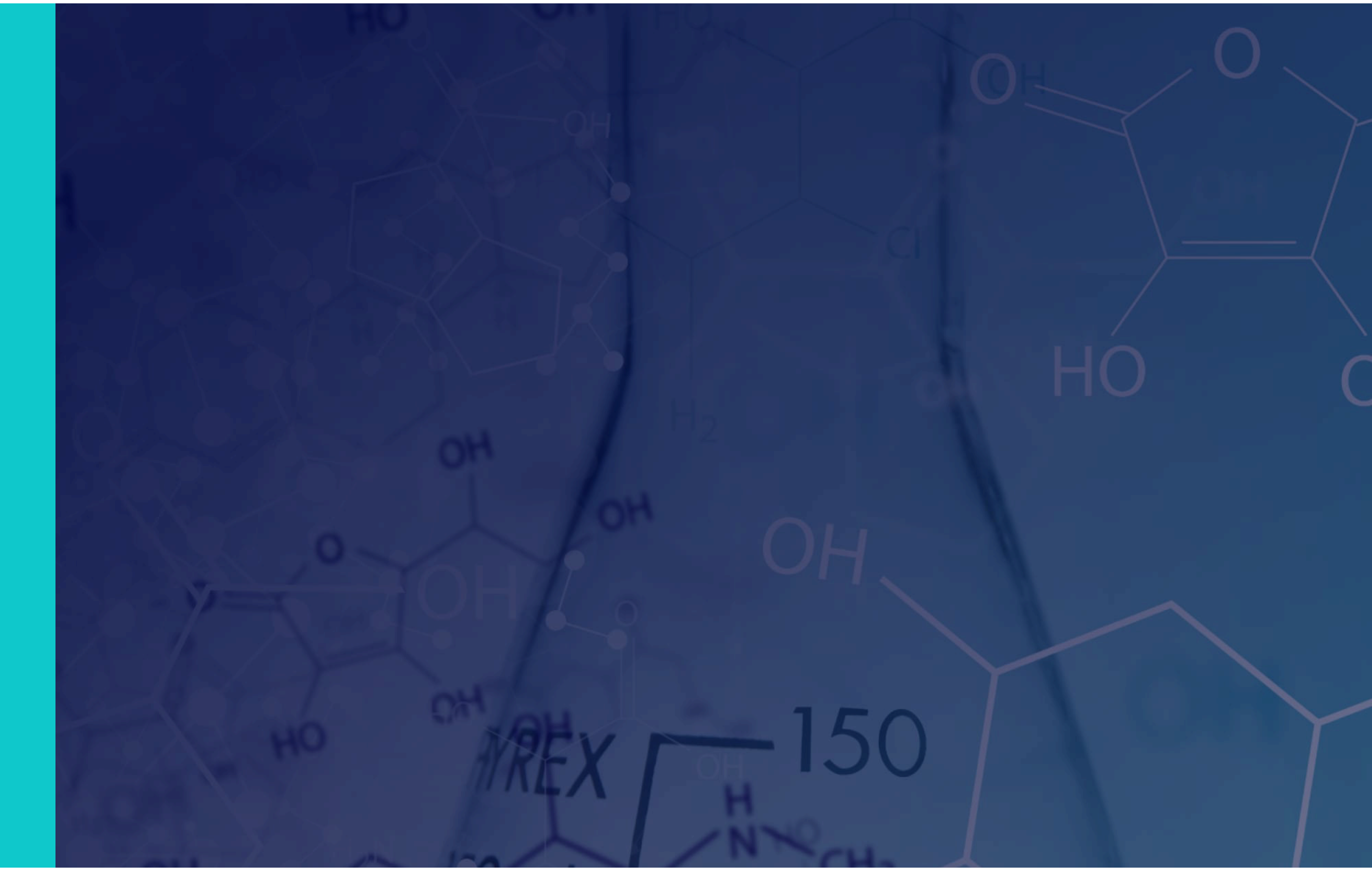


McGill

CHEM 110

Fall 2024, Chapter 2 Notes

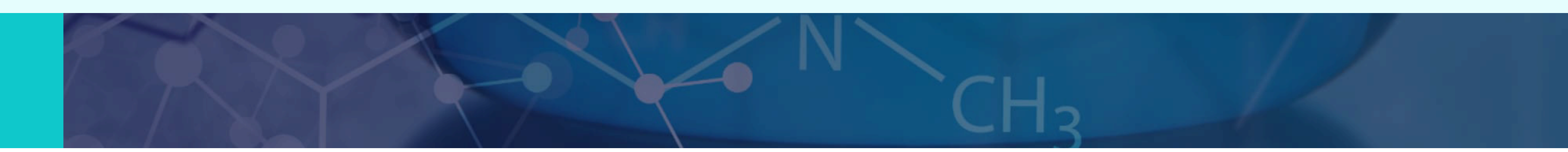


Table of Contents

2. Electron Configuration and Chemical Periodicity

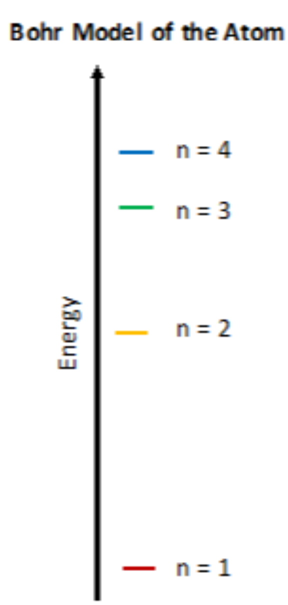
2.1 Orbital Filling Diagrams

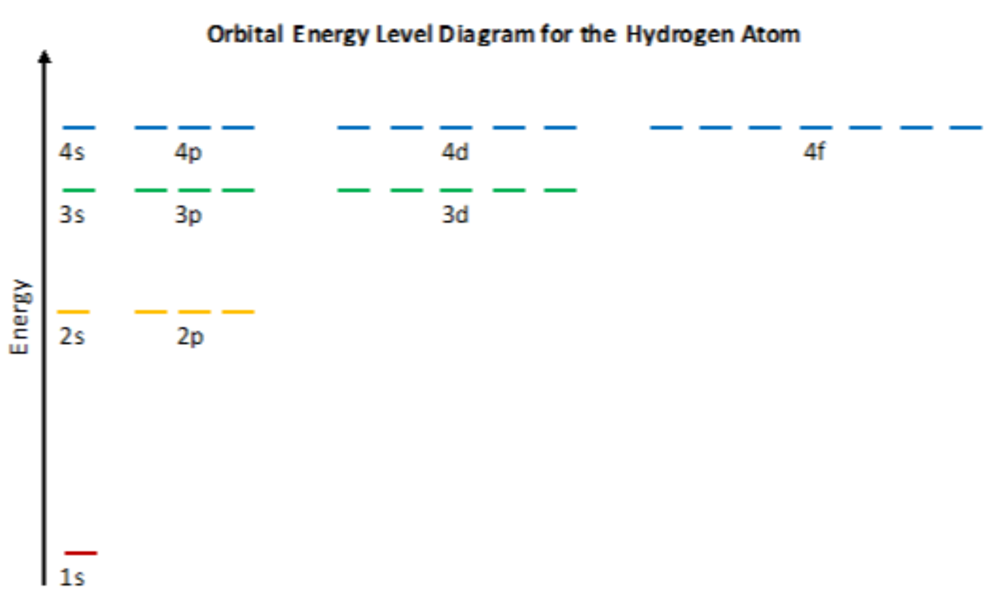
2.1.1

Relative Energies of Atomic Orbitals

One-Electron Species

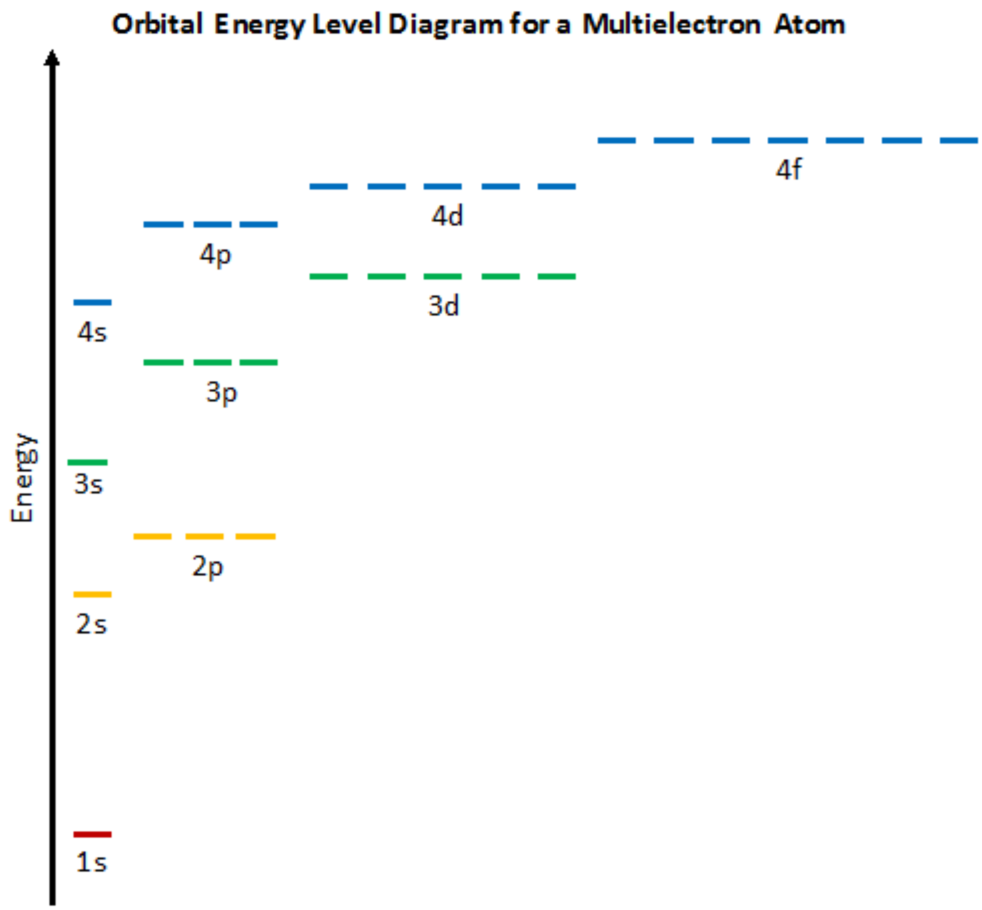
- For a one-electron species like the hydrogen atom, there is no electron-electron repulsion. Therefore, all the subshells of the same n are **degenerate.**
- Orbitals at the **same energy level** are called **degenerate.**





Multi-Electron Species

- With multi-electron species, there are interactions between electrons (electron-electron repulsion)
- The energy level diagram of the orbitals looks like this:



WIZE CONCEPT

- S subshells hold a maximum of ______ electrons
- P subshells hold a maximum of ______ electrons
- D subshells hold a maximum of ______ electrons
- F subshells hold a maximum of ______ electrons

$$\boxed{\text{The total number of orbitals in a given shell is given by : } n^2}$$

Example:

If n=2, what is the total number of orbitals in that shell?

- ________, circle the _____ orbitals that have n=2 above

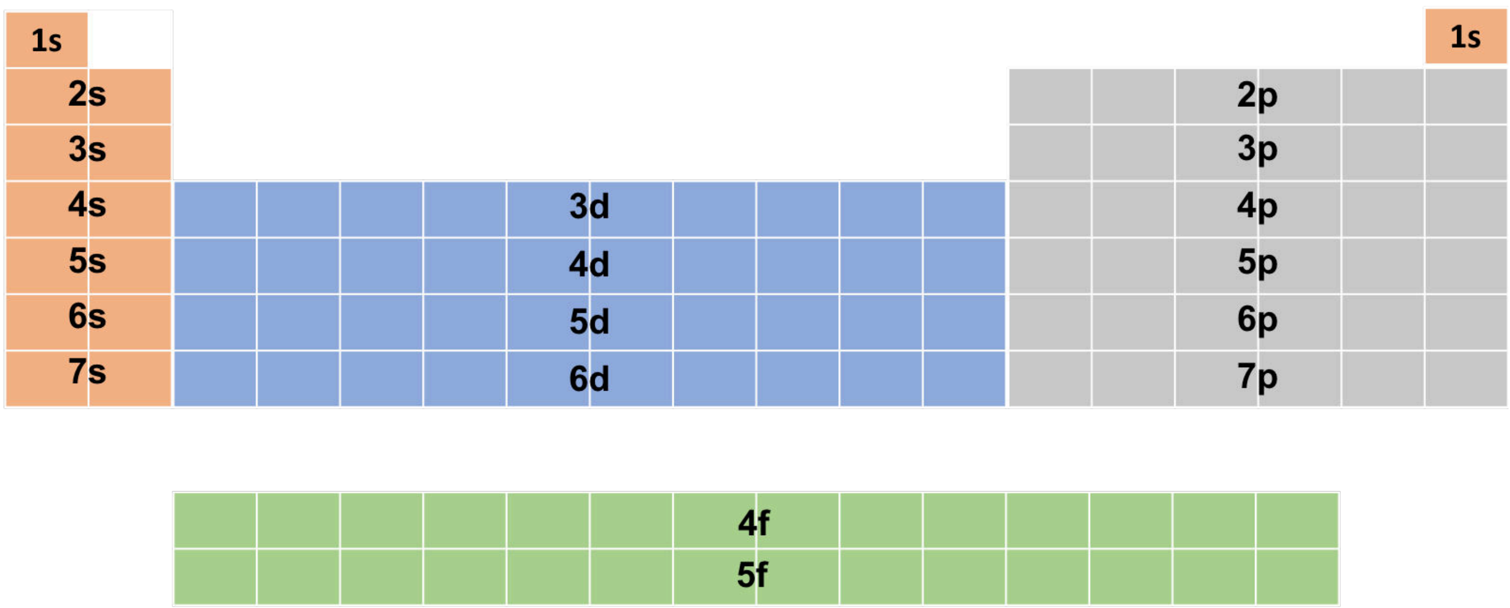
What if we wanted to know the maximum number of electrons we could have if n=2?

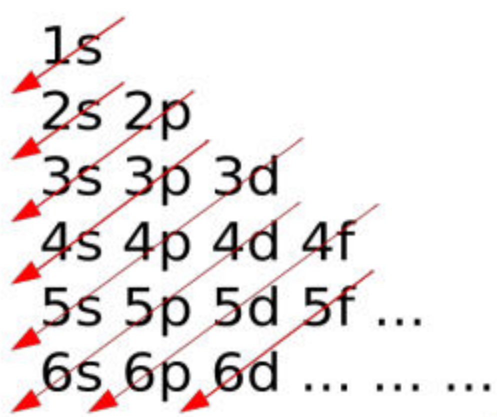
- We already found that total number of orbitals for this shell (_____)
- What is the maximum number of electrons we could have in each orbital? _____
- Therefore the maximum number of electrons possible when n=2 is _____

Rules for Orbital Filling

1) Aufbau Principle

Electrons will always occupy the **lowest available energy level first.**





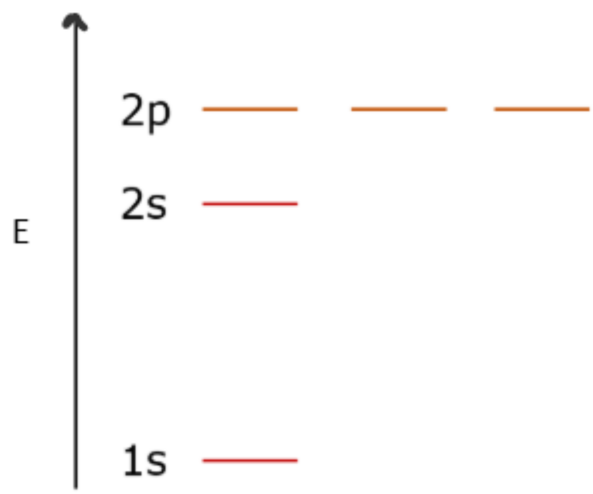
> **WATCH OUT!**
>
> **Memorize the order of orbital filling (or just familiarize yourself enough with the periodic table to know the order!):** 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p...

2) Hund's Rule

Due to electron-electron repulsion, electrons will **fill orbitals of the same energy singly before pairing up**

Electrons don't want to be next to each other unless they have to be!

Example: Fill out the following orbital diagram for C



3) Pauli Exclusion Principle

No two electrons in an atom will have the same set of **4 quantum numbers.**

Quantum Number	**Values**	**Interpretation**
m_s (the spin quantum number)	-1/2 or +1/2	An electron behaves like a magnet that has one of two possible orientations, aligned either with the magnetic field or against it

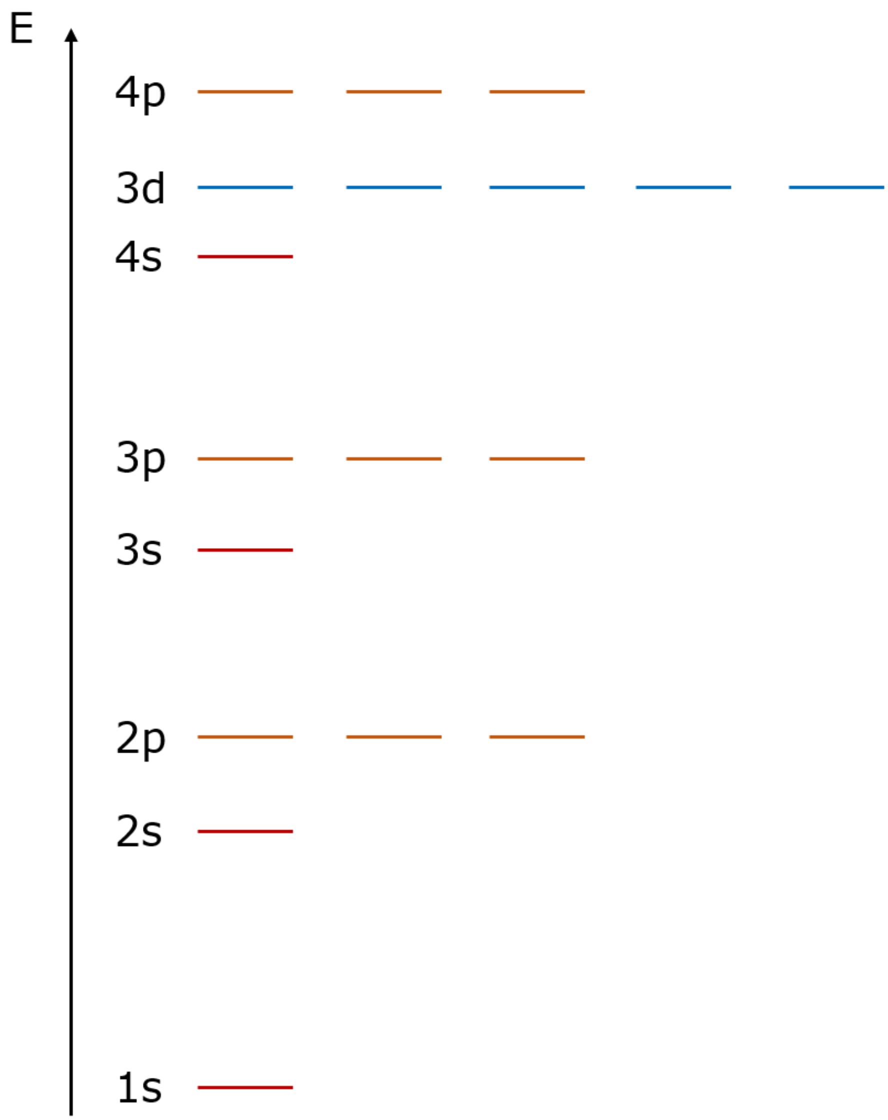
2.1.3

Example: Orbital Filling Diagrams

Draw the orbital diagram for oxygen.

Example: Orbital Filling Diagrams

Complete the atomic orbital diagram below for a neutral calcium atom.

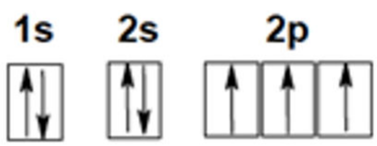


2.1.5

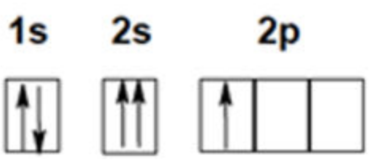
Practice: Orbital Filling Diagrams

Which of the orbital diagrams gives the correct electron configuration for an atom of boron, in the ground state?

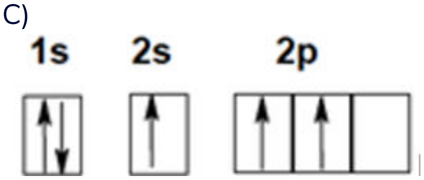
A)



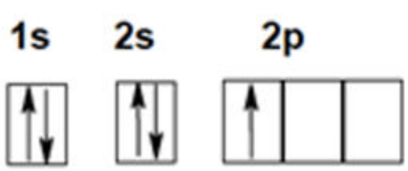
B)



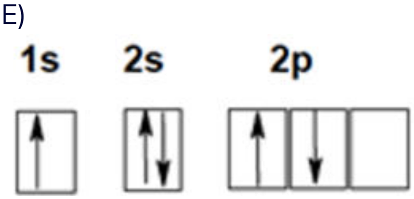
C)



D)



E)



2.1.6

Practice: Number of Orbitals

How many different atomic orbitals exist where n = 3?

- 3
- 6
- 9
- 12, which
- 18

2.2 Electronic Configuration for Atoms

2.2.1

Electron Configurations of Atoms

Group → ↓Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	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
5	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
6	55 Cs	56 Ba		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
7	87 Fr	88 Ra		104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Ms	116 Lv	117 Ts	118 Og

Lanthanides	57 La	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	89 Ac	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

- lable blocks
- describe the n number for each block

Simple Electron Configurations

Li:

C:

Ne:

Group → / ↓Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	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
5	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
6	55 Cs	56 Ba		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
7	87 Fr	88 Ra		104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Ms	116 Lv	117 Ts	118 Og

Lanthanides	57 La	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	89 Ac	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

Shorthand Notations

Write the name of the **previous noble gas and then fill in the rest of the electron configuration.**

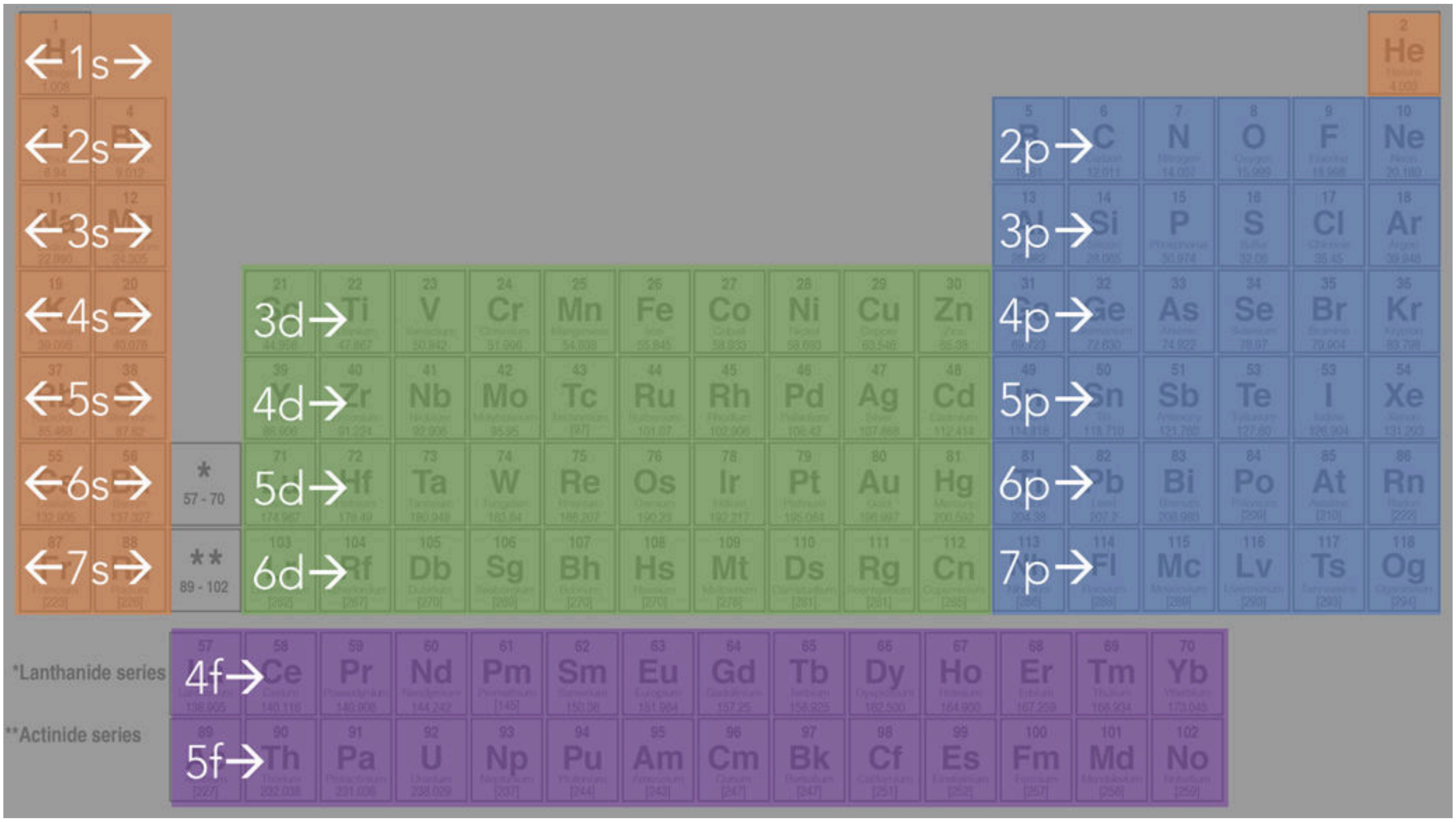
O:

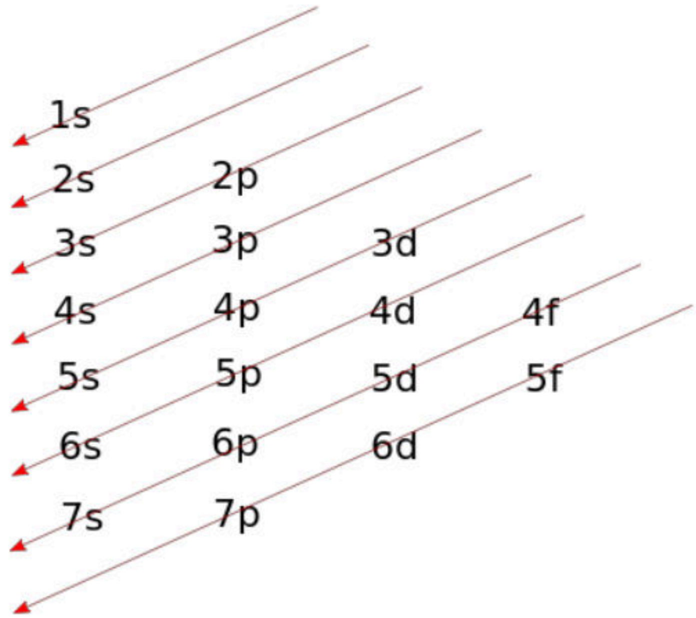
Fe:

Ge:

2.2.2

Electron Configurations Cheatsheet




1s
2s 2p
3s 3p 3d
4s 4p 4d 4f
5s 5p 5d 5f
6s 6p 6d
7s 7p

2.2.3

Electron Configuration Exceptions

Transition metals have their valence electrons in the d subshell.
It is a lot **more favourable** (and stable) to have **d shell either half full or totally full instead of just partially full.**

	Cr:	Cu:
Expected:	4s 3d [Ar] $4s^2 3d^4$	4s 3d [Ar] $4s^2 3d^9$
Seen:	4s 3d [Ar] $4s^1 3d^5$	4s 3d [Ar] $4s^1 3d^{10}$

WIZE TIP

The other exceptions that follow this same trend are: Mo, Ag, and Au! (Note they are in the same groups that Cr and Cu are in!)

Example: Writing Out Electron Configurations

Part 1:

Write the full electronic configurations for the following elements:

P:

Se:

Part 2:

Write the short electron configurations for the following elements:

Zn:

Zr:

The Periodic Table of the Elements

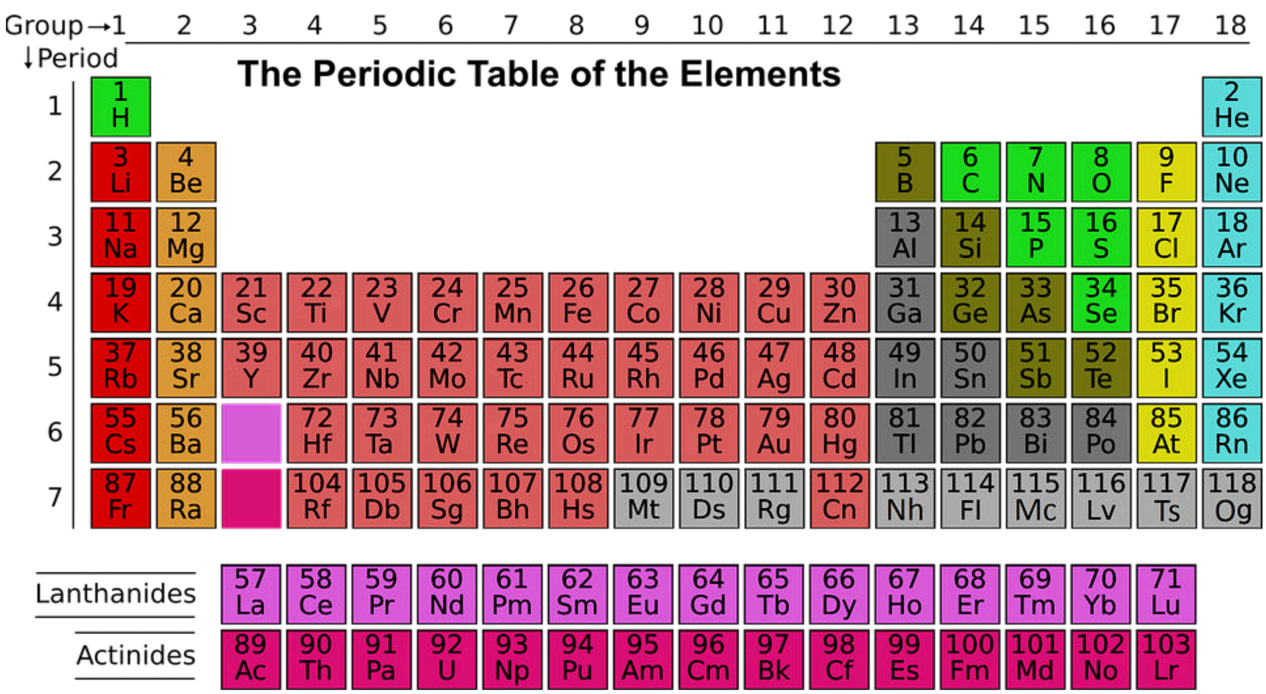
↓Period / Group→	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	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
6	55 Cs	56 Ba		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
7	87 Fr	88 Ra		104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Mc	116 Lv	117 Ts	118 Og

Lanthanides	57 La	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	89 Ac	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

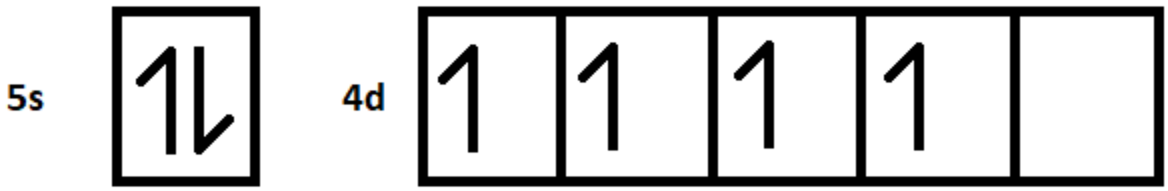
2.2.5

Example: Condensed Electron Configuration

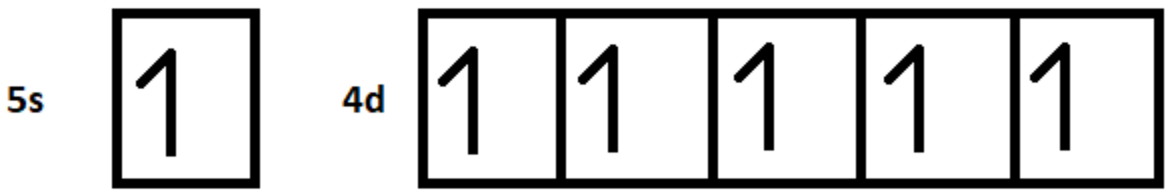
What is the condensed electronic configuration for Mo?



We ignore everything before the previous noble gas (Kr), which means we are only working with the **5s** and **4d** blocks. Let's fill them up:



But there is something special about Mo. It is in group 6. This group tends to **fill (or in this case half-fill) the d-block** at the expense of the s block.



Therefore, the electronic configuration is **[Kr]$5s^1 4d^5$.**

2.2.6

Practice: Identify the Neutral Element

What neutral element is represented in the example below?

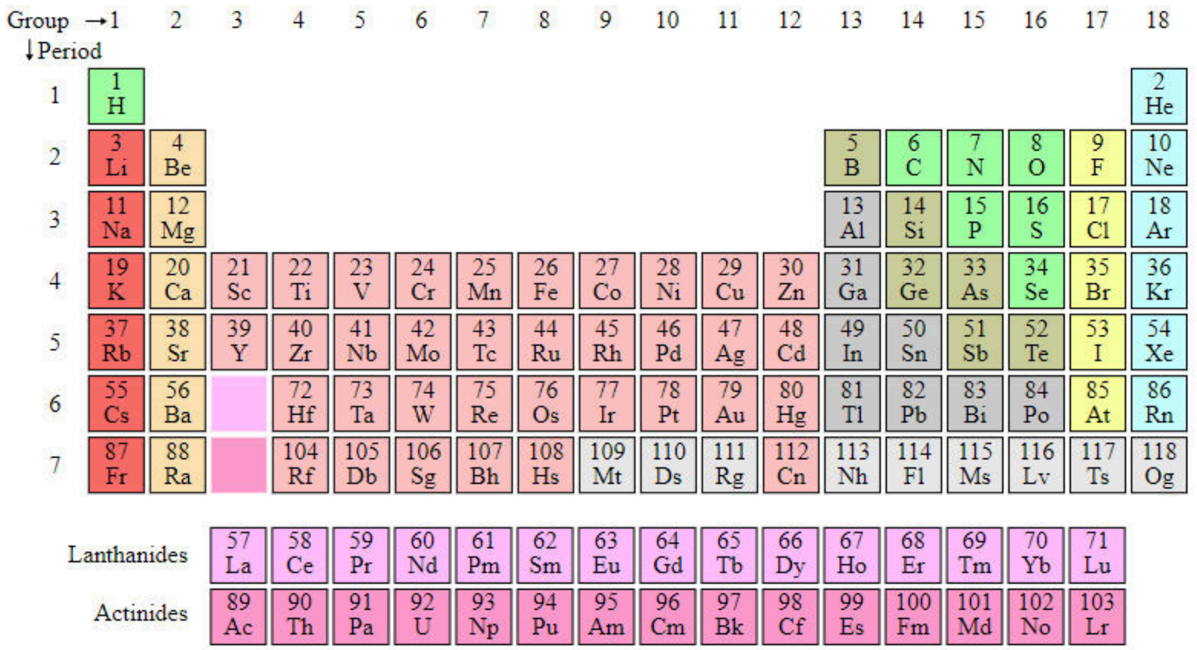
$1s^2 2s^2 2p^6 3s^2 3p^5$ became $[Ne]3s^2 3p^5$

- Al
- Ne
- P
- Cl

2.2.7

Practice: Electron Configuration

The ground state electronic configuration for At is



A) $[Xe]6s^2 4f^{14} 5d^{10} 6p^5$

B) $[Xe]6s^2 5d^{10} 6p^5$

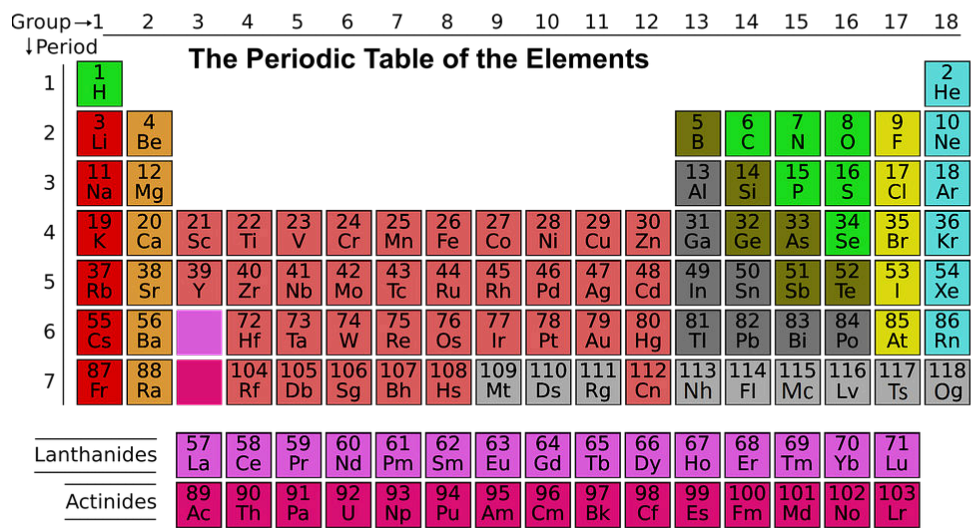
C) $[Xe]6s^2 6p^5$

D) $[Xe]6s^2 4f^{14} 5d^{10} 4p^3$

2.2.8 Practice

Practice: Impossible Electron Configuration

Which of the following electronic configurations is not possible?



- $[Ar]4s^2 4d^6$
- $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$
- $[Kr]5s^1$
- $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^8$

2.3 Electron Configurations for Ions

2.3.1

Electron Configuration of Ions of Main Group Elements

We will now look at how to write the electron configurations for ions!

Writing the electron configurations for anions (negative ions) are simple.

1. Write out the electron configuration for the **neutral element**
2. Fill in the desired number of electrons by **adding electron(s) to the highest energy subshell**

Writing the electron configurations for cations (positive ions) usually confuses students more, but there is a simple thing to remember to get these questions right every time!

1. Write out the electron configuration for the **neutral element**
2. Remove the desired number of electrons by **first removing electrons from the highest n level and highest energy subshell!**

WIZE CONCEPT

The higher the n level, the higher the energy.
Example: **2**s has a higher energy than **1**s

In order of **increasing energy for the subshells**, we write:
s < p < d < f
Example: 2**p** has a higher energy than 2**s**

F: $1s^2 2s^2 2p^5$ **[He] $2s^2 2p^5$**	2s 2p
F^-: $1s^2 2s^2 2p^6$ **[Ne]**	2s 2p

What is the electron configuration for F^+?

WIZE TIP

The most stable ions of atoms are **isoelectronic** with the noble gases or have filled shells. **Isoelectronic species** have the same number of electrons and the same electron configuration. The term "isoelectronic" is commonly seen on exams.

Example: F^- is isoelectronic with Ne!

2.3.2

Electron Configuration of Ions of Transition Elements

The same rules that we just learned apply for transition elements as well. Note that valence electrons are only removed, never core electrons.

WATCH OUT!

Transition metals can lose both “n” and “n-1” valence electrons, but “n” electrons are always lost first! Let's take a look!

Fe: [Ar] $4s^2 3d^6$	4s	3d
Fe^{1+}: [Ar] $4s^1 3d^6$	4s	3d
Fe^{2+}: [Ar] $3d^6$	4s	3d
Fe^{3+}: [Ar] $3d^5$	4s	3d

WIZE TIP

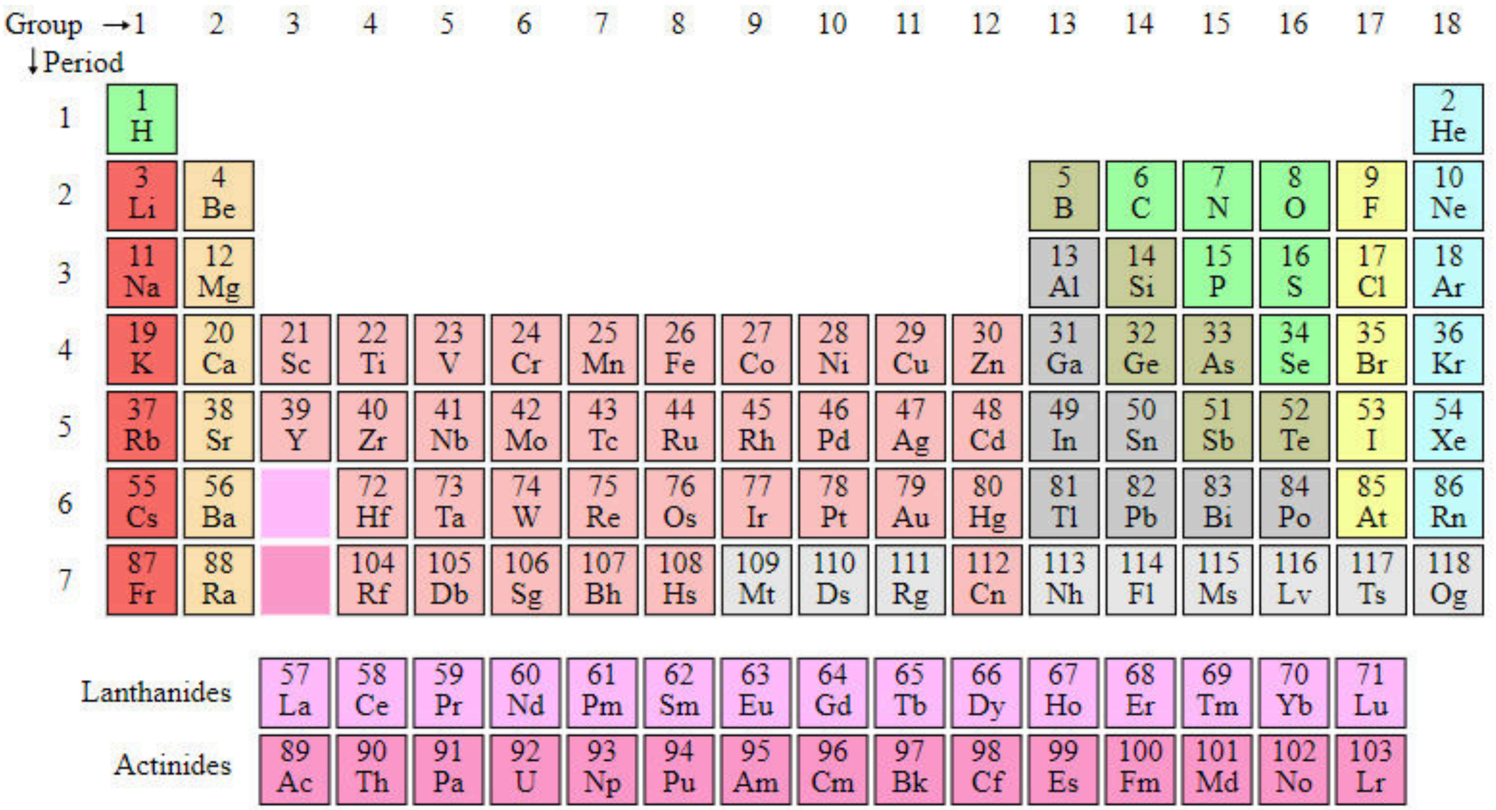
Always write out the **electron configuration for the neutral atom first.**
Then write out the **electron configuration for the charged element.**

This will help you avoid mistakes on the exam!

2.3.3

Example: Isoelectronic Species

Two atoms or ions are said to be isoelectronic if the electron configuration is the same for both species. Which of the following pairs are isoelectronic?



Group → ↓Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	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
5	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
6	55 Cs	56 Ba		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
7	87 Fr	88 Ra		104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Ms	116 Lv	117 Ts	118 Og

Lanthanides	57 La	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	89 Ac	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

i) N^+ and C

ii) C and B^+

2.3.4

Example: Writing Electron Configurations of Ions

Write the electronic configuration for the following species:

Group → / ↓Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	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
5	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
6	55 Cs	56 Ba		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
7	87 Fr	88 Ra		104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Ms	116 Lv	117 Ts	118 Og

Lanthanides	57 La	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	89 Ac	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

a) N^+

b) Ni^{2+}

2.3.5

Practice: Electron Configurations of Ions

For which of the atom/ions below, is the given electron configuration correct? (select all that apply)

	Species	Electron Configurations
i.	Cu	[Ar] $4s^2\ 3d^9$
ii.	Ti^{2+}	[Ar] $4s^2$
iii.	As^{5+}	[Ar] $4s^2\ 3d^8$

- i.
- ii.
- iii.
- none of the above

2.4 Diamagnetic vs Paramagnetic Electron Configurations

2.4.1

Diamagnetic vs Paramagnetic

You may see these terms come up when talking about electron configurations.

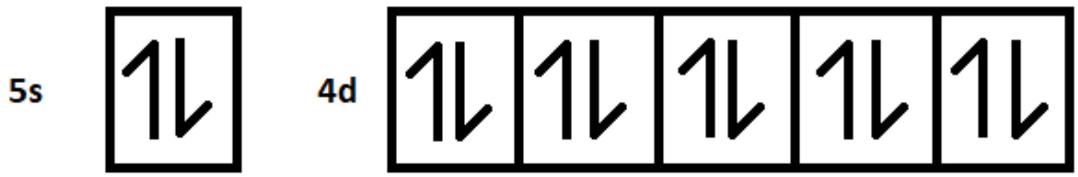
Diamagnetic - all electrons are **paired**

Paramagnetic - all electrons are **not paired**

Are the following electron configurations diamagnetic or paramagnetic?

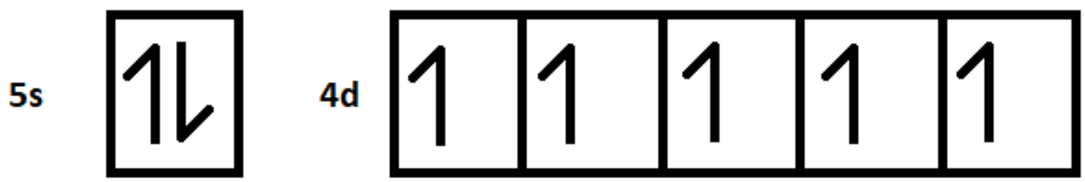
Cd

[Kr] $5s^2\ 4d^{10}$



Tc

[Kr] $5s^2\ 4d^5$



Group → ↓Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	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
5	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
6	55 Cs	56 Ba		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
7	87 Fr	88 Ra		104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Ms	116 Lv	117 Ts	118 Og
Lanthanides			57 La	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			89 Ac	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	

Consider the electron configuration for Na

Is it diamagnetic or paramagnetic? ______________

Consider the electron configuration for Si

Is it diamagnetic or paramagnetic? ______________

Consider the electron configuration for Be

WIZE CONCEPT

Diamagnetic-all electrons are **paired**. If something is diamagnetic it will also be **repelled** from externally produced magnetic fields.

Paramagnetic-One or more electrons are left **unpaired**. If something is paramagnetic it will be **attracted** to an externally produced magnetic field.

2.4.2

Example: Paramagnetic and Diamagnetic Species

Write the electronic configuration of the following species, and label them as **D** diamagnetic or **P** paramagnetic

Group → ↓Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	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
5	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
6	55 Cs	56 Ba		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
7	87 Fr	88 Ra		104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Ms	116 Lv	117 Ts	118 Og

Lanthanides	57 La	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	89 Ac	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

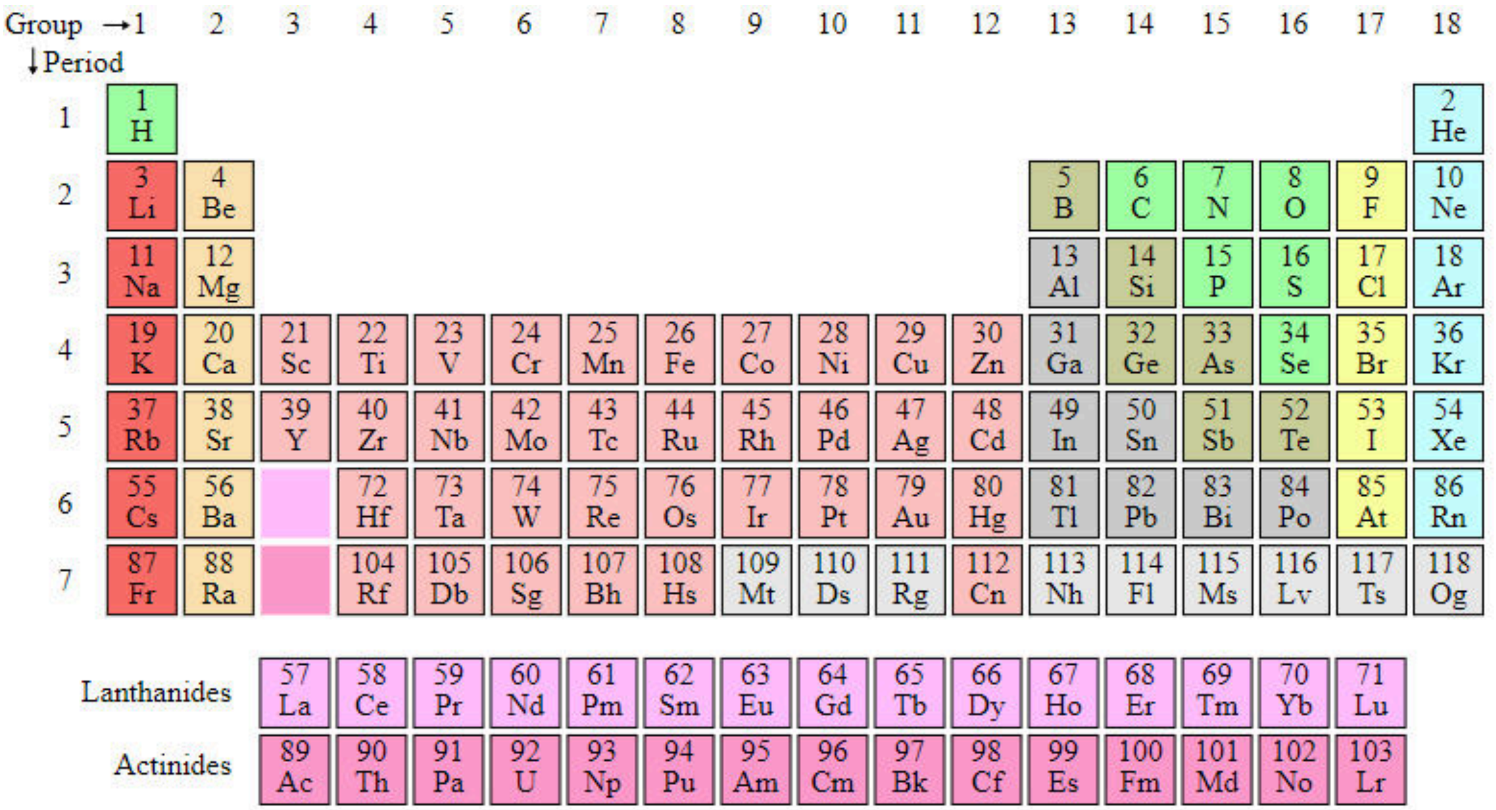
a) Sc^{2+}

b) Cr^{2+}

2.4.3

Practice: Diamagnetic

Which of the following are diamagnetic? (select all that apply in your answer)



- [ ] Cr
- [ ] Co
- [ ] Ru
- [ ] None of the above

2.5 Effective Nuclear Charge (Zeff)

2.5.1

Effective Nuclear Charge (Zeff)

To understand effective nuclear charge (Zeff), let's consider a concert:

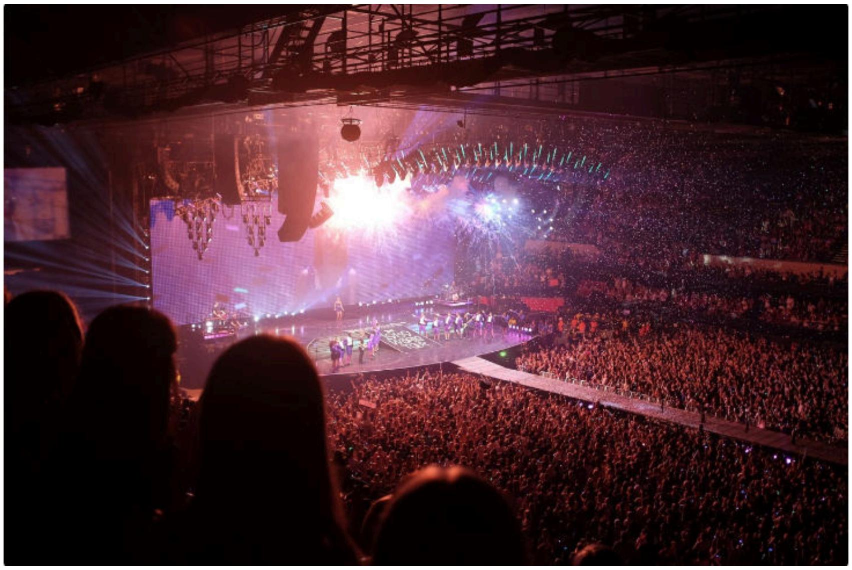


Photo by The Come Up Show / CC BY

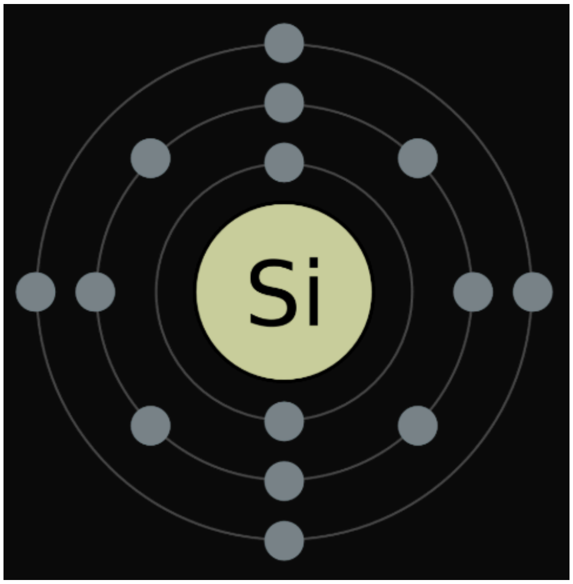
- People closer to the stage at a concert are going to be more into it. The music is louder and they really get a close connection with the artist. People in the highest rows in the stadium get less of that connection with the artist performing.
- This helps to explains the concept of nuclear shielding. **Inner electrons** (or people sitting in the front rows in our example) **shield the outer electrons** (or people sitting in the highest rows) from the **attractive force of the nucleus** (or artist in our example)

Effective Nuclear Charge

This is the **nuclear charge that is "felt"** by a valence electron.

Core Electrons Vs Valence Electrons

Label the **core** and **valence electrons** in the diagram below:



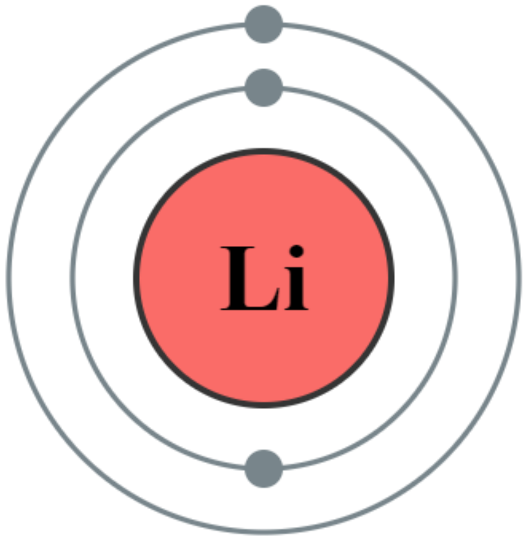
- Valence electrons are attracted to the positively charged nucleus BUT valence electrons are repelled by the core electrons
- Note that electrons in the same shell "feel" the same attraction to the nucleus (since they are they same distance from the nucleus, just like how the people in the same row would feel the same connection to the artist)

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

Zeff *is the* ***effective nuclear charge***
Z *is the* ***atomic number*** *(# of protons)*
S *is the # of* ***shielding electrons***

Are the shielding electrons core electrons or valence electrons? ________

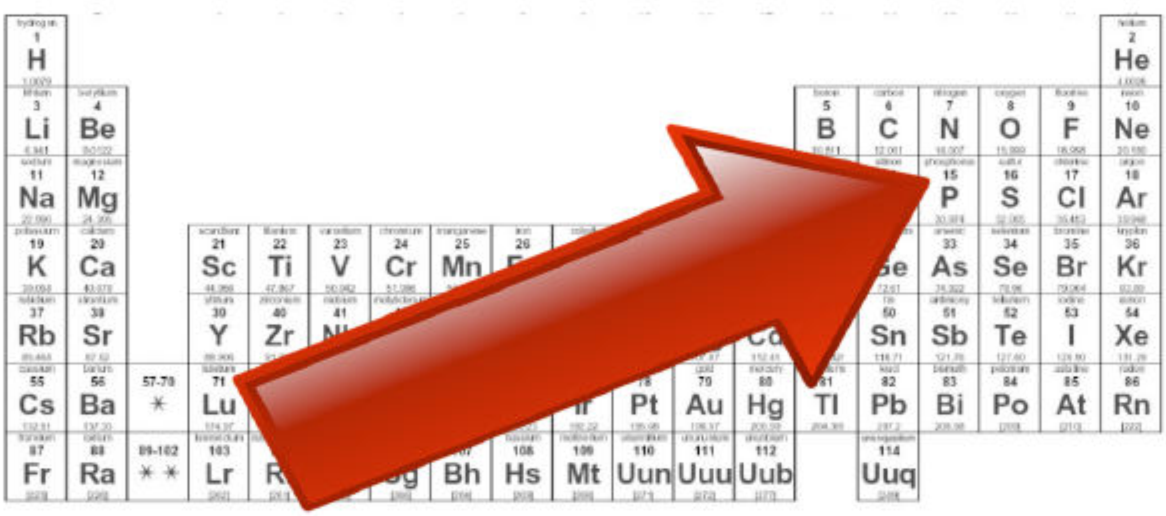
What would be the effective nuclear charge for Lithium?



The Periodic Trend

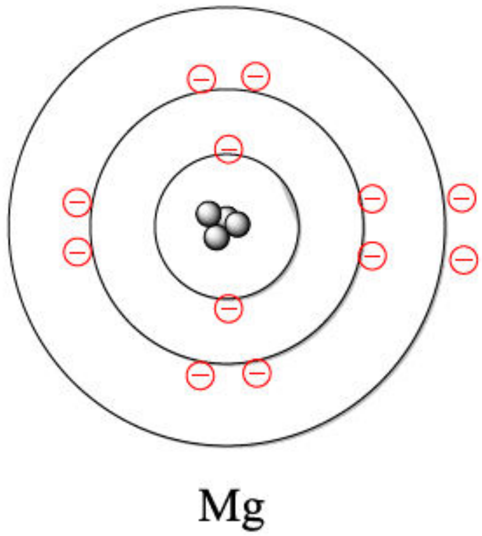
- As we move to the right across the periodic table, the # of core electrons stays the same but the # of protons increases. Therefore, Z_{eff} increases.
  - With protons, can pull the electrons in closer → Zeff
  - Effects of shielding are less with more protons

- As we move down a group of the periodic table, the # of core shells increases and the valence electrons get further from the nucleus. This increased distance from the nucleus leads to a smaller Z_{eff}.
  - a group → shells
  - shells → shielding
  - shielding → Zeff



Example: Estimating Zeff

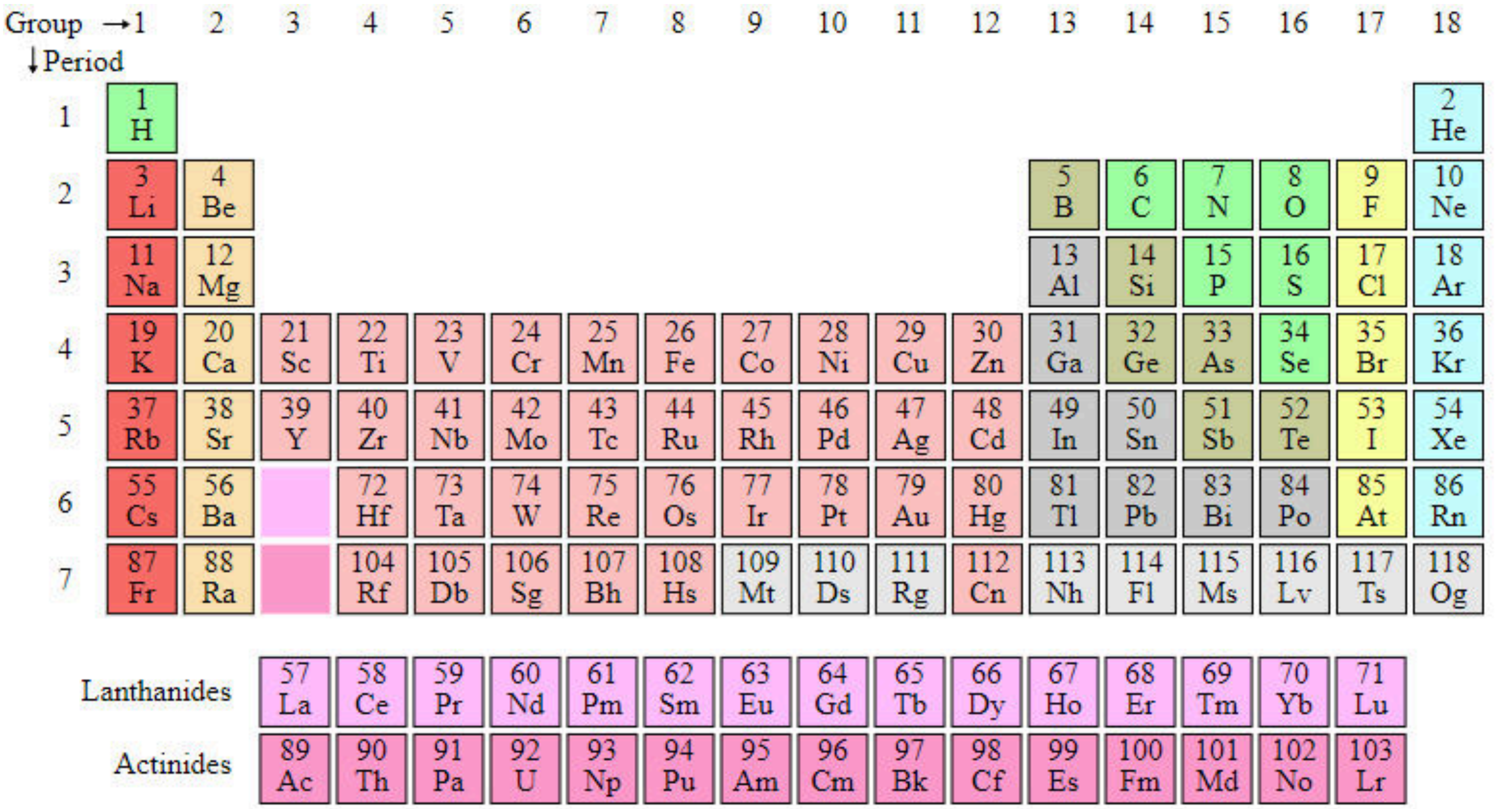
What is the Z_{eff} of an electron in the n=3 shell of magnesium (Mg)?



2.5.3 Practice

Practice: Estimating Zeff

Calculate the approximate effective nuclear charge of Se.



- 6
- 8
- 28
- 34

2.6 Atomic Radius Trend

2.6.1

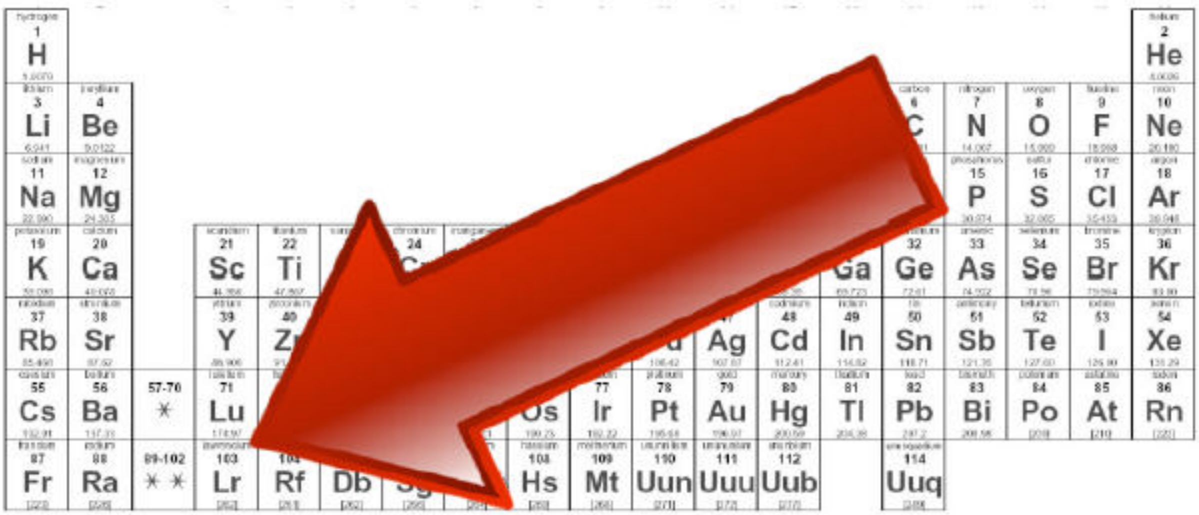
Atomic Radius

The atomic radius is the estimated **radius of an atom** (from the nucleus to the outermost valence electrons)

The Periodic Trend

- As we **move to the right** on the periodic table, the **Z_{eff} increases** and this “pulls” the electrons closer to
  the nucleus which **decreases the radii.**

- As we **move down a group** in the periodic table, the number of **electron shells increases** which makes the
  atom radii larger.
  - group → shells
  - shells → shielding and Zeff
  - With a Zeff → pull on outer electrons, leading to atomic radii



Element with the **largest atomic radii** in the periodic table: ______

Element with the **smallest atomic radii** in the periodic table: ______

How do ions change this?

- **Anions:** __________ due to increased electron-electron repulsion
- **Cations:** __________ due to decreased electron-electron repulsion
  - How does this relate to Zeff?

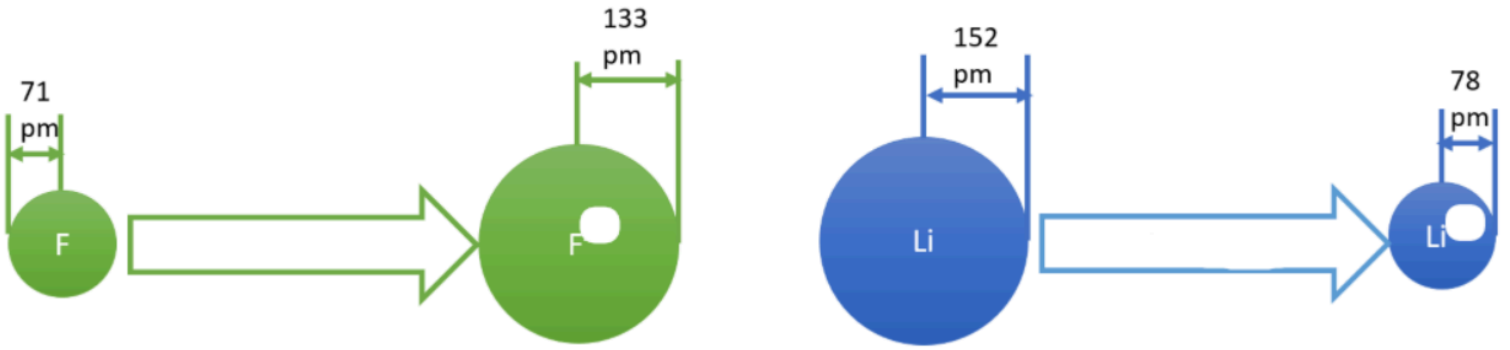
In general:

Ionic Radius: Three Scenarios

You could be asked to rank different atoms or ions according to the sizes of their ionic radii using the trends we discussed.

Here are three common scenarios you may encounter:

1. Same element different charge:



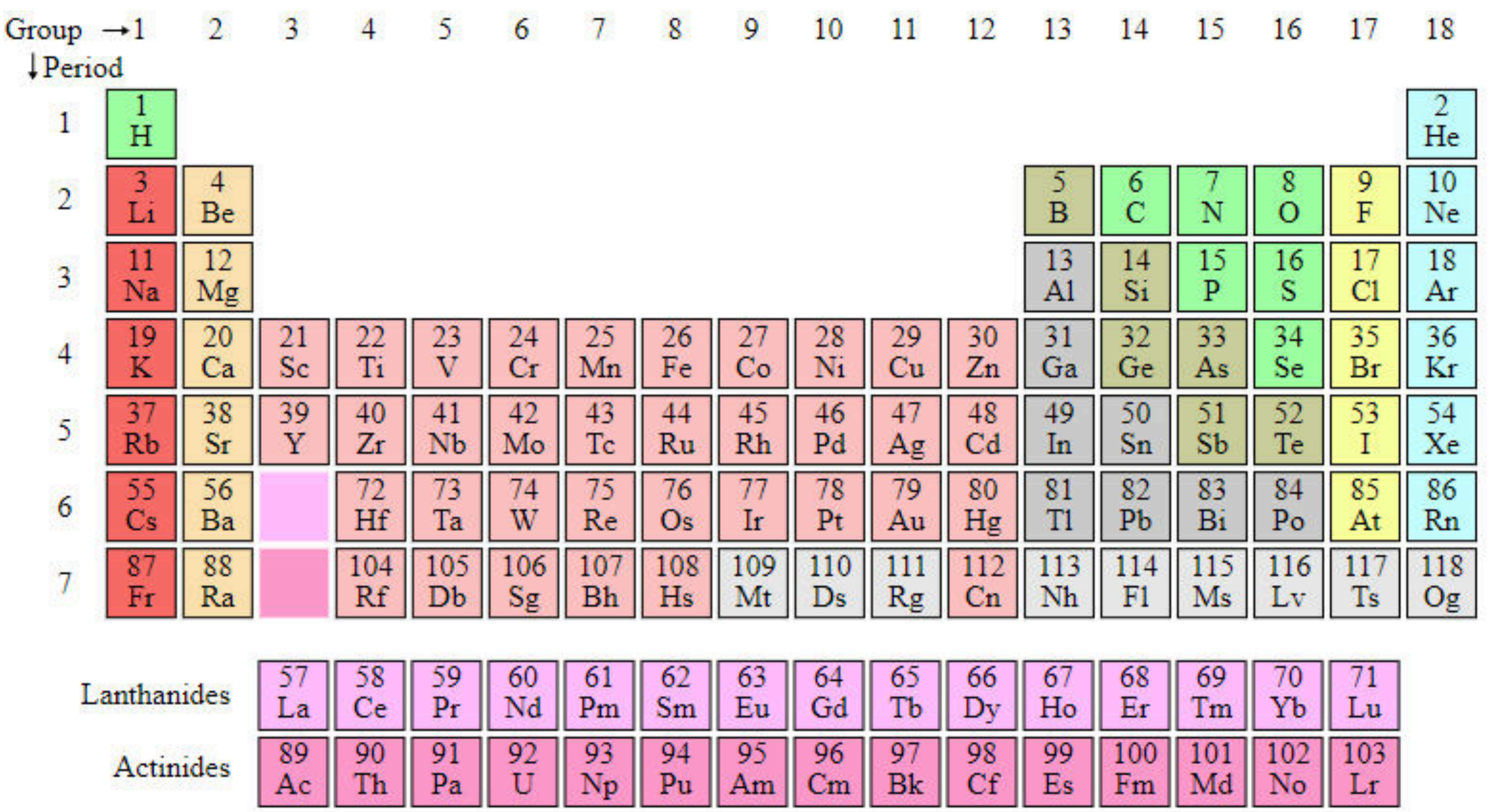
2. Different element same charge:

- Identical trend to atoms
  Example: Li < Na so $Li^+ < Na^+$

3. Different element different charge

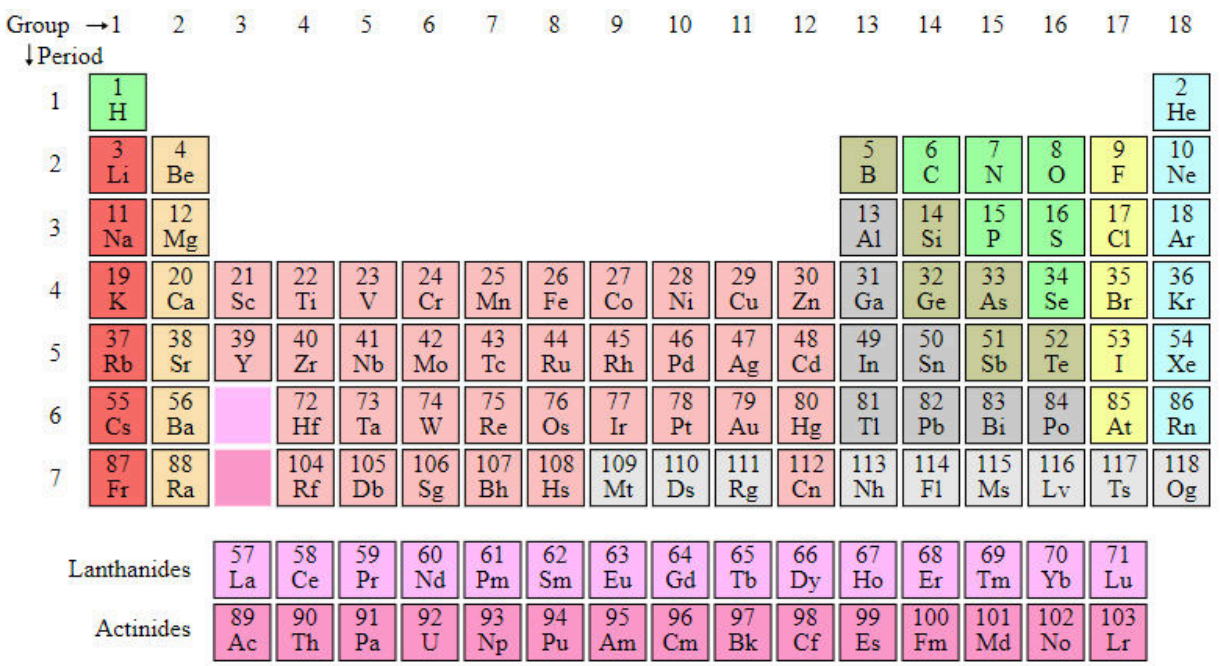
- Can only be assessed for **isoelectronic species** (same # of electrons)
- Compare the **proton to electron ratio.**
  - **More protons than electrons**, means stronger pull, **smaller radius.**
  - **Less protons than electrons**, means weaker pull, **larger radius.**

  Example: Compare the radii for the following: O^{2-} F^{-} Ne Na^{+} Mg^{2+}



Example: Ranking Size of Atoms

Rank the following atoms in order of increasing atomic radius: Se, Cs, Br, Ga, F, As.



2.6.4

Example: Ranking Size of Ions

Rank the following species in order of decreasing size. F^{-}, N^{3-}, Al^{3+}

Group → ↓Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	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
5	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
6	55 Cs	56 Ba		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
7	87 Fr	88 Ra		104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Ms	116 Lv	117 Ts	118 Og
Lanthanides			57 La	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			89 Ac	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	

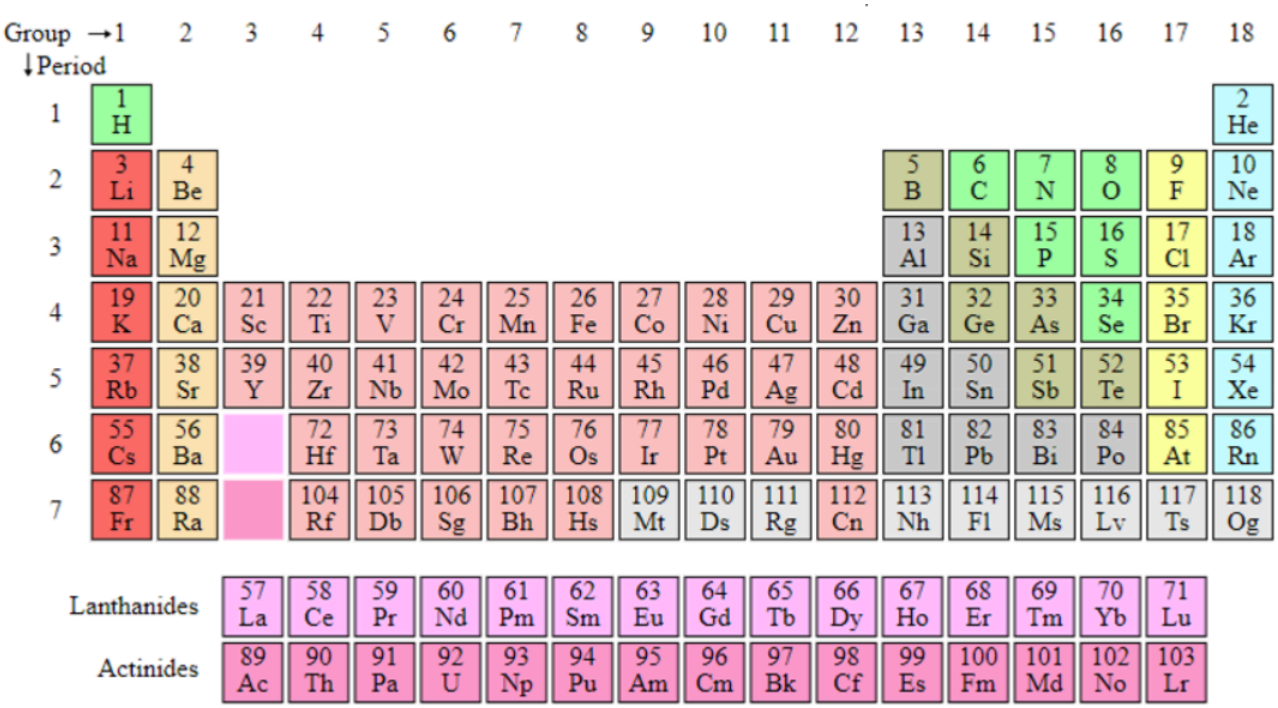
WIZE TIP

When comparing atoms that are **isoelectronic** (i.e. they have the same number of electrons), remember **more protons in the nucleus** means a stronger the pull of electrons toward the + charged nucleus, so the **smaller the atomic radius**

2.6.5

Practice: Determining the Largest Atom

Select the atom below which has the largest atomic radius?



- S
- I
- O
- In
- Rb

View Solutions on Wizeprep.com

Solutions to these questions, as well as step-by-step breakdowns of the answers at:

2.7 Ionization Energy

2.7.1

Ionization Energy (IE)

The ionization energy is the amount of **energy required to remove the outermost electron** from an atom or ion in the gas phase

- If removing an electron do we get an anion or a cation? ________

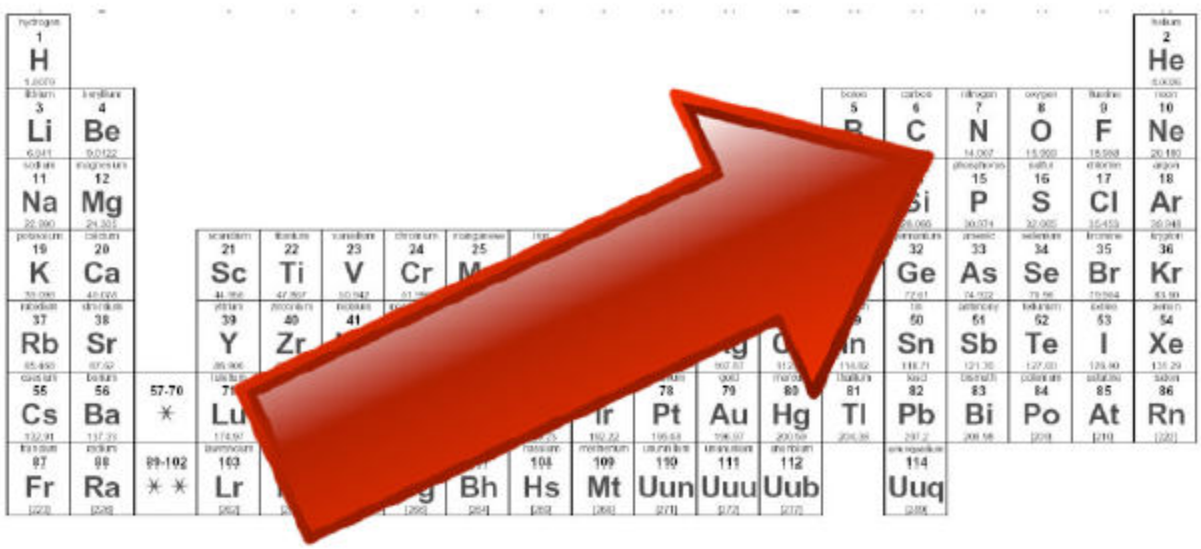
$$\mathbf{E} \rightarrow \mathbf{E}^{+} + \mathbf{e}^{-} \qquad \Delta \mathrm{E} = \text{Ionization Energy}$$

- **↑IE** means that the atom/ion requires **more energy** for a valence **electron to be removed** (has a tight pull on electron)

The Periodic Trend

- As we **move to the right** across a period, **Z_{eff} increases** and the electrons are held more tightly. Therefore, it takes more energy to remove an electron and the **IE increases** as we move to the right.

- As we **move down a group** and the valence electrons are further away from the nucleus, they are held more weakly (**lower Zeff)** and **IE decreases** (it becomes easier to remove the outermost electron)


H
He
Li
Be
C
N
O
F
Ne
Na
Mg
P
S
Cl
Ar
K
Ca
Sc
Ti
V
Cr
Ge
As
Se
Br
Kr
Rb
Sr
Y
Sn
Sb
Te
I
Xe
Cs
Ba
57-70
Ir
Pt
Au
Hg
Tl
Pb
Bi
Po
At
Rn
Fr
Ra
89-102
Lr
Bh
Hs
Mt
Uun
Uuu
Uub
Uuq

Comparing 1st, 2nd, and 3rd Ionization Energies

How does the first IE compare to the second IE?

$$First\ Ionization\ Energy\ (IE_1): \quad A\ (g)\ \rightarrow\ A^{+}\ (g)\ +\ e^{-}$$

$$Second\ Ionization\ Energy\ (IE_2): \quad A^{+}\ (g)\ \rightarrow\ A^{2+}\ (g)\ +\ e^{-}$$

$$Third\ Ionization\ Energy\ (IE_3): \quad A^{2+}\ (g)\ \rightarrow\ A^{3+}\ (g)\ +\ e^{-}$$

1st ionization is always smaller than the 2nd because removing a negatively charged electron from a cation is more difficult.

Example:

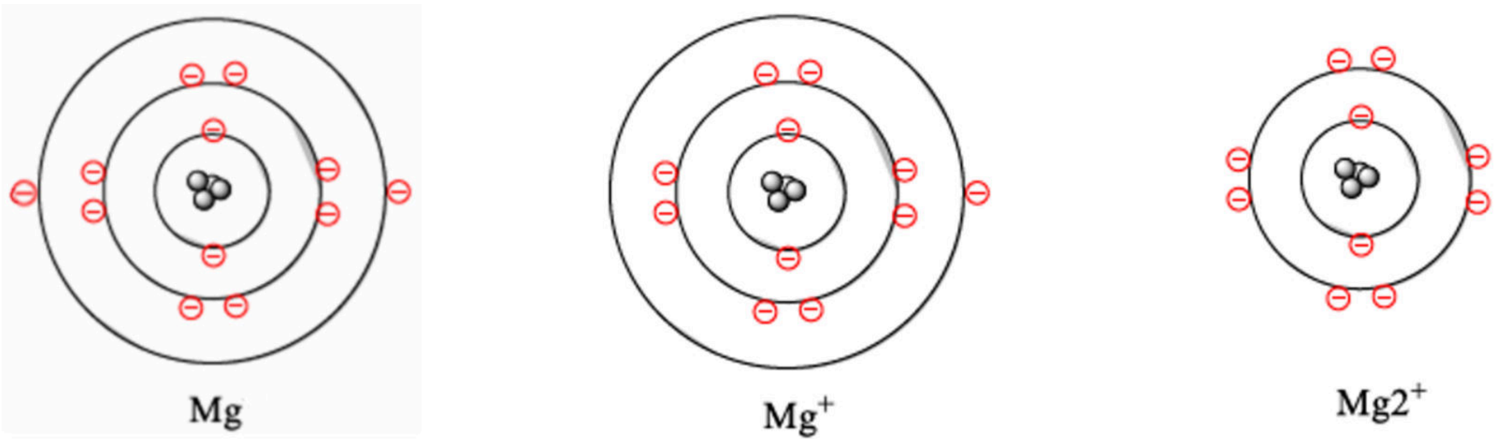
$$Ca\ (g)\ \rightarrow\ Ca^{+}\ (g)\ +\ e^{-} \quad IE_1\ =\ 590\ \frac{kJ}{mol}$$

$$Ca^{+}\ (g)\ \rightarrow\ Ca^{2+}\ (g)\ +\ e^{-} \quad IE_2\ =\ 1145\ \frac{kJ}{mol}$$

WIZE TIP

IE_1 < IE_2 < IE_3 < IE_4 etc

How does the IE_2 and IE_3 for magnesium compare?



A **core electron is much more difficult to remove than a valence electron**, for example: the 3rd ionization energy of magnesium is much, much, higher than the 2nd.

WIZE TIP

Although the **IE increases each time another electron is removed**, **the increase is not linear** and is usually **related to the electron configuration.**

Ionizing Mg^{2+} would be very difficult!

$$Mg \rightarrow \ Mg^{+} \ + \ e^{-} \quad IE_1 \ = \ 738 \ \frac{kJ}{mol}$$

$$Mg^{+} \rightarrow \ Mg^{2+} \ + \ e^{-} \quad IE_2 \ = \ 1451 \ \frac{kJ}{mol}$$

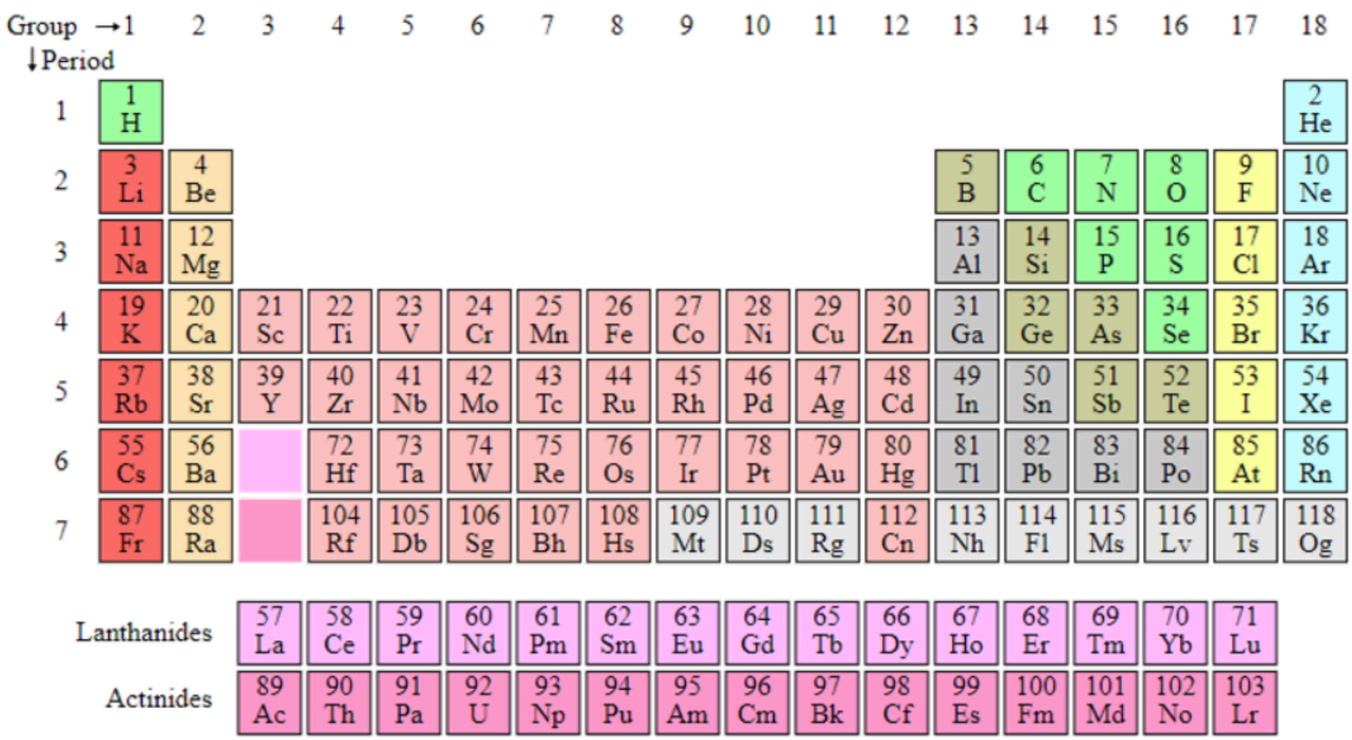
$$Mg^{2+} \rightarrow \ Mg^{3+} \ + \ e^{-} \quad IE_3 \ = \ 7733 \ \frac{kJ}{mol}$$

Needing so much energy to remove the 3rd electron indicates that when 2 were removed it was stable/unstable: ______ !

2.7.3

Practice: Highest 1st Ionization Energy

Which of the following species has the highest 1^{st} ionization energy?



- S
- F
- B
- N
- P

View Solutions on Wizeprep.com

Solutions to these questions, as well as step-by-step breakdowns of the answers at:

Practice: Trends in Ionization Energy

Rank the following atoms in order of increasing first ionization energy (1 = smallest; 6 = largest): Fe, O, Ba, Si, V, Zr

Group → / ↓Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	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
5	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
6	55 Cs	56 Ba		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
7	87 Fr	88 Ra		104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Ms	116 Lv	117 Ts	118 Og

Lanthanides	57 La	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	89 Ac	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

Fe	
O	
Ba	
Si	

V	
Zr	

2.7.5

Practice: Lowest Ionization Energy

Select the element with the lowest 3rd ionization energy.

Group → / ↓Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	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
5	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
6	55 Cs	56 Ba		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
7	87 Fr	88 Ra		104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Ms	116 Lv	117 Ts	118 Og

Lanthanides	57 La	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	89 Ac	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

A) Sc

B) Mg

C) Cl

D) I

E) C

2.8 Electron Affinity Trend

2.8.1

Electron Affinity (EA)

Electron affinity is the **amount of energy involved with adding an electron** to an atom or ion.

WIZE TIP

Don't confuse ionization energy (IE) with electron affinity (EA)!

- **Ionization energy** is when we **remove** the outermost electron
- **Electron affinity** is when we **add** an electron

Memory tip: Electron affinity has an "**a**" in it for adding an electron

When adding an electron, do we get a cation or an anion? ________!

$$\mathbf{E + e^- \rightarrow E^-} \qquad \mathbf{\Delta E = Electron\ Affinity}$$

- **If EA is negative** → energy is released/absorbed: ____________
  - If this is the case, would the atom go to a higher/lower E state? __________
  - This means the atom has become more/less ________ stable as a result of adding an electron
  - Did this atom like or dislike having an electron added to it? _______

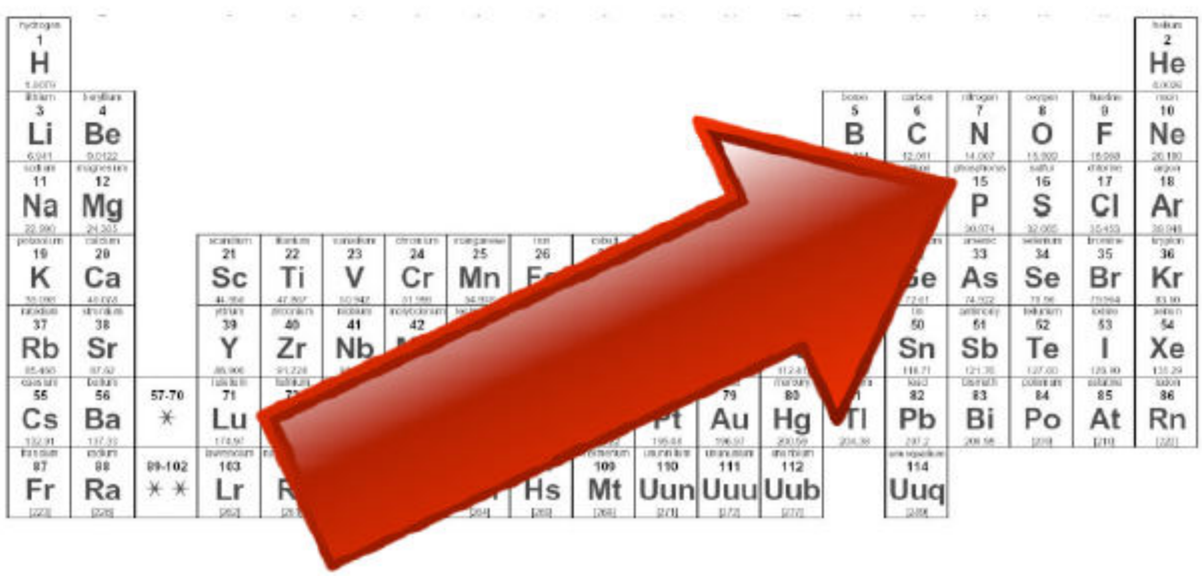
- **If EA is positive** → energy is released/absorbed: ____________
  - If this is the case, would the atom go to a higher/lower E state? __________
  - This means the atom has become more/less stable________ stable as a result of adding an electron
  - Did this atom like or dislike having an electron added to it? __________

> **WATCH OUT!**
>
> Use a **negative EA as a reference point**, unless the question states otherwise.

- This means that unless specified otherwise, when asked which element would have the largest electron affinity, assume they mean largest NEGATIVE electron affinity.
  - This would mean that the element really liked getting an electron and lost energy, resulting in a more stable form.

The Periodic Trend

In general, **as Z_{eff} increases**, the **incoming electron** experiences a greater electrostatic attraction and **stabilization** leading to a **greater (negative) electron affinity**



What about noble gases?

Do you think a noble gas would have a positive or negative electron affinity? ____________

- Are noble gases stable? ________
- Would a noble gas want another electron added? ________
- If you tried to add an electron to a noble gas would that require energy or release energy?

  ____

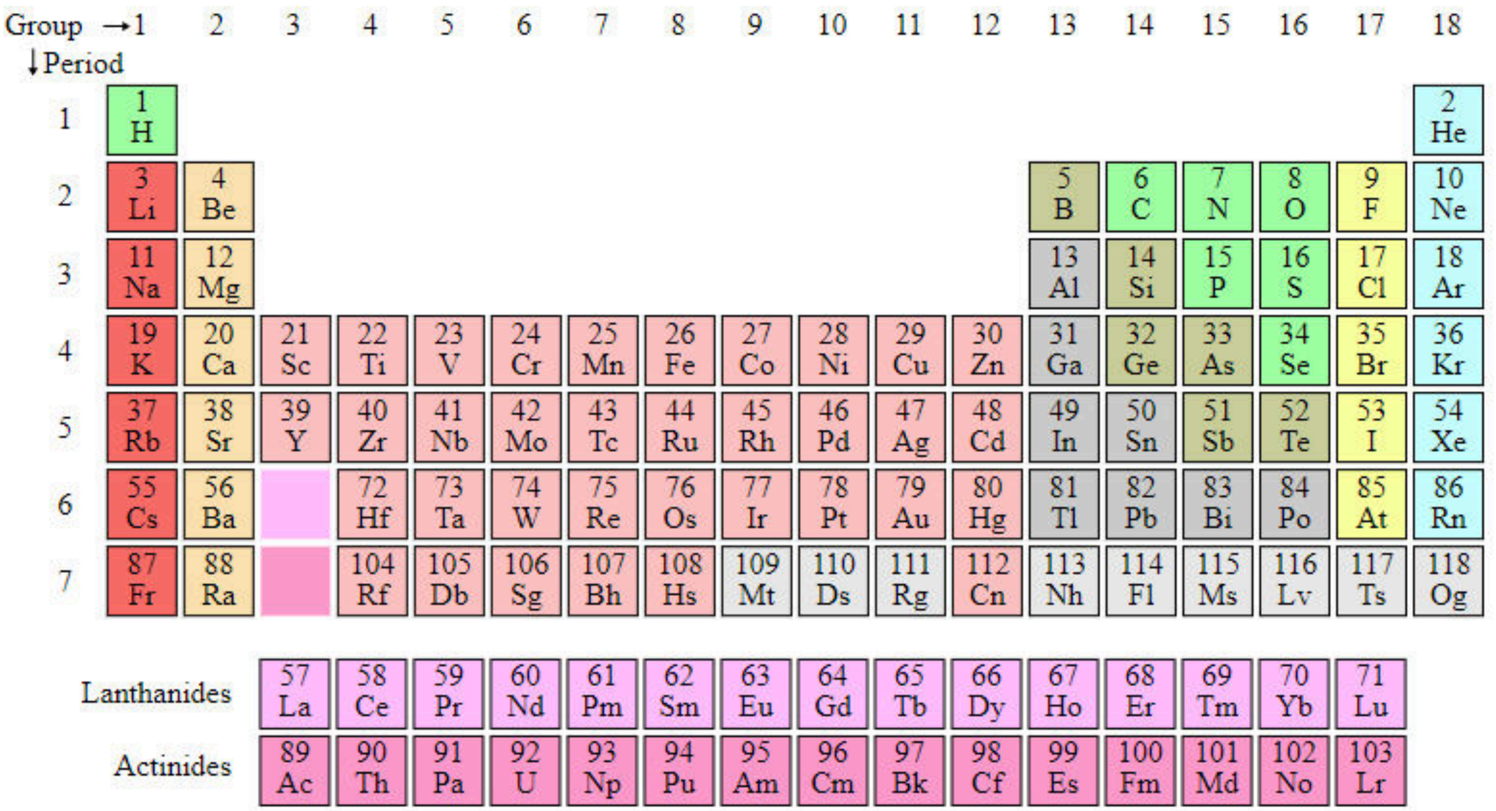
WATCH OUT!

This trend excludes noble gases. Noble gases have stable, completely filled shells. **Adding electrons to noble gases will break the noble gas configuration.**

2.8.2

Example: Increasing Electron Affinity

Rank the following atoms in increasing absolute energy of their first electron affinity: Br, Ag, Ba, Mo, Sb



2.8.3

Practice: Positive Electron Affinity

For most elements, the first electron affinity is negative; however, a few elements have a positive first electron affinity. Which atom has the most positive first electron affinity?

- F
- B
- Ne
- C
- O
- Na

2.8.4

Practice: Highest Electron Affinity

Which of the following has the highest electron affinity?

- N
- N+
- N-
- Not enough information is provided to determine this.

2.9 Electronegativity

2.9.1

Electronegativity (EN)

Electronegativity is a **measure of electron pull.** In a bond, the **more electronegative** element will **pull electrons closer towards itself**

Example: Draw the bond between H and F and show the difference in electronegativity.

Group → / ↓ Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	H 2.20																	He
2	Li 0.98	Be 1.57											B 2.04	C 2.55	N 3.04	O 3.44	F 3.98	Ne
3	Na 0.93	Mg 1.31											Al 1.61	Si 1.90	P 2.19	S 2.58	Cl 3.16	Ar
4	K 0.82	Ca 1.00	Sc 1.36	Ti 1.54	V 1.63	Cr 1.66	Mn 1.55	Fe 1.83	Co 1.88	Ni 1.91	Cu 1.90	Zn 1.65	Ga 1.81	Ge 2.01	As 2.18	Se 2.55	Br 2.96	Kr 3.00
5	Rb 0.82	Sr 0.95	Y 1.22	Zr 1.33	Nb 1.6	Mo 2.16	Tc 1.9	Ru 2.2	Rh 2.28	Pd 2.20	Ag 1.93	Cd 1.69	In 1.78	Sn 1.96	Sb 2.05	Te 2.1	I 2.66	Xe 2.60
6	Cs 0.79	Ba 0.89	*	Hf 1.3	Ta 1.5	W 2.36	Re 1.9	Os 2.2	Ir 2.20	Pt 2.28	Au 2.54	Hg 2.00	Tl 1.62	Pb 1.87	Bi 2.02	Po 2.0	At 2.2	Rn 2.2
7	Fr 0.7	Ra 0.9	**	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cn	Uut	Fl	Uup	Lv	Uus	Uuo
* Lanthanoids			La 1.1	Ce 1.12	Pr 1.13	Nd 1.14	Pm 1.13	Sm 1.17	Eu 1.2	Gd 1.2	Tb 1.1	Dy 1.22	Ho 1.23	Er 1.24	Tm 1.25	Yb 1.1	Lu 1.27	
** Actinoids			Ac 1.1	Th 1.3	Pa 1.5	U 1.38	Np 1.36	Pu 1.28	Am 1.13	Cm 1.28	Bk 1.3	Cf 1.3	Es 1.3	Fm 1.3	Md 1.3	No 1.3	Lr 1.3	

Electronegativity Values

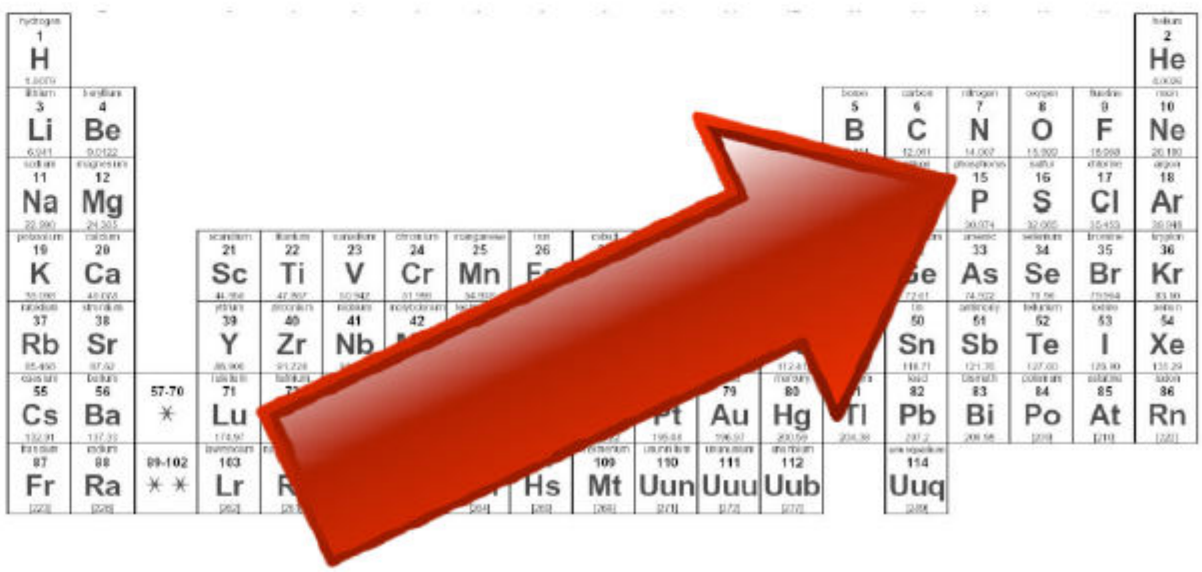
WIZE TIP

General order of EN may come in handy:

F > O > N > Cl > Br > I > S > C ~ H

The Periodic Trend

Electronegativity increases going up and to the right in the periodic table.



Which atom is the most electronegative? ____

Example: Increasing Electronegativity

Rank the following atoms in increasing electronegativity: O, Al, Sr, N, Si, Cs

Group → ↓Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	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
5	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
6	55 Cs	56 Ba		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
7	87 Fr	88 Ra		104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Ms	116 Lv	117 Ts	118 Og

Lanthanides	57 La	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	89 Ac	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

2.9.3

Practice: Ranking Electronegativity

Choose the option which correctly ranks the following elements in terms of electronegativity.

Group → / ↓Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	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
5	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
6	55 Cs	56 Ba		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
7	87 Fr	88 Ra		104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Ms	116 Lv	117 Ts	118 Og

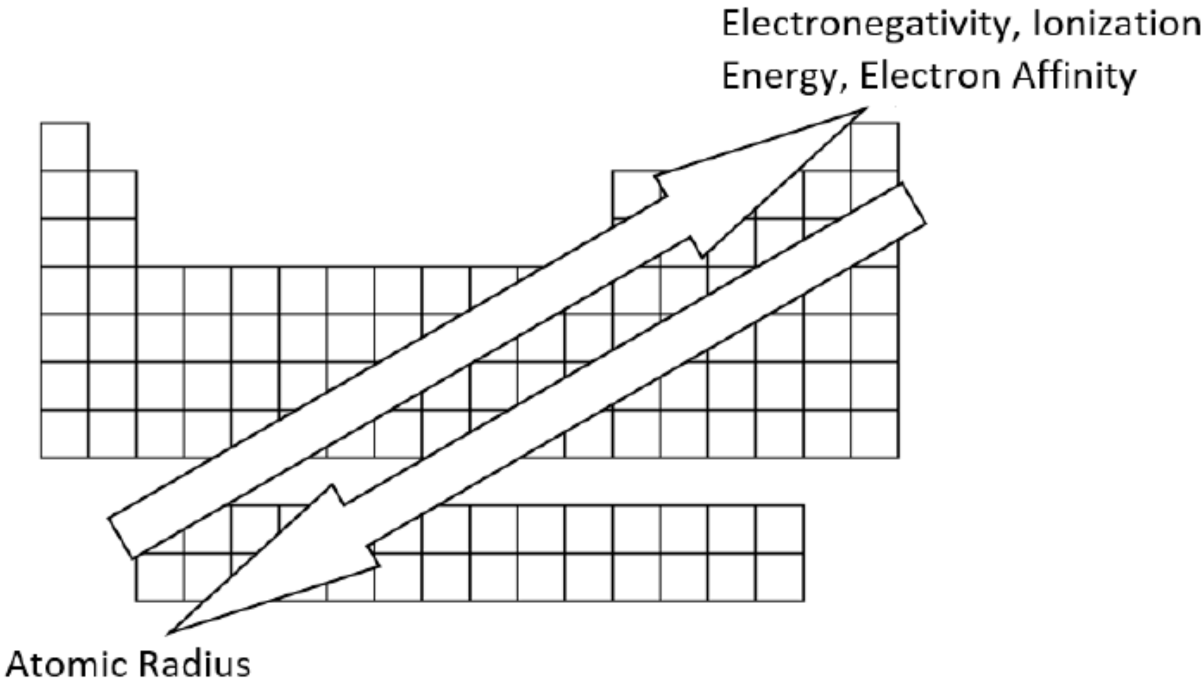
Lanthanides	57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
Actinides	89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

- Ti < Bi < Hg < Cs < Fr
- Na < Al < O < S < Cl
- Mg < P < C < O < F
- Fe < Co < Mo < F < Cl

2.10 Summary of Periodic Table Trends

2.10.1

Summary of Periodic Trends



Review: which direction does Zeff increase in? ____________________

2.10.2

Example: Periodic Trends

Label the following statements as either TRUE or FALSE

1. Ionization energy decreases when the atomic size decreases
2. As atomic size increases it gets easier to add an additional electron
3. Mg^{2+} is the same size as Ne since they are isoelectronic

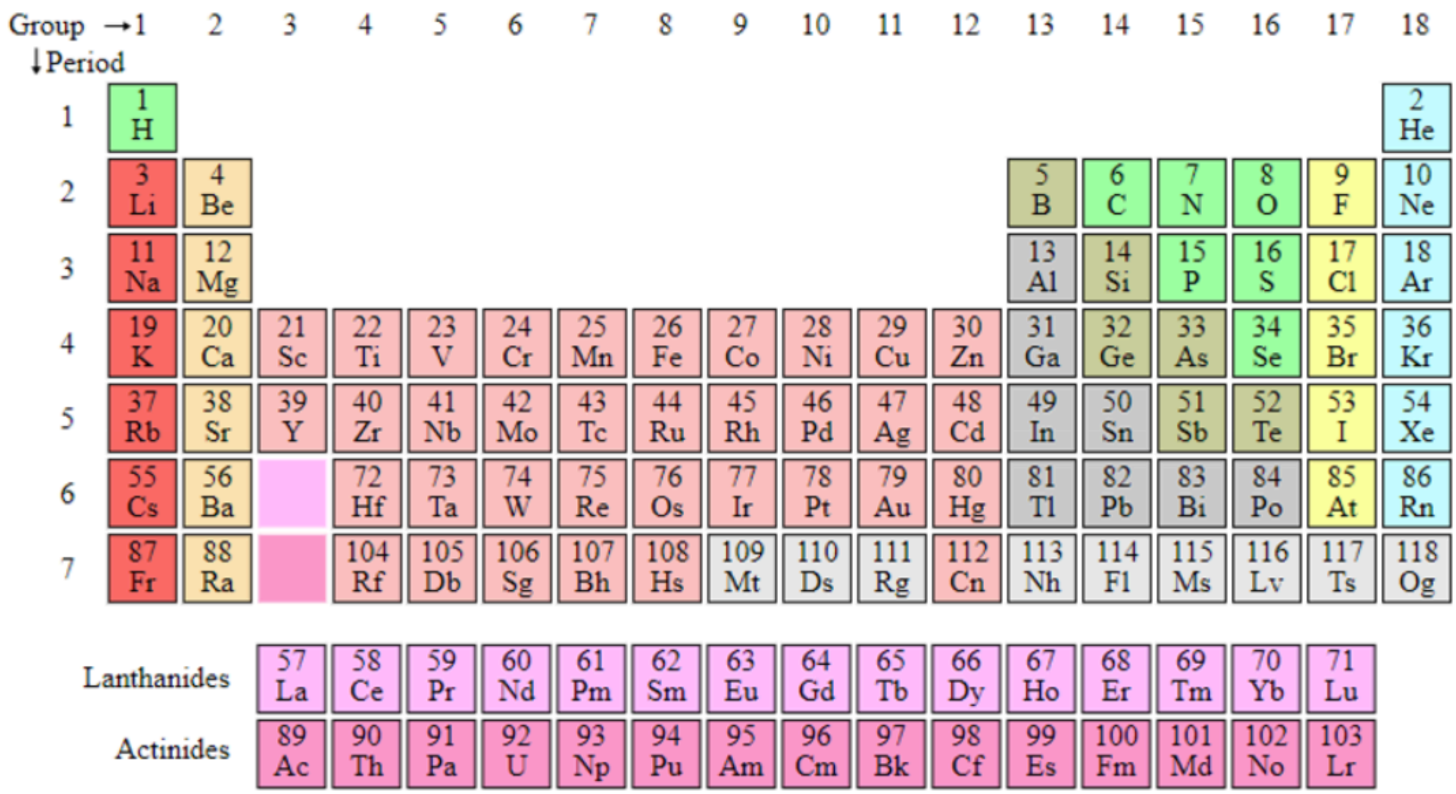
2.

3.

2.10.3

Practice: Periodic Trends

Complete the statements below by filling in the blanks and then choose which option below would help explain your answer.



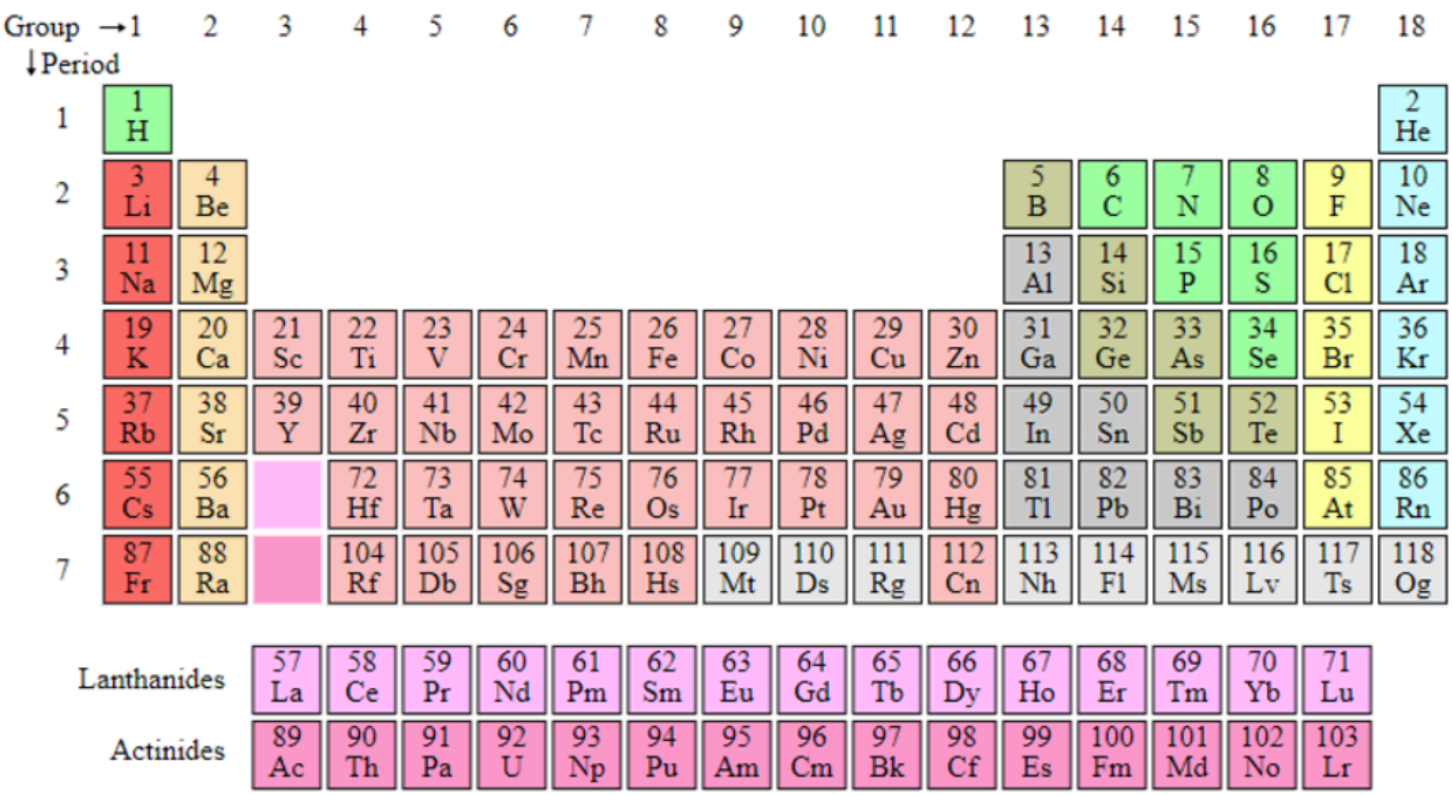
Part 1

N^{3-} has a (larger/smaller) ionic radius in comparison to Si^{4+} because ...

- ... of the difference in Z*
- ... of the difference in n
- ... it has a higher proton to electron ratio
- ... it has a higher electron to proton ratio

Practice: Periodic Trends

Complete the statements below by filling in the blanks and then choose which option below would help explain your answer.



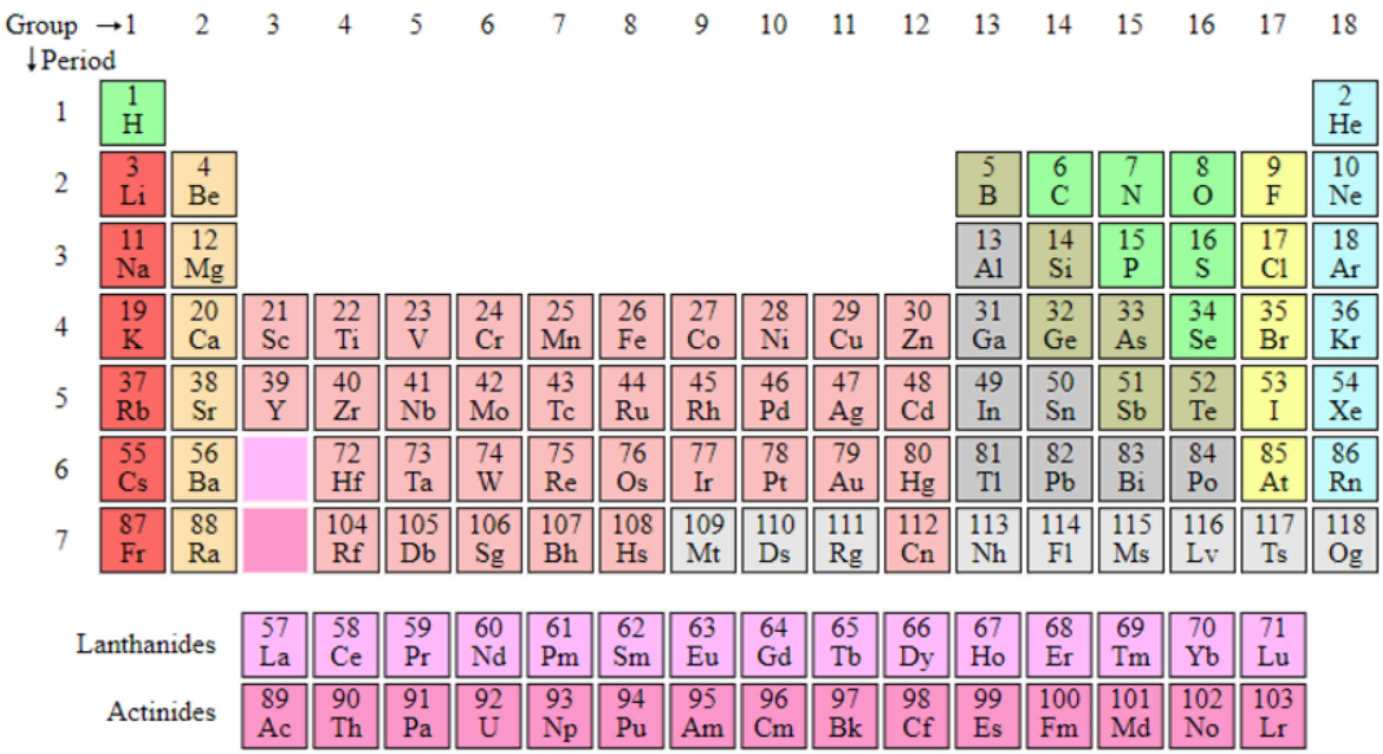
Part 2

Al has a (more negative/ less negative) EA than Si because ...

- ... of the difference in Zeff
- ... of the difference in n
- ... it has a higher proton to electron ratio
- ... it has a higher electron to proton ratio

Practice: Periodic Trends

Which of the following statements is correct about chlorine and phosphorus atoms?



A) The electron affinity of a chlorine atom is smaller than that of a phosphorus atom due to its smaller radius.

B) A chloride anion has higher electron affinity than a neutral chlorine atom.

C) A chlorine atom can easily gain and/or lose an electron since it has both a large electron affinity and a large ionization energy.

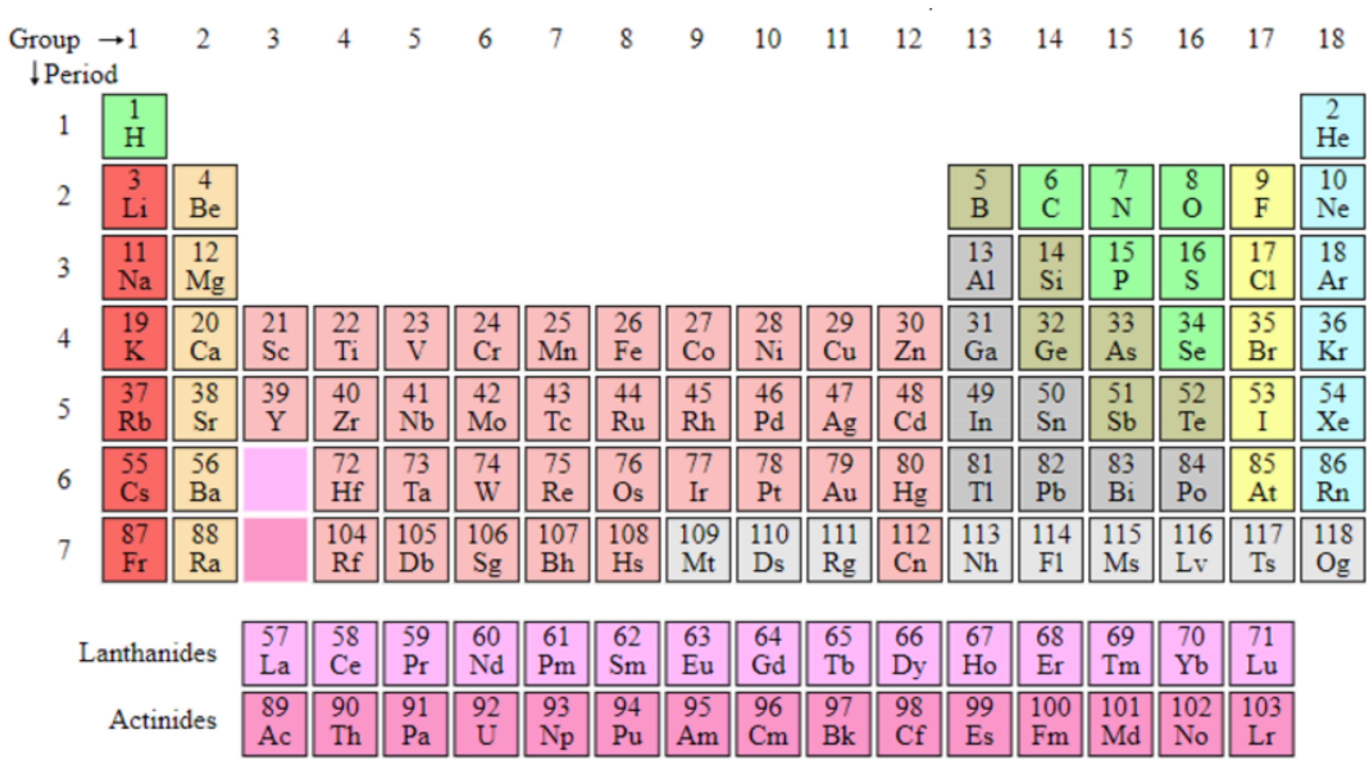
D) The effective nuclear charge experienced by the valence electrons in a chlorine atom is smaller than in a phosphorus atom.

E) All of the above statements are false.

2.10.5

Practice: Attracting Electrons

For an S-O bond, which atom will attract the electrons within the bond more strongly? Why?



- Sulfur, because it has a higher electronegativity
- Oxygen, because it has a higher electronegativity
- Sulfur, because it has a higher Zeff
- Oxygen, because it has more valence electrons
- Sulfur, because it has more protons
- Oxygen, because it is smaller

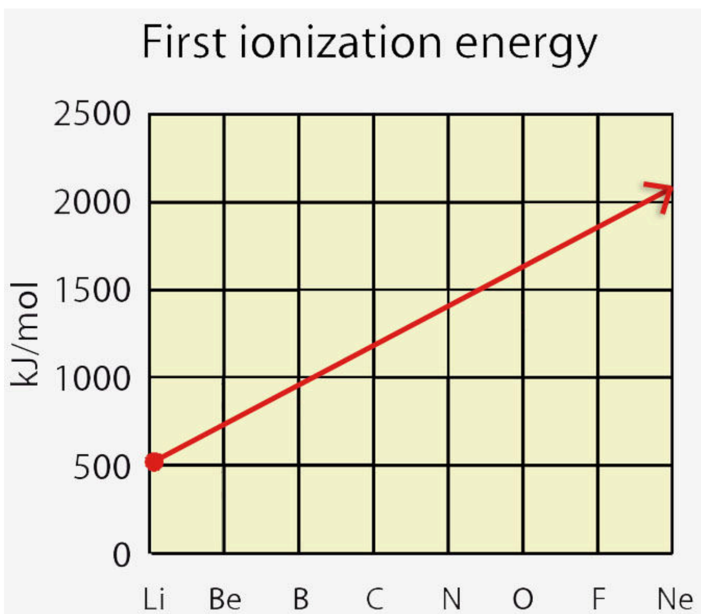
2.11 Exceptions to the Ionization Energy Trend

2.11.1

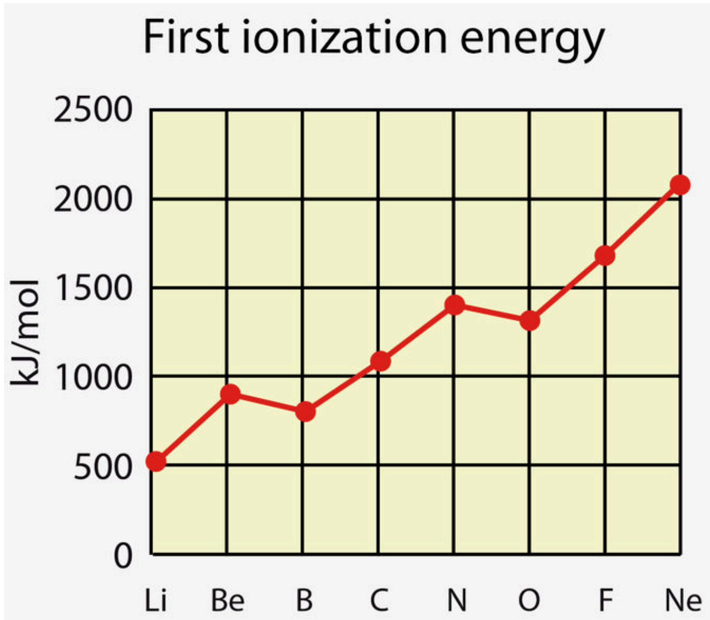
Exceptions to the Periodic Trends Introduction

There are a few exceptions for ionization energy and electron affinity that we should familiarize ourselves with.

Let's consider ionization energy first. **Based on what we currently know,** the trend for ionization energy should look like this:



However, it actually looks something like this:



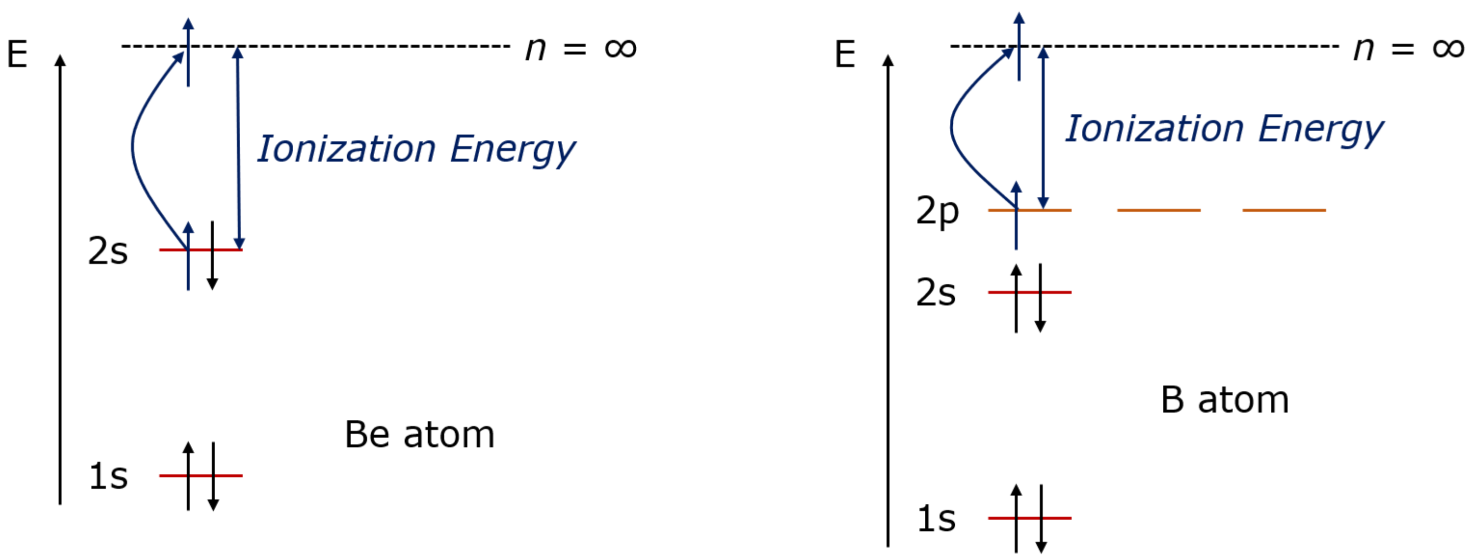
The differences have to do with **electron configurations.** Let's take a look.

2.11.2

Exceptions to the Periodic Trends Examples

Example #1:

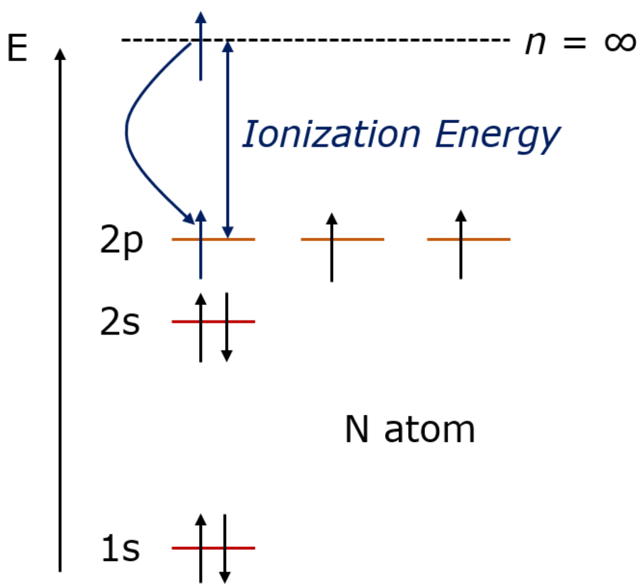
- Based on the periodic trend we learned so far, we would expect that Be or B would have the higher IE: _____

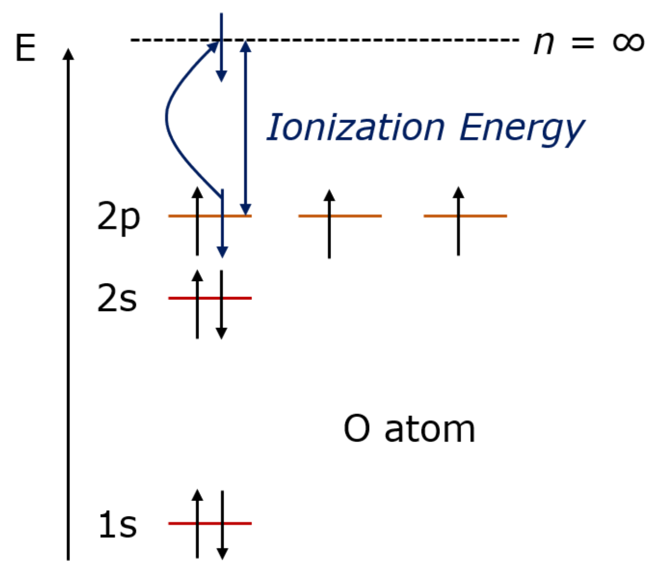


- Recall: which orbital is higher in energy, s or p? _____
- Based on the electron configurations and knowing that **B would LOVE to have its high energy p electron removed** so it can have its **valence electrons in the more stable 2s orbital**, should Be or B have the higher IE? _____

Example #2:

- Based on the periodic trend we learned so far, we would expect N or O to have the higher IE: ______

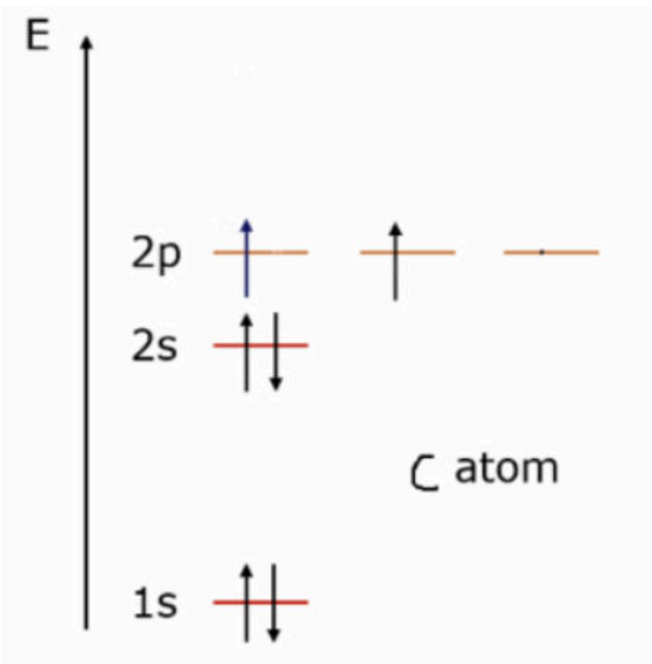


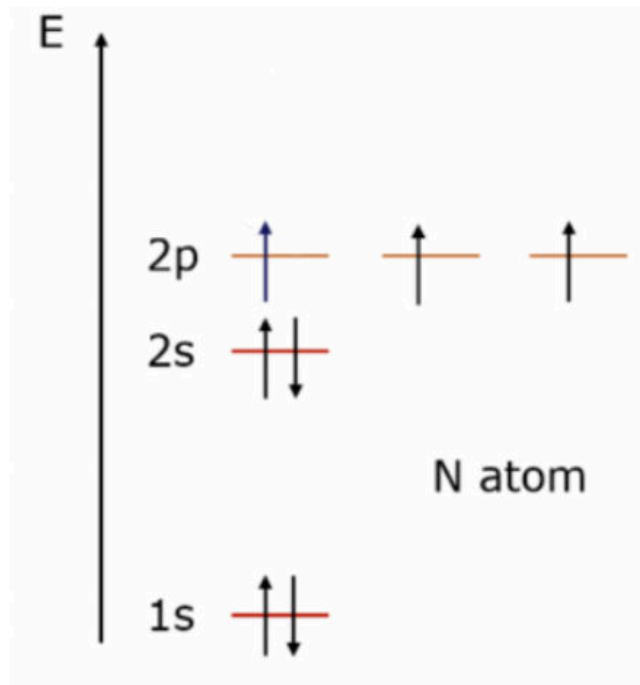


- Recall: are half-filled subshells more/less stable than partially filled subshells? ______________
- Based on the electron configurations and knowing that **N would HATE to have an outermost electron removed** since **it is currently very stable**, should N or O have the higher ionization energy? ______
- In addition, O has some electron-electron repulsion that it would prefer to get rid of. Thus, it makes sense that it has a higher/lower ________ IE

Example #3:

- Based on the periodic trend so far, we would expect C or N to have the higher negative EA: ______





- Based on the electron configurations and knowing that **N would HATE to have an electron added since it is currently very stable with a half-filled 2p subshell,** should N or C have the higher electron affinity? ______

2.11.3

Exceptions to the Periodic Trends Summary

WIZE TIP

The exceptions specifically happen in **two cases:**

- When we **switch from one type of orbital to another**
  Example: 2**s** to 2**p**; (Be to B is a jump from 2s to 2p)

- When we **start pairing electrons in a sub-shell**
  Example: $2p^3$ to $2p^4$ (N vs O)

These exceptions can be explained (and remembered) using the **relative energies of atomic orbitals (knowing that the p subshell is > in energy than the s subshell)** and the fact that electrons repel one another and this repulsion is highest for electrons sharing an orbital (**electrons would rather be half-filled in a subshell** instead of having some electrons paired and others unpaired

2.11.4

Example: Ionization Energy Exceptions

According to ionization energy (IE) trends, you would expect sulfur to have a higher IE than phosphorus; however, the actual ionization data shows the opposite: the IE of sulfur is 1000 kJ/mol and the IE of phosphorus is 1012 kJ/mol. Explain this phenomenon.

Practice: Ionization Energy Exception

Contrary to what we would predict based on $Z_{eff,}$ the ionization energy of aluminum is actually less than magnesium. This is because,

- The 3p orbitals of Magnesium are higher in energy than the 3p orbitals of Aluminum
- The 3s orbitals of Aluminum are higher in energy than the 3s orbitals or Magnesium
- The 3p orbital of Aluminum is higher in energy than the 3s orbital of Magnesium
- The 3s orbital of Magnesium is higher in energy than the 3p orbital of Aluminum
- The 3p orbital of Magneisum is higher in energy than the 3s orbital of Aluminum

2.12 The Periodic Table & Properties

2.12.1

The Periodic Table of Elements

The **periodic table** is something we are going to see a lot of in chemistry! It organizes the elements by their **atomic number (Z)** and is organized into **groups (columns)** and **periods (rows)**.

Elements in the **same group have very similar reactivity** which we will talk about more when we learn about things like valence electrons, bonding, and Lewis structures.

Group → ↓Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	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
5	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
6	55 Cs	56 Ba		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
7	87 Fr	88 Ra		104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Ms	116 Lv	117 Ts	118 Og

Lanthanides	57 La	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	89 Ac	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

You should be familiar with each of the following labels:

- Groups
- Periods
- Alkali Metals
- Alkaline Earth Metals
- Transition Metals
- Nitrogen group (aka Pnictogens)
- Oxygen group (aka Chalcogens)
- Halogens
- Noble (Inert) Gases
- Metals
- Non-metals
- Metalloids
- Lanthanides and Actinides (aka Rare Earth Metals)

Practice: Understanding Noble Gases

Which of the following statements is true?

- Noble gases are highly reactive since they want to obtain a full octet.
- Noble gases are highly reactive since they want to gain more electrons.
- Noble gases are unreactive since they already have a full octet.
- Noble gases are very stable because they want to gain more electrons