

CHEMISTRY



Periodic Table of Elements

1 H 1.008																	18 Ar 39.948																																																																																																				
2 He 4.0026																	36 Kr 83.80																																																																																																				
3 Li 6.941	4 Be 9.0122															54 Xe 131.29	86 Rn 222																																																																																																				
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1	ATOMIC NUMBER
H	SYMBOL
1.008	ATOMIC WEIGHT

- Elemental forms:**
 - Monoatomic:** Noble gases: helium, neon, argon, krypton, xenon, and radon; nonreactive
 - Diatomic:** 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 minerals:** 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

Periodic Trends

- Chemistry determined by the behavior of valence electrons
- Atoms or ions get larger as you go down the column; electron shells are added
- Cation and anions (same charge) decrease in size as the atomic number (Z) increases

- 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 radius:** Depends on charge, counter ion, and coordination (crystal type)
 - Atomic radius:** Based on diatomic gas data or metallic solid structural data

Development of the Periodic Table

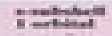
- Missing elements were predicted using periodicity
- Column 1:** Alkali metals
- Column 2:** Alkaline earth metals
- First transition series:** Titanium-Copper
- Column 17:** Halogens: fluorine, chlorine, bromine, iodine, and astatine
- Column 18:** Rare gases: helium, neon, argon, krypton, xenon, and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



Practical Application

- The colors of fireworks depend on the elements in the explosive

d-subshell
5 orbitals



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)

First Ionization Potential vs. Z



Atomic Radius vs. Z



Atomic Quantum Numbers & Orbitals

- n : principle**
 - n : magnetic (orbital direction)
 - n : electron spin
- For each n , the possible l values are $0, 1, \dots, n-1$
- For each l , the possible m_l values are $-l, \dots, 0, \dots, +l$
- m_s has two possible values: $+\frac{1}{2}$ and $-\frac{1}{2}$ (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
 - d-orbital: $l = 2$, 5 types
 - p-orbital: $l = 1$, 3 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:

Aufbau Order for Filling Subshells



n	l	m_l	m_s	element
1	0	0	$+\frac{1}{2}$	H
1	0	0	$-\frac{1}{2}$	He
2	0	0	$+\frac{1}{2}$	Li
2	0	0	$-\frac{1}{2}$	Be
2	1	-1	$+\frac{1}{2}$	B
2	1	0	$+\frac{1}{2}$	C
2	1	1	$+\frac{1}{2}$	N
2	1	-1	$-\frac{1}{2}$	O
2	1	0	$-\frac{1}{2}$	F
2	1	1	$-\frac{1}{2}$	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, $m_s = +\frac{1}{2}$, and one with spin down, $m_s = -\frac{1}{2}$
- 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

College Chemistry Study Guide

William Harrison Nebergall



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