

CHEMISTRY



Periodic Table of Elements

1																	2				
H 1.008 ATOMIC WEIGHT																	He 4.002 ATOMIC WEIGHT				
3	4															5	6	7	8	9	10
Li 6.941	Be 9.012															B 10.81	C 12.01	N 14.01	O 16.00	F 18.998	Ne 20.18
11	12															13	14	15	16	17	18
Na 22.99	Mg 24.31															Al 26.98	Si 28.09	P 30.97	S 32.06	Cl 35.45	Ar 39.95
19	20	21	22	23	24	25	26	27	28	29	30	31	32	33	34	35	36				
K 39.10	Ca 40.08	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr				
37	38	39	40	41	42	43	44	45	46	47	48	49	50	51	52	53	54				
Rb 85.47	Sr 87.62	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe				
55	56	57	58	59	60	61	62	63	64	65	66	67	68	69	70	71	72				
Cs 132.91	Ba 137.33	La	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu					
87	88	89	90	91	92	93	94	95	96	97	98	99	100	101	102	103	104				
Fr 223.02	Ra 226.02	Ac	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr					
119	120	121	122	123	124	125	126	127	128	129	130	131	132	133	134	135	136				
Uue	Uub	Uut	Uuq	Uur	Uus	Uuh	UuJ	Uuk	Uul	UuM	Uun	Uuo	UuP	UuQ	UuR	UuS	UuT				
119	120	121	122	123	124	125	126	127	128	129	130	131	132	133	134	135	136				
Uue	Uub	Uut	Uuq	Uur	Uus	Uuh	UuJ	Uuk	Uul	UuM	Uun	Uuo	UuP	UuQ	UuR	UuS	UuT				
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Uue	Uub	Uut	Uuq	Uur	Uus	Uuh	UuJ	Uuk	Uul	UuM	Uun	Uuo	UuP	UuQ	UuR	UuS	UuT				

1	ATOMIC NUMBER
H	SYMBOL
1.008	ATOMIC WEIGHT

- Elemental forms:**
 - Monoatomic:** 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 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

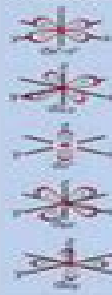
- Missing elements were predicted using periodicity
- Columns 1: Alkali metals
- Columns 2: Alkaline earth metals
- First transition series: Titanium-Copper
- Column 17: Halogens: fluorine, chlorine, bromine, iodine, and astatine
- Column 18: Rare gases: helium, neon, argon, krypton, xenon, and radon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



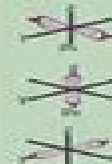
Practical Application

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

d-subshell
5 orbitals



p-subshell
3 orbitals



s-subshell
1 orbital



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**
 - angular momentum (orbital shape)
 - magnetic (orbital direction)
 - 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	0	$-\frac{1}{2}$	N
2	1	1	$+\frac{1}{2}$	O
2	1	1	$-\frac{1}{2}$	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

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