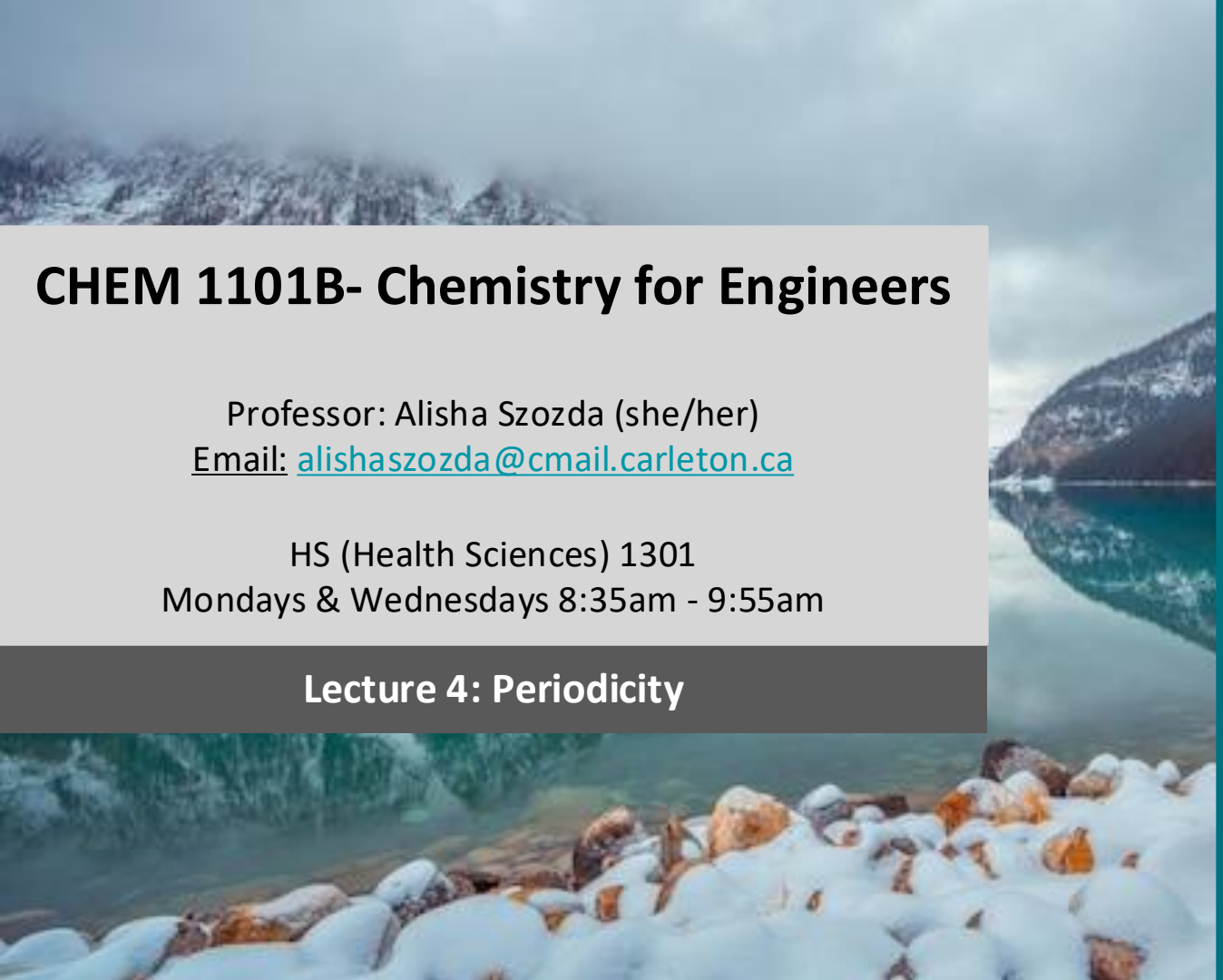


CHEM 1101B- Chemistry for Engineers

Professor: Alisha Szozda (she/her)
Email: alishaszozda@cmail.carleton.ca

HS (Health Sciences) 1301
Mondays & Wednesdays 8:35am - 9:55am

Lecture 4: Periodicity

Please feel free to introduce yourself to your neighbors– name, pronouns, a hobby, etc.

and/or

Answer the first question on Wooclap!

Learning outcomes Topic 4:
The ATOM – Periodicity

- Relate chemical and physical properties to electron configurations
- Predict relative atomic radius, ionization energies and electronegativities

Explaining the Periodic Table

Recall: valence electrons are electrons in the outermost shell (outermost principle quantum number) of an atom

- Elements in the same group have the same valence electron configuration
- Elements with the same valence electron configuration show similar chemical behaviour
- Electron configurations can be used to explain periodic trends

i.e.
Why is the **atomic radius** of potassium larger than that of sodium?

Why is the **first ionization energy** higher in beryllium than in lithium?

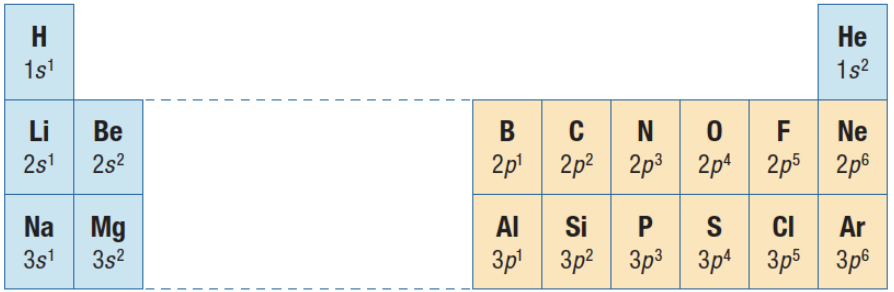
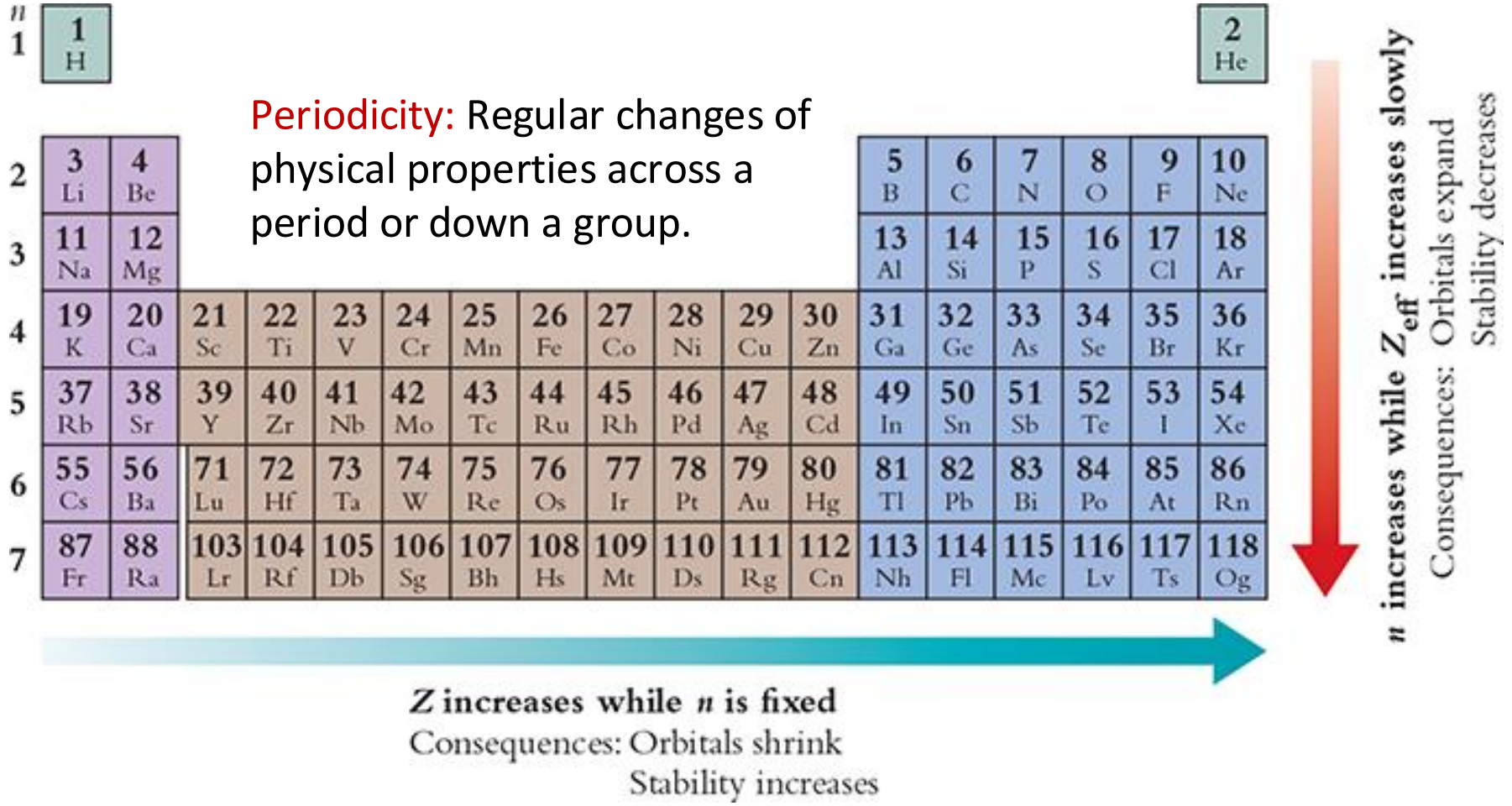


Figure 8 The valence electron configurations for the first 18 elements

Underlying Patterns of Periodicity



Recap: Underlying Patterns of Periodicity

Patterns are dependent on the:

(1) principal quantum number, $\boldsymbol{n}$

(2) atomic number, $\boldsymbol{Z}$

Across a row in the periodic table (left to right)

- $\boldsymbol{n}$ (number of shells) stays the same, $\boldsymbol{Z}$ increases
- Positive charge in the nucleus (nuclear charge) increases, pulling the outer valence electrons closer to the nucleus (stronger attraction), reducing the atom's size
- Electrons closer to the nucleus are energetically more stable

Down a column the periodic table

- $\boldsymbol{n}$ (number of shells) and $\boldsymbol{Z}$ both increase, however, $\boldsymbol{n}$ is the dominating factor for determining orbital size and stability within a column
- The atom is continuously becoming bigger as the valence electrons occupy a level that is farther from the nucleus
- Valence orbitals get larger and less stable

General Trend #1: Atomic Radius/Size

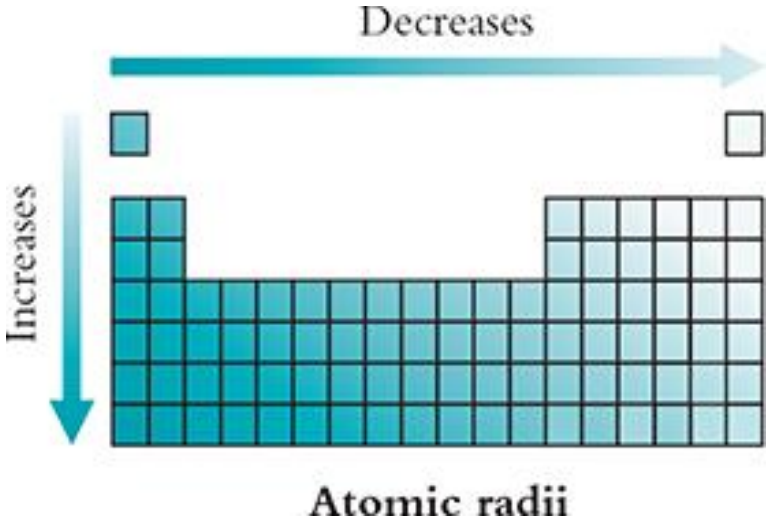
Atomic radius: The distance from the nucleus to the approximate outer boundary of electrons

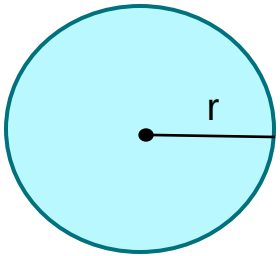
What dictates atomic radius/size?

- Outer valence electrons

- Moving down a group, the atomic radius ↑
  - As n ↑, atomic/valence orbitals become **larger** and less stable

- Across a period (left to right), the atomic radius ↓
  - As Z ↑, atomic orbitals become **smaller** and more stable





Atomic Radii/Size: Valence Electrons

		Z	$n_{e\text{- core}}$	$n_{e\text{- valance}}$
$_{3}Li$:	$1s^2, 2s^1$	3	2	1
$_{4}Be$:	$1s^2, 2s^2$	4	2	2
$_{5}B$:	$1s^2, 2s^2, 2p^1$	5	2	3
$_{6}C$:	$1s^2, 2s^2, 2p^2$	6	2	4
$_{7}N$:	$1s^2, 2s^2, 2p^3$	7	2	5
$_{8}O$:	$1s^2, 2s^2, 2p^4$	8	2	6
$_{9}F$:	$1s^2, 2s^2, 2p^5$	9	2	7
$_{10}Ne$:	$1s^2, 2s^2, 2p^6$	10	2	8
$_{11}Na$:	$1s^2, 2s^2, 2p^6, 3s^1$	11	10	1

Example: Trends in Atomic Radii/Size

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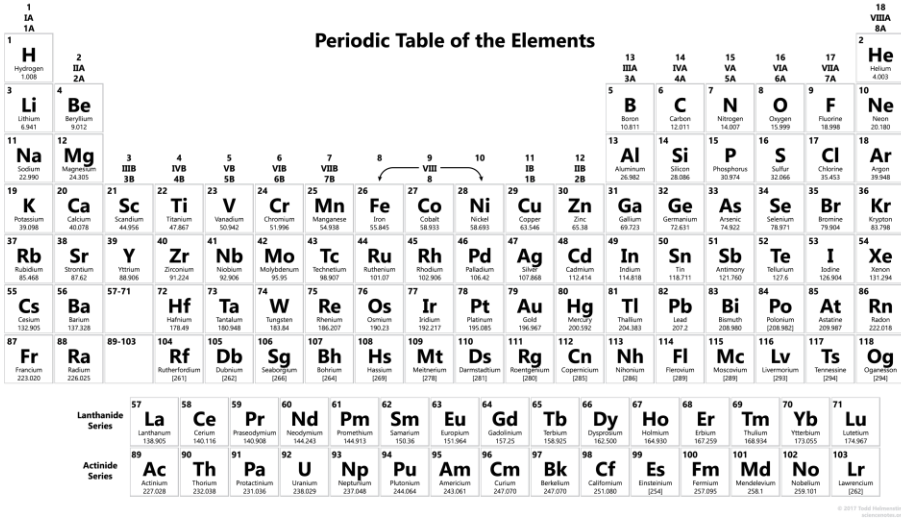
For each of the following, rank in order of increasing atomic size (smallest to biggest)

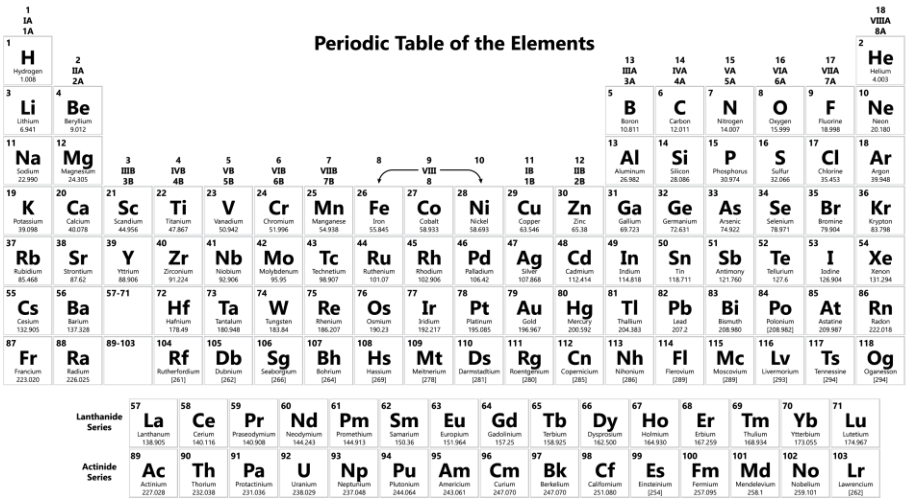
a) Mg, Si, S

b) As, N, P



What to do when both rules apply?

Arrange the following atoms in order of increasing size (smallest to biggest): **Al, C, S**

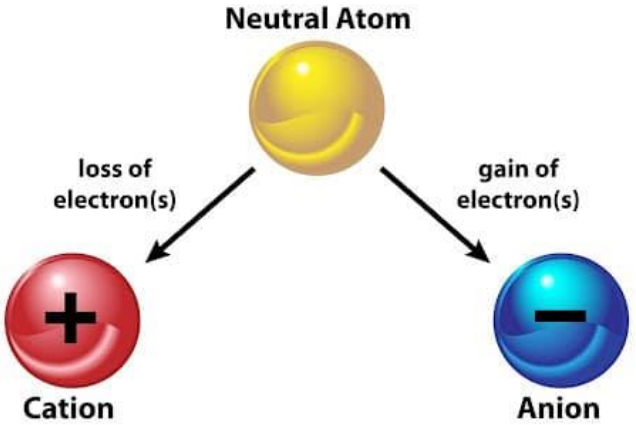


Recap: Ions

- Atoms tend to form **ions** (a neutral atom that gains or gives up electrons) in an attempt to achieve a **stable octet**
  - **Cations**— positive charged
  - **Anions**— negative charged

- **Metals** in the main group elements tend to **give up** electrons forming **cations** that have the same number of electrons as the nearest **noble gas**

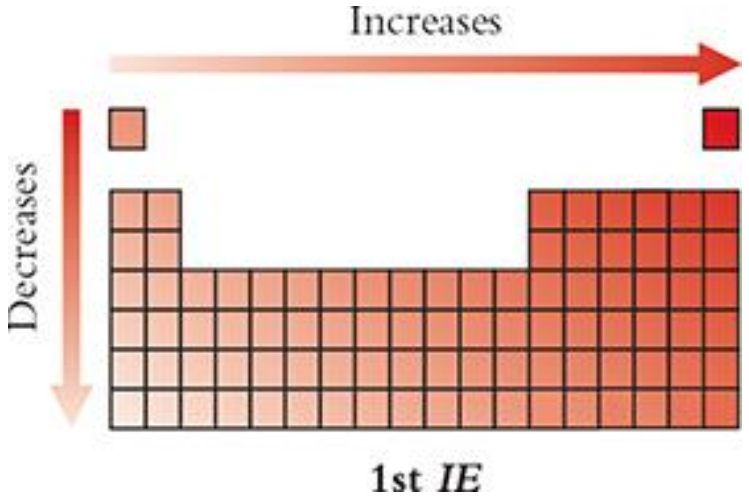
- **Non-metals** tend to **gain** electrons and form **anions** that have the same number of electrons as the nearest **noble gas**



General Trend: #2 Ionization Energy (IE)

Ionization E: The amount of energy required to remove an electron

- Moving down a group, the **IE ↓**
  - As n ↑, valence electrons are farther from the positive attractive force of the nucleus
  - **Less energy** is required to remove the electron (easier)



- Across a period (left to right), the **IE ↑**
  - As Z ↑, the attraction between the nucleus and valence electrons increases
  - **More energy** is required to remove the electron (harder)

2nd and 3rd Ionization Energies

IE_1= Removing the 1st electron
IE_2= Removing the 2nd electron
IE_3= Removing the 3rd electron

Process	Configurations	*IE*
$Mg(g) \rightarrow Mg^+(g) + e^-$	$[Ne]3s^2 \rightarrow [Ne]\ 3s^1$	739 kJ/mol
$Mg^+(g) \rightarrow Mg^{2+}(g) + e^-$	$[Ne]\ 3s^1 \rightarrow [Ne]$	1450 kJ/mol
$Mg^{2+}(g) \rightarrow Mg^{3+}(g) + e^-$	$[Ne] \rightarrow [He]\ 2s^2\ 2p^5$	7730 kJ/mol

When increasing the positive charge on an atom, this results in:

Increased energy required to remove each electron (more difficulty)

Example: Trends in Ionization Energy

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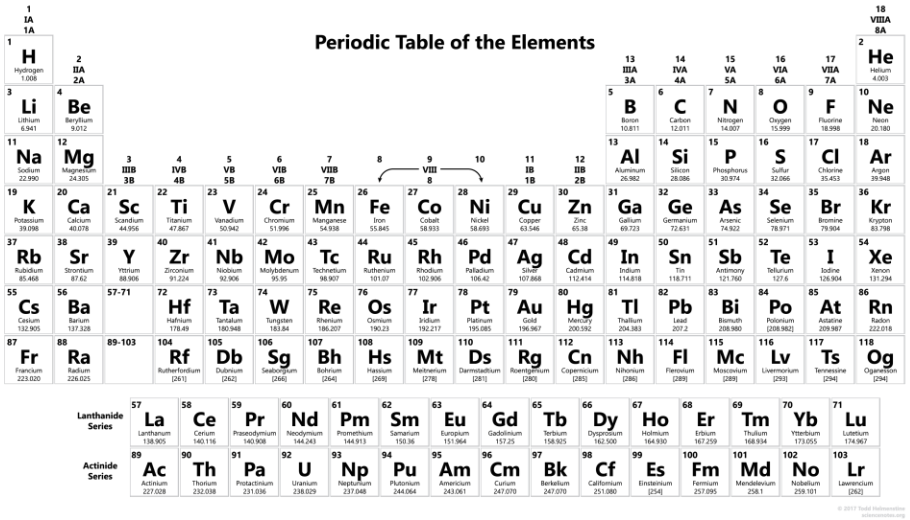
Rank the following atoms in order of increasing first IE (smallest to biggest)

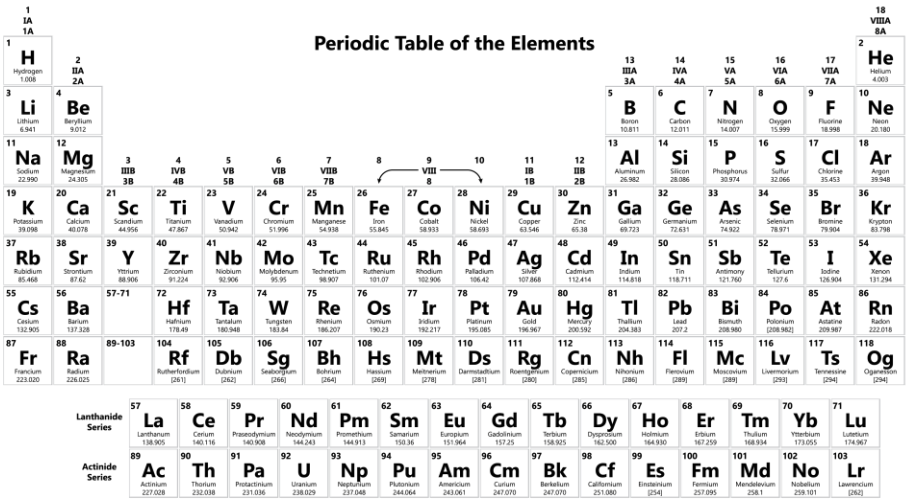
a) F, Li, C

b) Be, Ba, Mg



What to do when both rules apply?

Arrange the following atoms in order of decreasing first ionization energy: **Ar, Cl, Cs, K**



Exceptions to Ionization Energy Trend

Stability of electron configurations influences Ionization Energy

Neutral Atom		Ion: removal of 1 electron	
$_3$Li:	$1s^2, 2s^1$	Li^+:	$1s^2$
$_4$Be:	**$1s^2, 2s^2$**	**Be^+:**	**$1s^2, 2s^1$**
$_5$B:	**$1s^2, 2s^2, 2p^1$**	**B^+:**	**$1s^2, 2s^2$**
$_6$C:	$1s^2, 2s^2, 2p^2$	C^+:	$1s^2, 2s^2$
$_7$N:	**$1s^2, 2s^2, 2p^3$**	**N^+:**	**$1s^2, 2s^2, 2p^2$**
$_8$O:	**$1s^2, 2s^2, 2p^4$**	**O^+:**	**$1s^2, 2s^2, 2p^3$**

IE: Li < B < Be < C < O <N

Increasing I.E.

General Trend: #3 Electron Affinity

Electron affinity: the energy change when an electron is added to the atom

- Reverse principle of Ionization energy but follows the **same trend**

$$F \rightarrow F^+ + e^- \quad IE_1 = 1681\ kJ/mol$$
$$F + e^- \rightarrow F^- \quad EA = -322\ kJ/mol$$

- Moving down a group, the **electron affinity** ↓
  - As n ↑, there is a greater distance to the nucleus
  - Less attraction for incoming electrons

- Across a period (left to right), the **electron affinity** ↑
  - As Z ↑, the attraction between the nucleus and valence electrons increases
  - Stronger attraction for incoming electrons

General Trend: #4 Electronegativity

Electronegativity (EN): the tendency of an atom to pull electrons in a chemical bond toward itself

Only applicable when referring to properties of atoms involved in chemical bonding

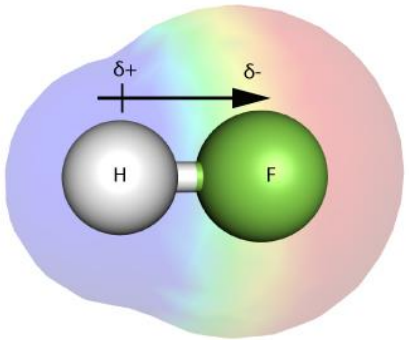
Increasing electronegativity

Decreasing electronegativity

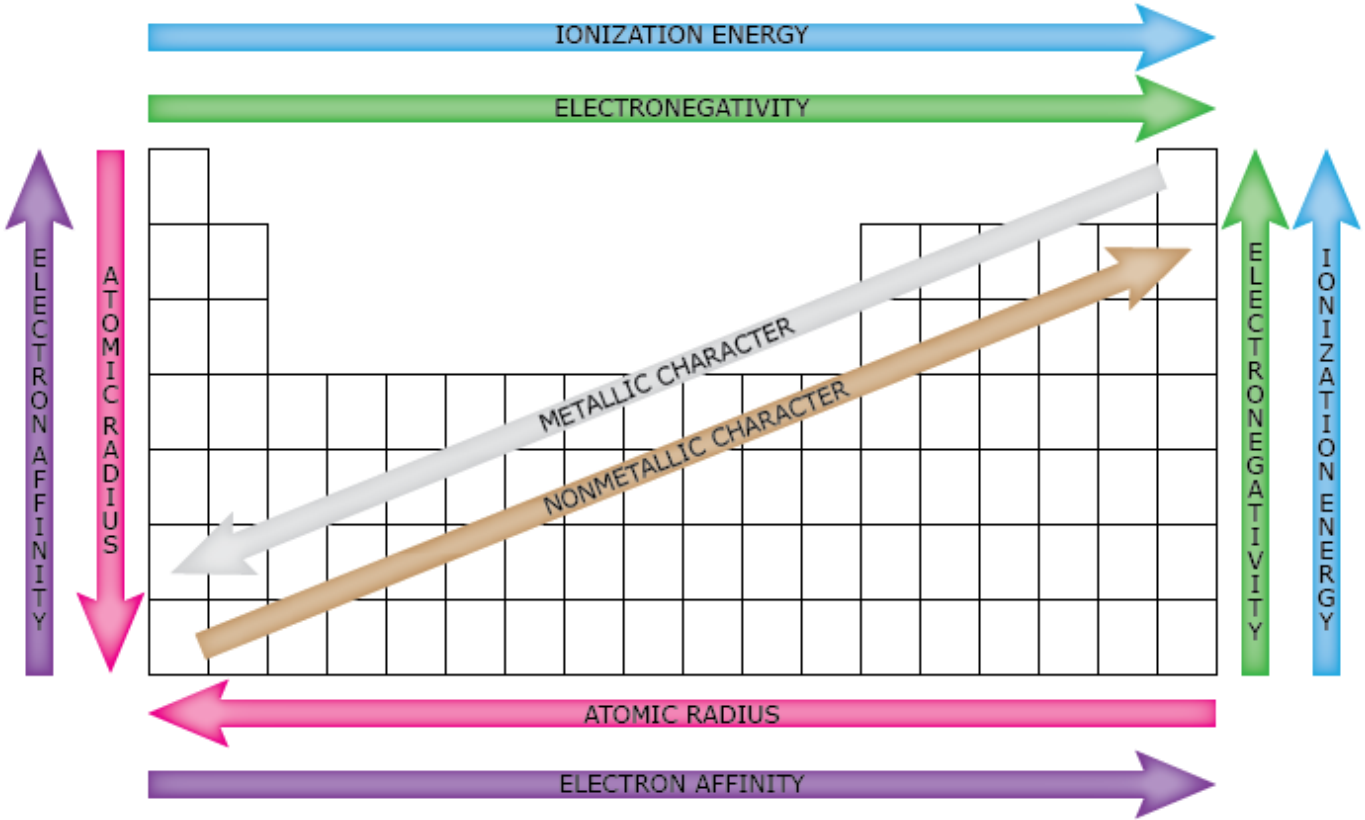
Increases across a row

Decreases down a group

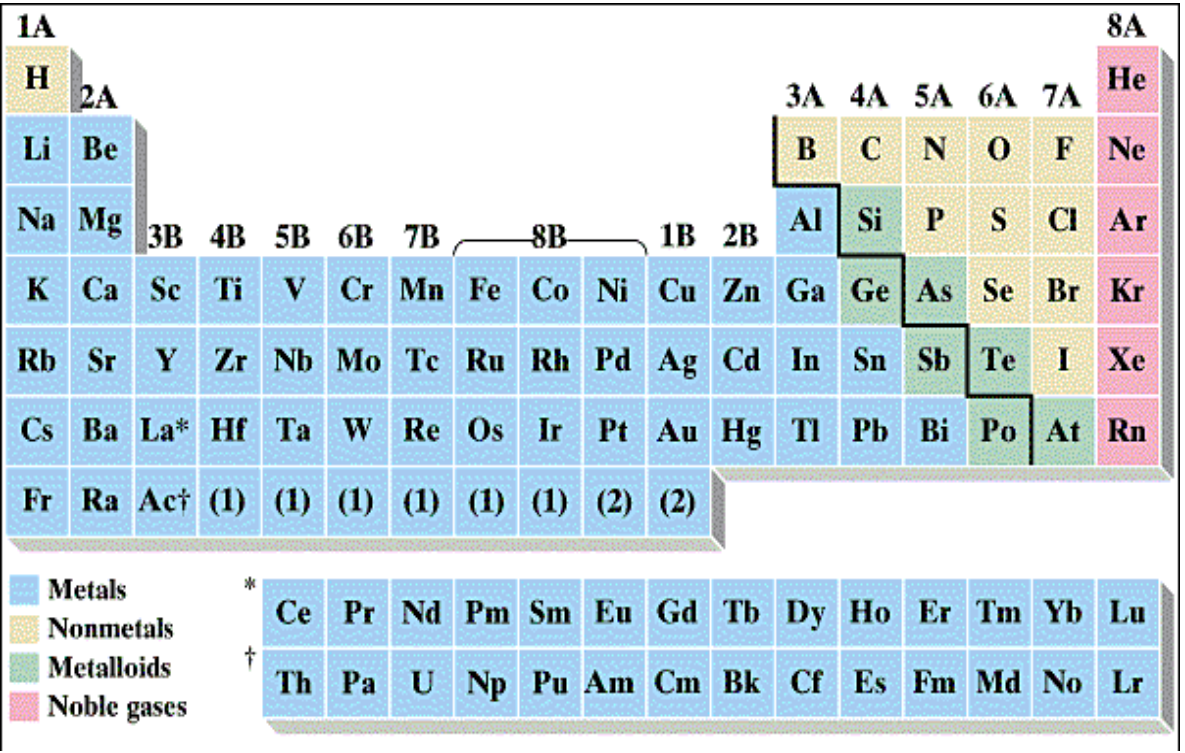
H 2.1																
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.9	Ni 1.9	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5
Cs 0.7	Ba 0.9	La–Lu 1.0–1.2	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.9	Bi 1.9	Po 2.0	At 2.2
Fr 0.7	Ra 0.9	Ac 1.1	Th 1.3	Pa 1.4	U 1.4	Np–No 1.4–1.3										



Overview of Periodic Table Trends



Rules for forming ions



Move electrons onto or off of the atom in order to:

- Fill the entire valence s and p subshells
- Empty the entire valence s and p subshells
- Empty the valence p subshell – only if n ≥ 4

NEVER add more than 3 electrons

NEVER remove more than:

3 e^- from small atoms (n ≤ 3)

4 e^- from large atoms (n ≥ 4)

Metals **NEVER** form negative ions

Rules for forming ions

When forming cations, *electrons are always lost **first** from the VALENCE subshells*, and only then will they be lost from what were originally inner subshells

Transition metals (d block elements) do not always follow predictable patterns for what stable ions they form (this is NOT the oxidation state).

- Most will form +2 ions M^{2+} or M(II)
- Some will form +1 ions M^{1+} or M(I)
- Some will form +3 ions M^{3+} or M(III)

Example: $_{23}V$: [Ar] $4s^2\ 3d^3$

$_{23}V(III)$: [Ar] $4s^0\ 3d^2$

Example(s): Predict the most likely ion(s) to form for an atom

Circle the possible ions for each atom

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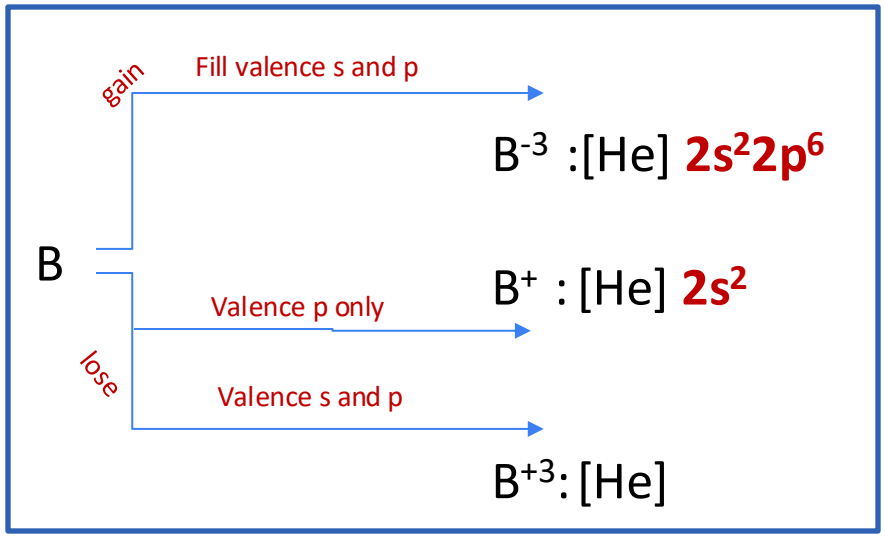
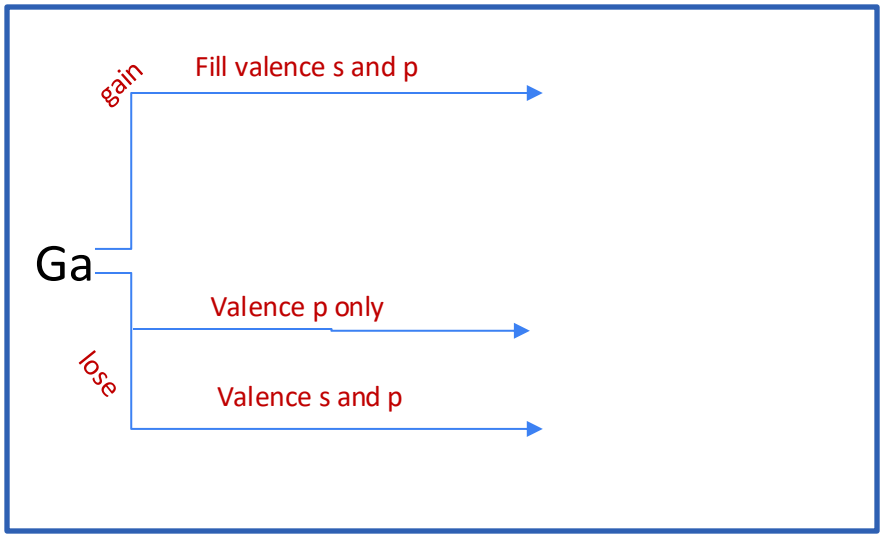
B: [He] $2s^2$ $2p^1$



1. Fill the entire valence s and p subshells
2. Empty the entire valence s and p subshells
3. Empty the valence p subshell – only if n ≥ 4

Ga: [Ar] $4s^2$ $3d^{10}$ $4p^1$



NEVER add more than 3 electrons

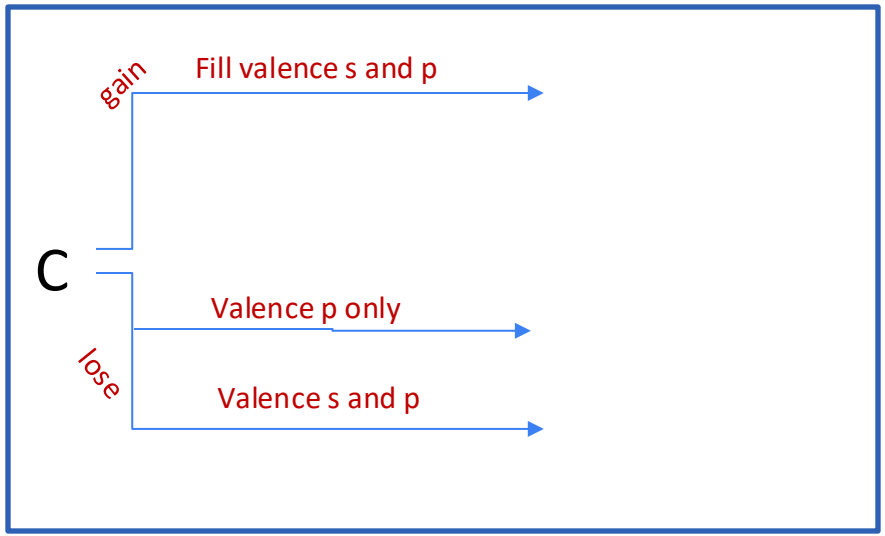
NEVER remove more than: 3 ($n \leq 3$) or 4 ($n \geq 4$)

Metals **NEVER** negative

Extra: Predict the most likely ion(s) to form for an atom

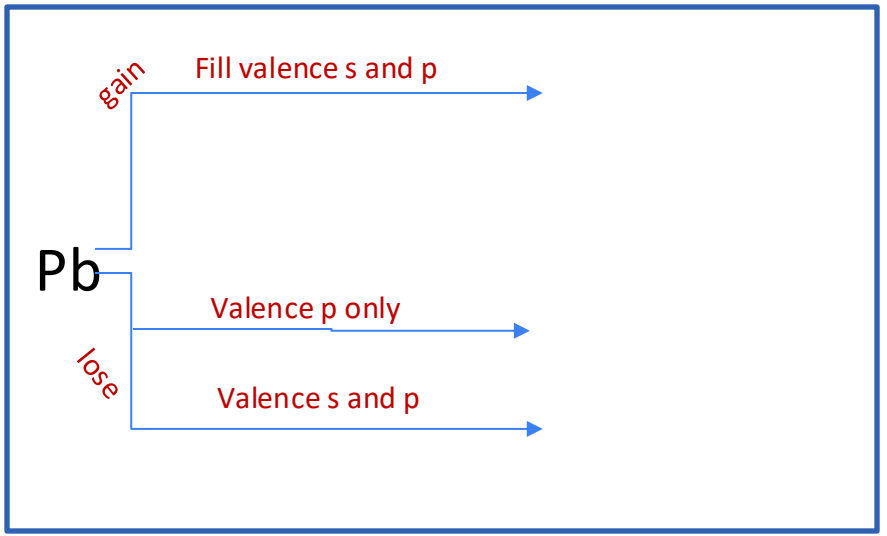
Circle the possible ions for each atom

C: [He] $\mathbf{2s^2\ 2p^2}$



1. Fill the entire valence s and p subshells
2. Empty the entire valence s and p subshells
3. Empty the valence p subshell – only if n ≥ 4

Pb: [Xe] $\mathbf{6s^2}$ $5d^{10}$ $4f^{14}$ $\mathbf{6p^2}$



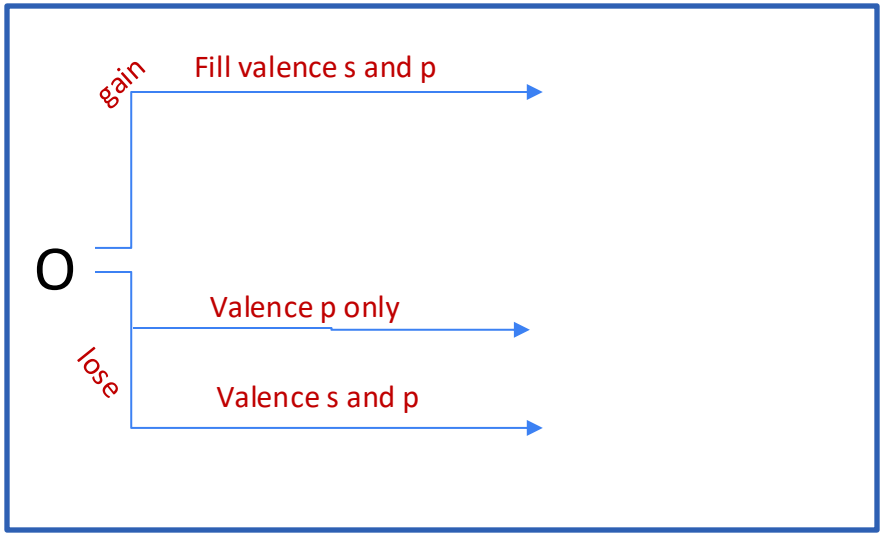
NEVER add more than 3 electrons

NEVER remove more than: 3 ($n \leq 3$) or 4 ($n \geq 4$)

Metals **NEVER** negative

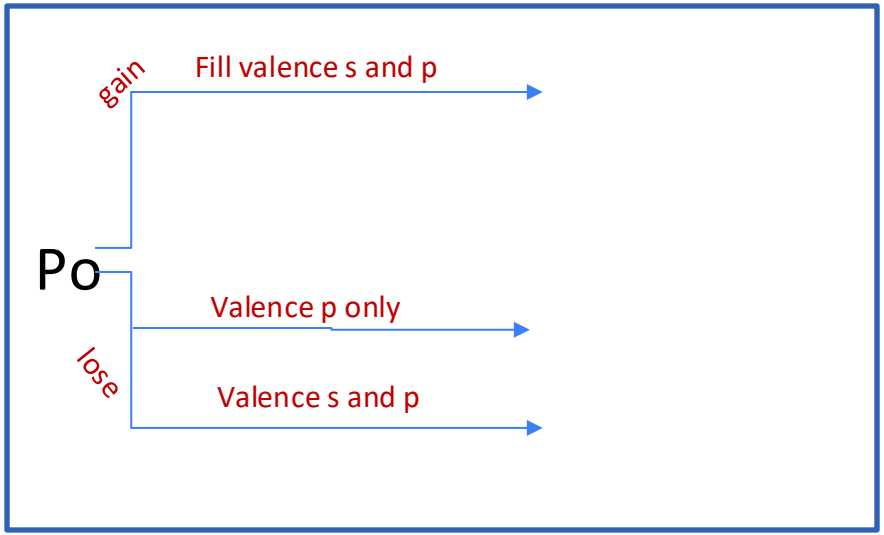
Extra: Predict the most likely ion(s) to form for an atom

O: [He] $2s^2 2p^4$



1. Fill the entire valence s and p subshells
2. Empty the entire valence s and p subshells
3. Empty the valence p subshell – only if n ≥ 4

Po: [Xe] $6s^2\ 5d^{10}\ 4f^{14}\ 6p^4$



NEVER add more than 3 electrons
NEVER remove more than: 3 (n≤3) or 4 (n≥4)
Metals **NEVER** negative

Writing electron configurations for Ions

1. Write configuration for neutral atom & count the number of electrons
2. If ion is negative, add electrons until ion's negative charge is reached
3. If ion is positive, remove electrons until positive charge reached
4. Fill orbitals to match the nearest noble gas of smaller atomic number
5. Add remaining electrons to the next filling orbitals according to Hund's rule
    - For neutral atoms and anions, place electrons in ns before (n-1)d
    - For cations, place electrons in (n-1)d before ns

Note: atoms and ions that have the same number of electrons are **isoelectronic**

e.g. Cr^{3+} configuration?