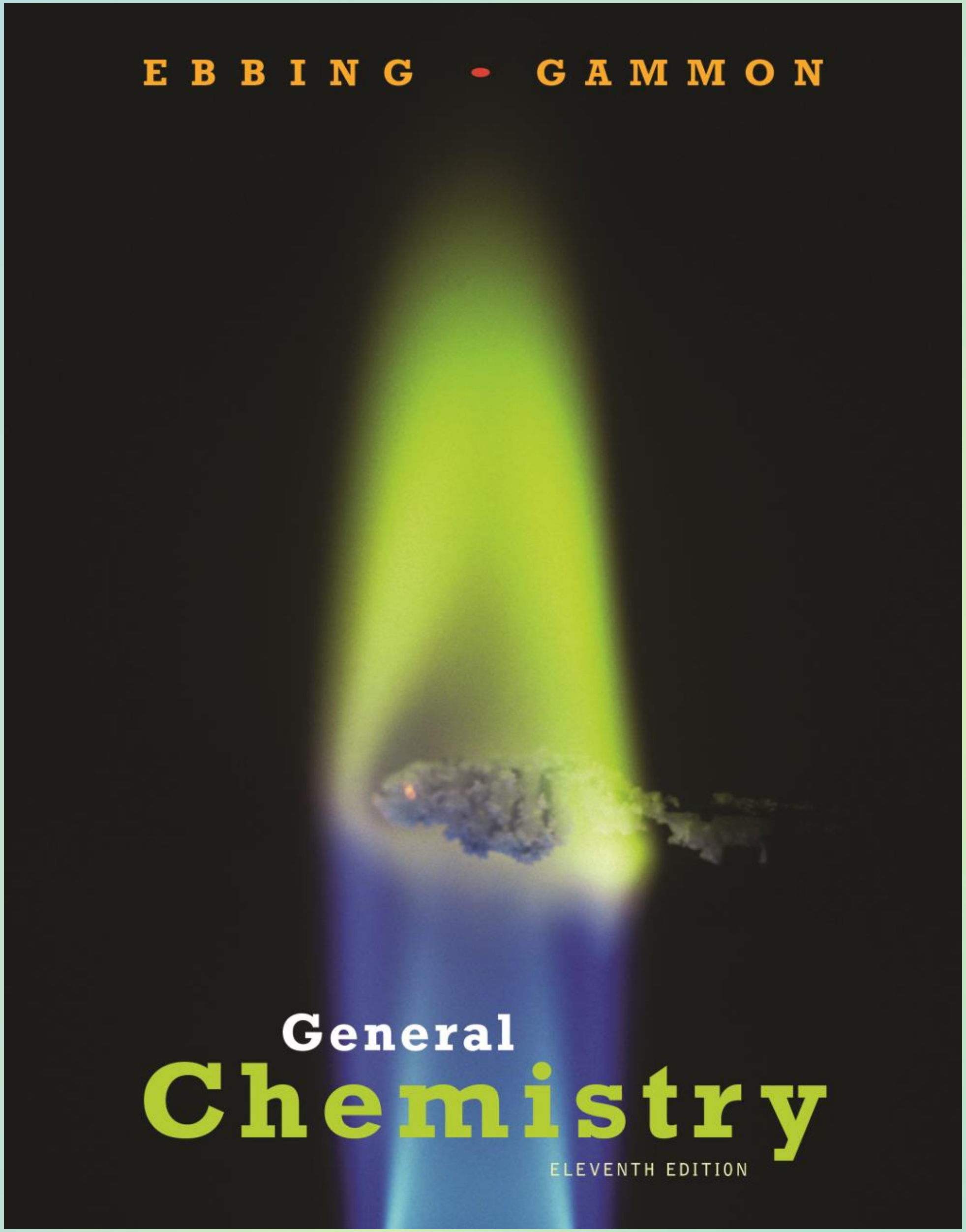
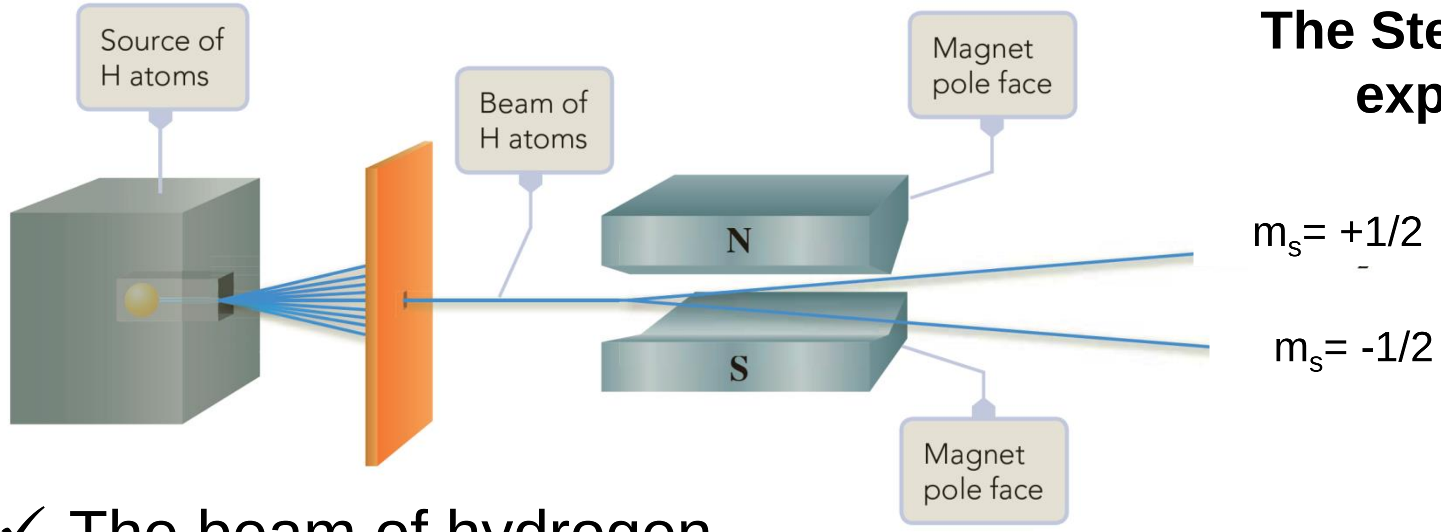


Electron Configurations and Periodicity

8.1 Electron Spin and the Pauli Exclusion Principle

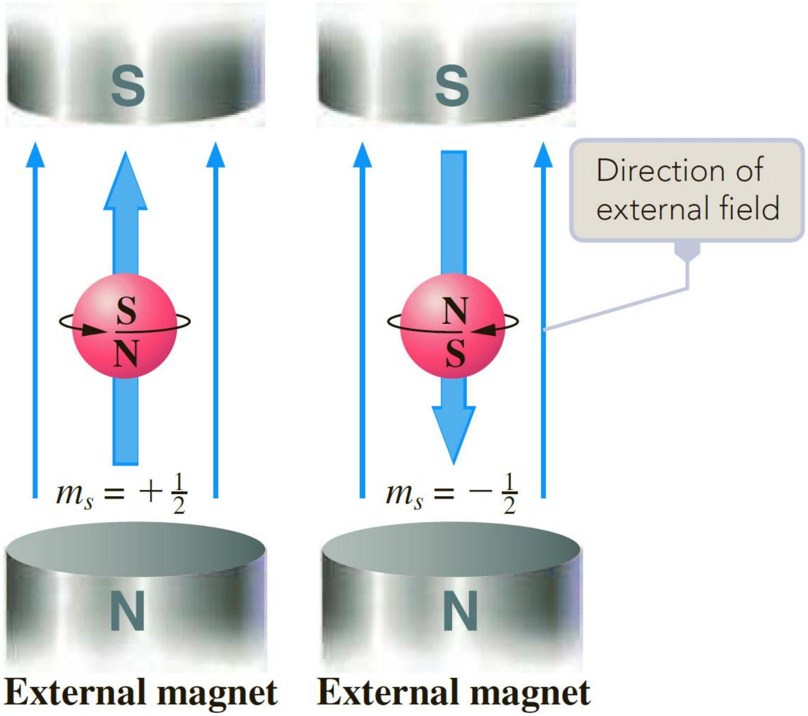
The Stern–Gerlach experiment



✓ The beam of hydrogen atoms is split into two by the magnetic field.

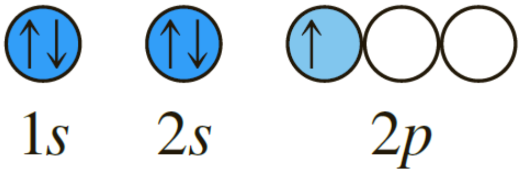
✓ Half of the atoms are bent in one direction and half in the other.

✓ The fact that the atoms are affected by the external magnet shows that they themselves act as magnets.



➢ Electron Configurations and Orbital Diagrams

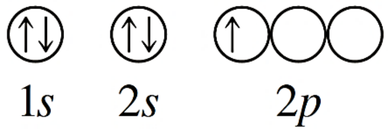
- ✓ An **electron configuration** of an atom is *a particular distribution of electrons among the available subshells.*
- ✓ A subshell consists of a group of orbitals having the same *n* and *l* quantum numbers but different m_l values.
- ✓ $_5B$: $1s^2 2s^2 2p^1$
- ✓ An electron in an orbital is shown by an arrow; the arrow points up when $m_s = +1/2$ and down when $m_s = -1/2$.
- ✓ The orbital diagram of B is:

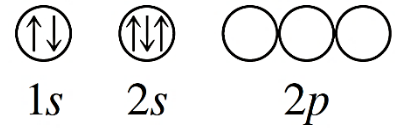


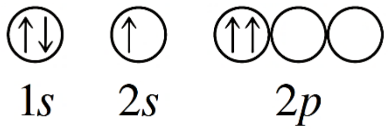
- ✓ **Pauli Exclusion Principle**

No two electrons in an atom can have the same four quantum numbers.

(Q) Which of the following orbital diagrams or electron configurations are possible and which are impossible, according to the Pauli exclusion principle? Explain

a. 

b. 

c. 

d. $1s^3 2s^1$

e. $1s^2 2s^1 2p^7$

f. $1s^2 2s^2 2p^6 3s^2 3p^6 3d^8 4s^2$

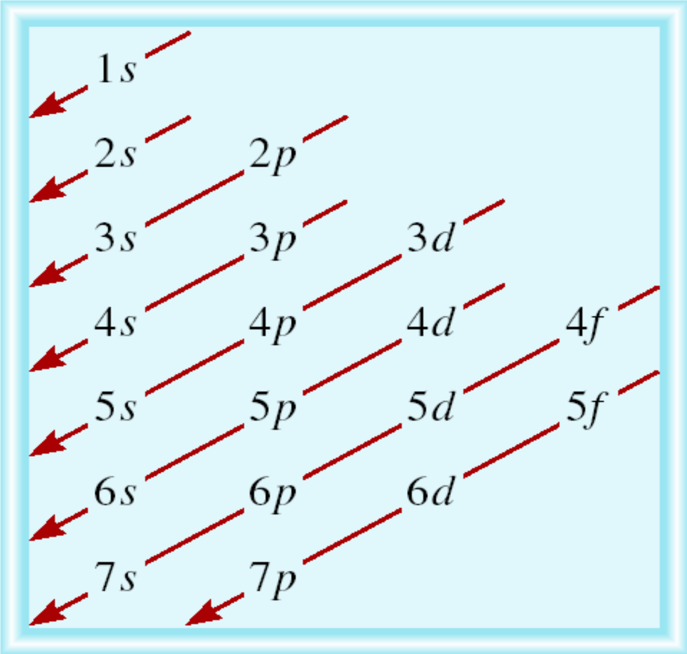
➢ Building-Up Principle (Aufbau Principle)

✓ *lowest energy orbitals are filled first* : 1s, then 2s, then 2p, then 3s, then 3p, etc.

✓ Following this principle, you obtain the electron configuration of an atom by successively filling subshells in the following order: 1*s*, 2*s*, 2*p*, 3*s*, 3*p*, **4*s***, **3*d***, 4*p*, **5*s***, **4*d***, 5*p*, 6*s*, 4*f*, 5*d*, 6*p*, 7*s*, 5*f*.

➢ Electron Configurations and the Periodic Table

Order of orbitals (filling) in multi-electron atom



helium	$1s^2$
neon	$1s^22s^22p^6$
argon	$1s^22s^22p^63s^23p^6$
krypton	$1s^22s^22p^63s^23p^63d^{10}4s^24p^6$

beryllium	$1s^22s^2$	or	$[He]2s^2$
magnesium	$1s^22s^22p^63s^2$	or	$[Ne]3s^2$
calcium	$1s^22s^22p^63s^23p^64s^2$	or	$[Ar]4s^2$

boron	$1s^22s^22p^1$	or	$[He]2s^22p^1$
aluminium	$1s^22s^22p^63s^23p^1$	or	$[Ne]3s^23p^1$
gallium	$1s^22s^22p^63s^23p^63d^{10}4s^24p^1$	or	$[Ar]3d^{10}4s^24p^1$

✓ **noble-gas core:** *an inner-shell configuration corresponding to one of the noble gases.*

✓ **valence electron**: *An electron in an atom outside the noble-gas or pseudo-noble-gas core.*

Main-Group Elements
s subshell fills

Main-Group Elements
p subshell fills

Transition Metals
d subshell fills

1
H
$1s^{1}$

Atomic number
Symbol
Valence-shell configuration

Period	1 (1A)	2 (2A)	3 (3B)	4 (4B)	5 (5B)	6 (6B)	7 (7B)	8	9 (8B)	10	11 (1B)	12 (2B)	13 (3A)	14 (4A)	15 (5A)	16 (6A)	17 (7A)	18 (8A)
1	1 **H** $1s^{1}$																	2 **He** $1s^{2}$
2	3 **Li** $2s^{1}$	4 **Be** $2s^{2}$											5 **B** $2s^{2}2p^{1}$	6 **C** $2s^{2}2p^{2}$	7 **N** $2s^{2}2p^{3}$	8 **O** $2s^{2}2p^{4}$	9 **F** $2s^{2}2p^{5}$	10 **Ne** $2s^{2}2p^{6}$
3	11 **Na** $3s^{1}$	12 **Mg** $3s^{2}$											13 **Al** $3s^{2}3p^{1}$	14 **Si** $3s^{2}3p^{2}$	15 **P** $3s^{2}3p^{3}$	16 **S** $3s^{2}3p^{4}$	17 **Cl** $3s^{2}3p^{5}$	18 **Ar** $3s^{2}3p^{6}$
4	19 **K** $4s^{1}$	20 **Ca** $4s^{2}$	21 **Sc** $3d^{1}4s^{2}$	22 **Ti** $3d^{2}4s^{2}$	23 **V** $3d^{3}4s^{2}$	24 **Cr** $3d^{5}4s^{1}$	25 **Mn** $3d^{5}4s^{2}$	26 **Fe** $3d^{6}4s^{2}$	27 **Co** $3d^{7}4s^{2}$	28 **Ni** $3d^{8}4s^{2}$	29 **Cu** $3d^{10}4s^{1}$	30 **Zn** $3d^{10}4s^{2}$	31 **Ga** $4s^{2}4p^{1}$	32 **Ge** $4s^{2}4p^{2}$	33 **As** $4s^{2}4p^{3}$	34 **Se** $4s^{2}4p^{4}$	35 **Br** $4s^{2}4p^{5}$	36 **Kr** $4s^{2}4p^{6}$
5	37 **Rb** $5s^{1}$	38 **Sr** $5s^{2}$	39 **Y** $4d^{1}5s^{2}$	40 **Zr** $4d^{2}5s^{2}$	41 **Nb** $4d^{4}5s^{1}$	42 **Mo** $4d^{5}5s^{1}$	43 **Tc** $4d^{5}5s^{2}$	44 **Ru** $4d^{7}5s^{1}$	45 **Rh** $4d^{8}5s^{1}$	46 **Pd** $4d^{10}$	47 **Ag** $4d^{10}5s^{1}$	48 **Cd** $4d^{10}5s^{2}$	49 **In** $5s^{2}5p^{1}$	50 **Sn** $5s^{2}5p^{2}$	51 **Sb** $5s^{2}5p^{3}$	52 **Te** $5s^{2}5p^{4}$	53 **I** $5s^{2}5p^{5}$	54 **Xe** $5s^{2}5p^{6}$
6	55 **Cs** $6s^{1}$	56 **Ba** $6s^{2}$	57-71 Lanthanides	72 **Hf** $5d^{2}6s^{2}$	73 **Ta** $5d^{3}6s^{2}$	74 **W** $5d^{4}6s^{2}$	75 **Re** $5d^{5}6s^{2}$	76 **Os** $5d^{6}6s^{2}$	77 **Ir** $5d^{7}6s^{2}$	78 **Pt** $5d^{9}6s^{1}$	79 **Au** $5d^{10}6s^{1}$	80 **Hg** $5d^{10}6s^{2}$	81 **Tl** $6s^{2}6p^{1}$	82 **Pb** $6s^{2}6p^{2}$	83 **Bi** $6s^{2}6p^{3}$	84 **Po** $6s^{2}6p^{4}$	85 **At** $6s^{2}6p^{5}$	86 **Rn** $6s^{2}6p^{6}$
7	87 **Fr** $7s^{1}$	88 **Ra** $7s^{2}$	89-103 Actinides	104 **Rf** $6d^{2}7s^{2}$	105 **Db** $6d^{3}7s^{2}$	106 **Sg** $6d^{4}7s^{2}$	107 **Bh** $6d^{5}7s^{2}$	108 **Hs** $6d^{6}7s^{2}$	109 **Mt** $6d^{7}7s^{2}$	110 **Uun** $6d^{8}7s^{2}$	111 **Rg** $6d^{9}7s^{2}$	112 **Cn** $6d^{10}7s^{2}$	113 **Uut** $7s^{2}7p^{1}$	114 **Uuq** $7s^{2}7p^{2}$	115 **Uup** $7s^{2}7p^{3}$	116 **Uuh** $7s^{2}7p^{4}$	117 **Uus** $7s^{2}7p^{5}$	118 **Uuo** $7s^{2}7p^{6}$

Inner Transition Metals
f subshell fills

57 **La** $5d^{1}6s^{2}$	58 **Ce** $4f^{1}5d^{1}6s^{2}$	59 **Pr** $4f^{3}6s^{2}$	60 **Nd** $4f^{4}6s^{2}$	61 **Pm** $4f^{5}6s^{2}$	62 **Sm** $4f^{6}6s^{2}$	63 **Eu** $4f^{7}6s^{2}$	64 **Gd** $4f^{7}5d^{1}6s^{2}$	65 **Tb** $4f^{9}6s^{2}$	66 **Dy** $4f^{10}6s^{2}$	67 **Ho** $4f^{11}6s^{2}$	68 **Er** $4f^{12}6s^{2}$	69 **Tm** $4f^{13}6s^{2}$	70 **Yb** $4f^{14}6s^{2}$	71 **Lu** $4f^{14}5d^{1}6s^{2}$
89 **Ac** $6d^{1}7s^{2}$	90 **Th** $6d^{2}7s^{2}$	91 **Pa** $5f^{2}6d^{1}7s^{2}$	92 **U** $5f^{3}6d^{1}7s^{2}$	93 **Np** $5f^{4}6d^{1}7s^{2}$	94 **Pu** $5f^{6}7s^{2}$	95 **Am** $5f^{7}7s^{2}$	96 **Cm** $5f^{7}6d^{1}7s^{2}$	97 **Bk** $5f^{9}7s^{2}$	98 **Cf** $5f^{10}7s^{2}$	99 **Es** $5f^{11}7s^{2}$	100 **Fm** $5f^{12}7s^{2}$	101 **Md** $5f^{13}7s^{2}$	102 **No** $5f^{14}7s^{2}$	103 **Lr** $5f^{14}7s^{2}7p^{1}$

Main-group elements
Transition metals
Inner transition metals

- **Exceptions to the Building-Up Principle**

chromium (Cr)(Z =24): [Ar]$3d^4 4s^2$ **expected**

[Ar]$3d^5 4s^1$ **experimental**

copper (Cu)(Z =29): [Ar]$3d^9 4s^2$ **expected**

[Ar]$3d^{10} 4s^1$ **experimental**

8.3 Writing Electron Configurations Using the Periodic Table

Kr(36): [Ar]$4s^2 3d^{10} 4p^6$ or [Ar]$3d^{10} 4s^2 4p^6$

Nb(41): [Kr]$4s^2 3d^{10} 4p^6$

Sb(51): [Kr]$5s^2 4d^{10} 5p^3$

(Q) Use the building-up principle to obtain the configuration for the ground state of the gallium atom ($Z = 31$). Give the configuration in complete form (do not abbreviate for the core). What is the valence-shell configuration?

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^1$ → $\mathbf{1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^1}$

→The valence-shell configuration is $\mathbf{4s^2 4p^1}$

(Q) What are the configurations for the outer electrons of:

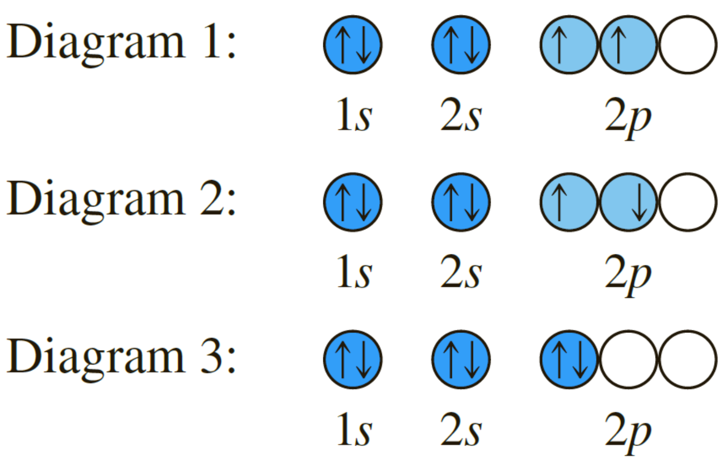
a. tellurium, $Z = 52$, and b. nickel, $Z = 28$?

Te : $[Kr]5s^2 4d^{10} 5p^4$ → $[Kr]\ 4d^{10}\ 5s^2 5p^4$ → $5s^2 5p^4$

Ni : $[Ar]4s^2 3d^8$ → $4s^2 3d^8$

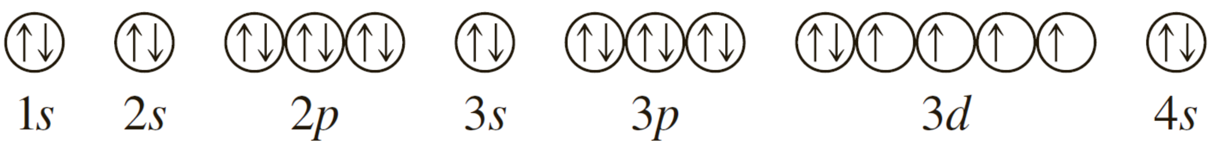
Exercise 8.4 The atom (X) has the ground-state configuration $[Xe]4f^{14}\ 5d^{10} 6s^2 6p^2$. Find the period and group for this element. From its position in the periodic table, would you classify lead as a main-group element, a transition element, or an inner transition element?

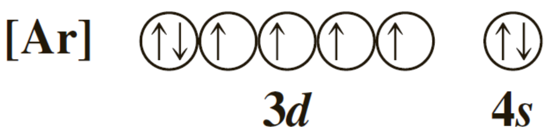
8.4 Orbital Diagrams of Atoms; Hund's Rule



Hund's rule states that *the lowest energy arrangement of electrons in a subshell is obtained by putting electrons into separate orbitals of the subshell with the same spin before pairing electrons*

(Q)Write an orbital diagram for the ground state of the iron atom.





✓ Magnetic Properties of Atoms

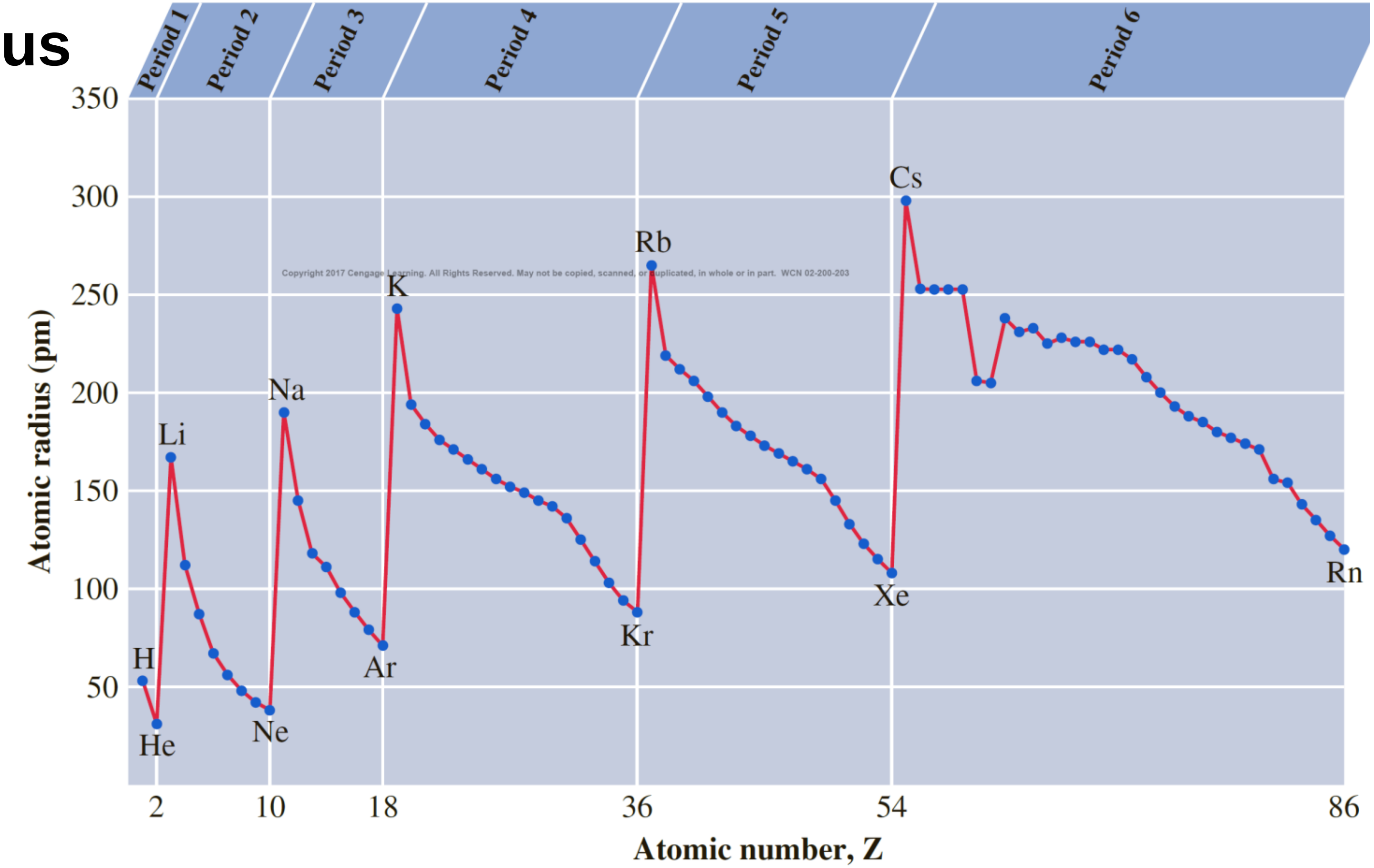
paramagnetic substance

At least one unpaired electron

diamagnetic substance

All electrons are paired

➢ Atomic Radius



1. Within each period (horizontal row), the atomic radius tends to **decrease** with increasing atomic number (nuclear charge).

2. Within each group (vertical column), the atomic radius tends to **increase** with the period number.

(Q) Arrange the following in order of increasing atomic radius:

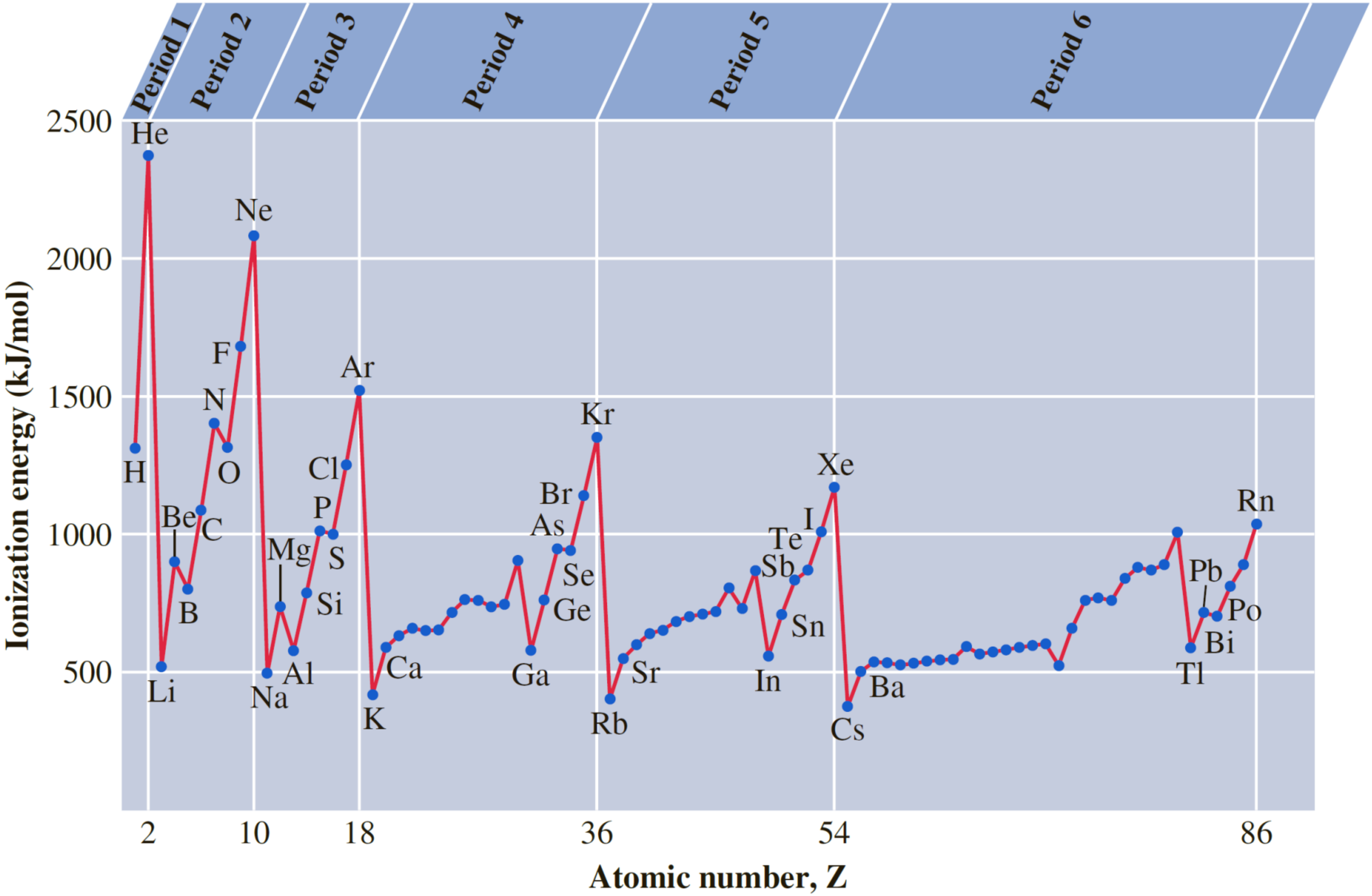
(1) Al, C, Si. **C < Si < Al**

(2) Na, Be, Mg **Be < Mg < Na**

➢ Ionization Energy

$Li \rightarrow Li^{+} + e^{-}$ First ionization energy = 520 kJ/mol

= 5.39 eV (1eV = 96.5 kJ/mol)



• high values of *first ionization energy* associated with the noble gases

•very low values of *first ionization energy* associated with the group 1 elements;

•general increase in values of *first ionization energy* as a given period is crossed

•Ionization energies tend to decrease going down any column. This is because atomic size increases going down the column.

Exceptions:

•3A element (B)(ns^2np^1) has smaller ionization energy than the preceding 2A element (Be)(ns^2). Or (Al vs. Mg)

•6A element (O)(ns^2np^4) has smaller ionization energy than the preceding 5A element (N) (ns^2np^3) . Or (S vs. P)

As a result of electron repulsion

Table 8.3 Successive Ionization Energies of the First Ten Elements (kJ/mol)*

Element	First	Second	Third	Fourth	Fifth	Sixth	Seventh
H	1312						
He	2372	5250					
Li	520	7298	11,815				
Be	900	1757	14,848	21,006			
B	801	2427	3660	25,026	32,827		
C	1086	2353	4620	6223	37,831	47,277	
N	1402	2856	4578	7475	9445	53,267	64,360
O	1314	3388	5300	7469	10,990	13,326	71,330
F	1681	3374	6050	8408	11,023	15,164	17,868
Ne	2081	3952	6122	9371	12,177	15,238	19,999

Exercise 8.7 The first ionization energy of the chlorine atom is 1251 kJ/mol. which of the following values would be the more likely ionization energy for the iodine atom. Explain.

a. 1000 kJ/mol.   b. 1400 kJ/mol.

8.26 Which of the following atoms, designated by their electron configurations, has the *highest* ionization energy?

a $[Ne]3s^2 3p^2$
b $[Ne]3s^2 3p^3$
c $[Ar]3d^{10} 4s^2 4p^3$
d $[Kr]4d^{10} 5s^2 5p^3$
e $[Xe]4f^{14} 5d^{10} 6s^2 6p^3$

8.27 When trying to remove electrons from Be, which of the following sets of ionization energy makes the most sense going from first to third ionization energy? Explain your answer.

a First IE 900 KJ/mol, second IE 1750 kJ/mol, third IE 15,000 kJ/mol
b First IE 1750 KJ/mol, second IE 900 kJ/mol, third IE 15,000 kJ/mol
c First IE 15,000 KJ/mol, second IE 1750 kJ/mol, third IE 900 kJ/mol
d First IE 900 KJ/mol, second IE 15,000 kJ/mol, third IE 22,000 kJ/mol
e First IE 900 KJ/mol, second IE 1750 kJ/mol, third IE 1850 kJ/mol

8.28 Consider the following orderings.

I. $Al < Si < P < S$

II. $Be < Mg < Ca < Sr$

III. $I < Br < Cl < F$

Which of these give(s) a correct trend in atomic size?

a) I only
b) II only
c) III only
d) I and II only
e) II and III only

➢ Electron Affinity

is defined as *the energy required to remove an electron from the atom's negative ion (in its ground state)*

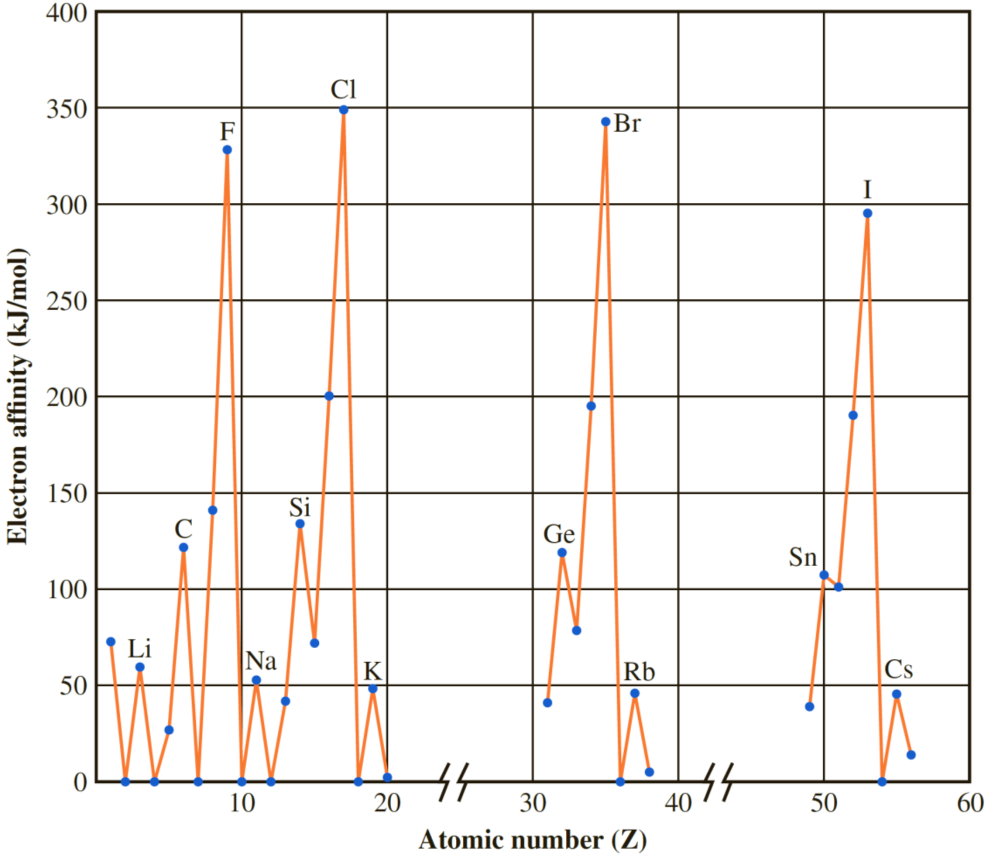
$$Cl([Ne]3s^2 3p^5) + e^- \longrightarrow Cl^-([Ne]3s^2 3p^6);\ \Delta E = -349\ kJ/mol$$

$$Cl^-([Ne]3s^2 3p^6) \longrightarrow Cl([Ne]3s^2 3p^5) + e^-;\ \Delta E = 349\ kJ/mol$$

$$Ge([Ar]4s^2 4p^2) + e^- \longrightarrow Ge^-([Ar]4s^2 4p^3);\ \Delta E = -119\ kJ/mol\ (EA = 119\ kJ/mol)$$

$$As([Ar]4s^2 4p^3) + e^- \longrightarrow As^-([Ar]4s^2 4p^4);\ \Delta E = -78\ kJ/mol\ (EA = 78\ kJ/mol)$$

✓ In the Group 5A element (arsenic, As, in the above list), the added electron must pair up with one of the *np* electrons since all of the *np* orbitals have one electron, whereas in the preceding element the extra electron goes into an empty *np* orbital. This pairing of electrons in an orbital requires some energy, resulting in a smaller electron affinity for the Group 5A element compared with the preceding 4A element.



- ✓ In a given period, the electron affinity rises from the Group 1A element to the Group 7A element but with sharp drops in the Group 2A and Group 5A elements.
- ✓ Group 8A elements (noble gases) have zero or small negative values (indicating unstable negative ions)
- ✓ Group 6A and Group 7A elements have the largest electron affinities of any of the other main-group elements