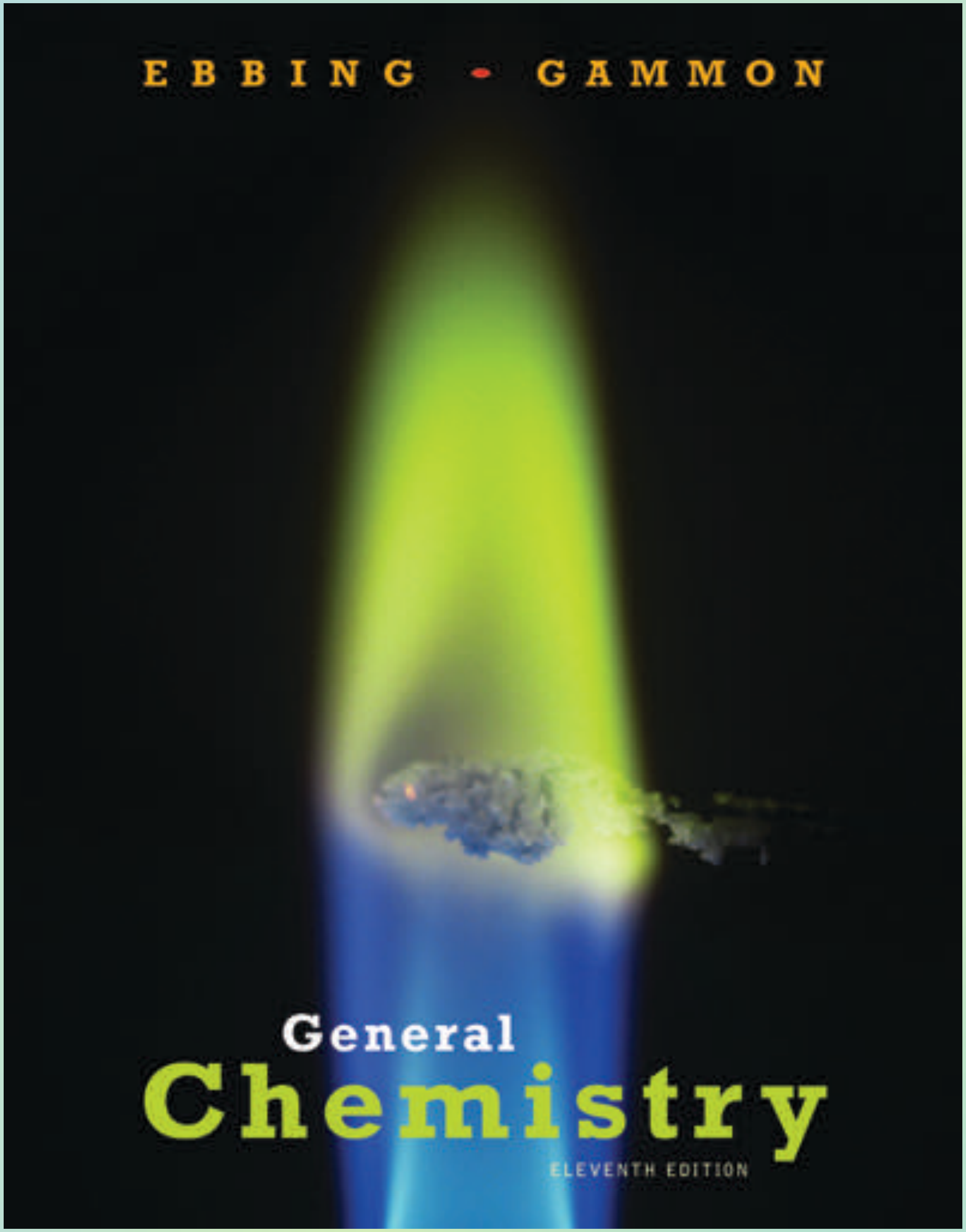
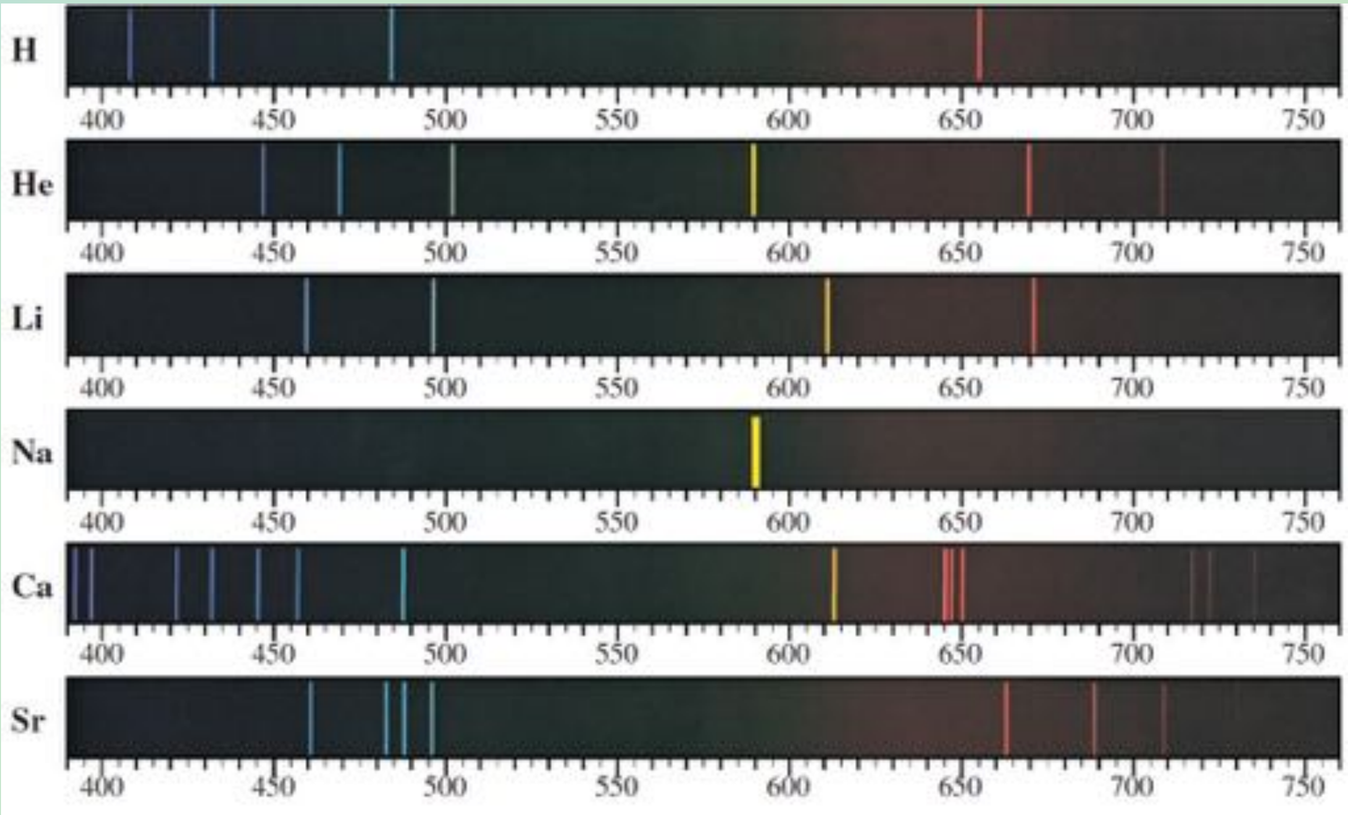
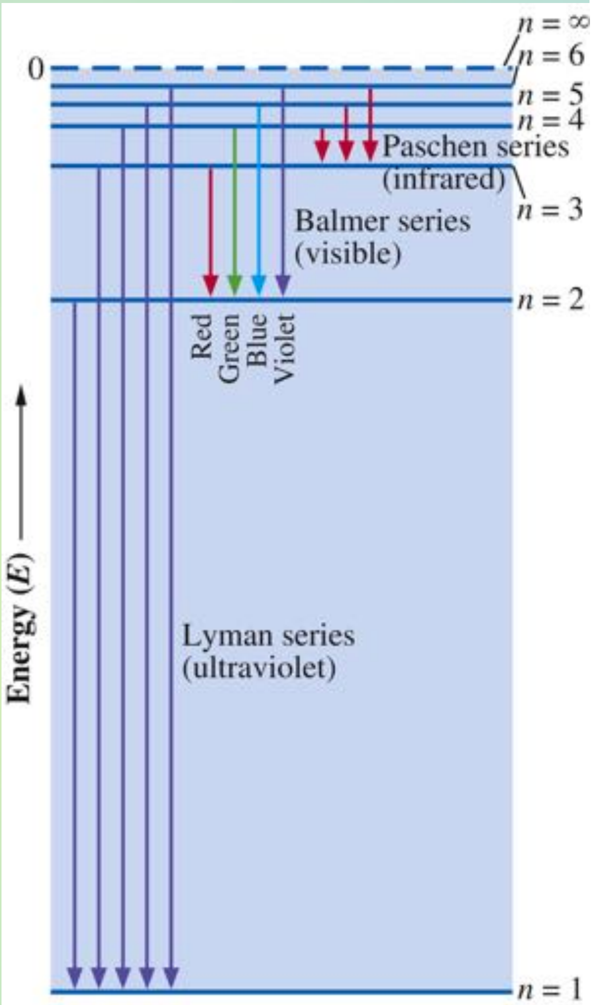


Chapter 8 Electron Configurations and Periodicity

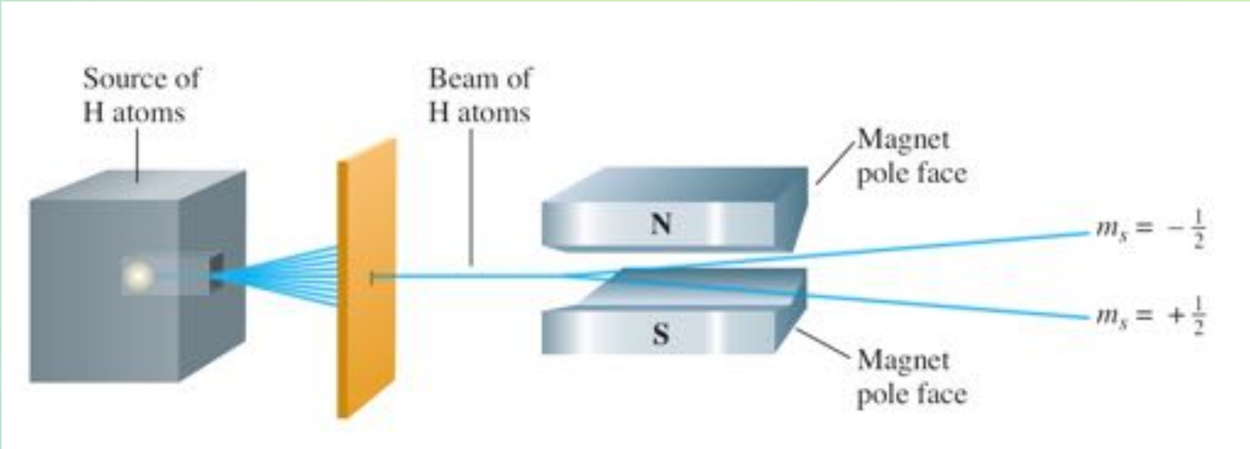
Excited atoms emit specific energies of electromagnetic radiation.

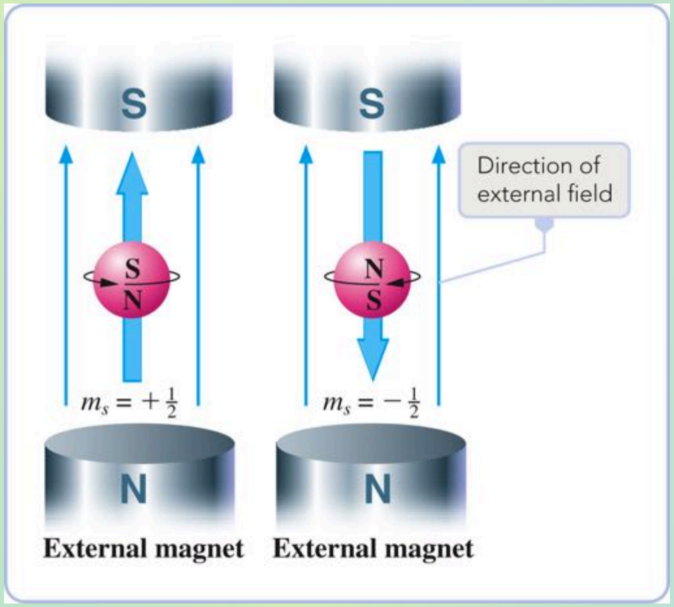




Electrons behave as though they are spinning.

- In 1921, Otto Stern and Walther Gerlach first observed electron spin magnetism. In the diagram below, a beam of hydrogen atoms divides in two while passing through a magnetic field.
- This leads to two values of m_s: +½ and -½.





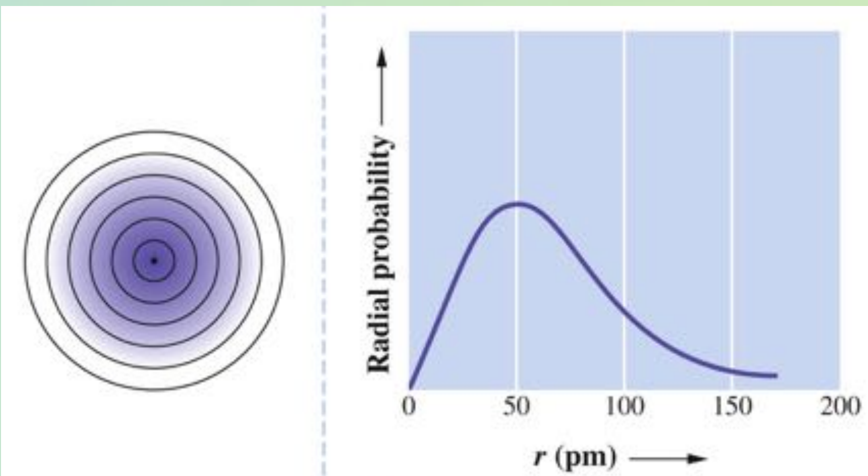
Atoms with unpaired electrons act as magnets.

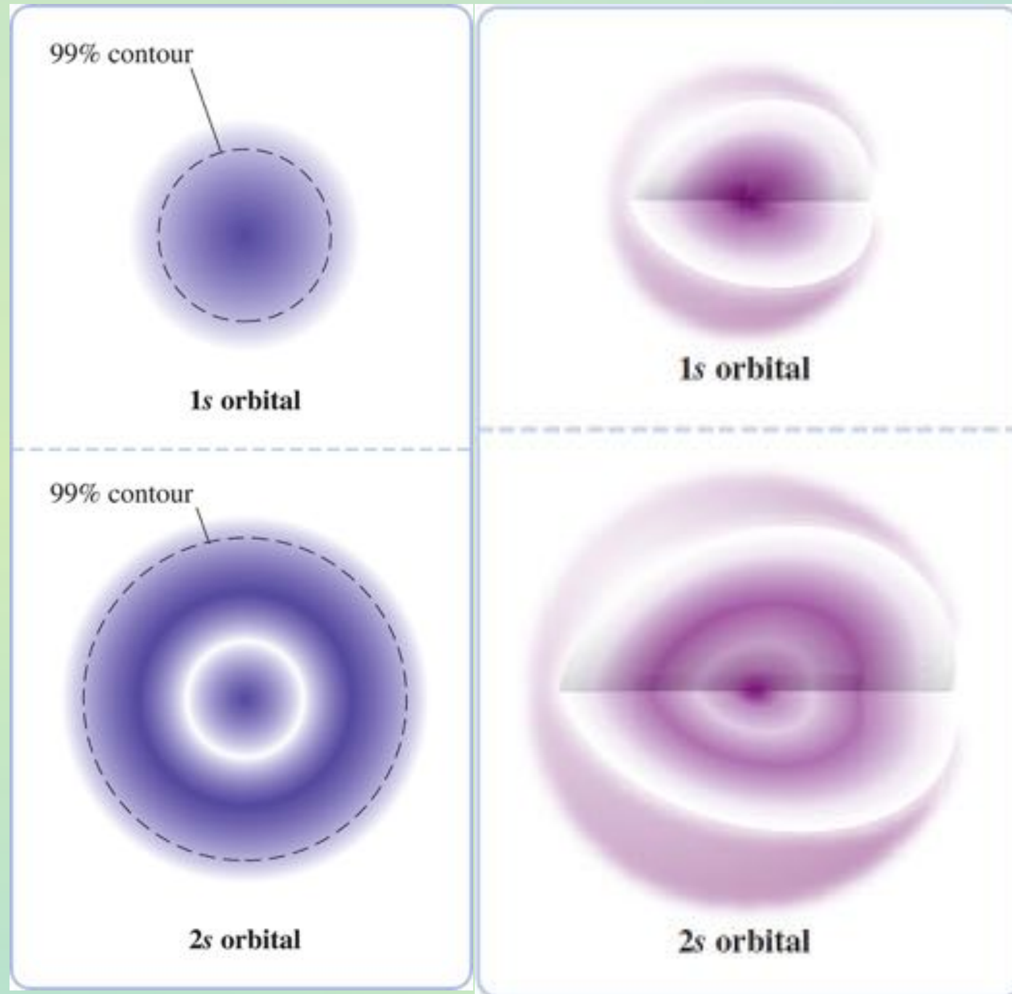
- Although an electron behaves like a tiny magnet, two electrons that are opposite in spin cancel each other.
- Only atoms with unpaired electrons exhibit magnetism.
- This allows for the classification of atoms based on their behavior in a magnetic field.

A **paramagnetic substance** is one that is weakly attracted by a magnetic field, usually as the result of unpaired electrons.

A **diamagnetic substance** is not attracted by a magnetic field generally because it has only paired electrons.

The structure of orbital wave functions is more complex than we draw them.





Capacities of Orbitals

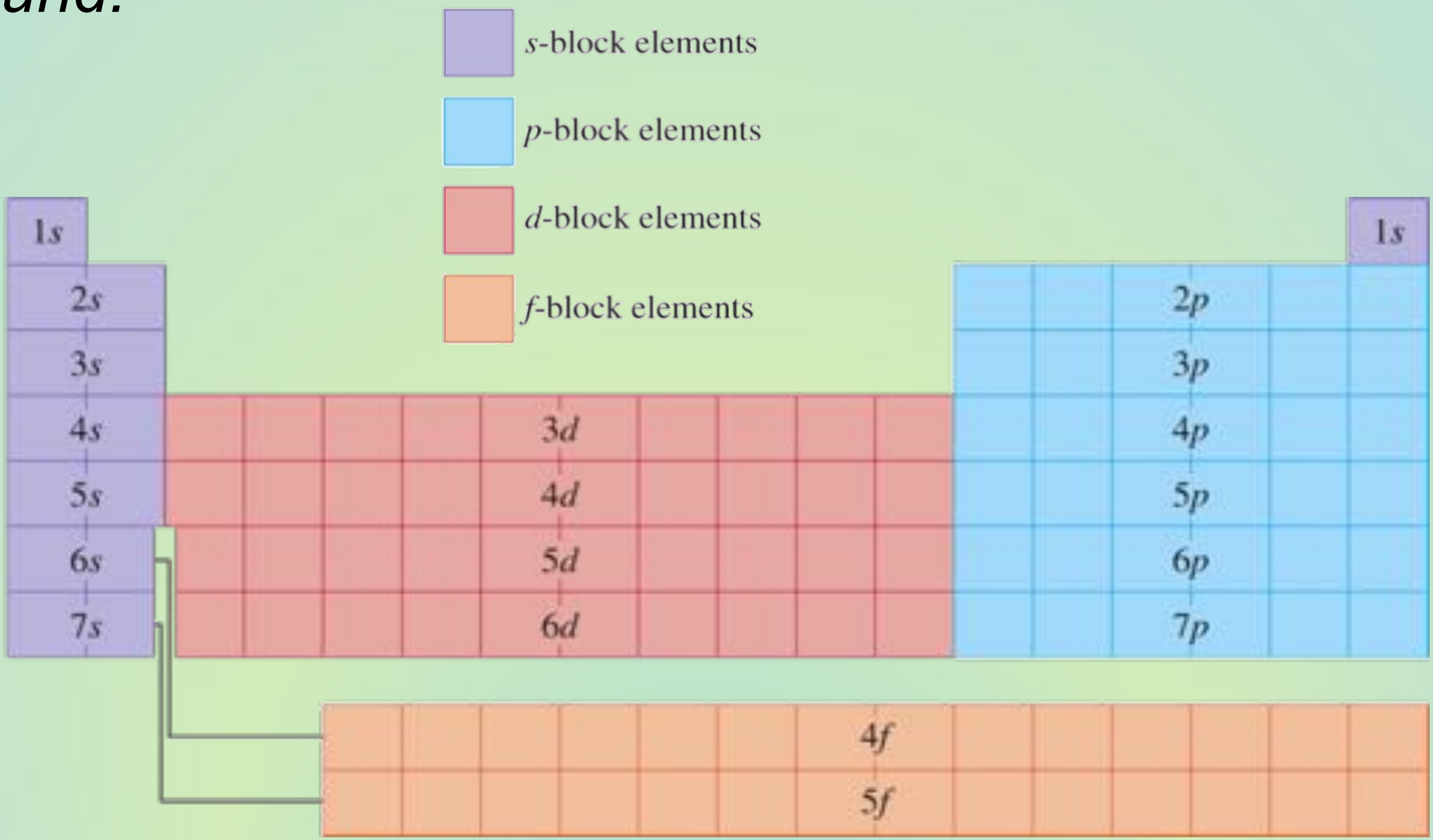
An *s* (l = 0) subshell, with one orbital, can hold a maximum of 2 electrons.

A *p* (l = 1) subshell, with three orbitals, can hold a maximum of 6 electrons.

A *d* (l = 2) subshell, with five orbitals, can hold a maximum of 10 electrons.

An *f* (l = 3) subshell, with seven orbitals, can hold a maximum of 14 electrons.

Orbitals are filled with electrons according to the section of the periodic table in which the element is found.



Practice

Determine which of the following electron configurations are possible and which are impossible. Provide explanations.

A. $1s^32s^1$

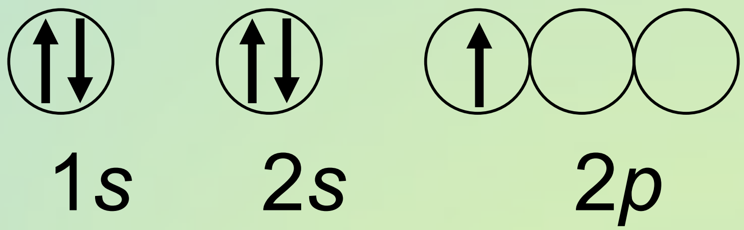
B. $1s^22s^12p^7$

C. $1s^22s^22p^63s^23p^63d^84s^2$

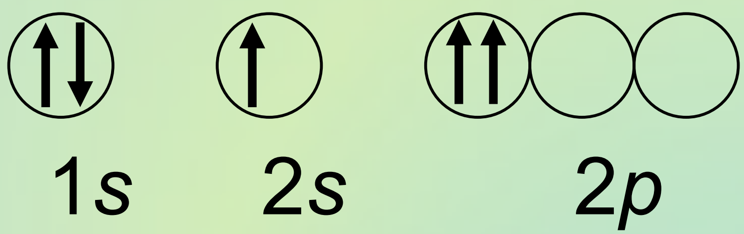
Practice

Determine which of the following orbital diagrams are possible and which are impossible. Provide explanations.

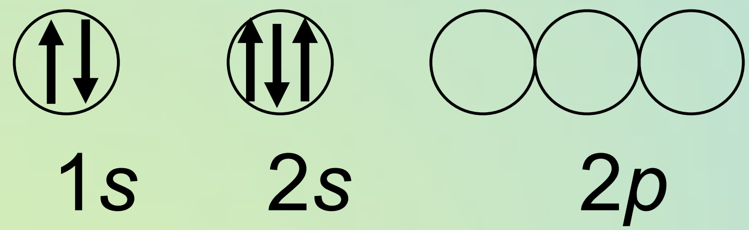
A.



B.



C.



Practice

Write the complete electron configuration of the arsenic atom, As, using the building-up principle.

For arsenic, As, $Z = 33$.

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

Practice

Write the complete and valence electron configuration for gallium.

	1*s*	2*s* 2*p*	3*s* 3*p*	4*s* 3*d* 4*p*
Period:	First	Second	Third	Fourth

Fill the subshells with electrons up to a total of 31 electrons. This gives:

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

The valence shell configuration is $4s^2 4p^1$

For main group elements, only the highest value of *n* counts as the valence shell.

Practice

Determine the configurations for the outer electrons for tellurium, whose atomic number is
$Z = 52$.

The location of Te in the periodic table is Period 5, Group 6A.

Te is a main-group element, and the outer subshells are 5*s* and 5*p.*

Since the group is 6A, the sub-shells contain six electrons.

The valence-shell configuration is $5s^2 5p^4$.

Periodic Law

When the elements are arranged by atomic number, their physical and chemical properties vary periodically.

Physical properties following the Periodic Law:

- Atomic radius
- Effective nuclear charge
- Ionization energy
- Electron affinity

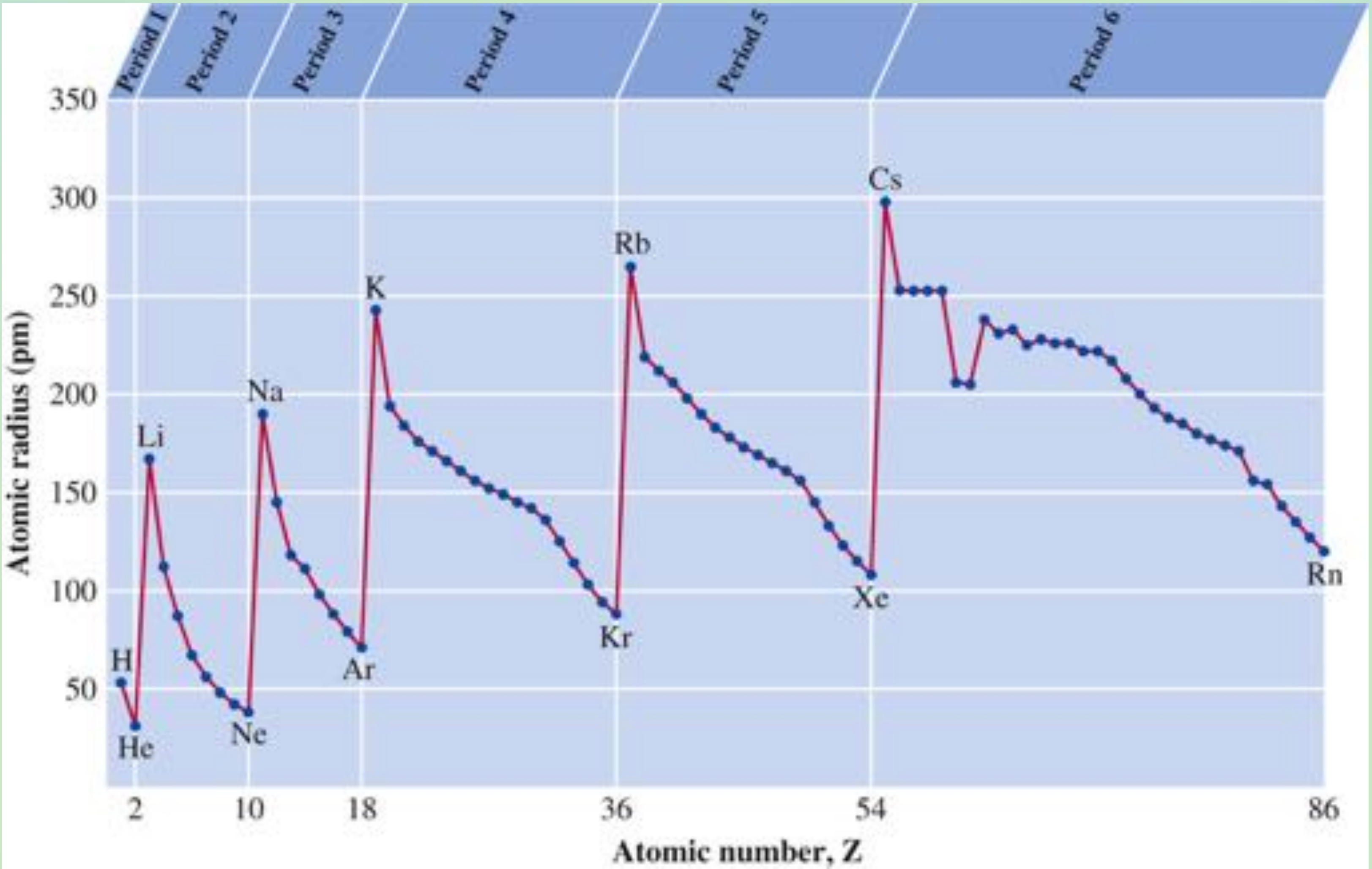
Effective Nuclear Charge

It is the positive charge that an electron experiences from the nucleus. It is equal to the nuclear charge, but is reduced by shielding or screening from any intervening electron distribution.

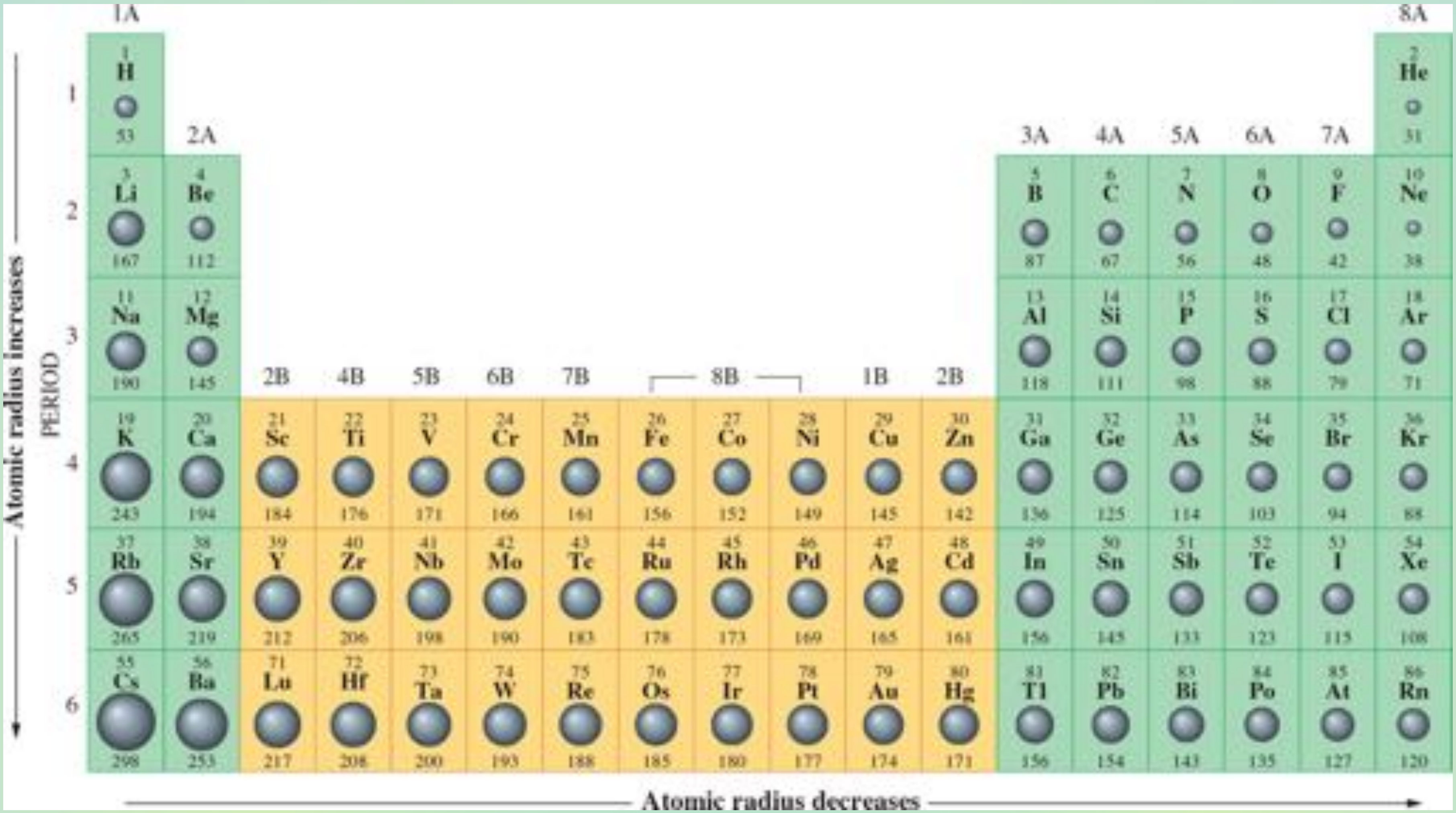
Effective nuclear charge increases across a period. Because the shell number (*n*) is the same across a period, each successive atom experiences a stronger nuclear charge.

Therefore, the size of the outermost orbital and the radius of the atom decreases with the increase in atomic number.

Atomic radius decreases across a period because the effective nuclear charge is increasing.



Atomic radii decrease from left to right, across a row/period.



Practice

Refer to a periodic table and arrange the following elements in order of increasing atomic radius: Ar, Se, S.

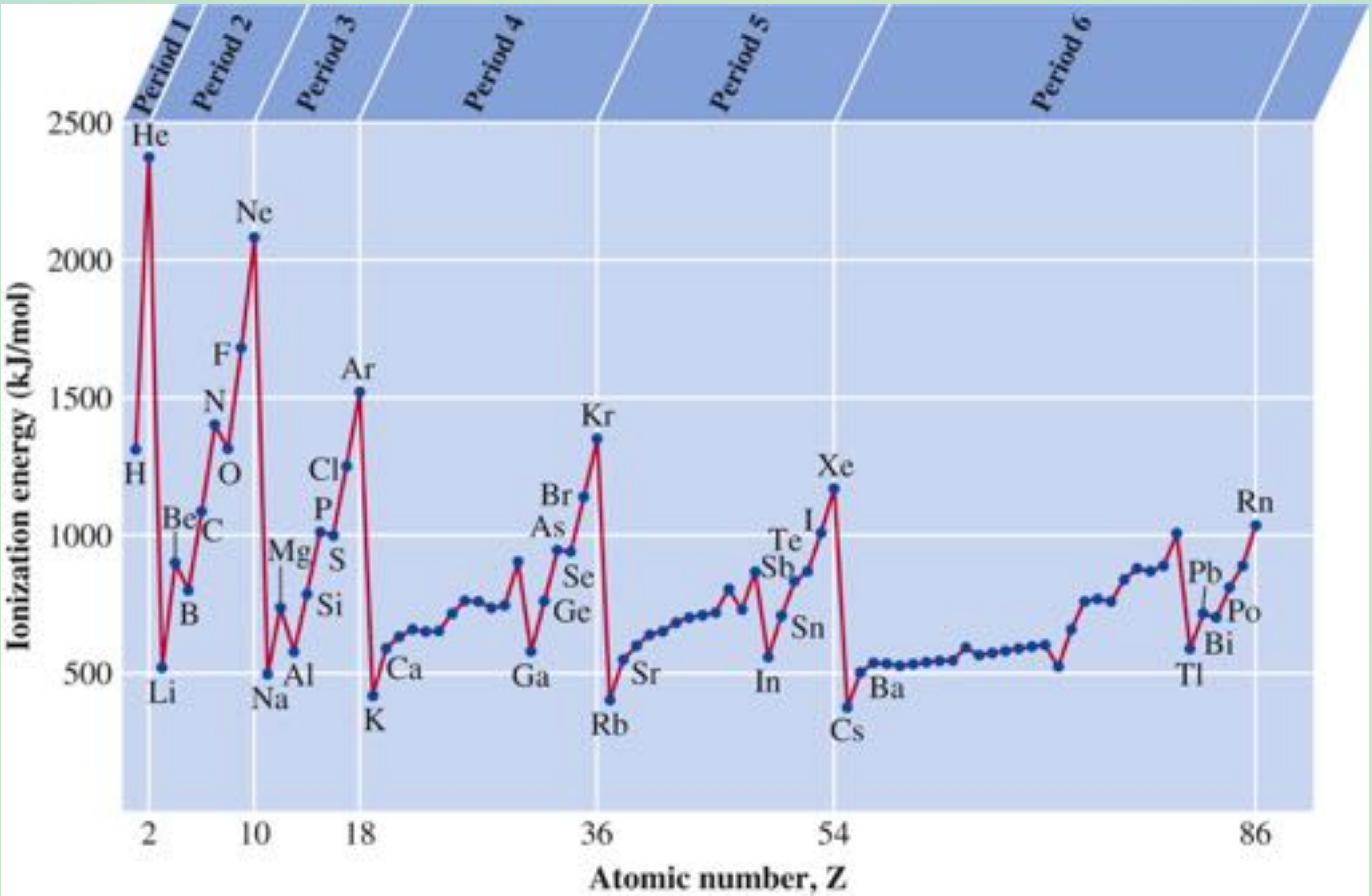
- Se is below S group in 6A. Hence, its ionization energy should be less than that of S.
- S and Ar are in the same period. Z increases from S to Ar. Hence, the ionization energy of S should be lesser than that of Ar.

Se < S < Ar

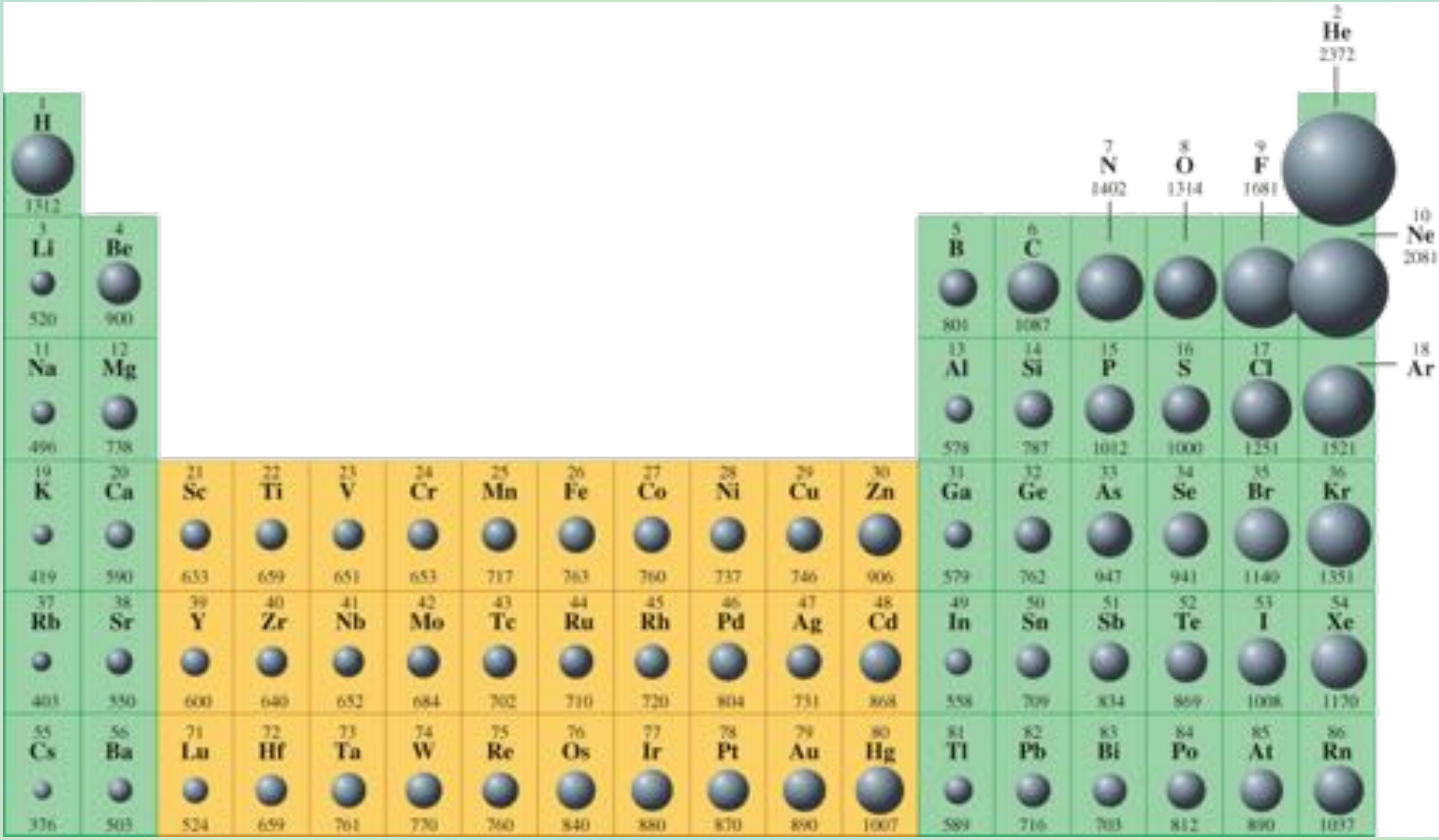
First Ionization Energy

- Also called first ionization potential
- Minimum energy needed to remove the highest-energy (outermost) electron from a neutral, gaseous atom
- Tend to increase with the increase in atomic number, but small deviations occur (between Groups 2A/3A, 5A/6A)
- Proportional to the effective nuclear charge divided by the average distance between the electron and the nucleus. Because the distance between the electron and the nucleus is inversely proportional to the effective nuclear charge, ionization energy is proportional to the square of the effective nuclear charge.

Trends in Ionization Energy



The size of each sphere indicates the size of the ionization energy.



Second and Third Ionization Energies

- Each successive ionization energy increases, because removing electrons from positive ions iof increasing charge becomes more difficult.
- Ionizations become difficult once a stable noble-gas configuration is obtained.

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

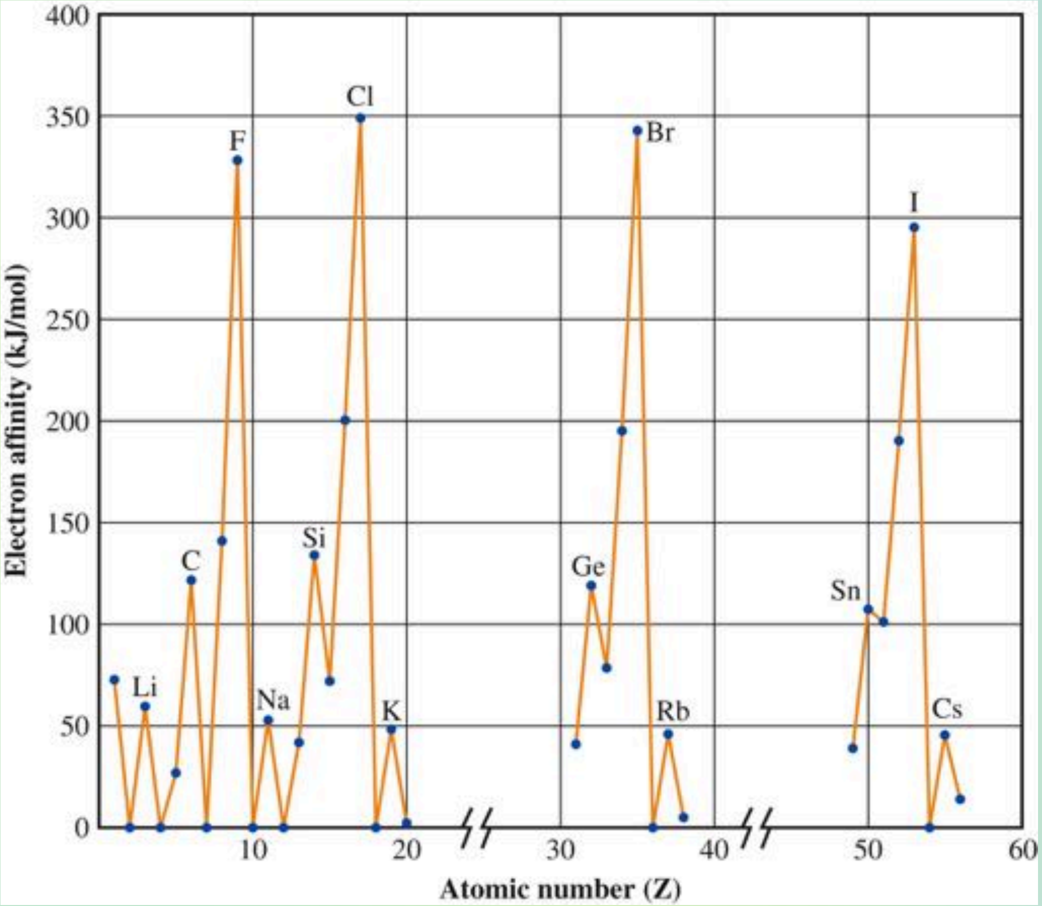
noble gas configuration above lines

*Ionization energies to the right of a vertical line correspond to removal of electrons from the core of the atom.

Electron Affinity (E.A.)

- Energy required to remove an electron from the atom's negative ion.

- It can also be defined as the negative energy obtained when the neutral atom picks up an electron. When a stable negative ion forms, the quantity is positive.



Electron affinities in the main-group elements show a periodic variation when plotted against atomic number, although this variation is somewhat more complicated than that displayed by ionization energies.

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.

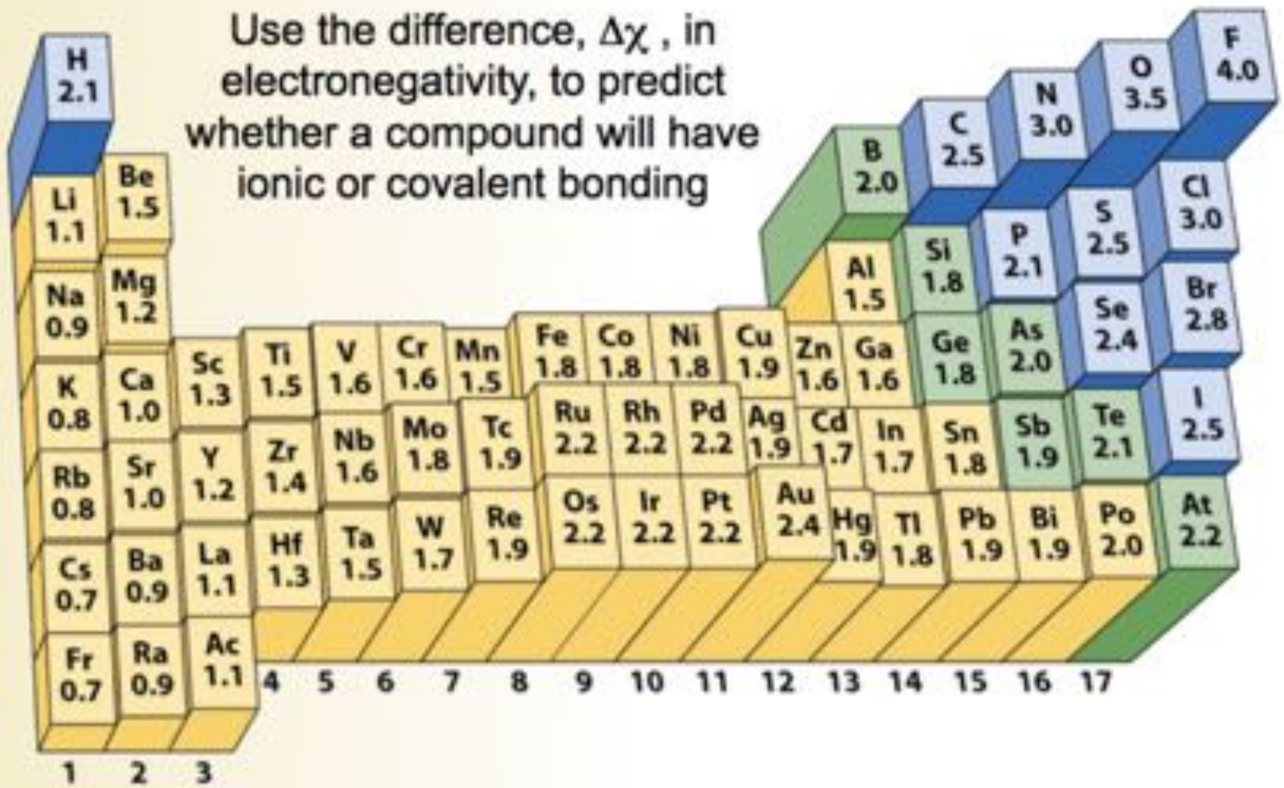
- All Group 1A elements have moderately positive electron affinities, greater than 2A.
- With the exception of the Group 5A element, the electron affinity tends to rise from the Group 2A element to the Group 7A element.
- No known stable negative ions of the Group 8A elements

Table 8.4 Electron Affinities of the Main-Group Elements (kJ/mol)

Period	1A	2A	3A	4A	5A	6A	7A	8A
1	H 73							He ≤0
2	Li 60	Be ≤0	B 27	C 122	N ≤0	O 141	F 328	Ne ≤0
3	Na 53	Mg ≤0	Al 44	Si 134	P 72	S 200	Cl 349	Ar ≤0
4	K 48	Ca 2	Ga 41	Ge 119	As 78	Se 195	Br 325	Kr ≤0
5	Rb 47	Sr 5	In 37	Sn 107	Sb 101	Te 190	I 295	Xe ≤0
6	Cs 46	Ba 14	Tl 36	Pb 35	Bi 91	Po 180	At 270	Rn ≤0

Electronegativity

- The concept of electronegativity sets guidelines for the types of bonds that form between two different elements.



Chemistry: The Science in Context 2/e Figure 8.5
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- ionic bonds between metals and non-metals
- covalent bonds between non-metals