Electron Configurations and the Periodic Table

- **The periodic table can be used as a guide for electron configurations.**
- **The period number is the value of *n*.**
- **Groups 1A and 2A have the *s*-orbital filled.**
- **Groups 3A - 8A have the *p*-orbital filled.**
- **Groups 3B - 2B have the *d*-orbital filled.**
- **The lanthanides and actinides have the *f*-orbital filled.**

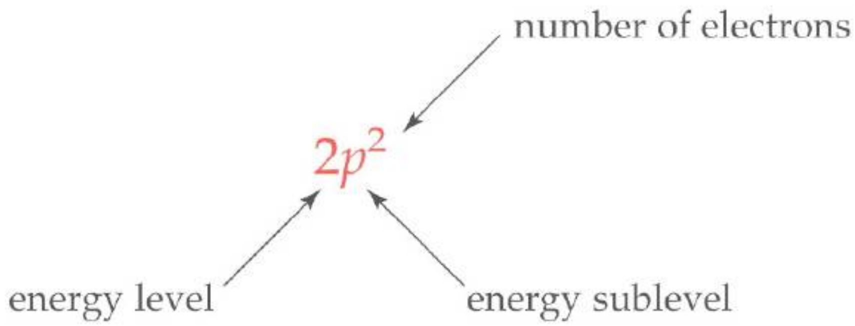
Electron Configuration

Na: 1s^2 2s^2 2p^6 3s^1

Na: [Ne] 3s^1

Electron Configurations

- The ***<u>electron configuration</u>*** of an atom is a shorthand method of writing the location of electrons by sublevel.
- The sublevel is written followed by a superscript with the number of electrons in the sublevel.
  - If the 2*p* sublevel contains 2 electrons, it is written $2p^2$



Electron Configurations and the Periodic Table

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- Groups 3B - 2B have the *d*-orbital filled.
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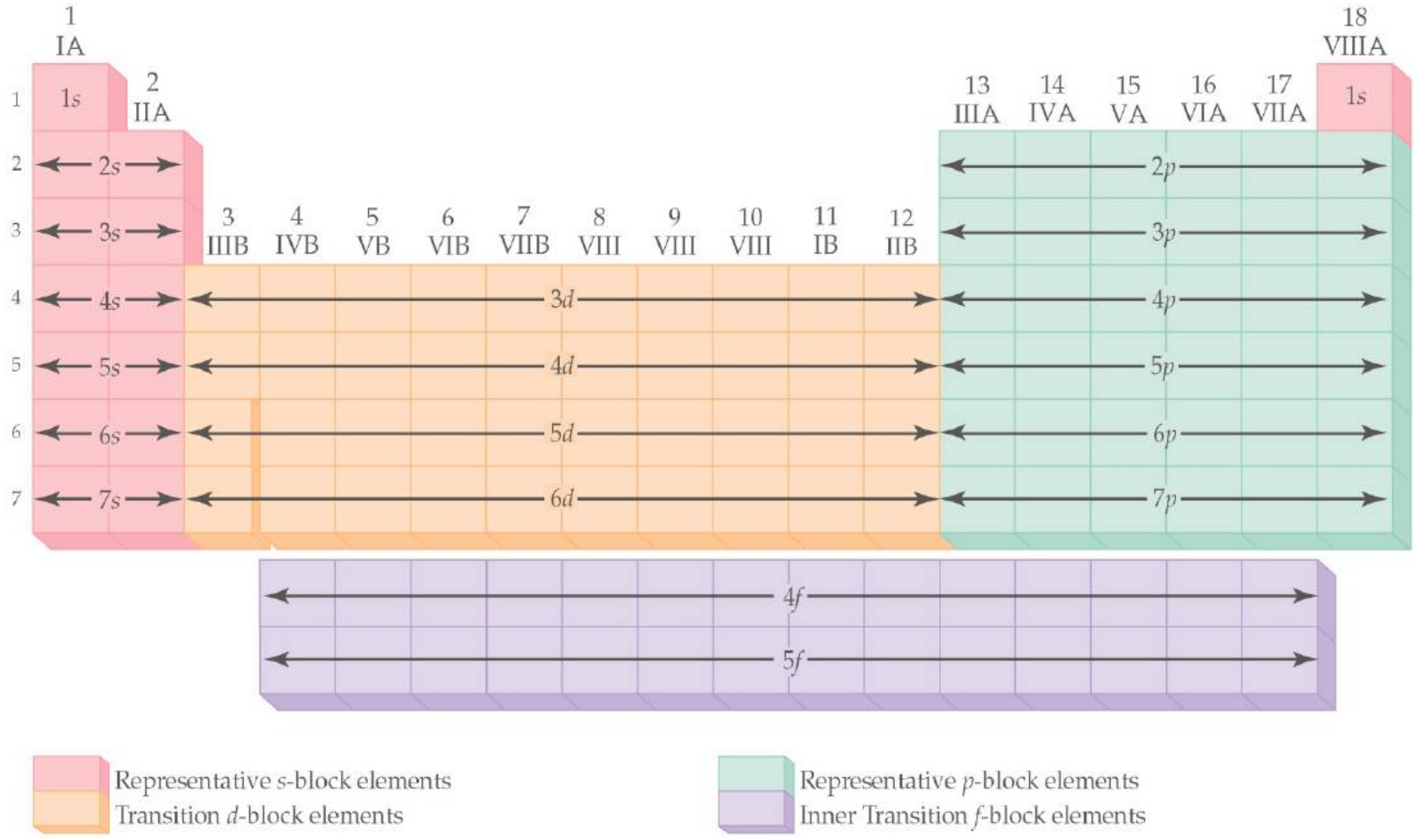
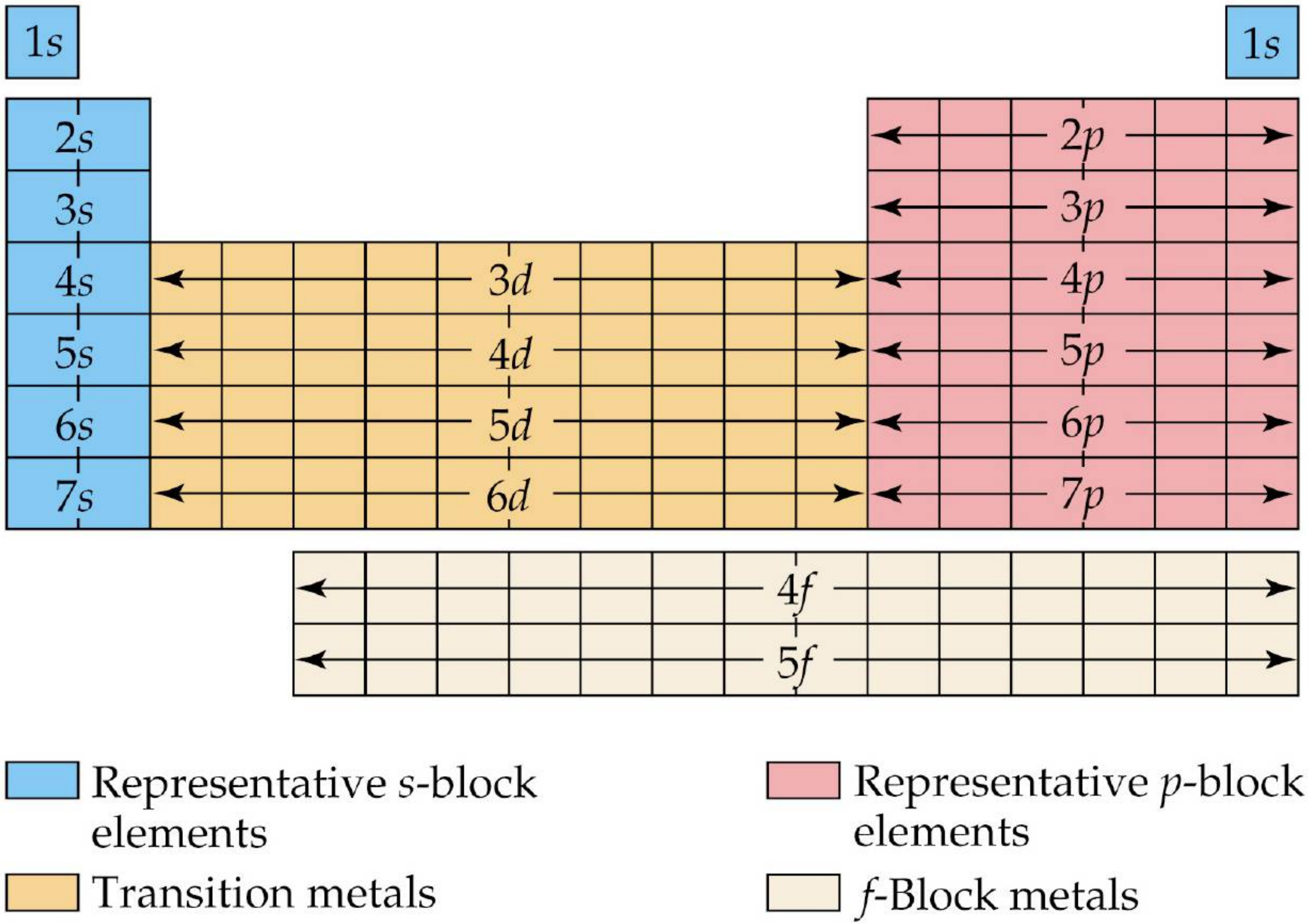
Blocks and Sublevels

- We can use the periodic table to predict which sublevel is being filled by a particular element.




1s
1s
2s
3s
4s
5s
6s
7s
3d
4d
5d
6d
2p
3p
4p
5p
6p
7p
4f
5f
Representative s-block elements
Transition metals
Representative p-block elements
f-Block metals

Noble Gas Core Electron Configurations

- Recall, the electron configuration for Na is:

  Na: $1s^2\ 2s^2\ 2p^6\ 3s^1$

- We can abbreviate the electron configuration by indicating the innermost electrons with the symbol of the preceding noble gas.

- The preceding noble gas with an atomic number less than sodium is neon, Ne. We rewrite the electron configuration:

  Na: [Ne] $3s^1$

Electron Configurations

Condensed Electron Configurations

- Neon completes the 2*p* subshell.
- Sodium marks the beginning of a new row.
- So, we write the condensed electron configuration for sodium as

$$Na: [Ne]\ 3s^1$$

- [Ne] represents the electron configuration of neon.
- **Core electrons:** electrons in [Noble Gas].
- **Valence electrons:** electrons outside of [Noble Gas].

Electron Configurations of Cations of Transition Metals

When a cation is formed from an atom of a transition metal, electrons are always removed first from the *ns* orbital and then from the (n – 1)d orbitals.

Fe: $[Ar]4s^2 3d^6$

Fe^{2+}: $[Ar]4s^0 3d^6$ or $[Ar]3d^6$

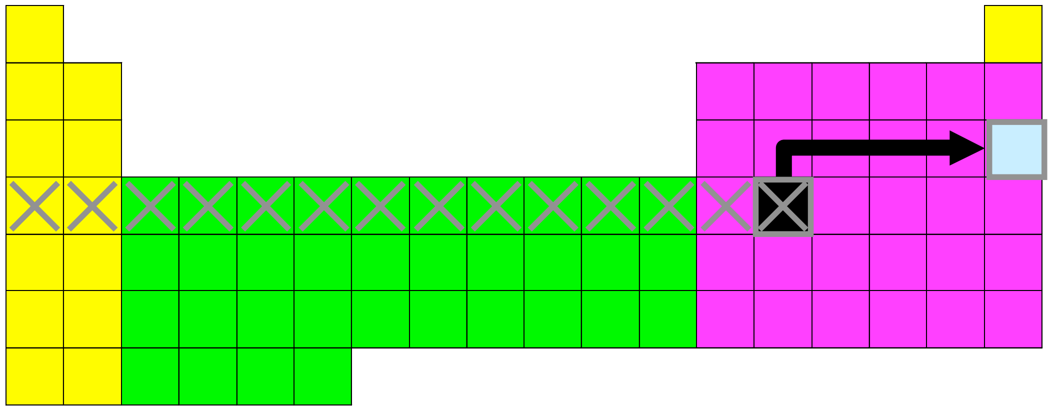
Fe^{3+}: $[Ar]4s^0 3d^5$ or $[Ar]3d^5$

Mn: $[Ar]4s^2 3d^5$

Mn^{2+}: $[Ar]4s^0 3d^5$ or $[Ar]3d^5$

C. Periodic Patterns

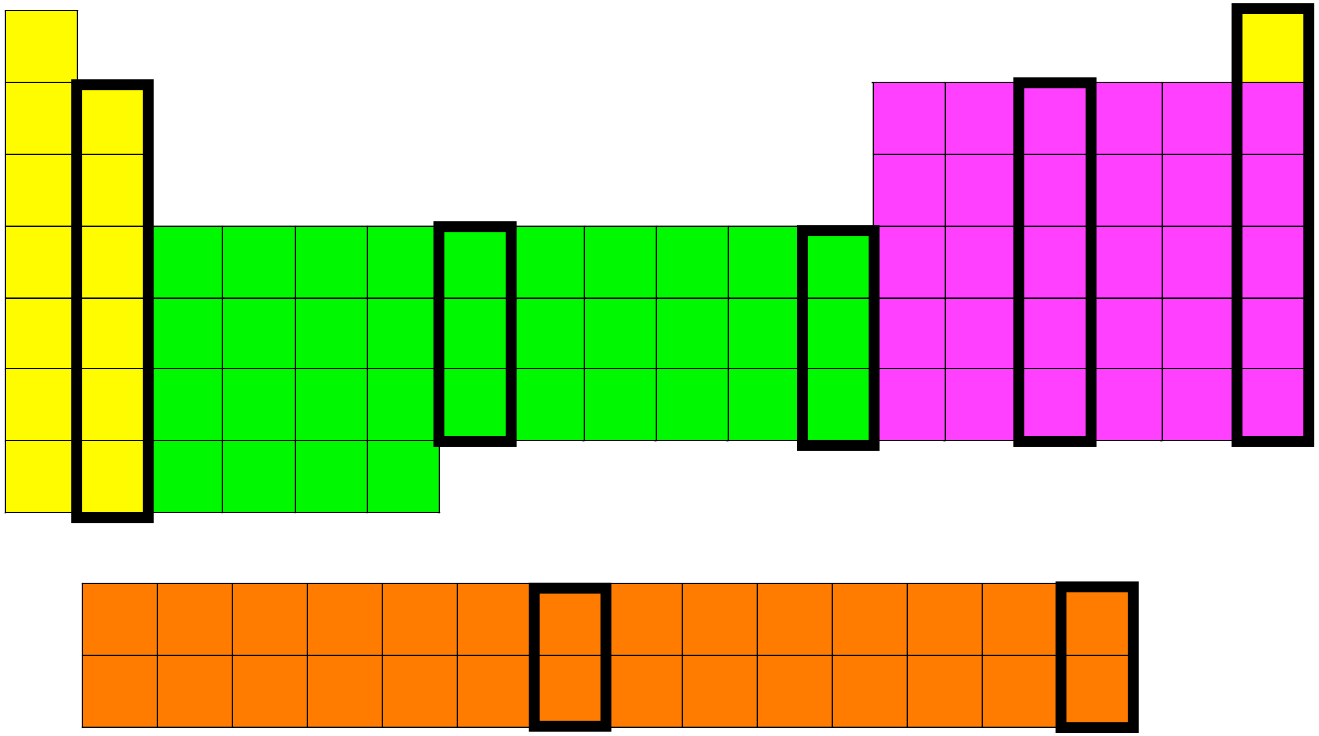
- **Example -** Germanium



$[Ar]\ 4s^2\ 3d^{10}\ 4p^2$

D. Stability

- **Full energy level**
- **Full sublevel (s, p, d, f)**
- **Half-full sublevel**



D. Stability

- **Electron Configuration Exceptions**
  - Copper

    EXPECT: [Ar] $4s^2\ 3d^9$

    ACTUALLY: [Ar] $4s^1\ 3d^{10}$

  - Copper gains stability with a full d-sublevel.

D. Stability

- **Electron Configuration Exceptions**
  – Chromium

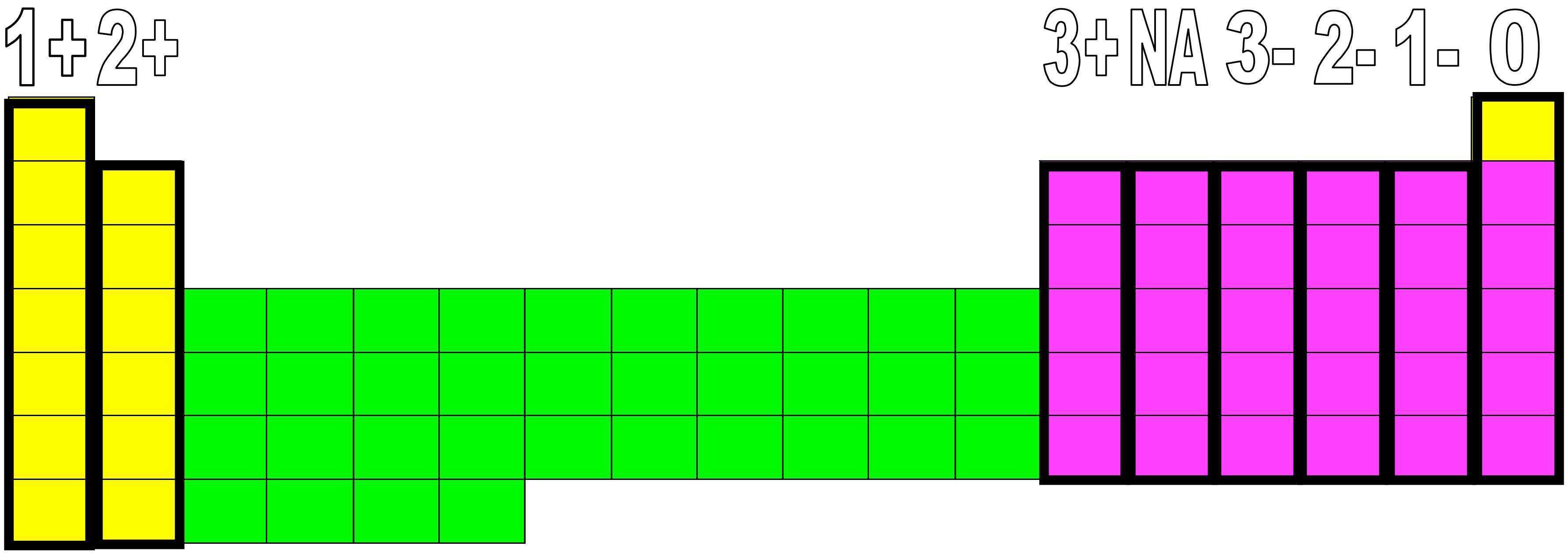
    EXPECT: $[Ar]\ 4s^2\ 3d^4$

    ACTUALLY: $[Ar]\ 4s^1\ 3d^5$

  – Chromium gains stability with a half-full d-sublevel.

D. Stability

- **Ion Formation**
  - Atoms gain or lose electrons to become more stable.
  - Isoelectronic with the Noble Gases.

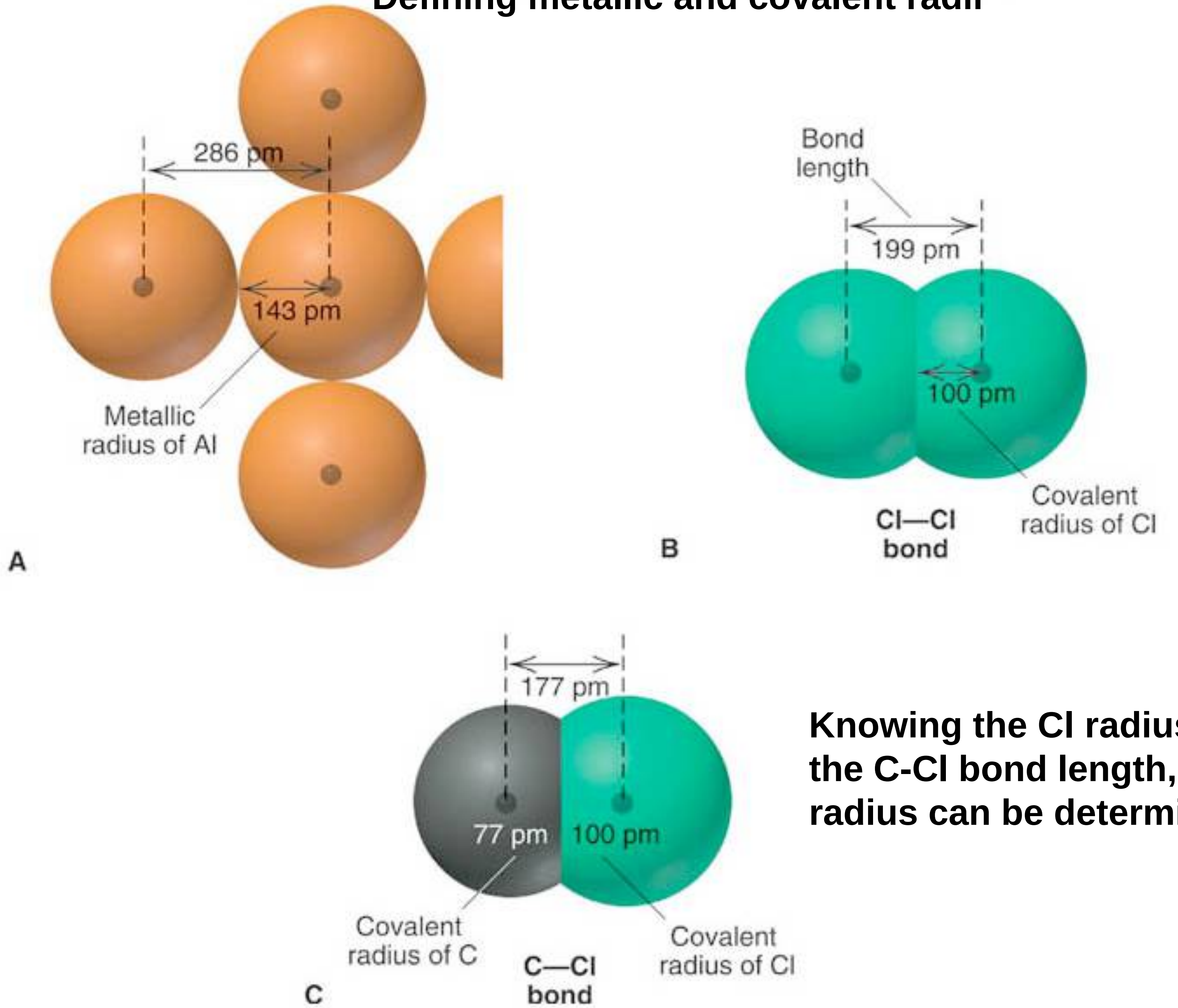


D. Stability

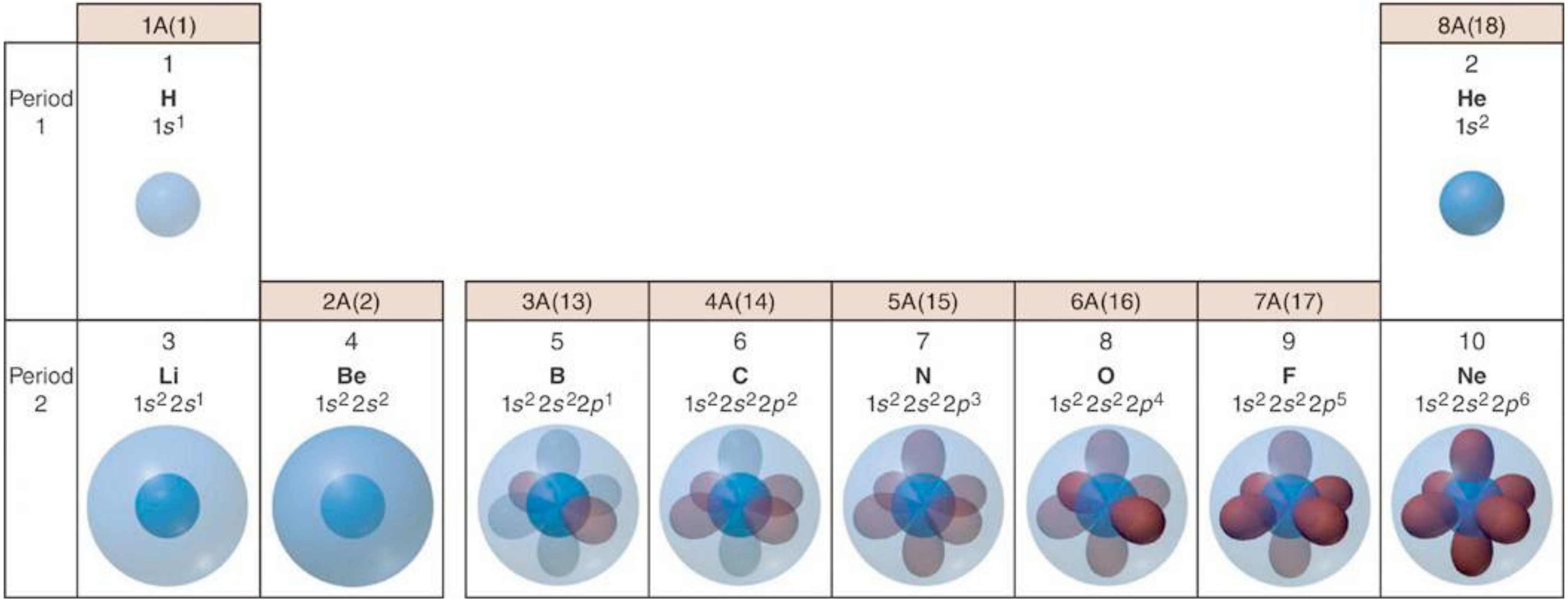
- **Ion Electron Configuration**
  - Write the e^- config for the closest Noble Gas
  - <u>EX</u>: Oxygen ion → O^{2-} ≡ Ne

O^{2-} $10e^-$ $[He]\ 2s^2\ 2p^6$

Defining metallic and covalent radii



Knowing the Cl radius and the C-Cl bond length, the C radius can be determined.


1A(1)
8A(18)
Period 1
1
H
$1s^1$
2
He
$1s^2$
2A(2)
3A(13)
4A(14)
5A(15)
6A(16)
7A(17)
Period 2
3
Li
$1s^2 2s^1$
4
Be
$1s^2 2s^2$
5
B
$1s^2 2s^2 2p^1$
6
C
$1s^2 2s^2 2p^2$
7
N
$1s^2 2s^2 2p^3$
8
O
$1s^2 2s^2 2p^4$
9
F
$1s^2 2s^2 2p^5$
10
Ne
$1s^2 2s^2 2p^6$

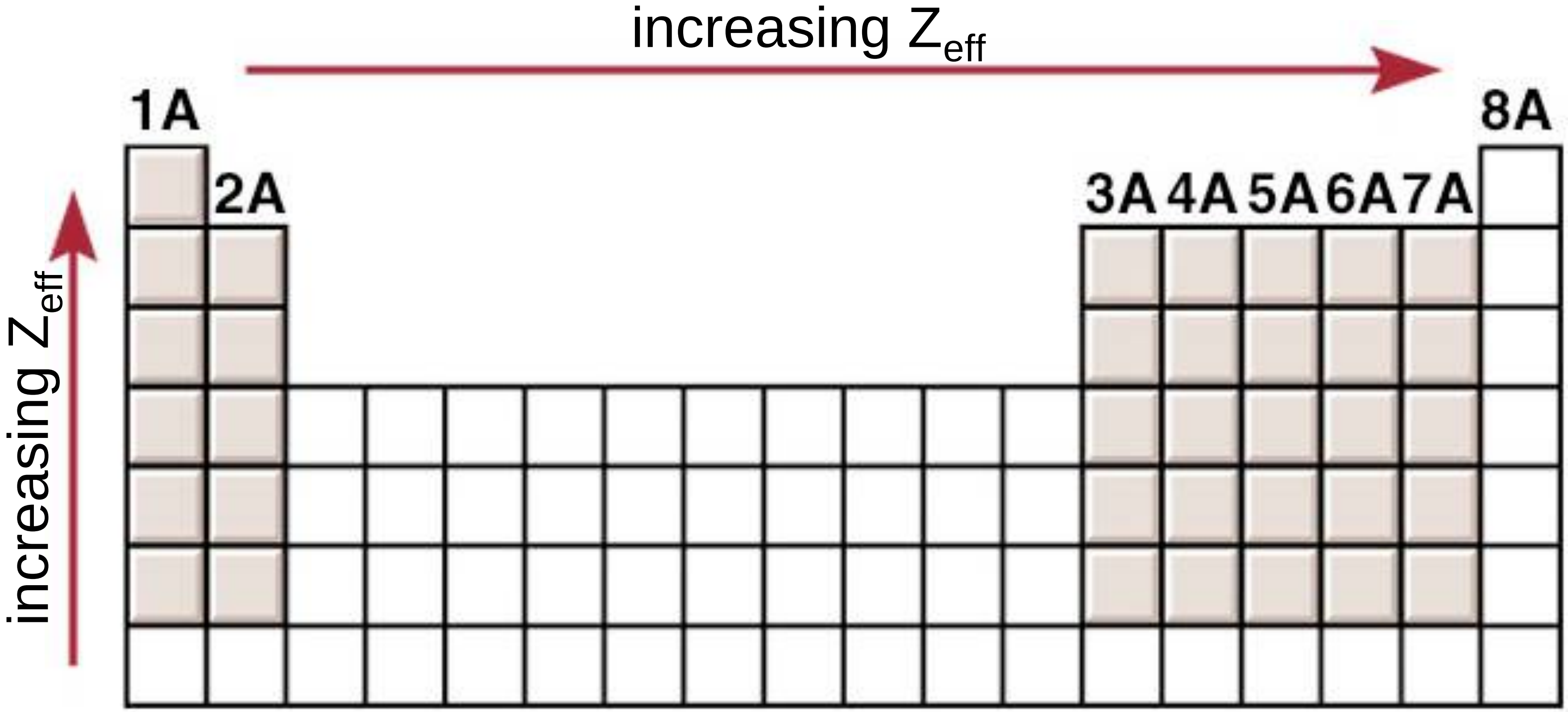
Effective nuclear charge (Z_{eff}) is the “positive charge” felt by an electron.

$Z_{eff} = Z - \sigma$ $0 < \sigma < Z$ (σ = shielding constant)

$Z_{eff} \approx Z$ – number of inner or core electrons

	Z	Core	Z_{eff}	Radius(pm)
Na	11	10	1	186
Mg	12	10	2	160
Al	13	10	3	143
Si	14	10	4	132

Effective Nuclear Charge (Z_{eff})



Trends in the Periodic Table

Electron Affinity:

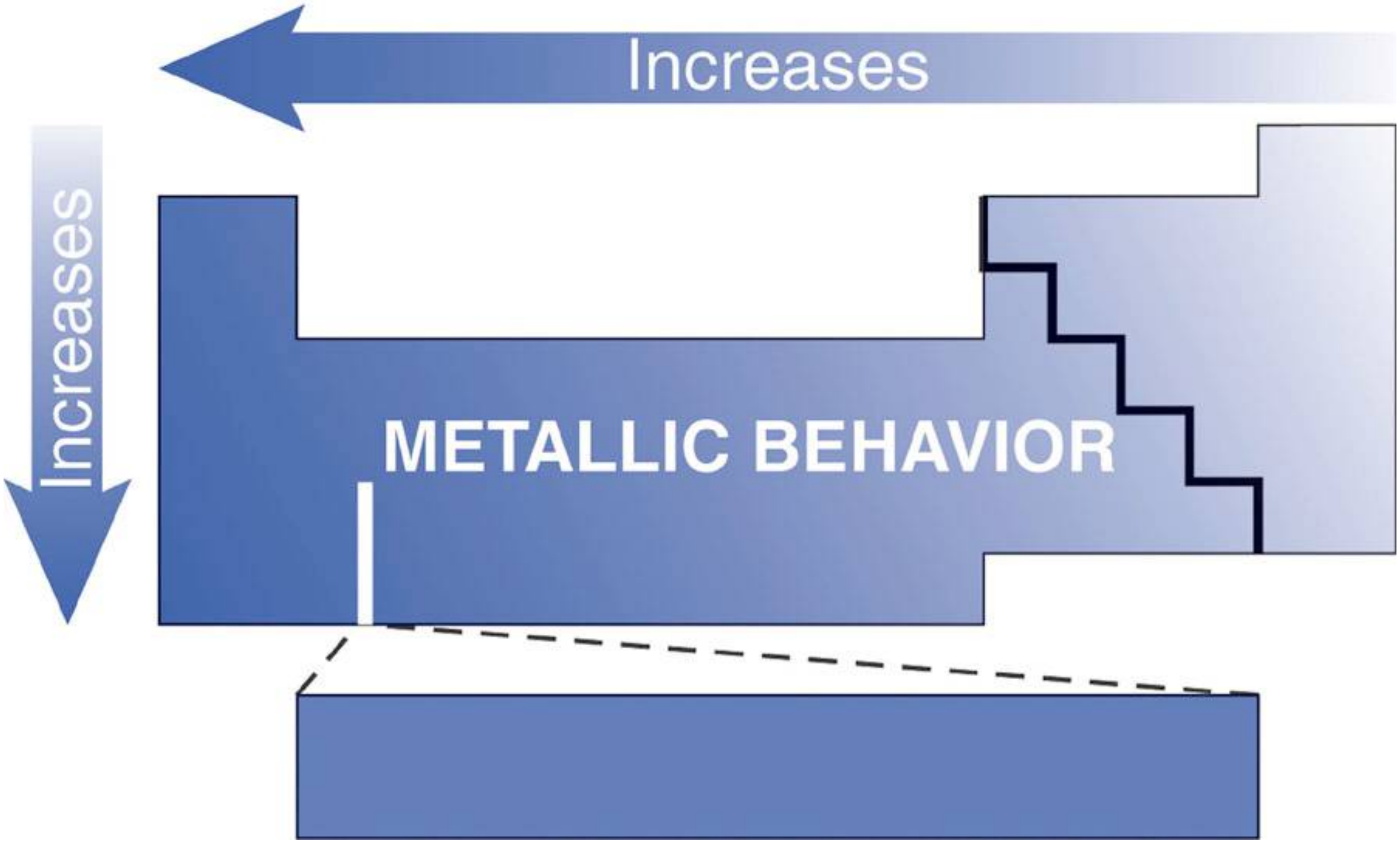
Energy change to add one electron.

In most cases, EA negative (energy released because electron attracted to nucleus

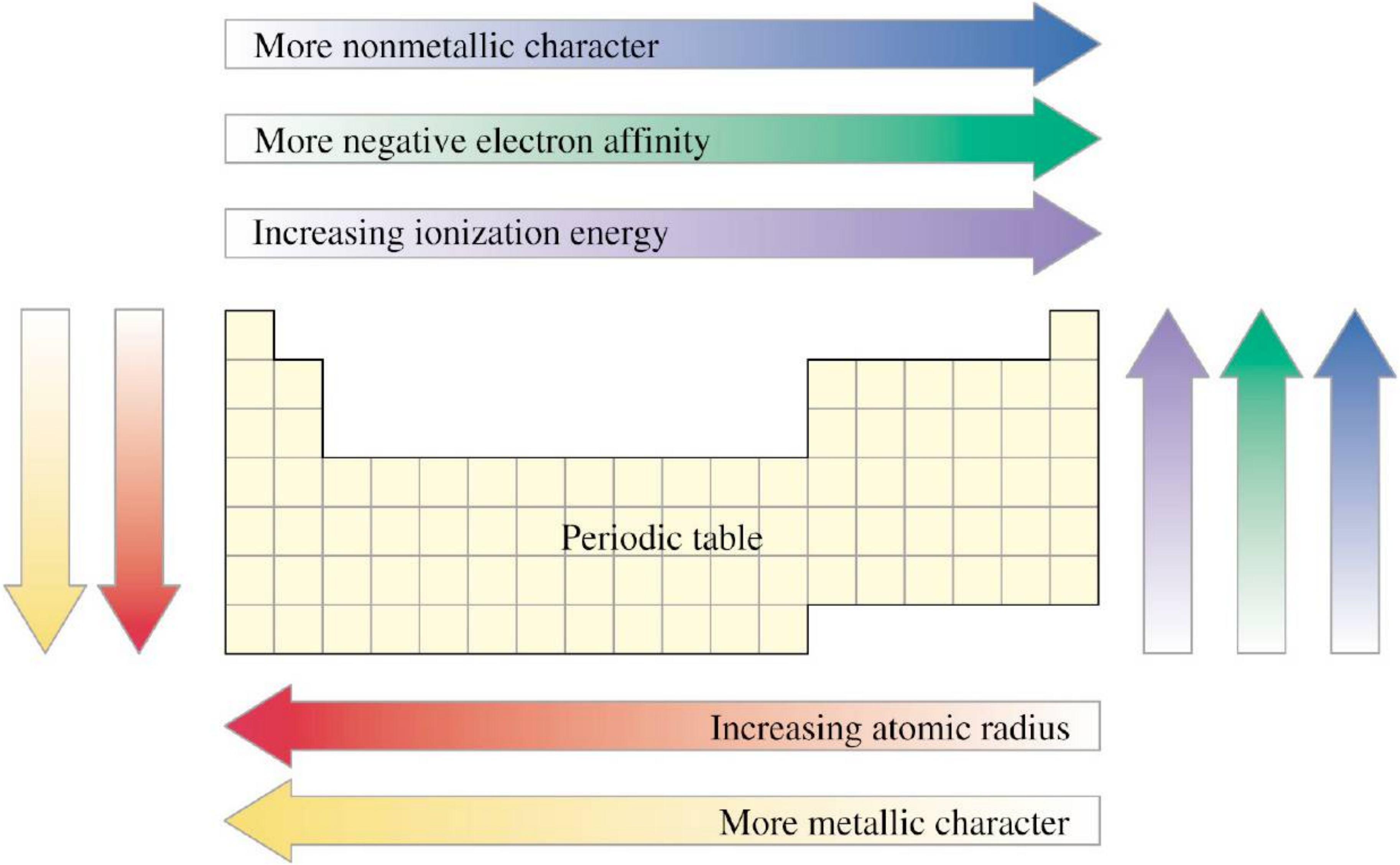
Electron affinities of the main-group elements

1A (1)	2A (2)	3A (13)	4A (14)	5A (15)	6A (16)	7A (17)	8A (18)
H −72.8							**He** (0.0)
Li −59.6	**Be** (+18)	**B** −26.7	**C** −122	**N** +7	**O** −141	**F** −328	**Ne** (+29)
Na −52.9	**Mg** (+21)	**Al** −42.5	**Si** −134	**P** −72.0	**S** −200	**Cl** −349	**Ar** (+35)
K −48.4	**Ca** (+186)	**Ga** −28.9	**Ge** −119	**As** −78.2	**Se** −195	**Br** −325	**Kr** (+39)
Rb −46.9	**Sr** (+146)	**In** −28.9	**Sn** −107	**Sb** −103	**Te** −190	**I** −295	**Xe** (+41)
Cs −45.5	**Ba** (+46)	**Tl** −19.3	**Pb** −35.1	**Bi** −91.3	**Po** −183	**At** −270	**Rn** (+41)

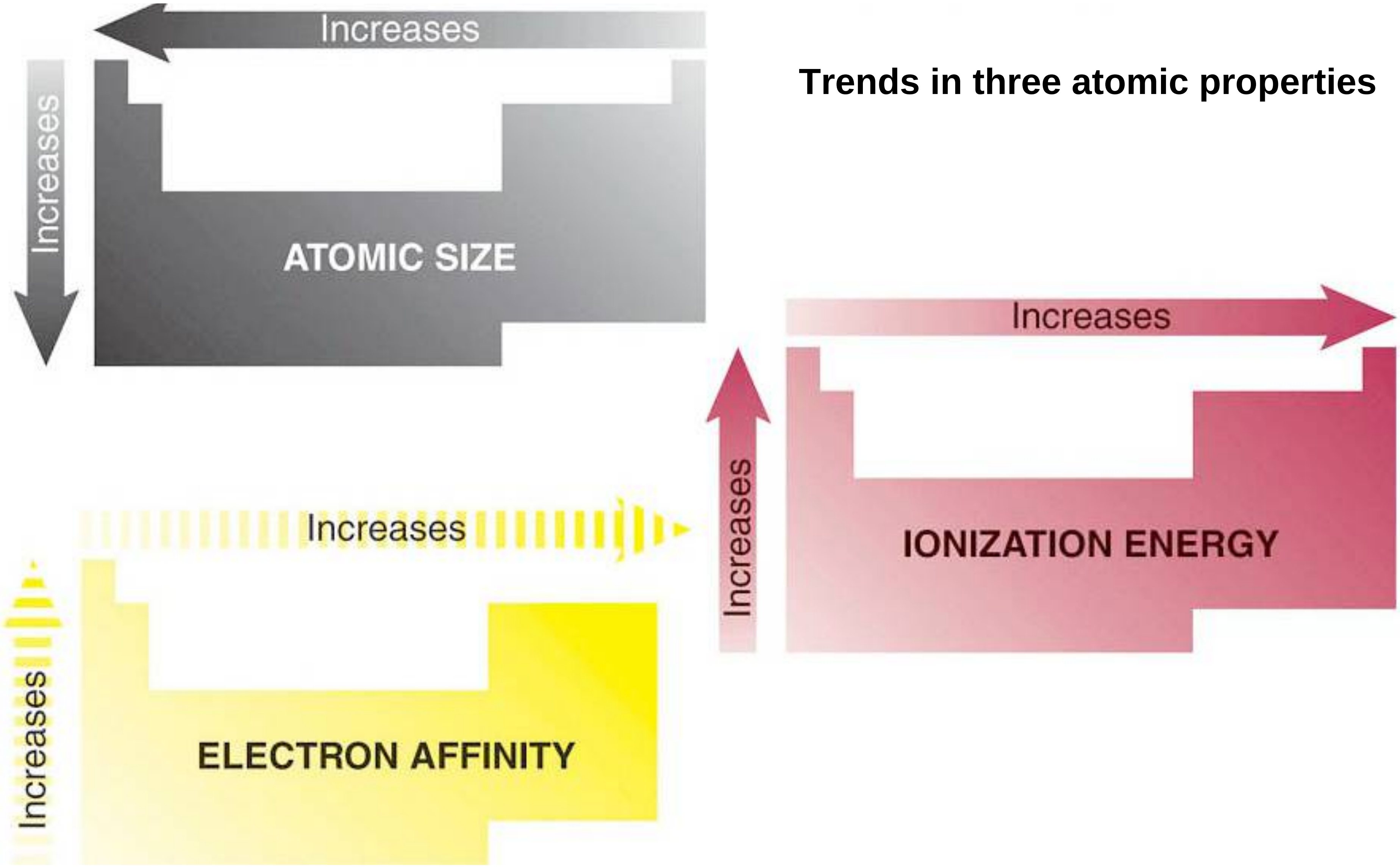
Trends in metallic behavior



A Summary of Periodic Trends



Trends in three atomic properties



Trends in Atomic Size

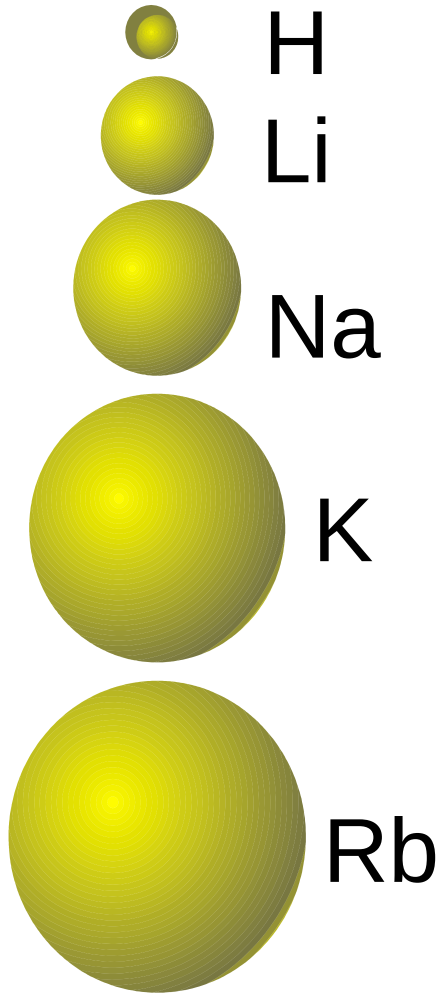
- Influenced by three factors:

  1. Energy Level
     - Higher energy level is further away.
  2. Charge on nucleus
     - More charge pulls electrons in closer.
- 3. Shielding effect

e <-> e repulsion

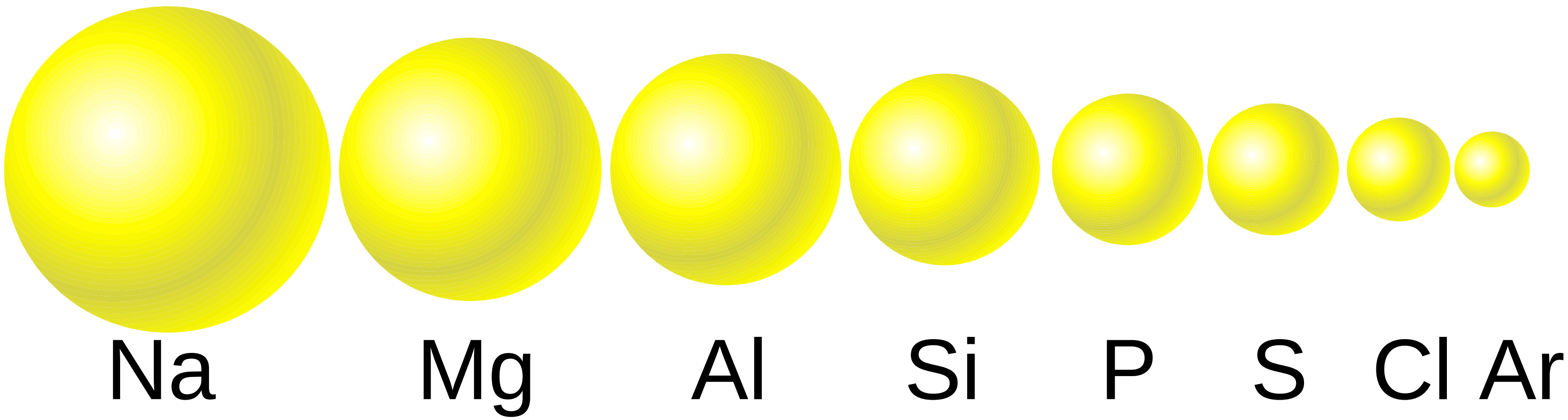
Group trends

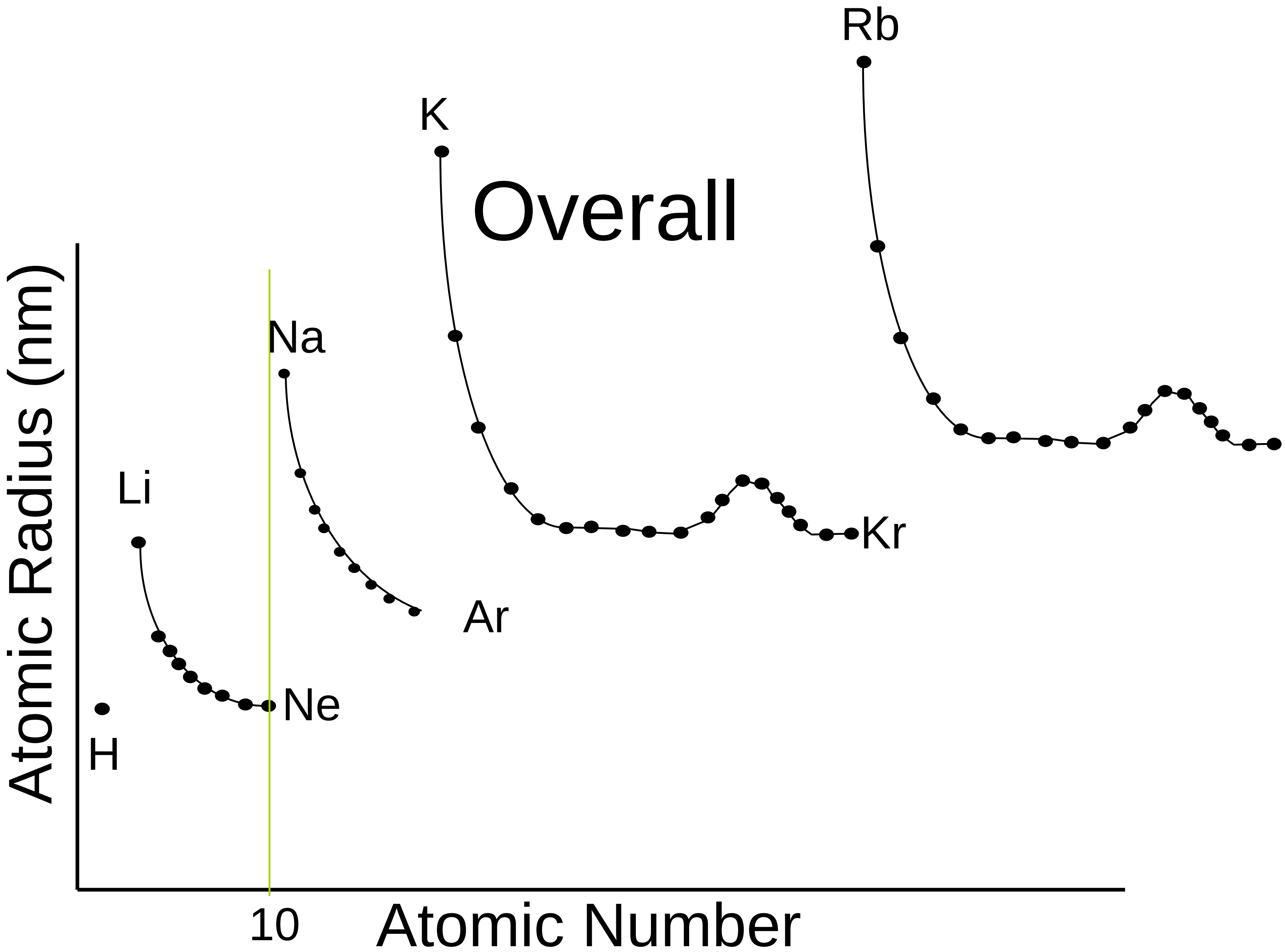
- As we go down a group...
- each atom has another energy level,
- so the atoms get bigger.



Periodic Trends

- As you go across a period, the radius gets smaller.
- Electrons are in same energy level.
- More nuclear charge.
- Outermost electrons are closer.





Overall
Atomic Radius (nm)
Atomic Number
10
H
Li
Ne
Na
Ar
K
Kr
Rb

Electron Configurations of Cations and Anions Of Representative Elements

Na	[Ne]$3s^1$	Na^+	[Ne]
Ca	[Ar]$4s^2$	Ca^{2+}	[Ar]
Al	[Ne]$3s^2 3p^1$	Al^{3+}	[Ne]

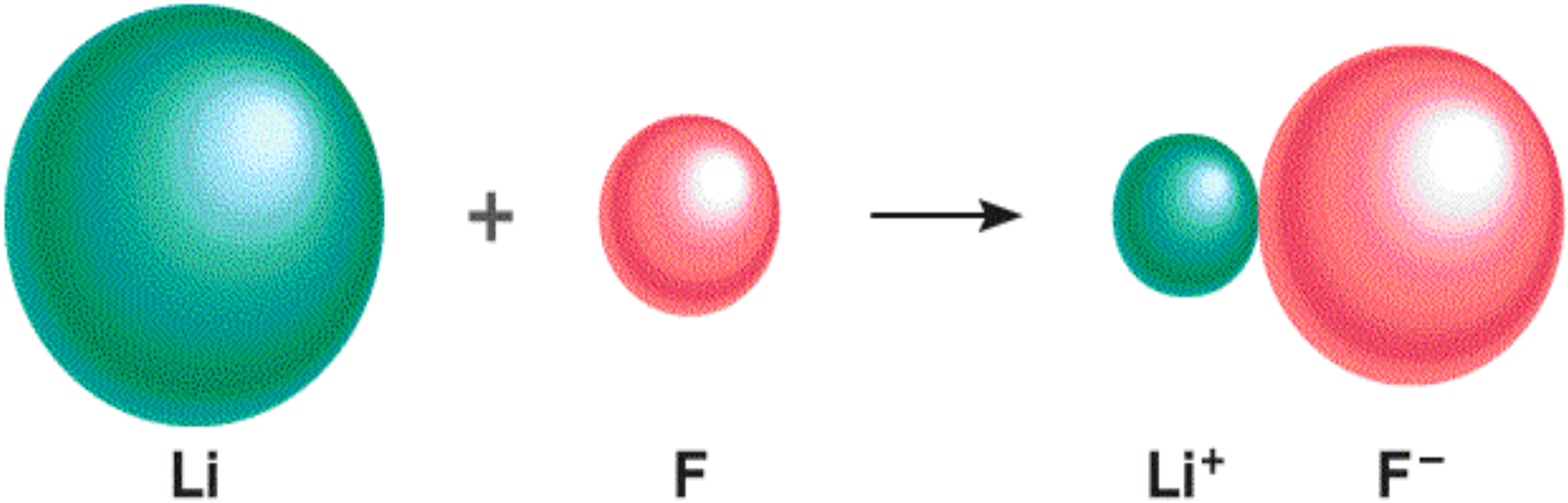
Atoms lose electrons so that cation has a noble-gas outer electron configuration.

Atoms gain electrons so that anion has a noble-gas outer electron configuration.

H	$1s^1$	H^-	$1s^2$ or [He]
F	$1s^2 2s^2 2p^5$	F^-	$1s^2 2s^2 2p^6$ or [Ne]
O	$1s^2 2s^2 2p^4$	O^{2-}	$1s^2 2s^2 2p^6$ or [Ne]
N	$1s^2 2s^2 2p^3$	N^{3-}	$1s^2 2s^2 2p^6$ or [Ne]

Trends in Ionization Energy

- The amount of energy required to completely remove a mole of electrons from a mole of gaseous atoms.
- Removing an electron makes a +1 ion.
- The energy required to remove (1 mole of) the first electron is called the <u>first ionization energy.</u>
  - When an atom or molecule gain or loses an electron it becomes an ion.
    - A **cation** has lost an electron and therefore has a positive charge
    - An **anion** has gained an electron and therefore has a negative charge.



Cation is always **smaller** than atom from which it is formed.
Anion is always **larger** than atom from which it is formed.

Ionization energy is the minimum energy (kJ/mol) required to remove an electron from a gaseous atom in its ground state.

$$I_1 + X_{(g)} \longrightarrow X^{+}_{(g)} + e^{-}$$ I_1 first ionization energy

$$I_2 + X_{(g)} \longrightarrow X^{2+}_{(g)} + e^{-}$$ I_2 second ionization energy

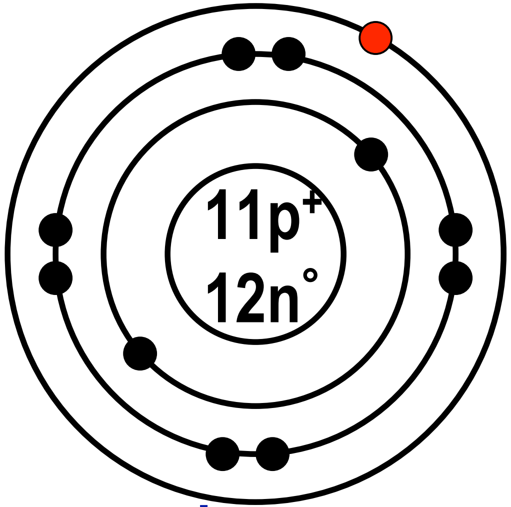
$$I_3 + X_{(g)} \longrightarrow X^{3+}_{(g)} + e^{-}$$ I_3 third ionization energy

$$\boxed{I_1 < I_2 < I_3}$$

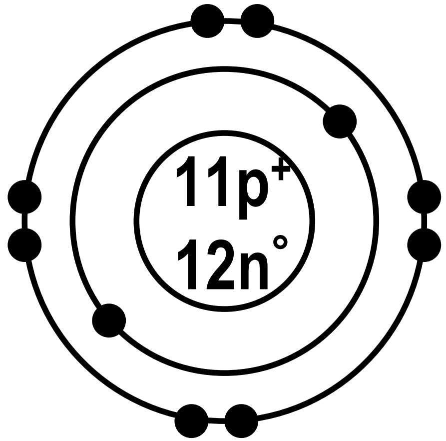
TABLE 8.2

The Ionization Energies (kJ/mol) of the First 20 Elements

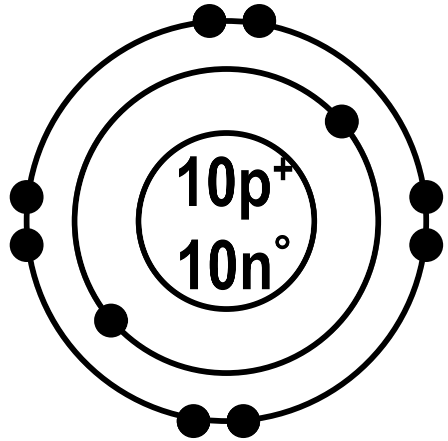
Z	Element	First	Second	Third	Fourth	Fifth	Sixth
1	H	1,312					
2	He	2,373	5,251				
3	Li	520	7,300	11,815			
4	Be	899	1,757	14,850	21,005		
5	B	801	2,430	3,660	25,000	32,820	
6	C	1,086	2,350	4,620	6,220	38,000	47,261
7	N	1,400	2,860	4,580	7,500	9,400	53,000
8	O	1,314	3,390	5,300	7,470	11,000	13,000
9	F	1,680	3,370	6,050	8,400	11,000	15,200
10	Ne	2,080	3,950	6,120	9,370	12,200	15,000
11	Na	495.9	4,560	6,900	9,540	13,400	16,600
12	Mg	738.1	1,450	7,730	10,500	13,600	18,000
13	Al	577.9	1,820	2,750	11,600	14,800	18,400
14	Si	786.3	1,580	3,230	4,360	16,000	20,000
15	P	1,012	1,904	2,910	4,960	6,240	21,000
16	S	999.5	2,250	3,360	4,660	6,990	8,500
17	Cl	1,251	2,297	3,820	5,160	6,540	9,300
18	Ar	1,521	2,666	3,900	5,770	7,240	8,800
19	K	418.7	3,052	4,410	5,900	8,000	9,600
20	Ca	589.5	1,145	4,900	6,500	8,100	11,000



Na has 11 electrons



Na^+ has 10 electrons



Ne has 10 electrons

electron configuration of Na^+ resembles Ne

Alkali metals become like noble gases

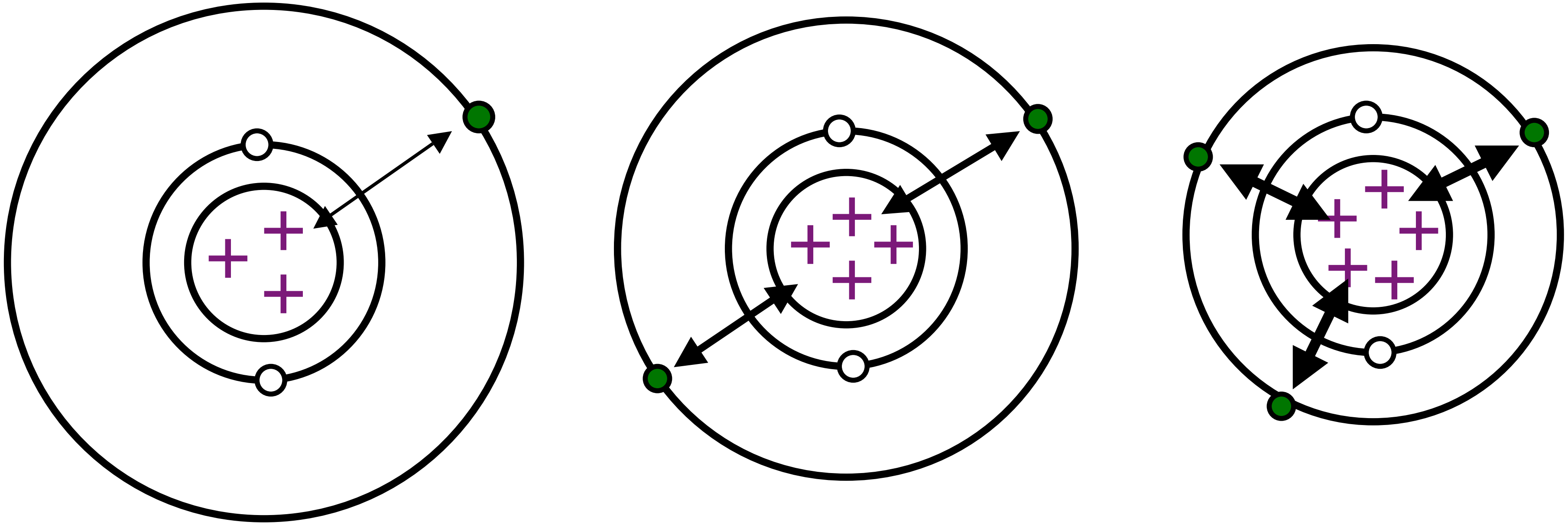
Radius increases because shells are added

Increased radius will make it easier to lose an electron because of greater distance between positive and negative charges

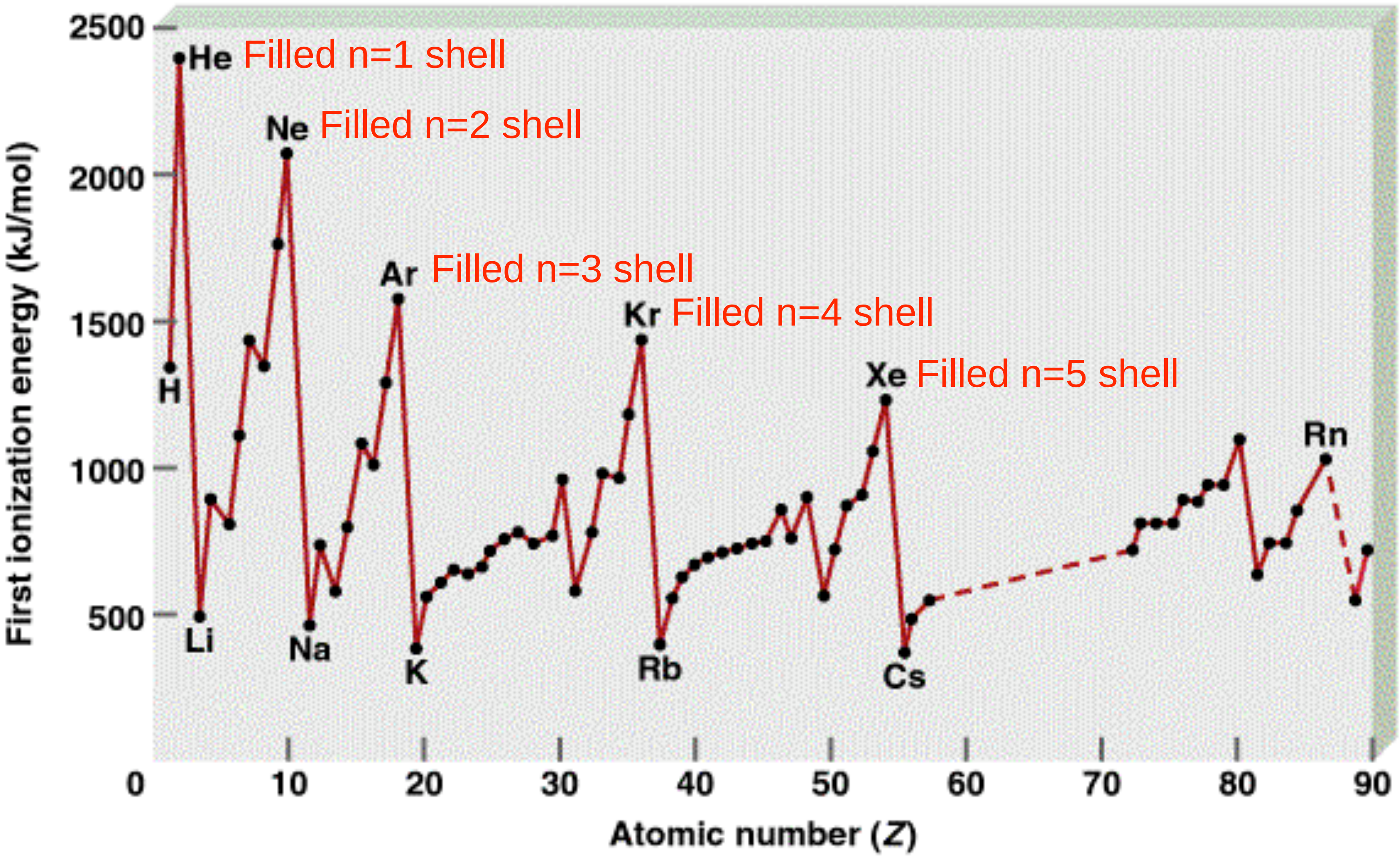
Proton # increases. More protons means greater attraction between nucleus and outer electron thus higher ionization energy.

The greater attraction also means that outer electrons are brought closer to the nucleus, thus smaller atomic radius results.

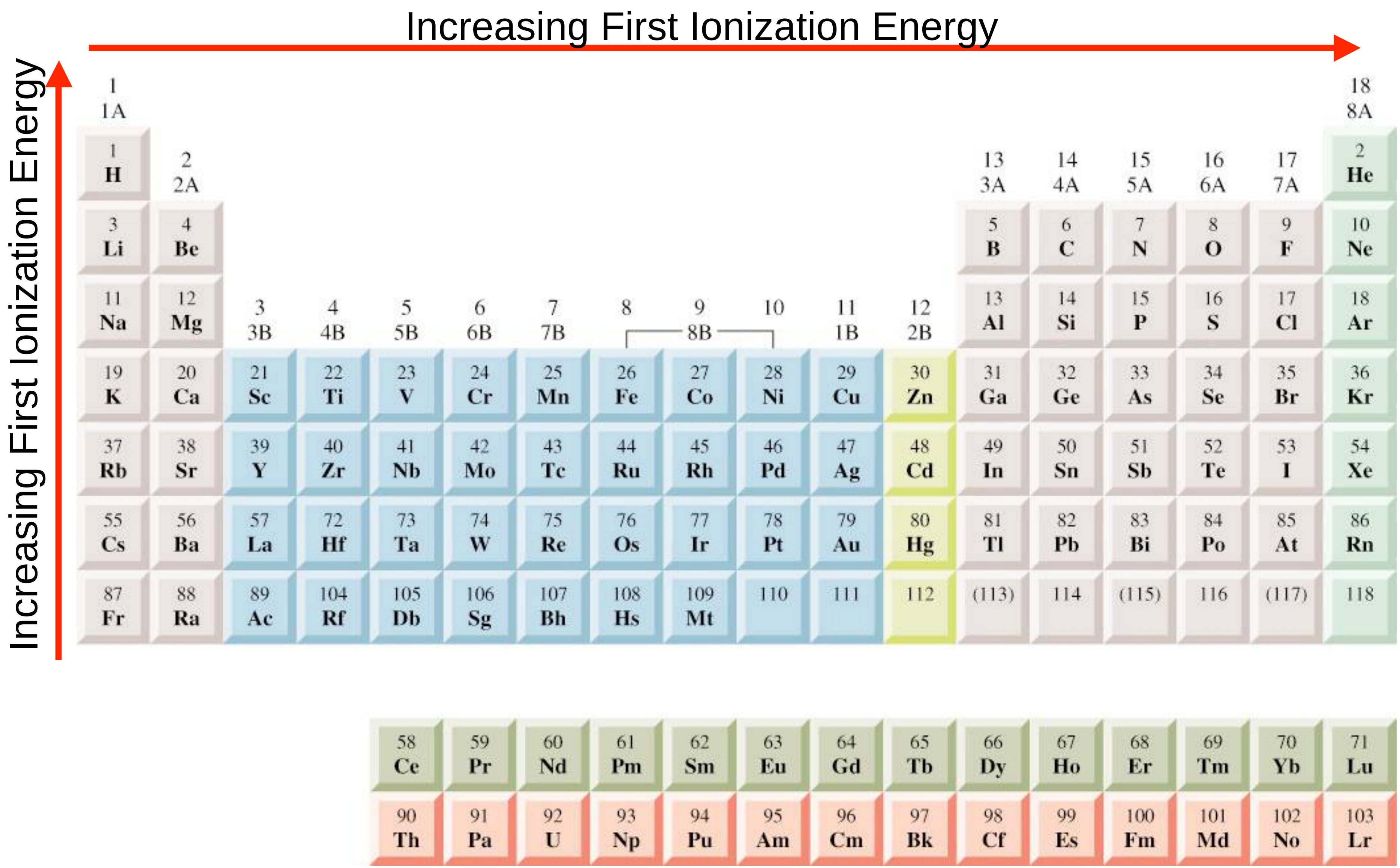
Li (enc = 1) Be (enc = 2) B (enc = 3)



Variation of the First Ionization Energy with Atomic Number



General Trend in First Ionization Energies



Electron affinity is the negative of the energy change that occurs when an electron is accepted by an atom in the gaseous state to form an anion.

$$X_{(g)} + e^{-} \longrightarrow X^{-}_{(g)}$$

$F_{(g)} + e^{-} \longrightarrow X^{-}_{(g)}$ $\Delta H = -328$ kJ/mol EA = +328 kJ/mol

$O_{(g)} + e^{-} \longrightarrow O^{-}_{(g)}$ $\Delta H = -141$ kJ/mol EA = +141 kJ/mol

EXAMPLE

How would you right the symbol for the sodium CATION?

$$Na^{+1}$$

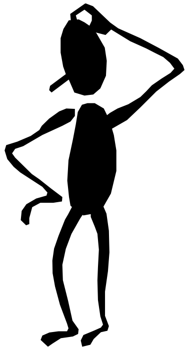
How many outer electrons does sodium have before it loses one?

It has 1…remember the group number!

Na^+: [Ne] Al^{3+}: [Ne] F^-: $1s^22s^22p^6$ or [Ne]

O^{2-}: $1s^22s^22p^6$ or [Ne] N^{3-}: $1s^22s^22p^6$ or [Ne]

Na^+, Al^{3+}, F^-, O^{2-}, and N^{3-} are all ***isoelectronic*** with Ne



What neutral atom is isoelectronic with H^- ?

H^-: $1s^2$ same electron configuration as He

SAMPLE PROBLEM 8.6 Writing Electron Configurations of Main-Group Ions

PROBLEM: Using condensed electron configurations, write reactions for the formation of the common ions of the following elements:

(a) Iodine (Z = 53) **(b)** Potassium (Z = 19) **(c)** Indium (Z = 49)

PLAN: Ions of elements in Groups 1A(1), 2A(2), 6A(16), and 7A(17) are usually isoelectronic with the nearest noble gas.

Metals in Groups 3A(13) to 5A(15) can lose their np or ns and np electrons.

SOLUTION:

(a) Iodine (Z = 53) is in Group 7A(17) and will gain one electron to be isoelectronic with Xe: I ($[Kr]5s^2 4d^{10} 5p^5$) + e^- $\longrightarrow$ I^- ($[Kr]5s^2 4d^{10} 5p^6$)

(b) Potassium (Z = 19) is in Group 1A(1) and will lose one electron to be isoelectronic with Ar: K ($[Ar]4s^1$) $\longrightarrow$ K^+ ([Ar]) + e^-

(c) Indium (Z = 49) is in Group 3A(13) and can lose either one electron or three electrons: In ($[Kr]5s^2 4d^{10} 5p^1$) $\longrightarrow$ In^+ ($[Kr]5s^2 4d^{10}$) + e^-

In ($[Kr]5s^2 4d^{10} 5p^1$) $\longrightarrow$ In^{3+}($[Kr]\ 4d^{10}$) + $3e^-$

SAMPLE PROBLEM 8.8 Ranking Ions by Size

PROBLEM: Rank each set of ions in order of *decreasing* size, and explain your ranking:

(a) Ca^{2+}, Sr^{2+}, Mg^{2+} **(b)** K^{+}, S^{2-}, Cl^{-} **(c)** Au^{+}, Au^{3+}

PLAN: Compare positions in the periodic table, formation of positive and negative ions and changes in size due to gain or loss of electrons.

SOLUTION:

(a) $Sr^{2+} > Ca^{2+} > Mg^{2+}$

These are members of the same Group (2A/2) and therefore decrease in size going up the group.

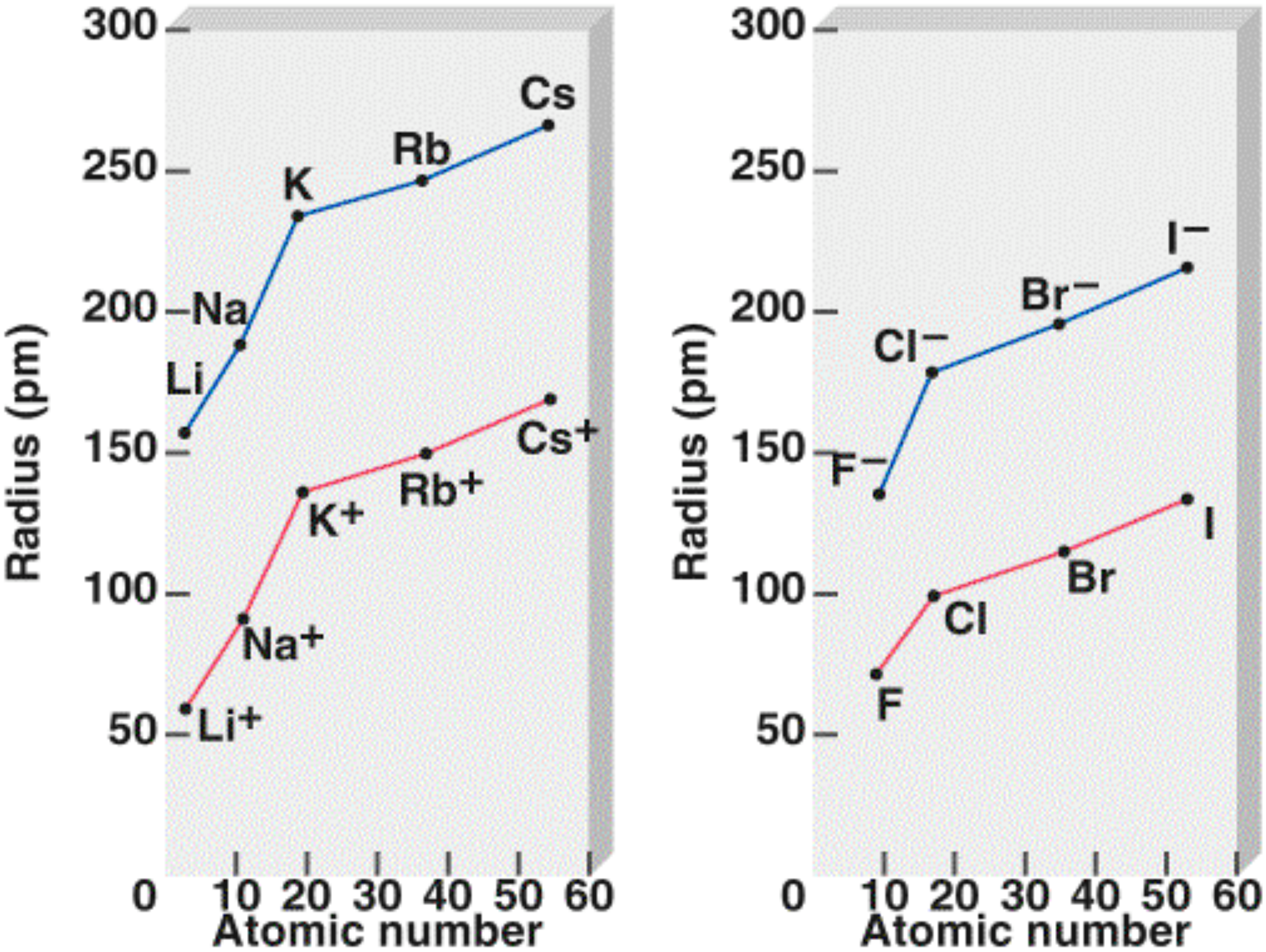
(b) $S^{2-} > Cl^{-} > K^{+}$

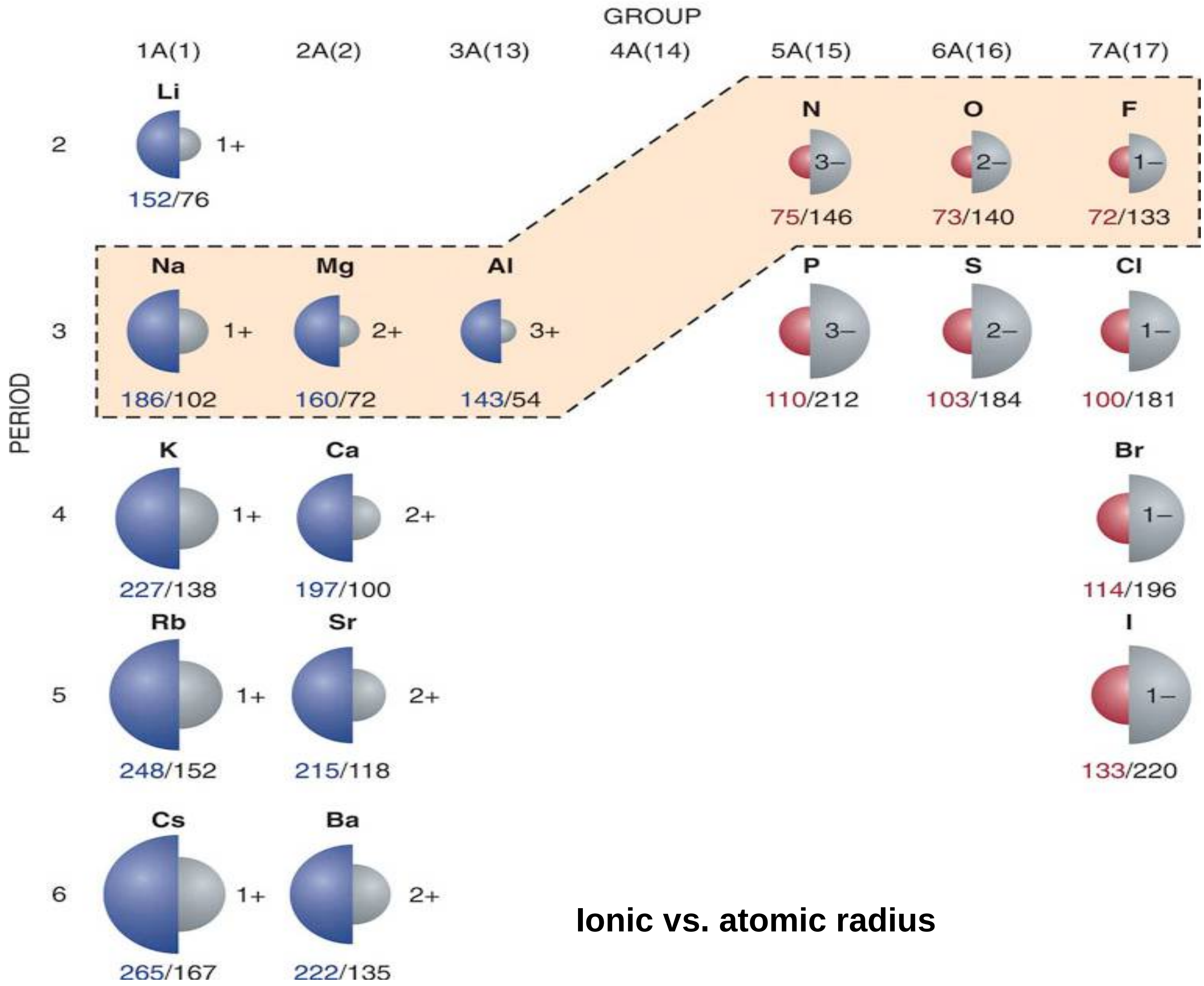
The ions are isoelectronic; S^{2-} has the smallest Z_{eff} and therefore is the largest while K^{+} is a cation with a large Z_{eff} and is the smallest.

(c) $Au^{+} > Au^{3+}$

The higher the + charge, the smaller the ion.

Comparison of Atomic Radii with Ionic Radii





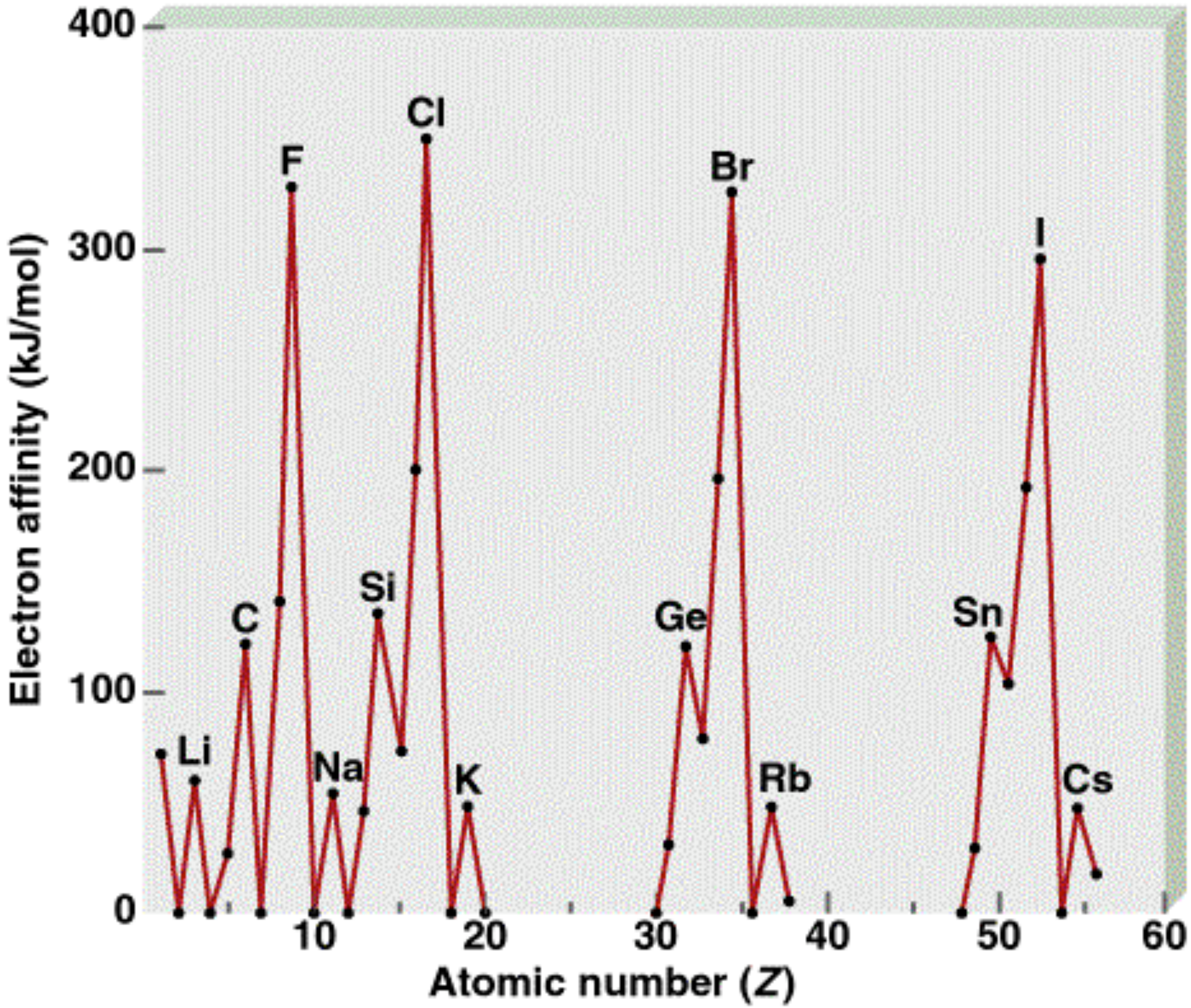
Ionic vs. atomic radius

TABLE 8.3 Electron Affinities (kJ/mol) of Some Representative Elements and the Noble Gases*

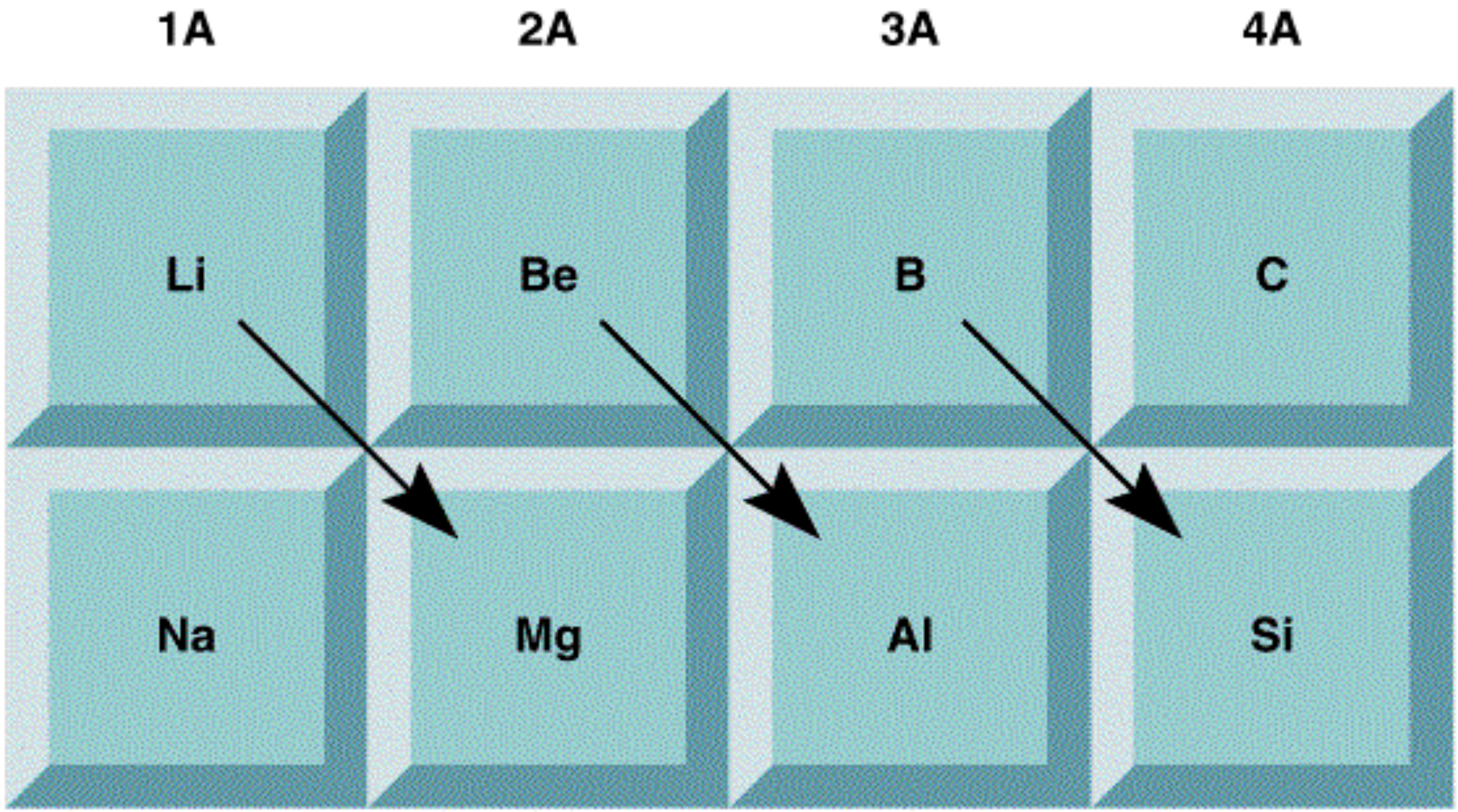
1A	2A	3A	4A	5A	6A	7A	8A
H							He
73							< 0
Li	Be	B	C	N	O	F	Ne
60	≤ 0	27	122	0	141	328	< 0
Na	Mg	Al	Si	P	S	Cl	Ar
53	≤ 0	44	134	72	200	349	< 0
K	Ca	Ga	Ge	As	Se	Br	Kr
48	2.4	29	118	77	195	325	< 0
Rb	Sr	In	Sn	Sb	Te	I	Xe
47	4.7	29	121	101	190	295	< 0
Cs	Ba	Tl	Pb	Bi	Po	At	Rn
45	14	30	110	110	?	?	< 0

* The electron affinities of the noble gases, Be, and Mg have not been determined experimentally, but are believed to be close to zero or negative.

Electron Affinity Versus Atomic Number



Diagonal Relationships in the Periodic Table

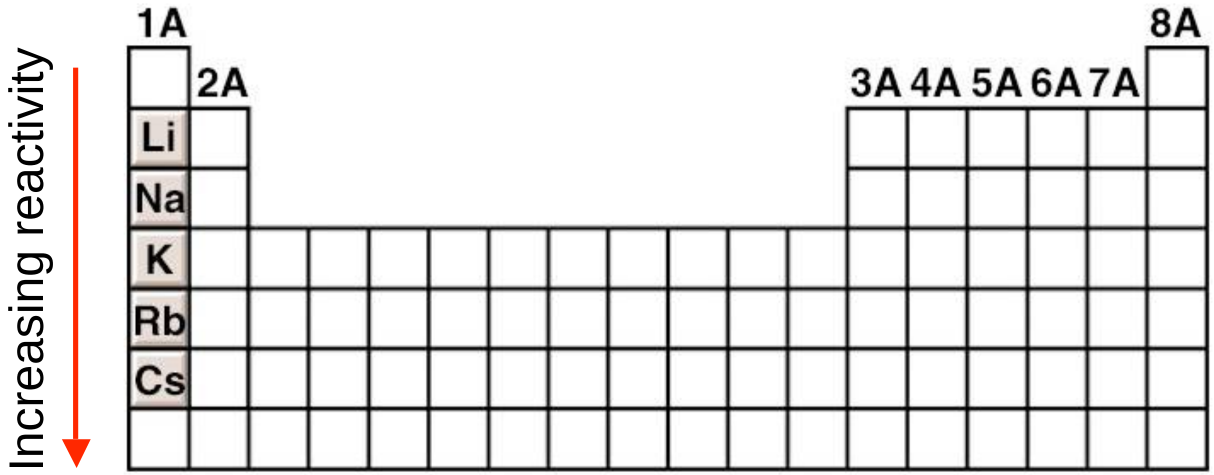


Group 1A Elements (ns^1, $n \geq 2$)

$$M \longrightarrow M^{+1} + 1e^-$$

$$2M_{(s)} + 2H_2O_{(l)} \longrightarrow 2MOH_{(aq)} + H_{2(g)}$$

$$4M_{(s)} + O_{2(g)} \longrightarrow 2M_2O_{(s)}$$



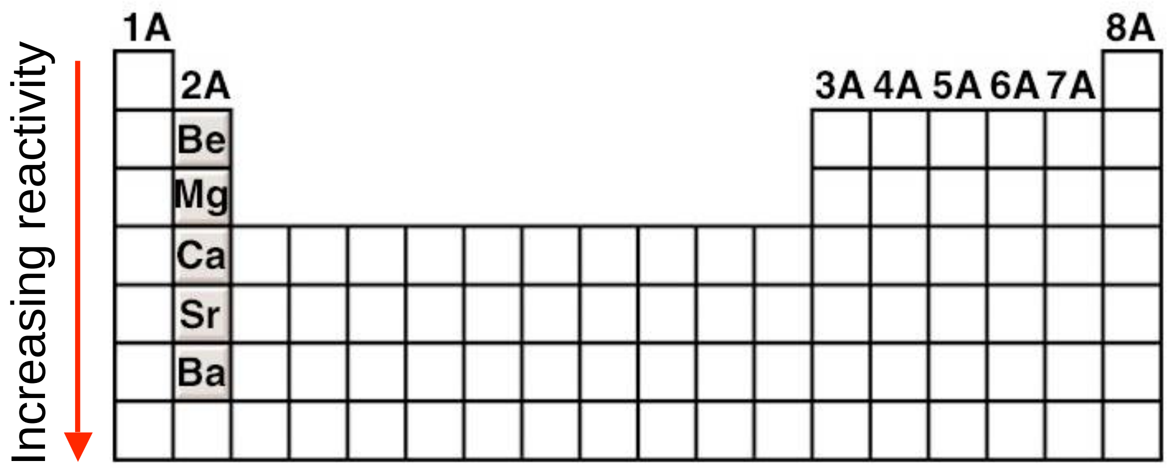
Group 2A Elements (ns^2, $n \geq 2$)

$$M \longrightarrow M^{+2} + 2e^-$$

$$Be_{(s)} + 2H_2O_{(l)} \longrightarrow \text{No Reaction}$$

$$Mg_{(s)} + 2H_2O_{(g)} \longrightarrow Mg(OH)_{2(aq)} + H_{2(g)}$$

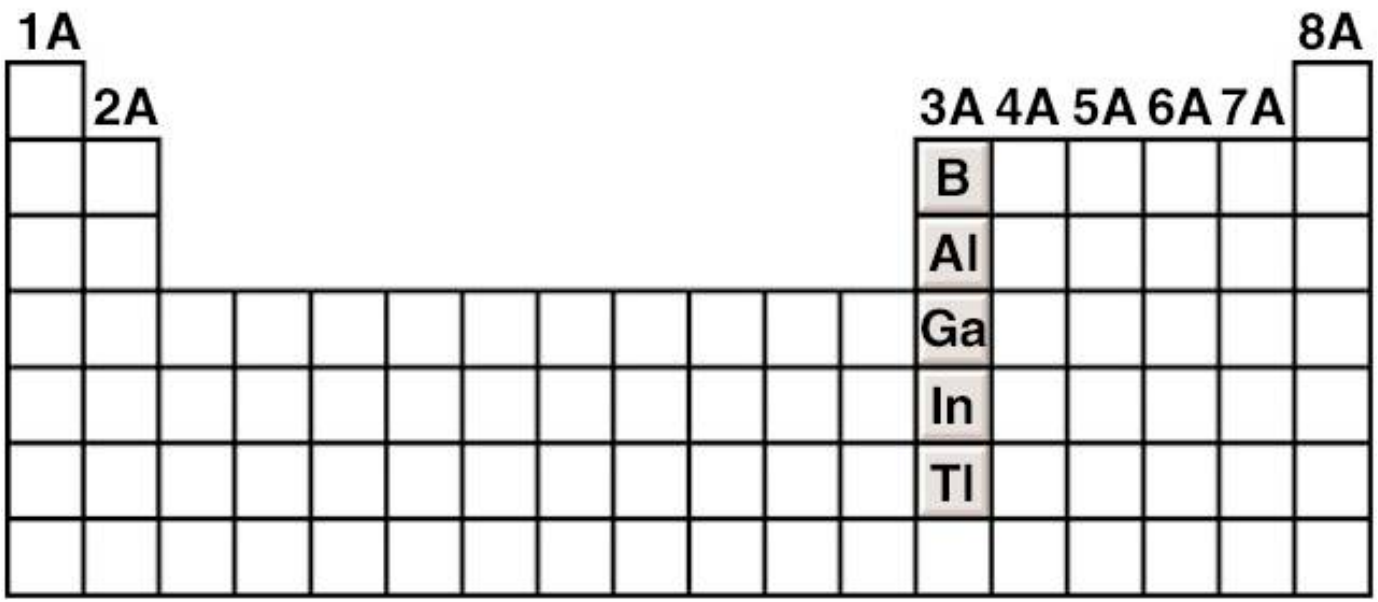
$$M_{(s)} + 2H_2O_{(l)} \longrightarrow M(OH)_{2(aq)} + H_{2(g)} \quad M = \text{Ca, Sr, or Ba}$$



Group 3A Elements (ns^2np^1, $n \geq 2$)

$$4Al_{(s)} + 3O_{2(g)} \longrightarrow 2Al_2O_{3(s)}$$

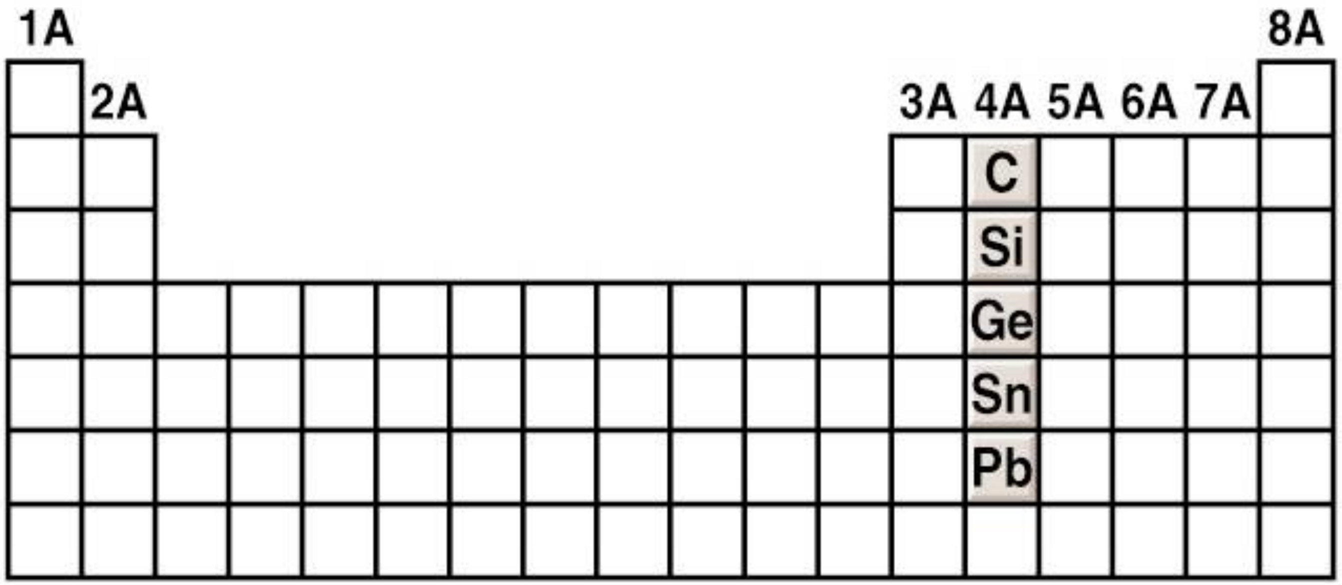
$$2Al_{(s)} + 6H^+_{(aq)} \longrightarrow 2Al^{3+}_{(aq)} + 3H_{2(g)}$$



Group 4A Elements (ns^2np^2, $n \geq 2$)

$$Sn_{(s)} + 2H^{+}_{(aq)} \longrightarrow Sn^{2+}_{(aq)} + H_{2\,(g)}$$

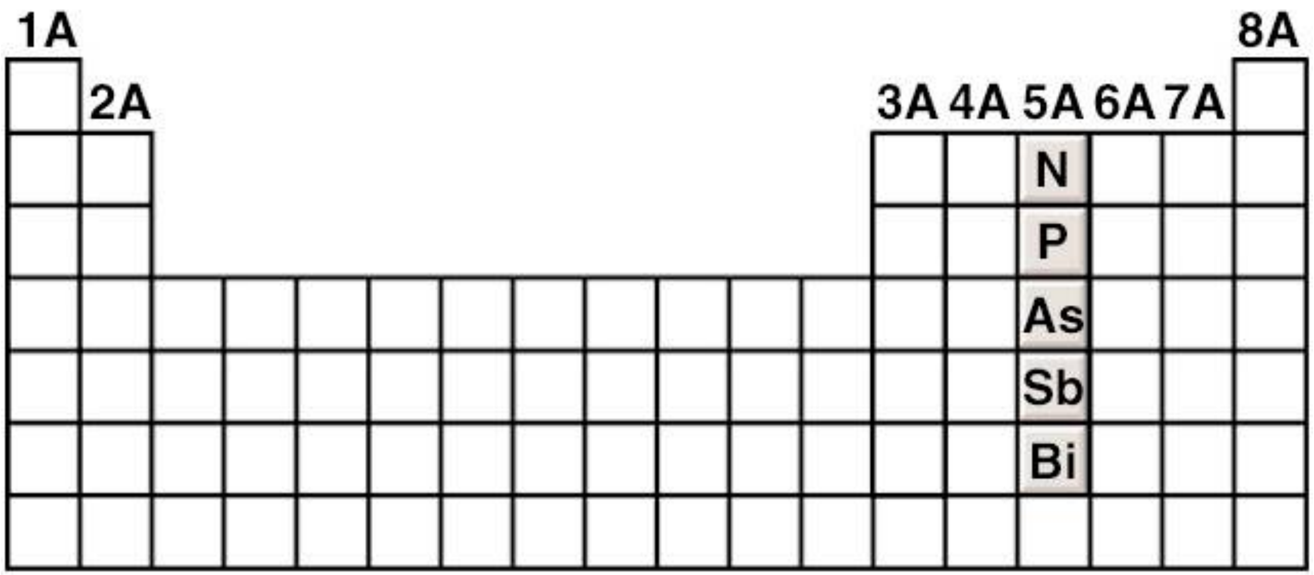
$$Pb_{(s)} + 2H^{+}_{(aq)} \longrightarrow Pb^{2+}_{(aq)} + H_{2\,(g)}$$



Group 5A Elements (ns^2np^3, $n \geq 2$)

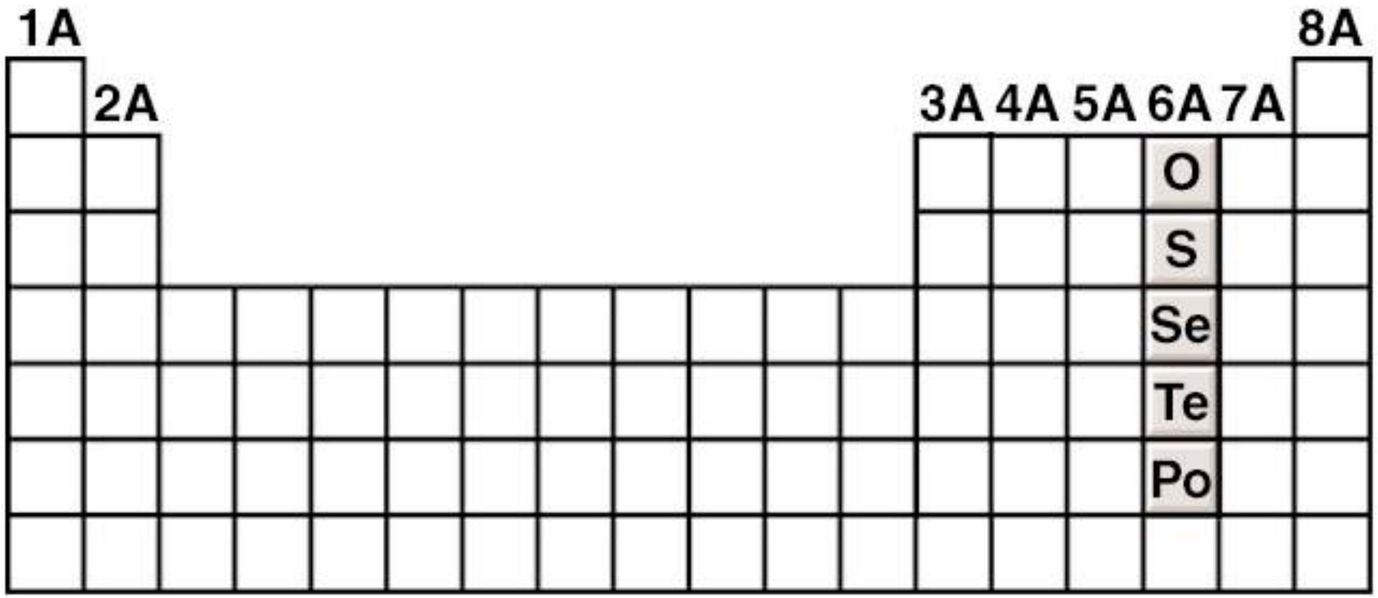
$$N_2O_{5(s)} + H_2O_{(l)} \longrightarrow 2HNO_{3(aq)}$$

$$P_4O_{10(s)} + 6H_2O_{(l)} \longrightarrow 4H_3PO_{4(aq)}$$



Group 6A Elements (ns^2np^4, $n \geq 2$)

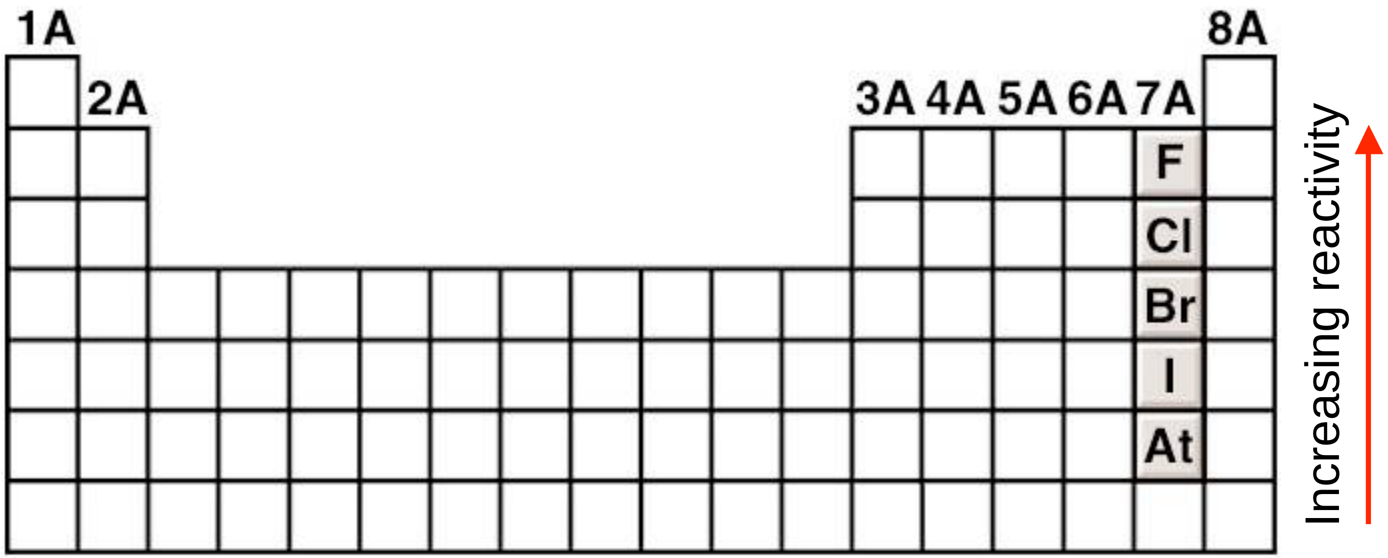
$$SO_{3(g)} + H_2O_{(l)} \longrightarrow H_2SO_{4(aq)}$$



Group 7A Elements (ns^2np^5, $n \geq 2$)

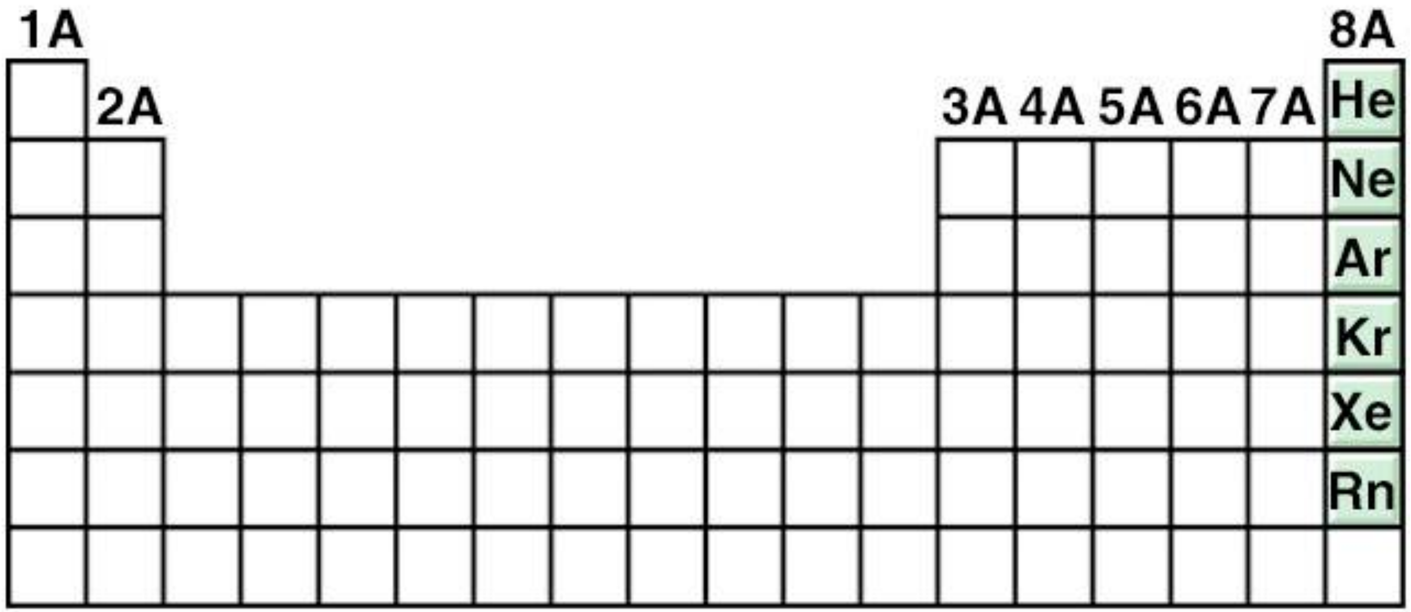
$$X + 1e^- \longrightarrow X^{-1}$$

$$X_{2(g)} + H_{2(g)} \longrightarrow 2HX_{(g)}$$

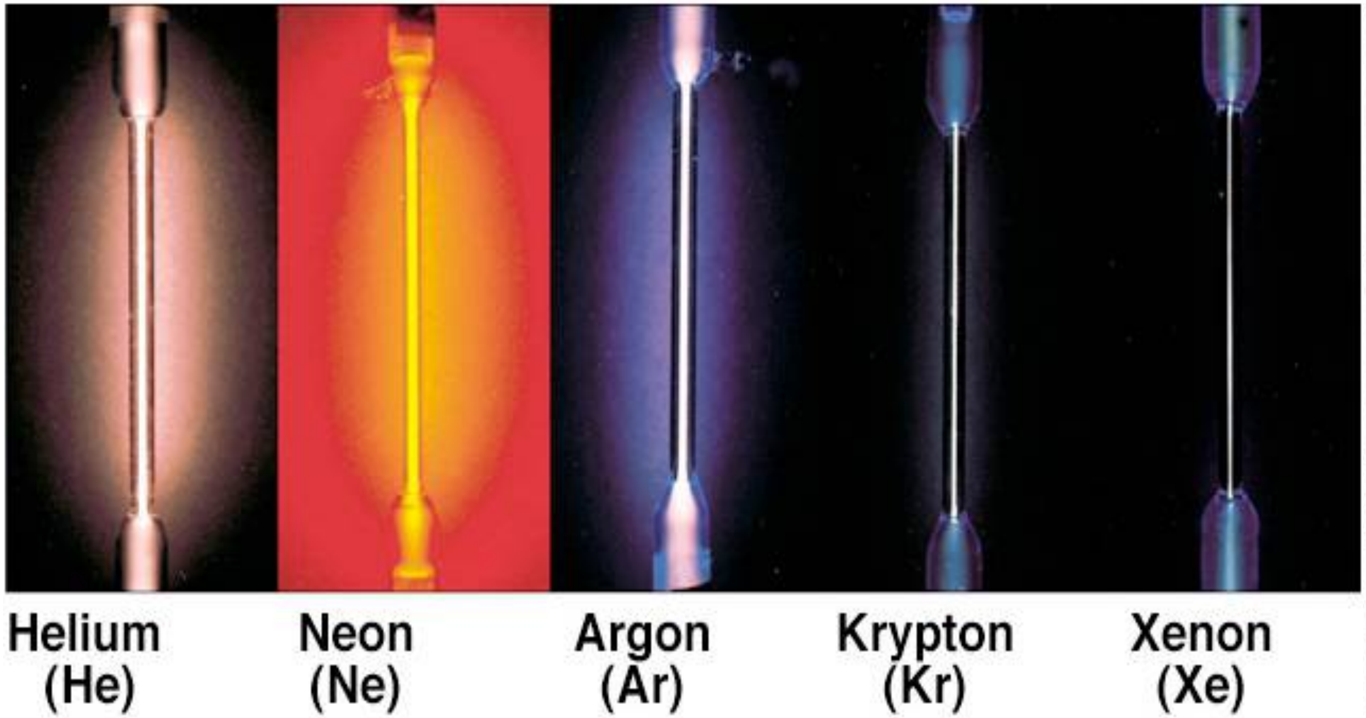


Group 8A Elements (ns^2np^6, $n \geq 2$)

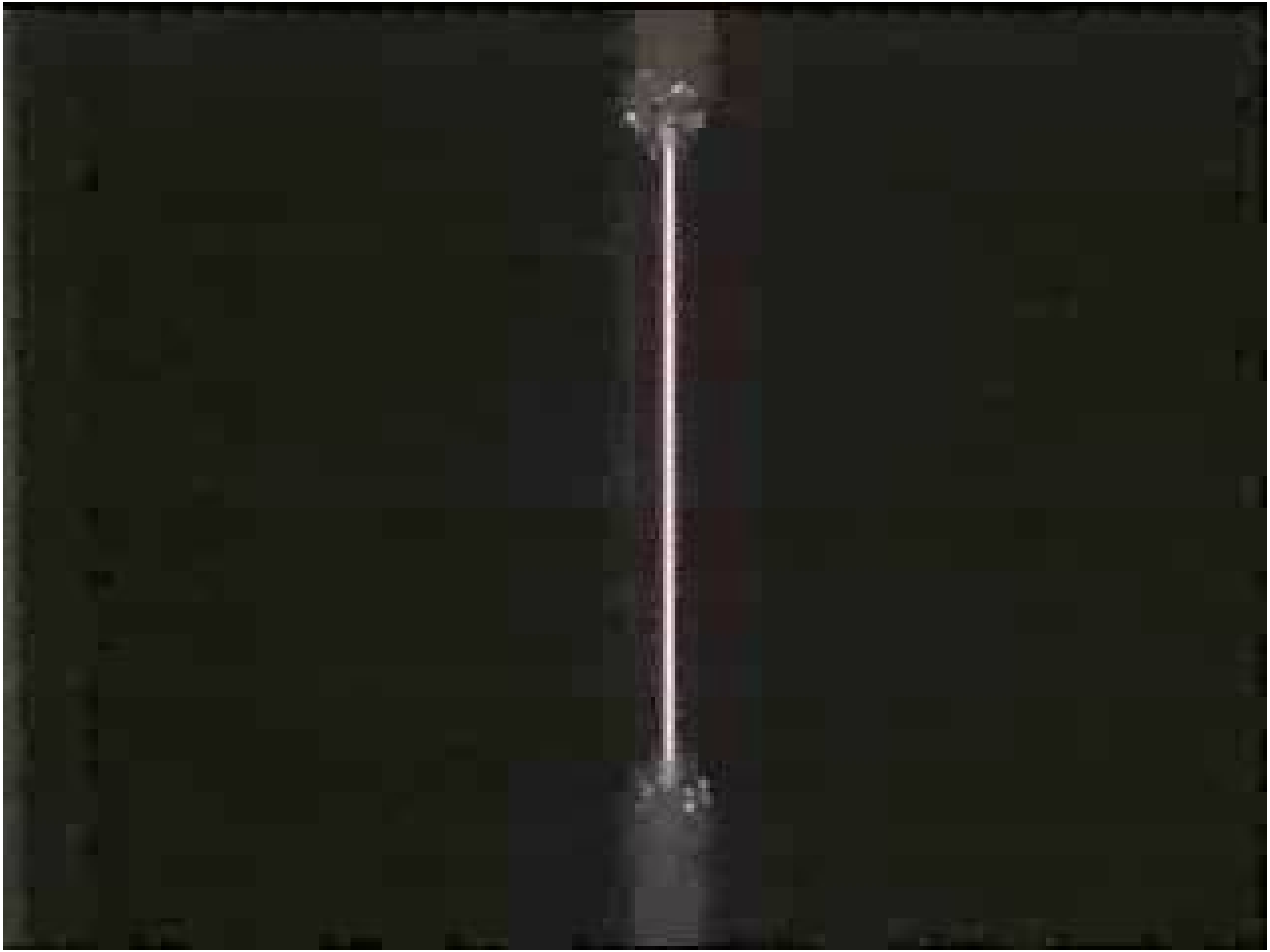
Completely filled *ns* and *np* subshells.
Highest ionization energy of all elements.
No tendency to accept extra electrons.



Chemistry in Action: Discovery of the Noble Gases



Sir William Ramsay



Noble gases are ignored

Electron affinity is energy associated with an atom gaining an electron. It is highest in the top right where atoms are smallest with the greatest number of protons

Electronegativity is a number that describes the relative ability of an atom (when bonded) to attract electrons. The trend is the same as affinity for the same reason

Properties of Oxides Across a Period

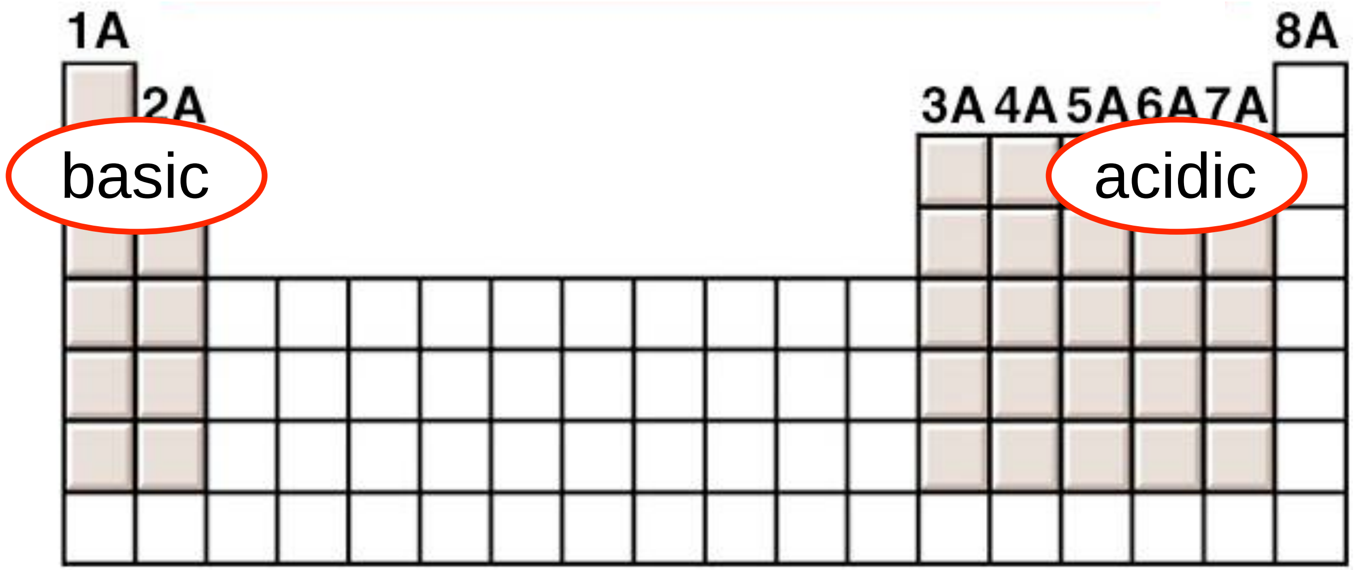


TABLE 8.4 Some Properties of Oxides of the Third-Period Elements

	Na_2O	MgO	Al_2O_3	SiO_2	P_4O_{10}	SO_3	Cl_2O_7
Type of compound	←	Ionic	→	←	Molecular		→
Structure	←	Extensive three-dimensional		→	← Discrete molecular units		→
Melting point (°C)	1275	2800	2045	1610	580	16.8	−91.5
Boiling point (°C)	?	3600	2980	2230	?	44.8	82
Acid-base nature	Basic	Basic	Amphoteric	←	Acidic		→