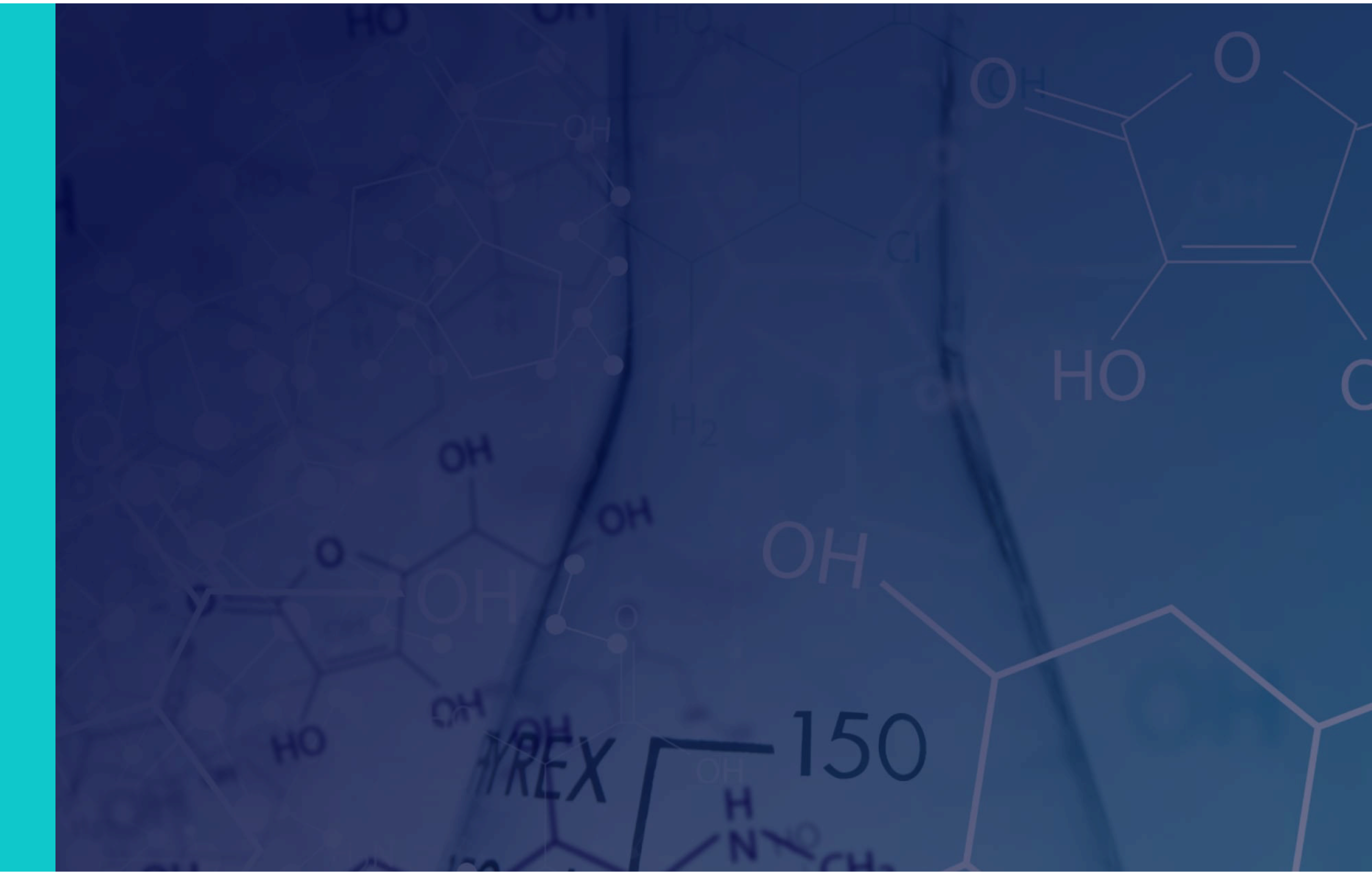


Mac

CHEM 1A03

Fall 2024, Chapter 3 Notes

Table of Contents

3. Unit 4: Periodic Table Trends

3.1 The Periodic Table & Properties

3.1.1

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)

3.1.2

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

3.2 Effective Nuclear Charge (Zeff)

3.2.1

Effective Nuclear Charge (Zeff)

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

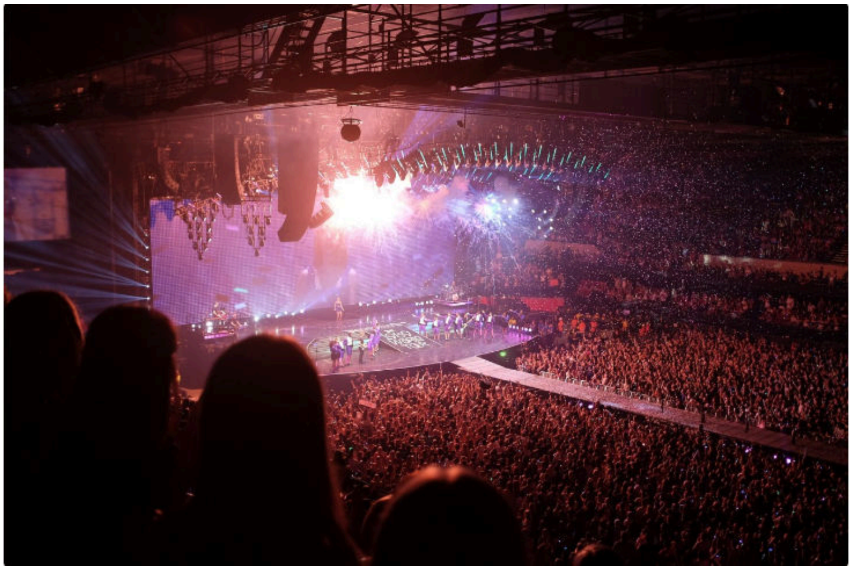


Photo by The Come Up Show / CC BY

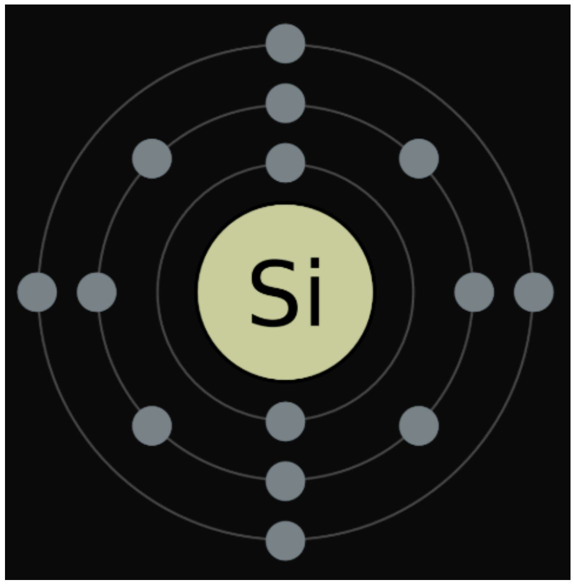
- People closer to the stage at a concert are going to be more into it. The music is louder and they really get a close connection with the artist. People in the highest rows in the stadium get less of that connection with the artist performing.
- This helps to explains the concept of nuclear shielding. **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?

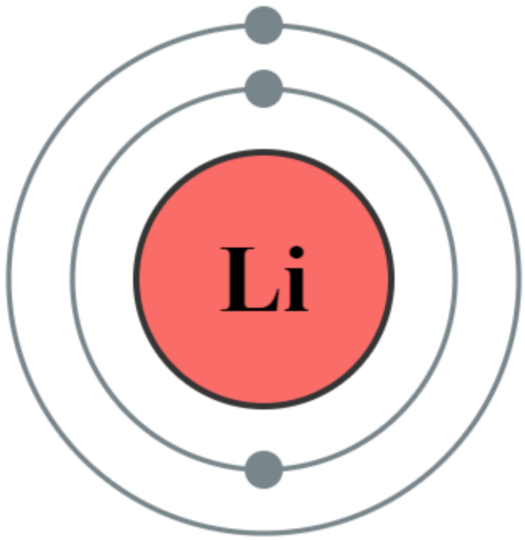
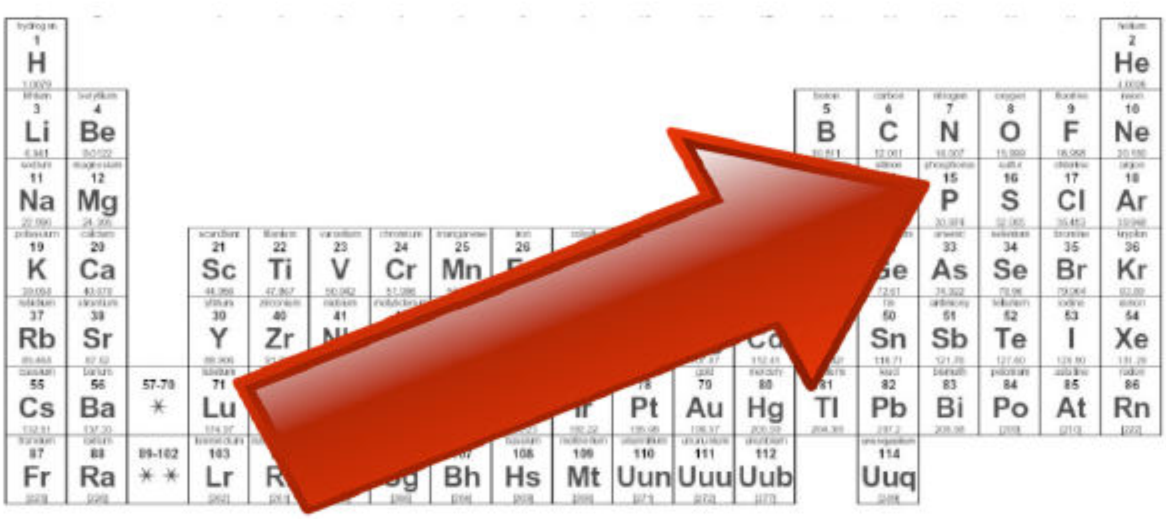


Photo by Greg Robson / CC BY

The Periodic Trend

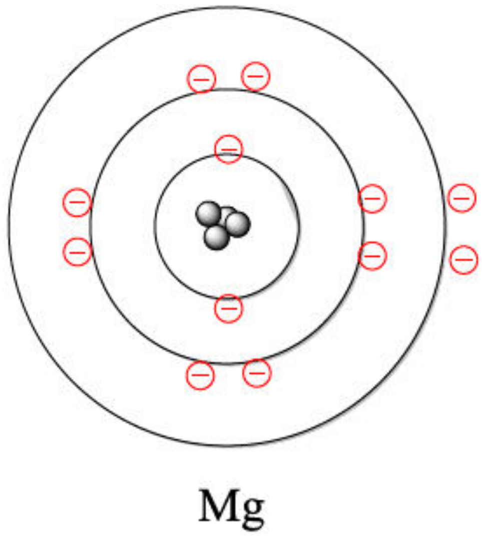
- As we **move to the right** across the periodic table, the # of core electrons stays the same but the **# of protons increases**. Therefore, **Z_{eff} increases.**
  - With protons, can pull the electrons in closer → Zeff
  - Effects of shielding are less with more protons

- 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

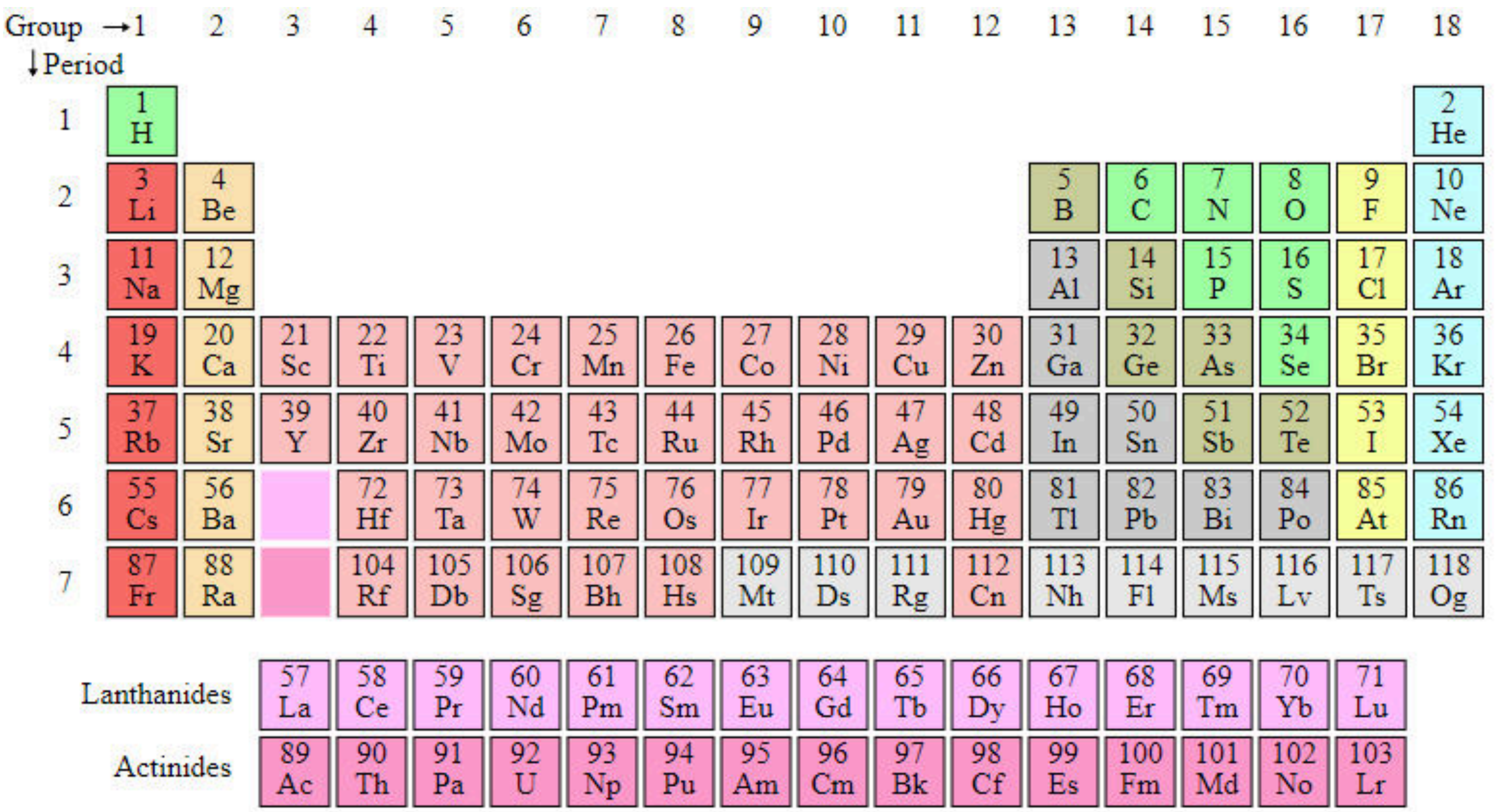
What is the Z_{eff} of an electron in the n=3 shell of magnesium (Mg)?



3.2.3 Practice

Practice: Estimating Zeff

Calculate the approximate effective nuclear charge of Se.



- 6
- 8
- 28
- 34

3.3 Atomic Radius Trend

3.3.1

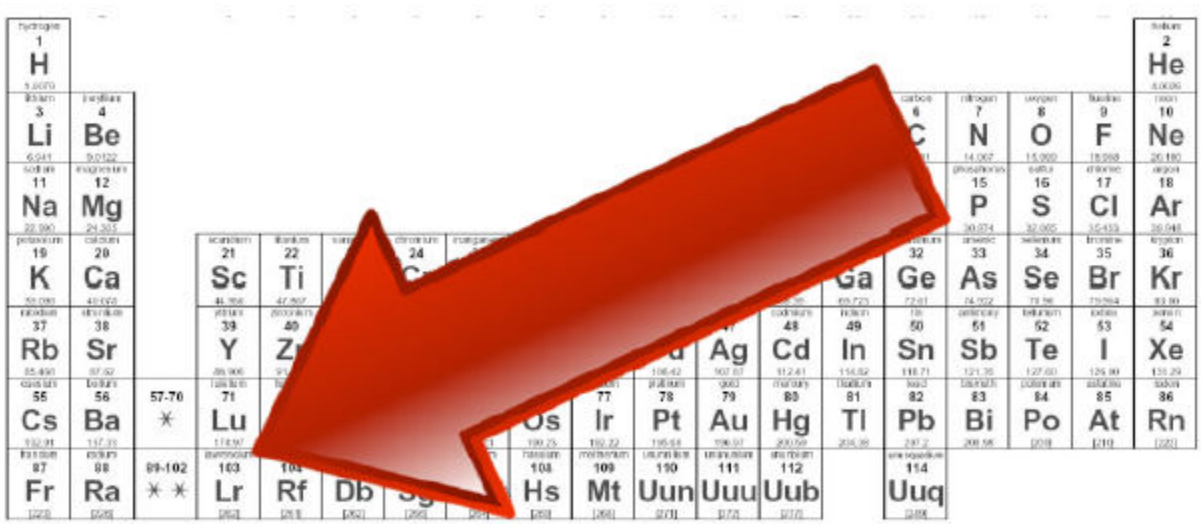
Atomic Radius

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:

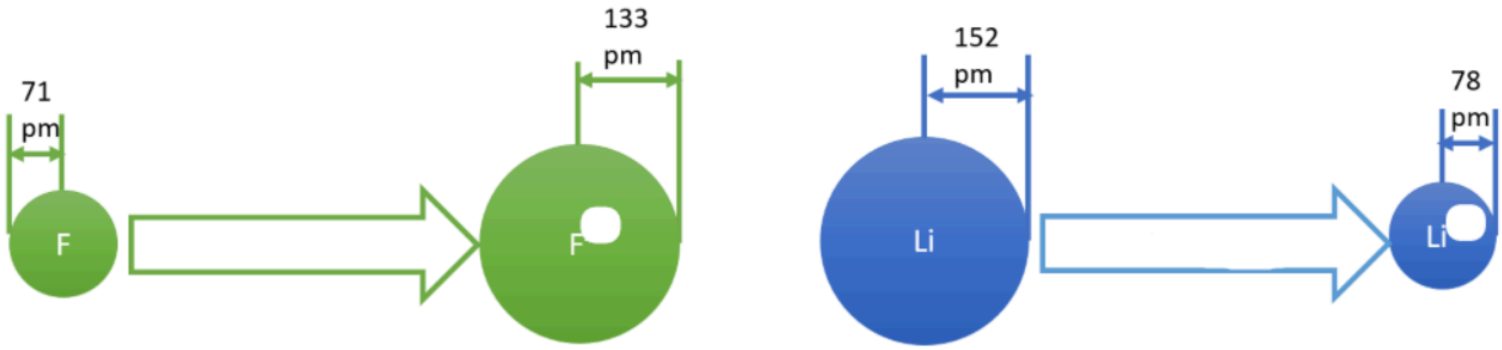
3.3.2

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+}

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
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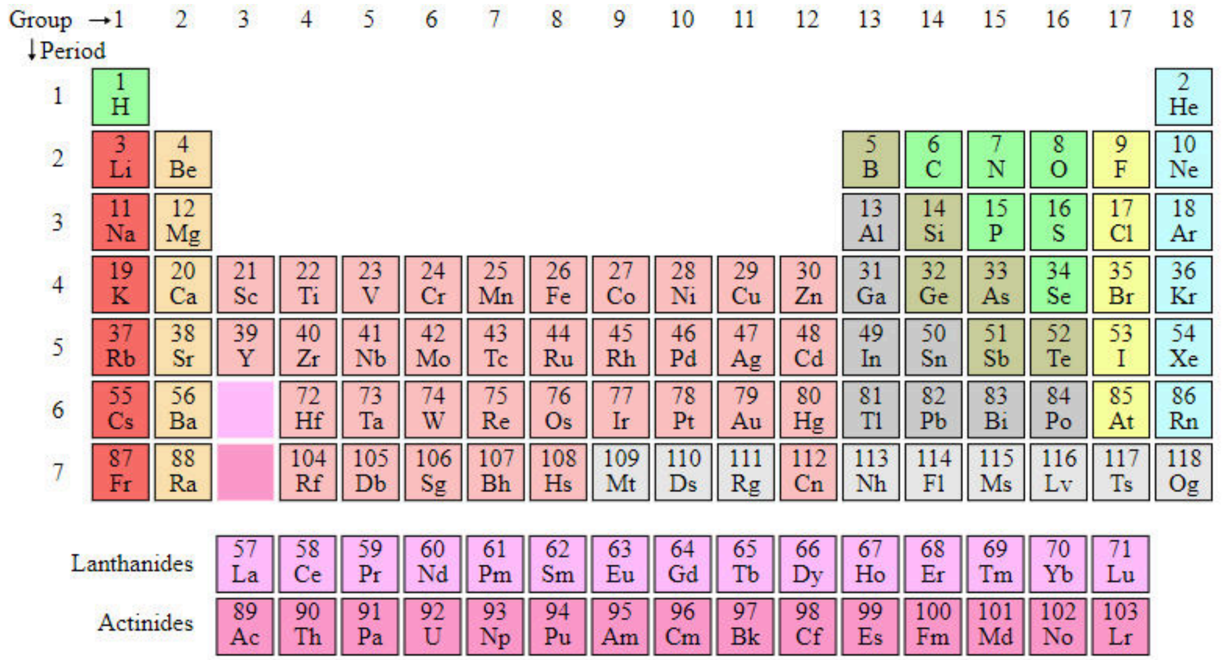
7	87 Fr	88 Ra		104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Ms	116 Lv	117 Ts	118 Og
Lanthanides			57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu	
Actinides			89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr	

Example: Ranking Size of Atoms

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



Group → ↓Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
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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

3.3.4

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

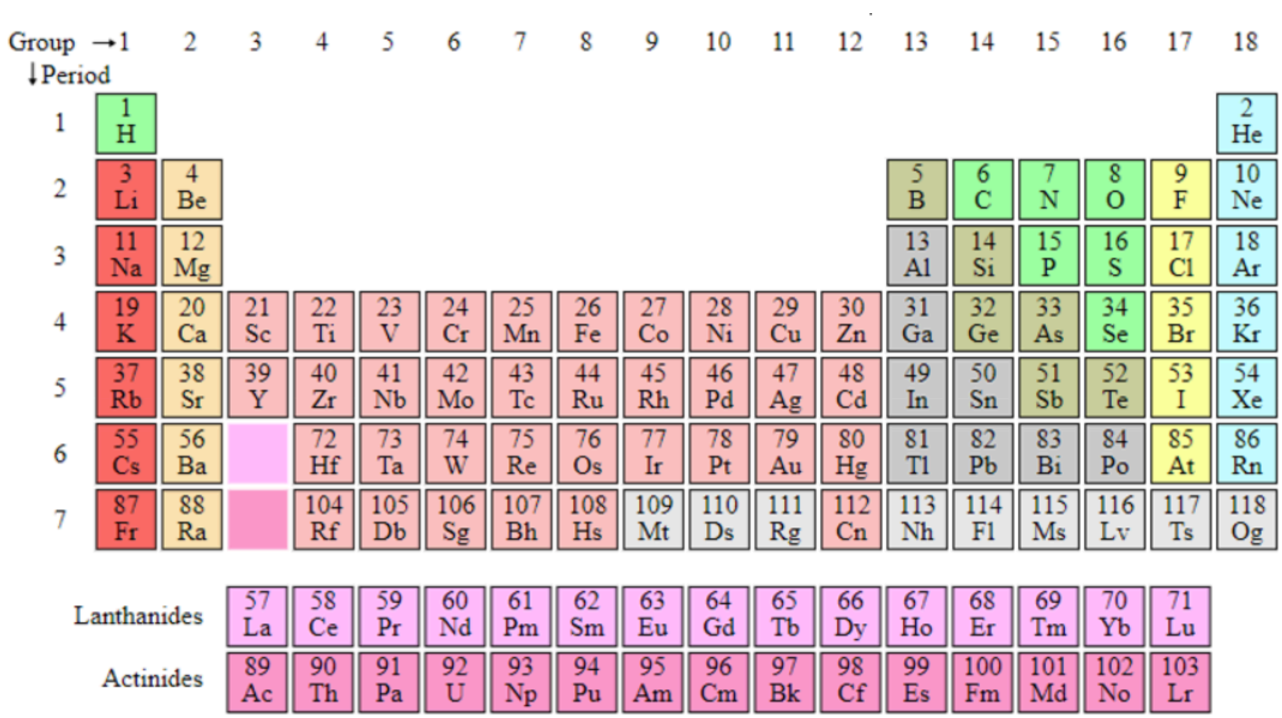
WIZE TIP

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**

3.3.5

Practice: Determining the Largest Atom

Select the atom below which has the largest atomic radius?



- S
- I
- O
- In
- Rb

3.4 Ionization Energy

3.4.1

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? ________

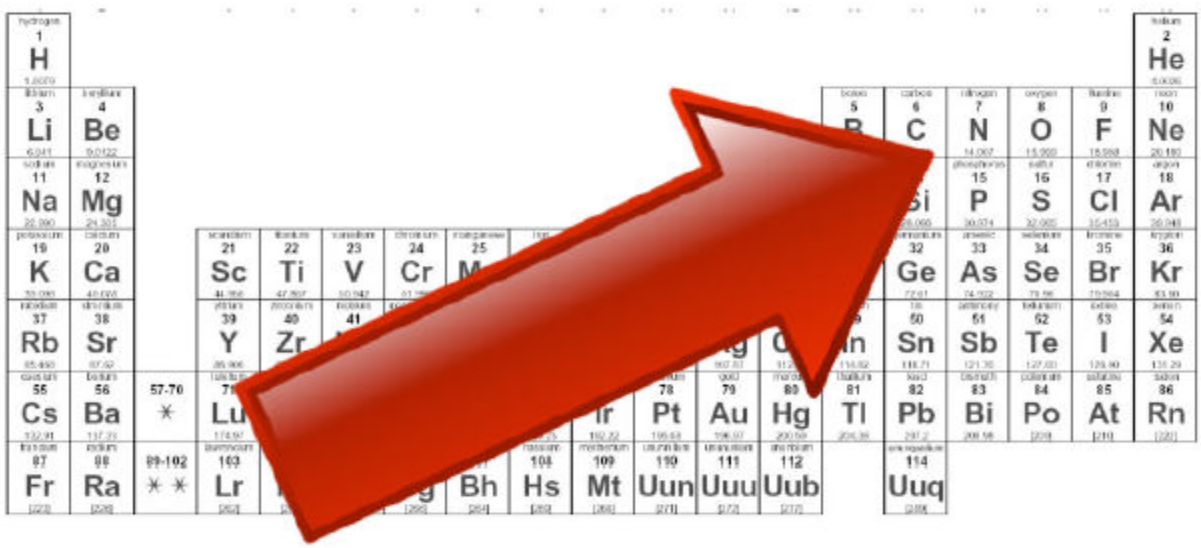
$$\mathbf{E} \rightarrow \mathbf{E}^{+} + \mathbf{e}^{-} \qquad \Delta \mathrm{E} = \text{Ionization Energy}$$

- **↑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
Pt
Au
Hg
Tl
Pb
Bi
Po
At
Rn
Fr
Ra
Lr
Bh
Hs
Mt
Uun
Uuu
Uub
Uuq

3.4.2

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

How does the first IE compare to the second IE?

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

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

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

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

Example:

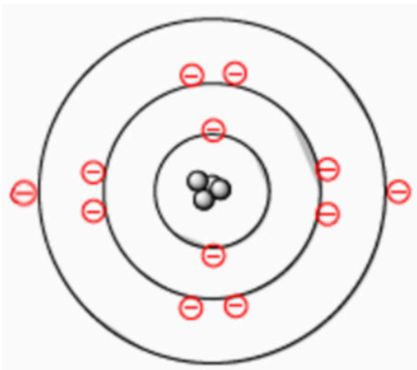
$$Ca\ (g)\ \rightarrow\ Ca^{+}\ (g)\ +\ e^{-}\quad IE_1\ =\ 590\ \frac{kJ}{mol}$$

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

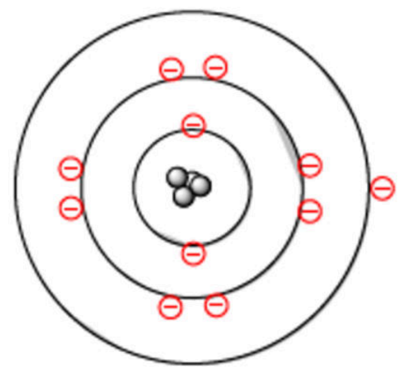
WIZE TIP

$IE_1 < IE_2 < IE_3 < IE_4$ etc

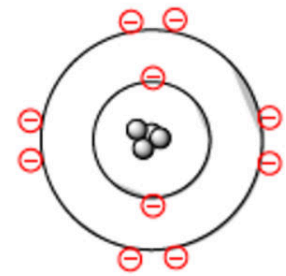
How does the IE_2 and IE_3 for magnesium compare?



Mg



Mg^+



$Mg2^+$

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

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: ________!

Practice: Highest 1st Ionization Energy

Which of the following species has the highest 1^{st} 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
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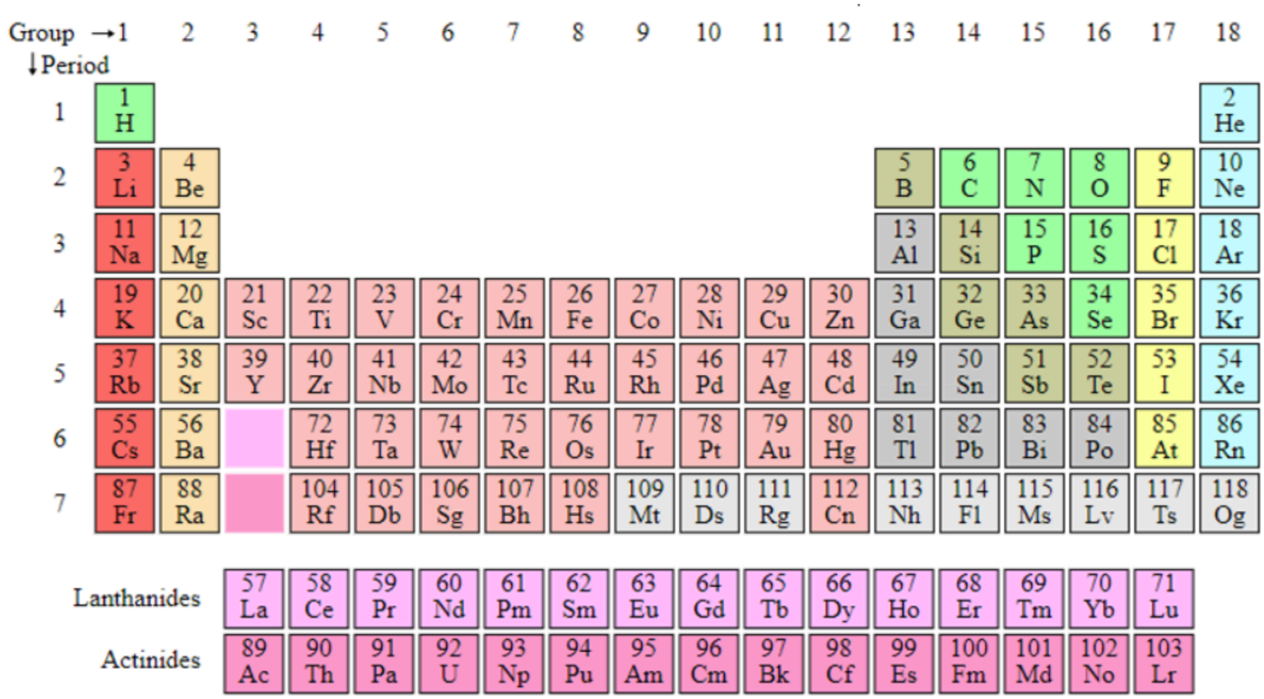
Actinides	89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

- S
- F
- B
- N
- P

3.4.4

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



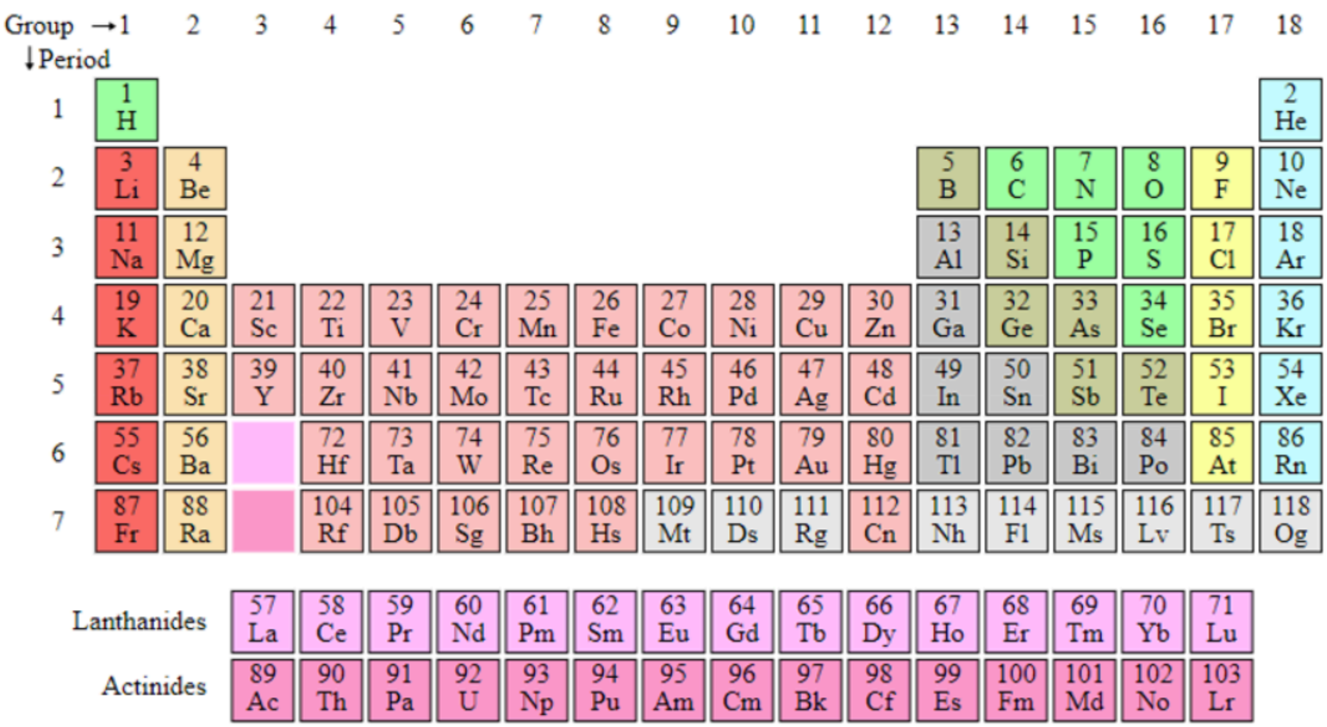
Fe	
O	
Ba	
Si	

V	
Zr	

3.4.5

Practice: Lowest Ionization Energy

Select the element with the lowest 3rd ionization energy.



A) Sc

B) Mg

C) Cl

D) I

E) C

3.5 Electron Affinity Trend

3.5.1

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? ________!

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

- **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? __________

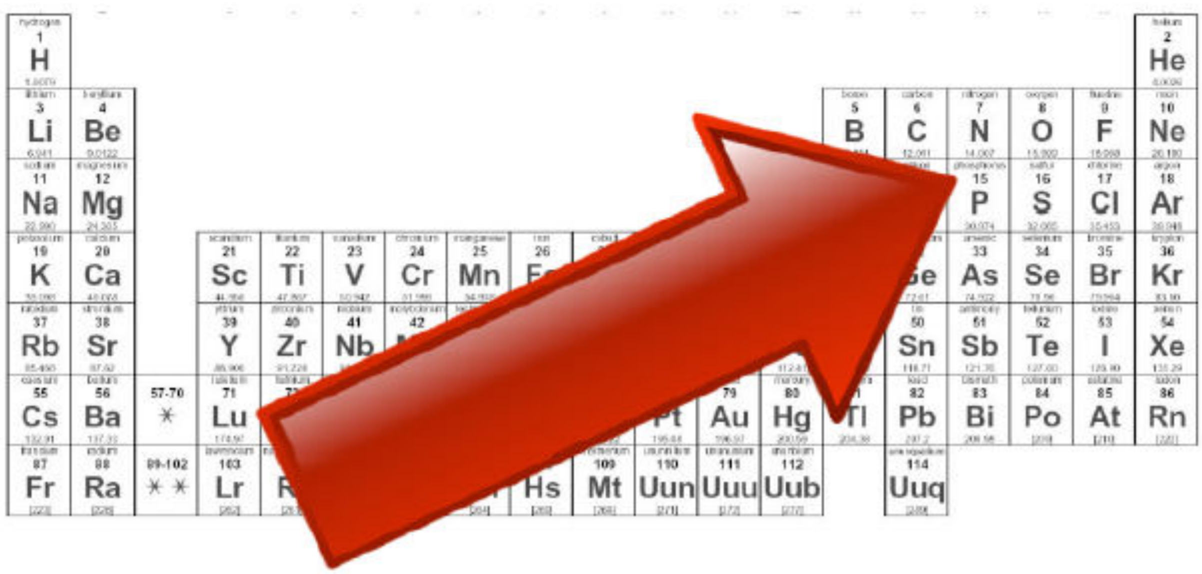
WATCH OUT!

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.**

3.5.2

Example: Increasing Electron Affinity

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

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
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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

3.5.4

Practice: Highest Electron Affinity

Which of the following has the highest electron affinity?

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

3.6 Electronegativity

3.6.1

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

Photo by DMacks / CC BY

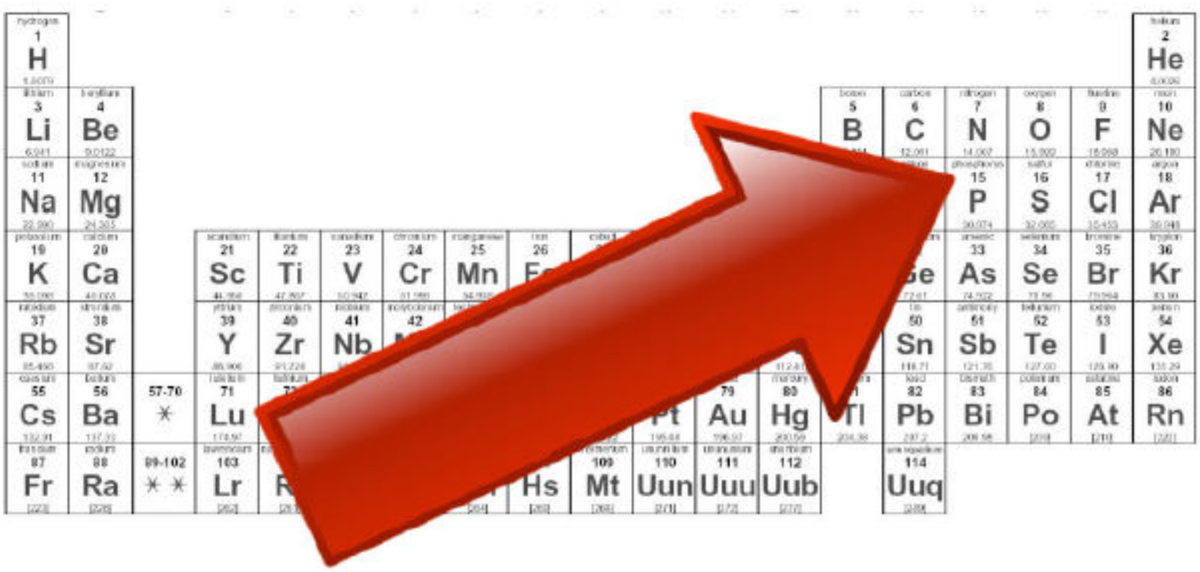
WIZE TIP

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.

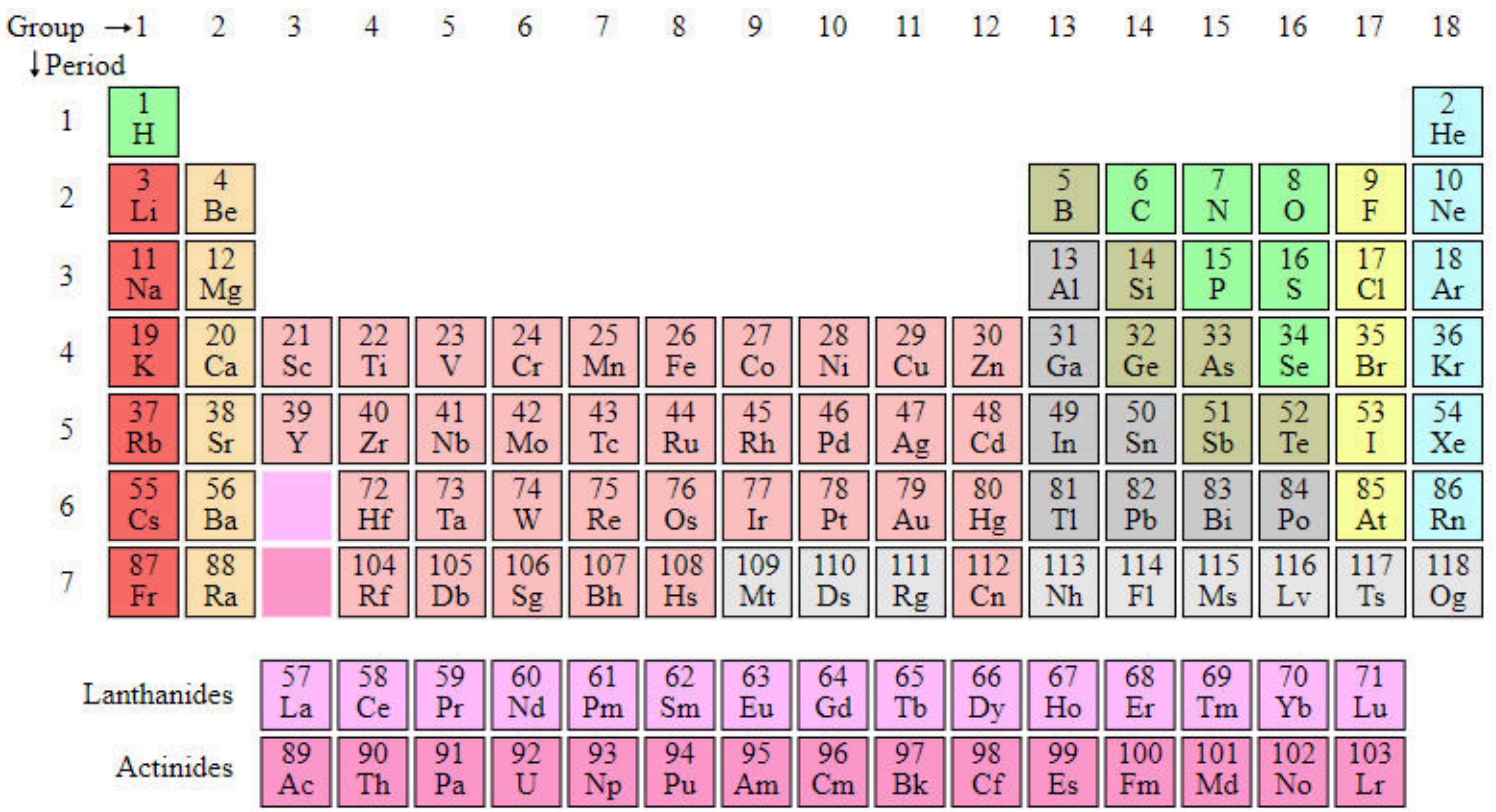


Which atom is the most electronegative? _____

3.6.2

Example: Increasing Electronegativity

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



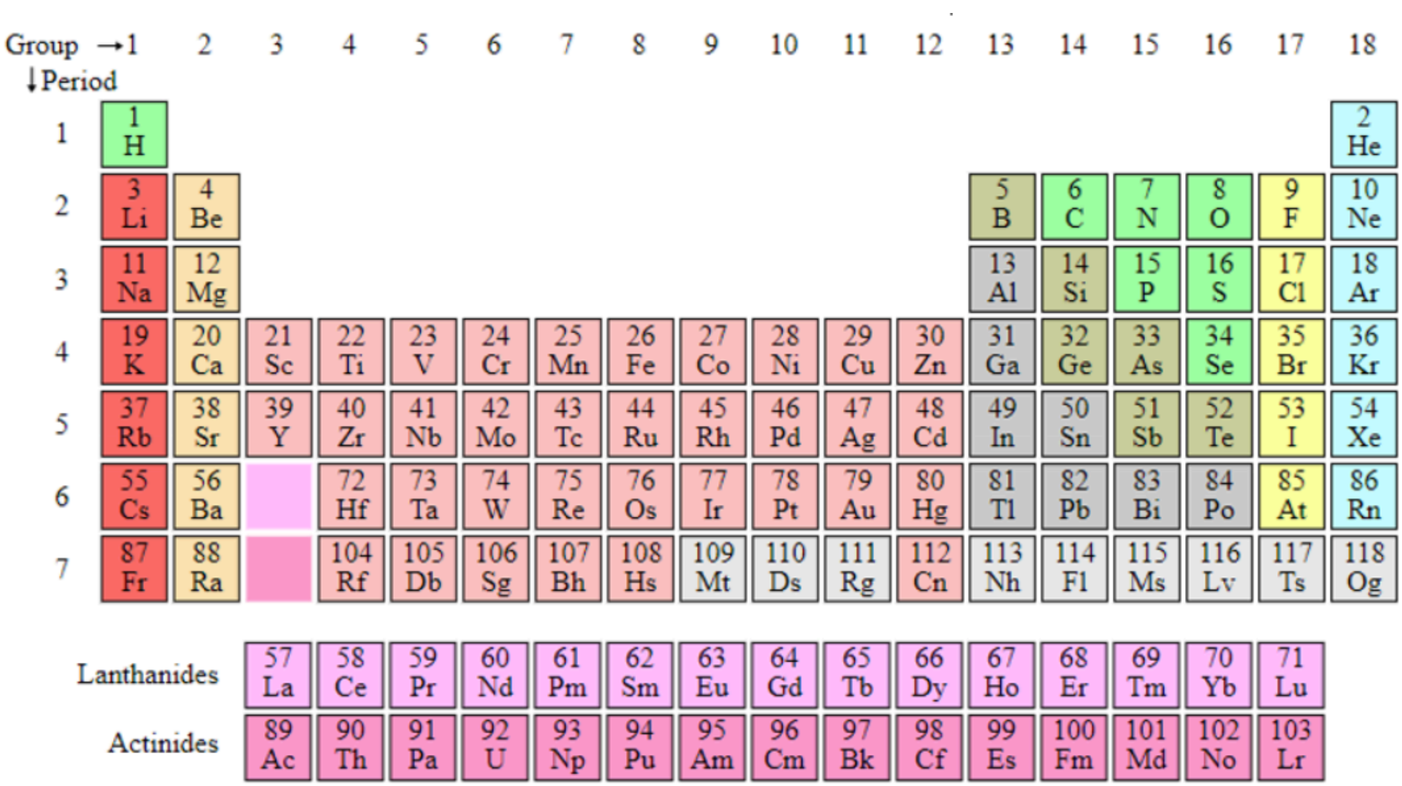
Group → / ↓Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 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

3.6.3

Practice: Ranking Electronegativity

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

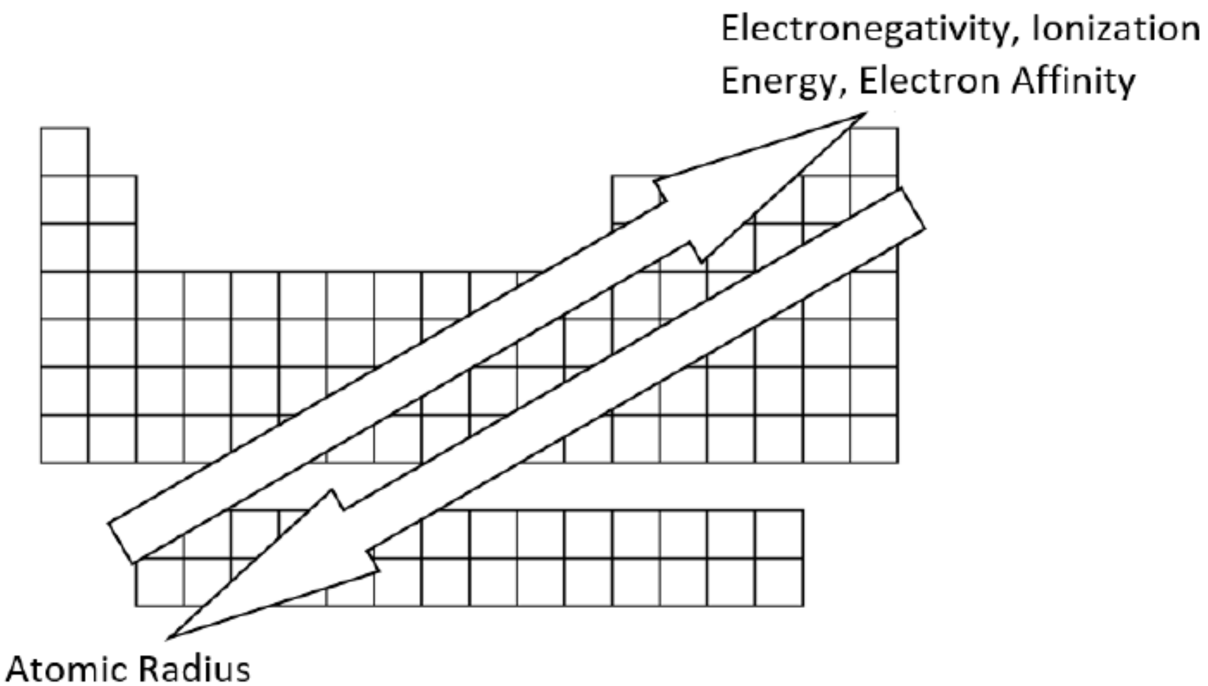


- Ti < Bi < Hg < Cs < Fr
- Na < Al < O < S < Cl
- Mg < P < C < O < F
- Fe < Co < Mo < F < Cl

3.7 Summary of Periodic Table Trends

3.7.1

Summary of Periodic Trends



Review: which direction does Zeff increase in? ____________________

3.7.2

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.

3.7.3

Practice: Periodic Trends

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

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

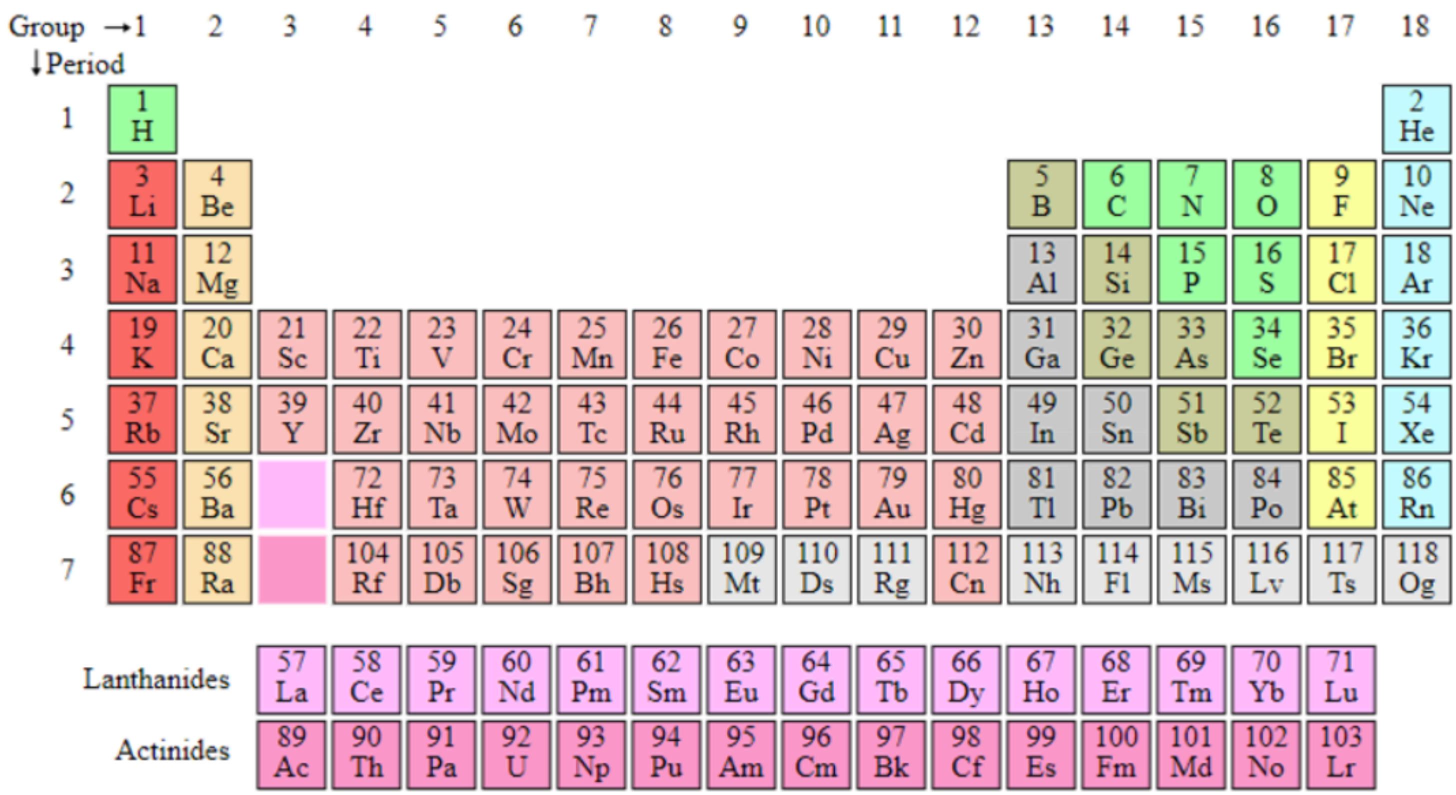
Part 1

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

3.7.4

Practice: Periodic Trends

Which of the following statements is correct about chlorine and phosphorus 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	

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.

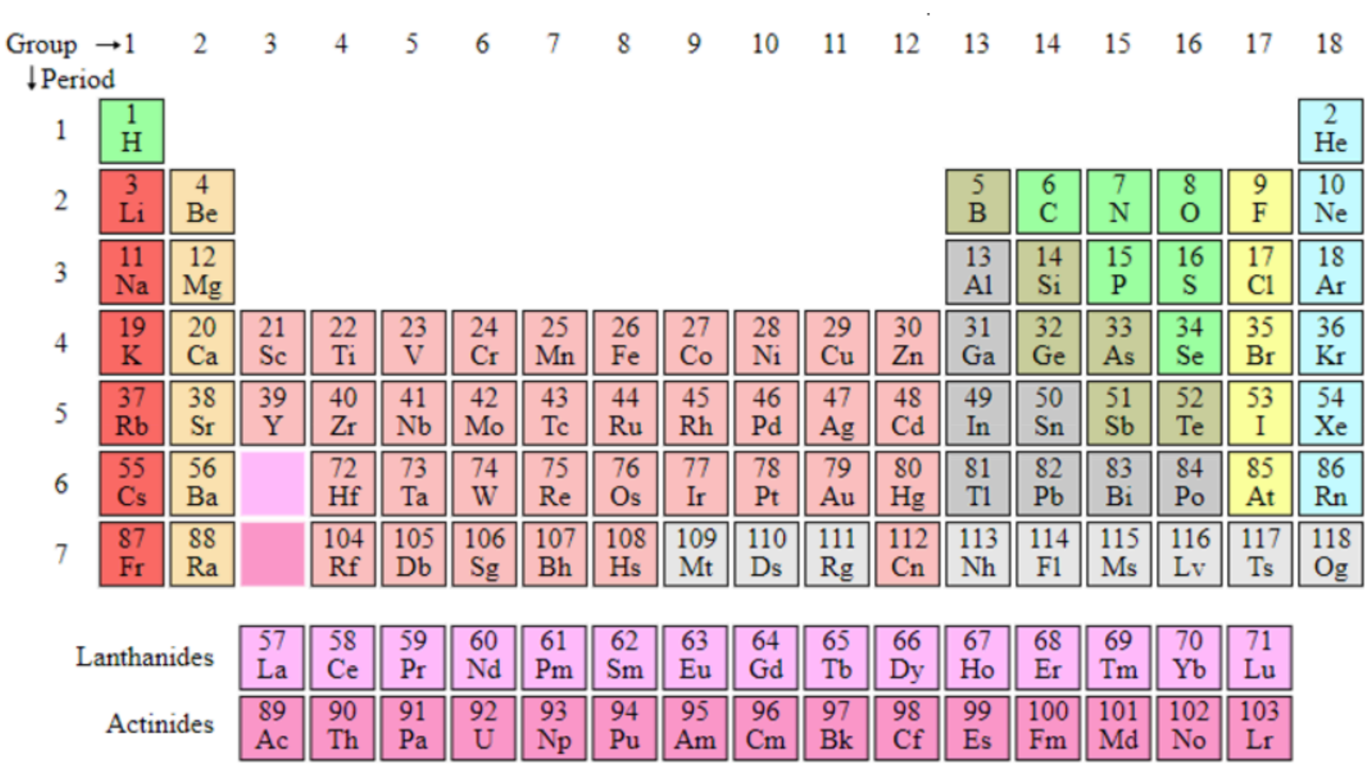
D) The effective nuclear charge experienced by the valence electrons in a chlorine atom is smaller than in a phosphorus atom.

E) All of the above statements are false.

3.7.5

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

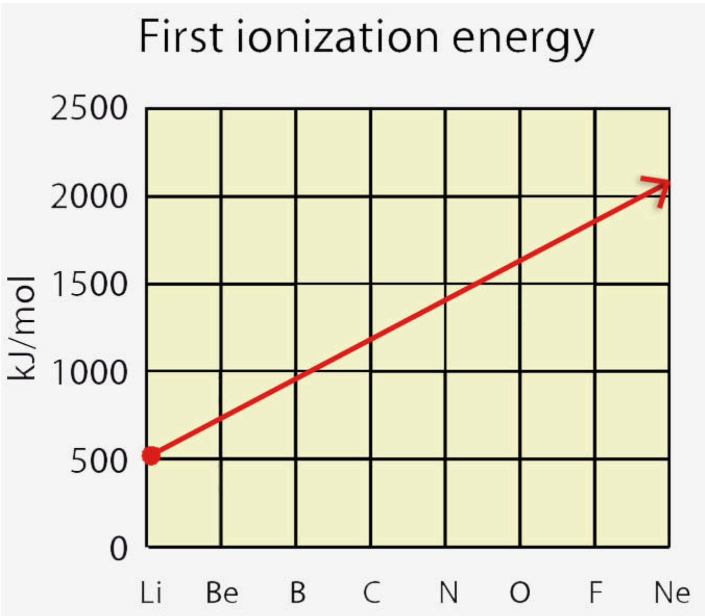
3.8 Exceptions to Periodic Trends

3.8.1

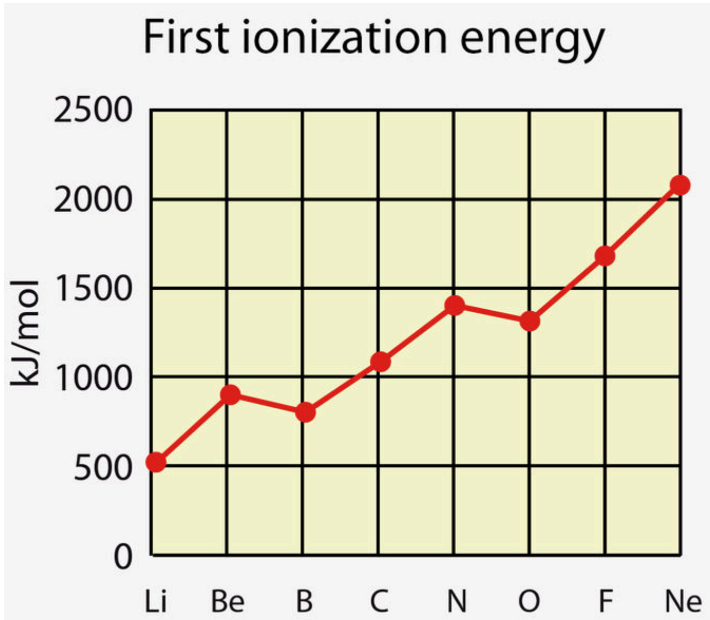
Exceptions to the Periodic Trends Introduction

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:

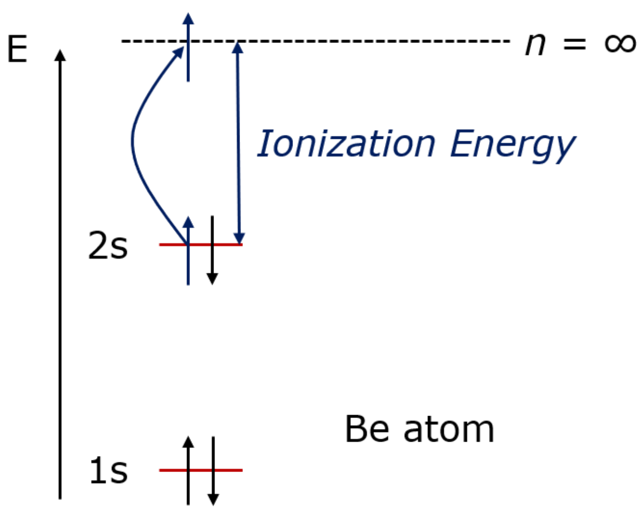


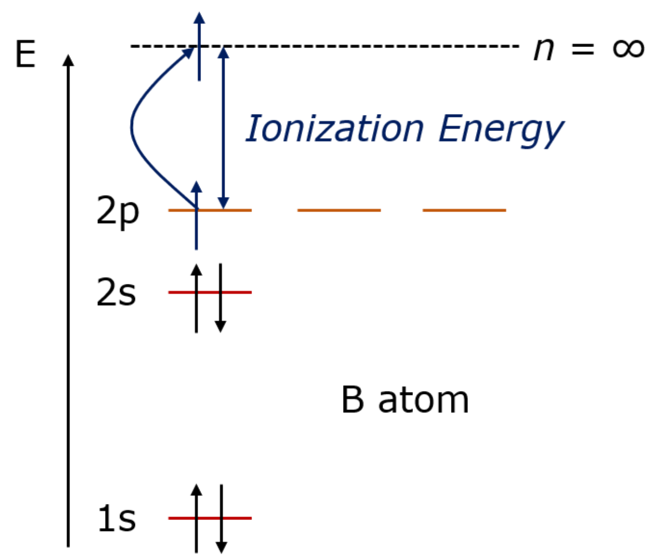
The differences have to do with **electron configurations.** Let's take a look.

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: ______

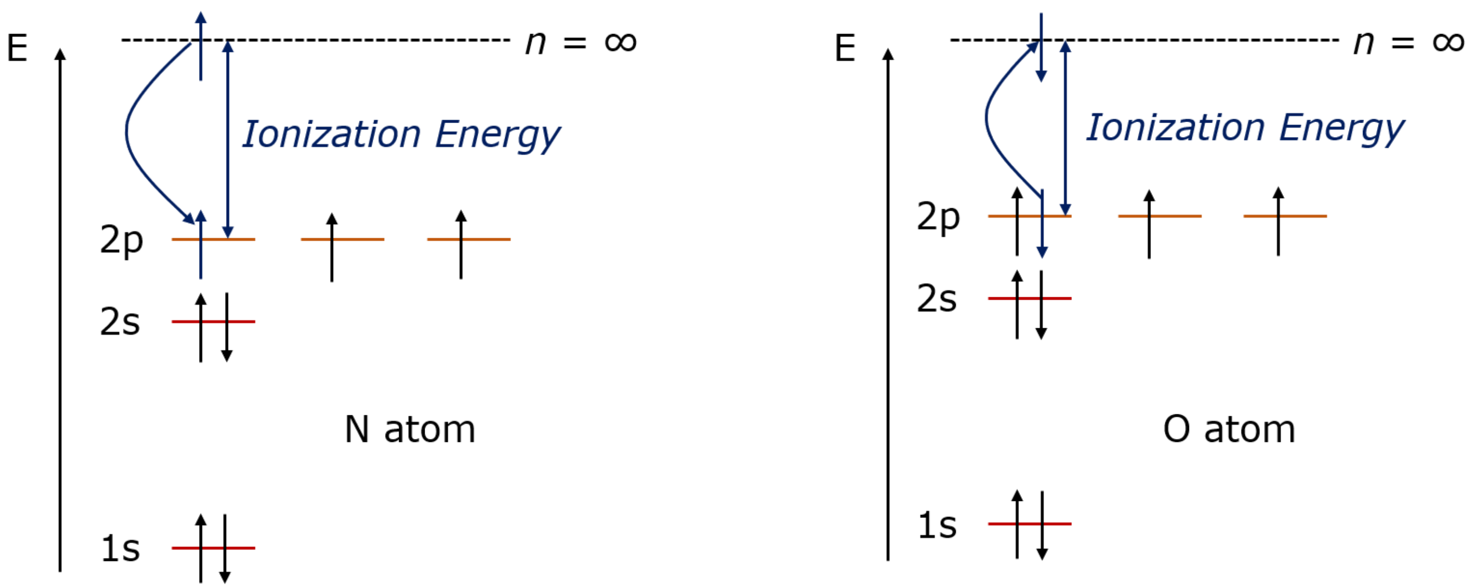




- Recall: which orbital is higher in energy, s or p? ______
- Based on the electron configurations and knowing that **B would LOVE to have its high energy p electron removed** so it can have its **valence electrons in the more stable 2s orbital**, should Be or B have the higher IE? ______

Example #2:

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

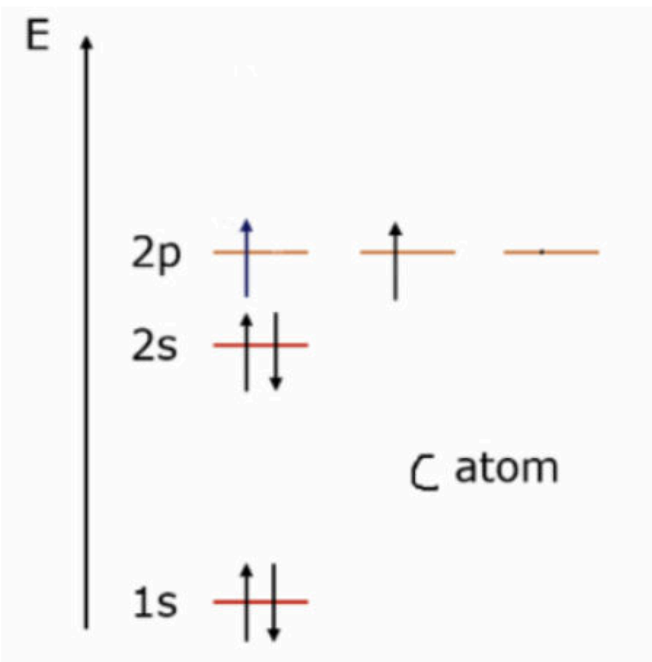


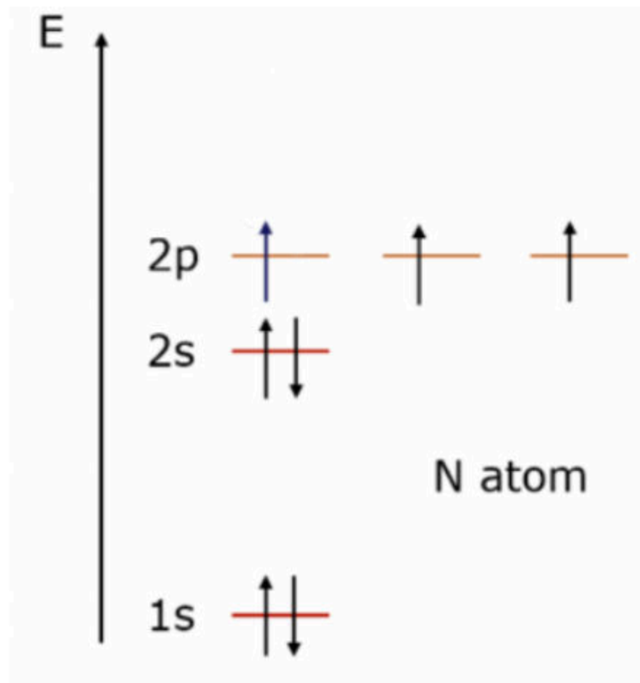
- Recall: are half-filled subshells more/less stable than partially filled subshells? ______________
- Based on the electron configurations and knowing that **N would HATE to have an outermost electron removed** since **it is currently very stable**, should N or O have the higher ionization energy? ______
- 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? ______

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

3.8.4

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.

3.8.5

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

3.9 Additional Info on Periodic Table and Reactions: s-Block

3.9.1

Where does Hydrogen Belong?

- Depending on the periodic table your looking at hydrogen is commonly included in group 1 or group 17, but you can imagine placing it in group 14 as well. So where should it go?
- Hydrogen doesn't fit well in any if these positions, it's a unique element.

	Group 1	Group 14	Group 17
Similarities	- Same electron configuration ns^1	-Similar electronegativity - Half filled valence shell	- Need one electron for a complete valence shell
Differences	-Not a metal -Very different electronegativity - Does not react like an Alkali metal	- Only one electron / orbital	- Very different electronegativity - Does not rect like a halogen

Element Hydrides

- Hydrogen forms element hydrides with all the main group elements with the exception of the noble gasses, Tl and In.
- These element hydrides are commonly divided up into ionic hydrides (metal hydrides) formed between hydrogen and the group 1 and group 2 metals and covalent hydrides formed between a p-block element and hydrogen.

	Metal Hydrides	Covalent Hydrides
Elements	Group 1 and group 2 with the exception of Be	The entire p-block with the exception of group 18, Tl and In.
Properties	- Ionic - White crystaline solids	- Covalent - Colorless gasses
Examples	CaH_2, KH, CsH, MgH_2	NH_3 CH_4 H_2S, HCl
H-Oxidation State	-1	+1

s-block Hydride Formation

Group 1 $$2\,M + H_2 \xrightarrow{\triangle} 2\,MH$$

Group 2 $$M + H_2 \xrightarrow{\triangle} MH_2$$

3.9.2

Alkali Metals

Elements of Group 1 (not including hydrogen) Li, Na, K, Rb, Cs, Fr

Physical Properties and Trends

- Generally soft metals
- Low mp/bp (decreases down group)
- Low denstiy (increases down group)
- Weak bonds (decreases down group)
- Large size (increases down group)
- Low ionization energy (decreases down a group)
- high reactivity (increases down a group)

Chemical Reactions

- Form ionic compounds with +1 oxidation states

$$2\ M_{(s)} + H_{2(g)} \longrightarrow 2\ MH_{(s)}$$

$$2\ M(s) + X_2(g) \longrightarrow 2\ MX(s)$$

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

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

3.9.3

Alkali Earth Metals

Elements of Group 2 Be, Mg, Ca, Sr, Ba, Ra

Physical Properties and Trends

- Very similar to the Alkali metals but everything is less extreme
- Low mp/bp (decreases down the group, but not as low as group 1)
- Low denstiy (increases down the group, but not as low as group 1)
- Weak bonds (weaker down the group, but stronger than group 1)
- Large size (increases down the group, but not as large as group 1)
- Low ionization energy (decreases down the group, but not as low as group 1)
- Reactivity increases down a group

Chemical Reactions

- Form ionic compounds with +2 oxidation states

$$M_{(s)} \quad + \quad H_{2(g)} \longrightarrow MH_{2(s)}$$

$$M(s) \quad + \quad X_2(g) \longrightarrow MX_{2(s)}$$

$$M_{(s)} \quad + \quad O_{2(g)} \longrightarrow 2\ MO_{(s)}$$

$$M_{(s)} \quad + \quad H_2O_{(l)} \longrightarrow 2\ M(OH)_{2(aq)} \quad + \quad H_{2(g)}$$ - Except Be

Importance of Magnesium

- Mg is an essential element in many living systems, including humans.
- In biological systems magnesium exists as Mg^{2+}
- The role of Mg^{2+} is often to keep small organic molecules or proteins rigid.
- Example: One Mg^{2+} ion keeps chlorophyll rigid

3.9.4

Group 1 vs Group 2

Physical Properties

- Group 1 elements have lower Ionization Energy
- Group 1 elements have lower mp/bp
- Group 1 elements are more reactive
- Group 1 elements have lower density
- Group 1 elements form weaker bonds
- Group 1 elements are larger

- Group 2 elements form compounds which have much larger lattice energies because they have greater positive charges and are smaller than group 1 ions. This leads to:
  - Group 1 salts are more soluble than group 2 salts
  - Group 2 salts have stronger ion dipole forces. This is why group 2 salts commonly exist as hydrates, ex. $MgSO_4 \cdot 7\ H_2O$

Beryllium and Lithium

- The lightest entry for the alkali and alkali earth metals behave slightly differently than their heavier congeners.
- For example Be does not react with H_2O like the rest of the elements in group 2 and while Li does react with H_2O it is much less violent than the other elements of group 1
- Be actually forms covalent bonds unlike all the other elements in group 2
- Whats going on with the 2nd row?

The Diagonal Relationship

- The elements in the 2nd row actually behave more like elements one row down and one group right due to their similar size with those elements. Eg. Li and Mg, Be and Al, B and Si
- This is why Li reacts less violently (like a Alkali earth metal) with water and why Be forms mostly covalent bonds like Al

3.9.5

Rationalize the fact that although sodium is a lustrous (shiny) metal sodium often appears dull and grey in the laboratory.

- Sodium is a group 1 metal, these metals are very soft and easily scuffed which gives a dull appearance
- Sodium has a low melting point so the surface of sodium melts and re-solidifies frequently, this produces a rough surface which makes the sodium appear grey.
- Sodium is prone to oxidation and when it self-ionizes it produces Na^+ cations that are grey in color and coat the surface
- The surface of sodium metal can react with oxygen and water in the air to form Na_2O and NaOH which are white solids
- Sodium can display a variety of colors which depend on the temperature of the room it is being stored in

3.9.6

Sodium sparks and smokes when it comes into contact with water where as potassium reacts much more violently and immediately catches fire? The difference between these two metals can be best explained by

- Potassium is substantially more reactive because of its electron configuration
- Potassium is one row lower on the periodic table, it has a lower ionization energy so it reacts faster and more violently
- Potassium is larger than sodium, therefore the electrons are orbiting faster and react more vigorously
- Sodium is less reactive because it has a lower ionization energy
- Sodium is less reactive because it forms a more stable covalent bond with oxygen than potassium

3.10 Additional Info on Periodic Table and Reactions: p-Block

3.10.1

General Trends in p-block Bonding

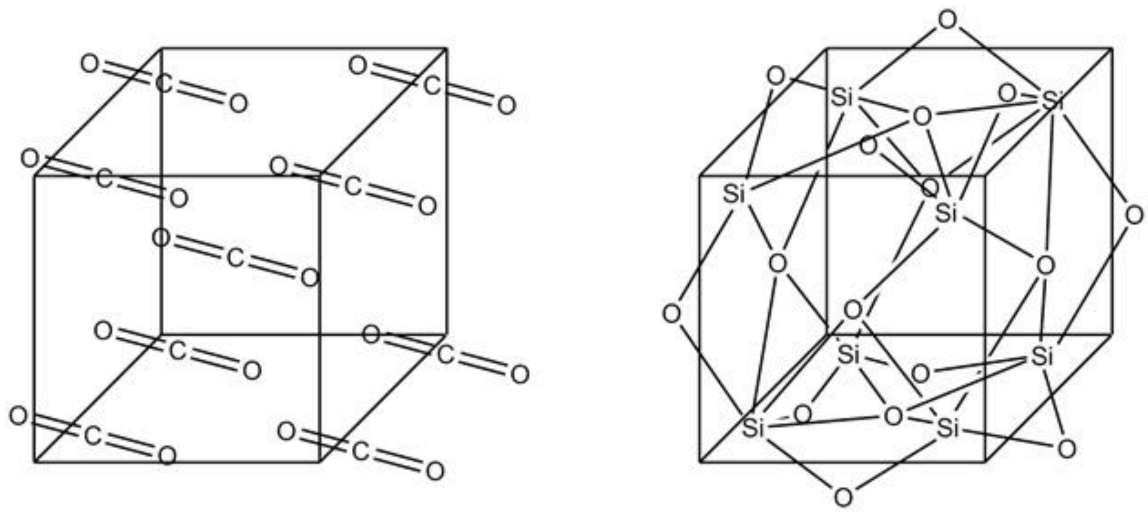
- Period 2 elements strictly follow the octet rule (Boron can have 6) meaning neutral carbon always forms 4 bonds.
- Period 3 and lower can expand there octet by using their empty 3d orbitals
- Larger elements can form more bonds because there is more room around the atom.

Ex.

Most stable Lewis Structures for $NOCl_3$ and $POCl_3$

- Period 2 elements are smaller than Period 3 and lower elements and therefore there bond lengths are shorter.
- Elements in period 2 commonly form double bonds and elements in period 3, which are held further apart due to their size, have very little orbital overlap when trying to form a πbond.

Ex.



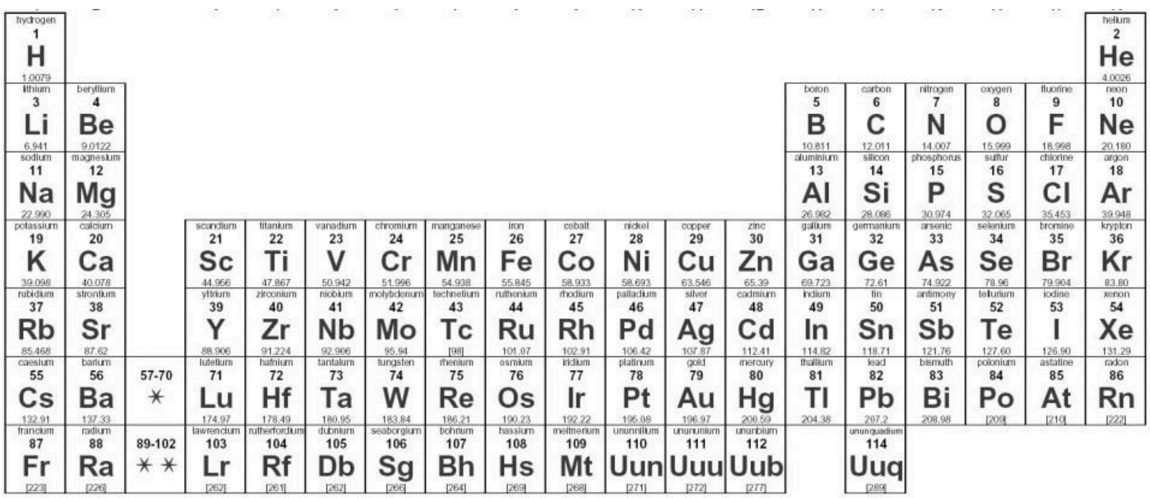
CO_2 vs SiO_2

Group 13

Elements of Group 13: B, Al, Ga, In, Tl, Uut

Physical Properties and Trends

- No trend in mp
- bp decreases down the group
- density increases down the group
- Atomic size generally increases down the group
- There is more of a difference in atomic size between group 2 and group 13 for Ga, In, Tl, Uut because of the d and f block elements which separate them. There are many more protons in the nucleus of these elements and the d/f electrons do not shield well. Therefore we see a greater contraction from group 2 to 13 for elements in rows 4, 5, 6 and 7

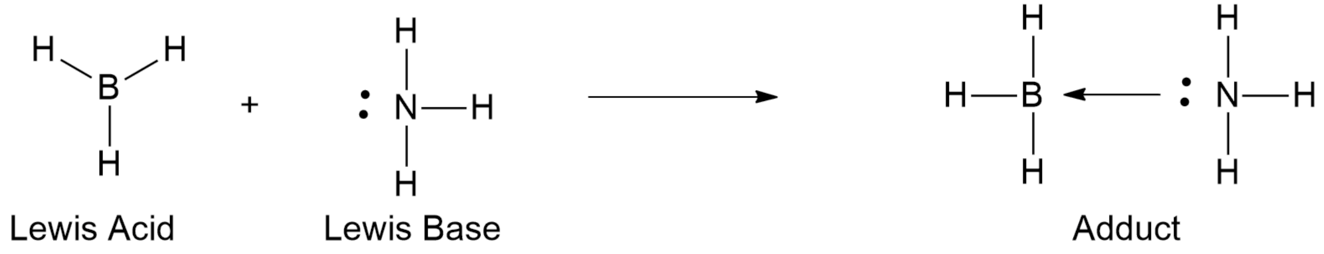


Oxide Formation

- Form compounds with +1 or +3 oxidation states
  - B_2O_3 - Acidic
  - Al_2O_3 - Amphoteric
  - Ga_2O_3 - Basic
  - In_2O_3 or In_2O - Both Basic
  - Tl_2O - Basic

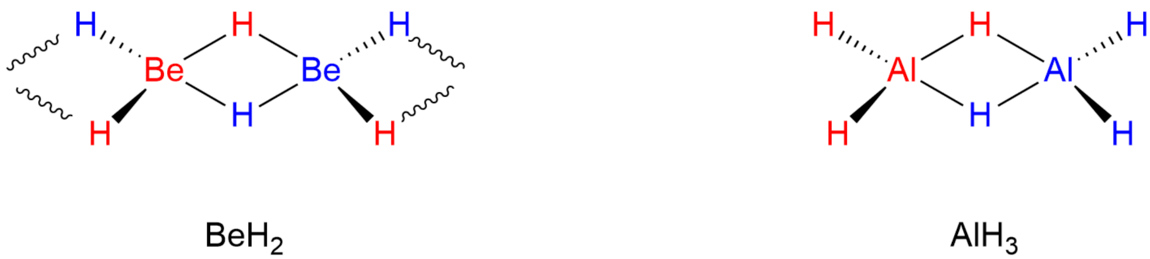
Boron as a Lewis Acid

Unlike the other elements from group 13 Boron and aluminum forms covalent bonds. Commonly boron has 6 electrons and is neutral, because boron is electrons deficient it can act as a Lewis acid accepting electrons from a Lewis base.



Diagonal Relationship between Be and Al

- Both form $M(OH)_4^{n-}$ salts in strong base { $Be(OH)_4^{2-}$ and $Al(OH)_4^{-}$}
- Primarily covalent bonds with non-metals
- Forms bridging bonds



3.10.3

Group 14

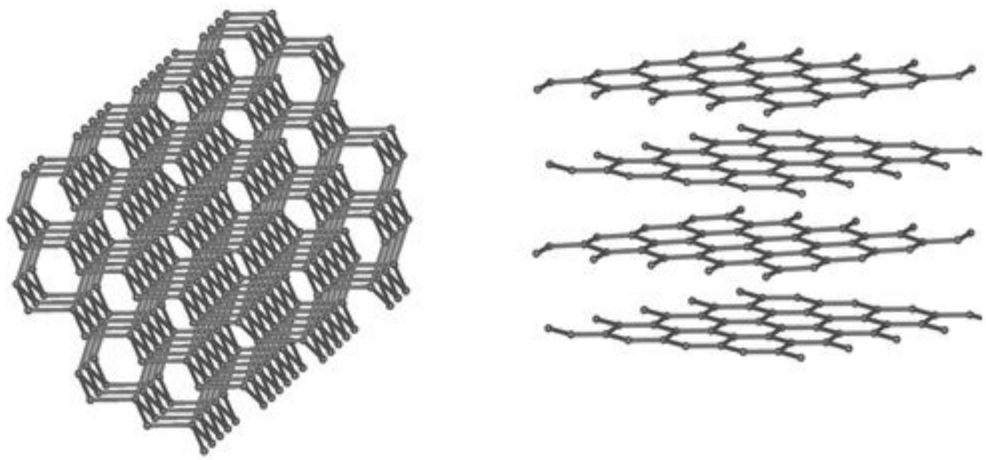
Elements of Group 14: C, Si, Ge, Sn, Pb

Physical Properties and Trends

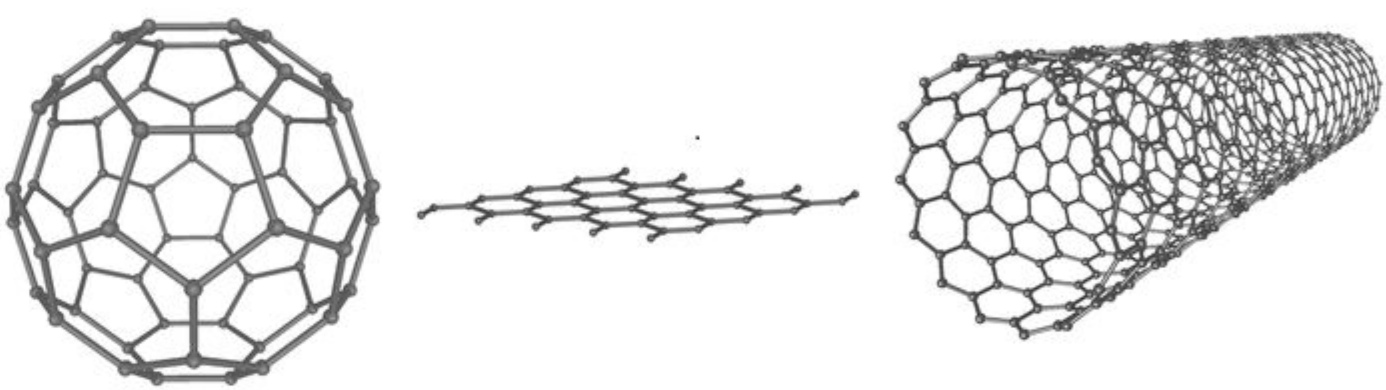
- mp/bp decreases down the group
- density increases down the group
- Atomic size increases down the group
- IE are higher than for group 13 and are similar up and down the group
- Common oxidation states are +4, -4 and +2

Allotropes of Carbon

- An allotrope is different form of the same element in the solid phase.
- Sn has two allotropes white and grey. Grey tin is more stable below room temperature and is brittle, white tin is more stable at room temperature and above and is malleable
- Carbon has several allotropes: Graphite, Diamond, nanotubes, buckminsterfullerene, graphene



Graphite (left) and Diamond (right)



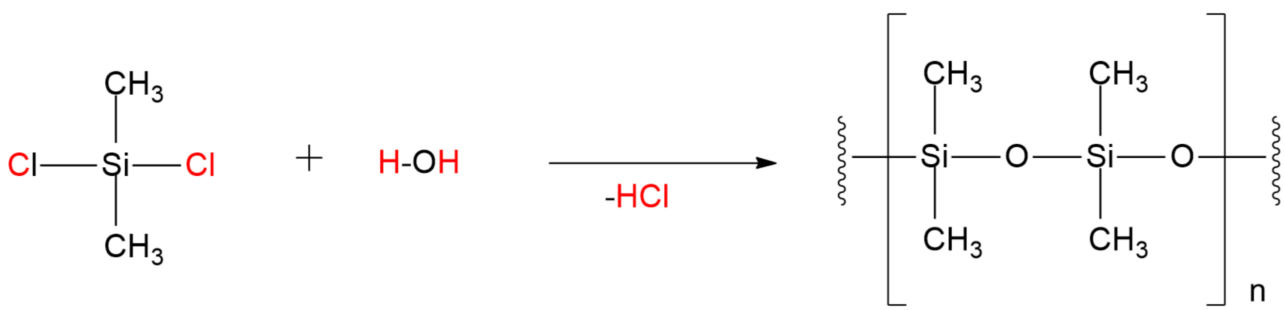
Buckminsterfullerene (left) graphene (top right) and a carbon nanotube (bottom right)

Carbon an Anomaly

- Carbon is the only non-metal in group 14
- Carbon displays very strong bonds to itself
- Carbon displays a wide array of bond geometries
- There is a whole field of chemistry devoted to the study of carbon compounds, Organic Chemistry!

Silicon, a very useful element

- Silicon has some very important applications.
- Silicon forms very weak bonds to itself (unlike carbon) but very strong bonds to oxygen
- Silicon compounds can form silicones as shown below



- Silicon and silicates are semi-conductors and are used in nearly every computer chip on the market!
- **Diagonal Relationship:** boron and borosilicates are also semi-conductors and can be used in electronics

3.10.4

Group 15

Elements of Group 15: N, P, As, Sb, Bi

Physical Properties and Trends

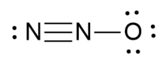
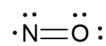
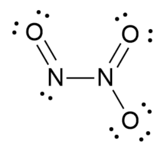
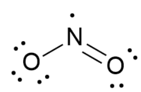
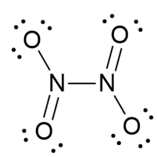
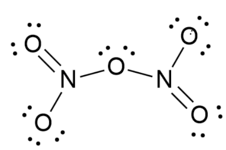
- mp increases up to As then decreases down the group
- bp increases down the group
- density increases down the group
- Atomic size increases down the group
- IE decrease down the group
- Common oxidation states are +5, -3 and +3

Group 15 Hydrides

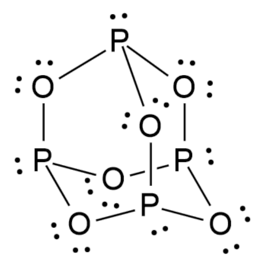
- NH_3 - Oxidation state -3
- PH_3 - Oxidation state -3
- AsH_3 - Oxidation state -3
- SbH_3 - Oxidation state +3
- BiH_3 - Oxidation state +3

Anomalous Nitrogen

- Nitrogen gas makes up 70% of the atmosphere and is inert
- Nitrogen forms many compounds with oxygen with a wide array of oxidation states that play important roles in nature.

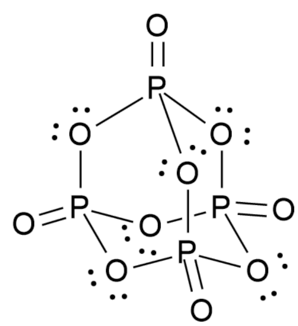
Empirical Formula	Lewis Structures (No F.C.)	Nitrogen Oxidation State
N_2O		+1 (0, +2)
NO		+2
N_2O_3		+3 (+2, +4)
NO_2		+4
N_2O_4		+4
N_2O_5		+5

Oxides of Phosphorous and Oxoacids of Group 15



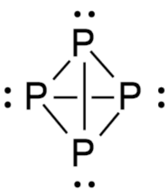
$$P_{4(s)} + O_{2(g)} \longrightarrow P_4O_{6(s)} \longrightarrow H_3PO_3$$

$$P_{4(s)} + O_{2(g)} \longrightarrow P_4O_{10(s)} \longrightarrow H_3PO_4$$

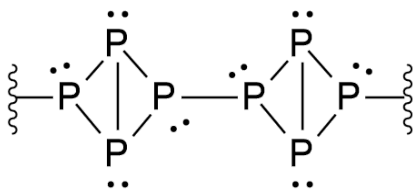


Allotropes of Elemental Phosphorous

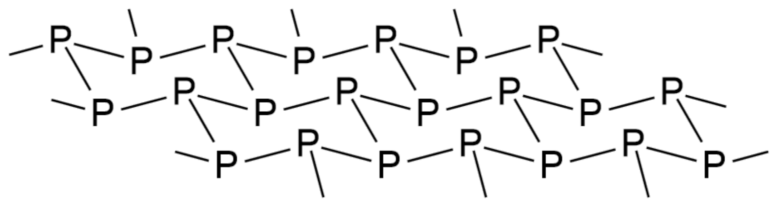
- An allotrope is different form of the same element in the solid phase.
- Phosphorous has three allotropes, red, white, and black
- The more strain the more reactive!



White



Red



Black

- White phosphorous is a chemical weapon. White phosphorous munitions result in horrific burns.
- Red phosphorous is used on strike anywhere matches
- Black phosphorous is inert

Importance of Phosphates

- Phosphate, PO_4^{3-} is found throughout our body, and can form polyphosphates like diphosphate $(O_3POPO_3)^{4-}$
- Most importantly they form the backbone of DNA and a triphosphate $(-OPO_3PO_3PO_3O-)^{5-}$ is attached to the end of ATP, a molecule used to move cellular energy around the body

Group 16

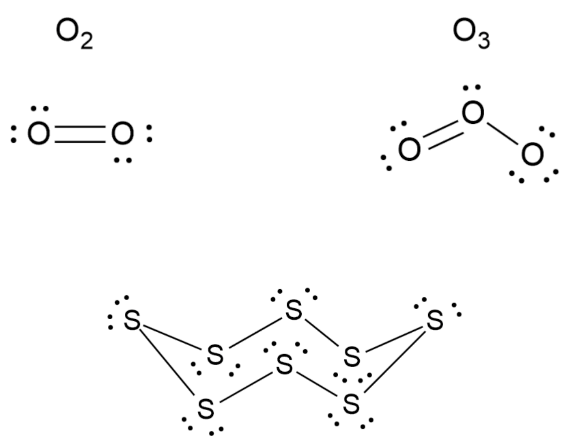
Elements of Group 16: O, S, Se, Te, Po, Uuh

Physical Properties and Trends

- mp increases up to Te then decreases down the group
- bp increases down the group
- density increases down the group
- Atomic size increases down the group
- IE decrease down the group
- Common oxidation states are -2, +4, +6

Allotropes of Oxygen and Sulfur

- Oxygen has two allotropes O_2 and O_3
- Sulfur has one naturally occuring allotrope S_8



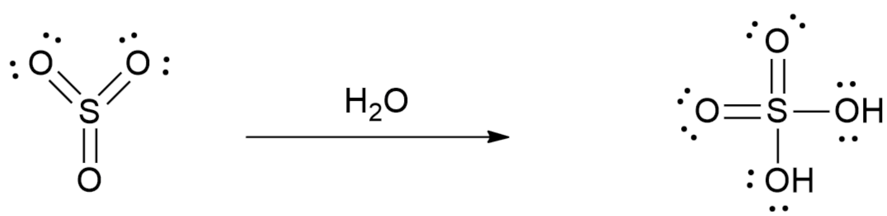
Hydrides of Group 16

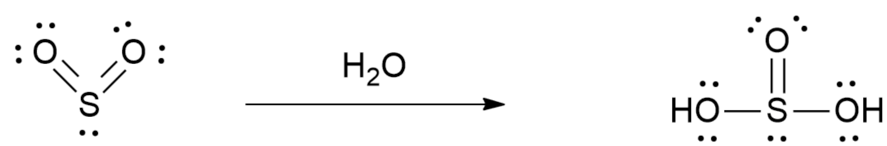
- Oxygen forms two hydrides H_2O and H_2O_2 with oxidation states of -2 and -1 respectively.
- The heavier elements of group 16 form simple hydrides H_2E which are foul smelling gasses

Anamolous oxygen

- non-metal
- strongly oxidizing (hence the name....)
- forms strong bonds to all elements except He, Ne and Ar

Sulfur Oxides and Oxoacids





3.10.6

Group 17

Elements of Group 17: F, Cl, Br, I, At, Uus

Physical Properties and Trends

- mp increases down the group
- bp increases down the group
- density increases down the group
- Atomic size increases down the group
- IE decrease down the group, very high in general
- Common oxidation state is -1 but oxidation states between -1 and +7 are possible

Reactivity trends

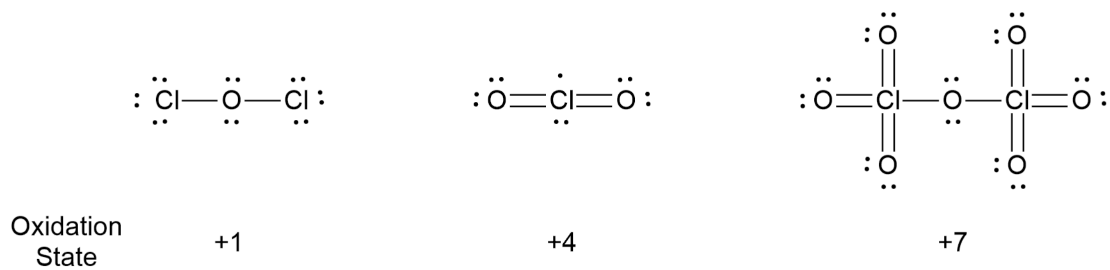
- From Cl downwards the bond strength decreases as the bond length increases (F is weird)
- Oxidizing ability of the X_2 molecule increases moving up the group
- Reducing ability of X^- increases down a a group.

Hydrides of Group 17

- All the elements of group 17 form simple HX hydrides and are acidic
- The X-H bond length and therefore acidity increases down the group.
- HF is the only weak acid (F is weird)

Oxides of Chlorine

- Chlorine can form oxides with an oxidation state ranging from +1 to +7
- Bromine and Iodine can form similar oxides



Halogen Oxoacids

	Hypo__acid	___ous acid	___ic acid	Per___ic acid
F	HOF	-	-	-
Cl	HOCl	HOClO	$HOClO_2$	$HOClO_3$
Br	HOBr	HOBrO	$HOBrO_2$	$HOBrO_3$
I	HOI	-	$HOIO_2$	$HOIO_3$

3.10.7

Group 18

Elements of Group 18: Ne, He, Ar, Kr, Xe, Rn

Physical Properties and Trends

- mp increases down the group
- bp increases down the group
- density increases down the group
- Atomic size increases down the group, smallest elements on the periodic table
- IE decrease down the group, highest on the periodic table
- Common oxidation state is 0 because they are mostly inert but oxidation states between +2 and +8 are possible for Kr - Rn

Reactivity trends

- Noble Gasses are inert and do not commonly participate in reactions or form compounds.
- The heavier noble gasses (Kr and lower) can form compounds under extreme conditions and the oxidation states range from +2 to +8

3.10.8

Which of the following elements is most likely to form an element hydride EH_3, when reacting with hydrogen?

- C
- P
- Si
- Na
- S

3.10.9

Chlorine oxides form a variety of inorganic acids. What is the oxidation state of chlorine in Chloric Acid?

- +1
- +3
- +5
- +7
- 0