons that fill orbitals is the number of electrons. Electrons are added to orbitals in order of increasing energy. The order of filling of orbitals is given by the Aufbau Principle.

NUCLEAR CHEMISTRY

Nuclear reactions alter the number, particles present or chemical behavior of atoms and ions. Energy (E) is related to mass (m) by the Einstein equation: $E = mc^2$. Nuclear reaction: ${}^A_Z X \rightarrow {}^A_Z Y + \dots$. Types of processes: Radioactive decay, Nuclear fission, Nuclear fusion.



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