CHEMISTRY
A Molecular Approach

Fourth Edition

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Lecture 04

tuần hoàn **Periodic Properties of the Elements**

bảng tuần hoàn

Mendeleev (1834–1907)

- Ordered elements by atomic mass
- Saw a repeating pattern of properties
- **Periodic law**: When the elements are arranged in order of increasing atomic mass, certain sets of properties recur periodically.
- Put elements with similar properties in the same column
- Used pattern to predict properties of undiscovered elements
- Where atomic mass order did not fit other properties, he reordered by other properties.
  - Te and I
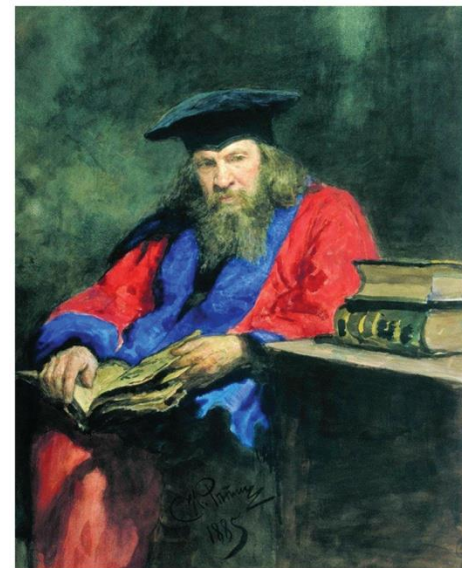
Mendeleev's Predictions

Gallium (eka-aluminum)

	Mendeleev's predicted properties	Actual properties
Atomic mass	About 68 amu	69.72 amu
Melting point	Low	29.8 °C
Density	5.9 g/cm^3	5.90 g/cm^3
Formula of oxide	X$_2$O$_3$	Ga$_2$O$_3$
Formula of chloride	XCl$_3$	GaCl$_3$

Germanium (eka-silicon)

	Mendeleev's predicted properties	Actual properties
Atomic mass	About 72 amu	72.64 amu
Density	5.5 g/cm^3	5.35 g/cm^3
Formula of oxide	XO$_2$	GeO$_2$
Formula of chloride	XCl$_4$	GeCl$_4$

What versus Why

- Mendeleev's periodic law allows us to predict *what* the properties of an element will be based on its position on the table.

- It doesn't explain why the pattern exists.

- Quantum mechanics is a theory that explains *why* the periodic trends in the properties exist.
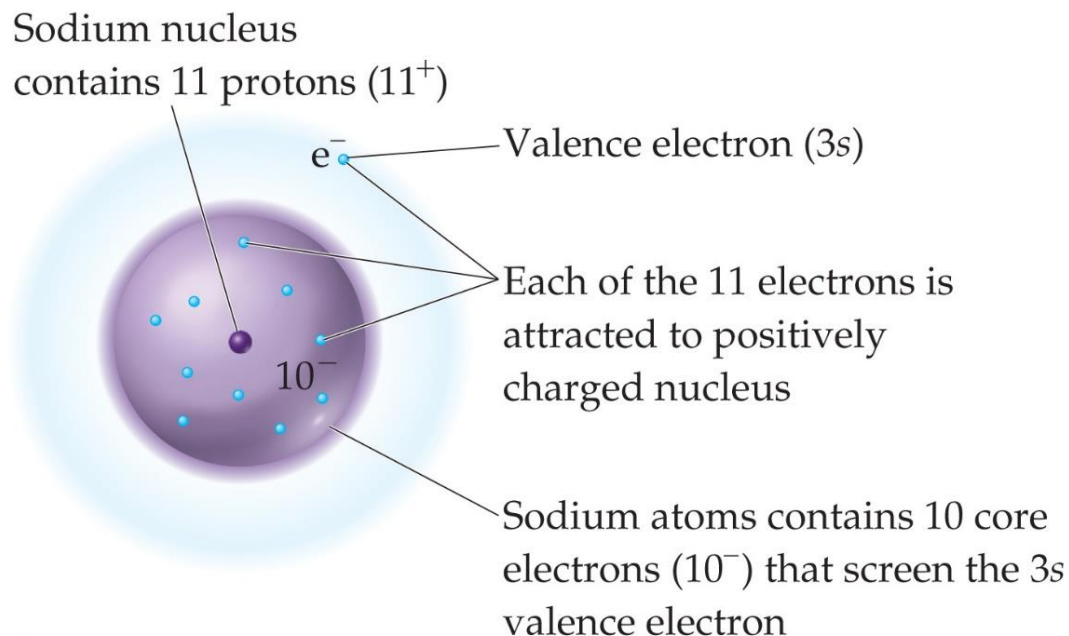 - Knowing *why* allows us to predict *what*.

Coulomb's Law

$$E = \frac{1}{4\pi\varepsilon_0} \frac{q_1 q_2}{r}$$

- Coulomb's law describes the attractions and repulsions between charged particles.

- For like charges, the potential energy (E) is positive and decreases as the particles get farther apart as r increases.

- For opposite charges, the potential energy is negative and becomes more negative as the particles get closer together.

- The strength of the interaction increases as the size of the charges increases.

  - Electrons are more strongly attracted to a nucleus with a 2+ charge than a nucleus with a 1+ charge.

Shielding and Effective Nuclear Charge

- Each electron in a multielectron atom experiences both the attraction to the nucleus and the repulsion by other electrons in the atom.

- These repulsions cause the electron to have a net reduced attraction to the nucleus; it is **shielded** from the nucleus. *(lực đẩy)* *(chắn)*

- The total amount of attraction that an electron feels for the nucleus is called the **effective nuclear charge** of the electron.

Shielding and Effective Nuclear Charge

Sodium nucleus contains 11 protons (11$^+$)

e^- — Valence electron (3s)

Each of the 11 electrons is attracted to positively charged nucleus

10$^-$

Sodium atoms contains 10 core electrons (10$^-$) that screen the 3s valence electron

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The effective nuclear charge, Z_{eff}, is found this way:

$$Z_{eff} = Z - S$$

where Z is the atomic number and S is a screening constant, usually close to the number of inner electrons.

S: calculated from the Slater's rules (next slides)

Slater's Rules

1) Write the electron configuration for the atom using the following design: $(1s)(2s,2p)(3s,3p)\,(3d)\,(4s,4p)\,(4d)\,(4f)\,(5s,5p)$

2) Any electrons to the right of the electron of interest contributes no shielding.

3) All other electrons in the same group as the electron of interest shield to an extent of 0.35 nuclear charge units

4) **If the electron of interest is an s or p electron**: All electrons with one less value of the principal quantum number (n-1 shell) shield to an extent of 0.85 units of nuclear charge. All electrons with two less values of the principal quantum number (n-2 shell) shield to an extent of 1.00 units.

5) **If the electron of interest is an d or f electron**: All electrons to the left shield to an extent of 1.00 units of nuclear charge.

6) Sum the shielding amounts from steps 2 through 5 and subtract from the nuclear charge value to obtain the effective nuclear charge.

Shielding and Effective Nuclear Charge

Slater's Rules: Calculation of shielding value (S)
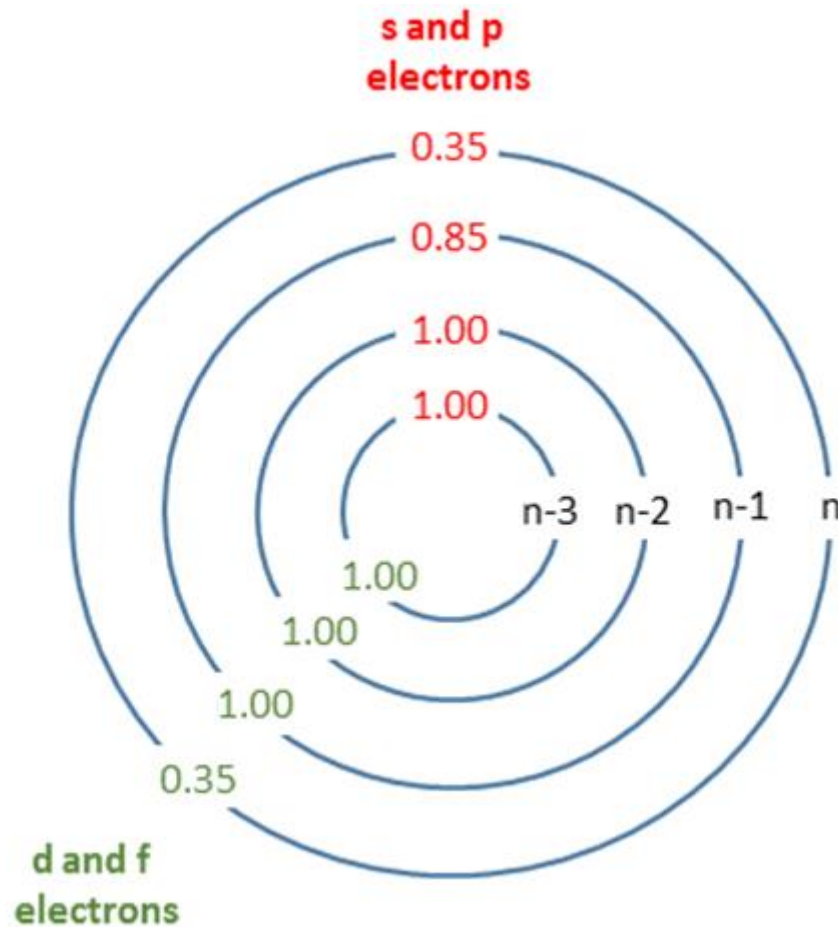
Shielding and Effective Nuclear Charge

Slater's Rules: Calculation of shielding value (S)

Element: N

Configuration: $1s^2 2s^2 2p^3$

n-value grouping: $(1s)^2(2s2p)^5$

Groupings via Slater: $(1s)^2(2s2p)^{5-1=4}$

Slater's shielding value: $2 \times 0.85 + 4 \times 0.35 = 3.1$

Z_{eff}: $7 - 3.1 = 3.9$

Element: Zn

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

n-value grouping: $(1s)^2(2s2p)^8(3s3p)^8(3d)^{10}4s^2$

Groupings via Slater: $(1s)^2(2s2p)^8(3s3p)^8(3d)^{10}4s^{2-1=1}$

Slater's shielding value: $2 \times 1 + 8 \times 1 + 18 \times 0.85 + 1 \times 0.35 = 25.65$

Z_{eff}: $30 - 25.65 = 4.35$

Penetration

- The closer an electron is to the nucleus, the more attraction it experiences.

- The better an outer electron is at **penetrating** through the electron cloud of inner electrons, the more attraction it will have for the nucleus.

- The degree of penetration is related to the orbital's radial distribution function.

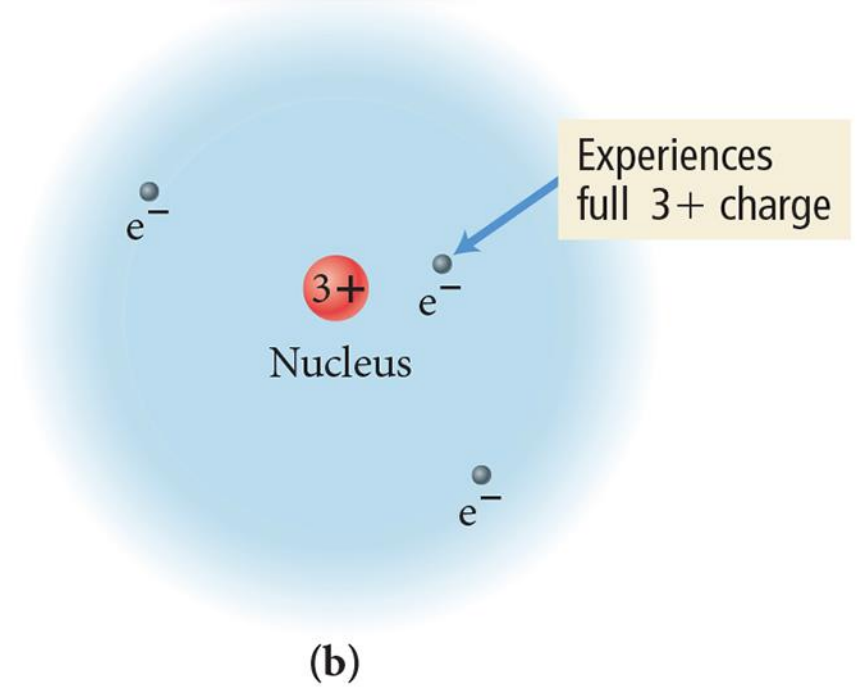 – In particular, the distance the maxima of the function are from the nucleus
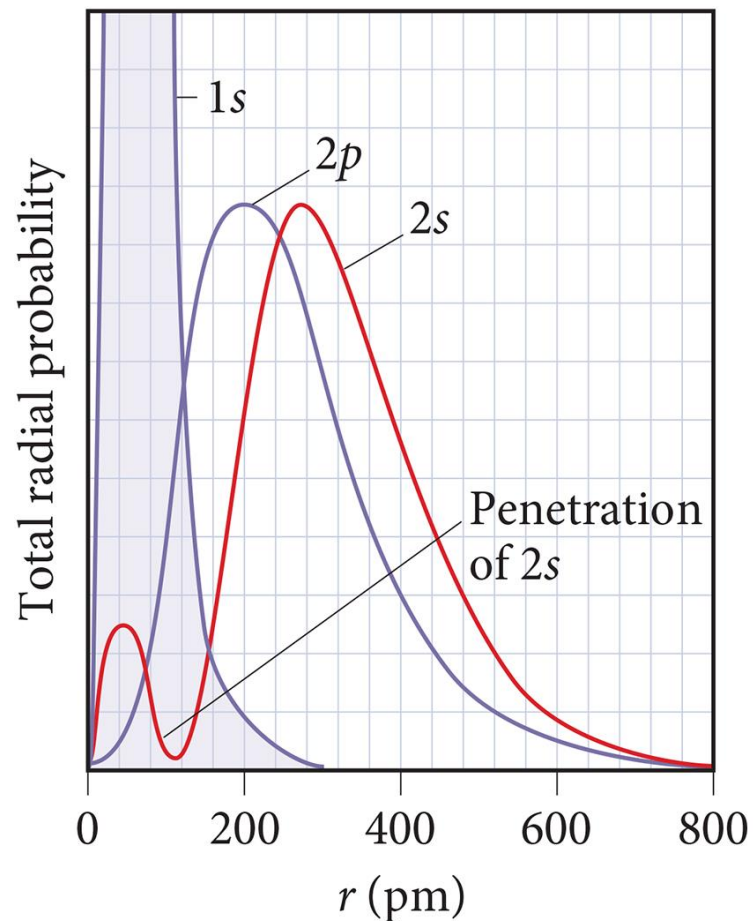
Shielding and Penetration

Shielding
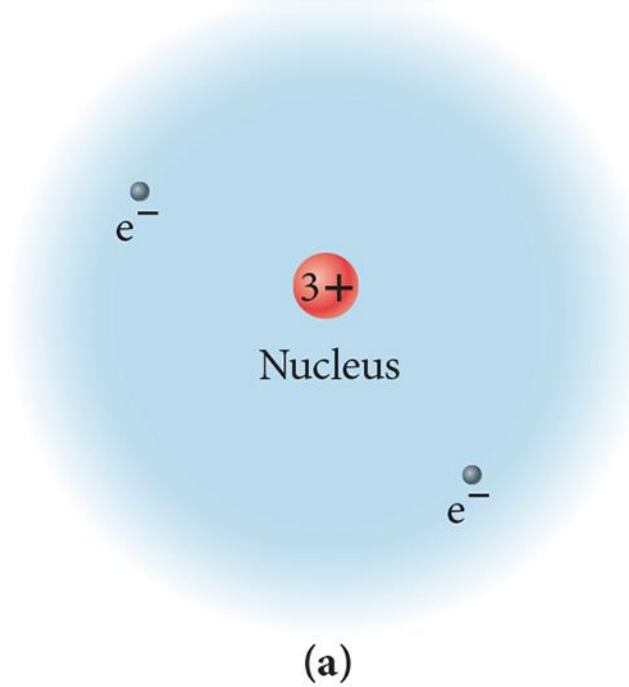
Penetration

Experiences net charge of about 1+

Experiences full 3+ charge

e^-

$3+$

Nucleus

e^-

(a)

e^-

$3+$

e^-

Nucleus

e^-

(b)

Penetration and Shielding

- The radial distribution function shows that the 2*s* orbital penetrates more deeply into the 1*s* orbital than does the 2*p.*

- The weaker penetration of the 2*p* sublevel means that electrons in the 2*p* sublevel experience more repulsive force; they are more shielded from the attractive force of the nucleus.

- The deeper penetration of the 2*s* electrons means electrons in the 2*s* sublevel experience a greater attractive force to the nucleus and are not shielded as effectively.

Effect of Penetration and Shielding

- Penetration causes the energies of sublevels in the same principal level to not be degenerate.

- In the fourth and fifth principal levels, the effects of penetration become so important that the *s* orbital lies lower in energy than the *d* orbitals of the previous principal level.

- The energy separations between one set of orbitals and the next become smaller beyond the 4*s*.

  - The ordering can therefore vary among elements, causing variations in the electron configurations of the transition metals and their ions.

General Energy Ordering of Orbitals for Multielectron Atoms

Energy

5s

4s

4p   4p   4p

3s

3p   3p   3p

2s

2p   2p   2p

1s

4d   4d   4d   4d   4d

3d   3d   3d   3d   3d
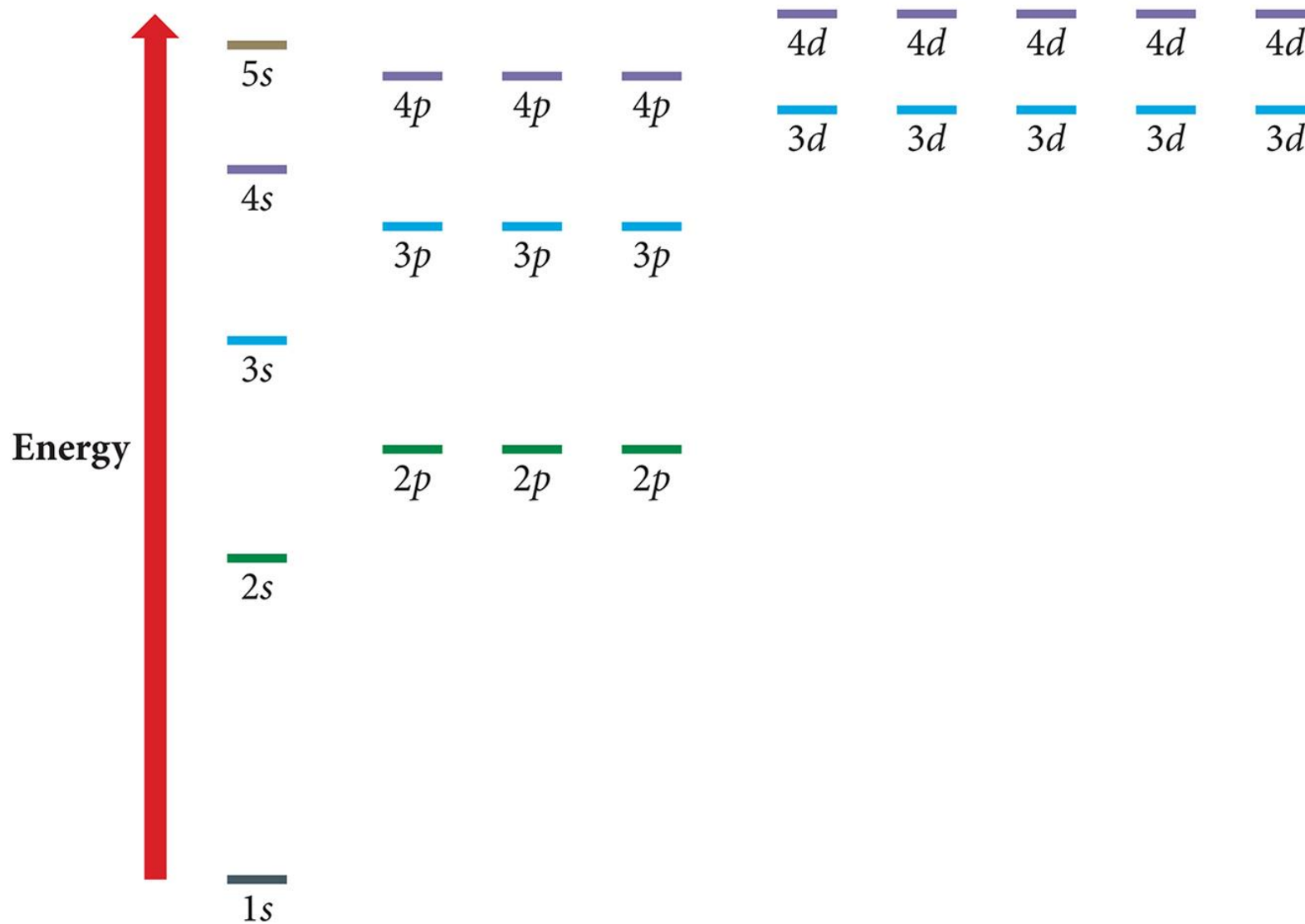
Valence Electrons

- The electrons in all the sublevels with the highest principal energy shell are called the **valence electrons**.

- Electrons in lower energy shells are called **core electrons**.

- One of the most important factors in the way an atom behaves, both chemically and physically, is the number of valence electrons.

Outer Electron Configurations of Elements 1–18

1A							8A
1 **H** $1s^1$	2A	3A	4A	5A	6A	7A	2 **He** $1s^2$
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** $2s^2 2p^6$
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** $3s^2 3p^6$

Electron Configuration and the Periodic Table

- The group number corresponds to the number of valence electrons.

- The length of each "block" is the maximum number of electrons the sublevel can hold.

- The period number corresponds to the principal energy level of the valence electrons.

Orbital Blocks of the Periodic Table

Legend: s-block elements · p-block elements · d-block elements · f-block elements

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

Lanthanides

58 **Ce** $6s^24f^15d^1$	59 **Pr** $6s^24f^3$	60 **Nd** $6s^24f^4$	61 **Pm** $6s^24f^5$	62 **Sm** $6s^24f^6$	63 **Eu** $6s^24f^7$	64 **Gd** $6s^24f^75d^1$	65 **Tb** $6s^24f^9$	66 **Dy** $6s^24f^{10}$	67 **Ho** $6s^24f^{11}$	68 **Er** $6s^24f^{12}$	69 **Tm** $6s^24f^{13}$	70 **Yb** $6s^24f^{14}$	71 **Lu** $6s^24f^{14}5d^1$

Actinides

90 **Th** $7s^26d^2$	91 **Pa** $7s^25f^26d^1$	92 **U** $7s^25f^36d^1$	93 **Np** $7s^25f^46d^1$	94 **Pu** $7s^25f^6$	95 **Am** $7s^25f^7$	96 **Cm** $7s^25f^76d^1$	97 **Bk** $7s^25f^9$	98 **Cf** $7s^25f^{10}$	99 **Es** $7s^25f^{11}$	100 **Fm** $7s^25f^{12}$	101 **Md** $7s^25f^{13}$	102 **No** $7s^25f^{14}$	103 **Lr** $7s^25f^{14}6d^1$

Inner electron configuration

8A

1A

1 2A

2

3 $3s^2$ 3A 4A 5A 6A 7A Ne

4 s $3p^5$ Cl

5   d p

6

7

f
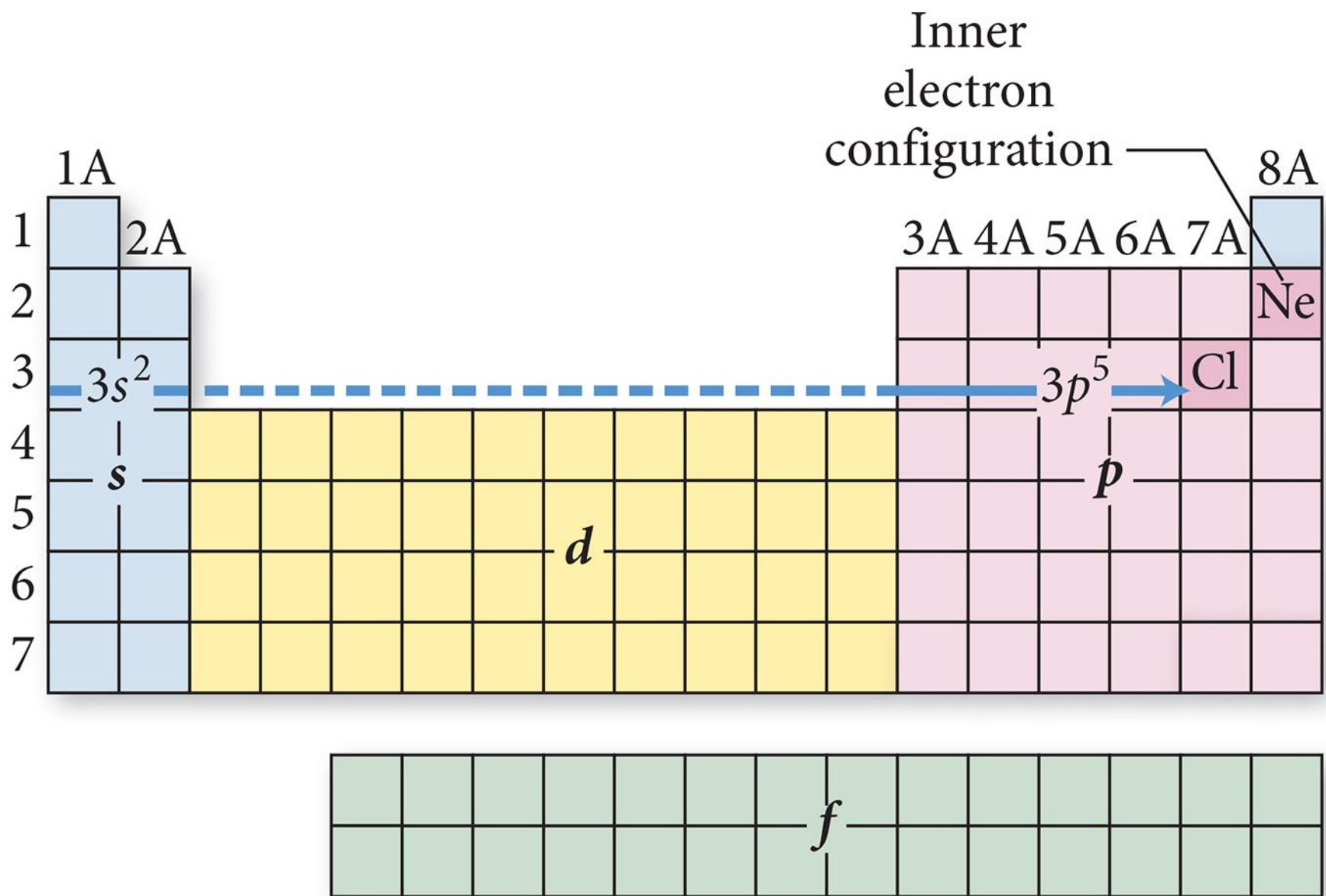
Irregular Electron Configurations

- We know that, because of sublevel splitting, the 4s sublevel is lower in energy than the 3d; therefore, the 4s fills before the 3d.

- But the difference in energy is not large.

- Some of the transition metals have irregular electron configurations in which the ns only partially fills before the $(n-1)d$ or doesn't fill at all.

- Therefore, their electron configuration must be found experimentally.

Irregular Electron Configurations

- Expected
- Cr = [Ar]$4s^2 3d^4$
- Cu = [Ar]$4s^2 3d^9$
- Mo = [Kr]$5s^2 4d^4$
- Ru = [Kr]$5s^2 4d^6$
- Pd = [Kr]$5s^2 4d^8$

- Found experimentally
- Cr = [Ar]$4s^1 3d^5$
- Cu = [Ar]$4s^1 3d^{10}$
- Mo = [Kr]$5s^1 4d^5$
- Ru = [Kr]$5s^1 4d^7$
- Pd = [Kr]$5s^0 4d^{10}$

Properties and Electron Configuration

- The properties of the elements follow a periodic pattern.
  - Elements in the same column have similar properties.
  - The elements in a period show a pattern that repeats.

- The quantum-mechanical model explains this because the number of valence electrons and the types of orbitals they occupy are also periodic.

8A	
2 **He** $1s^2$	
10 **Ne** $2s^2 2p^6$	
18 **Ar** $3s^2 3p^6$	
36 **Kr** $4s^2 4p^6$	
54 **Xe** $5s^2 5p^6$	
86 **Rn** $6s^2 6p^6$	

Noble gases

1A	
3 **Li** $2s^1$	
11 **Na** $3s^1$	
19 **K** $4s^1$	
37 **Rb** $5s^1$	
55 **Cs** $6s^1$	
87 **Fr** $7s^1$	

Alkali metals

2A	
4 **Be** $2s^2$	
12 **Mg** $3s^2$	
20 **Ca** $4s^2$	
38 **Sr** $5s^2$	
56 **Ba** $6s^2$	
88 **Ra** $7s^2$	

Alkaline earth metals

7A	
9 **F** $2s^2 2p^5$	
17 **Cl** $3s^2 3p^5$	
35 **Br** $4s^2 4p^5$	
53 **I** $5s^2 5p^5$	
85 **At** $6s^2 6p^5$	

Halogens

The Noble Gas Electron Configuration

- The noble gases have eight valence electrons.
  - Except for He, which has only two electrons

- They are especially nonreactive.
  - He and Ne are practically inert.

- The reason the noble gases are so nonreactive is that the electron configuration of the noble gases is especially stable.

8A
2 **He** $1s^2$
10 **Ne** $2s^2 2p^6$
18 **Ar** $3s^2 3p^6$
36 **Kr** $4s^2 4p^6$
54 **Xe** $5s^2 5p^6$
86 **Rn** $6s^2 6p^6$

Noble gases

khí hiếm

The Alkali Metals

- The **alkali** metals have one more electron than the previous noble gas.

- In their reactions, the alkali metals tend to lose one electron, resulting in the same electron configuration as a noble gas.
  - Forming a cation with a 1+ charge

1A

| 3 Li $2s^1$ |
| 11 Na $3s^1$ |
| 19 K $4s^1$ |
| 37 Rb $5s^1$ |
| 55 Cs $6s^1$ |
| 87 Fr $7s^1$ |

Alkali metals

The Halogens

- Have one fewer electron than the next noble gas

- In their reactions with metals, the halogens tend to gain an electron and attain the electron configuration of the next noble gas, forming an anion with charge 1−.

- In their reactions with nonmetals, they tend to share electrons with the other nonmetal so that each attains the electron configuration of a noble gas.

7A

9   **F** $2s^2 2p^5$
17   **Cl** $3s^2 3p^5$
35   **Br** $4s^2 4p^5$
53   **I** $5s^2 5p^5$
85   **At** $6s^2 6p^5$

Halogens

Electron Configuration and Ion Charge

- We have seen that many metals and nonmetals form one ion and that the charge on that ion is predictable based on its position on the periodic table.

  - Group 1A = 1+, group 2A = 2+, group 7A = 1−, group 6A = 2−, etc.

- These atoms form ions that will result in an electron configuration that is the same as the nearest noble gas.

Elements That Form Ions with Predictable Charges

	1A	2A	3B	4B	5B	6B	7B	⌐— 8B —⌐			1B	2B	3A	4A	5A	6A	7A	8A
1	Li^+															N^{3-}	O^{2-}	F^-	
2	Na^+	Mg^{2+}												Al^{3+}			S^{2-}	Cl^-	
3	K^+	Ca^{2+}															Se^{2-}	Br^-	
4	Rb^+	Sr^{2+}															Te^{2-}	I^-	
5	Cs^+	Ba^{2+}																	

Electron Configuration of Anions in Their Ground State

- Anions are formed when nonmetal atoms gain enough electrons to have eight valence electrons.
  - Filling the *s* and *p* sublevels of the valence shell

- The sulfur atom has six valence electrons.

  S atom = $1s^2 2s^2 2p^6 3s^2 3p^4$

- To have eight valence electrons, sulfur must gain two more.

  S^{2-} anion = $1s^2 2s^2 2p^6 3s^2 3p^6$

Electron Configuration of Cations in Their Ground State

- Cations are formed when a metal atom loses all its valence electrons, resulting in a new lower energy level valence shell.

  - However, the process is always endothermic.

- The magnesium atom has two valence electrons.

  Mg atom = $1s^2 2s^2 2p^6 3s^2$

- When magnesium forms a cation, it loses its valence electrons.

  Mg^{2+} cation = $1s^2 2s^2 2p^6$

Trend in Atomic Radius: Main Group

- There are several methods for measuring the radius of an atom, and they give slightly different numbers.
  - Van der Waals radius = nonbonding
  - Covalent radius = bonding radius
  - Atomic radius is an average radius of an atom based on measuring large numbers of elements and compounds.

van der Waals radius

2 × Krypton radius

Krypton solid

Covalent radius

Br   Br

228 pm

$$Br\ radius = \frac{228\ pm}{2} = 114\ pm$$
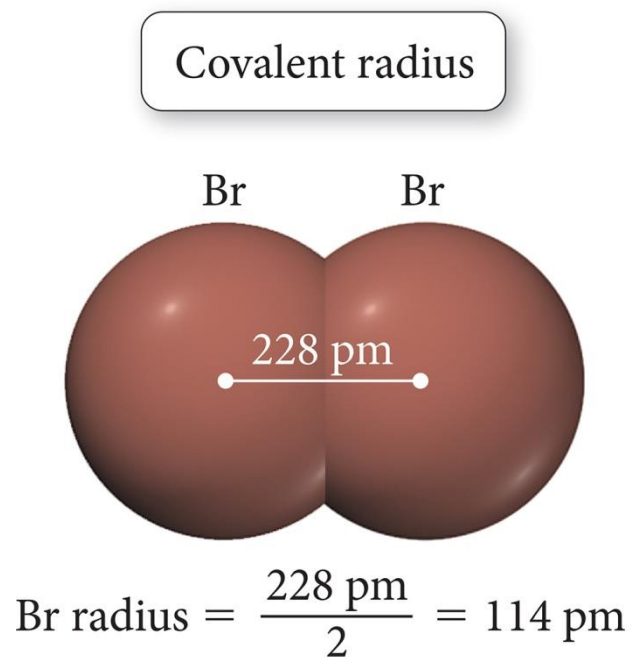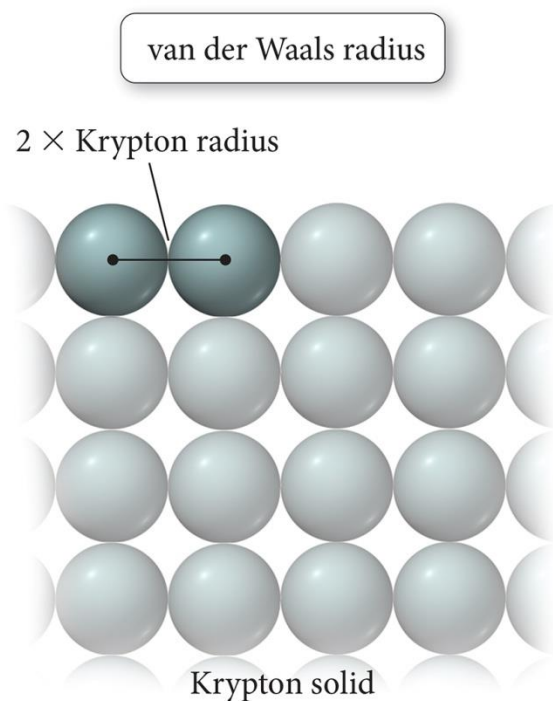
Trend in Atomic Radius: Main Group

- Atomic radius decreases across period (left to right).
  - Adding electrons to same valence shell
  - Effective nuclear charge increases
  - Valence shell held closer

Atomic Radii
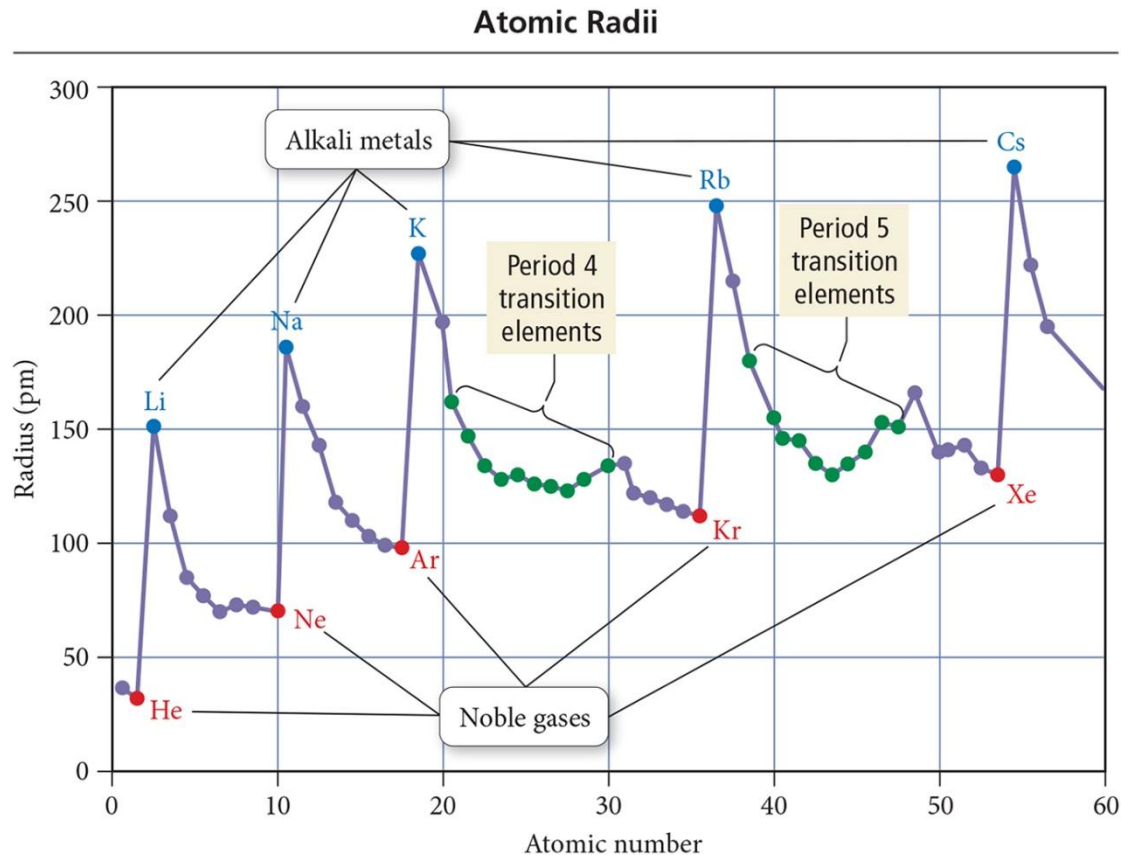
Shielding and Effective Nuclear Charge

Electron config.	Z_{eff} (Slater's rules)
Li $(1s)^2(2s)^1$	$3 - [2(0.85)] = 1.30$
Be $(1s^2)(2s)^{2-1=1}$	$4 - [2(0.85) + 1(0.35)] = 1.95$
B $(1s)^2(2s2p)^{3-1=2}$	$5 - [2(0.85) + 2(0.35] = 2.60$
C $(1s)^2(2s2p)^{4-1=3}$	$6 - [2(0.85) + 3(0.35] = 3.25$
N $(1s)^2(2s2p)^{5-1=4}$	$7 - [2(0.85) + 4(0.35)] = 3.90$
O $(1s)^2(2s2p)^{6-1=5}$	$8 - [2(0.85) + 5(0.35)] = 4.55$
F $(1s)^2(2s2p)^{7-1=6}$	$9 - [2(0.85) + 6(0.35)] = 5.20$
Ne $(1s)^2(2s2p)^{8-1=7}$	$10 - [2(0.85) + 7(0.35)] = 5.85$

Shielding and Effective Nuclear Charge

$$Z_{eff} = Z - S$$

Periodic Trends in Atomic Radius

	1 IA	2 IIA		13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA
	Li 0.152	Be 0.111	Transition elements show many exceptions to the general trends.	B 0.080	C 0.077	N 0.070	O 0.066	F 0.064
	Na 0.186	Mg 0.160		Al 0.143	Si 0.118	P 0.111	S 0.104	Cl 0.099
	K 0.227	Ca 0.197		Ga 0.122	Ge 0.123	As 0.125	Se 0.116	Br 0.115
	Rb 0.248	Sr 0.215		In 0.163	Sn 0.141	Sb 0.145	Te 0.143	I 0.133
	Cs 0.266	Ba 0.217		Tl 0.170	Pb 0.175	Bi 0.155		

DECREASE

Trend in Atomic Radius: Main Group

- Atomic radius increases down group.
  - Valence shell farther from nucleus
  - Effective nuclear charge fairly close

Atomic Radii

Shielding and Effective Nuclear Charge

Electron config.   **Z_{eff} (Slater's rules)**

1
IA

Li   Li $(1s)^2(2s)^1$ $3 - [2(0.85)] = 1.30$

0.152

Na   Na $(1s)^2(2s2p)^8(3s)^1$ $11 - [2(1)+8(0.85)] = 2.2$

0.186

K   K $(1s)^2(2s2p)^8(3s3p)^8(4s)^1$ $19 - [2(1)+8(1)+8(0.85] = 2.2$

0.227

Rb   Rb $(1s)^2(2s2p)^8(3s3p3d)^{18}(4s4p)^8(5s)^1$ $37 - [2(1)+8(1)+18(1)+8(0.85)] = 2.2$

0.248

Cs   Cs $(1s^2)(2s^22p^6)(3s^23p^63d^{10})(4s^24p^64d^{10})(5s^25p^6)(6s^1)$

0.266
                                         $55 - [2(1)+8(1)+18(1)+18(1)+8(0.85] = 2.2$

Periodic Trends in Atomic Radius

|	1 IA	2 IIA		13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA
	Li 0.152	Be 0.111	Transition elements show many exceptions to the general trends.	B 0.080	C 0.077	N 0.070	O 0.066	F 0.064
	Na 0.186	Mg 0.160		Al 0.143	Si 0.118	P 0.111	S 0.104	Cl 0.099
	K 0.227	Ca 0.197		Ga 0.122	Ge 0.123	As 0.125	Se 0.116	Br 0.115
	Rb 0.248	Sr 0.215		In 0.163	Sn 0.141	Sb 0.145	Te 0.143	I 0.133
	Cs 0.266	Ba 0.217		Tl 0.170	Pb 0.175	Bi 0.155		
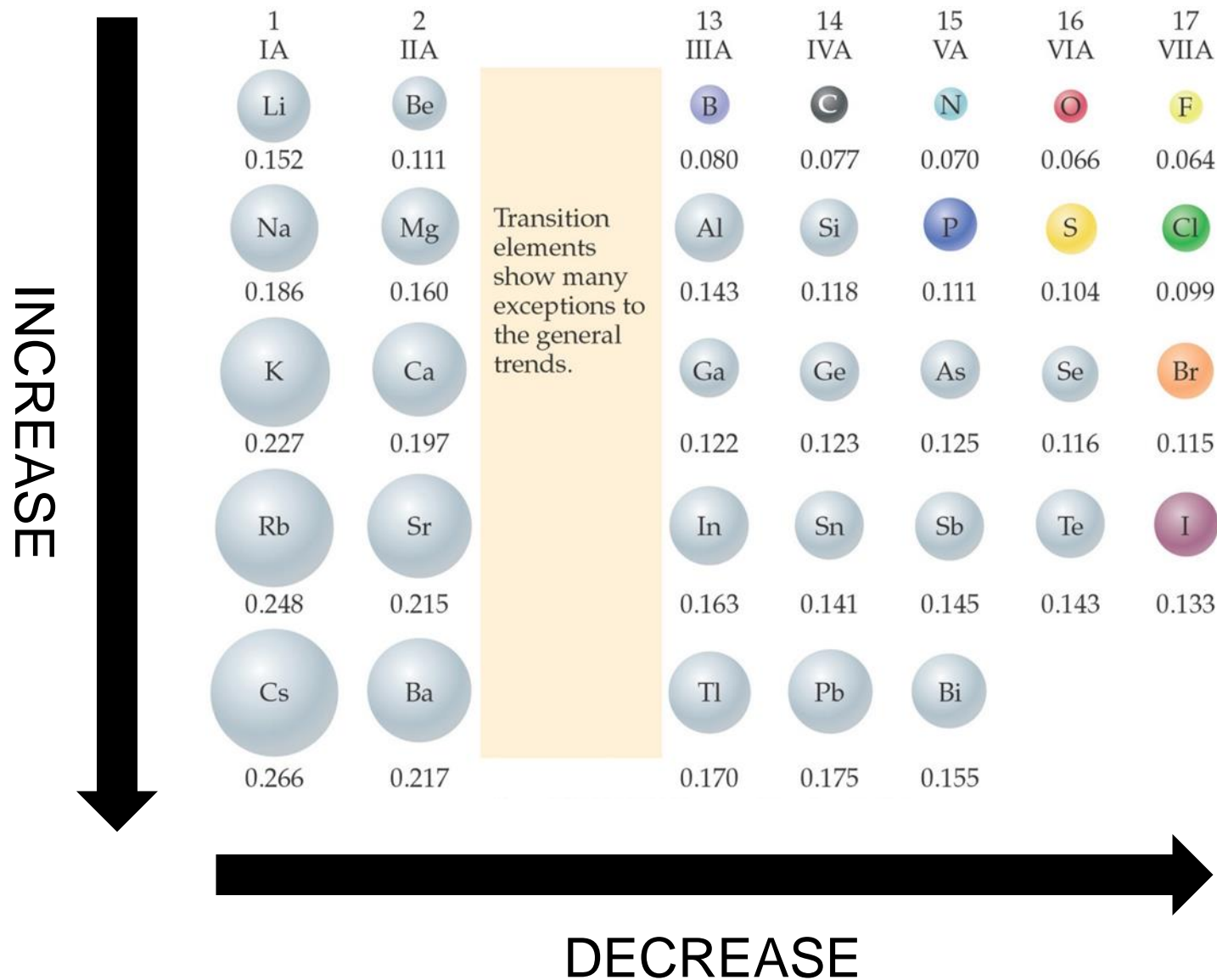
INCREASE

DECREASE

Trends in Ionic Radius

- Ions in the same group have the same charge.
- Ion size increases down the column.
  - Higher valence shell, larger ion
- Cations are smaller than neutral atoms; anions are larger than neutral atoms.
- Cations are smaller than anions.
  - Except Rb^+ and Cs^+, bigger or same size as F^- and O^{2-}
- Larger positive charge = smaller cation
  - For isoelectronic species
  - Isoelectronic = same electron configuration
- Larger negative charge = larger anion
  - For isoelectronic species

Periodic Trends in Ionic Radius

- Ionic size depends upon
  - The nuclear charge.
  - The number of electrons.
  - The orbitals in which electrons reside.
- Cations are smaller than their parent atoms
- Anions are larger than their parent atoms"

Group 1A	Group 2A	Group 3A	Group 6A	Group 7A
Li^+ 0.90	Be^{2+} 0.59	B^{3+} 0.41	O^{2-} 1.26	F^- 1.19
Li 1.34	Be 0.90	B 0.82	O 0.73	F 0.71
Na^+ 1.16	Mg^{2+} 0.86	Al^{3+} 0.68	S^{2-} 1.70	Cl^- 1.67
Na 1.54	Mg 1.30	Al 1.18	S 1.02	Cl 0.99
K^+ 1.52	Ca^{2+} 1.14	Ga^{3+} 0.76	Se^{2-} 1.84	Br^- 1.82
K 1.96	Ca 1.24	Ga 1.26	Se 1.16	Br 1.14
Rb^+ 1.66	Sr^{2+} 1.32	In^{3+} 0.94	Te^{2-} 2.07	I^- 2.06
Rb 2.11	Sr 1.92	In 1.44	Te 1.35	I 1.33

= cation   = anion   = neutral atom

Periodic Trends in Ionic Radius

- In an **isoelectronic series**, ions have the same number of electrons.

- Ionic size decreases with an increasing nuclear charge.

Explanation for the Trends in Anion Radius

- When atoms form anions, electrons are added to the valence shell.

- These "new valence electrons" experience a smaller effective nuclear charge than the "old valence electrons," increasing the size.

- The result is that the anion is larger than the atom.

Explanation for the Trends in Anion Radius

- Traversing down a group increases the *n* level, causing the anions to get larger.

- Traversing to the right across a period decreases the effective nuclear charge for isoelectronic anions, causing the anions to get larger.

Ionization Energy (IE)

NANG LUONG ION hoa

- Minimum energy needed to remove an electron from an atom or ion
  - Gas state
  - Endothermic process
  - Valence electron easiest to remove, lowest IE
  - $M(g) + IE_1 \rightarrow M^{1+}(g) + 1\ e^-$
  - $M^{+1}(g) + IE_2 \rightarrow M^{2+}(g) + 1\ e^-$
    - First ionization energy = energy to remove electron from neutral atom
    - Second IE = energy to remove electron from 1+ ion, etc.

General Trends in First Ionization Energy

- The larger the effective nuclear charge on the electron, the more energy it takes to remove it.
- The farther the most probable distance the electron is from the nucleus, the less energy it takes to remove it.
- First IE *decreases* down the group.
  - Valence electron farther from nucleus
- First IE generally *increases* across the period.
  - Effective nuclear charge increases

First Ionization Energies

Trends in First Ionization Energy

Ionization energy (kJ/mol)

Increasing ionization energy

Decreasing ionization energy

1A	2A		3A	4A	5A	6A	7A	8A
H 1312								He 2372
								Ne 2081
	Be 899		B 801	C 1086	N 1402	O 1314	F 1681	Ar 1521
Li 520	Mg 738		Al 578	Si 786	P 1012	S 1000	Cl 1251	Kr 1351
Na 496	Ca 590		Ga 579	Ge 762	As 947	Se 941	Br 1140	Xe 1170
K 419	Sr 549		In 558	Sn 709	Sb 834	Te 869	I 1008	Rn 1037
Rb 403	Ba 503		Tl 589	Pb 716	Bi 703	Po 812		
Cs 376								
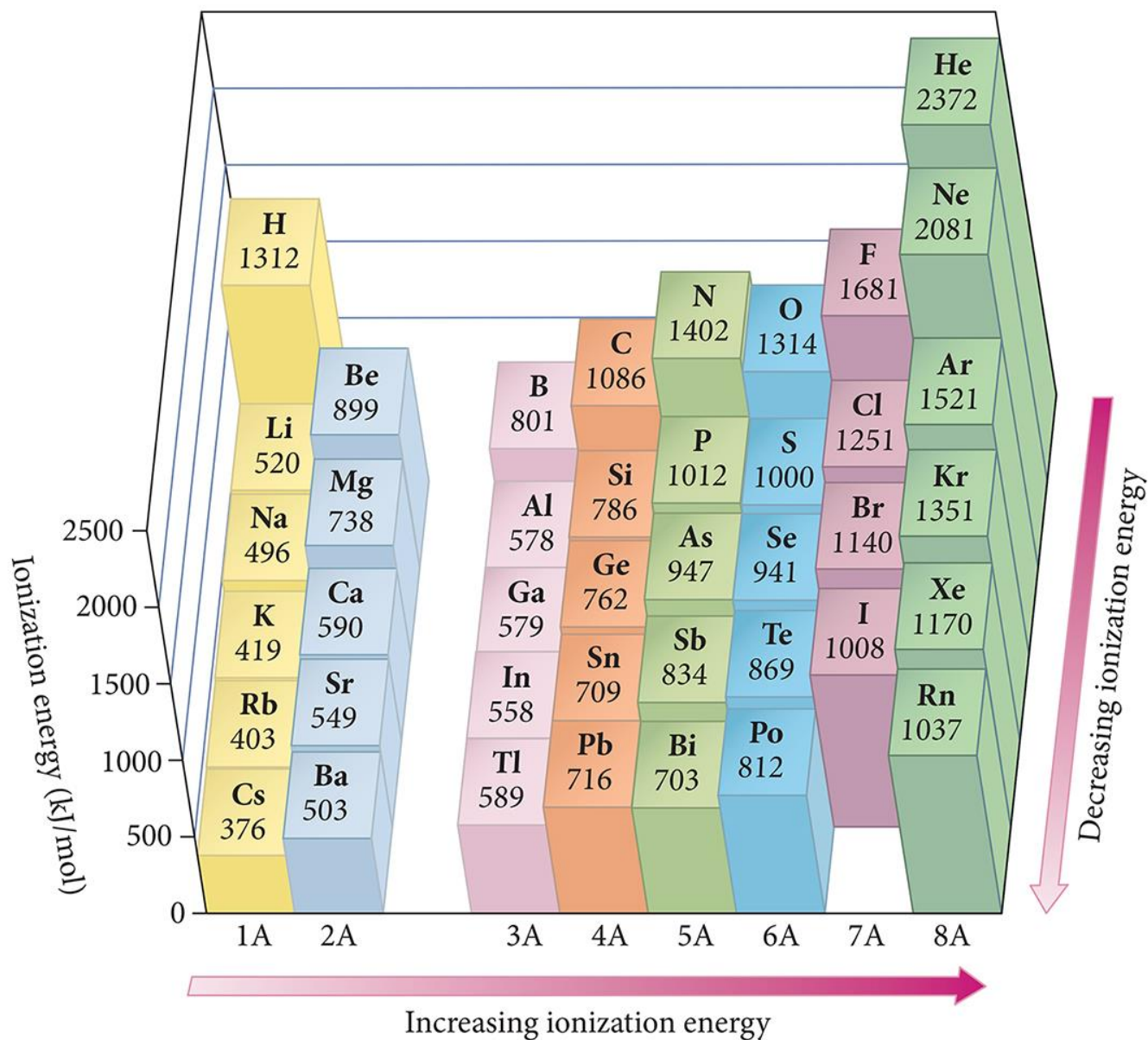
General Trends in First Ionization Energy

- Down a column, less energy is required to remove the first electron.

- In the same group, Z_{eff} is the same, but the valence electrons are farther from the nucleus.
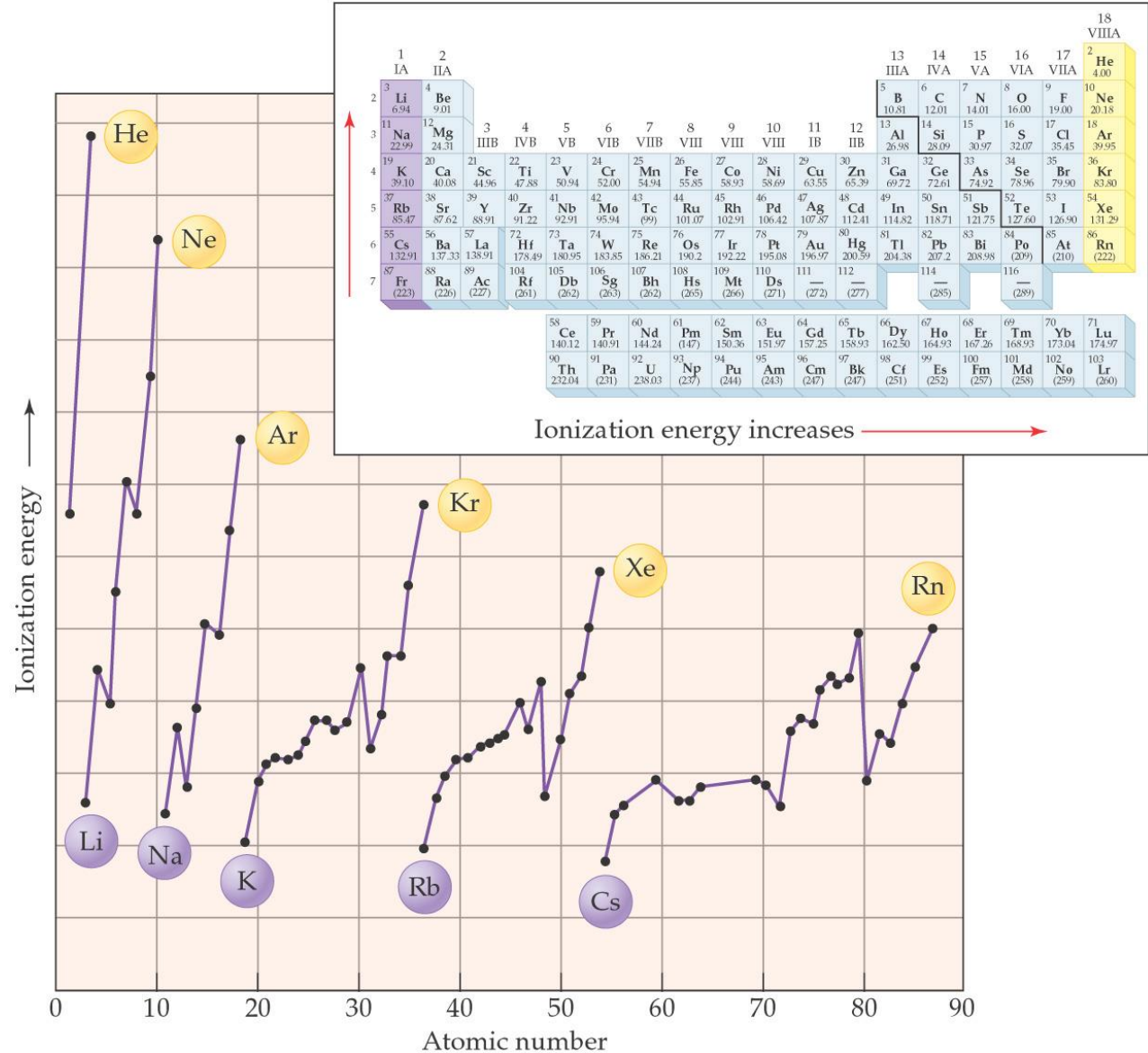
General Trends in First Ionization Energy

- Across a row, it gets harder to remove an electron.

- As you go from left to right, Z_{eff} increases.

- However, there are two apparent discontinuities in this trend.
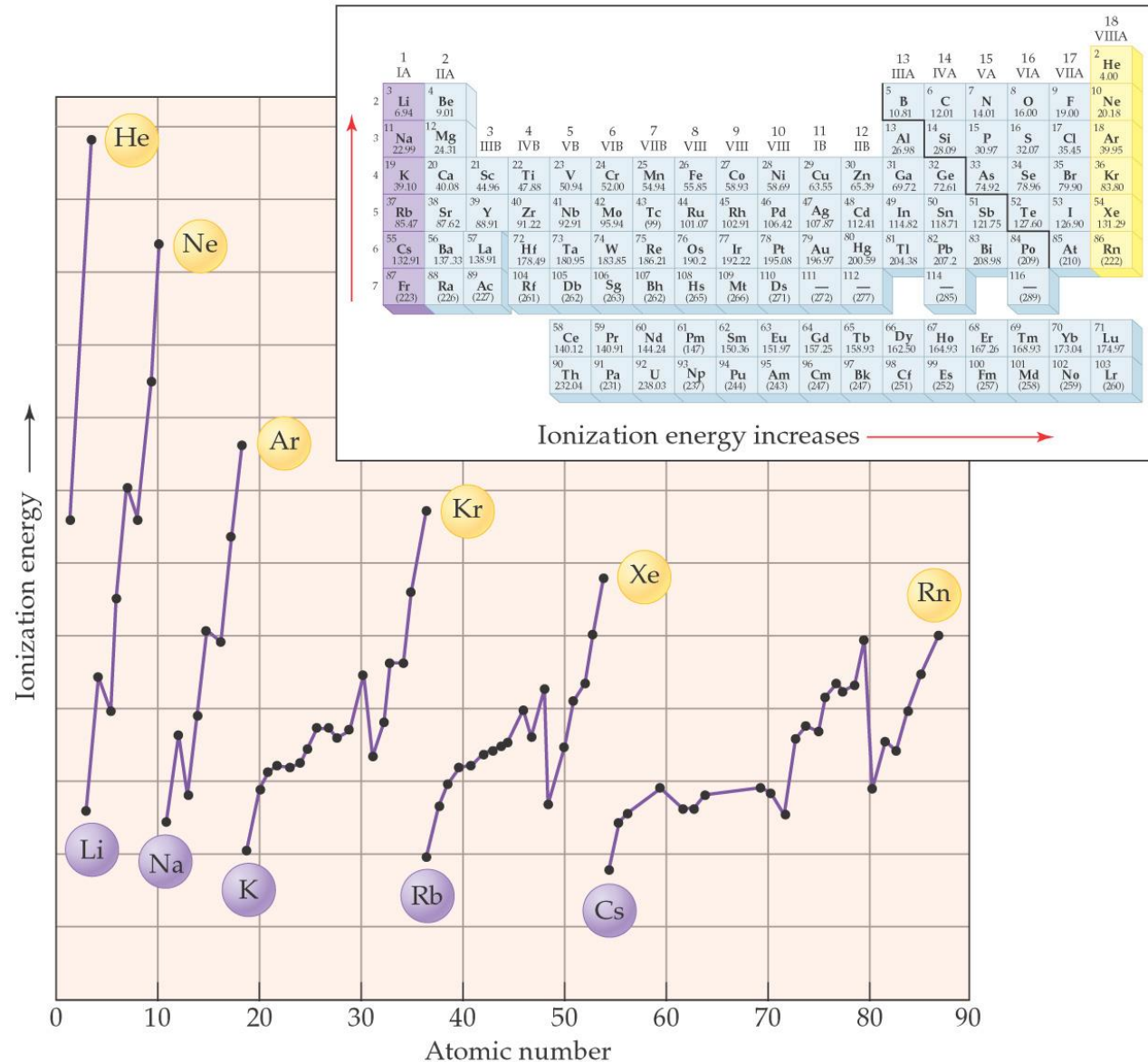
Exceptions in the First IE Trends

- First ionization energy generally increases from left to right across a period.

- Except from 2A to 3A and 5A to 6A

N $1s^2 2s^2 2p^3$ 1s   2s   2p

O $1s^2 2s^2 2p^4$ 1s   2s   2p
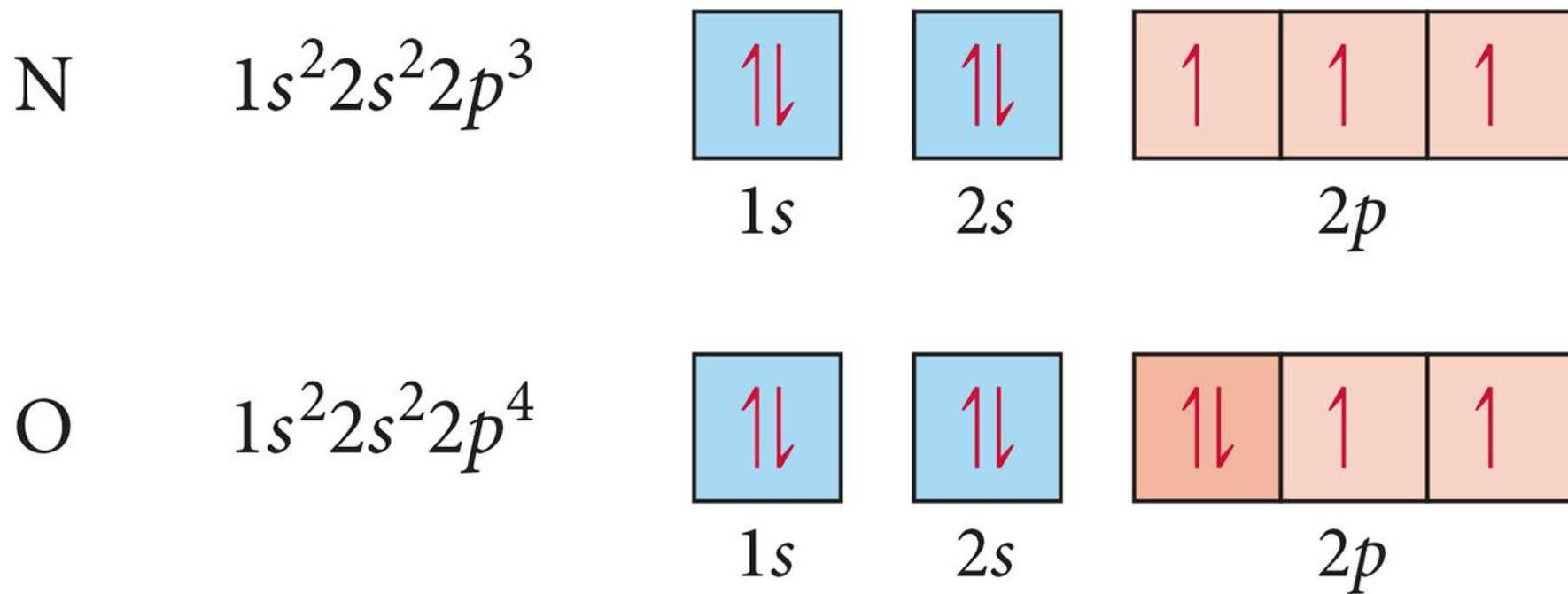
Exceptions in the First Ionization Energy Trends, N and O

N $1s^2 2s^2 2p^3$

1s	2s	2p		
↑↓	↑↓	↑	↑	↑
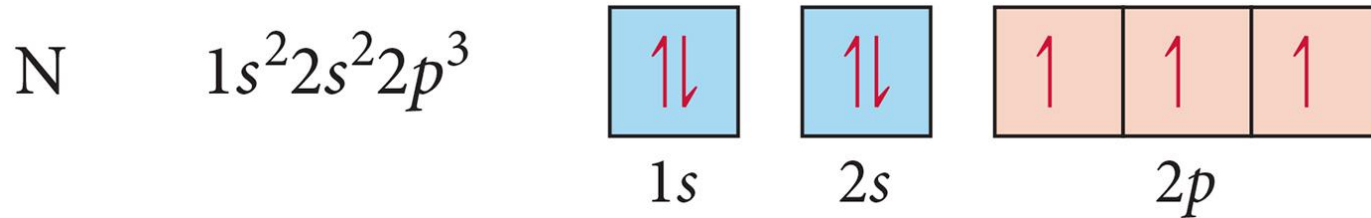
To ionize N, you must break up a half-full sublevel, which costs extra energy.

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

1s	2s	2p		
↑↓	↑↓	↑↓	↑	↑
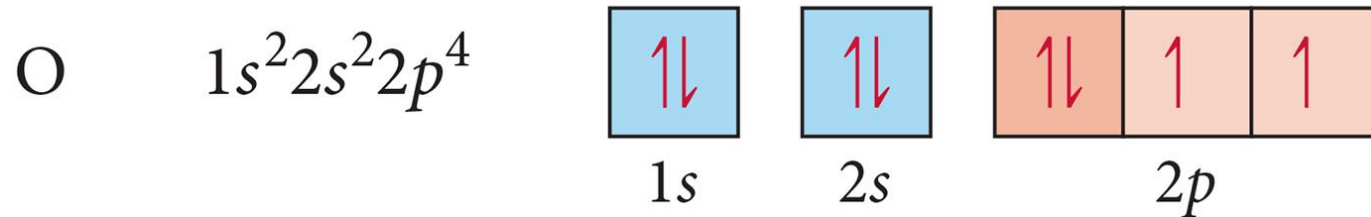
When you ionize O, you get a half-full sublevel, which costs less energy.

Trends in Successive Ionization Energies

- Removal of each successive electron costs more energy.
  - Shrinkage in size due to having more protons than electrons
  - Outer electrons closer to the nucleus; therefore harder to remove
- There's a regular increase in energy for each successive valence electron.
- There's a large increase in energy when core electrons are removed.
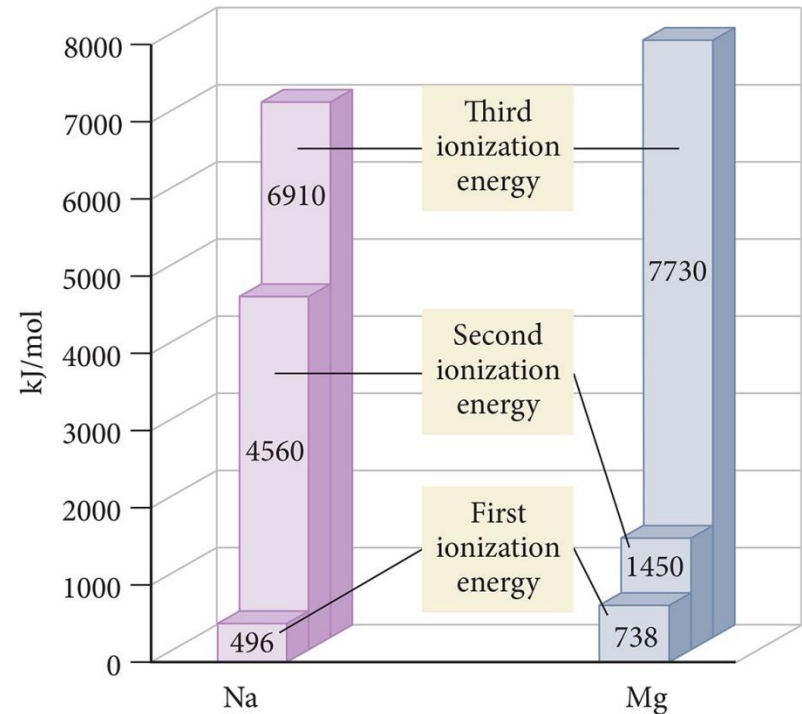
Trends in Second and Successive Ionization Energies

TABLE 8.1 Successive Ionization Energies for the Elements Sodium through Argon (kJ/mol)

Element	IE_1	IE_2	IE_3	IE_4	IE_5	IE_6	IE_7
Na	496	4560					
Mg	738	1450	7730		Core electrons		
Al	578	1820	2750	11,600			
Si	786	1580	3230	4360	16,100		
P	1012	1900	2910	4960	6270	22,200	
S	1000	2250	3360	4560	7010	8500	27,100
Cl	1251	2300	3820	5160	6540	9460	11,000
Ar	1521	2670	3930	5770	7240	8780	12,000

Electron Affinity

- Energy is released when a neutral atom gains an electron.
  - Gas state
  - $M(g) + 1e^- \rightarrow M^{1-}(g) + EA$
- Electron affinity is defined as exothermic (−) but may actually be endothermic (+).
  - Some alkali earth metals and all noble gases are endothermic. Why?
- The more energy that is released, the larger the electron affinity.
  - The more negative the number, the larger the EA.

Trends in Electron Affinity

- Alkali metals decrease electron affinity down the column.
  - But not all groups do
  - Generally irregular increase in EA from second period to third period
- "Generally" increases across period
  - Becomes more negative from left to right
  - Not absolute
  - Group 5A generally lower EA than expected because extra electron must pair
  - Groups 2A and 8A generally very low EA because added electron goes into higher energy level or sublevel
- Highest EA in any period = halogen

Electron Affinities (kJ/mol)

1A	2A		3A	4A	5A	6A	7A	8A
H −73								**He** >0
Li −60	**Be** >0		**B** −27	**C** −122	**N** >0	**O** −141	**F** −328	**Ne** >0
Na −53	**Mg** >0		**Al** −43	**Si** −134	**P** −72	**S** −200	**Cl** −349	**Ar** >0
K −48	**Ca** −2		**Ga** −30	**Ge** −119	**As** −78	**Se** −195	**Br** −325	**Kr** >0
Rb −47	**Sr** −5		**In** −30	**Sn** −107	**Sb** −103	**Te** −190	**I** −295	**Xe** >0

Properties of Metals and Nonmetals

- Metals
  - Malleable and ductile
  - Shiny, lustrous, reflect light
  - Conduct heat and electricity
  - Most oxides basic and ionic
  - Form cations in solution
  - Lose electrons in reactions—**oxidized**
- Nonmetals
  - Brittle in solid state
  - Dull, nonreflective solid surface
  - Electrical and thermal insulators
  - Most oxides are acidic and molecular
  - Form anions and polyatomic anions
  - Gain electrons in reactions—**reduced**

Metallic Character

- **Metallic character** is how closely an element's properties match the ideal properties of a metal.
  - More malleable and ductile, better conductors, and easier to ionize
- Metallic character *decreases left to right* across a period.
  - Metals found at the left of the period, and nonmetals to the right
- Metallic character *increases down* the column.
  - Nonmetals found at the top of the middle main-group elements, and metals found at the bottom

Trends in Metallic Character

Metallic character decreases →

Metallic character increases ↓

Metals | Metalloids | Nonmetals

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

Periods

Lanthanides	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
Actinides	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

Explanation for the Trends in Metallic Character

- Metals generally have smaller first ionization energies, and nonmetals generally have larger electron affinities.
  - Except for the noble gases
- ∴ quantum mechanics predicts the atom's metallic character should increase down a column because the valence electrons are not held as strongly.
- ∴ quantum mechanics predicts the atom's metallic character should decrease across a period because the valence electrons are held more strongly and the electron affinity increases.

Trends in the Alkali Metals

- Atomic radius increases down the column.
- Ionization energy decreases down the column.
- Very low ionization energies
  - Good reducing agents; easy to oxidize
  - Very reactive; not found uncombined in nature
  - React with nonmetals to form salts
  - Compounds generally soluble in water $\therefore$ found in seawater
- Electron affinity decreases down the column.
- Melting point decreases down the column.
  - All very low MP for metals
- Density increases down the column.
  - Except K
  - In general, the increase in mass is greater than the increase in volume.

Alkali Metals

TABLE 8.2 Properties of the Alkali Metals*

Element	Electron Configuration	Atomic Radius (pm)	IE$_1$ (kJ/mol)	Density at 25 °C (g/cm^3)	Melting Point (°C)
Li	[He] 2s^1	152	520	0.535	181
Na	[Ne] 3s^1	186	496	0.968	102
K	[Ar] 4s^1	227	419	0.856	98
Rb	[Kr] 5s^1	248	403	1.532	39
Cs	[Xe] 6s^1	265	376	1.879	29

*Francium is omitted because it has no stable isotopes.

Reactions of the Alkali Metals with Water

Lithium

Sodium

Potassium

Trends in the Halogens

- Atomic radius increases down the column.

- Ionization energy decreases down the column.

- Very high electron affinities
  - Good oxidizing agents; easy to reduce
  - Very reactive; not found uncombined in nature
  - React with metals to form salts
  - Compounds generally soluble in water $\therefore$ found in seawater

Trends in the Halogens

- Reactivity increases down the column.

- They react with hydrogen to form HX, acids.

- Melting point and boiling point increase down the column.

- Density increases down the column.
  - In general, the increase in mass is greater than the increase in volume.

Halogens

TABLE 8.3 Properties of the Halogens*

Element	Electron Configuration	Atomic Radius (pm)	EA (kJ/mol)	Melting Point (°C)	Boiling Point (°C)	Density of Liquid (g/cm^3)
F	[He] $2s^2 2p^5$	72	−328	−219	−188	1.51
Cl	[Ne] $3s^2 3p^5$	99	−349	−101	−34	2.03
Br	[Ar] $4s^2 4p^5$	114	−325	−7	59	3.19
I	[Kr] $5s^2 5p^5$	133	−295	114	184	3.96

*At is omitted because it is rare and radioactive.

Chlorine

Bromine

Iodine
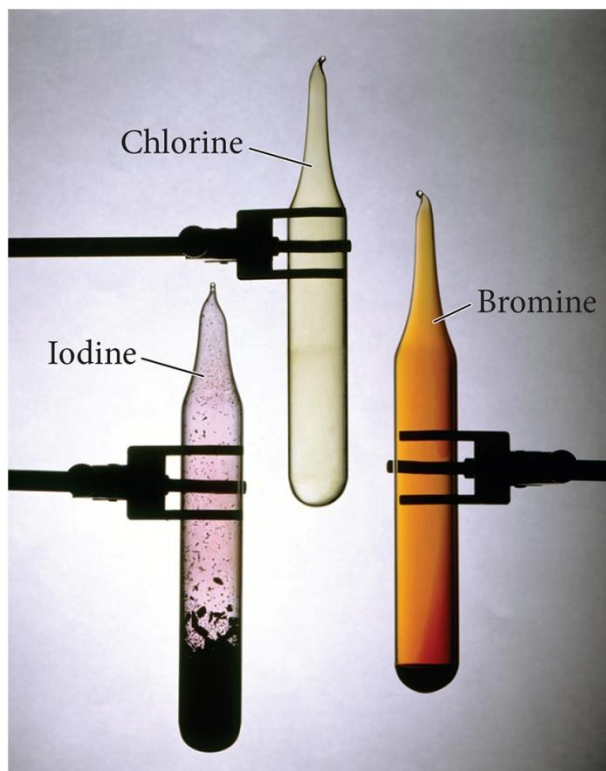
Reactions of Alkali Metals with Halogens

- Alkali metals are oxidized to the 1+ ion.

- Halogens are reduced to the 1− ion.

- The ions then attach together by ionic bonds.

- The reaction is exothermic.

Reactions of Alkali Metals with Water

- Alkali metals are oxidized to the 1+ ion.
- H_2O is split into $H_2(g)$ and OH^- ion.
- The Li, Na, and K are less dense than the water, so they float on top.
- The ions then attach together by ionic bonds.
- The reaction is exothermic, and often the heat released ignites the $H_2(g)$.

Reactions of the Alkali Metals with Water

Lithium   Sodium   Potassium

Trends in the Noble Gases

- Atomic radius increases down the column.

- Ionization energy decreases down the column.
  - Very high IE

- Very unreactive
  - Only found uncombined in nature
  - Used as "inert" atmosphere when reactions with other gases would be undesirable

Trends in the Noble Gases

- Melting point and boiling point increase down the column.
  - All gases at room temperature
  - Very low boiling points

- Density increases down the column.
  - In general, the increase in mass is greater than the increase in volume.

Noble Gases

TABLE 8.4 Properties of the Noble Gases*

Element	Electron Configuration	Atomic Radius (pm)**	IE_1(kJ/mol)	Boiling Point (K)	Density of Gas (g/L at STP)
He	$1s^2$	32	2372	4.2	0.18
Ne	[He] $2s^2 2p^6$	70	2081	27.1	0.90
Ar	[Ne] $3s^2 3p^6$	98	1521	87.3	1.78
Kr	[Ar] $4s^2 4p^6$	112	1351	119.9	3.74
Xe	[Kr] $5s^2 5p^6$	130	1170	165.1	5.86

*Radon is omitted because it is radioactive.

**Since only the heavier noble gases form compounds, covalent radii for the smaller noble gases are estimated.