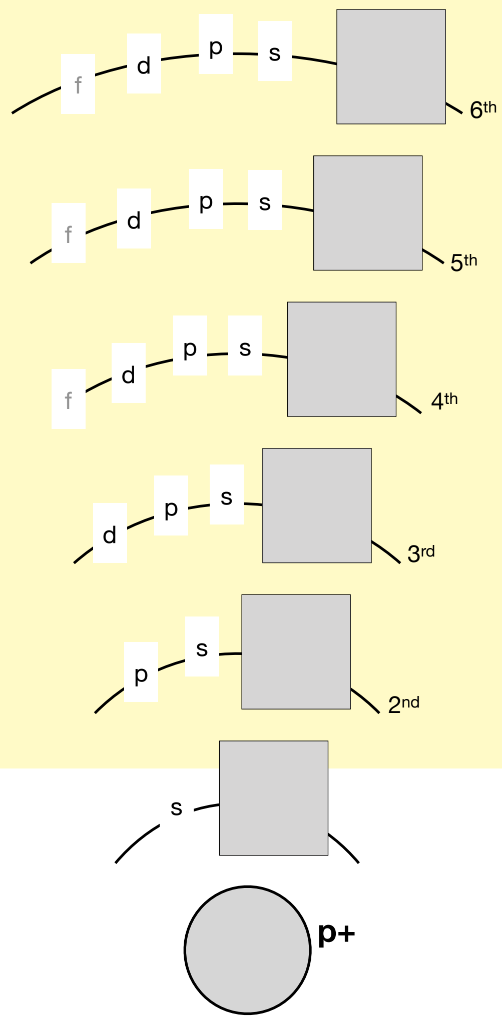
Electrons

Unit H

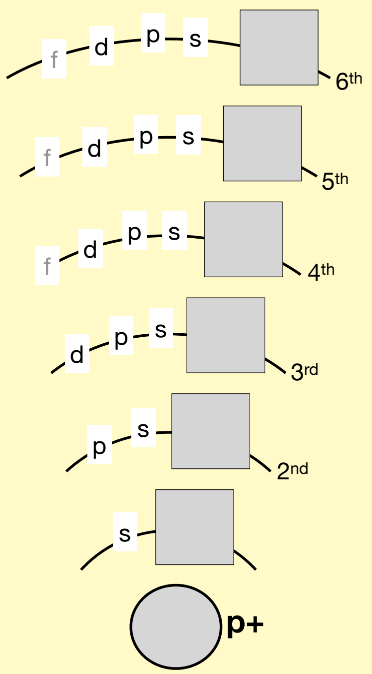
Break out the Whiteboards
and your Periodic Table
& Clickers



1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 **H** 1.008																	2 **He** 4.00
3 **Li** 6.94	4 **Be** 9.01											5 **B** 10.81	6 **C** 12.01	7 **N** 14.01	8 **O** 16.00	9 **F** 19.00	10 **Ne** 20.18
11 **Na** 22.99	12 **Mg** 24.30											13 **Al** 26.98	14 **Si** 28.09	15 **P** 30.97	16 **S** 32.06	17 **Cl** 35.45	18 **Ar** 39.95
19 **K** 39.10	20 **Ca** 40.08	21 **Sc** 44.96	22 **Ti** 47.87	23 **V** 50.94	24 **Cr** 52.00	25 **Mn** 54.94	26 **Fe** 55.85	27 **Co** 58.93	28 **Ni** 58.69	29 **Cu** 63.55	30 **Zn** 65.38	31 **Ga** 69.72	32 **Ge** 72.63	33 **As** 74.92	34 **Se** 78.97	35 **Br** 79.90	36 **Kr** 83.80
37 **Rb** 85.47	38 **Sr** 87.62	39 **Y** 88.91	40 **Zr** 91.22	41 **Nb** 92.91	42 **Mo** 95.95	43 **Tc**	44 **Ru** 101.07	45 **Rh** 102.91	46 **Pd** 106.42	47 **Ag** 107.87	48 **Cd** 112.41	49 **In** 114.82	50 **Sn** 118.71	51 **Sb** 121.76	52 **Te** 127.60	53 **I** 126.90	54 **Xe** 131.29

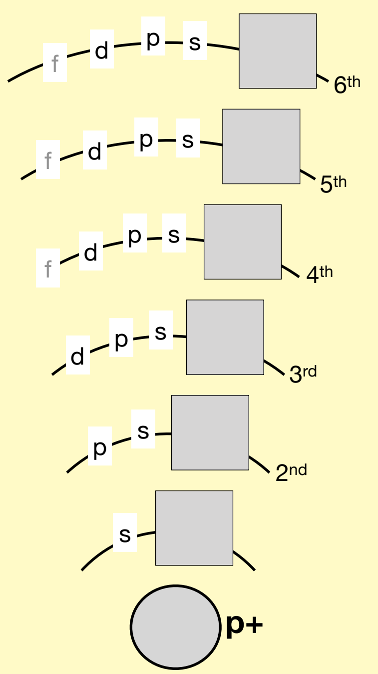
Write the *entire* electron configuration for $_{35}Br$

1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe



Write the entire electron configuration for ${}_{35}Br$

Turn this into a *condensed* electron configuration.



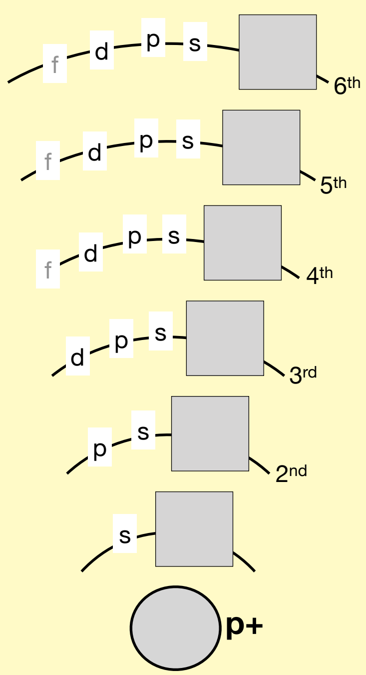
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe

- $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^5$

Write the entire electron configuration for $_{35}Br$

Turn this into a *condensed* electron configuration.

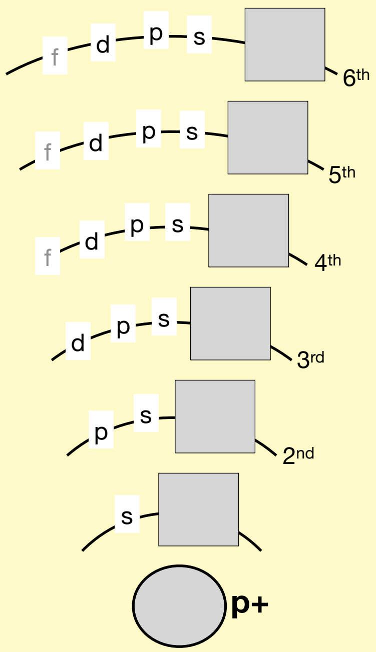
- $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^5$
- $[Ar]\ 4s^2\ 3d^{10}\ 4p^5$



1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe

Write the *condensed* electron configuration for $_{33}\text{As}$

1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe



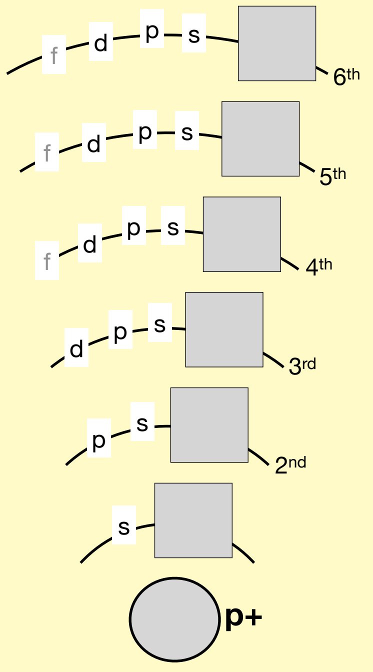
Here is the the condensed electron configuration for $_{33}As$

Sketch an *orbital notation* for this.

How many unpaired electrons in As?

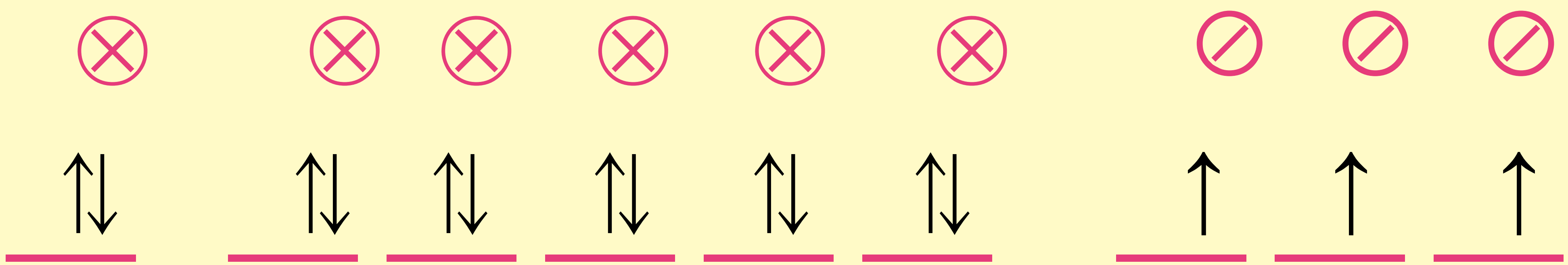
- [Ar] $4s^2\ 3d^{10}\ 4p^3$

1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe



Write the condensed electron configuration for $_{33}As$
Sketch an *orbital notation* for this.

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

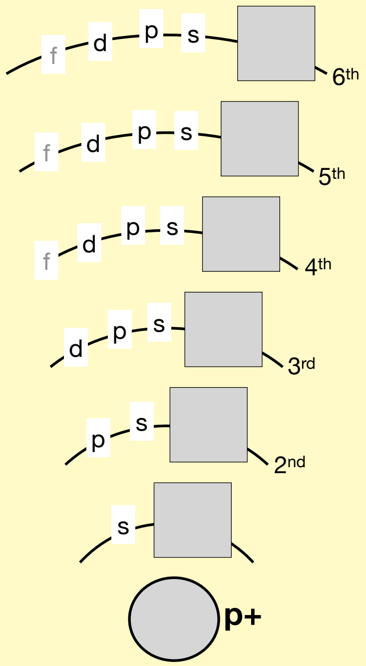


- 3 unpaired electrons

Writing orbital notation reminds us that electrons will spread out within a sublevel. Electrons won't pair up until each orbital has at least one electron.

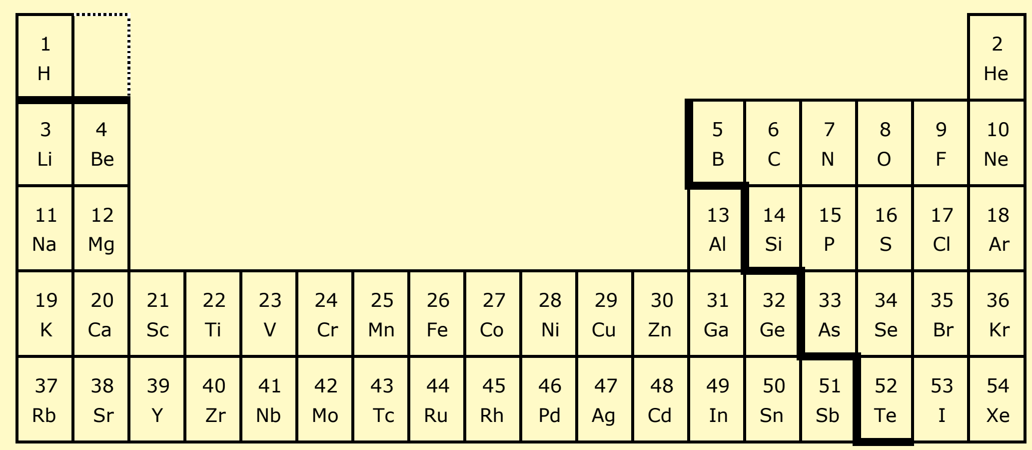
Write condensed electron configuration for $_{50}$Sn in its *ground state*.

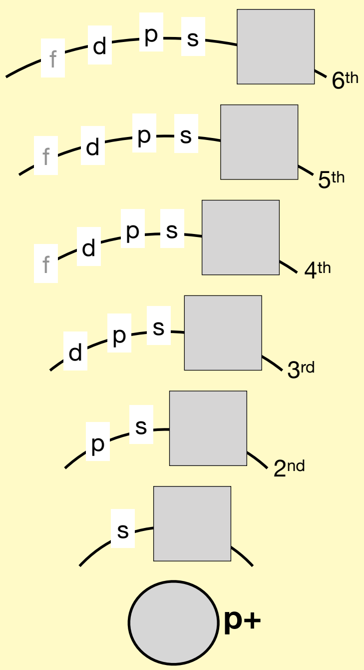
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe



Write condensed electron configuration for $_{50}$Sn in its *ground state*.

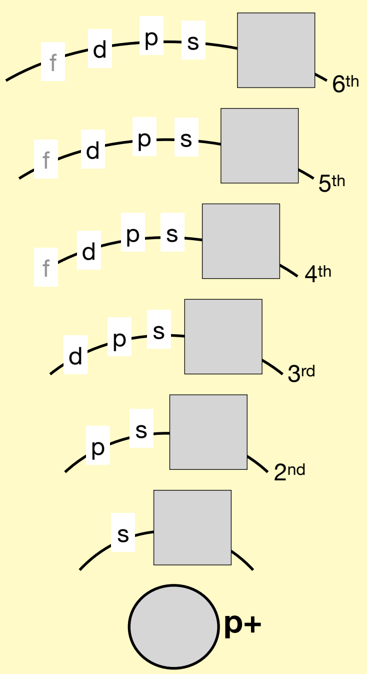
- $[Kr]5s^2\, 4d^{10}\, 5p^2$





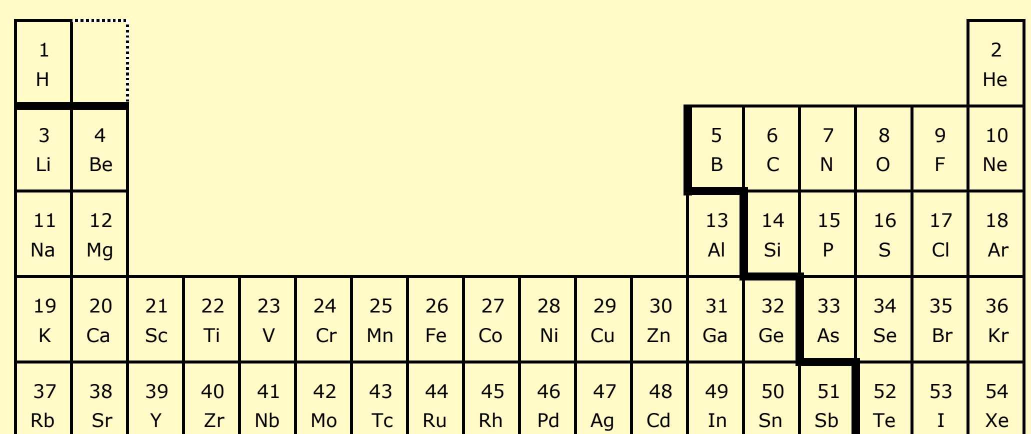
Write condensed electron configuration that describes the $_{35}Br$ atom and then the $_{35}Br^{-}$ ion both in their *ground state*.

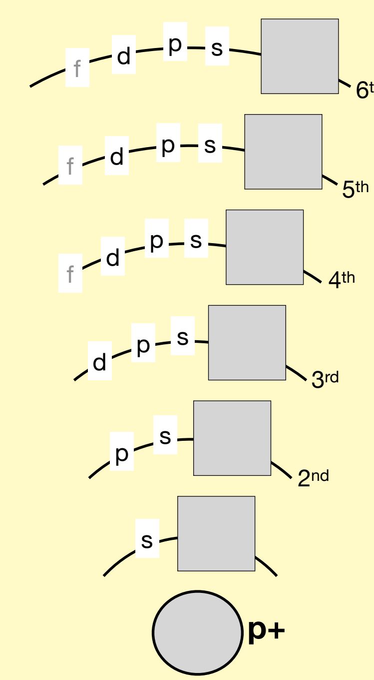
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe



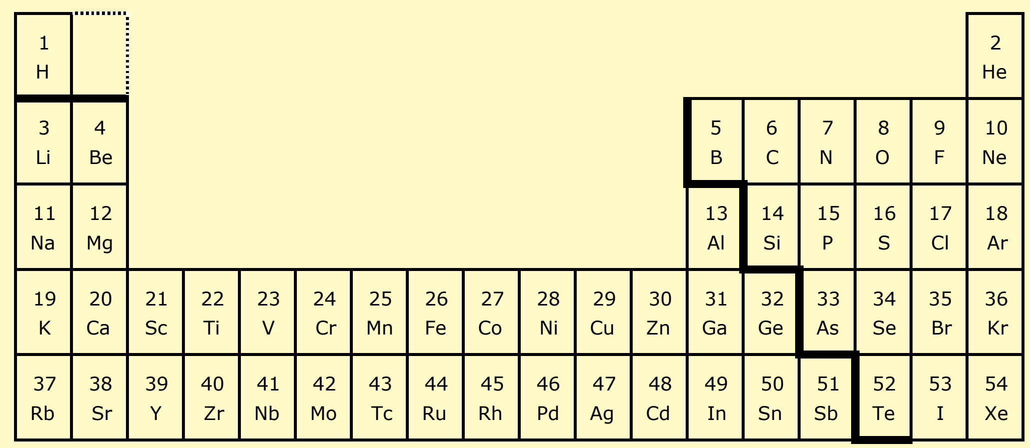
Write condensed electron configuration that describes the $_{35}Br$ atom and then the $_{35}Br^{-}$ ion both in their *ground state*.

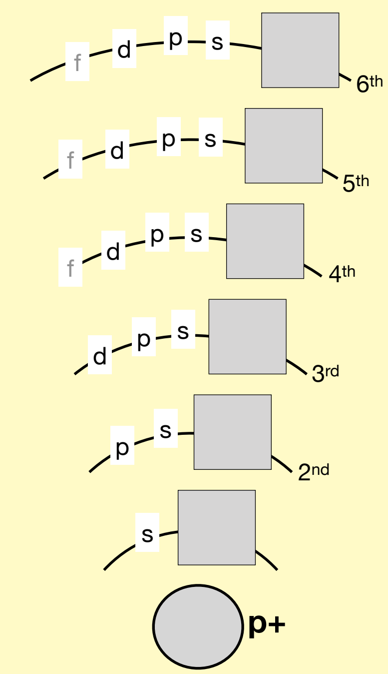
- Br atom: [Ar] $4s^2\ 3d^{10}\ 4p^5$
- Br^{-} ion: [Ar] $4s^2\ 3d^{10}\ 4p^6$





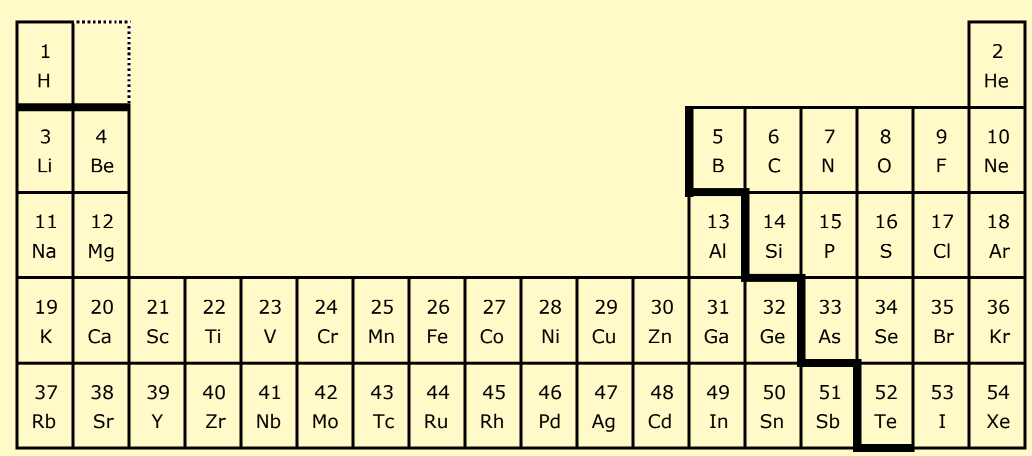
Write entire electron configuration for $_{12}Mg$ atom and then the $_{12}Mg^{2+}$ ion in their *ground state*.

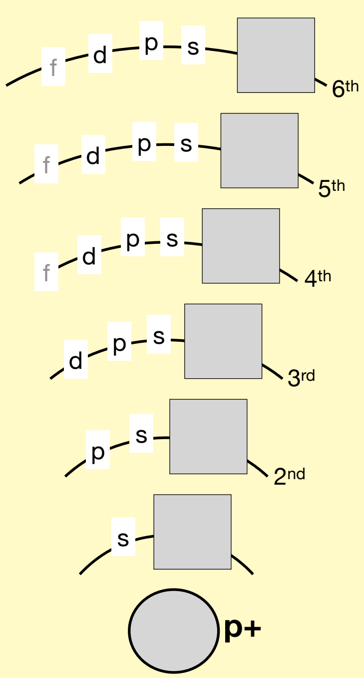




Write entire electron configuration for $_{12}Mg$ atom
and then the $_{12}Mg^{2+}$ ion
in their *ground state*.

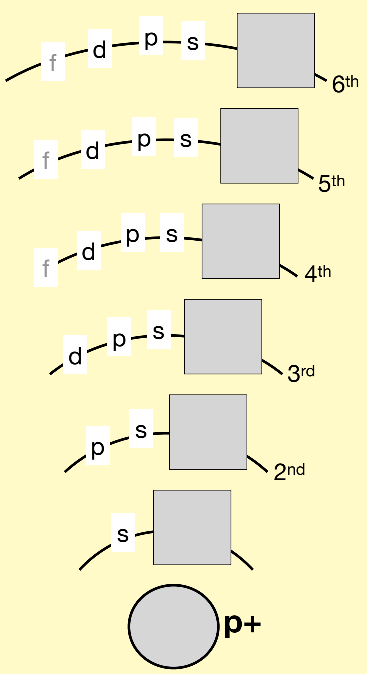
- Mg atom: $1s^2\ 2s^2\ 2p^6\ 3s^2$
- Mg^{2+} ion: $1s^2\ 2s^2\ 2p^6$



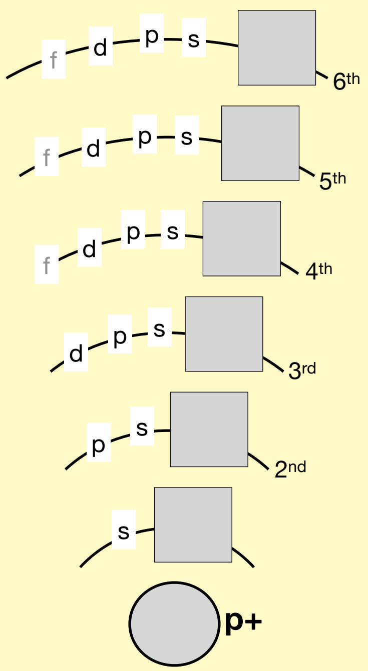


Write *condensed* electron configuration that describes $_{28}$Ni atom, and then for the $_{28}Ni^{2+}$ ion both in their *ground state*.

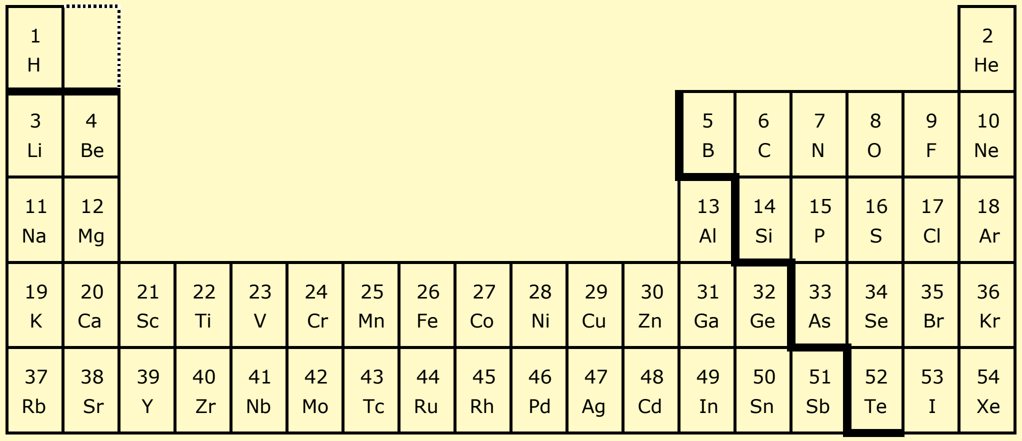
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe



Write *condensed* electron configuration that describes $_{28}Ni$ atom, and then for the $_{28}Ni^{2+}$ ion both in their *ground state*.



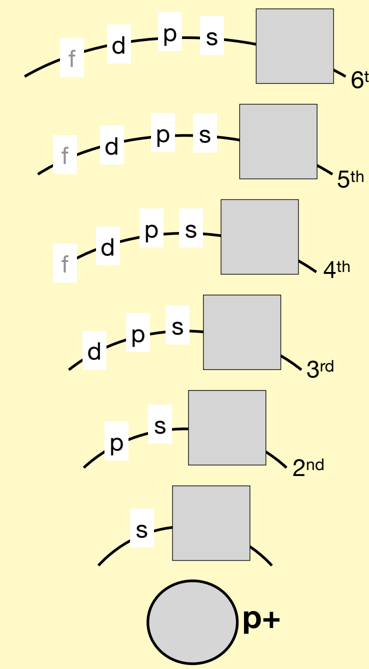
- Ni: [Ar] $4s^2$ $3d^8$
- Ni^{2+}: [Ar] $3d^8$
- ~~NOT [Ar] $4s^2$ $3d^6$~~



- While the d electrons are higher energy, and in theory, easier to remove, the s electrons are on the outside and get knocked off first.

Name the element that is described by the condensed electron configuration of for an atom.

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

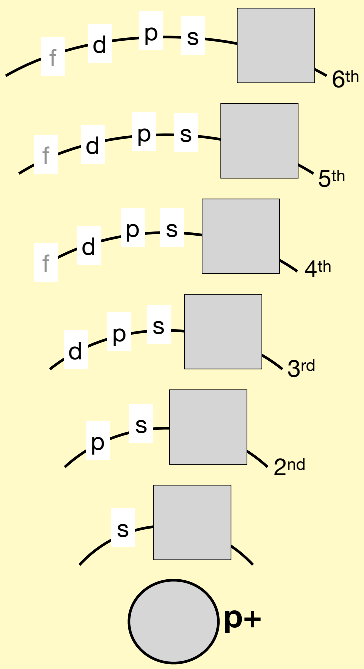


1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe

Name the element that is described by the condensed electron configuration of for an atom.

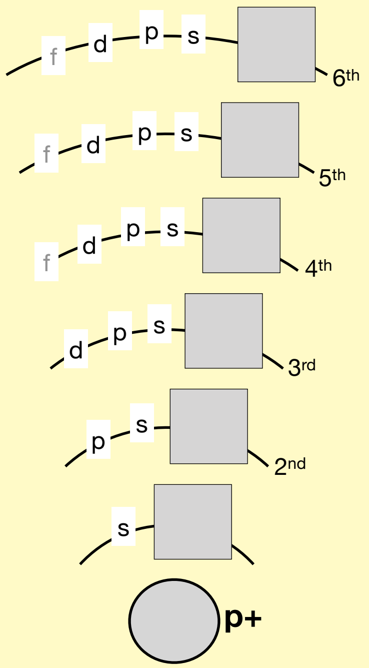
[Kr] $5s^2\ 4d^4$

- $_{42}$Mo



1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe

Write a generic *valence* electron configuration that could describe any halogen.



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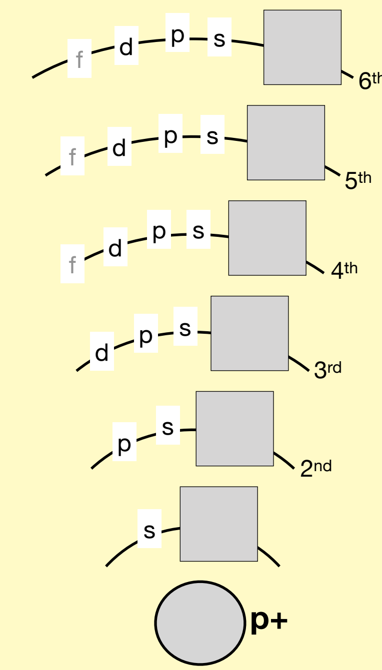
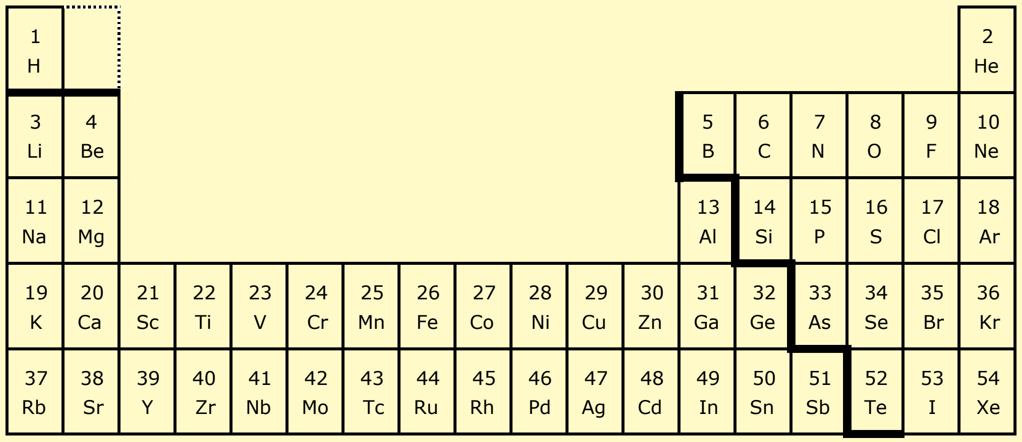
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe

Write a generic *valence* electron configuration that could describe any halogen.

- $ns^2\,np^5$ $\#s^2\,\#p^5$ $xs^2\,xp^5$ $s^2\,p^5$
- in most other contexts, n means mole, but in reference to electrons, n means “principal quantum number” = energy level

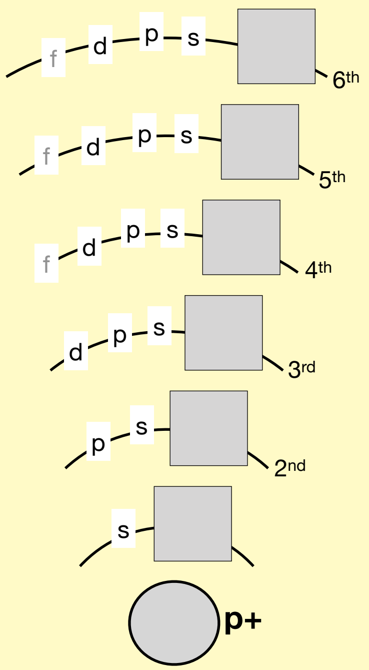
Name three (common) naturally occurring particles (atoms or ions) that are *isoelectronic* with Cl^{-}





Name three (common) naturally occurring particles (atoms or ions) that are *isoelectronic* with Cl^-

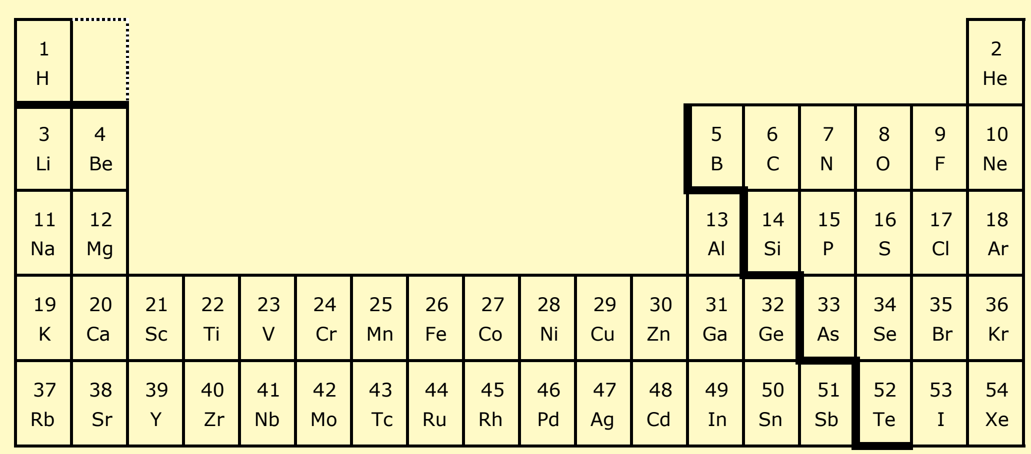
...means, same number of electrons

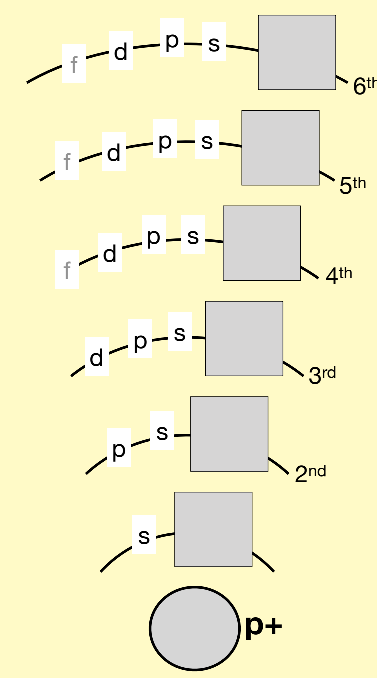


1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe

Name three (common) naturally occurring particles (atoms or ions) that are *isoelectronic* with Cl^{-}

- same number of electrons, 18
- Ar, S^{2-}, K^{+}, Ca^{2+}, V^{5+}, and others





What is the electron configuration of the *valence* electrons of selenium?

1. $4s^2$
2. $4p^4$
3. $4d^4$
4. $4s^2\ 4p^4$
5. $4s^2\ 3d^{10}\ 4p^4$
6. $4s^2\ 4d^{10}\ 4p^4$

1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	71 Lu	72 Hf	73 Ta	73 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

What is the electron configuration of the *valence* electrons of selenium?

1. $4s^2$
2. $4p^4$
3. $4d^4$
4. $4s^2\ 4p^4$
5. $4s^2\ 3d^{10}\ 4p^4$
6. $4s^2\ 4d^{10}\ 4p^4$

Some textbooks consider the d electrons to be valence electrons, but most do not. AP will refer only to s and p as valence electrons.

1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	71 Lu	72 Hf	73 Ta	73 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

How many unpaired electrons does selenium have?

1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	71 Lu	72 Hf	73 Ta	73 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

How many unpaired electrons does selenium have?

- 2 unpaired electrons
- $4s^2\ 4p^4$
- orbital notation: ⊗ ⊗ ⊘ ⊘
- orbital notation: $\underline{\uparrow\downarrow}$ $\underline{\uparrow\downarrow}$ $\underline{\uparrow\ }$ $\underline{\uparrow\ }$

1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	71 Lu	72 Hf	73 Ta	73 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

Magnetism

Ferromagnetic, Paramagnetic Diamagnetic

What information does magnetism give us about the arrangement of electrons?

Ferromagnetic

- This is the noticible force that you know and ove from magnets on your refrigerator.
- These materials that can maintain a magnetic field in the absence of an external magnetic source
- They are made from a variety of elements.
  - ✓ iron
  - ✓ aluminum, nickel, and cobalt
  - ✓ samarium
  - ✓ neodymium
- We will not discuss these any further

paramagnetic and diamagnetic

- These two types of magnesim are very weak and are more difficult to detect, but give us great insight to the arrangement of electrons within atoms
- paramagnetism
  - ✓ The tendency of a species align itself in a magnetic field (attracted) as a result of having unpaired electrons.
- diamagnetism
  - ✓ The tendency of a species to turn at right angle to a magnetic field (repelled) as a result of all of its electrons being paired.

Evidence of electron configurations

- *Paramagnetic studies are used to provide additional evidence of proposed electron configurations*
  - ✓ Experimental evidence shows that both Ti and Ti^{2+} are both paramagnetic, indicating which electrons the atom has lost to become an ion.
    - Ti [Ar] $4s^2 3d^2$
    - ⊗ ⊘⊘○○○
    - Ti^{2+} [Ar] $3d^2$
    - ○ ⊘⊘○○○
  - ✓ If the d orbital electrons that were removed (they are not), from the Ti atom to produce the Ti^{2+} ion, then the ion would not be paramagnetic (yet the ion is paramagnetic).

Evidence of electron configurations

- *Paramagnetic studies are used to provide additional evidence of proposed electron configurations*
  - ✓ Iron atoms are paramagnetic
    - Fe [Ar] $4s^2 3d^6$
    - ⊗ ⊗⊘⊘⊘⊘
  - ✓ Fe^{2+} ions do *not* demonstrate an increase or decrease in paramagnetism
    - Fe^{2+} [Ar] $3d^6$
    - ○ ⊗⊘⊘⊘⊘
  - ✓ Fe^{3+} ions demonstrate increased paramagnetism
    - Fe^{3+} [Ar] $3d^5$
    - ○ ⊘⊘⊘⊘⊘

Shapes of

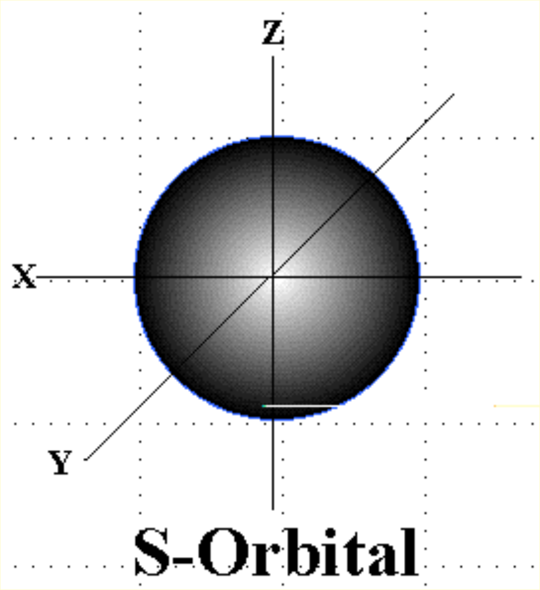
You do not need to know much about shapes for AP.

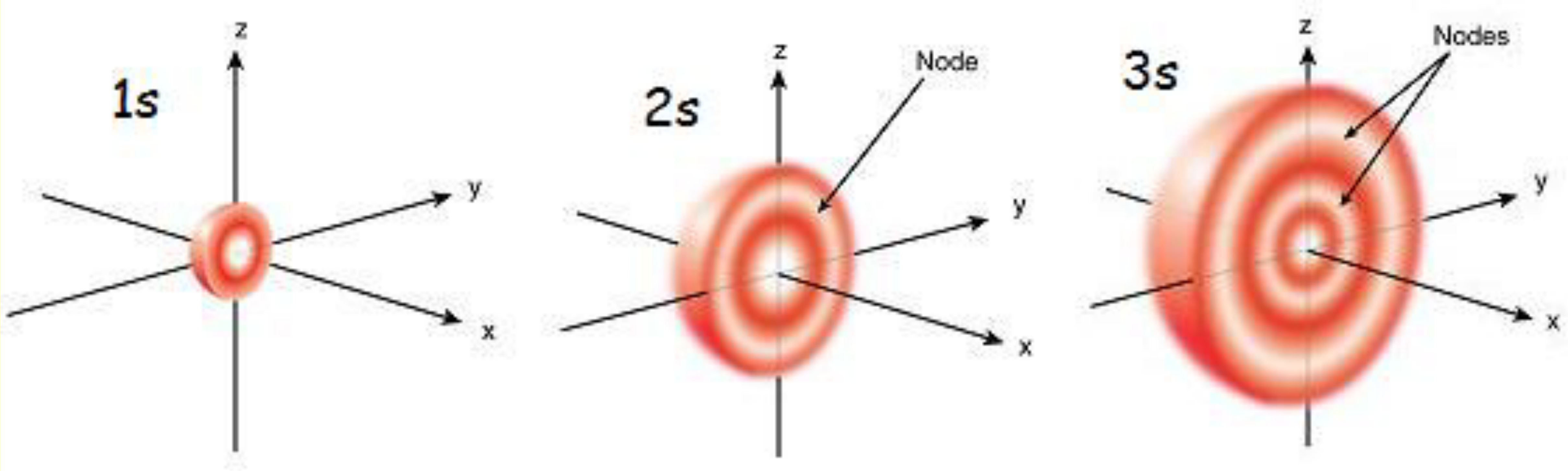
It will be more than enough to know about only s and p orbitals.

I'll show you the d's & f's only for comparison.

Shape of s Orbitals

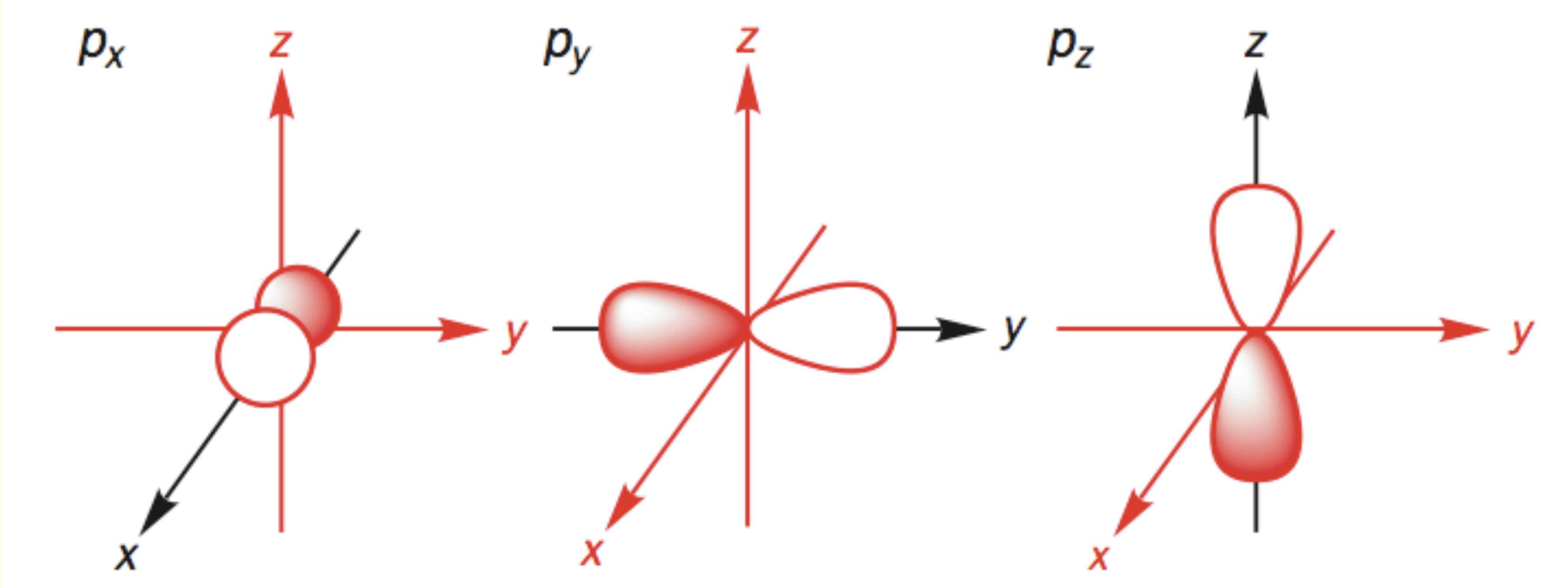
Not particularly helpful to answer any AP questions



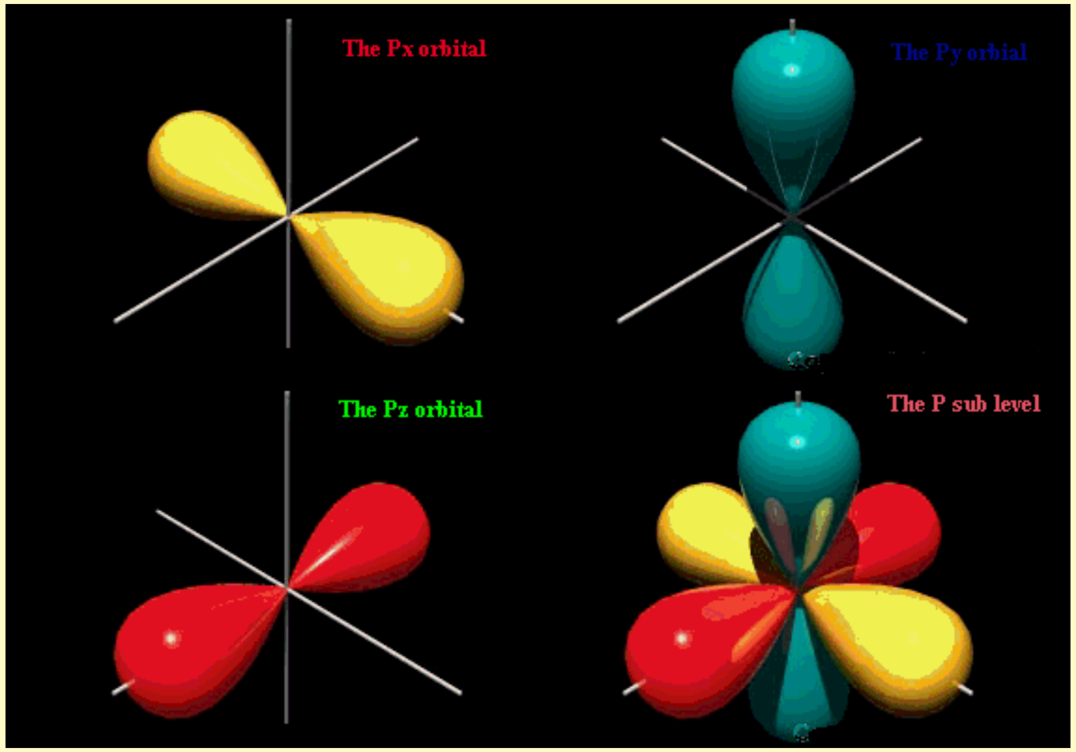


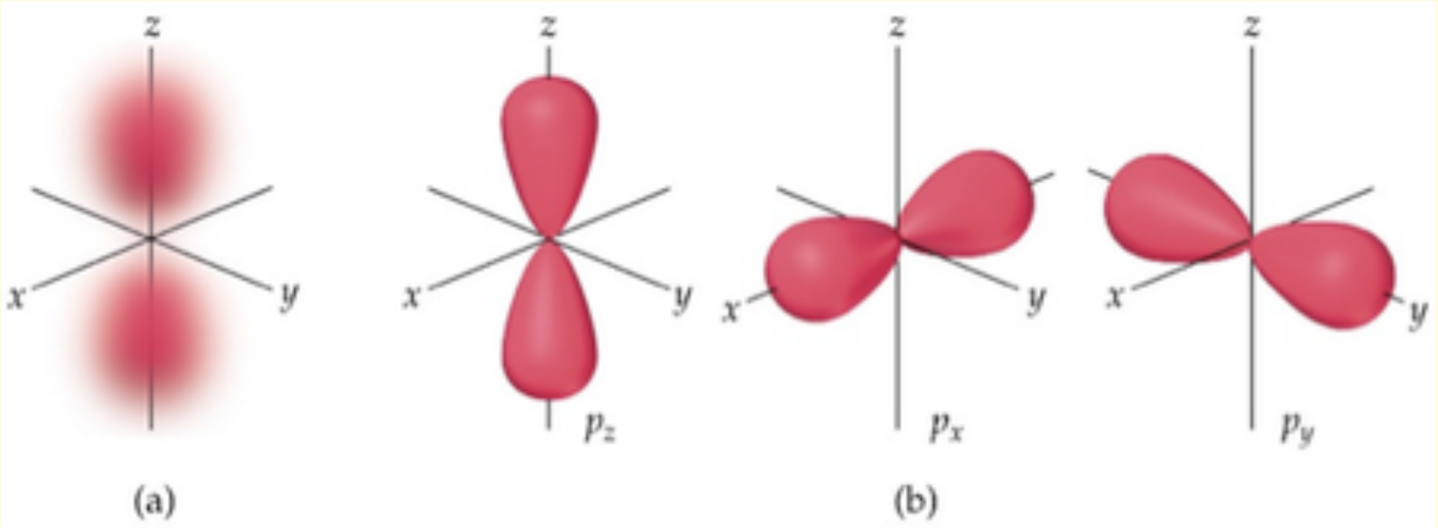
Shape of p Orbitals

Maybe worth knowing that the three p orbitals are oriented at right angles to each other. This might be helpful in a bonding situation. However this VERY rarely shows up on the AP exam.



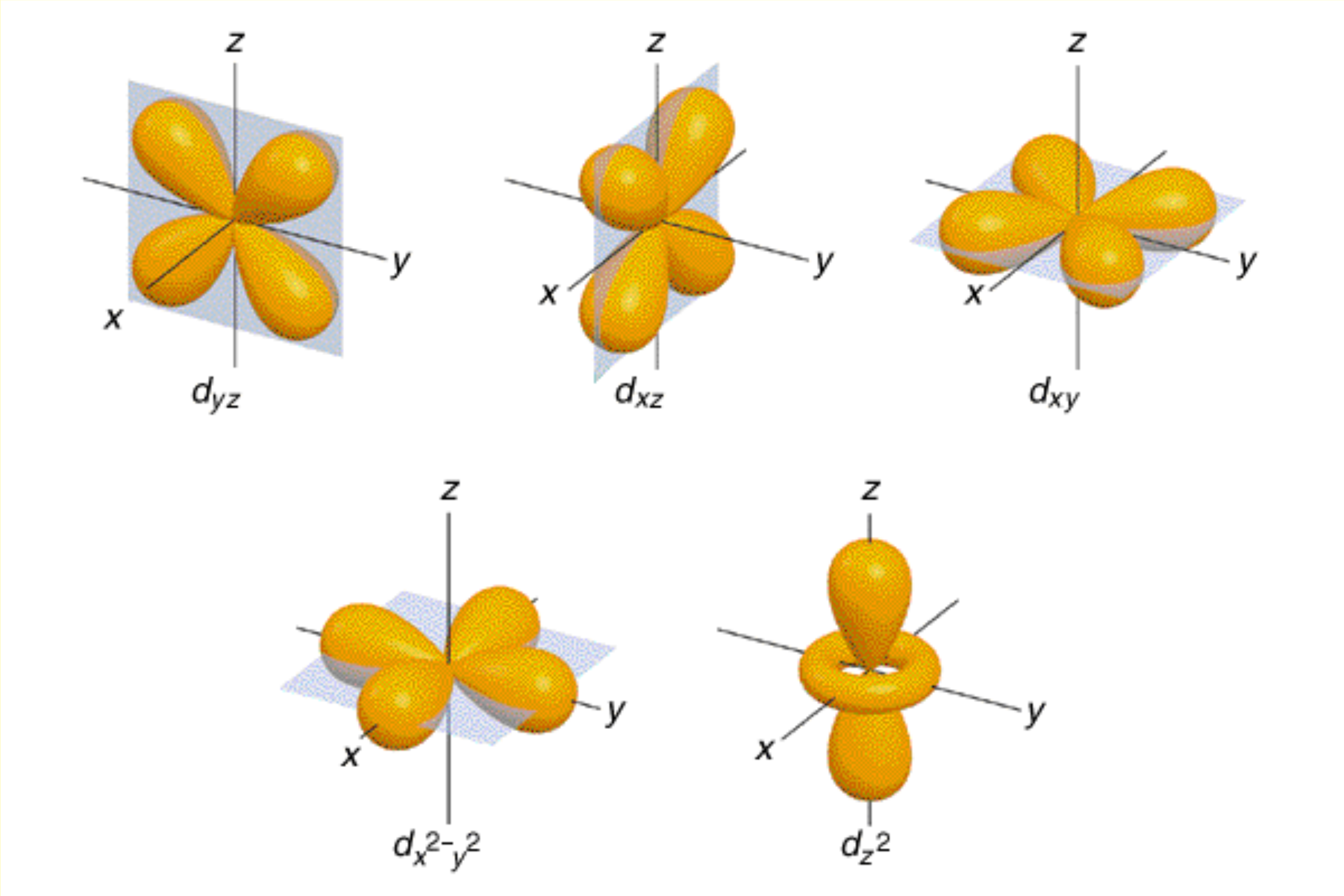
the three degenerate p orbitals are aligned along perpendicular axes





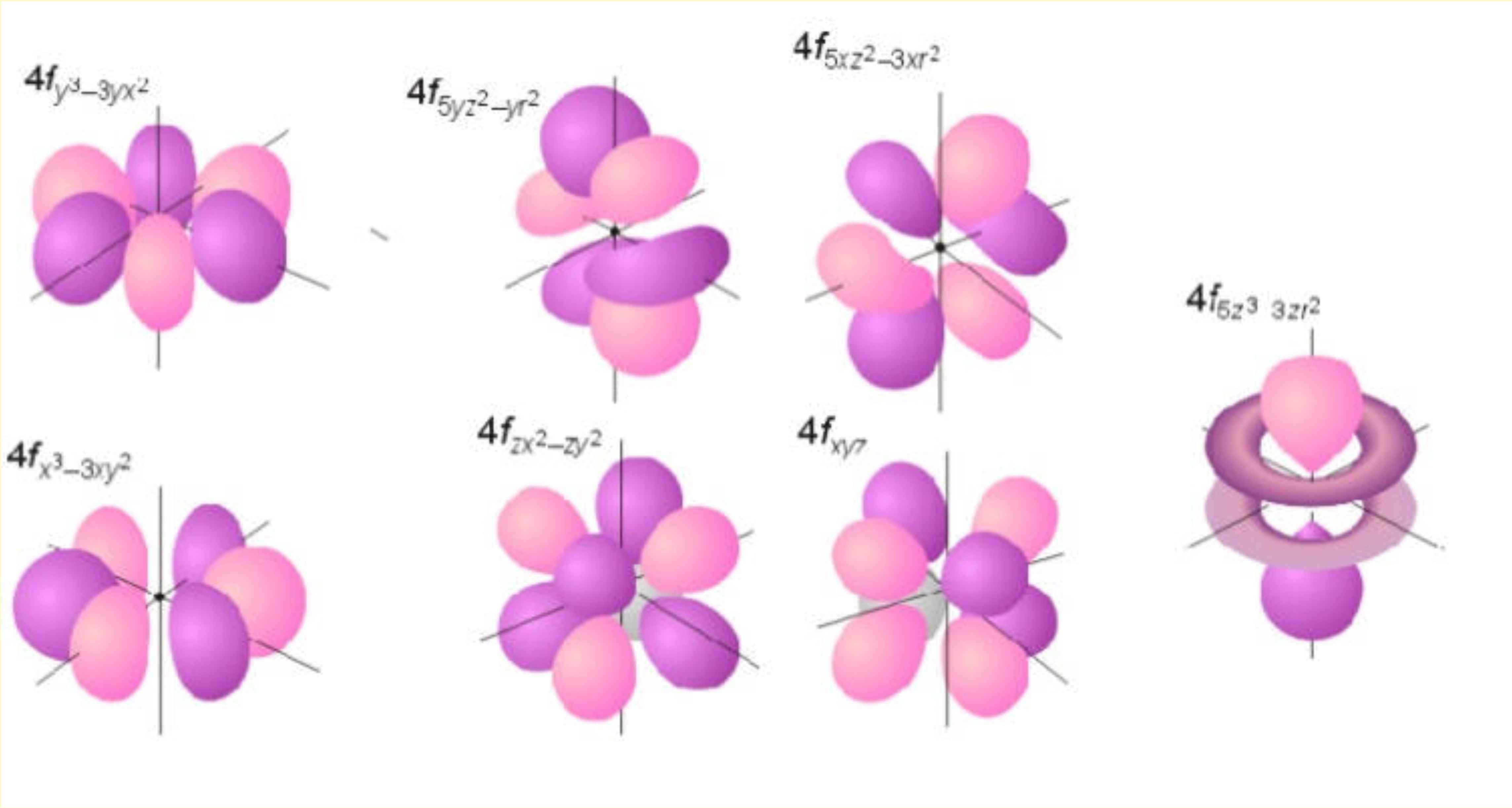
Shape of d Orbitals - whoa!

so not on the AP exam



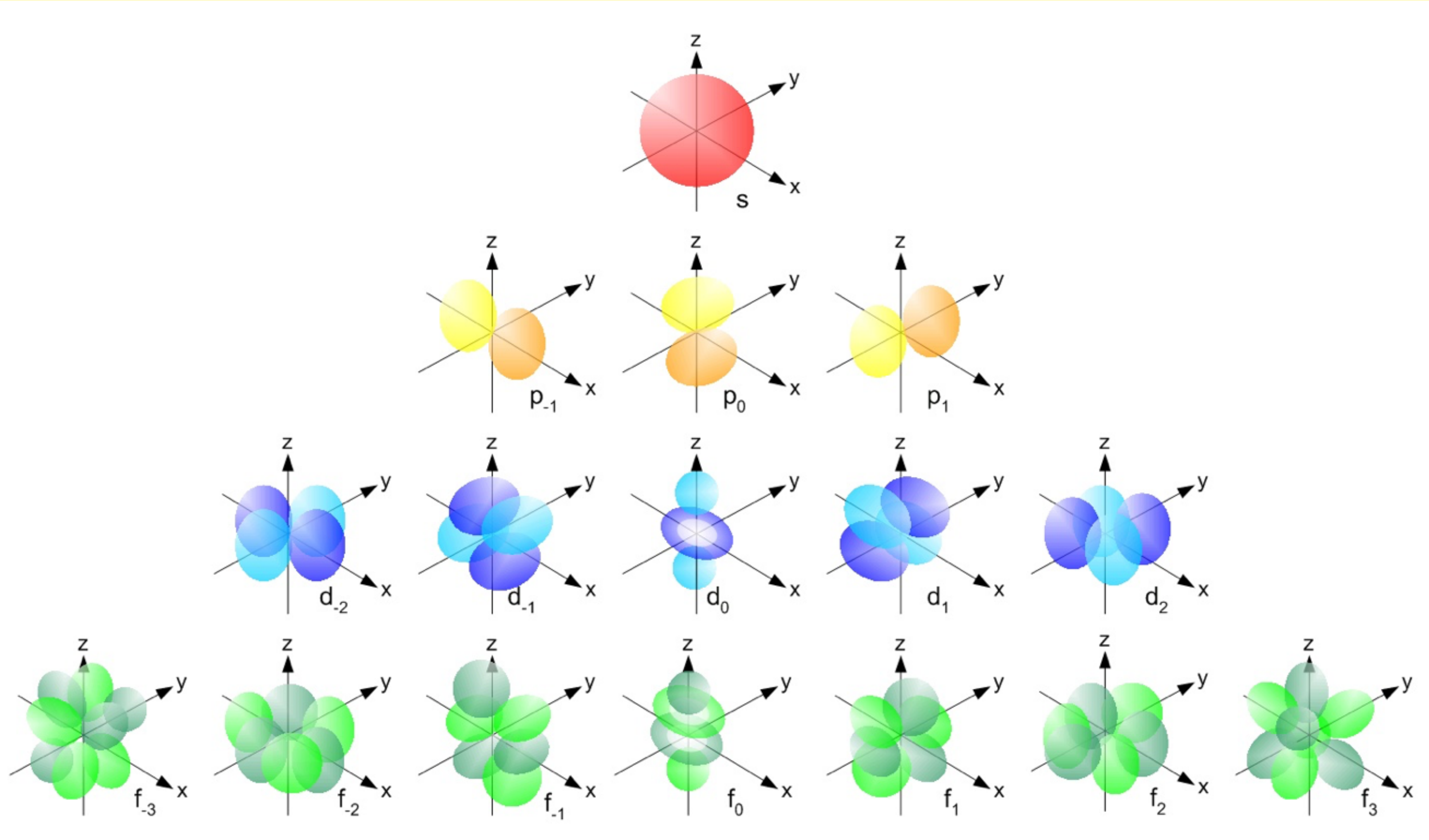
Shape of f Orbitals - Yikes !!

so not on the AP exam



Shape of Orbitals

so not on the AP exam



Periodicity

Reoccurring trends in the physical properties of the elements in the periodic table.

We must be able to identify the underlying cause(s) that explains these properties. Often we will refer to Coulombs Law in our justification.

Coulombs Law

As it pertains to periodicity

Proportionality constant that we do NOT need to worry about

proton charge

electron charge

$$F = k\frac{(Q^{+})(Q^{-})}{d^2}$$

distance between nucleus and electron

Let's be clear that we will never need to put numbers into this formula, but only use it in support of any statements we make.

- Coulomb's Law states that the force of attraction between two oppositely charged particles is
  - ✓ directly proportional to the magnitude of the charges (Q)
    (In particular the nuclear charge transmitted to any particular electron)
  - ✓ inversely proportional to the distance between those charges (d)
    (In particular, the distance between the nucleus and any particular electron of interest)

What is *Effective Nuclear Charge*? (Z_{eff} or *enc*)

- enc is an approximate determination of the nuclear charge that is transmitted to the valence electron(s)
- Some of the total nuclear charge is blocked or shielded by the inner core electrons

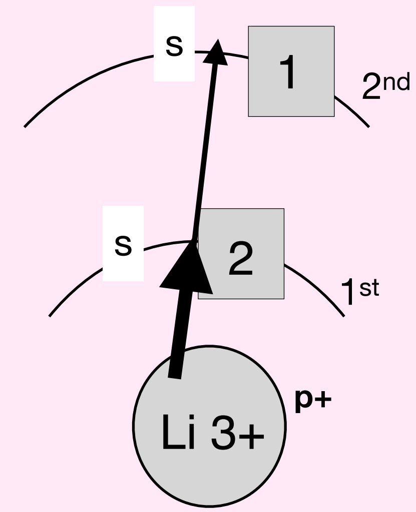
$$F = k\frac{(Q^{+})(Q^{-})}{d^{2}}$$

1	2											13	14	15	16	17	18
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe

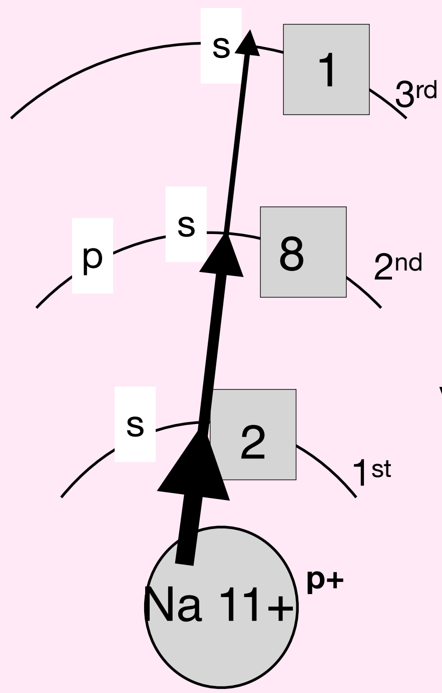
- \# protons − inner core electrons = enc
- In an atom, this value will be ~equal to the number of valence electrons of that atom

Group 1 *Effective Nuclear Charge* (Z_{eff} or *enc*)

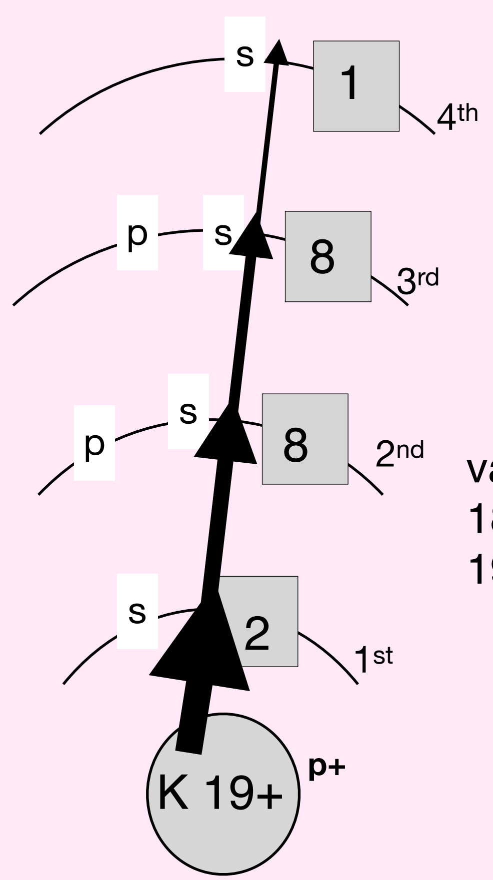
In effect, what is the nuclear charge experienced by the valence electrons?



valence e– experiences enc of 1+
2 inner core e– screening
3+ total



valence e– experiences enc of 1+
10 inner core e– screening
11+ total

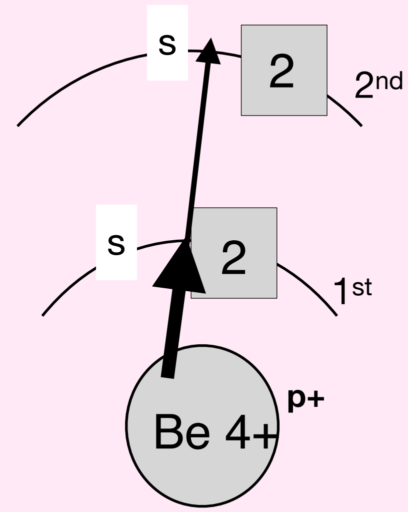


valence e– experiences enc of 1+
18 inner core e– screening
19+ total

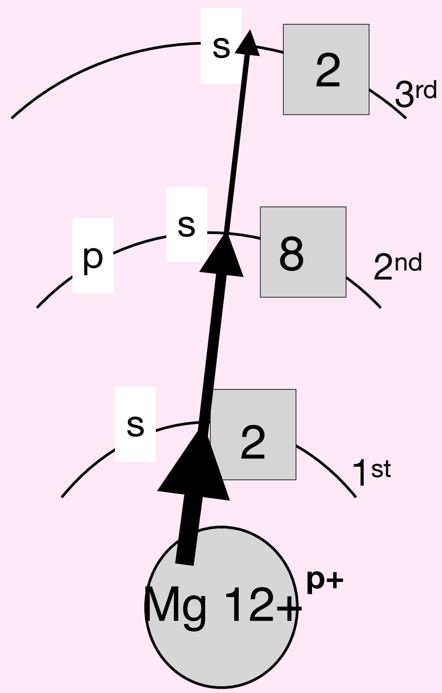
1	2											13	14	15	16	17	18
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe

Group 2 *Effective Nuclear Charge* (Z_{eff} or *enc*)

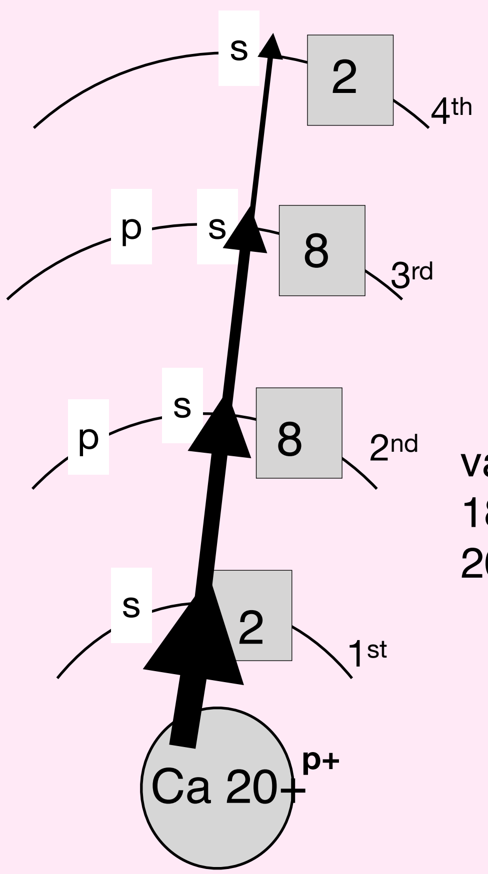
In effect, what is the nuclear charge experienced by the valence electrons?



valence e– experiences enc of 2+
2 inner core e– screening
4+ total



valence e– experiences enc of 2+
10 inner core e– screening
12+ total

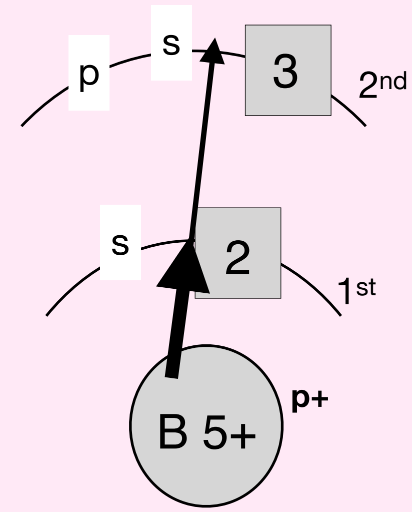


valence e– experiences enc of 2+
18 inner core e– screening
20+ total

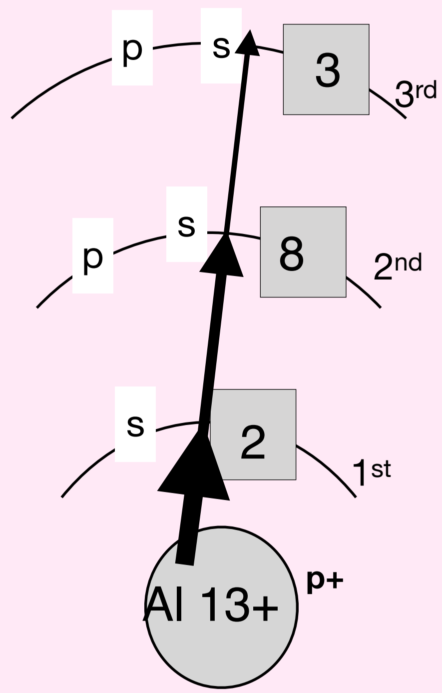
1	2											13	14	15	16	17	18
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe

Group 3 *Effective Nuclear Charge* (Z_{eff} or *enc*)

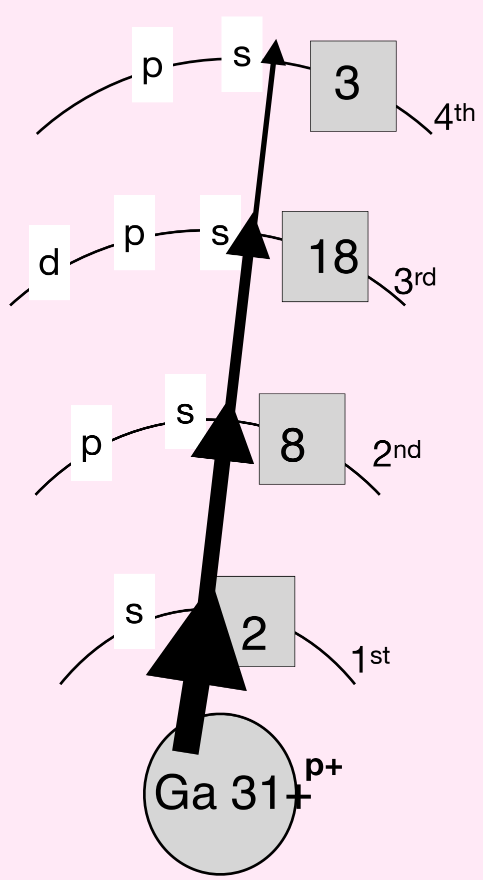
In effect, what is the nuclear charge experienced by the valence electrons?



valence e– experiences enc of 3+
2 inner core e– screening
5+ total



valence e– experiences enc of 3+
10 inner core e– screening
13+ total



valence e– experiences enc of 3+
28 inner core e– screening
31+ total

ENC increases across the periodic table, and is approximately equal to the number of valence electrons of the atom

~1	~2											~3	~4	~5	~6	~7	~8
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe

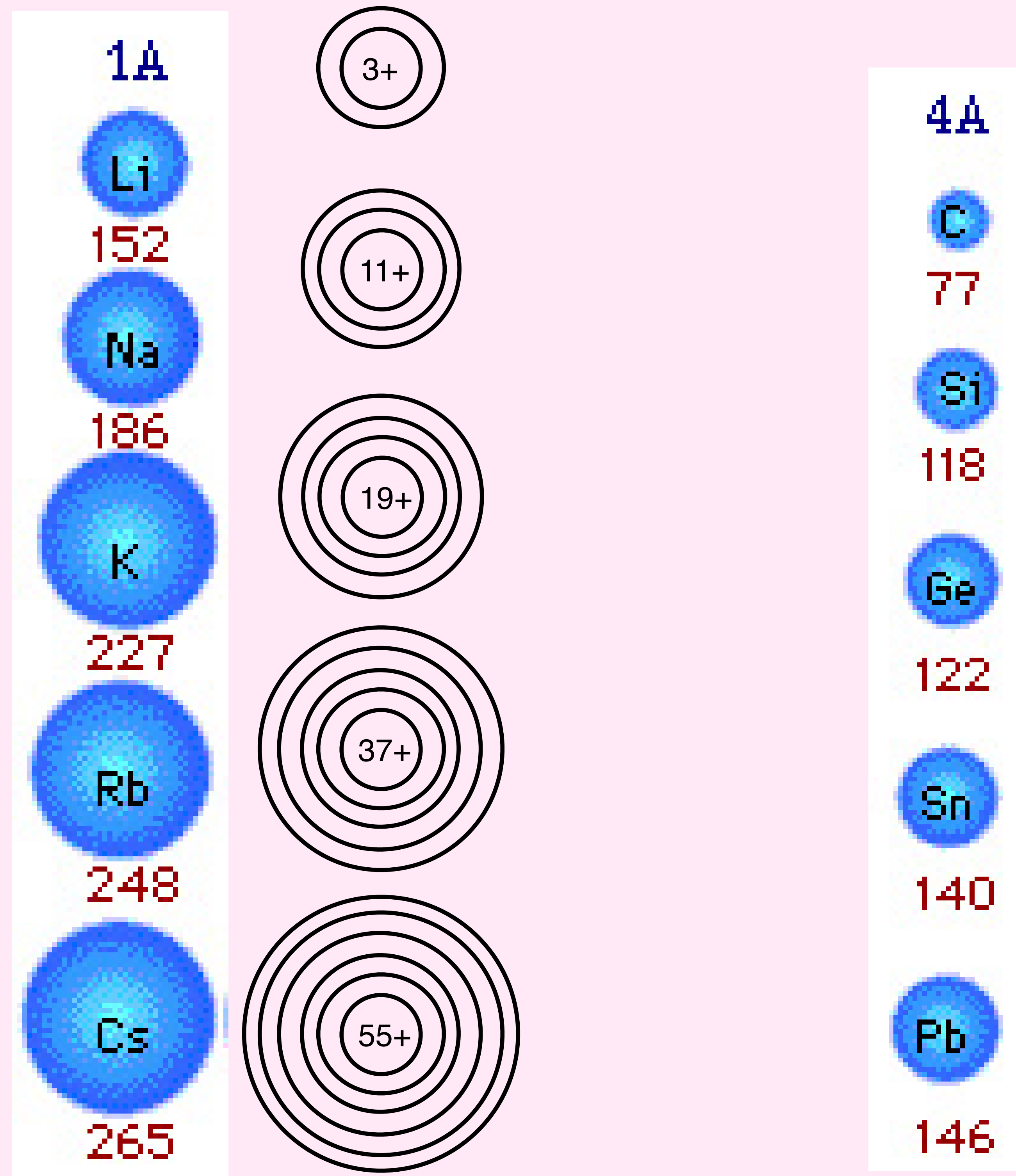
Atomic

and

Ionic Radii

Atomic Radii (pm)

- Increases down the periodic table
  - ✓ because of increased *occupied* energy levels, causing valence electrons to be further from the nucleus.

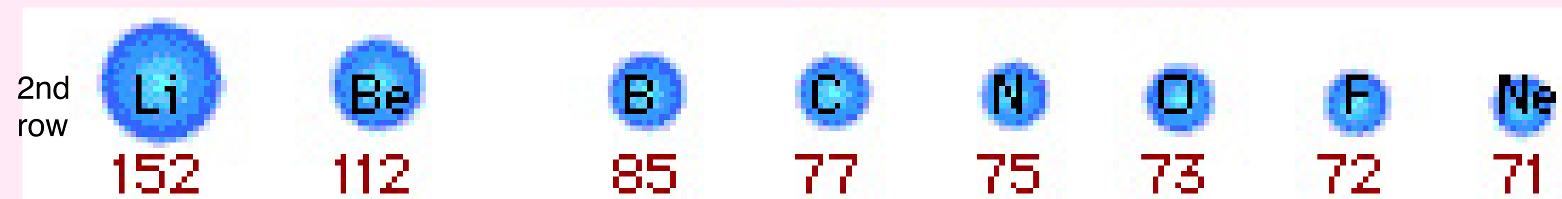


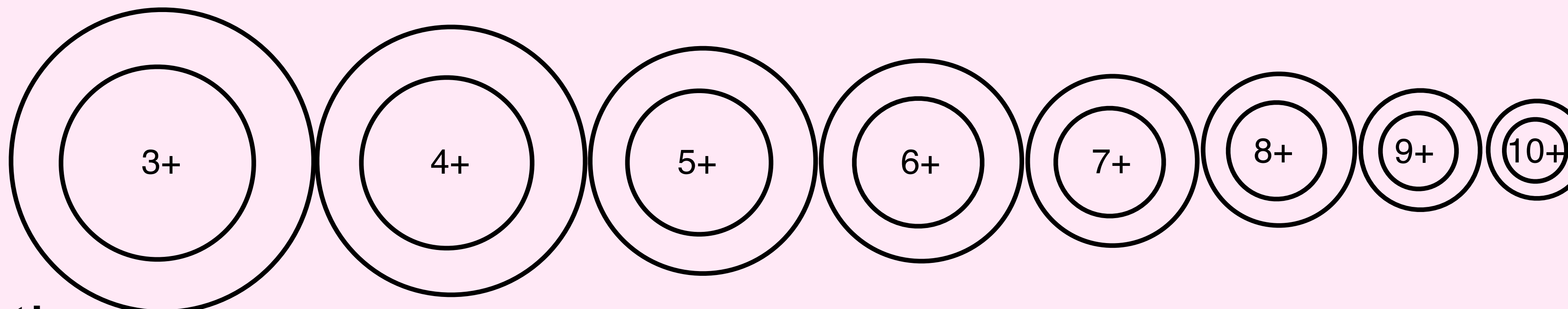
Atomic Radii (pm)

Decrease across the periodic table (L-R)

- ✓ more protons (+) attract valence electrons that are occupying the same energy level.
- ✓ increased effective nuclear charge (*enc*) acting on electrons in the same energy level drawing electrons in closer.
- ✓ Coulombs Law

$$F = k\frac{(Q^{+}\uparrow)(Q^{-})}{d^{2}}$$





Based on atomic radii, which of the following molecules would you *expect* to have the shortest bonds?

1. F_2
2. Cl_2
3. Br_2

This may initally appear to be a bonding question, but really this is a periodicity question. Rephrase the question into "which atom is smallest; F, Cl, Br? The answer to that will result in the diatomic molecule with the smallest bonds.

1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe

Based on atomic radii, which of the following molecules would you expect to have the shortest bonds?

1. F_2

- Fluorine is the smallest of these atoms (since F has the fewest number of occupied energy levels), the atoms can "skootch" closer and the fluorine molecule would be have shortest bonds.

2. Cl_2
3. Br_2

1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe

Which of the following is(are) expected to be a correct order of atomic radius from smallest to largest?

Select all that are correct

1. $_{12}Mg < _{13}Al < _{18}Ar$
2. $_{16}S < _{14}Si < _{11}Na$
3. $_{37}Rb < _{19}K < _{3}Li$
4. $_{15}P < _{33}As < _{51}Sb$

1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe

Which of the following is(are) expected to be a correct order of atomic radius from smallest to largest? *Select all that are correct.*

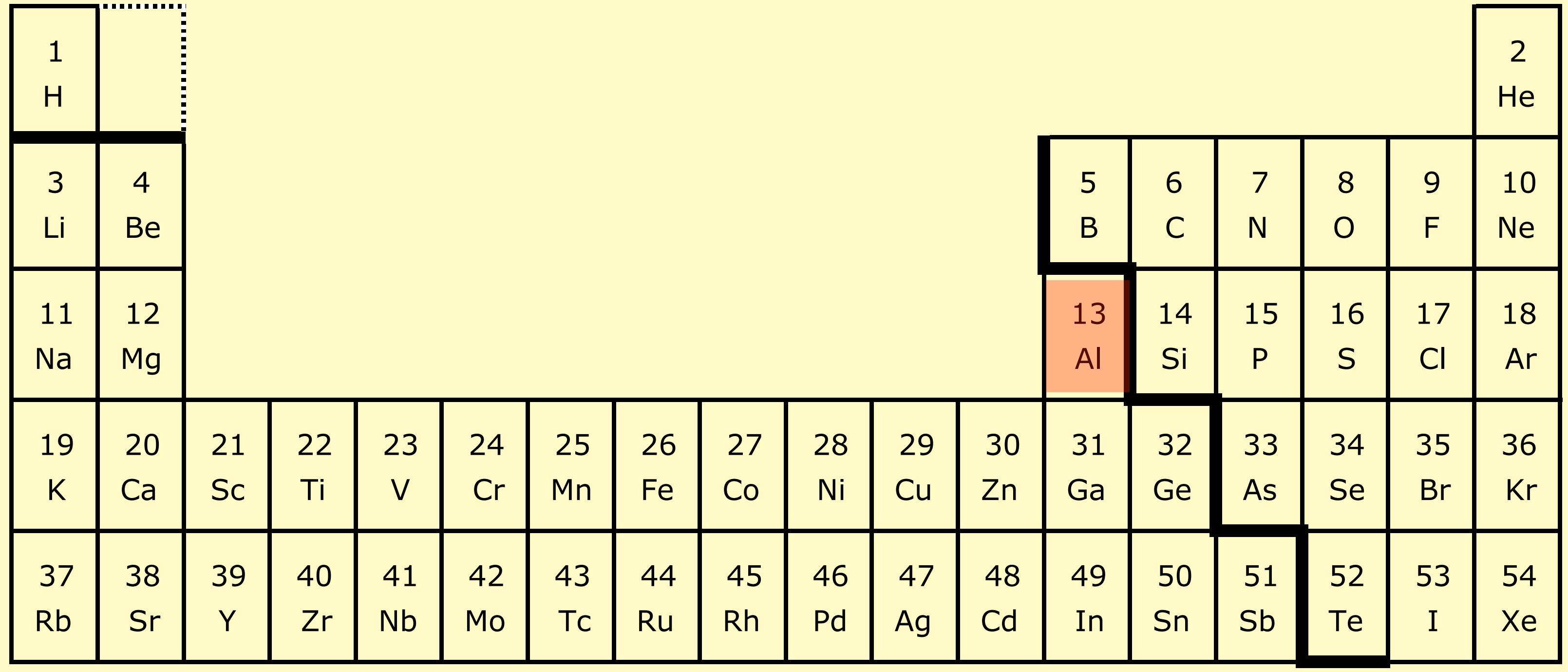
1. Mg < Al < Ar
2. S < Si < Na — *In the SAME energy level (same period in PT)* As you proceed from right to left, the smaller total proton charge has less attraction on the electron cloud. Coulombs Law.

1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe

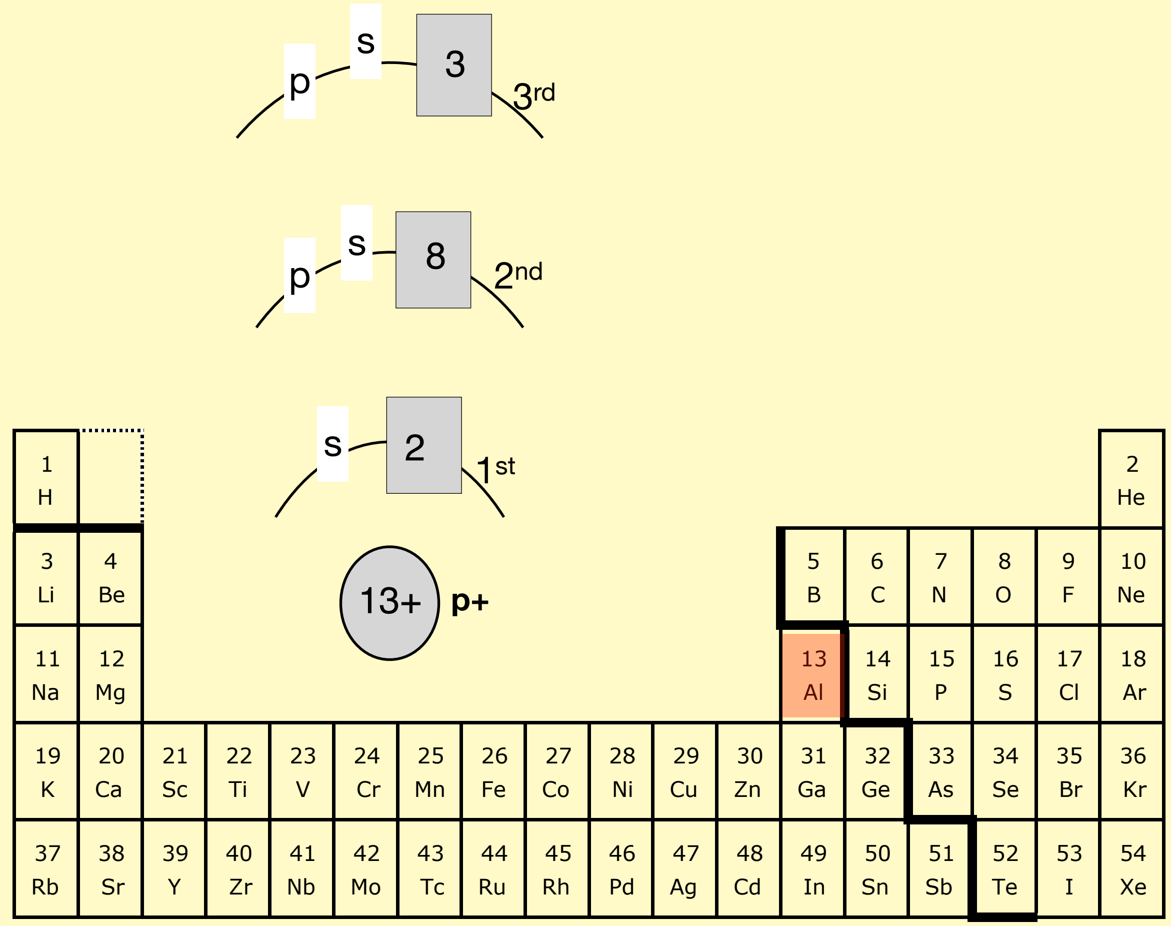
3. Rb < K < Li
4. P < As < Sb — Down the table in a column, the number of occupied energy levels increase making for larger atoms.

Which has larger radius?

1. Al
2. Al^{3+}



Which has larger radius?



1. Al
2. Al^{3+}

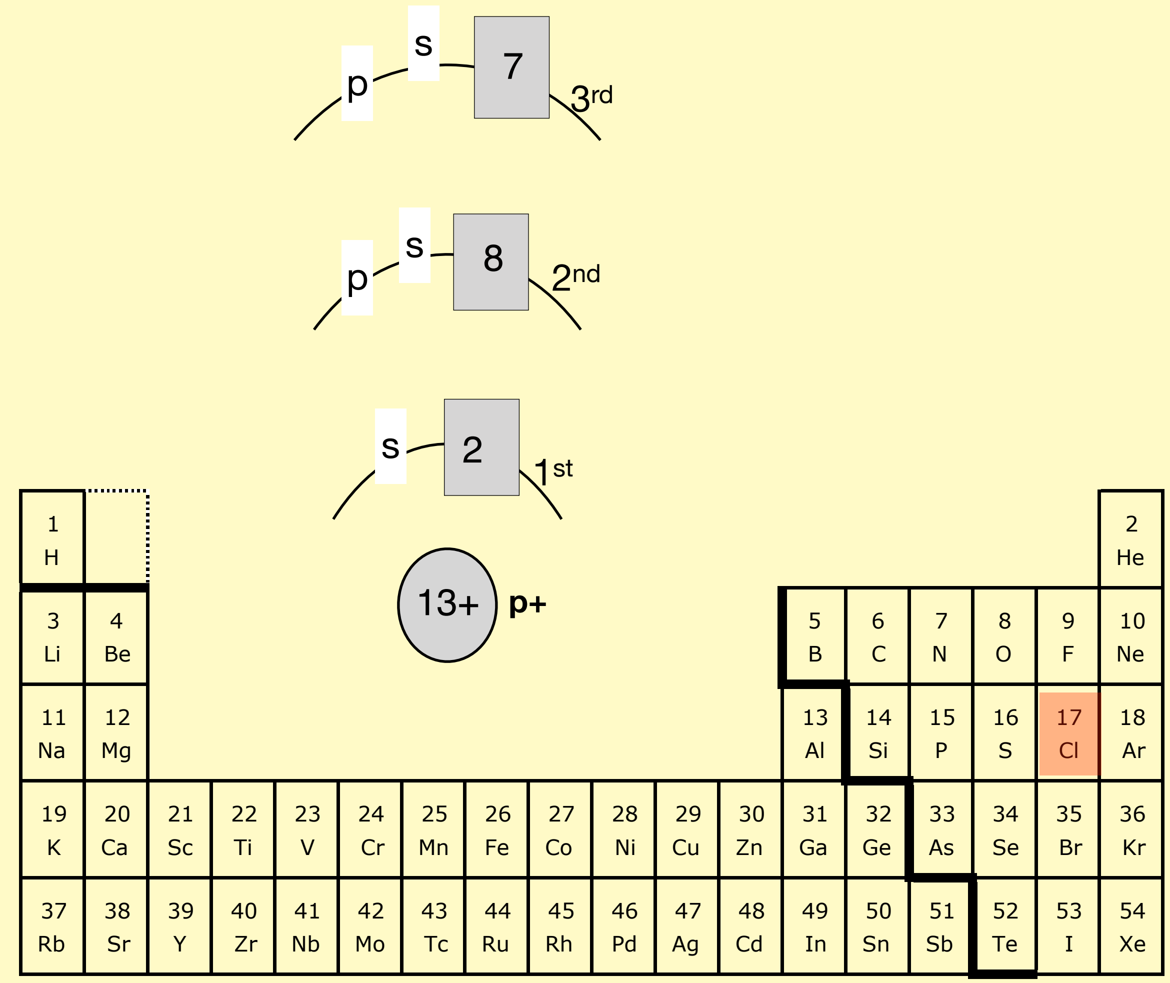
- same number of protons
- the entire 3rd **valence energy level has been removed** to make Al^{3+}, thus the valence electrons are one energy level closer to the nucleus.
- *All positive ions are smaller than their parent atom.*

Which has larger radius?

1. Cl
2. Cl^{-}

1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe

Which has larger radius?



1. Cl
2. Cl^-

- same # protons
- but the added electron **increases repulsion** within the valence shell increasing the size of the ion.
- All negative ions are larger than their parent atom

Which of the following gives the correct ordering of their ionic size (from smallest to largest)?

1. They are all the same size
2. $F^- < Ne < Na^+$
3. $F^- < Na^+ < Ne$
4. $Ne < F^- < Na^+$
5. $Na^+ < Ne < F^-$
6. $Na^+ < F^- < Ne$

1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	71 Lu	72 Hf	73 Ta	73 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

Which of the following gives the correct ordering of their ionic size (from smallest to largest)?

1. They are all the same size
2. $F^- < Ne < Na^+$
3. $F^- < Na^+ < Ne$
4. $Ne < F^- < Na^+$
5. $Na^+ < Ne < F^-$

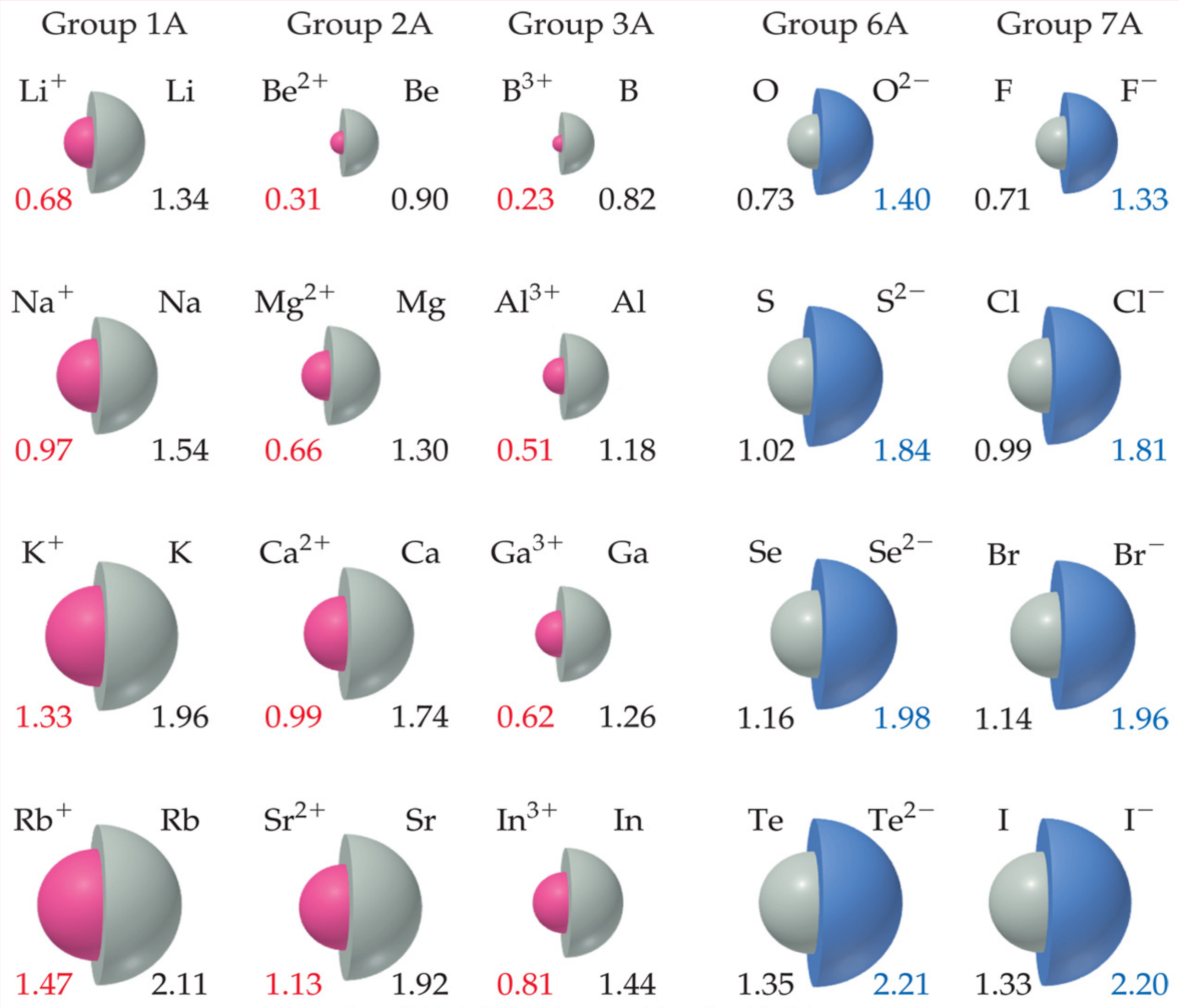
$$^{10e}_{11+}Na \quad ^{10e}_{10+}Ne \quad ^{10e}_{9+}F$$

- These particles are all isoelectronic, which might make you think they are all the same size, but you must consider the number of protons.
- **Na^+ has the most protons** thus greater nuclear attraction pulling in the valence electrons.

6. $Na^+ < F^- < Ne$

Size of Atoms and their Ions

- Cations (pink)
  - ✓ the atom lost electrons
  - ✓ the ion's valence electrons are one energy level closer to nucleus
  - ✓ are always smaller than their parent atom
- Anions (blue)
  - ✓ the atom gains e− increasing repulsion
  - ✓ are always larger than their parent atom



Ionization Energy

Some electrons go easy and others go hard, but none goes for free. (Munowitz)

What can we learn from measuring the amount of energy required to make an ion?

Ionization Energy

- The amount of energy required to forcibly remove an electron from a gaseous atom.
- Energy added as
  - light (electromagnetic radiation)
  - heat
  - electricity
    - Equation: $X + IE \rightarrow X^{+} + e^{-}$

First Ionization Energy (kJ/mole)

First *meaning the energy to remove one electron, which would be the easiest electron to remove.*

	1	2	3 (13)	4 (14)	5 (15)	6 (16)	7 (17)	8 (18)	
1	H 1311							He 2370	1
2	Li 521	Be 899	B 799	C 1087	N 1404	O 1314	F 1682	Ne 2080	2
3	Na 496	Mg 737	Al 576	Si 786	P 1052	S 1000	Cl 1245	Ar 1521	3
4	K 419	Ca 590	Ga 579	Ge 762	As 944	Se 941	Br 1140	Kr 1351	4
5	Rb 403	Sr 550	In 558	Sn 709	Sb 832	Te 869	I 1009	Xe 1170	5
6	Cs 376	Ba 503	Tl 589	Pb 716	Bi 703	Po 812	At	Rn 1037	6
7	Fr	Ra							7

First Ionization Energy

- Decreases down the chart
  - ✓ meaning easier to remove, less force required
  - ✓ due to valence electrons occupying energy levels **further from the attractive force of the protons in the nucleus**

Coulombs Law

$$\downarrow F = k\frac{(Q^{+})(Q^{-})}{\uparrow d^{2}}$$

	1	2	3 (13)	4 (14)	5 (15)	6 (16)	7 (17)	8 (18)	
1	H 1311							He 2370	1
2	Li 521	Be 899	B 799	C 1087	N 1404	O 1314	F 1682	Ne 2080	2
3	Na 496	Mg 737	Al 576	Si 786	P 1052	S 1000	Cl 1245	Ar 1521	3
4	K 419	Ca 590	Ga 579	Ge 762	As 944	Se 941	Br 1140	Kr 1351	4
5	Rb 403	Sr 550	In 558	Sn 709	Sb 832	Te 869	I 1009	Xe 1170	5
6	Cs 376	Ba 503	Tl 589	Pb 716	Bi 703	Po 812	At	Rn 1037	6
7	Fr	Ra							7

First Ionization Energy

- Increases across the chart (L-R)
  - ✓ meaning harder to remove, more force required
  - ✓ due to **increased proton charge** acting on electrons in same energy level.
  - ✓ ever increasing effective nuclear charge across the periodic table

The anomalies in groups 3(13) and 6(16) will be addressed in slides further along.

	1	2	3 (13)	4 (14)	5 (15)	6 (16)	7 (17)	8 (18)	
			oops!			*oops!*			
1	H 1311							He 2370	1
2	Li 521	Be 899	B 799	C 1087	N 1404	O 1314	F 1682	Ne 2080	2
3	Na 496	Mg 737	Al 576	Si 786	P 1052	S 1000	Cl 1245	Ar 1521	3
4	K 419	Ca 590	Ga 579	Ge 762	As 944	Se 941	Br 1140	Kr 1351	4
5	Rb 403	Sr 550	In 558	Sn 709	Sb 832	Te 869	I 1009	Xe 1170	5
6	Cs 376	Ba 503	Tl 589	Pb 716	Bi 703	Po 812	At	Rn 1037	6
7	Fr	Ra							7

Anomalies in groups 3(13) and 6(16) are addressed further in the next video.

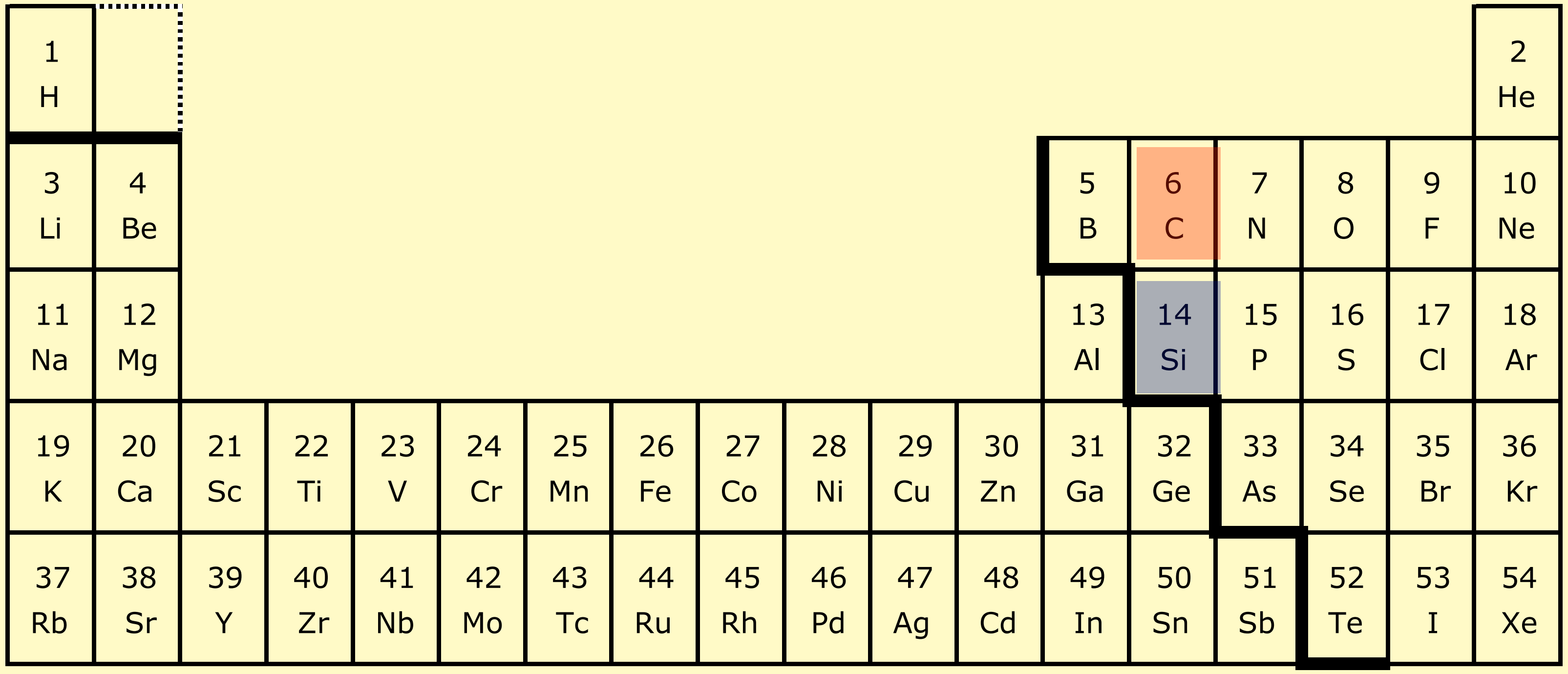
Coulombs Law

$$\uparrow F = k\frac{(\uparrow Q^{+})(Q^{-})}{d^{2}}$$

First *meaning the energy to remove one electron, and of course the easiest electron to remove*

Is the *first* ionization energy of carbon higher, lower, or the same as silicon?

1. higher
2. lower
3. same as



Is the *first* ionization energy of carbon larger, smaller or the same as silicon?

Coulombs Law

$$F = k\frac{(Q^+)(Q^-)}{d^2}$$

1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe

1. higher
2. lower
3. same as

- C has fewer occupied energy levels thus the valence electrons are closer to the nucleus resulting in greater nuclear attraction on the electron, making the ionization energy higher.

First *meaning the energy to remove one electron, and of course the easiest electron to remove*

Is the *first* ionization energy of carbon higher, lower or the same as nitrogen?

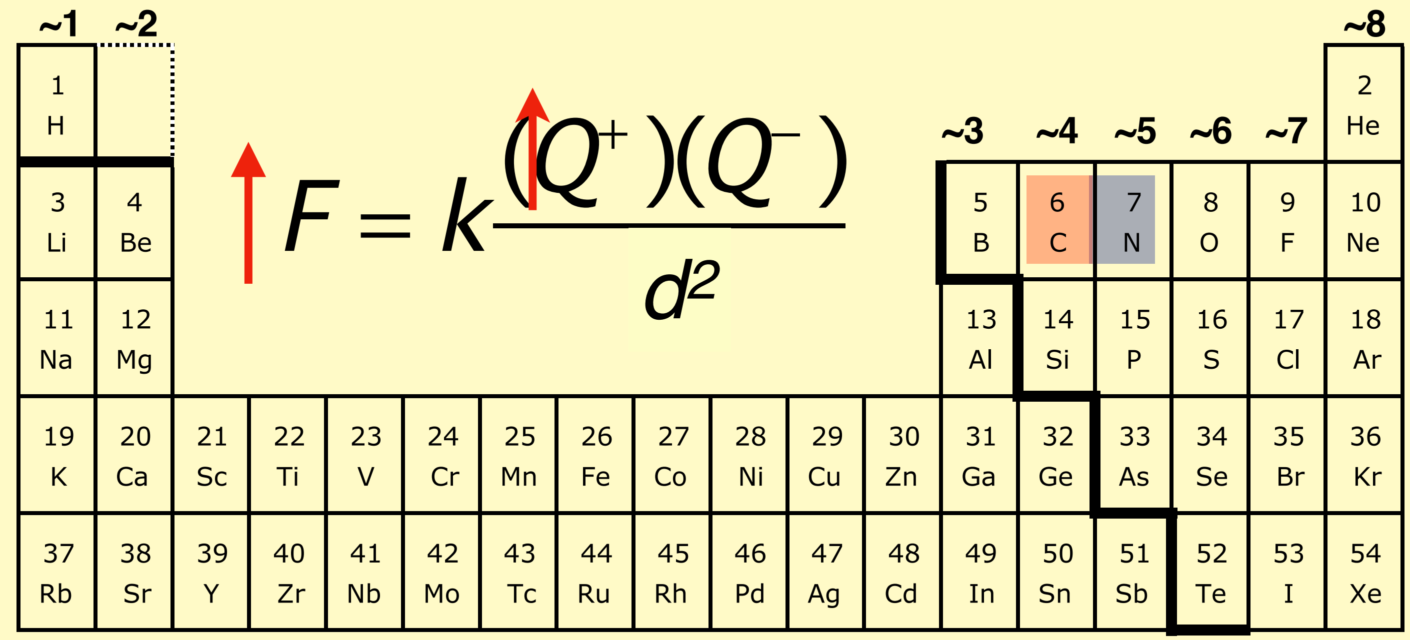
1. higher
2. lower
3. the same as

1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe

First meaning the energy to remove one electron, and of course the easiest electron to remove

Is the **first** ionization energy of nitrogen larger, smaller or the same as carbon?

ENC increases across the periodic table, and is approximately equal to the number of valence electrons of the atom



$$F = k\frac{(Q^{+})(Q^{-})}{d^2}$$

1. higher
2. lower

- N has one more proton provides greater nuclear attraction on electrons in the SAME energy level, making the valence electron harder to remove, and thus higher ionization energy.
- N has greater effective nuclear charge (enc).

3. the same as

Successive Ionization Energy

If you can learn something from knocking off one electron….what could we learn from knocking off another electron?

Successive Ionization Energy

- The amount of energy required to **repeatedly** remove electrons *one at a time, one after the other*.

- First: $X + IE \rightarrow X^{+} + e^{-}$

- Second: $X^{+} + IE \rightarrow X^{2+} + e^{-}$

- Third: $X^{2+} + IE \rightarrow X^{3+} + e^{-}$

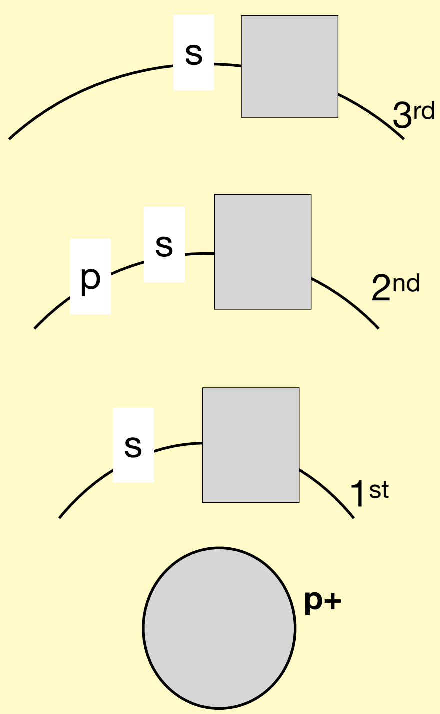
  ✓ Etc, etc, etc.

Which ionization energy will be larger?

1. 1st IE of Mg
2. 2nd IE of Mg

1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe

Hint: Fill in a Bohr Diagram if you are not sure.

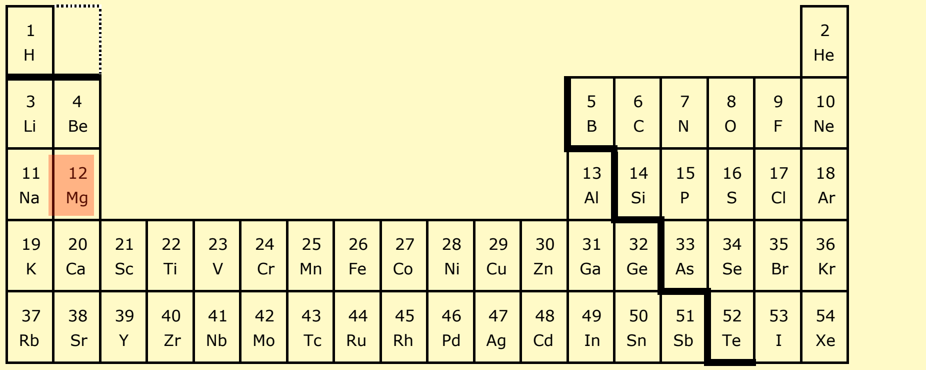


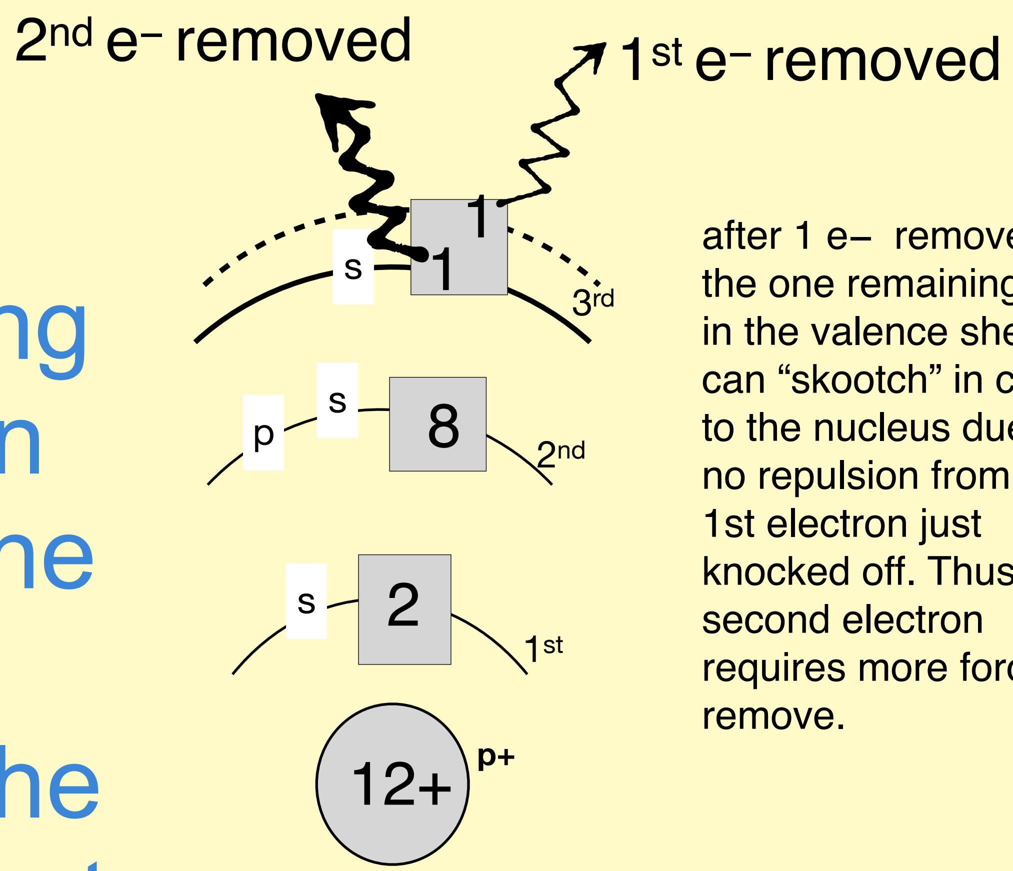
Which ionization energy will be larger?

Coulombs Law

$$\uparrow F = k\frac{(Q^{+})(Q^{-})}{\downarrow d^2}$$

1. 1st IE of Mg
2. 2nd IE of Mg
3. The loss of one electron in the valence shell reduces (eliminates) **the** repulsion between the two s valence electrons allowing the remaining valence electron to “scootch” in closer to the nucleus making the force required to remove the second electron greater than the force required to remove the first electron.





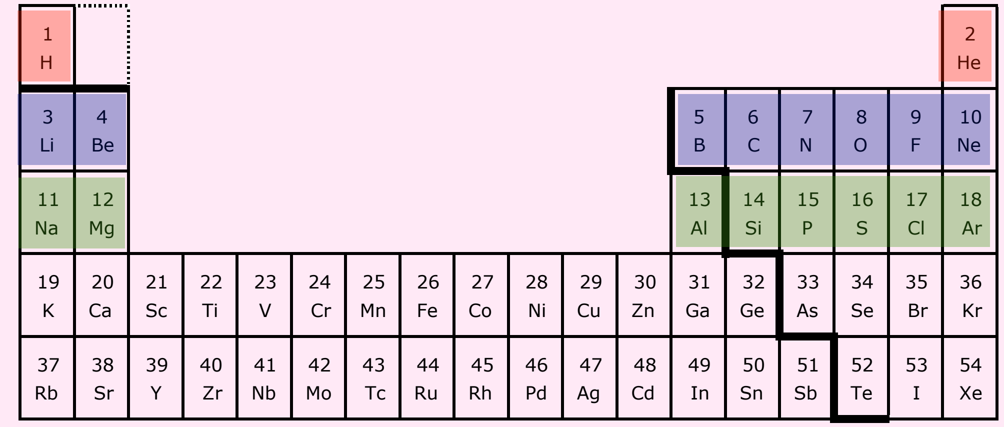
after 1 e− removed, the one remaining e− in the valence shell can “skootch” in closer to the nucleus due to no repulsion from the 1st electron just knocked off. Thus the second electron requires more force to remove.

Every successive ionization energy will be greater…. but how much greater?

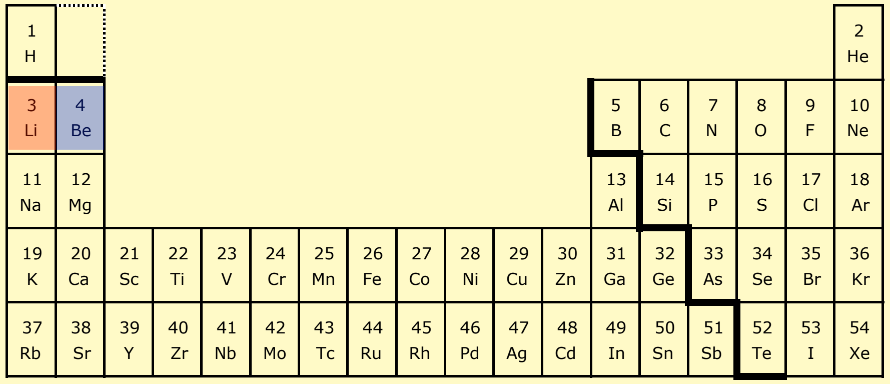
Successive Ionization Energy

Energy in kJ/mol

	1st	2nd	3rd	4th	5th	6th	7th	8th
H	1311							
He	2370	5220						
Li	521	7304	11752					
Be	899	1756	14849	20899				
B	799	2422	3657	25019	32660			
C	1087	2393	4622	6223	37822	46988		
N	1404	2856	4573	7468	9446	53250	63970	
O	1314	3396	5297	7468	10990	13325	71312	83652
F	1682	3367	6050	8423	11028	15167	17869	91950
Ne	2080	3946	6165	9301	12138	15148	19972	22963
Na	496	4564	6918	9542	13373	16644	20175	25501
Mg	737	1447	7738	10546	13624	18033	21767	25742
Al	576	1814	2750	11578	14820	18361	23465	27575
Si	786	1582	3232	4361	16007	19693	23658	29110
P	1052	1901	2914	4959	6272	21516	25858	30489
S	1000	2258	3387	4544	6947	8500	27112	31734
Cl	1245	2287	3850	5162	6542	9359	11028	33442
Ar	1521	2653	3927	5886	7526	8587	11964	13778

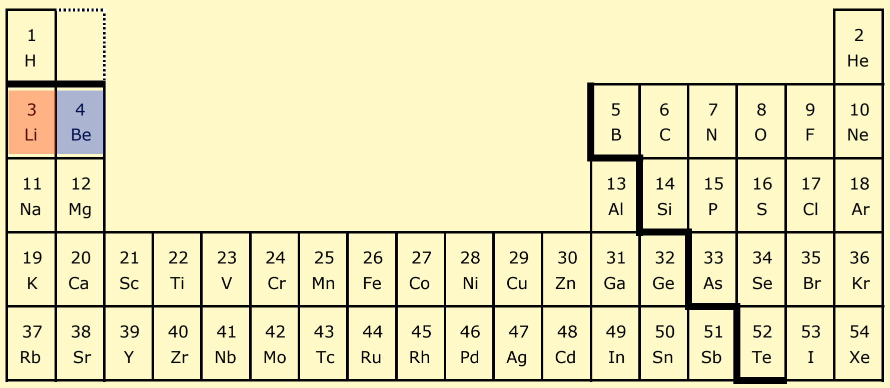


Why is Li’s 2nd I.E. so much larger than 1st, but Be’s 3rd I.E. so much larger than 2nd?

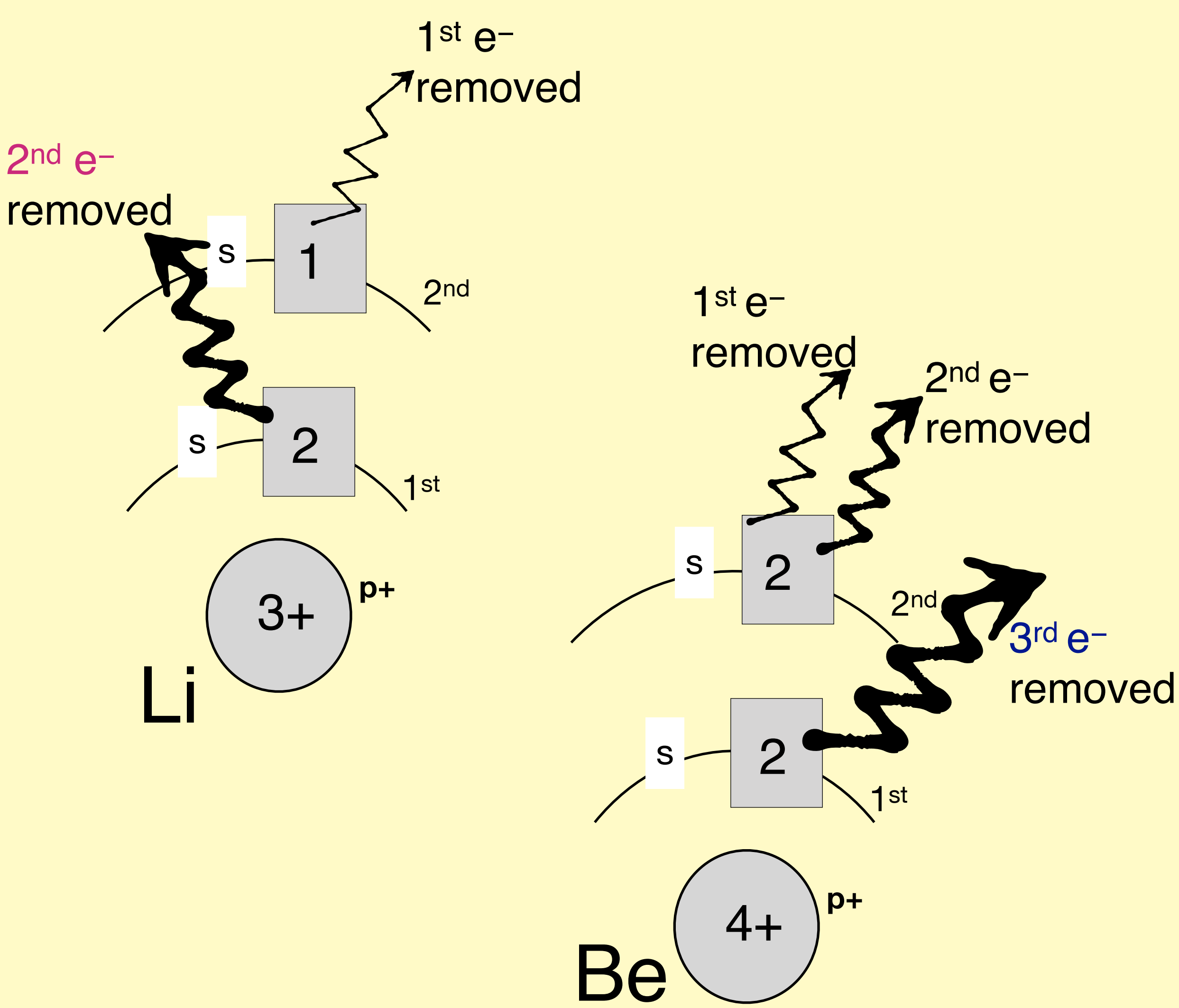


Hint: Fill in a Bohr Diagram might help.

Why is Li's 2nd I.E. so much larger than 1st, but Be's 3rd I.E. so much larger than 2nd?



Hint: Fill in a Bohr Diagram if you are not sure.



- The second e⁻ of Li is removed from the next closer energy level and thus experiences greater nuclear attractive force.
- The third e⁻ of Be is removed from the next closer energy level and thus experiences greater nuclear attractive force.

For which successive ionization energy in Ga would we expect to see an extremely large increase (compared to the previous IE)

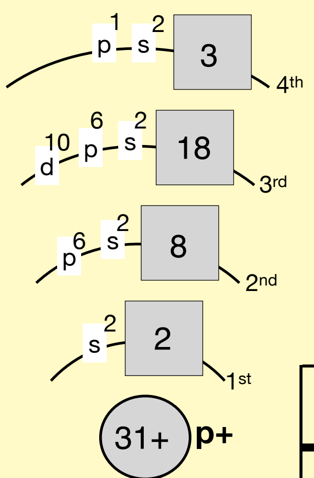
1. none any different...successive ionization energies increase steadily with each electron removed.
2. 2nd
3. 3rd
4. 4th
5. 5th
6. 13th

1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe

For which successive ionization energy in Ga would we expect to see an extremely large increase (compared to the previous IE)

4. 4th

- At that point, the electron being removed is **one energy level closer to the nucleus** (attractive force) than the previous electron. The very large increase always occurs at one more than the number of valence electrons.



Coulombs Law

$$F = k\frac{(Q^+)(Q^-)}{d^2}$$

1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe

- The electrons in the closer energy level will **experience a much greater effective nuclear charge**.
- *The large increase always occurs when removing the electron that is one more than the number of valence electrons.*

Anomalies in the IE trend

Groups 3(13) and 6(16)

First Ionization Energy

- In general increases across the chart (L-R)

Anomalies in the trend in groups 3(13) and 6(16)
What's going on?!!?

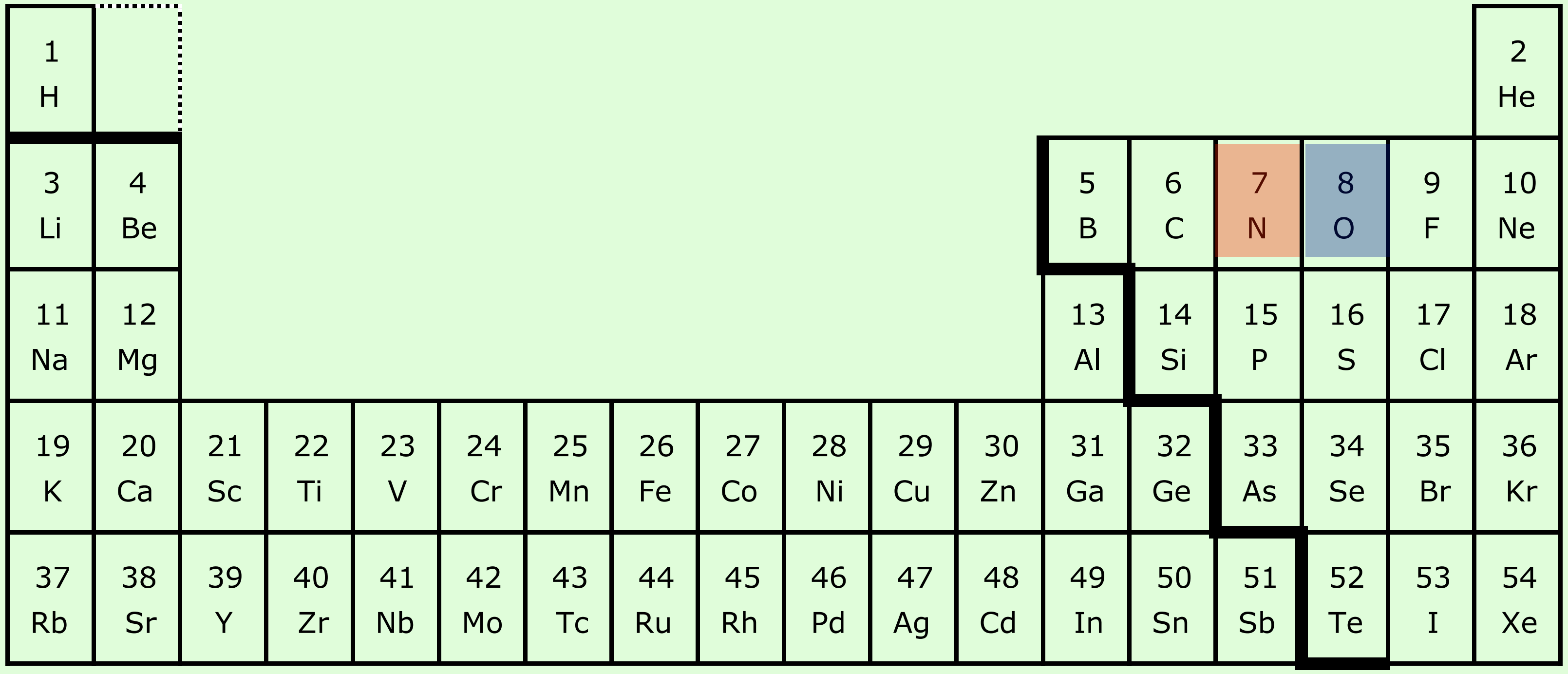
	1	2	*oops!* 3 (13)	4 (14)	5 (15)	*oops!* 6 (16)	7 (17)	8 (18)	
1	H 1311							He 2370	1
2	Li 521	Be 899	B 799	C 1087	N 1404	O 1314	F 1682	Ne 2080	2
3	Na 496	Mg 737	Al 576	Si 786	P 1052	S 1000	Cl 1245	Ar 1521	3
4	K 419	Ca 590	Ga 579	Ge 762	As 944	Se 941	Br 1140	Kr 1351	4
5	Rb 403	Sr 550	In 558	Sn 709	Sb 832	Te 869	I 1009	Xe 1170	5
6	Cs 376	Ba 503	Tl 589	Pb 716	Bi 703	Po 812	At	Rn 1037	6
7	Fr	Ra							7

The forces on electrons

- Protons pull on electrons. As you know from Coulombs Law, higher nuclear charge attract more strongly than lower nuclear charge.
- In multi-electron atoms, a second force comes into play, the repulsion from other electrons. This repulsion can counteract the nuclear attractions.
- Explaining the force required to remove electrons requires a consideration of the attractive force of the nucleus the repulsive forces of other electrons.

Is the first ionization energy of oxygen larger, smaller or the same as nitrogen?

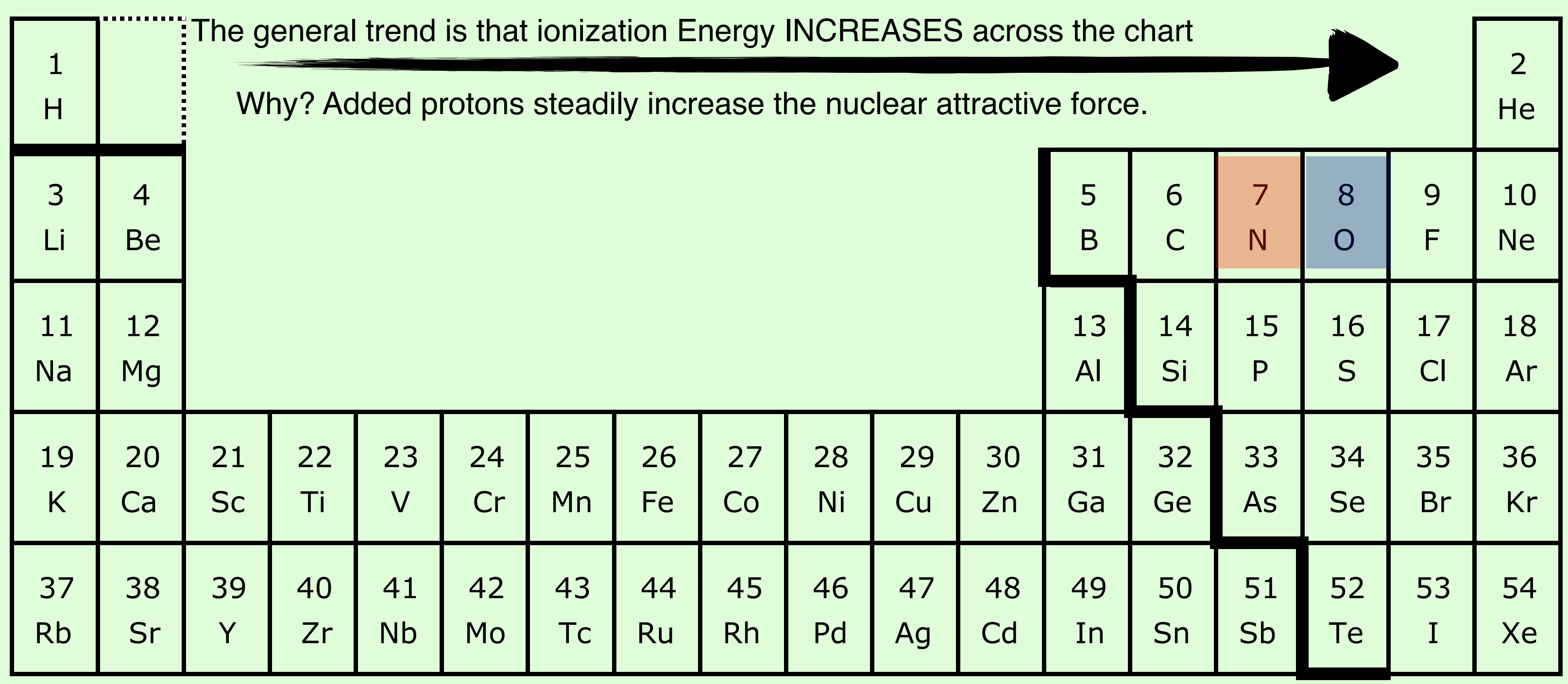
1. larger
2. smaller
3. the same as



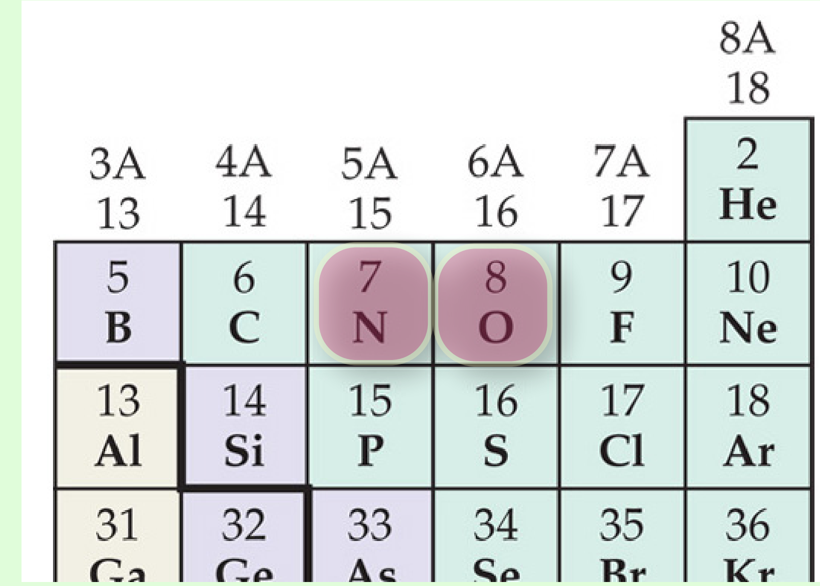
The general trend of increasing IE might make you think the I.E. of oxygen is larger, however the I.E. of oxygen is lower than for nitrogen.

Why?

While the original question implies that you would have to remember that the oxygen group is an anomaly, I am confident that in a free response (maybe not in MC) AP would tell you oxygen has a lower IE than nitrogen.



The general trend of increasing I.E. might make you think the I.E. of oxygen is larger, however the I.E. of oxygen is lower than for nitrogen.



Even though there is one added proton which increased the nuclear attraction, the *added repulsion* of 4th *p* electron that is *paired* in the *p* orbital offsets the increased nuclear attraction.

This repulsion raises the orbital energy and the I.E. required to remove that paired *p* electron for oxygen will be lower than for nitrogen

Orbital notation: N ⊗ ⊘⊘⊘ O ⊗ ⊗⊘⊘

In fluorine and neon, the added proton(s) restore the increasing trend.

Is the first ionization energy of boron larger, smaller or the same as beryllium?

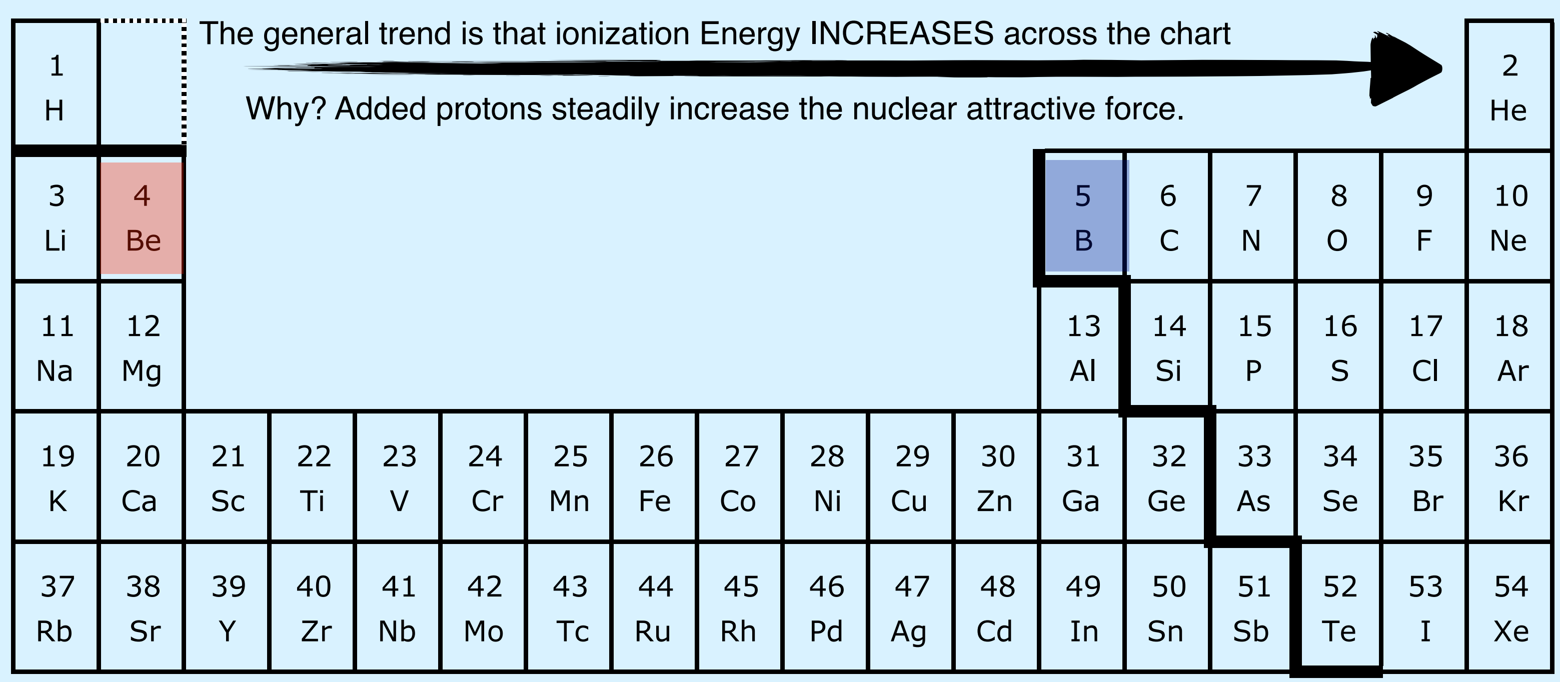
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe

1. larger
2. smaller
3. the same as

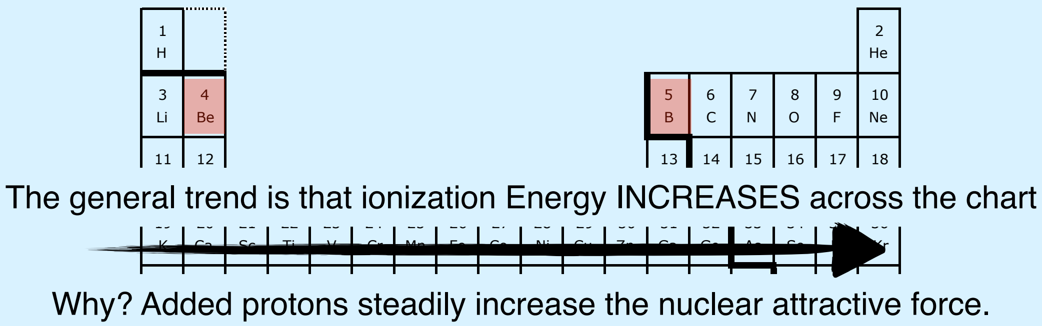
The general trend of increasing IE might make you think the I.E. of beryllium is larger, however the I.E. of boron is lower than for berylium.

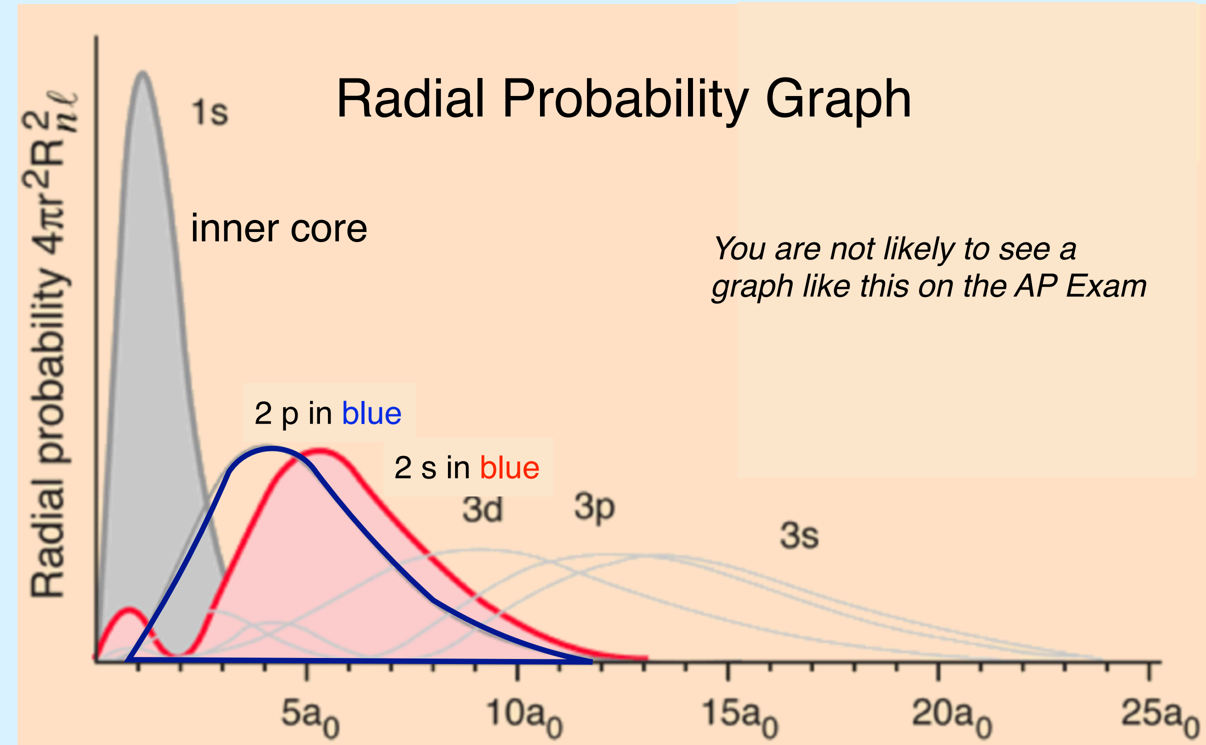
Why?

While the original question implies that you would have to remember that the oxygen group is an anomaly, I am quite confident that in a free response (maybe not in MC) AP would tell you boron has a lower IE than beryllium.



B’s I.E. is lower than Be’s. Why?

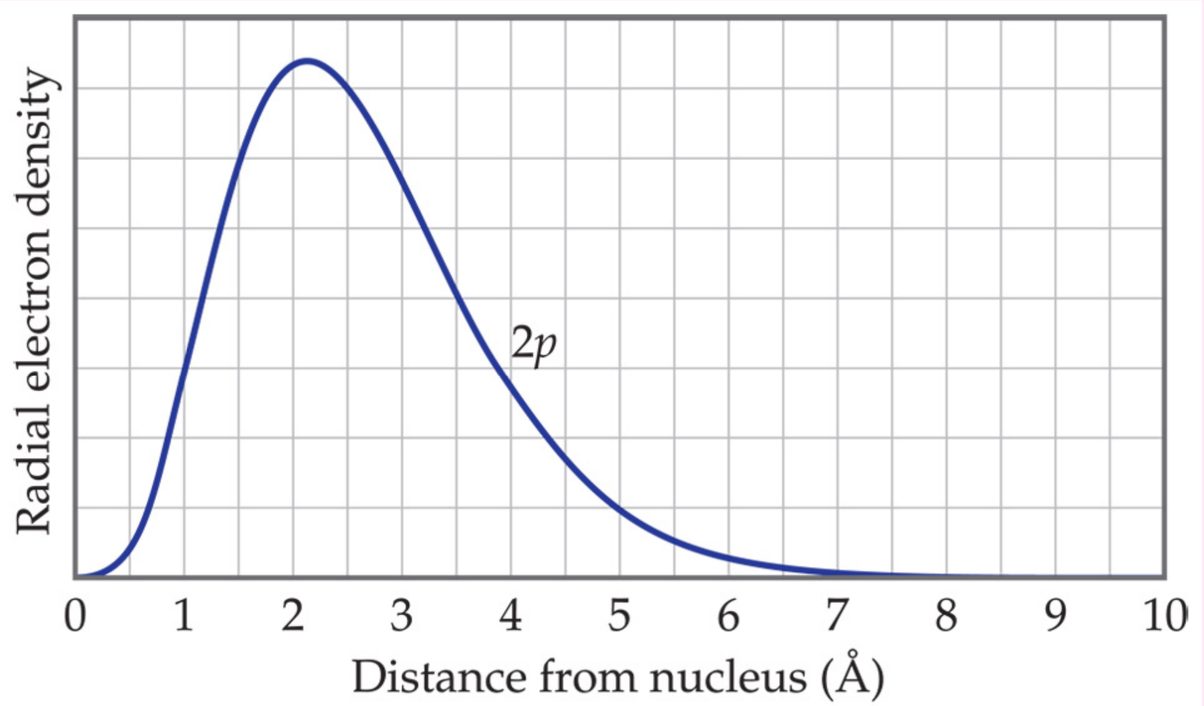


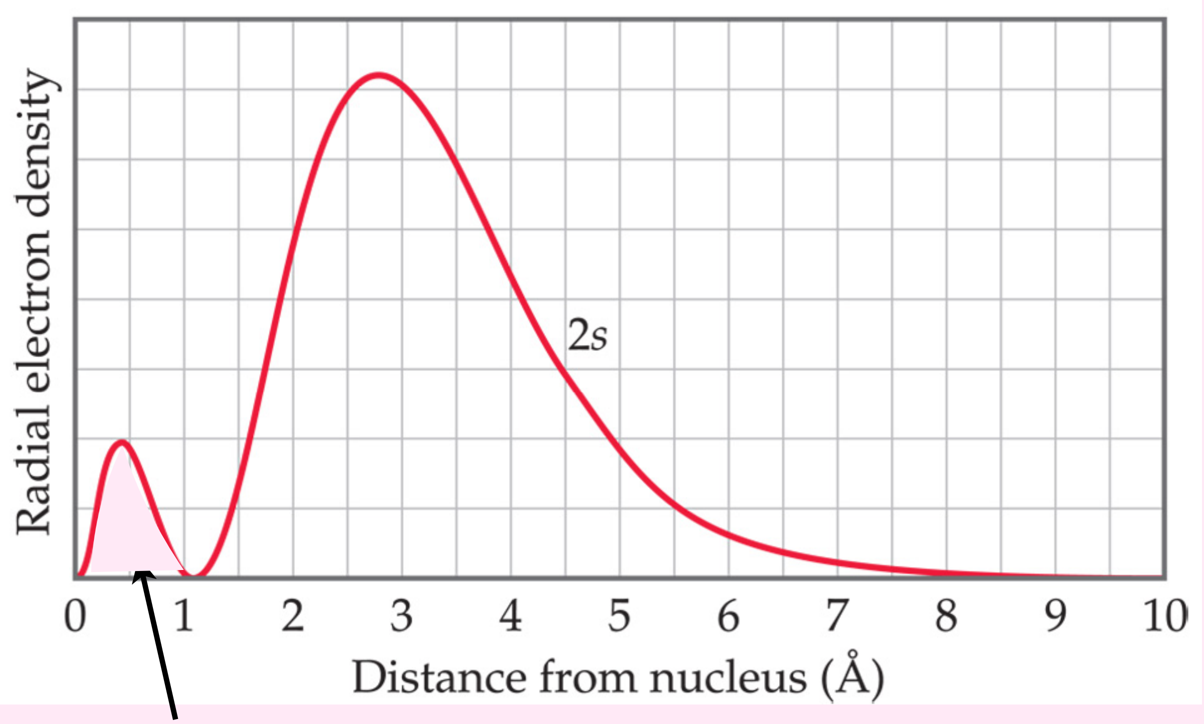


- The trend of increasing I.E. due to increased nuclear charge when proceeding to the right across the periodic table would indicate B’s I.E. should be larger, however the electron to be removed in B is in the 2p orbital (compared to 2s for Mg) and 2p has **less penetration** into the inner core, nearer to the nucleus, making the Coulombic attractive force on the 2p electron in B smaller.
- In the next elements, C and N, added proton(s) restore the increasing trend.

Orbital Penetration & Nuclear Force

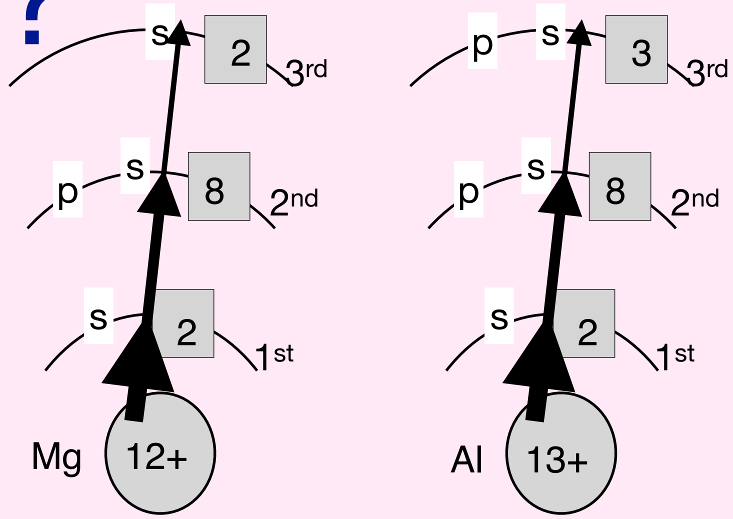
- The radial probability function for the 2s orbital of the hydrogen atom (red curve) shows an area of probability close to the nucleus, whereas the 2p orbital (blue curve) does not have this.
- As a result, an electron in the 2s orbital for a many-electron atom experiences more of the nuclear charge than does an electron in the 2p orbital.
- This difference leads to the observation that in a many-electron atom the orbitals for a given *n* value increase in energy s < p < d < f, that is, orbital *n*s is lower energy than *n*p, which is lower energy than *n*d.

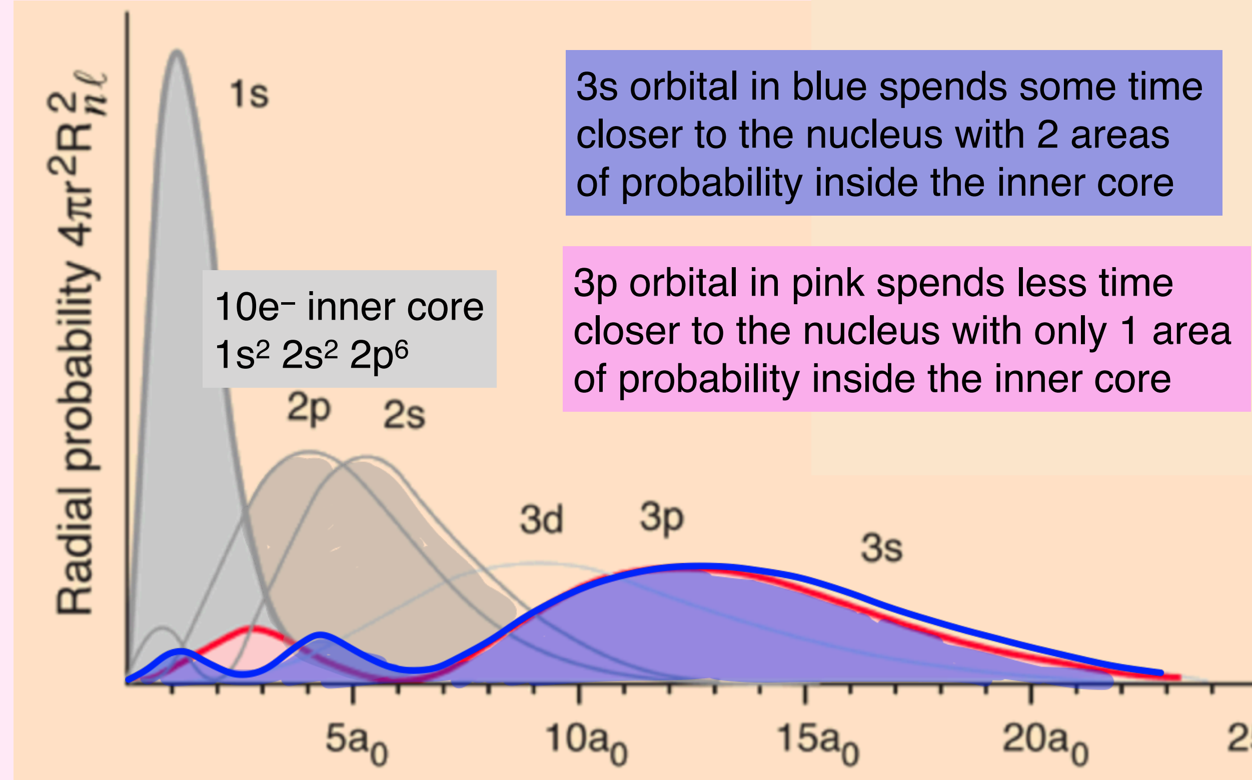




This extra area of probability of the 2s orbital close to the nucleus makes 2s electrons lower energy and thus more stable than 2p electrons.

Why is Al's I.E. lower than Mg's?





- The $3p^1$ electron in pink on the graph has one area of probability inside the inner core. As a consequence of this "less penetration," the inner core is not quite as effective in shielding the $3p^1$ electron from the nucleus.
- Thus the ionization energy of the Al atom is lower than what otherwise might be expected.
- The growing nuclear charge restores the trend of increasing ionization energy for Si and P
  - Historically AP has not put these graphs on the AP exam.

PES

Photo Electron Spectroscopy

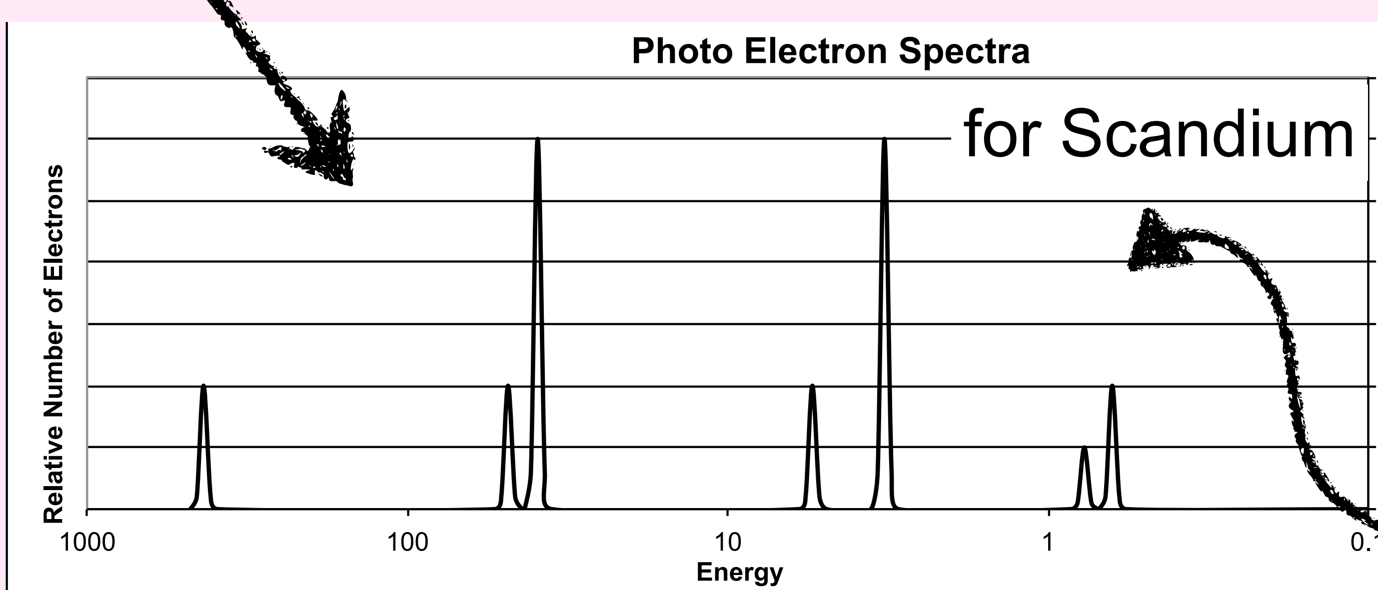
Evidence for the Validity of the Shell Orbital Model

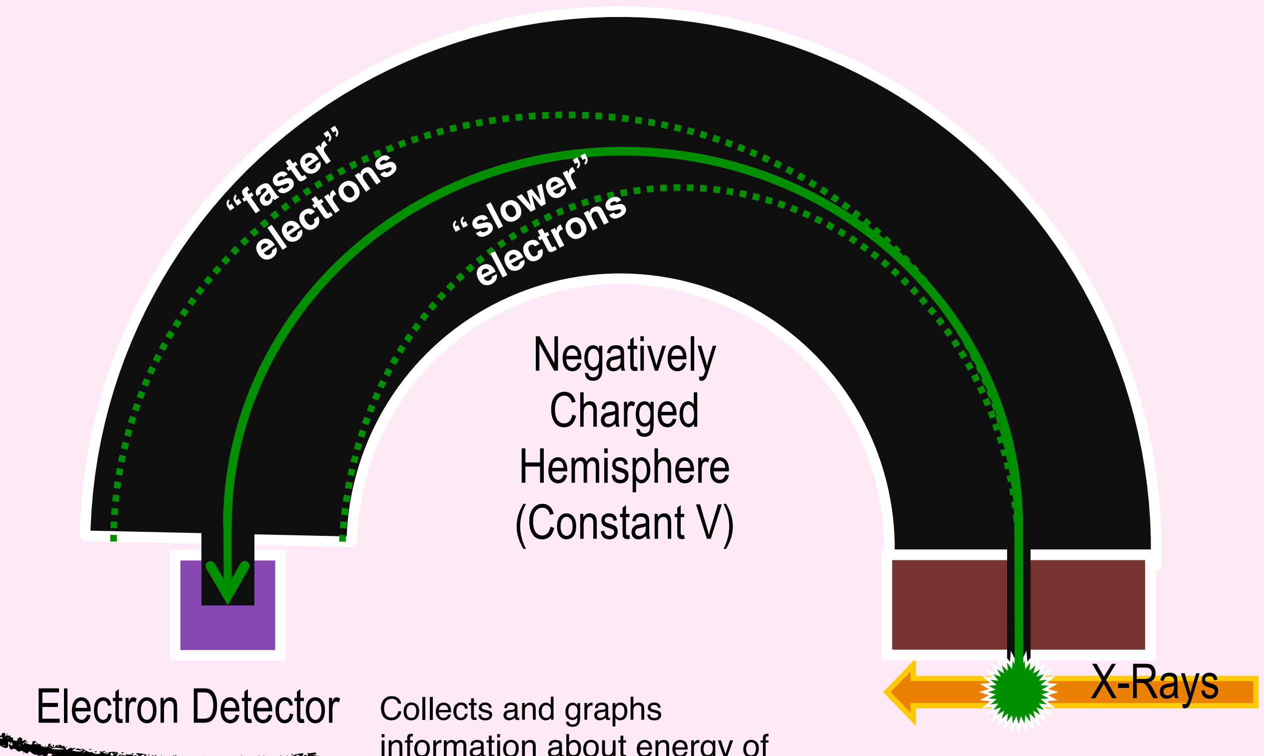
PES – Measuring Binding Energy

- X-Rays of a particular energy ($E_{x\text{-}ray}$) are blasted at atoms
- Electrons from each energy level of that atom are blown away with varying amounts of excess kinetic energy from the X-ray
- The charged hemisphere bends that spray of electrons
  - ✓ “faster” (more KE) electrons bend less, “slower” (Less KE) electrons bend more
- Detector tells us how many of each electron of a particular energy ($KE_{detected}$).

You do NOT need to explain this measurement process, you simply need to be able to interpret the resulting graphs of spectra.

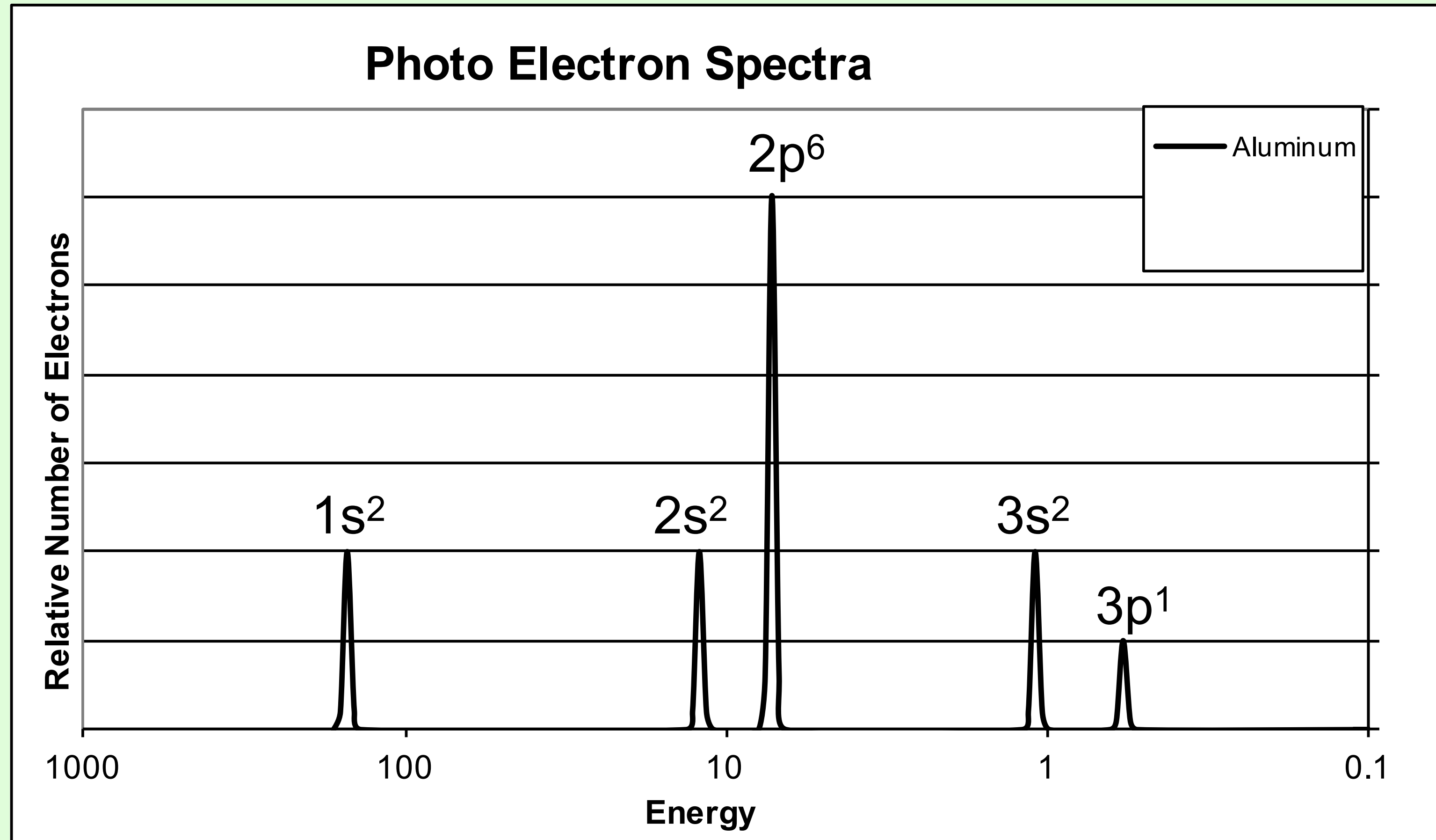
- Binding energy ($BE_{electron}$) can be calculated
- $BE_{electron} = E_{x\text{-}ray} - KE_{detected}$





PES of Aluminum

- Note that the numbers on the x-axis are larger to smaller.
- This is by design so the left side can represent the nucleus



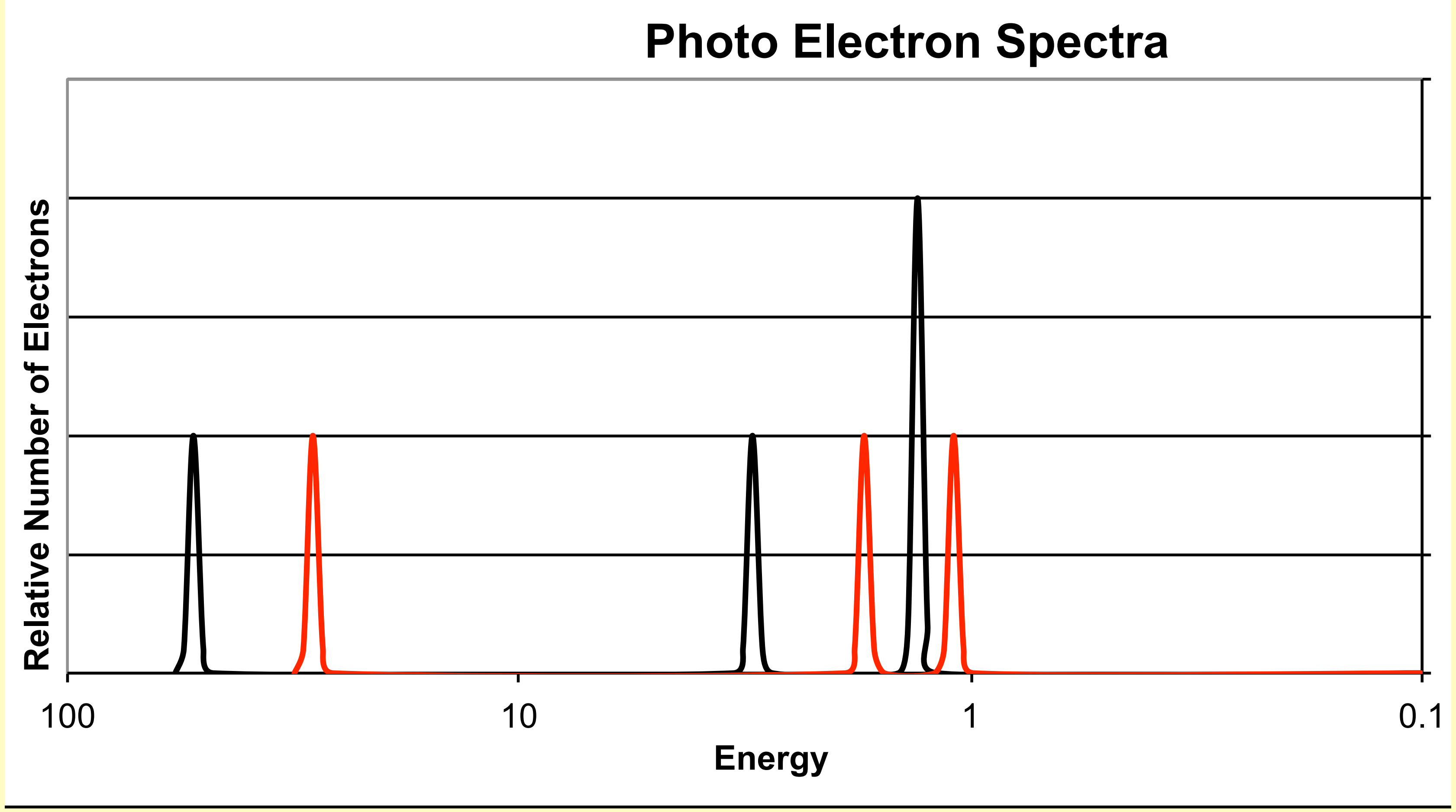
Binding energy, the energy required to remove the electrons represented by the peaks shown

- It is very easy to apply an electron configuration to the set of peaks.
- Each peak represents a different sublevel
- Height of the peak indicates the number of electrons within that sublevel, with the most left peak showing the height of 2 electrons (or 1 electron if the PES is for hydrogen.)

These two separate (black and red) spectra represent two different elements overlayed on the same graph.

Which element is represented by the black peaks shown below.

1. lithium
2. carbon
3. oxygen
4. neon



These two separate (black and red) spectra represent two different elements overlayed on the same graph.

Which element is represented by the black peaks shown below.

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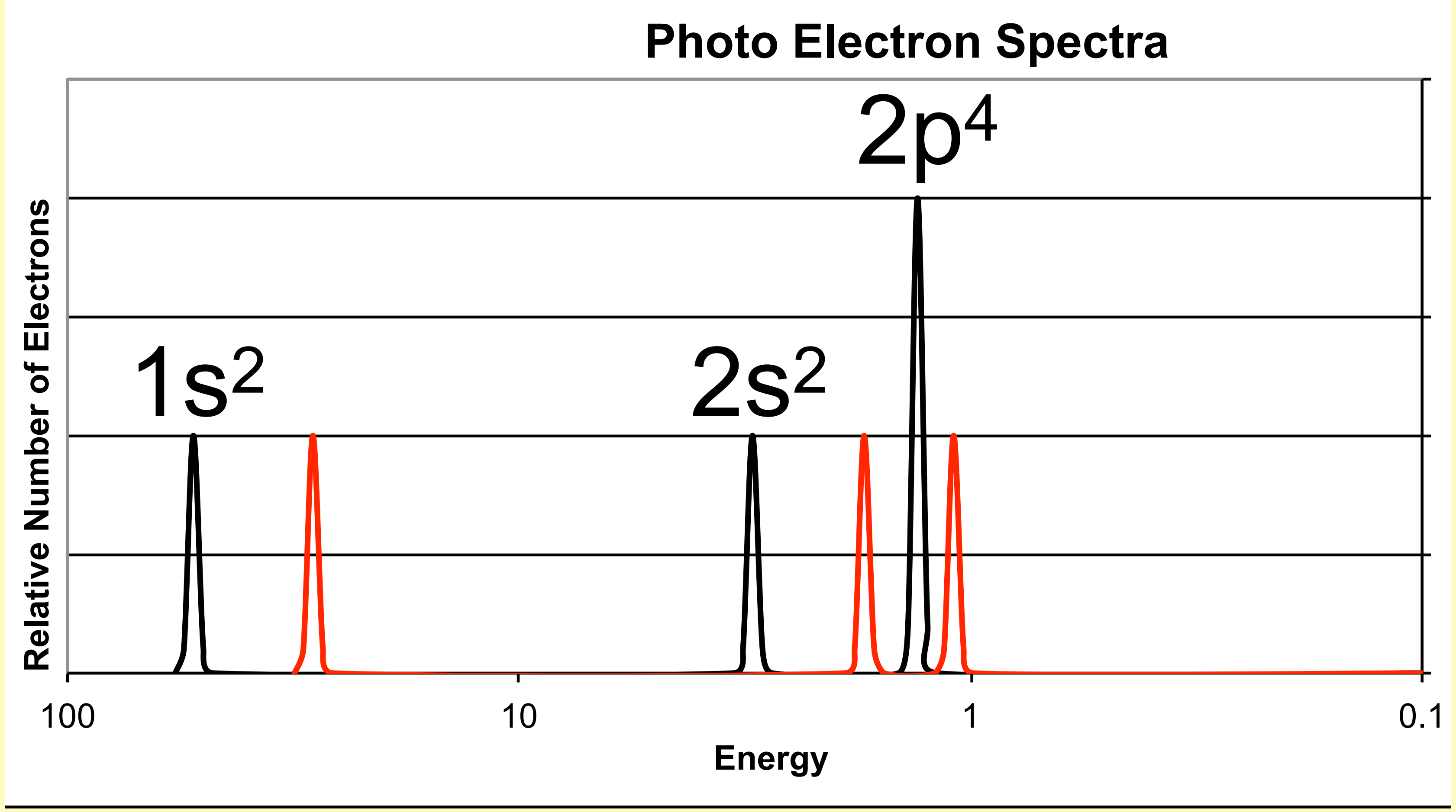
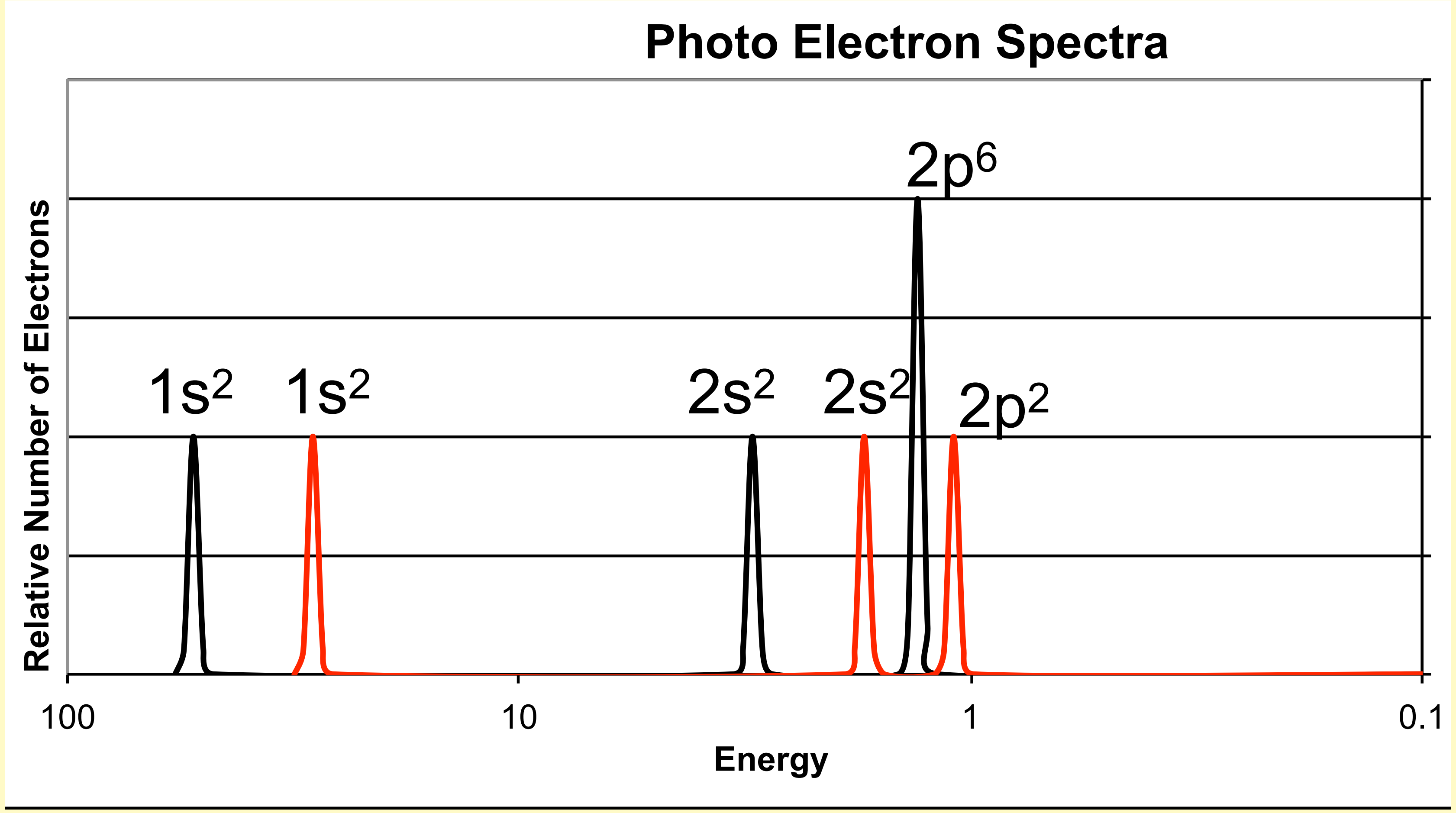


And the red peaks represent which element?

Which element is represented by the PES red peaks shown below.

1. lithium
2. carbon
3. oxygen
4. neon

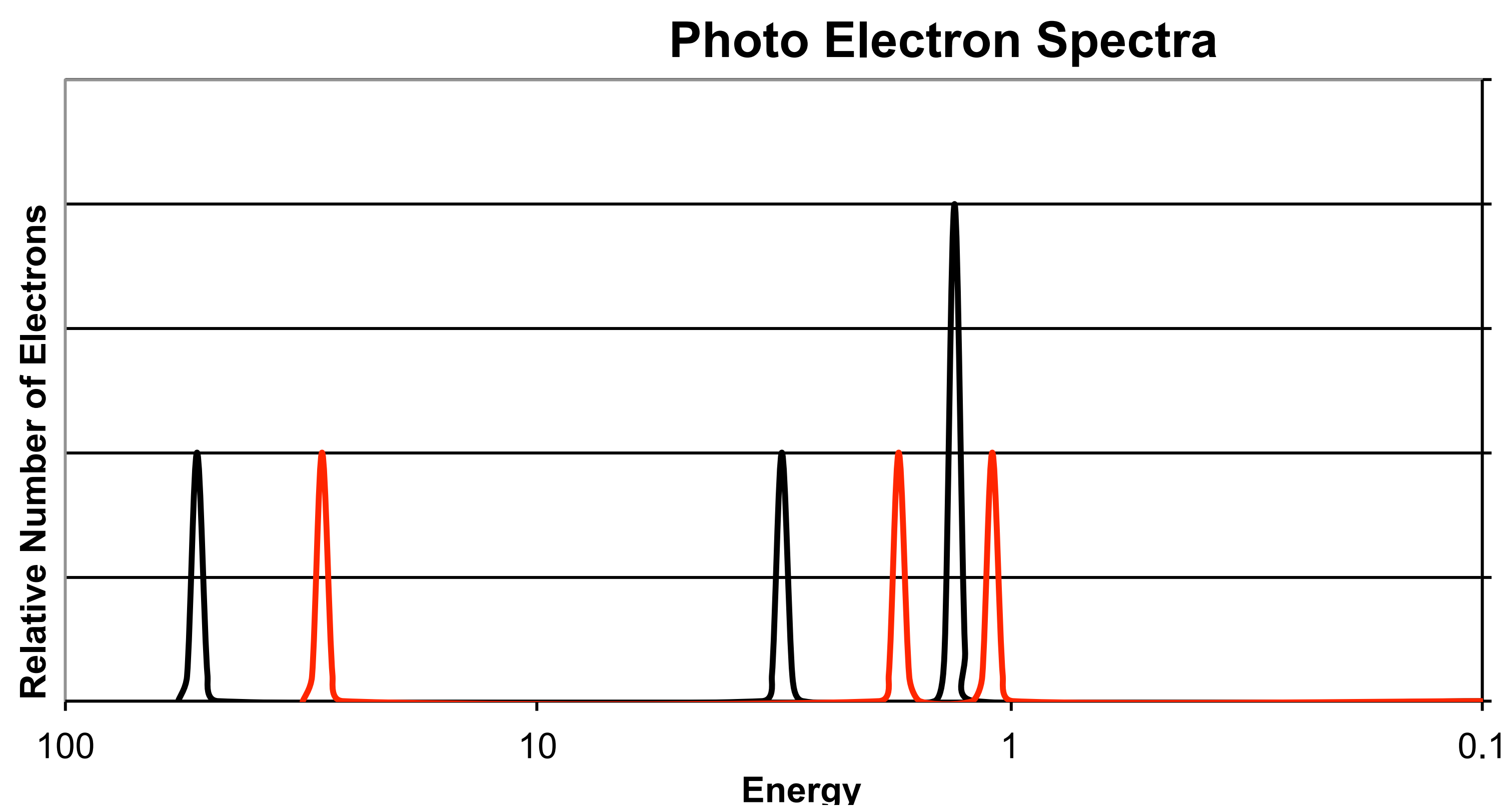


Explain to your mate why the black peaks of oxygen are to the left (and not in the same place) as the red peaks of carbon.

The black peaks represent oxygen
The red peaks represent carbon

2. carbon 6 protons
3. oxygen 8 protons

- The increased nuclear charge in oxygen makes the electrons in oxygen, held more tightly.



Binding energy, the energy required to remove the electrons represented by the peaks shown

Need more Practice? Go to Practice H1 – PES

- thus more binding energy required to remove those oxygen electrons

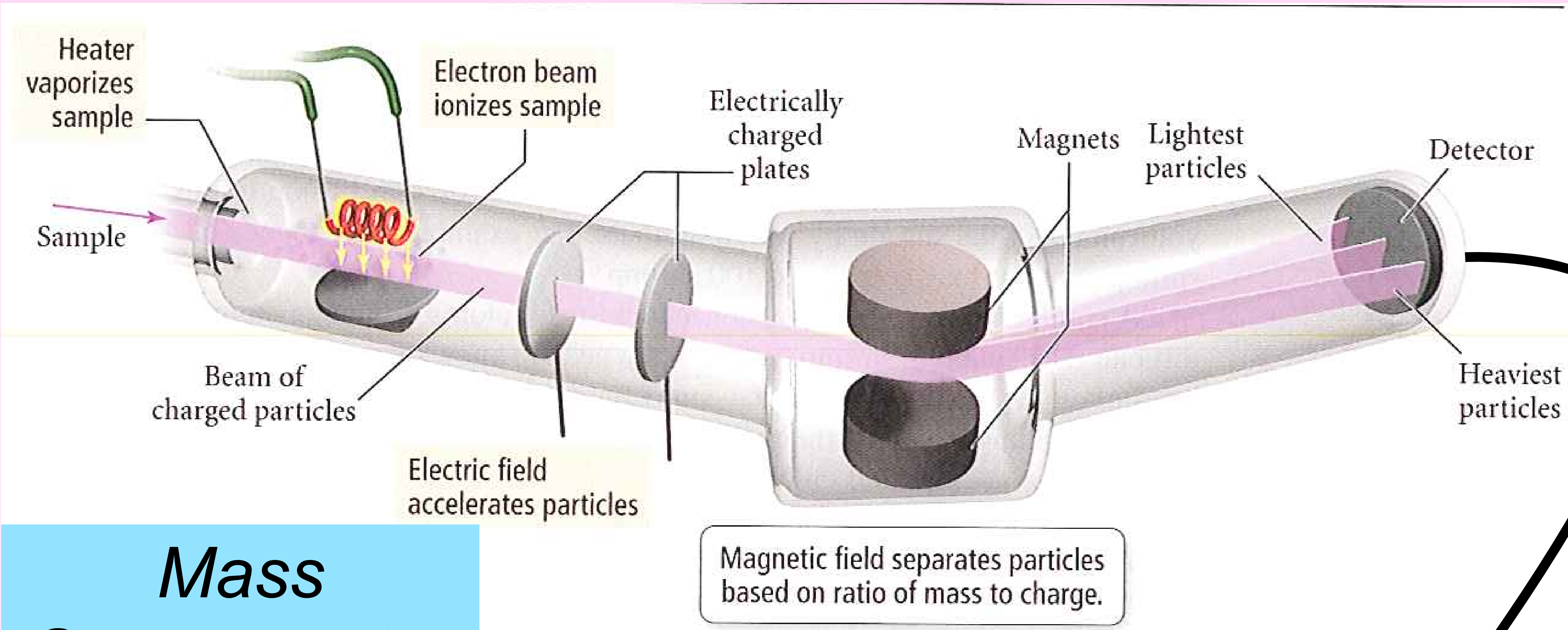
Isotopes

Different versions of the same element.

Equal amount of protons but different number of neutrons

Same Chemical Properties
Different Masses

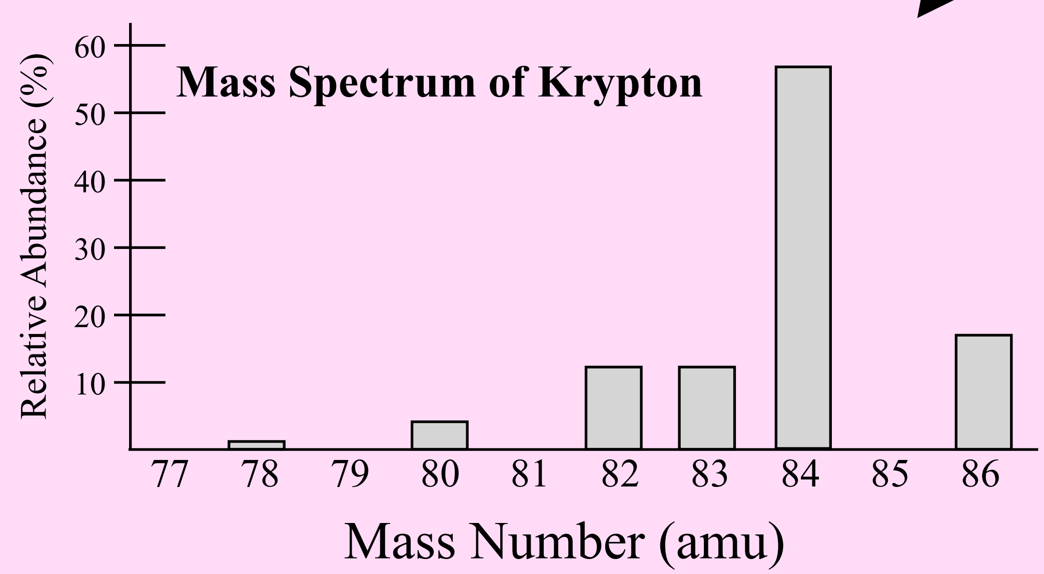
Mass Spectrometry; Determining masses and percentages of isotopes.



Mass Spectrometer

You do NOT need to explain this intrument nor measurement process, you simply need to be able to interpret the resulting graphs of spectra.

Graphical Readout from a Mass Spectrometer



Percent Abundance determined from signal intensity

^{78}Kr	0.356%
^{80}Kr	2.27%
^{82}Kr	11.6%
^{83}Kr	11.5%
^{84}Kr	57.0%
^{86}Kr	17.3%

Determine the average atomic mass for zinc from the four major isotopes that naturally occur.
(Five isotopes exist, but one is a very small %)

Of course you can look us the average molar mass in the periodic table, but let's calculate that value

Ignore sig figs right now, and report your answer to the nearest 10^{ths} place.

isotope	% abundance
^{64}Zn	49.6%
^{66}Zn	27.8%
^{67}Zn	4.1%
^{68}Zn	18.5%

Determine the average atomic mass for zinc from the four major isotopes that naturally occur.
(Five isotopes exist, but one is a very small %)

Of course you can look us the average molar mass in the periodic table, but let's calculate that value

Ignore sig figs right now, and report your answer to the nearest 10^{ths} place.

This is a "weighted "average

$$(64 \times 0.496) + (66 \times 0.278) + (67 \times 0.041) + (68 \times 0.185) = 65.4$$

This matches the periodic table value of 65.39 g/mol

isotope	% abundance
^{64}Zn	49.6%
^{66}Zn	27.8%
^{67}Zn	4.1%
^{68}Zn	18.5%

You would like to calculate the % abundance found in nature of chlorine, Cl. There are only two naturally occurring isotopes: ^{35}Cl and ^{37}Cl.

- Remember you need one more bit of info to solve this problem. Go to the periodic table to get that info.
- Report the % number to the nearest 10^{ths} place for the isotope that occurs in highest quantity.
- Report just the # no need to put % on the end of number.

You would like to calculate the % abundance found in nature of chlorine, Cl. There are only two naturally occurring isotopes: ^{35}Cl and ^{37}Cl.

- Report the % to the nearest 10^{ths} place for the isotope that occurs in highest quantity.
- 25% is ^{37}Cl
- 75% is ^{35}Cl

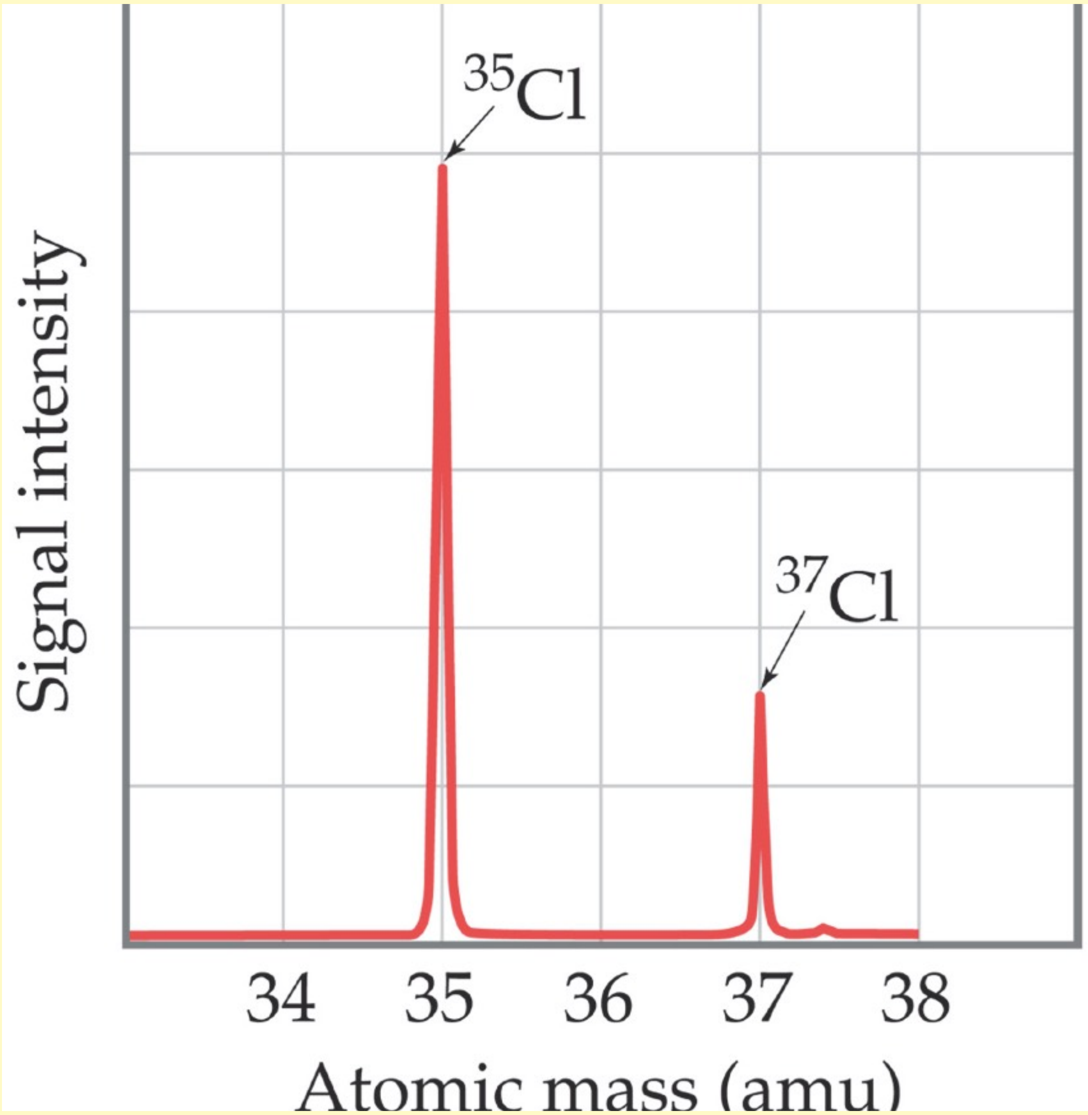
$$37x + 35(1 - x) = 35.5$$
$$37x + 35 - 35x = 35.5$$
$$2x = 0.5$$
$$2x = 0.25$$

$$x + y = 100\%$$
$$x + y = 1.0$$
$$\mathrm{y} = 1.0 - x$$

Chlorine has two isotopes, ^{35}Cl (~75% abundant) and ^{37}Cl (~25% abundant). How many different molar masses are possible for a Cl_2 molecule?

No Calculator

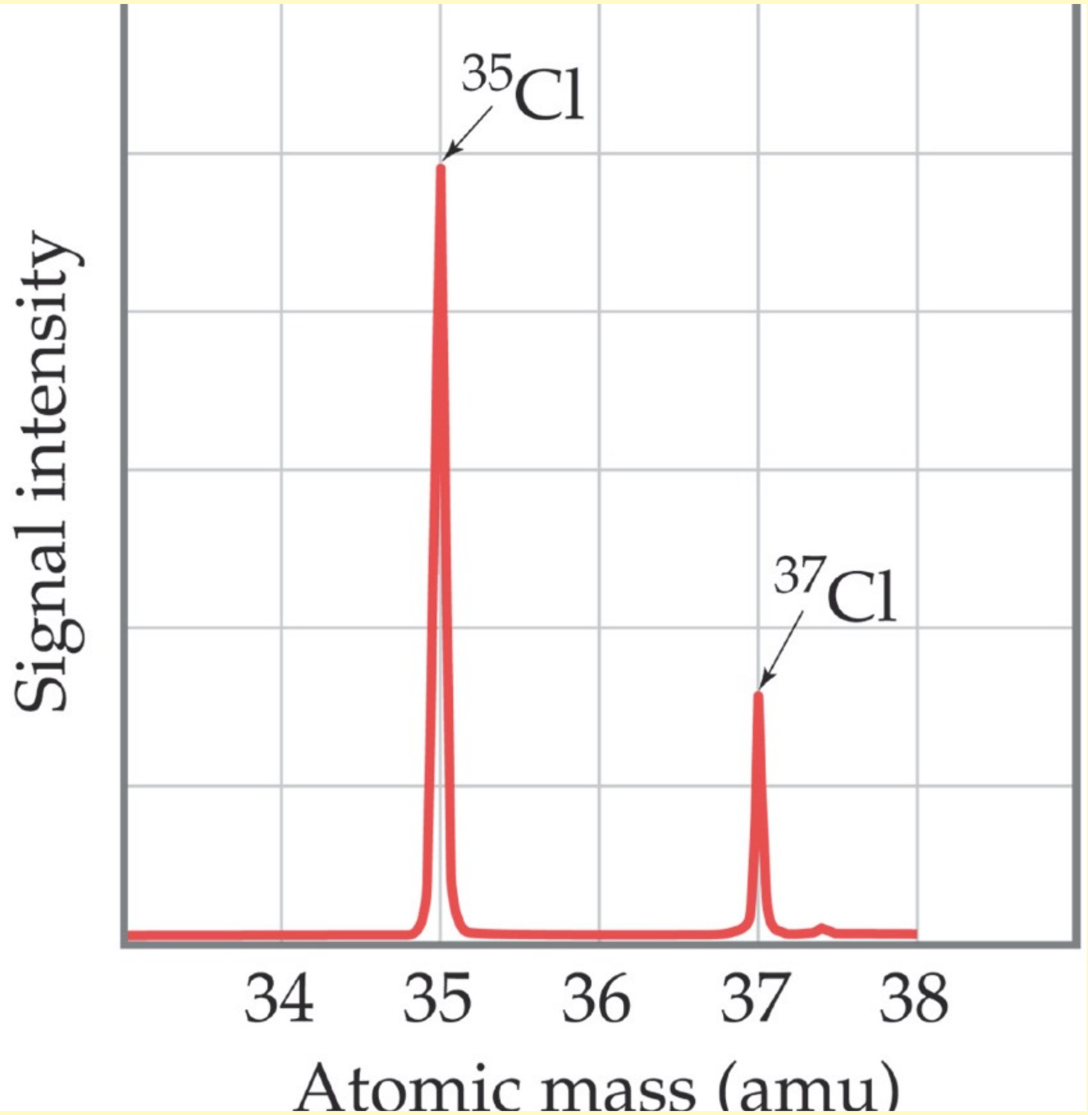
1. 1
2. 2
3. 3
4. 4
5. 5
6. Not enough information to determine



Graphical Readout from a Mass Spectrometer

Chlorine has two isotopes, ^{35}Cl (~75% abundant) and ^{37}Cl (~25% abundant). How many different molar masses are possible for a Cl_2 molecule?

1. 1
2. 2
3. 3
   - 35+35, 35+37, 37+37
   - the percentages in this problem are distractors
4. 4
5. 5
6. Not enough information to determine



Graphical Readout from a Mass Spectrometer

Both chlorine and bromine exist as two naturally-occurring isotopes, distributed as shown to the right. % natural occurrence is based on the distribution of isotopes. Chlorine reacts with bromine to form ClBr. How many different possible molar masses are there for ClBr?

No Calculator

isotope	% natural occurrence
chlorine-35	76%
chlorine-37	24%
bromine-79	51%
bromine-81	49%

Both chlorine and bromine exist as two naturally-occurring isotopes, distributed as shown to the right. % natural occurrence is based on the distribution of isotopes. Chlorine reacts with bromine to form ClBr. How many different possible molar masses are there for ClBr?

Again, the percentages are a distractor

isotope	% natural occurrence
chlorine-35	76%
chlorine-37	24%
bromine-79	51%
bromine-81	49%

Three possible molar masses:

- 35+79=114
- 35+81=116, (37+79=116, repeat)
- 37+81=118

Rubidium has only two naturally occurring isotopes. Estimate the ~% abundance of the heavier isotope

isotope	% abundance
^{85}Rb	
^{87}Rb	?

1. ~75
2. ~50
3. ~25
4. ~10
5. Can not be determined

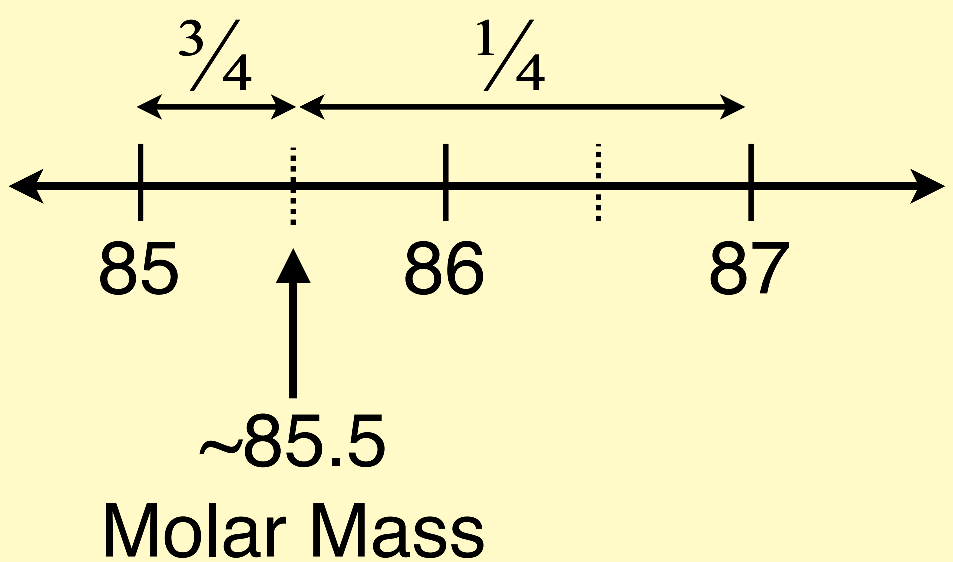
No Calculator

Rubidium has only two naturally occurring isotopes. Estimate the ~% abundance of the heavier isotope

isotope	% abundance
^{85}Rb	
^{87}Rb	?

No Calculator

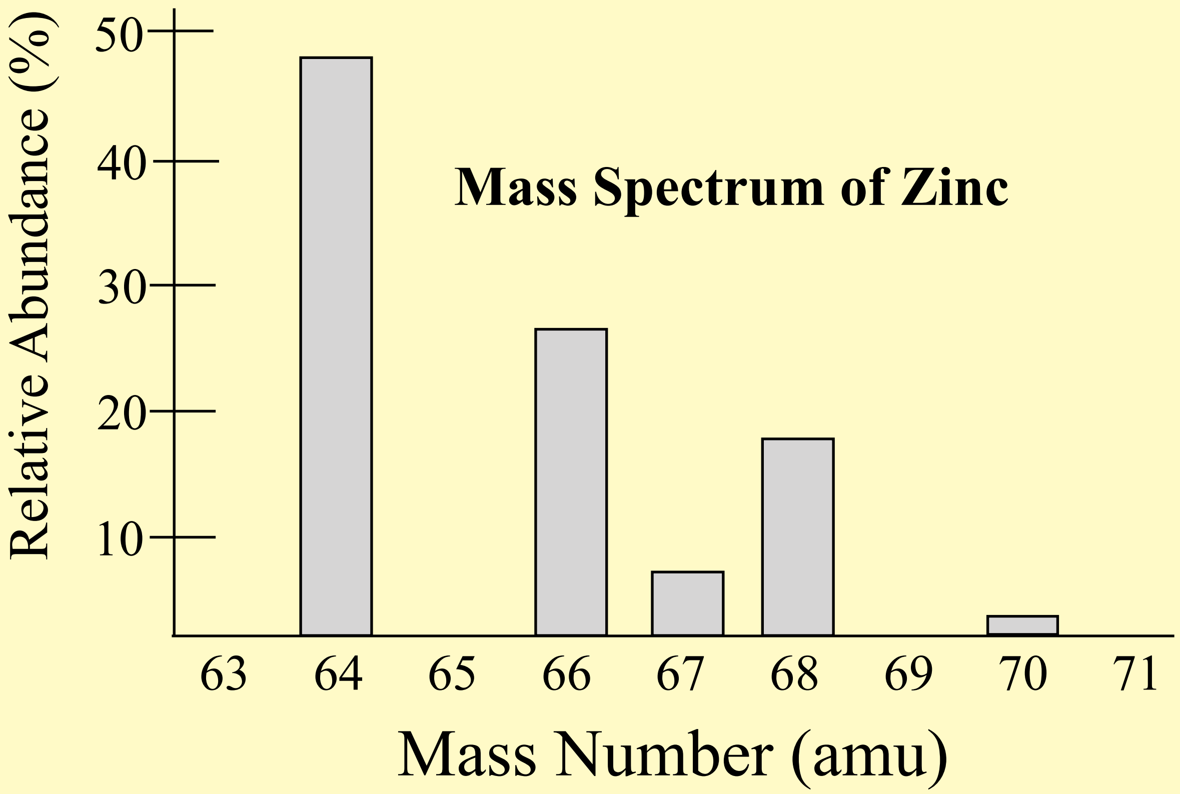
Use a number line to help you estimate



1. ~75
2. ~50
3. ~25
4. ~10
5. Can not be determined

A sample of zinc was run through a mass spectrometer. Which of the following statements can be concluded from the spectrograph shown below.

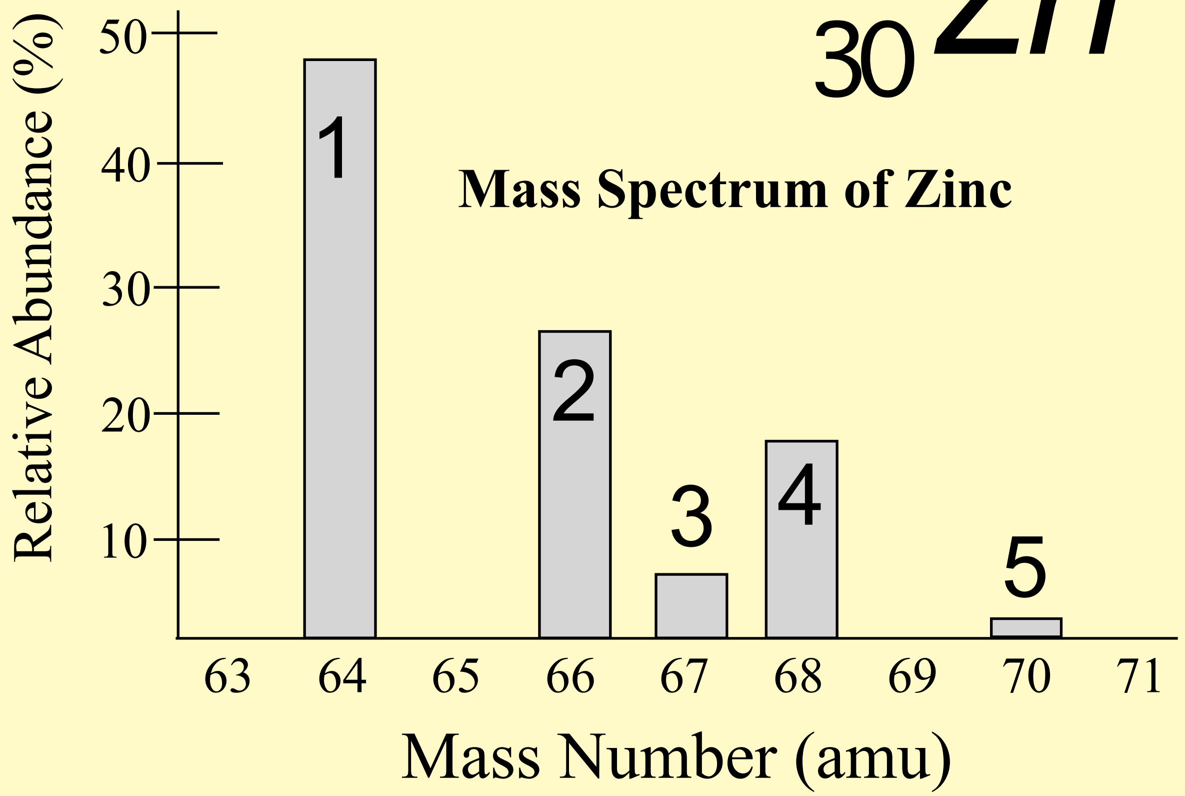
1. Zn has 5 oxidation states.
2. Naturally occurring zinc has an isotope with 35 neutrons.



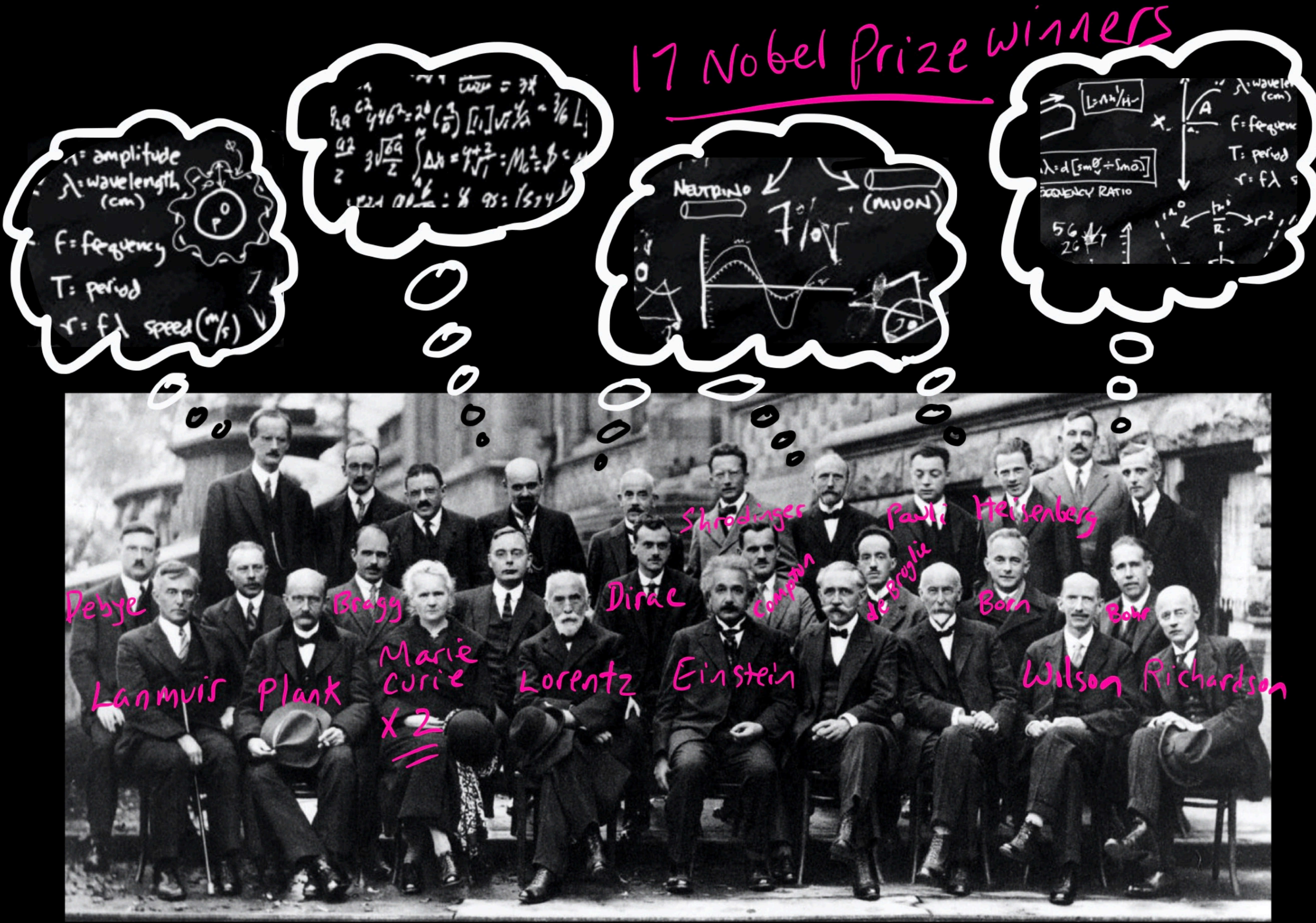
3. There are 5 naturally occurring isotopes.
4. The most common isotope is ^{65}Zn.
5. Zinc most commonly forms Zn^{2+}.

A sample of zinc was run through a mass spectrometer. Which of the following statements can be concluded from the spectrograph shown below.

$_{30}Zn$



1. Zn has 5 oxidation states.
   - not true, this graph gives no information about oxidation states
2. Naturally occurring zinc has an isotope with 35 neutrons.
   - Isotope with 35 neutrons would have a mass of 65, only isotopes ^{64}Zn, ^{66}Zn, ^{67}Zn, ^{68}Zn, ^{70}Zn
3. There are 5 naturally occurring isotopes.
   - The five lines indicate five isotopes
4. The most common isotope is ^{65}Zn.
   - Even though the molar mass rounds to a mass number of 65, in fact there is no isotope ^{65}Zn, and the most common one is ^{64}Zn
5. Zinc most commonly forms Zn^{2+}.
   - While this is true, the mass spectrograph does give any indication of this



Solvay Conference in Brussels in 1927…Lots of Calculations Happening

Calculations

AP may ask you to perform.
Plug and chug. Calculator Gymnastics

ATOMIC STRUCTURE

$$E = h\nu$$
$$c = \lambda\nu$$

E = energy
ν = frequency
λ = wavelength

Planck's constant, $h = 6.626 \times 10^{-34}$ J s
Speed of light, $c = 2.998 \times 10^{8}$ m s^{-1}
Avogadro's number $= 6.022 \times 10^{23}$ mol^{-1}
Electron charge, $e = -1.602 \times 10^{-19}$ coulomb

Find these formulae and constants on your AP Equations and Constants Sheet

The energy of radiant light (EMR) is capable of breaking molecular bonds. The longest wavelength of light with enough energy capable of breaking the bond in a *single* oxygen molecule is 242 nm.

1. Calculate the frequency, ν , in s^{-1} ("cycles"/sec) of this light.
2. Calculate the energy of a photon of the light, units.
3. Calculate the minimum energy, in kJ mol^{-1}, of the oxygen bond.

The energy of radiant light is sometimes capable of breaking molecular bonds. The longest wavelength of light with enough energy capable of breaking the bond in a single oxygen molecule is 242 nm.

$$E = h\nu$$

$$c = \lambda\nu$$

E = energy
ν = frequency
λ = wavelength

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Speed of light, $c = 2.998 \times 10^{8}$ m s^{-1}
Avogadro's number $= 6.022 \times 10^{23}$ mol^{-1}
Electron charge, $e = -1.602 \times 10^{-19}$ coulomb

1. Calculate the frequency, in s^{-1} of this light.

$$\nu = \frac{c}{\lambda}$$

$$\nu = \frac{2.998 \times 10^{8} m/sec}{242 nm} \times \frac{10^{9} nm}{1 m} = 1.24 \times 10^{15} \sec^{-1}$$

2. Calculate the energy of a photon of this light. Be sure and put units on your answer?

The energy of radiant light is sometimes capable of breaking molecular bonds. The longest wavelength of light with enough energy capable of breaking the bond in a single oxygen molecule is 242 nm.

$$E = h\nu$$

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E = energy
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Electron charge, $e = -1.602 \times 10^{-19}$ coulomb

1. The frequency, in s^{-1} of the light.

$$\nu = 1.24 \times 10^{15}\ \text{sec}^{-1}$$

2. Calculate the energy of a photon of this light. Be sure and put units on your answer?

$$E = h\nu$$

$$E = \left(6.63 \times 10^{-34}\ J\,\text{sec}\right) \times \left(1.24 \times 10^{15}\ \text{sec}^{-1}\right) = 8.22 \times 10^{-19}\ J$$

3. Calculate the minimum energy, in kJ mol^{-1}, of the oxygen bond.

The energy of radiant light is sometimes capable of breaking molecular bonds. The longest wavelength of light with enough energy capable of breaking the bond in a single oxygen molecule is 242 nm.

$$E = h\nu$$

$$c = \lambda\nu$$

E = energy
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Planck's constant, $h = 6.626 \times 10^{-34}$ J s
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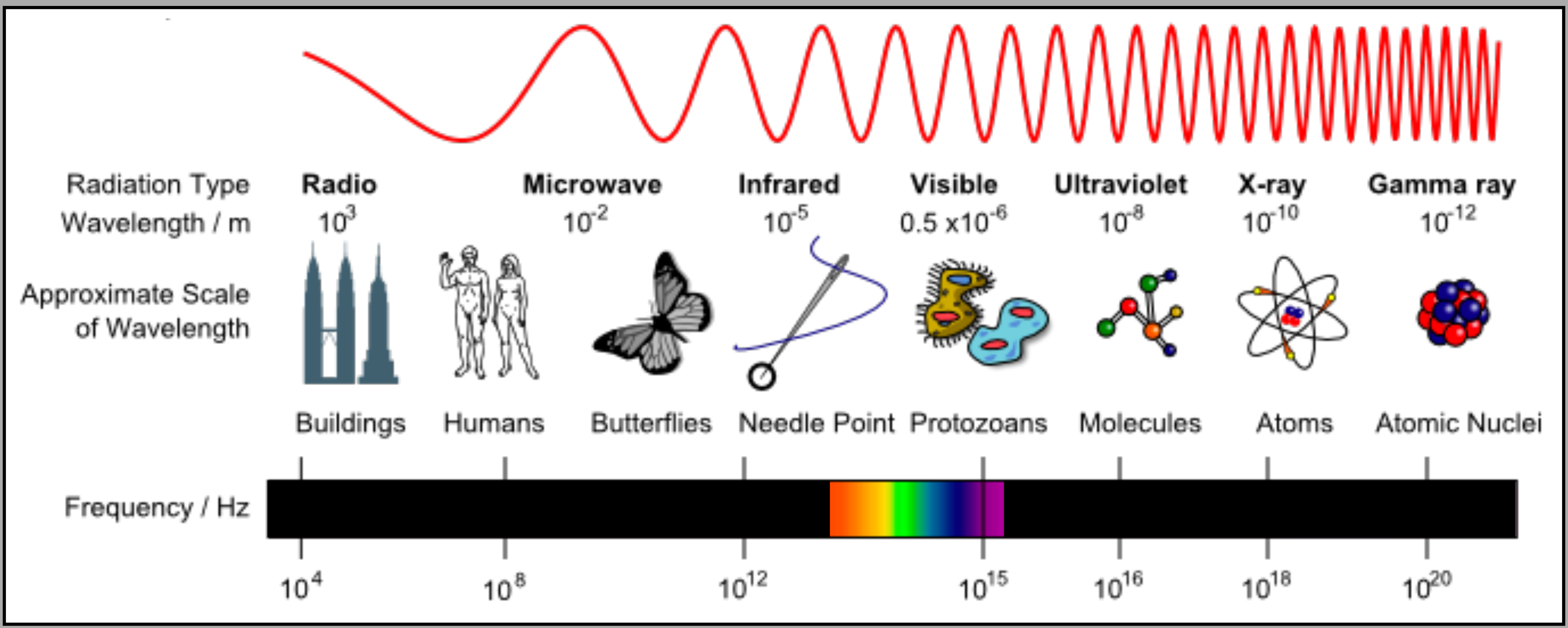
Avogadro's number $= 6.022 \times 10^{23}$ mol^{-1}
Electron charge, $e = -1.602 \times 10^{-19}$ coulomb

1. The frequency, in s^{-1} of this light. $\nu = 1.24 \times 10^{15}\ \text{sec}^{-1}$
2. The energy, in J, of a photon of this light.
   $$E = 8.22 \times 10^{-19}\ J$$
3. The minimum energy, in kJ mol^{-1}, of the oxygen bond.

$$8.22 \times 10^{-19}\ J \times \frac{1kJ}{1000J} \times \frac{6.022 \times 10^{23}\ bonds}{1mol} = 495\ kJmol^{-1}$$

EMR
ElectroMagneticRadiation

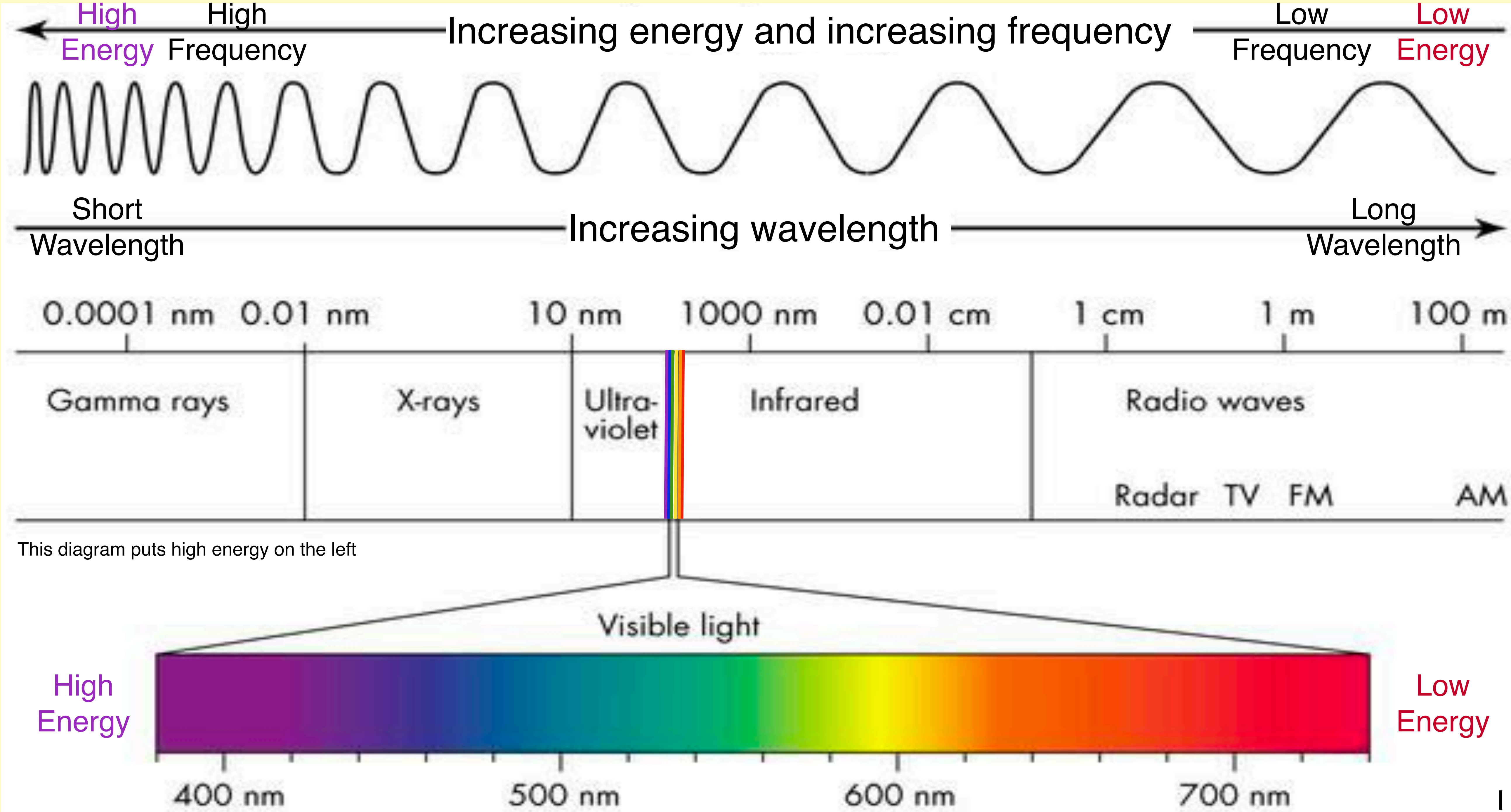
Tools to Inspect Atoms and molecules



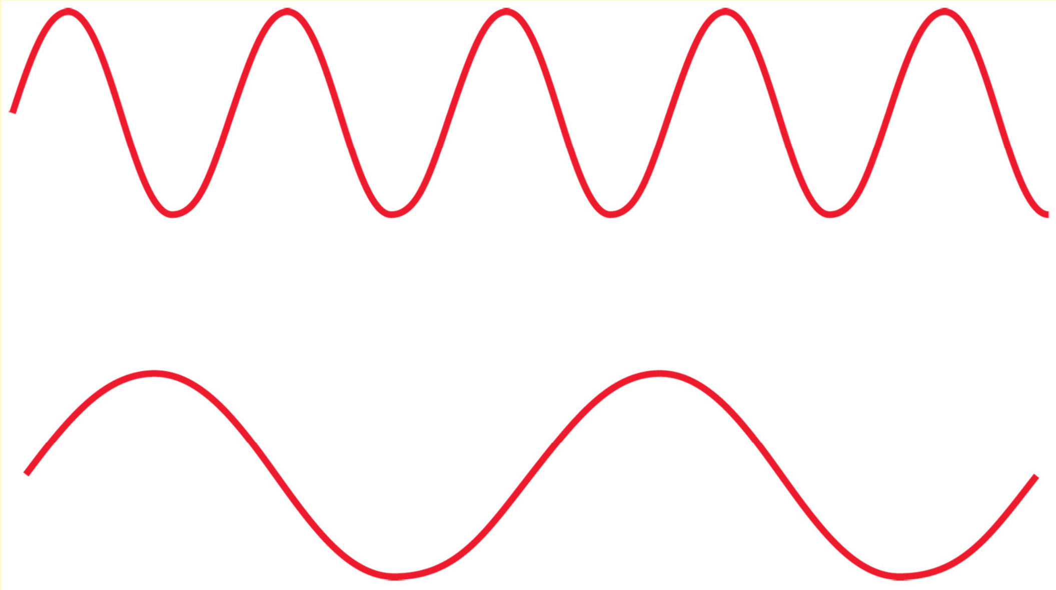
This diagram puts high energy on the right

Appreciate that wavelength (λ) is inversely proportional to both frequency (ν or *f*) and energy (*E*)

$$E = h\nu$$
$$c = \lambda\nu$$



Which statement(s) about the top *EMR* wave compared to the bottom *EMR* wave is (are) true? *Select as many as apply.*

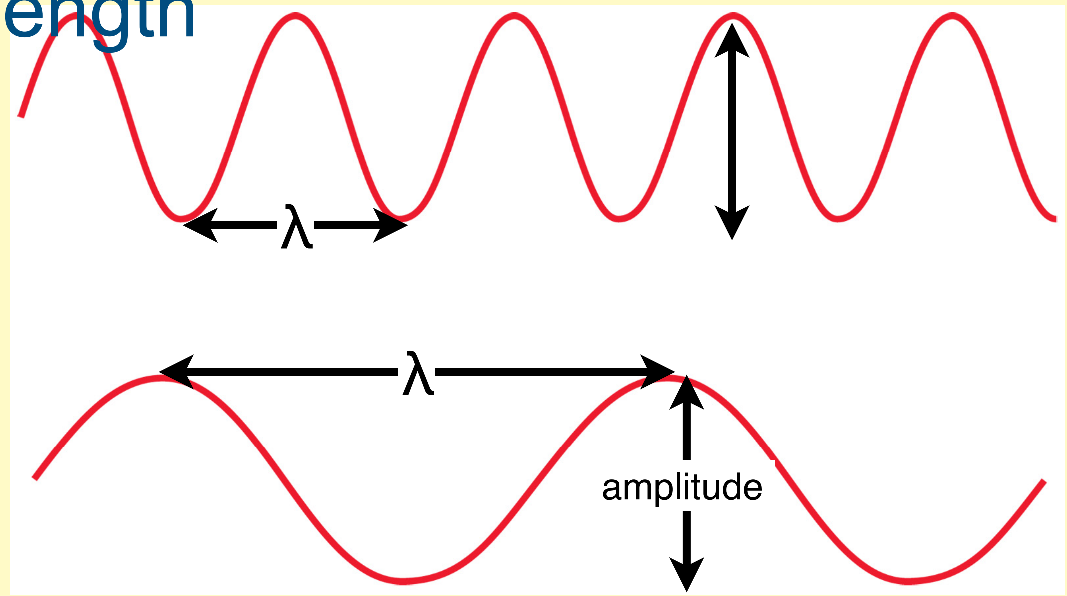


1. Top is higher energy.
2. Top is longer λ (wavelength)
3. Top is higher ν (frequency)
4. Top is the same amplitude.
5. Top could be red light if the lower were be blue.
6. Top moves faster than bottom.
7. Top could be ultraviolet if the bottom were infrared.

Which statement(s) about the top *EMR* wave compared to the bottom *EMR* wave is (are) true?

Select as many as apply.

1. Top is higher energy
   - high energy = high frequency, short wavelength
2. Top is longer λ (wavelength)
3. Top is higher ν (frequency)
   - ν = how frequent does a wavelength pass
4. Top is the same amplitude.
   - AP will not ask about amplitude (how intense the light is, height of wave)
5. Top could be red light (low energy) if the lower were be blue (high energy)
6. Top moves faster than bottom.
   - All EMR travels at the “speed of light” 3x10−8 m/sec
7. Top could be ultraviolet if the bottom were infrared.



Which radiation below has the lowest energy?

1. infrared
2. microwaves
3. ultraviolet
4. visible
5. x-rays

Which radiation below has the highest energy?

1. infrared
2. microwaves

- in order from lowest to highest energy
  - microwave < IR < visible < UV < x-rays

Low Energy	High Energy
Low Frequency	High Frequency
Long Wavelength	Short Wavelength

3. ultraviolet
4. visible
5. x-rays

Which radiation below is used in PES measurements as well as the analysis of bond length and atomic arrangement in crystal structures?

1. infrared
2. microwaves
3. ultraviolet
4. visible
5. x-rays

Which radiation below is used in the analysis of bond length and atomic arrangement in crystal structures, as well as PES measurements?.

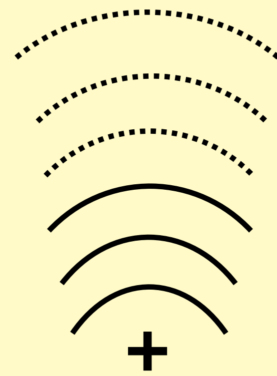
1. infrared
2. microwaves
3. ultraviolet
4. visible
5. x-rays

- x-rays have enough energy to “remove” all of the electrons in an atom as needed for PES.
- x-ray crystallography is used to measure bond lengths and radii of atoms since the nucleus’s show up as more dense on the x-ray image

Which radiation below can be used to analyze the concentration of colored solutions?

1. infrared
2. microwaves
3. visible
4. x-rays

Which radiation below can be used to analyze the concentration of colored solutions?



UV and visible light can cause electron transitions.

1. infrared
2. microwaves
3. visible (from our lab work)

- Color is a manifestation of visible light which is related to electron transitions. We used the spectrophotometer to measure the intensity of color of a blue Cu^{2+} solution, and crystal violet in kinetics ~~(a red $FeSCN^{+}$ solution)~~

4. x-rays (of course can eliminate electrons as in PES)

Which radiation below is used to study the **vibrational motion** of molecules to study the *bending and stretching of bonds*?

1. infrared
2. microwaves
3. ultraviolet
4. visible
5. x-rays

Which radiation below is used to study the vibrational motion of molecules to study the bending and stretching of bonds?

1. infrared
2. microwaves
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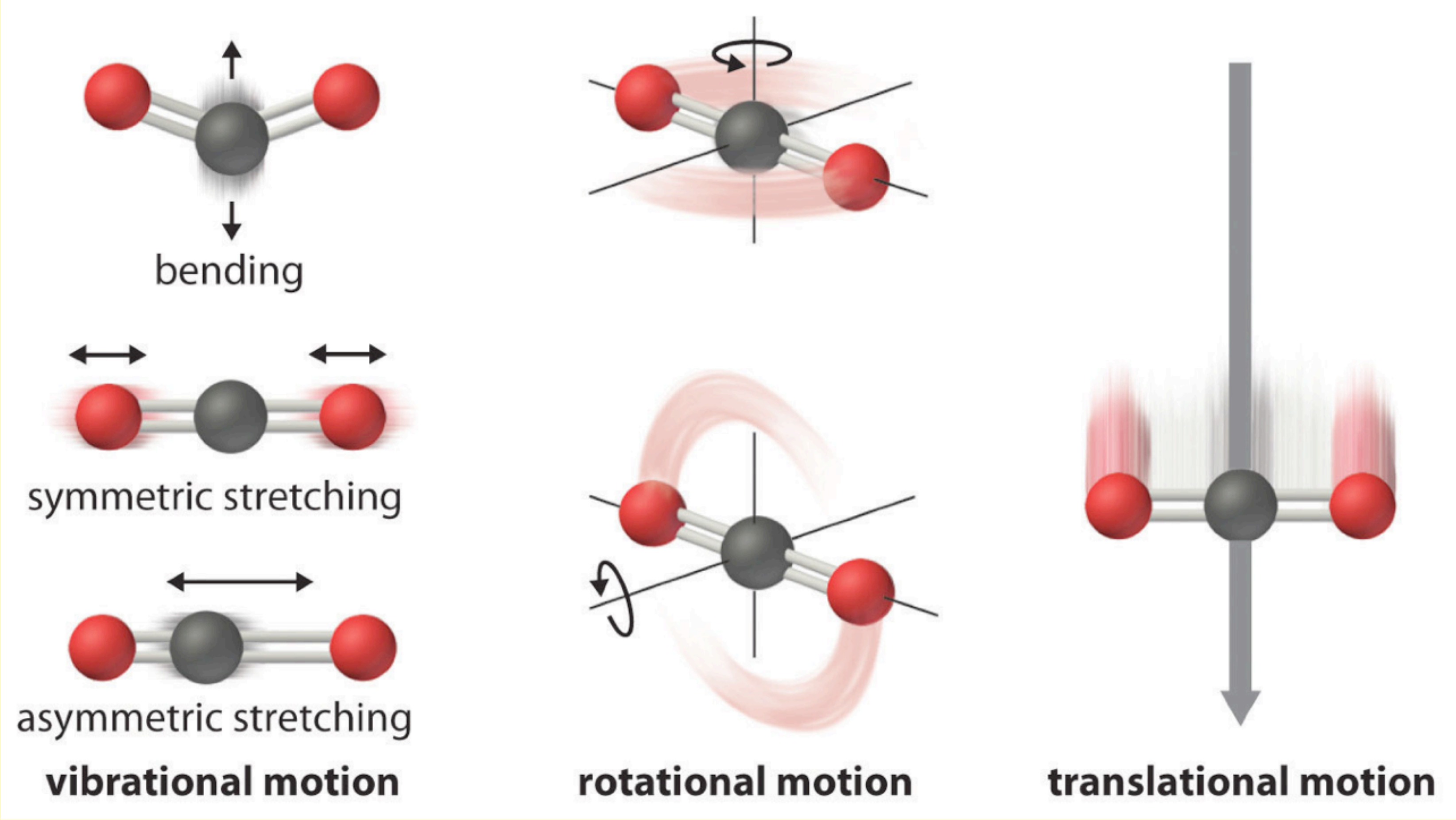

infrared

movement of bonds

microwave

movement of molecules (especially water)

Summary of Radiation as a tool in chemistry

Type	ν / λ	Relative Energy	Effect on atom or molecule	Information Acquired	Name of Spectroscopy
X-Ray	higher ν shorter λ	very high	removes core e^- can exam nuclei arrangement	binding energy, how tightly electrons are held	PES and Xray crystal-ography
UV		high	excites valence electrons	emission spectrum, identity of elements	UV
Visible		medium	excites valence electrons	Beers Law, concentration of colored species	visible
Infrared		low	changes bond vibrations (vibrational motion)	types of bonds / atoms / functional groups in molecules	IR
Micro-wave	lower ν longer λ	quite low	changes movement (rotational & translational motion)	location of H atoms in molecules	microwave

That’s it for now….