Ch 3 Atomic Structure and the Periodic Table

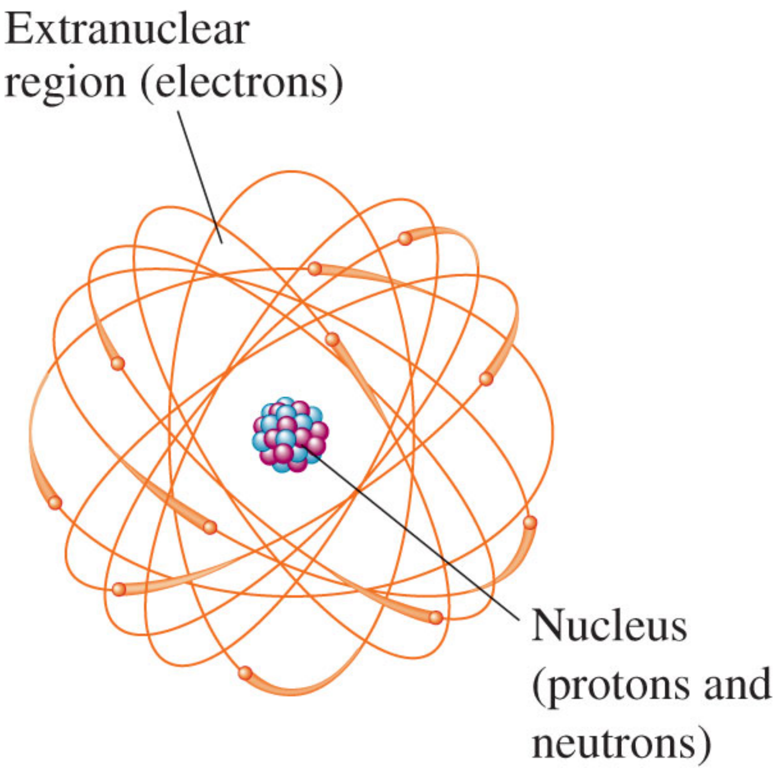
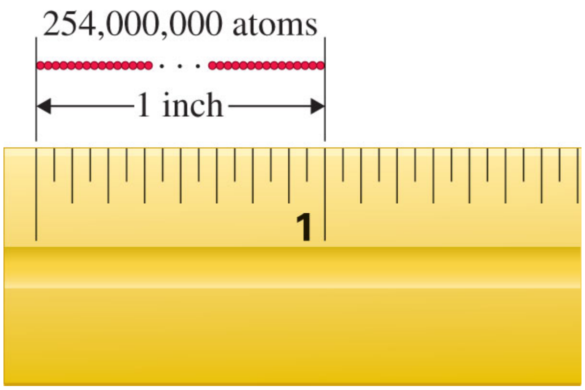


Figure 3.1 size relationship is not to scale, ratio of average diameters atom/nucleus = 10^5

Atoms are very small and spherical.

Radii Range

$$0.9 \times 10^{-10} \text{ to } 2.4 \times 10^{-10} \text{ m}$$
$$90 - 240 \text{ pm}$$



Recall Figure 1.12

Atomic Mass

example: 1 H atom = 1.673×10^{-24} g

1atomic mass unit (amu) = 1.6605×10^{-24} g

For H: $1.673 \times 10^{-24} \text{ g} \times \dfrac{1 \text{ amu}}{1.6605 \times 10^{-24} \text{ g}}$

= 1.008 amu

Ch 3.1 Internal Structure of an Atom

Subatomic particles

An atom is characterized by the number of protons (**p**) that it contains. A proton has a positive electrical charge.

p charge = 1.60×10^{-19} coulombs = +1 relative charge

p mass = 1.6726×10^{-24} g = ________ amu

A neutron (**n**) has no charge associated with it.

n mass = 1.6750×10^{-24} g = ________ amu

An electron (**e**) has a negative electrical charge.

e charge = -1.60×10^{-19} coulombs = -1 relative charge

e mass = 9.109×10^{-28} = ________ amu

mass of p ~ mass of n >>> mass of e^-

mass of e^- is often ignored

Ch 3.2 Atomic Number and Mass Number

Atomic number = Z = # of protons (unique physical property of each element)

Mass number = A = # of protons + # of neutrons

# of neutrons = ____________

For a neutral atom, net electrical charge = zero

# of electrons = # of protons = ______

$$^{A}_{Z}E$$ complete chemical symbol notation

$${}^{A}_{Z}E \qquad {}^{23}_{11}Na$$

$\# p = Z =$ ______

$\# n = A - Z =$ ____________ $=$ ______

$\# e^- = \# p =$ ______

Ch 3.3 Isotopes and Atomic Masses

Isotopes are atoms of an element that have the same number of protons and electrons but different numbers of neutrons.

$$^{12}_{6}\mathrm{C} \qquad ^{13}_{6}\mathrm{C} \qquad ^{14}_{6}\mathrm{C}$$

- isotopes have nearly identical chemical properties
- most elements have at least two isotopes

Table 3.2 Elements with Z = 1 through 12

Z	Element	Isotope	Mass	Abundance
1	Hydrogen	$^{1}_{1}H$	1.008 amu	99.985%
		$^{2}_{1}H$	2.014 amu	0.015%
		$^{3}_{1}H$	3.016 amu	trace
2	Helium	$^{3}_{2}He$	3.016 amu	trace
		$^{4}_{2}He$	4.003 amu	100%
3	Lithium	$^{6}_{3}Li$	6.015 amu	7.42%
		$^{7}_{3}Li$	7.016 amu	92.58%
4	Beryllium	$^{9}_{4}Be$	9.012 amu	100%
5	Boron	$^{10}_{5}B$	10.013 amu	19.6%
		$^{11}_{5}B$	11.009 amu	80.4%
6	Carbon	$^{12}_{6}C$	12.000 amu	98.89%
		$^{13}_{6}C$	13.003 amu	1.11%
		$^{14}_{6}C$	14.003 amu	trace
7	Nitrogen	$^{14}_{7}N$	14.003 amu	99.63%
		$^{15}_{7}N$	15.000 amu	0.37%
8	Oxygen	$^{16}_{8}O$	15.995 amu	99.759%
		$^{17}_{8}O$	16.999 amu	0.037%
		$^{18}_{8}O$	17.999 amu	0.204%
9	Fluorine	$^{19}_{9}F$	18.998 amu	100%
10	Neon	$^{20}_{10}Ne$	19.992 amu	90.92%
		$^{21}_{10}Ne$	20.994 amu	0.26%
		$^{22}_{10}Ne$	21.991 amu	8.82%
11	Sodium	$^{23}_{11}Na$	22.990 amu	100%
12	Magnesium	$^{24}_{12}Mg$	23.985 amu	78.70%
		$^{25}_{12}Mg$	24.986 amu	10.13%
		$^{26}_{12}Mg$	25.983 amu	11.17%

The atomic mass of an element is the calculated weighted average mass for the isotopes.

Chem 101 Grading Scheme

Quizzes	180 points	180/1000 = 0.18
Labs	160 points	160/1000 = 0.16
Exams	660 points	660/1000 = 0.66
	1000	1.00

Sample calculation of weighted average

Quizzes	85 %	x	0.18	=	15.3 %
Labs	95 %	x	0.16	=	15.2 %
Exams	55 %	x	0.66	=	36.3 %
Average	____ %				____ % Weighted Average

$^{35}_{17}Cl$ $^{37}_{17}Cl$

34.96885 amu 36.96590 amu

75.77 % 24.23 % natural abundance

atomic mass of Cl = weighted average mass for the isotopes

34.96885 x 75.77/100 = 25.60

36.96590 x 24.23/100 = 8.957

35.46 amu

$^{63}_{29}Cu$ $^{65}_{29}Cu$

62.9298 amu 64.9278 amu

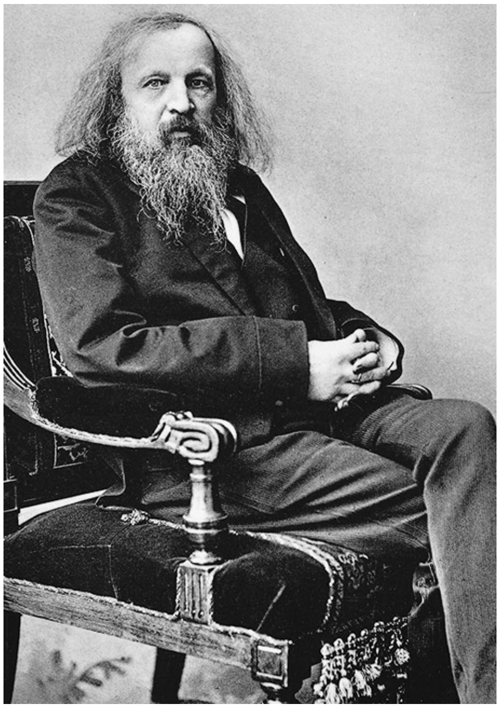
Atomic weight Cu = 63.546 amu

Which isotope is the more abundant?

Answer:

Ch. 3.4 The Periodic Law and the Periodic Table

Figure 3.2 Dmitri Ivanovich Mendeleev (1834-1907)



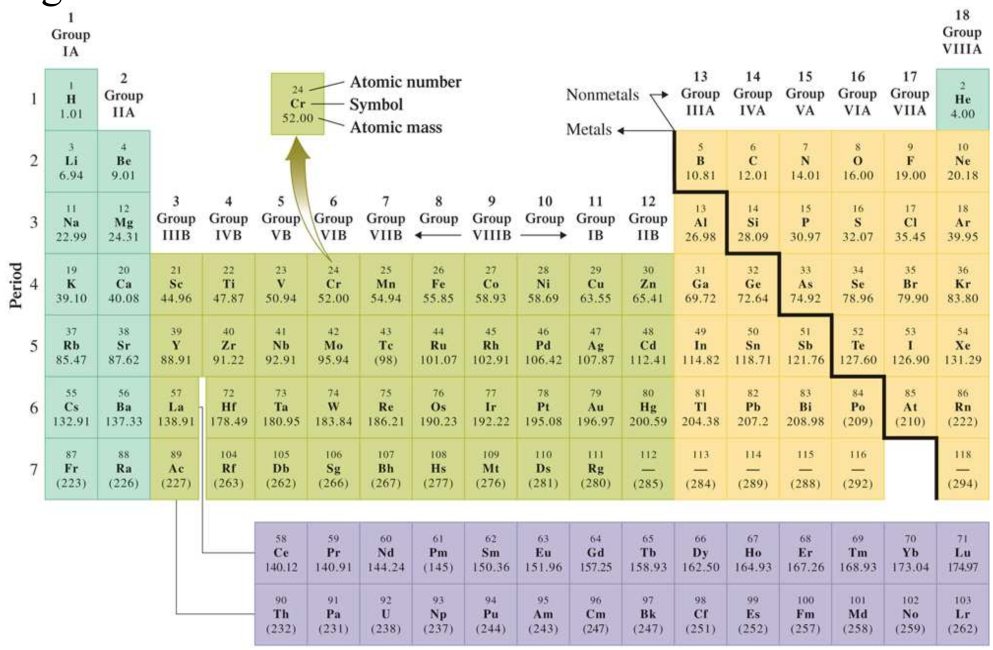
In 1869 grouped
63 known elements
in 1st periodic table
based on atomic weights and
similar properties within a
group.

Modern periodic law - when elements are arranged in order of increasing atomic number, elements with similar chemical properties occur at periodic intervals.

Figure 3.4

1																															2
3	4																									5	6	7	8	9	10
11	12																									13	14	15	16	17	18
19	20	21															22	23	24	25	26	27	28	29	30	31	32	33	34	35	36
37	38	39															40	41	42	43	44	45	46	47	48	49	50	51	52	53	54
55	56	57	58	59	60	61	62	63	64	65	66	67	68	69	70	71	72	73	74	75	76	77	78	79	80	81	82	83	84	85	86
87	88	89	90	91	92	93	94	95	96	97	98	99	100	101	102	103	104	105	106	107	108	109	110	111	112	113	114	115	116		118

Periods of Elements
Figure 3.3



Periods	# of elements	
1	2	H, He
2 + 3	8	
4 + 5	18	
6 + 7	> 18	

Groups of Elements (Ch 3.4 & 3.9)

Groups		
Groups	IA – VIIIA	representative elements
	IA	alkali metals
	IIA	alkaline earth metals
	VIIA	halogens
	VIIIA	noble gases
	IB – VIIIB	transition metals
	Outside rows	inner transition elements lanthanides & actinides

Note: The noble gases are also representative elements.

Ch. 3.5 Metals and Nonmetals

Periodic Table

1A	2A	3B	4B	5B	6B	7B	8B	8B	8B	1B	2B	3A	4A	5A	6A	7A	8A
1 H 1.008																	2 He 4.003
3 Li 6.941	4 Be 9.012											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
11 Na 23.00	12 Mg 24.31											13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.06	17 Cl 35.45	18 Ar 39.95
19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.90	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.70	29 Cu 63.55	30 Zn 65.38	31 Ga 69.72	32 Ge 72.59	33 As 74.92	34 Se 78.96	35 Br 79.90	36 Kr 83.80
37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.94	43 Tc (98)	44 Ru 101.1	45 Rh 102.9	46 Pd 106.4	47 Ag 107.9	48 Cd 112.4	49 In 114.8	50 Sn 118.7	51 Sb 121.8	52 Te 127.6	53 I 126.9	54 Xe 131.3
55 Cs 132.9	56 Ba 137.3	57 La 138.9	72 Hf 178.5	73 Ta 180.9	74 W 183.9	75 Re 186.2	76 Os 190.2	77 Ir 192.2	78 Pt 195.1	79 Au 197.0	80 Hg 200.6	81 Tl 204.4	82 Pb 207.2	83 Bi 209.0	84 Po (209)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra 226.0	89 Ac 227.0	104 Rf (261)	105 Ha (262)	106 Unh (263)	107 Uns (262)		109 Une (267)									

Lanthanides	58 Ce 140.1	59 Pr 140.9	60 Nd 144.2	61 Pm (145)	62 Sm 150.4	63 Eu 152.0	64 Gd 157.3	65 Tb 158.9	66 Dy 162.5	67 Ho 164.9	68 Er 167.3	69 Tm 168.9	70 Yb 173.0	71 Lu 175.0
Actinides	90 Th 232.0	91 Pa 231.0	92 U 238.0	93 Np 237.0	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (260)

Metals Semimetals/Metalloids Nonmetals

Semimetals/Metalloids

Used in semiconductors. They have physical properties closer to those of metals, whereas their chemical properties are closer to those of nonmetals.

(a) Metals

Fig 3.5

solids at RT, except Hg
metallic luster
malleable & ductile
high thermal & electrical
____

Al Pb Sn Zn
clockwise from left

(b) Nonmetals

Fig 3.5

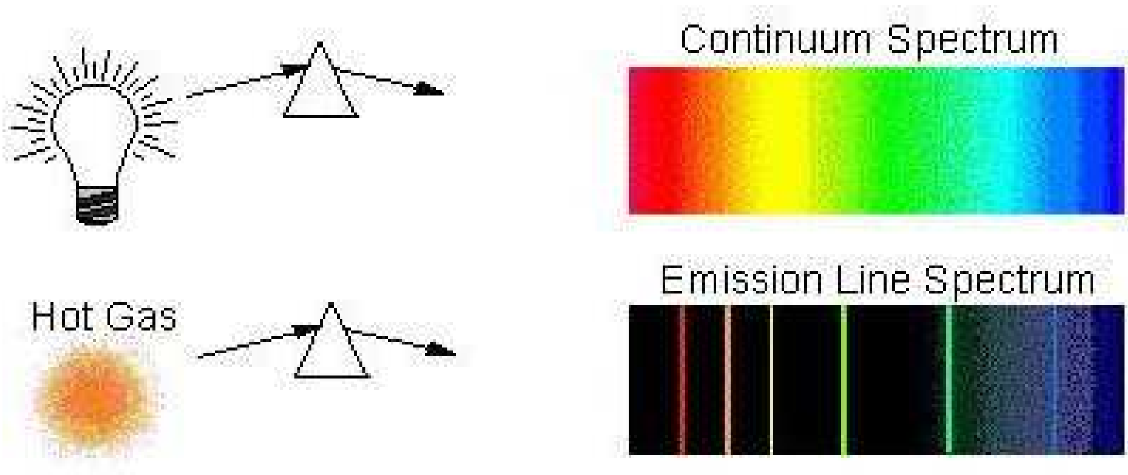
gas or solid at RT, except Br_2
variety of colors
solids are brittle
poor __________ (except graphite)
good __________ (except diamond)
nonductile

S_8 Br_2 P_4

Ch 3.6 Electron Arrangements within Atoms

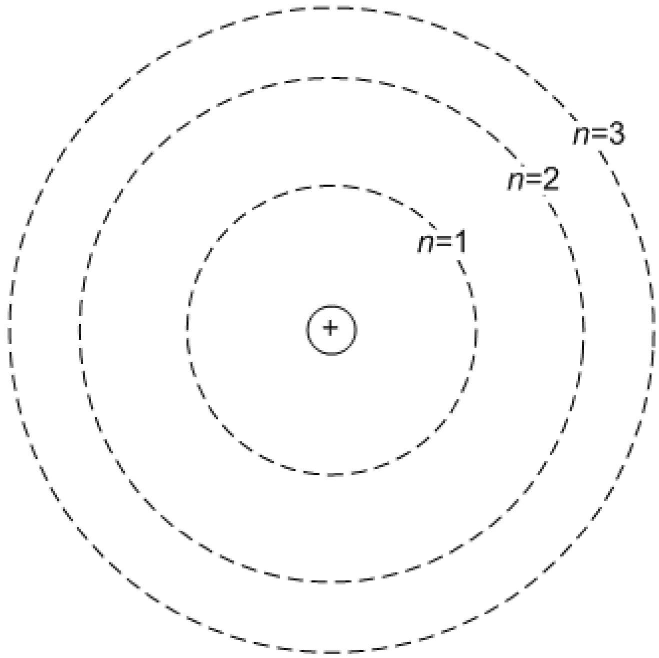
Supplemental material: Line Spectra and the Bohr Atom

Each element has a unique line spectrum

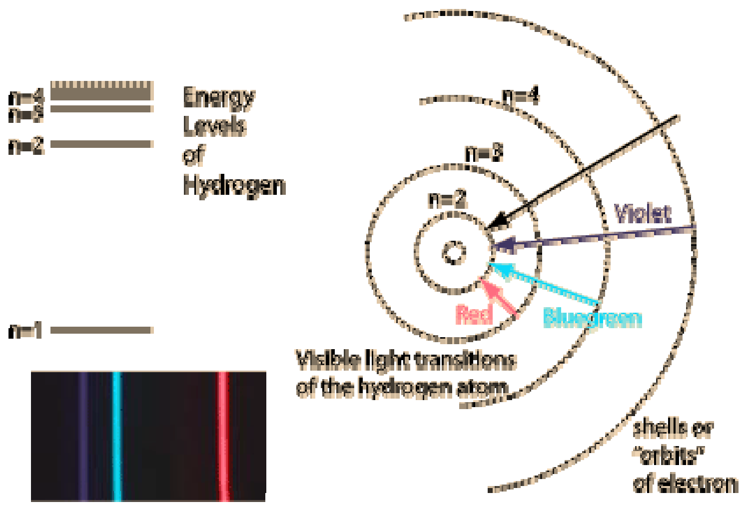


demo: noble gas discharge tubes

Bohr Model



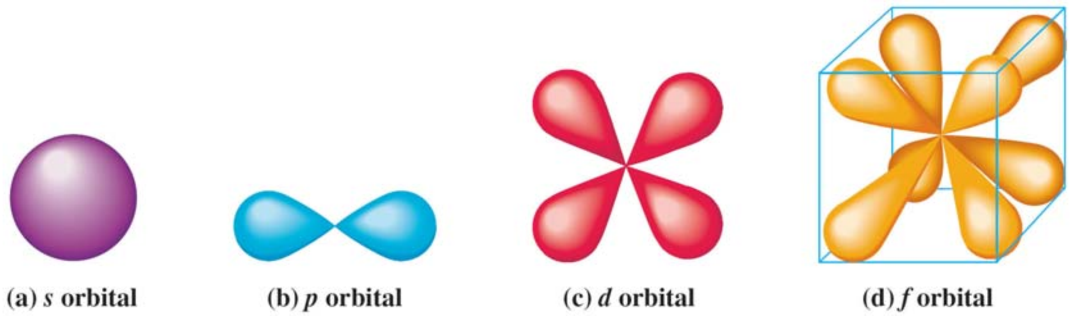
Energies of electrons are quantized = limited to certain values.



An "excited" atom releases energy in form of light when an electron "falls" back to its lower orbit (ground state).

It is now known that electrons do not exist in planet-like orbits but rather they occupy regions of space about the nucleus called orbitals.

Figure 3.8 shapes of orbitals



The space in which electrons move rapidly about a nucleus is divided into shells, subshells, and orbitals.

Electron Shells

- are specific energy levels

  n = 1, 2, 3, 4.........

  $\longrightarrow$

  increasing average distance from nucleus
  increasing average energy of shells

- electrons occupy the lowest shell available

- max # of electrons allowed in a shell = ________

n = 1	2 e^-
n = 2	8 e^-
n = 3	18 e^- etc.

Drill question: Where are the electrons of Mg located?

Z = 12 12 p = 12 e-

1^{st} shell (n = 1) = 2 e
2^{nd} shell (n = 2) = 8 e
3^{rd} shell (n = 3) = ____ e

Electron Subshells

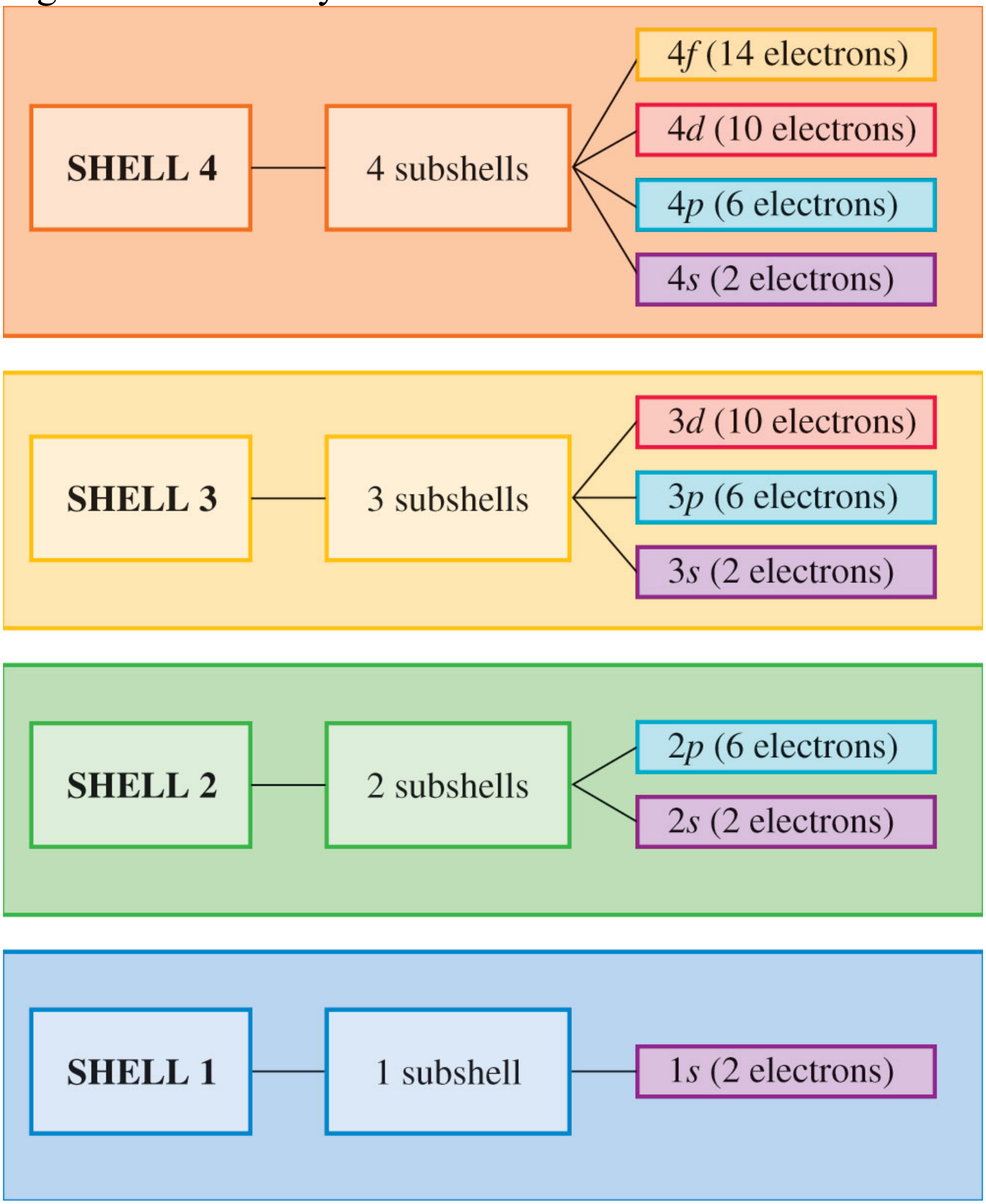
- Each shell (n) consists of subshells
- Subshells are designated s, p, d, f …..
- Energy levels increase $s < p < d < f$
- Max # of subshells = n (shell)

1st shell	n = 1	s
2nd shell	n = 2	s + p
3rd shell	n = 3	s + p + d
4th shell	n = 4	s + p + d + f

- Max # of electrons in subshells

s	= 2 e
p	= 6 e
d	= 10 e
f	= 14 e

Figure 3.7 Summary so far



Drill Problems

1. What is the max # of electrons in a 5 d subshell?

   Shell # (n) → 5 d ← subshell

   max # for d always ______ e, regardless of n

2. What is the max # of electrons in the 4^{th} shell?

   s = 2 p = 6 d = 10 f = 14   total: 32 e

   also can use: $2\,n^2 = 2\,(4)^2 = 32$ e

Electron Orbital is a region of space within a subshell where an electron with a specific energy is most likely to be found.

max of 2 e can be found in 1 orbital

subshell	# of e	# of orbitals
s	2	1
p	6	3
d	10	5
f	14	7

Characteristics of Orbitals:

- Each type has a specific shape.

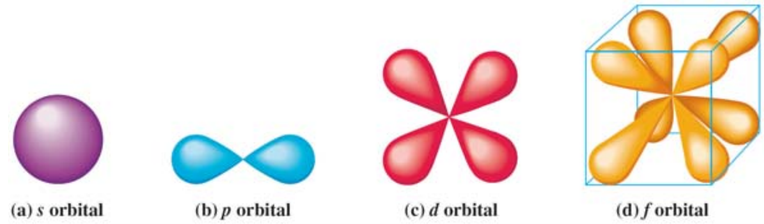


Fig 3.8

- Within the same subshell, orbitals differ mainly in orientation.

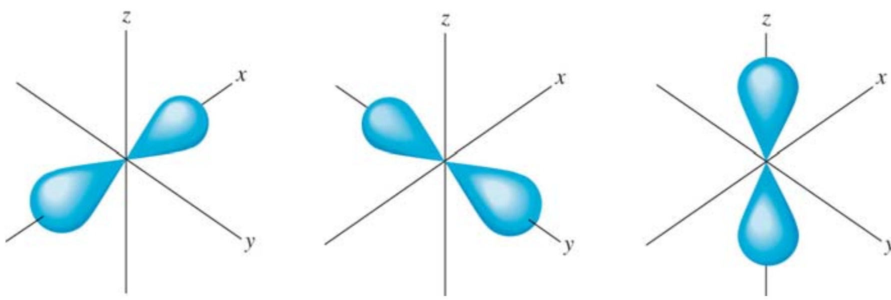


Fig 3.9

- Within a given subshell each orbital has the same _______________.
- Volume, average distance, and energy increase with increasing shell number (n).
- Electrons move rapidly and "occupy" the entire orbital volume.
- A pair of electrons in the same orbital must be of opposite _______.

Electron Spin

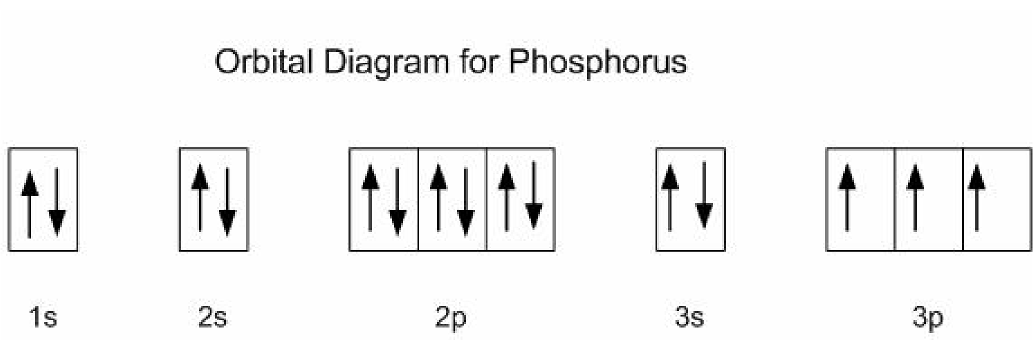
Clockwise "spin up" ↑

Counterclockwise "spin down" ↓

A pair of electrons of opposite spin can occupy the same orbital but electrons prefer to be alone – Hund's Rule

Ch 3.7 Electron Configurations and Orbital Diagrams

- fill lower energy levels first
- for a given subshell, first place single electrons
- pair electrons only with opposite spins



$1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^3$

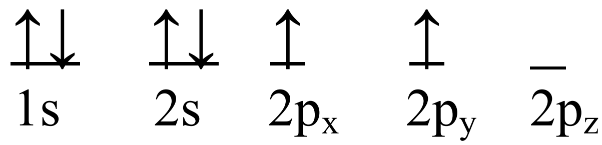
Electron Configuration for P: $1s^22s^22p^63s^23p^3$

Shorthand representation for P:

$1s^22s^22p^63s^23p^3$ = $[Ne]3s^23p^3$

Electron-Dot Symbols

${}_6C$

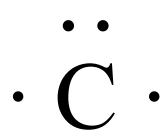


$1s^2 2s^2 2p^2$

unpaired electrons

Electrons in the s & p orbitals of the outermost shell (valence shell) are called valence electrons and can be shown in an electron-dot symbol or Lewis symbol (also see Chapter 4.2)

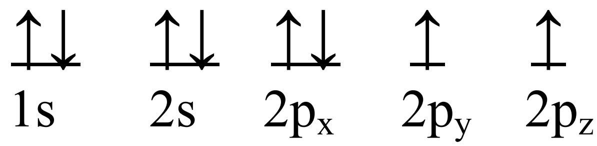
For the isolated (non-bonded) carbon atom:
4 valence electrons in the 2nd shell
2 valence e in s are paired, 2 e in p are unpaired



- all 4 sides in the symbol are equal
- unpaired electrons must be shown as single dots
- paired electrons must be shown as pairs

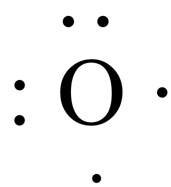
$_8O$

$1s^2 2s^2 2p^4$



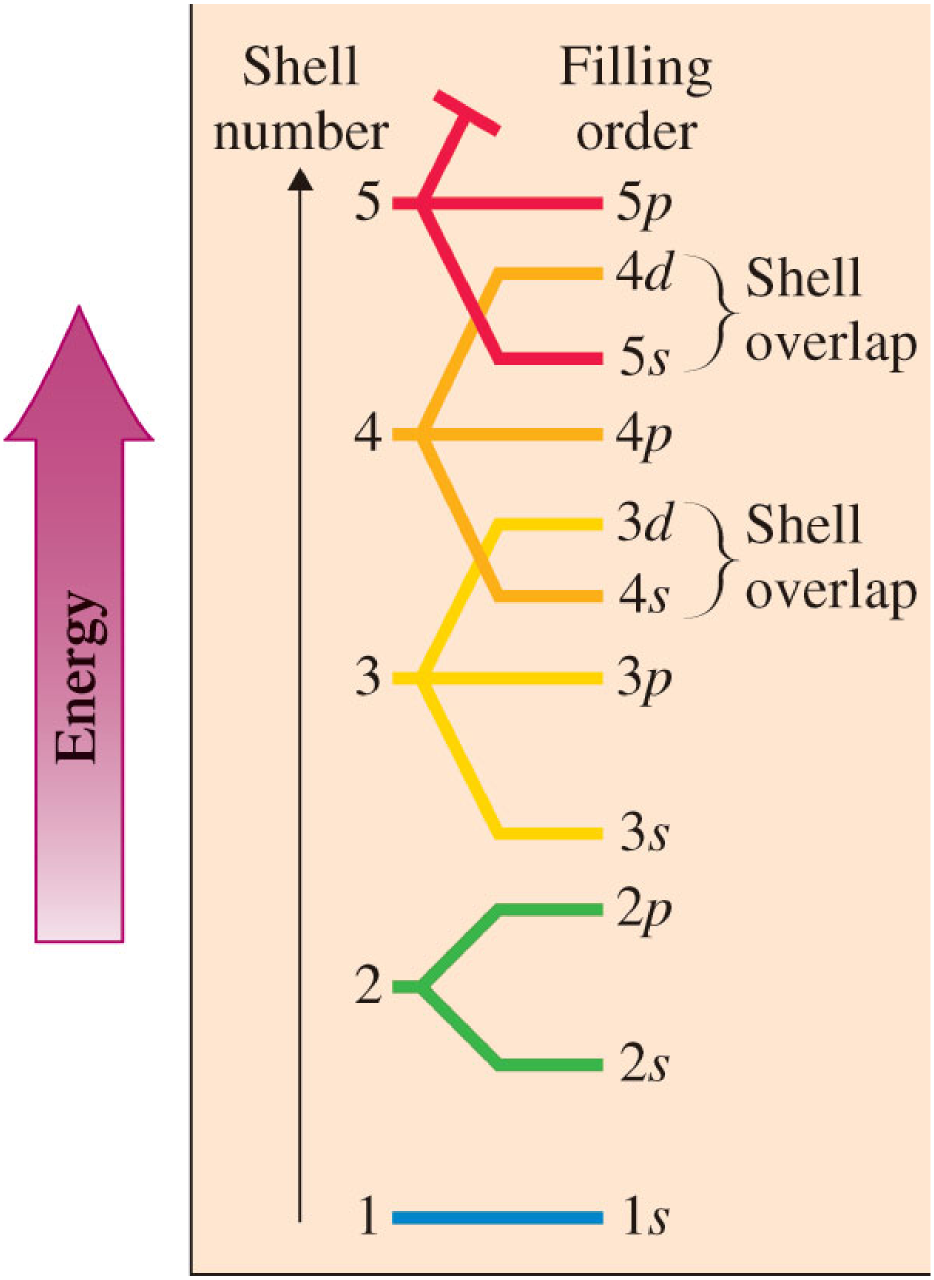
2 unpaired electrons

For the isolated (non-bonded) oxygen atom:
___ valence electrons in the 2^{nd} shell
___ valence e paired, ___ e unpaired



Note: The group number of the representative elements corresponds to the ___________ of valence electrons in the element

Fig 3.10 Order of filling electron subshells



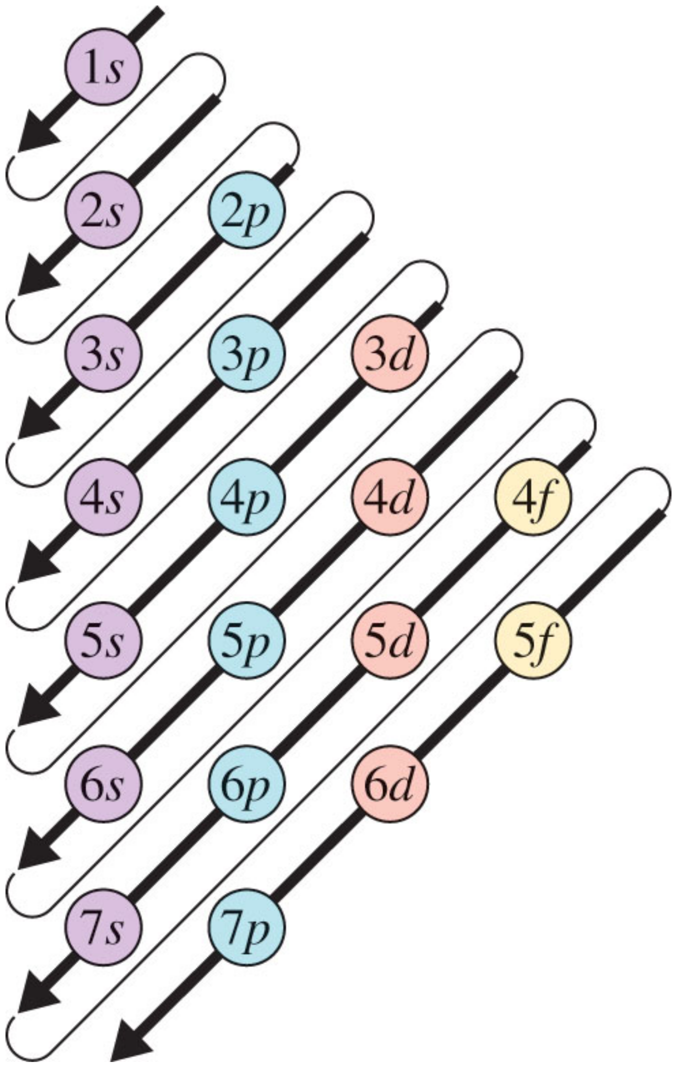


Figure 3.11
Mnemonic device for remembering subshell filling order

Ch 3.8 The Electronic Basis for the Periodic Law and the Periodic Table

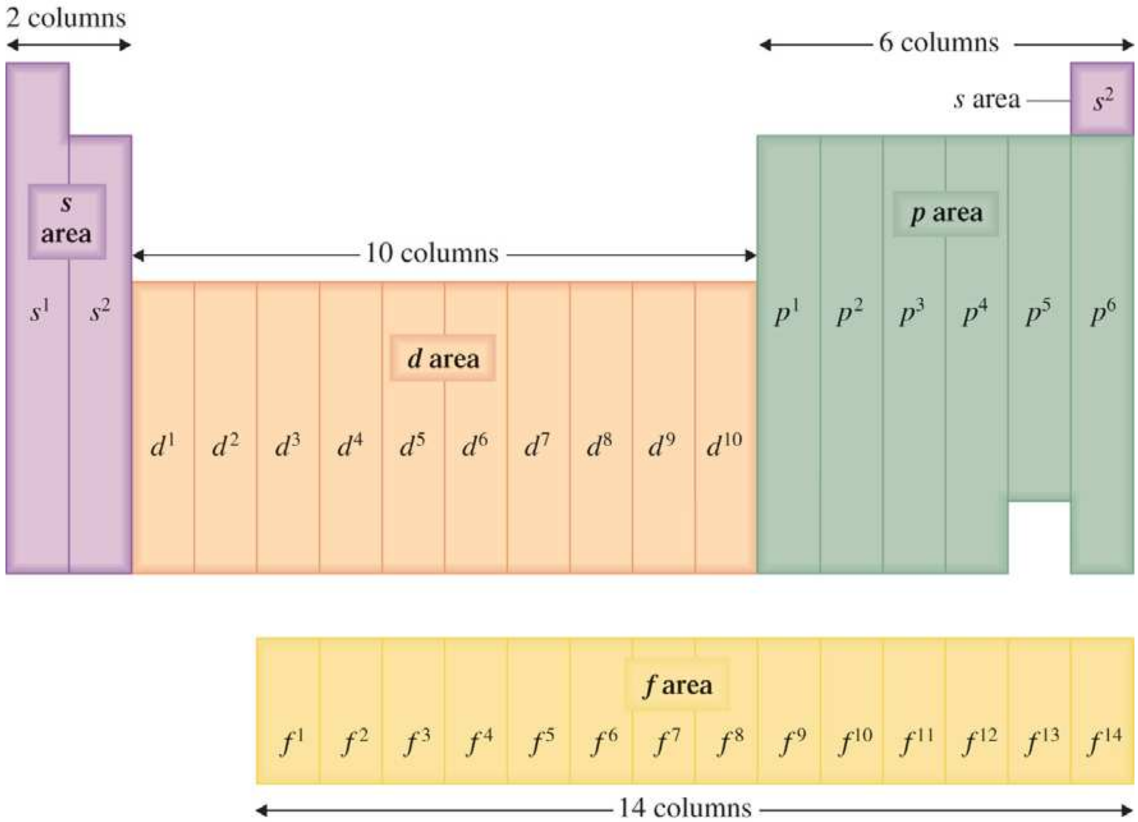
$_{3}$Li $1s^2 2s^1$
$_{11}$Na $1s^2 2s^2 2p^6 3s^1$
$_{19}$K $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$
$_{37}$Rb $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^1$

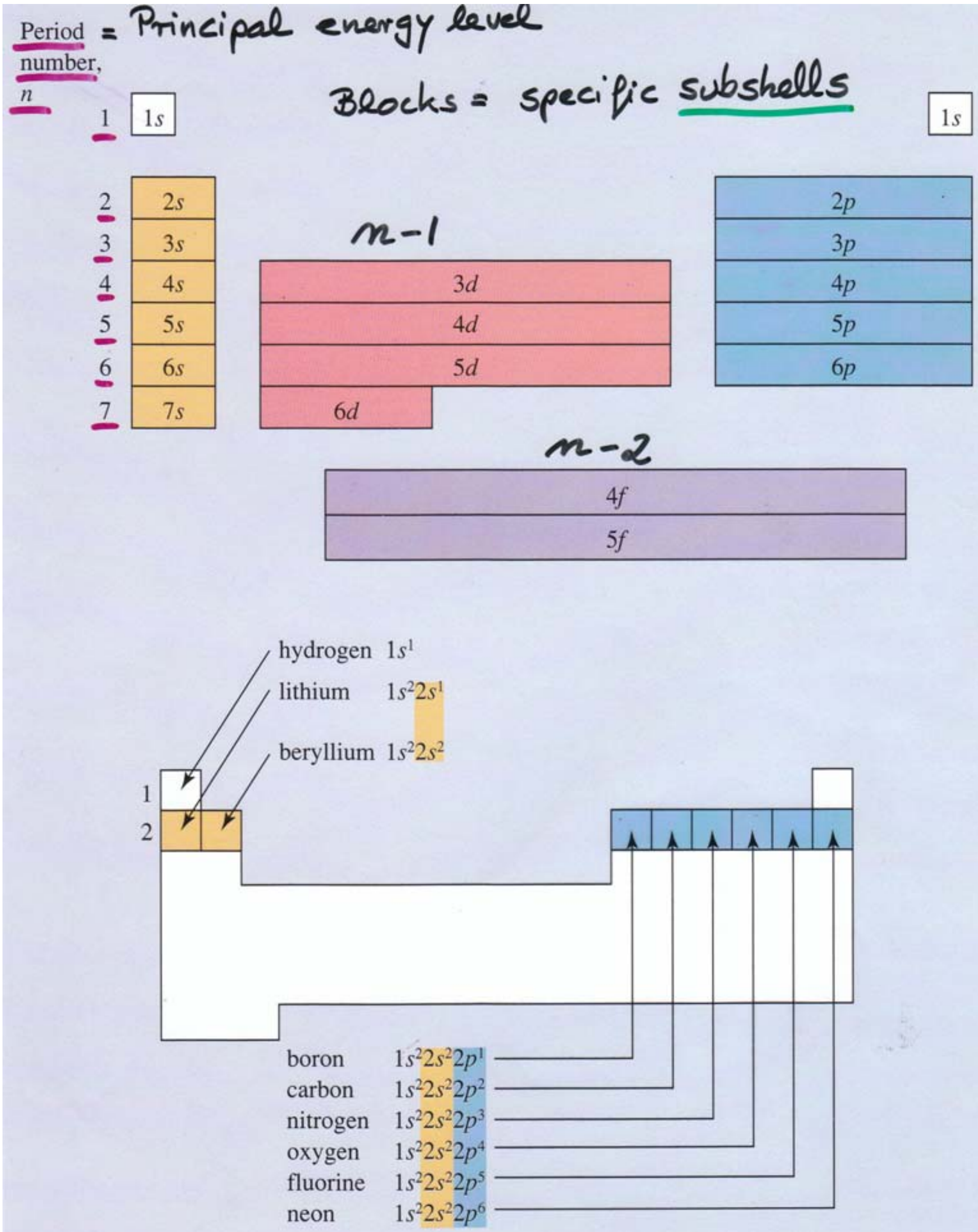
Group IA ____ valence electron in s orbital

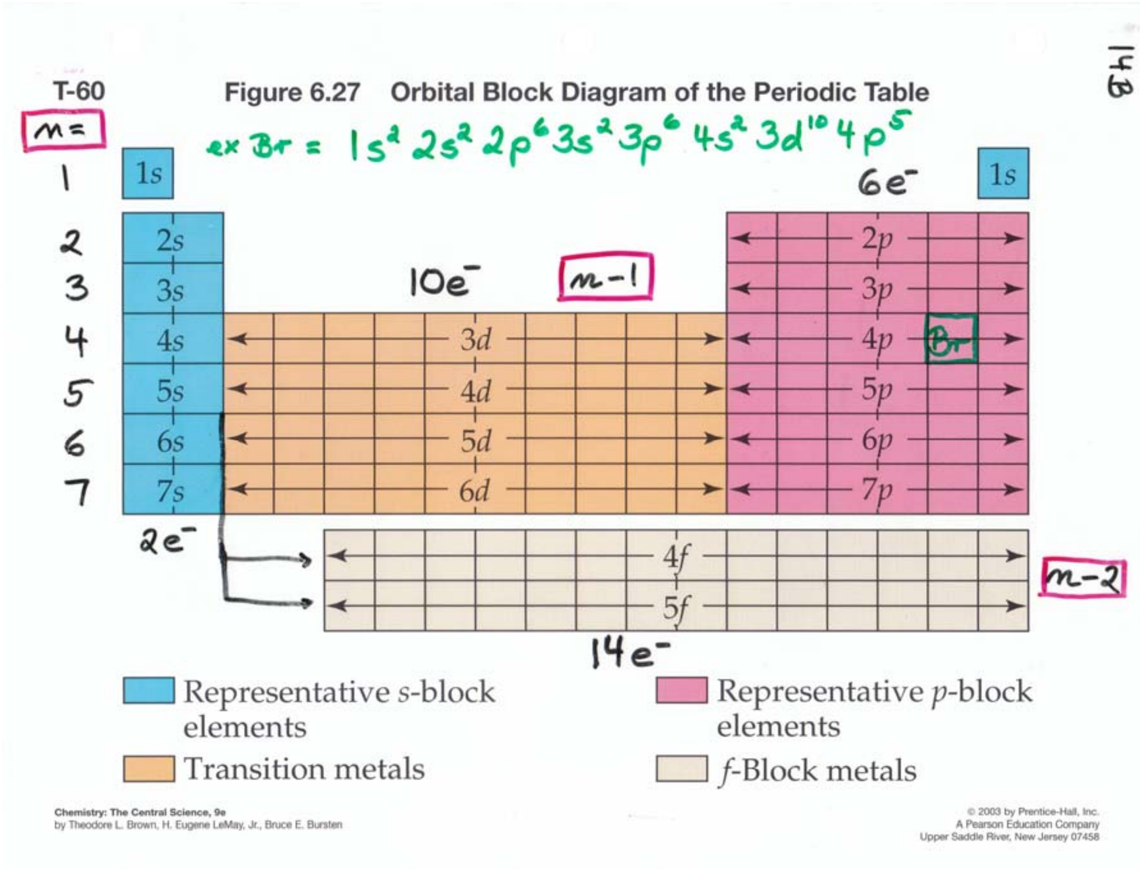
$_{9}$F $1s^2 2s^2 2p^5$
$_{17}$Cl $1s^2 2s^2 2p^6 3s^2 3p^5$
$_{35}$Br $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$
$_{53}$I $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^5$

Group VIIA ____ valence electrons s^2 & p^5

Figure 3.12 Electron configuration and the position of elements in the periodic table.




Period number, n = Principal energy level
Blocks = specific subshells
1 1s
2 2s
3 3s
4 4s
5 5s
6 6s
7 7s
1s
n-1
3d
4d
5d
6d
2p
3p
4p
5p
6p
n-2
4f
5f
hydrogen 1s1
lithium 1s22s1
beryllium 1s22s2
1
2
boron 1s22s22p1
carbon 1s22s22p2
nitrogen 1s22s22p3
oxygen 1s22s22p4
fluorine 1s22s22p5
neon 1s22s22p6


T-60
Figure 6.27 Orbital Block Diagram of the Periodic Table
n =
ex Br = 1s2 2s2 2p6 3s2 3p6 4s2 3d10 4p5
1s
2s
3s
4s
5s
6s
7s
3d
4d
5d
6d
2p
3p
4p
5p
6p
7p
4f
5f
10e-
6e-
2e-
14e-
n-1
n-2
Br
Representative s-block elements
Transition metals
Representative p-block elements
f-Block metals
Chemistry: The Central Science, 9e
by Theodore L. Brown, H. Eugene LeMay, Jr., Bruce E. Bursten
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Supplemental Material & Ch 5.9

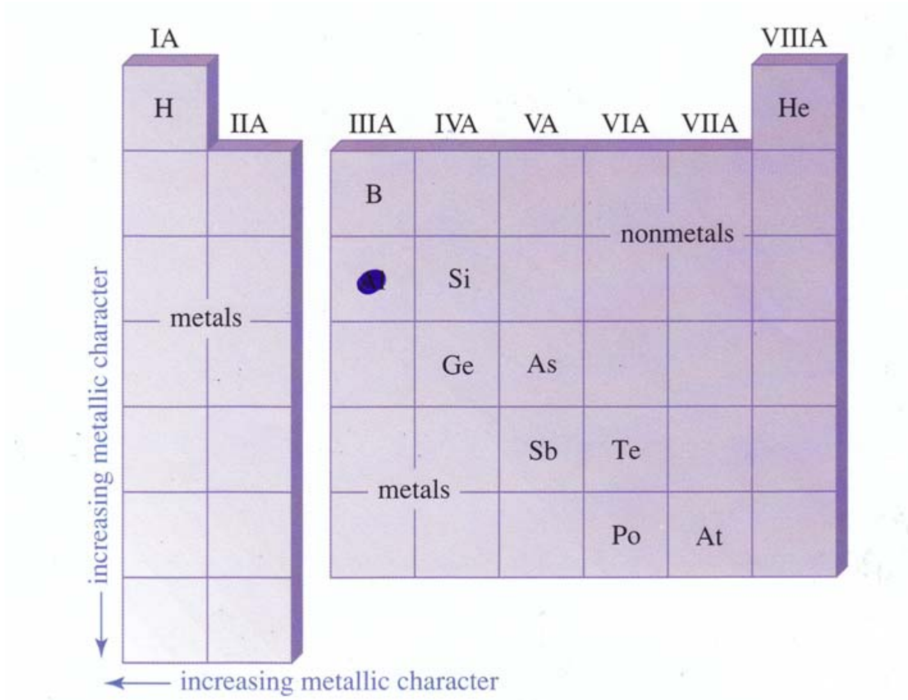
Properties of Atoms and the Periodic Table

A. Chemical properties of an element are related to the position of an element in the periodic table.

demo: compare reactivity of Li, Na, K, Ca, Mg with H_2O

same group	=	same valence electrons	=	similar chemical properties

B. Periodic trends in metallic character for representative elements.



- Metallic character increases with increasing shell number (n) in a group.
- It decreases with increasing number of valence electrons.

C. Periodic trends in atomic radii (pm) for representative elements.

IA	IIA	IIIA	IVA	VA	VIA	VIIA	VIIIA
H 37							He 50
Li 152	Be 111	B 88	C 77	N 70	O 66	F 64	Ne 70
Na 186	Mg 160	Al 143	Si 117	P 110	S 104	Cl 99	Ar 94
K 231	Ca 197	Ga 122	Ge 122	As 121	Se 117	Br 114	Kr 109
Rb 244	Sr 215	In 162	Sn 140	Sb 141	Te 137	I 133	Xe 130
Cs 262	Ba 217	Tl 171	Pb 175	Bi 146	Po 150	At 140	Rn 140

↓ Increasing radius (down a group)

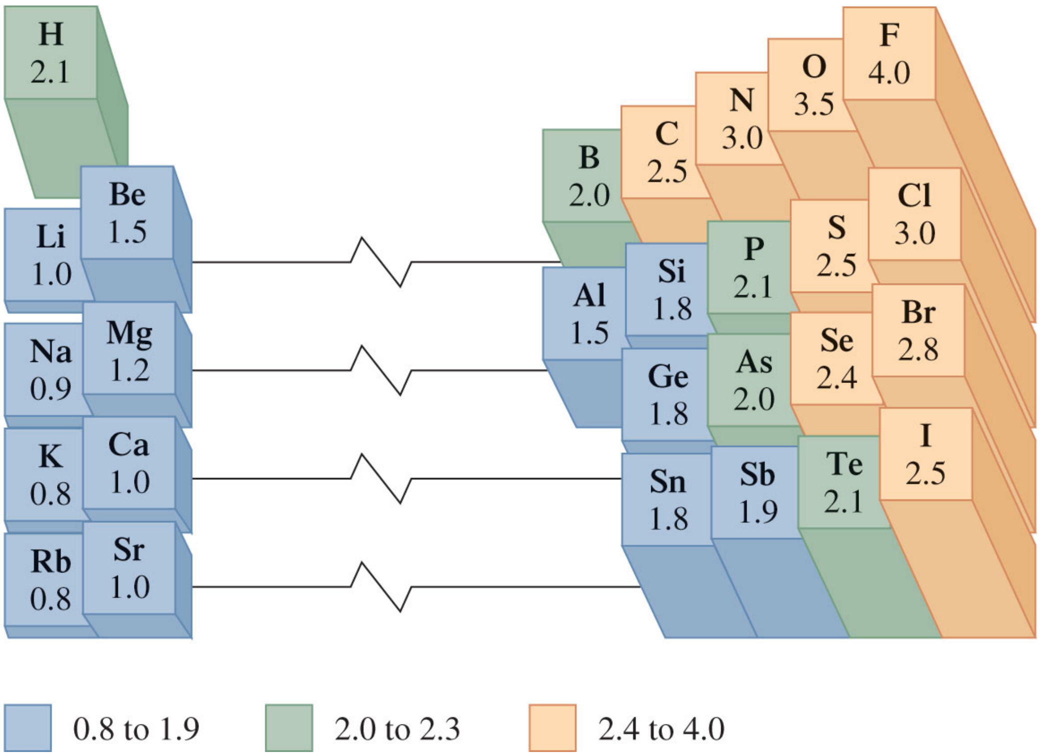
← Increasing radius (across a period)

- Within a group, radii increase with increasing shell number (i.e. increasing distance from the nucleus).
- Across a period, radii decrease as the number of protons in the nucleus increase (i.e. increasing nuclear charge).

D. The type and number of bonds an element forms are related to its position in the periodic table.

Electronegativity is a measure of the relative attraction that an atom has for the shared electrons in a covalent bond.

Figure 5.11 in Stoker



very different values → __________ bonds Chapter 4

similar values → __________ bonds Chapter 5

E. Periodic trends in ionization energies for representative elements.

$$\text{Li}\cdot \quad \rightarrow \quad \text{Li}^{+} \quad + \quad e^{-}$$

neutral atom → positive ion + electron

cation

Ionization of atoms requires Ionization Energy (I.E.)

- Within a period I.E. increases with increasing Z.
- Down a group I.E. decreases with increased atomic radius.