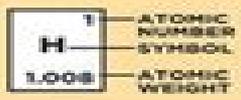


CHEMISTRY



Periodic Table of Elements

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Na 22.99	Mg 24.31	Al 26.98	Si 28.09	P 30.97	S 32.07	Cl 35.45	Ar 39.95	K 39.10	Ca 40.08	Sc 44.96	Ti 47.88	V 50.94	Cr 52.00	Mn 54.94	Fe 55.85	Co 58.93	Ni 58.71	Cu 63.55	Zn 65.38	Ga 69.72	Ge 72.64	As 74.92	Se 78.96	Br 79.90	Kr 83.80																																																																																									
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- Elemental forms:**
 - Monatomic:** Noble gases; helium, neon, argon, krypton, xenon, and radon; nonreactive
 - Diatoms:** Gases: fluorine, chlorine, hydrogen, oxygen, and nitrogen; Liquid: bromine; Solid: iodine
 - Molecular solids:** Sulfur
 - Covalent solids:** Carbon (diamond, graphite)
 - Metals:** Most common type of element; high coordination numbers; very dense
- Types of compounds:** Large variety due to ionic, covalent, and metallic bonding options
 - Ionic:** Electron exchange
 - Covalent:** Shared electrons
 - Polar covalent: Unequal sharing
 - Ion complex:** Central ion + solvent
 - Transition metal compounds:**
 - Coordination compounds: Non-bonding ligands
 - Organometallics: Metal-C bond
 - Complex, mineral:** Metal oxides and silicates
- Periodicity:** The chemical properties of an element depend on valence electrons
 - Family of elements:** Column in periodic table
 - The addition of d-orbitals produces an expanded octet

- Ionization energies:** Noble gases are the most stable; alkali metals are the easiest to ionize; electron screening lessens Z-effect on orbital energy
- Atomic and ionic radii:** Not real properties; derive from structural data; divide bond into radii using a consistent scheme
 - Ionic radii:** Depends on charge, counter ion, and coordination (crystal type)
 - Atomic radii:** Based on diatomic gas data or metallic solid structural data

Development of the Periodic Table

- Missing elements were predicted using periodicity
- Column 1:** Alkali metals
- Column 2:** Alkaline earth metals
- First transition series:** Titanium-Copper
- Column 17:** Halogens: fluorine, chlorine, bromine, iodine, and astatine
- Column 18:** Rare gases: helium, neon, argon, krypton, xenon, and radon
- Astroids:** Z = 90 or above; radioactive
- Lanthanides:** f-subshell elements; starts with Z = 58



Practical Application
The colors of fireworks depend on the elements in the pyrotechnic.



Atomic Structure

- Atomic number (Z):** The number of protons in the nucleus; for a neutral atom, Z = number of electrons; for an anion of charge -q, Z = q electrons; for a cation of charge +q, Z = q electrons
- Mass number (A):** A = Z + N, where N = the number of neutrons in the nucleus
- Isotopes:** Atoms with the same Z, different A; an element can have one or more isotopes; any sample contains a number of isotopes; in practical work, use the average mass number or atomic weight (weighted average of natural isotopes for a given element)



Atomic Quantum Numbers & Orbitals

- n:** principle quantum number (orbital distance)
- l:** magnetic (orbital direction)
- m_l:** electron spin
- m_s:** spin quantum number (orbital shape)
- For each n, the possible l values are 0, 1, ..., n - 1
- For each l, the possible m_l values are -l, ..., 0, ..., +l
- m_s has two possible values: +1/2 and -1/2 (spin up, spin down)
- Each value of n denotes a "shell" in the atomic structure; each l denotes a subshell
 - 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 derived for the hydrogen atom are used to describe all many-electron atoms and ions; a more rigorous treatment determines exact energy levels and orbital properties; the main result is that subshells have different energies; the filling of levels is guided by the Aufbau principle:



n	l	m _l	m _s	element
1	0	0	+1/2	H
1	0	0	-1/2	He
2	0	0	+1/2	Li
2	0	0	-1/2	Be
2	1	-1	+1/2	B
2	1	0	+1/2	C
2	1	1	+1/2	N
2	1	-1	-1/2	O
2	1	0	-1/2	F
2	1	1	-1/2	Ne

- Electron configuration:** Atomic orbital occupancy for an atom
- Pauli exclusion principle:** Each electron has a unique set of quantum numbers; an orbital may hold up to two electrons; one with spin up, m_s = +1/2, and one with spin down, m_s = -1/2
- Hund's rule:** Electrons fill p, d, and f subshells in a manner to maximize the number of unpaired electrons (half-filled orbitals); spin is maximized
- Ionization potential (IP):** Energy required to remove electron from an atom or ion
 - First IP:** Removal of first valence electron
- Electronegativity:** Tendency of an atom to attract electrons in a chemical bond; 0-4 scale; 0 for rare gas

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