

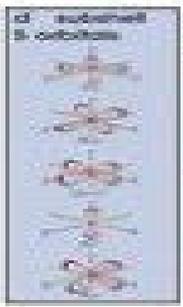
CHEMISTRY

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PERIODIC TABLE OF THE ELEMENTS

Diagram showing the periodic table of elements, color-coded by groups. Hydrogen (H) is highlighted with a callout box:

- 1** - Atomic Number
- H** - Atomic Symbol
- 1.008** - Atomic Weight



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 Number, A , is $Z + N$, the number of nucleons in the nucleus. Isotopes atoms with the same Z , different N , the changed sign have a number of protons, say Z , and a different number of neutrons, as practical work use the average Atomic Weight or Atomic Weight (weighted average of isotopes) instead for a given element.

NUCLEAR CHEMISTRY

Nuclear reactions alter the nucleus, particles emitted or absorbed. Release energy and mass energy. Energy (E) is related to mass (m) using the Einstein's equation, $E = mc^2$.
Nucleus is built from protons, p , neutrons, n , symbolized by ${}^1_1\text{H}$ and ${}^1_0\text{n}$ (eg. Deuterium, ${}^2_1\text{H}$ or ${}^2\text{D}$).
Types of processes:
- Radioactivity: Two α -particles made from particle emissions.
- Fusion: Small nuclei combine to form a larger nucleus.
- Fission: Large nuclei split.

Radioactive decay:
- α decay: $\Delta A = 4, \Delta Z = 2$
- β decay: $\Delta A = 0, \Delta Z = \pm 1$
- γ decay: $\Delta A = 0, \Delta Z = 0$
- Positron emission: $\Delta A = 0, \Delta Z = -1$
- Electron capture: $\Delta A = 0, \Delta Z = -1$
- Spontaneous fission: $\Delta A = 0, \Delta Z = 0$



ATOMIC QUANTUM NUMBERS AND ORBITALS

Quantum numbers:
- n : principal quantum number (shell)
- l : azimuthal quantum number (subshell)
- m_l : magnetic quantum number (orbital)
- m_s : spin quantum number (electron spin)
- $l = n - 1, n - 2, \dots, 0$
- $m_l = l, l - 1, \dots, 0, \dots, -(l - 1), -l$
- $m_s = +\frac{1}{2}, -\frac{1}{2}$
- $n = 1, 2, 3, \dots$
- $l = 0, 1, 2, \dots, n - 1$
- $m_l = 0, \pm 1, \pm 2, \dots, \pm l$
- $m_s = \pm \frac{1}{2}$

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 atoms 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

Atomic configuration: Shows orbital occupancy. Pauli Exclusion Principle: Each orbital has a unique set of quantum numbers, no orbital may hold up to two electrons, one with spin up, $m_s = +\frac{1}{2}$, and spin down, $m_s = -\frac{1}{2}$.
Hund's Rule: Electrons fill orbitals of equal energy, one electron per orbital, maximize unpaired electrons. Pauli Exclusion Principle: No two electrons can have the same set of quantum numbers. Electrons fill orbitals in order of increasing energy. Hund's Rule: Orbitals of equal energy fill singly, with parallel spins, before pairing up. Aufbau Principle: Orbitals are filled in order of increasing energy.



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generation often lived together in the same house or at least close to each other

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