C08 Periodic Relationships Among the Elements

Chemistry by Chang and Goldsby Chapter 8: Periodic Relationships ~~8.1 Development of the Periodic Table~~ 8.2 Periodic Classification of the Elements 8.3 Periodic Variation of Physical Properties 8.4 Ionization Energy 8.5 Electron Affinity 8.6 Chemical Properties of Representative Elements	College textbooks include far more material than is required for the AP Chemistry exam. I have posted a condensed version of Chang's Chapter 8 on the classroom webpage that you can use with this summary.

8.1 Development of the Periodic Table

Historical background material – not in the AP syllabus but an interesting story.

<u>8.2 Periodic Classification of the Elements</u>

Ignore the A and B designations and explanations they will not be used in AP Chem.
Instead use the simple 1-18 group numbers on the AP Periodic Table

Valence electrons are the electrons in the **top energy level** and are the major contributor to the chemical reactions of the elements.

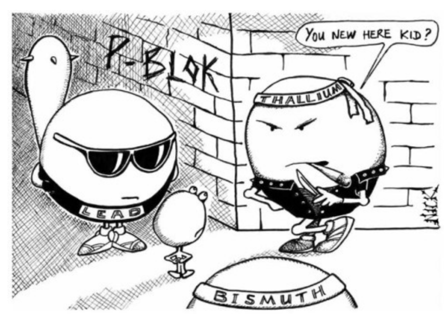
Representative Elements

s-block elements

s^1	alkali metals	1 valence electron
s^2	alkaline earth	2 valence electrons

p-block elements

$s^2 p^1$	boron family	3 valence electrons
$s^2 p^2$	carbon family	4 valence electrons
$s^2 p^3$	nitrogen family	5 valence electrons
$s^2 p^4$	oxygen family	6 valence electrons
$s^2 p^5$	halogens	7 valence electrons
$s^2 p^6$	noble gases	8 valence electrons



Unwittingly, and against his mother's advice, Vince the first-row Transition Metal had been lured far away from home, and now found himself surrounded by heavier elements of the P-Block.

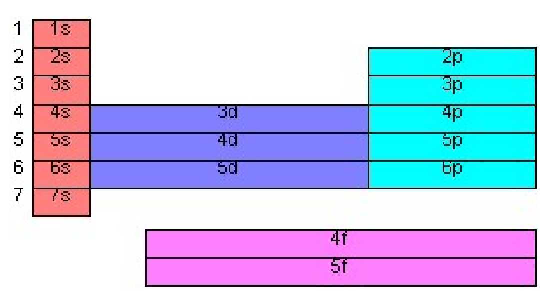
Transition Elements: 2 or 1 valence electrons

Only the upper energy level *s* subshell electrons are valence electrons.

The lower energy level *d* orbital electrons are not valence electrons.

Important! The transition metals will only have 1 or 2 valence electrons.

Be able to determine the electron configuration of an element just by looking at the periodic table.



Electron Configuration of Cations and Anions

Cations of representative elements become isoelectric with noble gas configurations by losing valence electrons to become ns^2np^6 which can produce a lower potential energy system.

e.g.	Na [Ne] $3s^1$	loses $1e^-$ to become:	Na^+ [He] $2s^22p^6$
	Mg [Ne] $3s^2$	loses $2e^-$ to become:	Mg^{2+} [He] $2s^22p^6$
	Al [Ne] $3s^23p^1$	loses $3e^-$ to become:	Al^{3+} [He] $2s^22p^6$

Anions of representative elements become isoelectric with noble gas configurations by gaining electrons to form ns^2np^6

e.g.	F [He] $2s^22p^5$	gains $1e^-$ to become:	F^- [He] $2s^22p^6$
	O [He] $2s^22p^4$	gains $2e^-$ to become:	O^{2-} [He] $2s^22p^6$
	N [He] $2s^22p^3$	gains $3e^-$ to become:	N^{3-} [He] $2s^22p^6$

Isoelectric series Na^+, Mg^{2+}, Al^{3+}, F^-, O^{2-}, N^{3-}, Ne

N.B. You can never just give a abbreviated configuration with just the noble gas.

e.g. Na [Ne] $3s^1$ loses $1e^-$ to become: Na^+ [Ne]

Even though the sodium cation is isoelectric with Ne,

[Ne] would not receive the point on an AP Chemistry FRQ

Noble Gases do not form ions.

Have filled *p* subshells ns^2np^6 (exception He is $1s^2$). The noble gases do not form ions.

Transition metals form Cations

Important!

Transition elements lose their top energy level *s* subshell electrons before their lower *d* subshell electrons.

Only after the upper principal energy level *s* electrons are removed can you remove the lower principal energy level *d* subshell electrons.

Using the misleading "*aufbau*" electronic structure with the *d* subshell after the *s* will lead to errors in ionizing the transition metals by incorrectly removing *d* subshell electrons before the *s*.

Don't make that mistake!
Always write electron configurations in energy level order.

e.g. Mn [Ar] $3d^5$ $4s^2$ ionizes to Mn^{2+} [Ar] $3d^5$

Most transition metals form more than one cation and are rarely isoelectric with noble gases.

The stability of the ns^2np^6 structure is because this valence electron configuration produces the lowest potential energy arrangement.

When explaining why the ns^2np^6 structure is stable you must use the **low potential energy** explanation.

DO NOT use statements like the "atoms are happy as octets" or stable octets.

These statements will always lose a point on the AP FRQ.

<u>8.3 Periodic Variation in Physical Properties</u>

It is as important to understand the reasons behind the trends in atomic size and ionization energy, as it is to memorize the trends themselves.

PERIODIC TABLE OF THE ELEMENTS

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 47.90	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 65.39	31 Ga 69.72	32 Ge 72.59	33 As 74.92	34 Se 78.96	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 95.94	43 Tc (98)	44 Ru 101.1	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 121.75	52 Te 127.60	53 I 126.91	54 Xe 131.29
55 Cs 132.91	56 Ba 137.33	57 *La 138.91	58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm (145)	62 Sm 150.4	63 Eu 151.97	64 Gd 157.25	65 Tb 158.93	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.04	71 Lu 174.97	72 Hf 178.49	73 Ta 180.95	74 W 183.85	75 Re 186.21	76 Os 190.2	77 Ir 192.2	78 Pt 195.08	79 Au 196.97	80 Hg 200.59	81 Tl 204.38	82 Pb 207.2	83 Bi 208.98	84 Po (209)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra 226.02	89 †Ac 227.03	90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np (237)	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (262)	104 Rf (261)	105 Db (262)	106 Sg (266)	107 Bh (264)	108 Hs (277)	109 Mt (268)	110 Ds (271)	111 Rg (272)							

The periodic table lists the elements in atomic number order. A new row is started when the size of the atoms of an element undergoes a dramatic increase in size because of the formation of a new principal energy level.

The next principal energy level is farther away from the nucleus resulting in lower Coulombic attractions to the nucleus.
The atoms of each principal energy level are larger than the atoms in the principal energy level before it.

The inner transition metals and transition metals show very small changes in properties between members of their groups. The AP Chem curriculum focuses on the representative elements where the changes are significant.

While you can use the position of elements in the periodic table to predict relative atom sizes, **you cannot use the periodic table position as a justification or an explanation in answering FRQ's.**

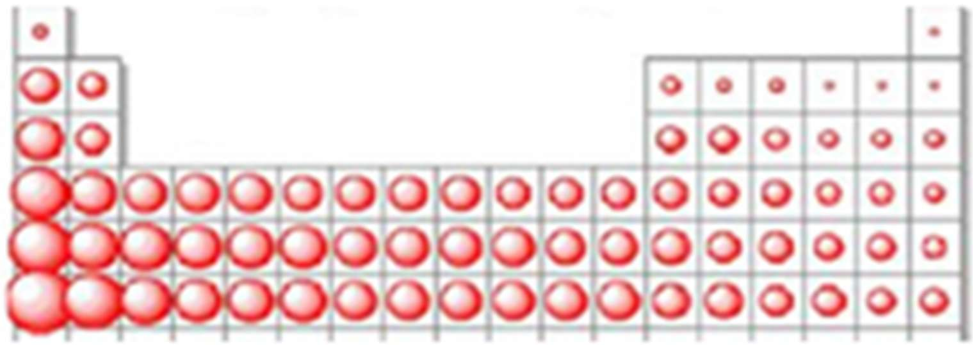
To explain relative atomic sizes, you must use Coulombic forces or differences in energy levels for explanations.

(1) Protons attract electrons making the atoms smaller.

(2) The further the electrons are from the nucleus, the smaller the attractions to the nucleus.

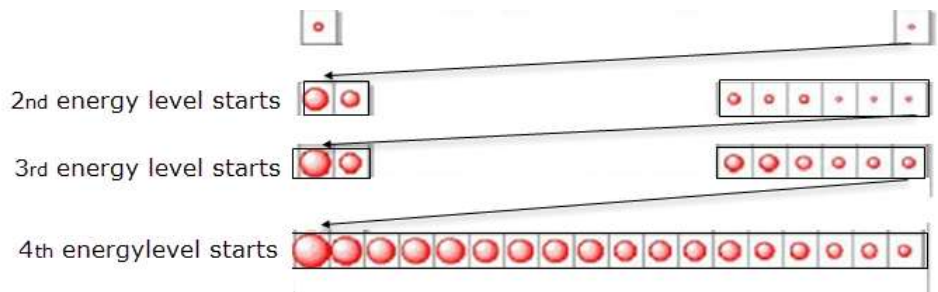
(3) There is repulsion between electrons in shells, subshells, and orbitals.

This diagram shows the relative sizes of atoms based on the positions in the periodic table:



Each period in the periodic table starts a new principal energy level, *n*.

Atoms dramatically increase in size at the start of each new period↓ because electrons are in a shell that is farther from the nucleus.



Atomic size dramatically increases with the start of each new principal energy level↓

because the new principal energy level is further from the nucleus.

Atoms get smaller across a period →

because of the increasing number of protons increases the attraction to the electrons.

Shielding

Do not use shielding explanations when dealing with atomic size.

Z = atomic charge of the atom. Atomic charge attracts electrons making a smaller atom.

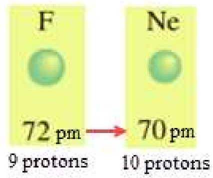
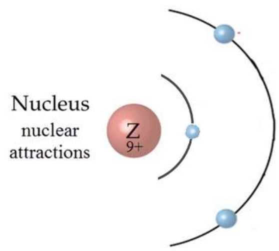
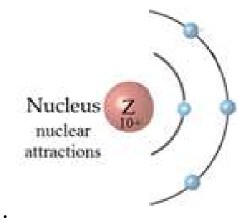
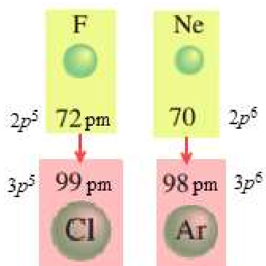
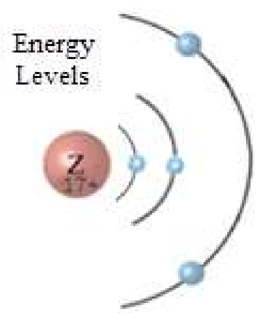
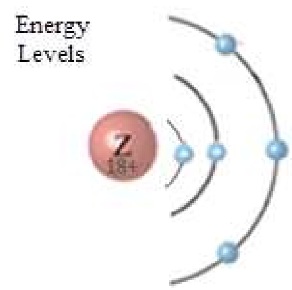
σ (sigma) = "shielding" factor is the repulsions of outer electrons from one shell and subshell to the next higher shell or subshell. The term shield is misleading because electrons do not block other electrons attraction to protons.

Effective Nuclear Charge $Z_{effective} = Z - \sigma$… net attraction to the nucleus after shielding (electron repulsion) is taken into account.
However, shielding and $Z_{effective}$ are complex and easily misused so do not use this in explanations. AP readers say that the misuse of shielding explanations result loss of points.

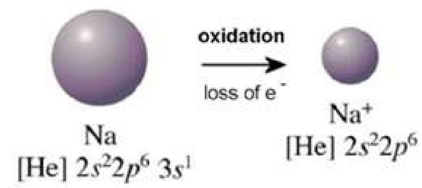
Bottom line: For FRQ's: Do not use the term "shielding" in FRQ explanations or justifications.

Atomic Radius and Ionization Energy

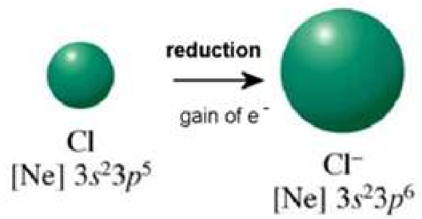
Ionization energy is the energy required to remove electrons from an atom.

 **Atomic sizes decreases** The increase in number of protons increases the attraction of all the electrons to the nucleus making the atom smaller. **Ionization Energy increases** The increase in number of protons make makes the atom more difficult to ionize. The 10 proton argon atom is more difficult to ionize than the 9 proton nitrogen atom.	Fluorine 9 protons 72 picometers $1s^2)\ 2s^2 2p^5$  **Ionization Energy = 1,680 kJ/mol**	Neon 10 protons 70 picometers $1s^2)\ 2s^2 2p^6$  **Ionization Energy= 2,080 kJ/mol The increase is because of extra protons attracting electrons in the same energy level.**
 **Atomic sizes increase as you move down the periodic table.** Cl and Ar atoms are larger than F and Ne atoms because the valence electrons of Cl and F are in the 3rd principal energy level which is further from nucleus than the 2nd period electrons of F and Ne. **Ionization Energy decreases** Ionization energies of 3rd period Cl and Ar atoms is lower than ionization energy of 2nd period F and Ne atoms because the 3rd period energy level electrons are further from the nucleus than the 2nd energy level electrons.	Chlorine 17 protons 99 picometers $1s^2)\ 2s^2 2p^6)\ \mathbf{3s^2 3p^5}$  **Ionization Energy = 1,251 kJ/mol**	Argon 18 protons 98 picometers $1s^2)\ 2s^2 2p^6)\ \mathbf{3s^2 3p^6}$  **Ionization Energy = 1,520 kJ/mol The increase is because of extra protons attracting electrons in the same energy level.**

Ionic Radius

<table>
<tr><td>
</td><td>Oxidation is the loss of electrons, e^-. (mnemonic LEO)

Loss of electrons, the formation of cations, results in ions that are smaller in size than the original atoms because of the removal of outer shell electrons.

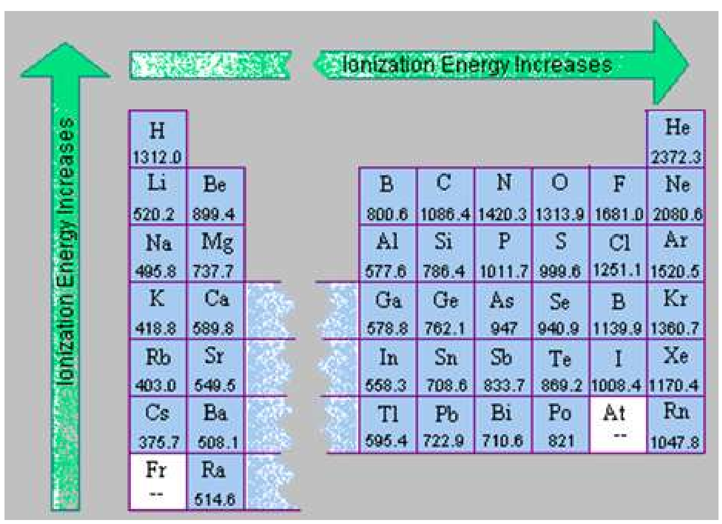
Metals form cations by losing electrons in the upper energy level. Since the atom has lost electrons from the outer energy level the remaining electrons are in a lower energy level and thus the cation is smaller than the atom</td></tr>
<tr><td>
</td><td>Reduction is the gain of electrons e^-. (mnemonic GER)

Gain of electrons, the formation of anions will result in ions that are larger than the original atoms because of the mutual repulsion of extra electrons within an energy level.</td></tr>
</table>

8.4 Ionization Energy

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

Trends in ionization energy are important. The smaller the atom, the greater the attraction of the electrons to the nucleus, and the higher the energy required to remove the electron.



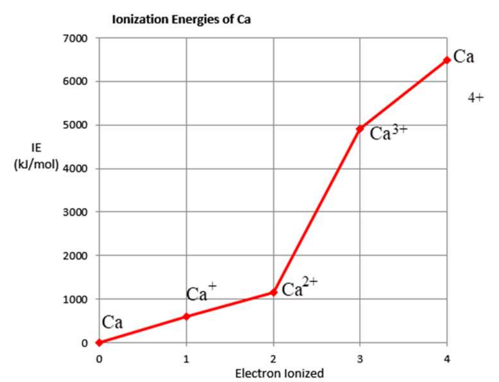
http://chemed.chem.purdue.edu/genchem/topicreview

$X(g) \rightarrow X^+(g) + 1e^-$ **First ionization energy ΔH = + always endothermic (breaking a bond)**
First ionization energies are typically hundreds of kJ/mol

$X^+(g) \rightarrow X^{2+}(g) + 1e^-$ Second ionization energy $\Delta H_{2nd} > \Delta H_{1st}$
Second ionization energies are typically thousands of kJ/mol

$$\Delta H_{\text{ionize 1st}} < \Delta H_{\text{ionize 2nd}} < \Delta H_{\text{ionize 3rd}} < \Delta H_{\text{ionize 4th}}$$

Ionization energies of ions dramatically increase once the top energy level electron has been removed, and the atom becomes smaller.

Atom or Ion to be ionized	Ionization Energy to remove electron from species	
Ca [Ar] $4s^2$	1st IE 590 kJ/mol $Ca \rightarrow Ca^+ + e^-$	
Ca^+ [Ar] $4s^1$	2nd IE 1,140 kJ/mol $Ca^+ \rightarrow Ca^{2+} + e^-$	
Ca^{2+} [Ar] Ca^{2+} [Ne] $3s^1 3p^6$	**3rd IE 4,900 kJ/mol** $Ca^{2+} \rightarrow Ca^{3+} + e^-$	
Ca^{3+} [Ne] $3s^1 3p^5$	4th IE = 6,500 kJ/mol $Ca^{3+} \rightarrow Ca^{4+} + e^-$	

Successive Ionization energies and ionization energy graphs can be used to identify elements.

Note the sudden jump in the ionization energy for the 3rd electron in this table and in the slope of graph.

Removing the 3rd electron is far more difficult than removing the 2nd electron. This is because the 3rd electron is in a lower principal energy level closer to the nucleus.

The structure of the Ca atom before each ionization:

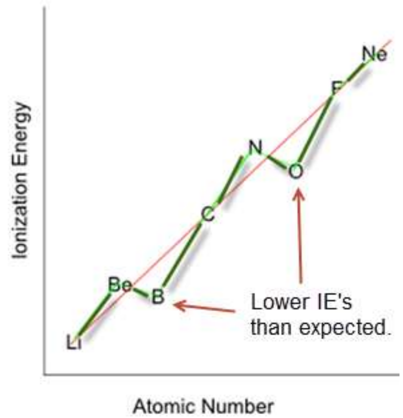
Ca	IE:	$1s^2)\ 2s^2 2p^6)\ 3s^2 3p^6)\ 4s^2$	590 kJ/mol to remove the 4*s* electron
Ca^+	IE:	$1s^2)\ 2s^2 2p^6)\ 3s^2 3p^6)\ 4s^1$	1,140 kJ/mol to remove the 4*s* electron
Ca^{2+}	IE:	$1s^2)\ 2s^2 2p^6)\ 3s^2 3p^6$	**4,900 kJ/mol to remove the 3*p* electron**

The $4s^2$ electrons are within the outermost energy level and are relatively easy to ionize.
The third electron is in the $3p^6$, a lower energy level, closer to the atom and much more difficult to remove.
When dealing with ionization problems write out the electron configuration to see the shells.

Bottom line:

In ionization sequences the IE gradually increases until all the valence electrons are removed.
Then there is a dramatic increase in IE as the electron in the next lowest energy level must be ejected.
In the case of Ca there are two valence electrons (590 kJ and 1,140 kJ) but the third IE (4,900 kJ) is from the lowered energy level.

Now for two anomalies in Ionization Energy trends:



First Ionization Energies of Period 2 Elements (kJ/mol)

Trend in First Ionization Energies →

Increasing IE is caused by the increase in number of protons with the **exception of boron and oxygen**

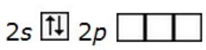
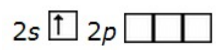
Except for Boron and Oxygen

Li	Be	**B**	C	N	**O**	F	Ne
520	899	**801?**	1086	1400	**1314?**	1680	2080

The Be to B sequence and the N to O sequence have "unexpected" drops in IE.

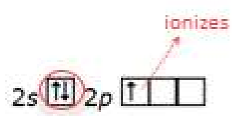
Be	**B**
899	**801?**

Why with more protons is the boron atom easier to ionize than beryllium?
To explain the anomaly, go to the subshell orbital diagrams.

Be  Be → $Be^+ + e^-$ 899 kJ 

The two electrons in the *s* subshell do repel one another so it is easier to ionize one of them.

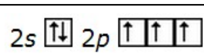
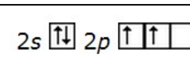
B  B → $B^+ + e^-$ 801 kJ 



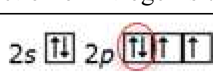
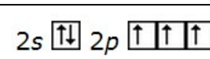
The repulsion of the 2*s* electrons against the electron in the 2*p* subshell makes it easier to remove the 2*p* electron.

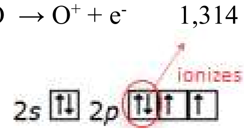
N	**O**
1400	**1314?**

Why with more protons does oxygen have a lower ionization energy than nitrogen?
To explain the anomaly, go to the subshell orbital diagrams.

N  N → $N^+ + e^-$ 1,400 kJ 

The attraction of nitrogen's 7 protons on the electron to be ionized acts to prevent ionization.

O  O → $O^+ + e^-$ 1,314 kJ 



The repulsion of the paired electrons within the *p* orbital of oxygen makes it easier to ionize the electron in the orbial.

From the AP Chemistry Curriculum guide.

"If given an exception on the AP Exam, students will be responsible for providing possible reasons for the exceptions based on theory."

If an atom has a lower than expected ionization energy (lower attractions)

Reason: Mutual electron repulsion resulting from either.

(1) Repulsion by lower subshell energy level electrons.

(2) Paired electrons within an orbital repelling one another.

8.5 Electron Affinity

Opposite of Ionization Energy.

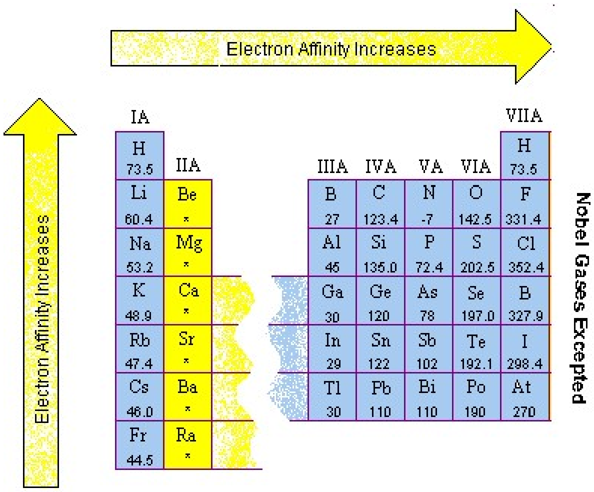
Outside electrons can be attracted to the protons in neutral atoms. This is usually an exothermic reaction.

X(g) + e^- $\rightarrow$ X^-(g) usually exothermic, ΔH = -

The trend has smaller atoms with a greater attraction to an electron.

The trend has exceptions but generally metals have a very low electron affinity (in some cases zero) and nonmetals as you would expect have a higher electron affinity.

Noble gases usually are not included in electron affinity charts because each noble gas's energy level is already filled and any extra electrons would have to be placed in a higher energy level.



8.6 Variation in Chemical Properties

Families (vertical columns) of the periodic table have similar chemical reactions.

You know quite a lot about the elements already but memorize the colors and states of the halogens.

Metal oxides react with water to make bases.

e.g. $Na_2O + H_2O \rightarrow 2\ NaOH$

Nonmetal oxides react with water to make acids.

e.g. $SO_2 + H_2O \rightarrow H_2SO_3$ (sulfurous acid)

Bottom line: Metal oxides are basic and nonmetal oxides are acidic

Aluminum[1] is an amphoteric oxide and reacts with both acids and bases.

You won't have to memorize the reactions. Just realize that metal oxides tend to produce hydroxides which are bases, and nonmetal oxides produce acids.

Environmental examples of nonmetal oxides reacting with water are worthy of note:

Cars with internal combustion engines produce nonmetal oxides in the form of nitrogen oxides.

The high pressures and temperatures in the combustion chambers are good conditions for the combination of nitrogen and oxygen. These nitrogen oxides react with water to form nitrous acid and nitric acid which are pollutants since nonmetal oxides react with water to form acids. The catalytic converter allows the nitrogen oxides to decompose back into nitrogen and oxygen so that no nitrogen oxides are released.

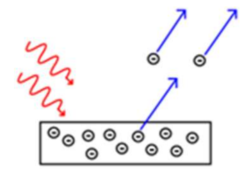
Coal often contains sulfur impurities which oxidize into sulfur oxides. Sulfur dioxide and sulfur trioxide form sulfurous acid which changes to sulfuric acid in the atmosphere creating acid rain.

Coal burning plants scrub the combustion gases with water to form the acid before it reaches the atmosphere. Then acid wash is mixed with metal oxides to neutralize the acid that is formed. The amount of waste sludge from this process is considerable.

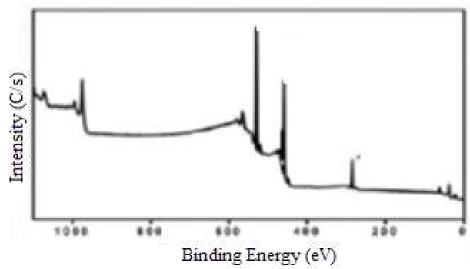
Increased levels of atmospheric carbon dioxide (a nonmetal oxide) dissolve in water to make the oceans slightly more acidic.

[1] While you may think of aluminum as an excellent metal, it does have some nonmetal properties. Note that it is close to the metalloids and its chemistry reflects that.

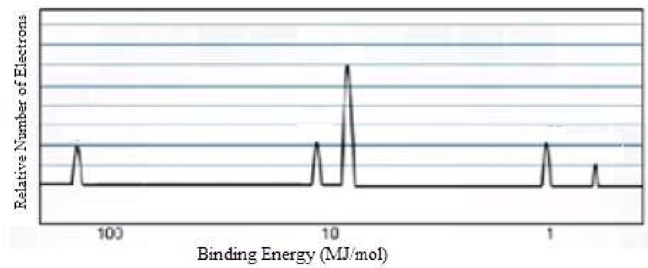
Photoelectron Spectroscopy



This topic is not covered in most college textbooks but is part of the AP Chem curriculum.
PES is used to analyze the electrons in atoms. Very high energy photons (x-rays) are used to bombard a substance. When the energy of the photons exceeds the attraction of the electrons to the nucleus, the electrons are ejected, and the number and energies of the electrons are recorded producing a PES graph. Most of the PES graphs used on the AP Chemistry Exam are simplified. Real PES data is more complex.

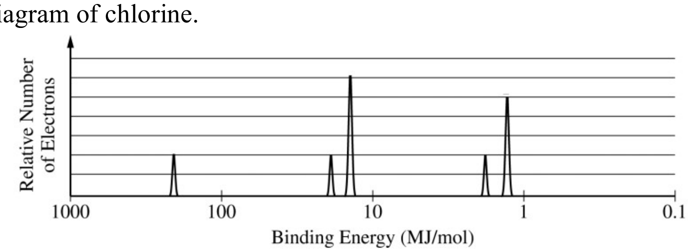


PES Data

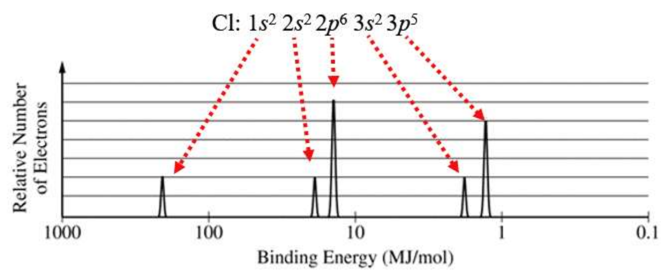


Simplified PES Graph

AP Chemistry PES data graphs show the number of electrons in subshells on the y-axis and the binding energy (attraction of the electrons to the nucleus) in the subshell on the x-axis.
This is the simplified PES diagram of chlorine.



The electron configuration (or a portion of it) of an element is graphically represented as peaks in a PES graph.
For example, notice how the electron configuration of chlorine is matched to its PES Spectrum.



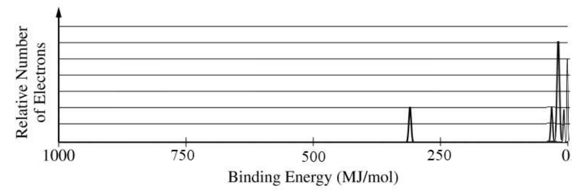
The *y*-axis height of the peak indicates the number of electrons in a subshell and the *x*-axis location of the subshells indicates the attraction (binding energy) of the electrons in the subshell to the protons in the nucleus of the atom.
There are three things to note on the *x*-axis of a PES diagram. The binding energy values are in megajoules per mol of electrons or joules per electron**.**

Note the x axis values are the reverse of what you would expect.

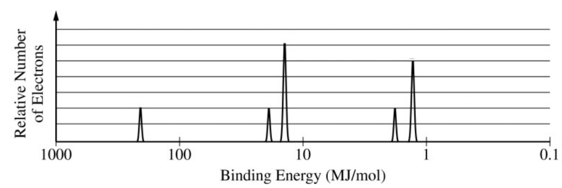
The x axis binding energy values decrease as you go to the right.

Therefore, the subshells on the right-hand side of the graph have lower attractions to the nucleus than those to the left.
Finally, the decreasing binding energies scale is usually shown as a logarithmic scale in which each marked interval is one-tenth of the next (1000, 100, 10, …).

The 1st energy level electrons of chlorine are so close to the nucleus that their binding energies are in the hundreds of megajoules. The 3rd energy level electrons are so much further from the nucleus that their binding energies are just barely over 1 megajoule. In a linear graph, the 2nd and 3rd energy levels are jammed together. The logarithmic scale, on the other hand, allows the three energy levels to be more easily discerned.

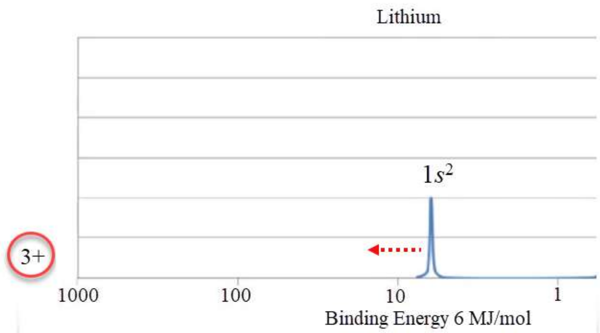


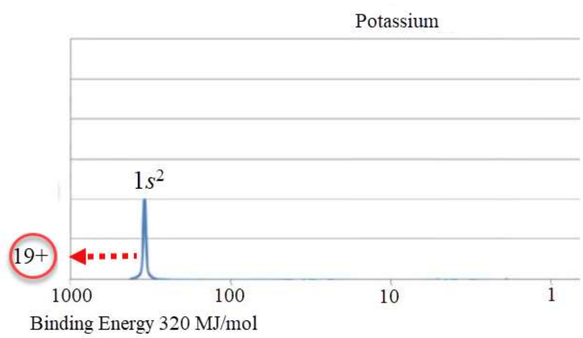
Linear scale



Logarithmic scale

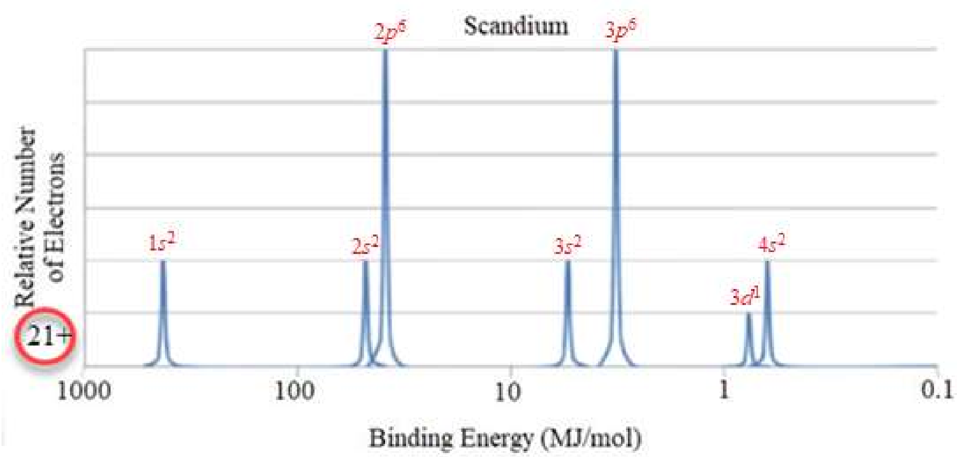
PES diagrams are good at showing the degree of attraction of electrons to the protons in the nucleus. As the number of protons increases, the electron attraction increases. For example, here are the $1s^2$ subshells of Li and K showing the $1s^2$ subshell's attraction to the nucleus.





Lithium, with only 3 protons, has a much weaker attraction to its $1s^2$ electrons than potassium which has 19 protons that pull in the $1s^2$ electrons with greater force. When thinking of the PES diagrams, imagine the attraction of the protons in the nucleus as being to the left of the graph pulling in on the electrons represented in the peaks of the PES diagram.

PES diagrams are also used to show the number of electrons in each subshell.
PES diagrams also show each of the subshells, their attraction to the nucleus, and the number of electrons within each subshell. For example, the electron configuration of scandium, a transition metal, is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^1\ 4s^2$



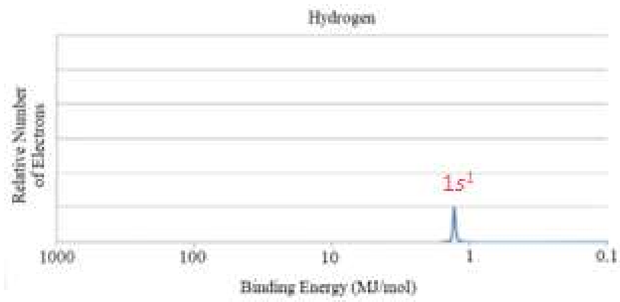
The above diagram shows that there is only one electron in the 3*d* subshell and two electrons in the 4*s* subshell that the $4s^2$ valence electrons with the lower binding energies would be the first to ionize.

Carefully look over the atomic radius and PES diagram changes in the periods → and families ↓ in the periodic table on the next page.

Two of the major themes of this chapter are atomic size and ionization energy trends. The first ionization energy of an atom will be very close to the binding energy of the electrons in the outermost energy level of the atom. Go through each sequence of the PES spectra to correlate the PES spectra with these concepts.

First, see how both atomic size and binding energies change as you move down a family.

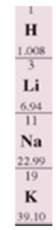
Idealized Photoelectron Spectrum	Electron Configurations and Binding Energy in MJ/mol

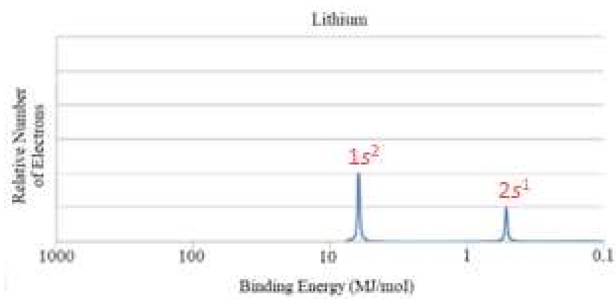


H 1 proton, 1 energy level

atomic size 37 pm

$1s^1$					
1.3					

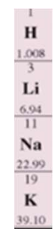




Li 3 protons, 2 energy levels

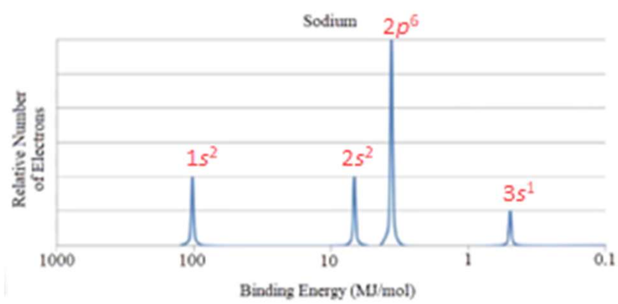
atomic size 152 pm

$1s^2$	$2s^1$				
6.26	0.52				



The 1st energy level of Li has a higher Binding Energy than the 1st energy level of H because of Li's attraction to 3 protons.

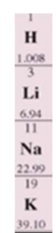
The 2nd energy level of Li has a lower Binding Energy than H because Li's higher energy levels are at a greater distance from the nucleus.



Na 11 protons, 3 energy levels.

atomic size 186 pm

$1s^2$	$2s^2$	$2p^6$	$3s^1$		
104	6.8	3.7	0.50		



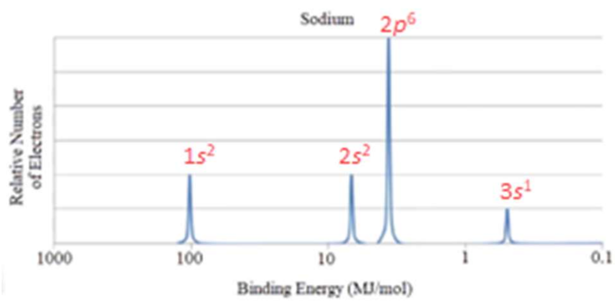
The 1st energy level of Na has a much higher Binding Energy than the 1st energy level of Li because of the attraction to Na's 11 protons.

The 3rd energy level of Na has a lower Binding Energy because of its distance from the nucleus.

Next, look at the changes in atomic size and binding energies as you move across a period.

11	12	13	14	15	16	17	18
Na	**Mg**	**Al**	**Si**	**P**	**S**	**Cl**	**Ar**
22.99	24.30	26.98	28.09	30.97	32.06	35.45	39.95

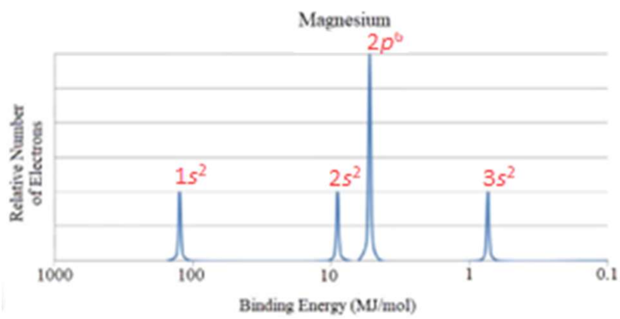
Idealized Photoelectron Spectrum	Electron Configuration and Binding Energy in MJ/mol



Na 11 protons 

atomic radius 186 pm

$1s^2$	$2s^1$	$2p^6$	$3s^1$
104			0.50

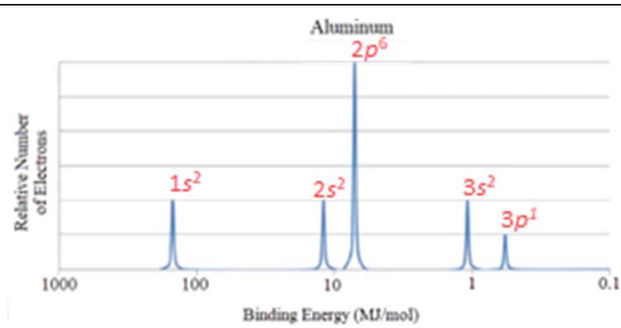


Mg 12 protons 

atomic radius 160 pm

$1s^2$	$2s^1$	$2p^6$	$3s^2$
126			0.74

Every subshell is moved left ← with higher Binding Energies because of the extra proton in the nucleus. Note the $3s^2$ peak is higher with the extra electron.



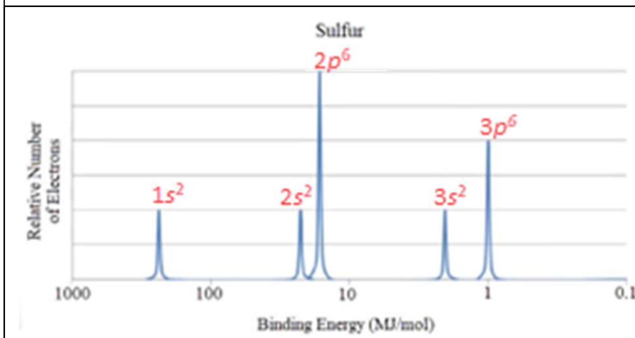
Al 13 protons 

atomic radius 143 pm

$1s^2$	$2s^1$	$2p^6$	$3s^2$	$3p^1$
151			1.1	0.58

The $3p$ subshell is a new subshell.

All the lower subshells are moved ← with higher Binding Energies because of the extra proton.

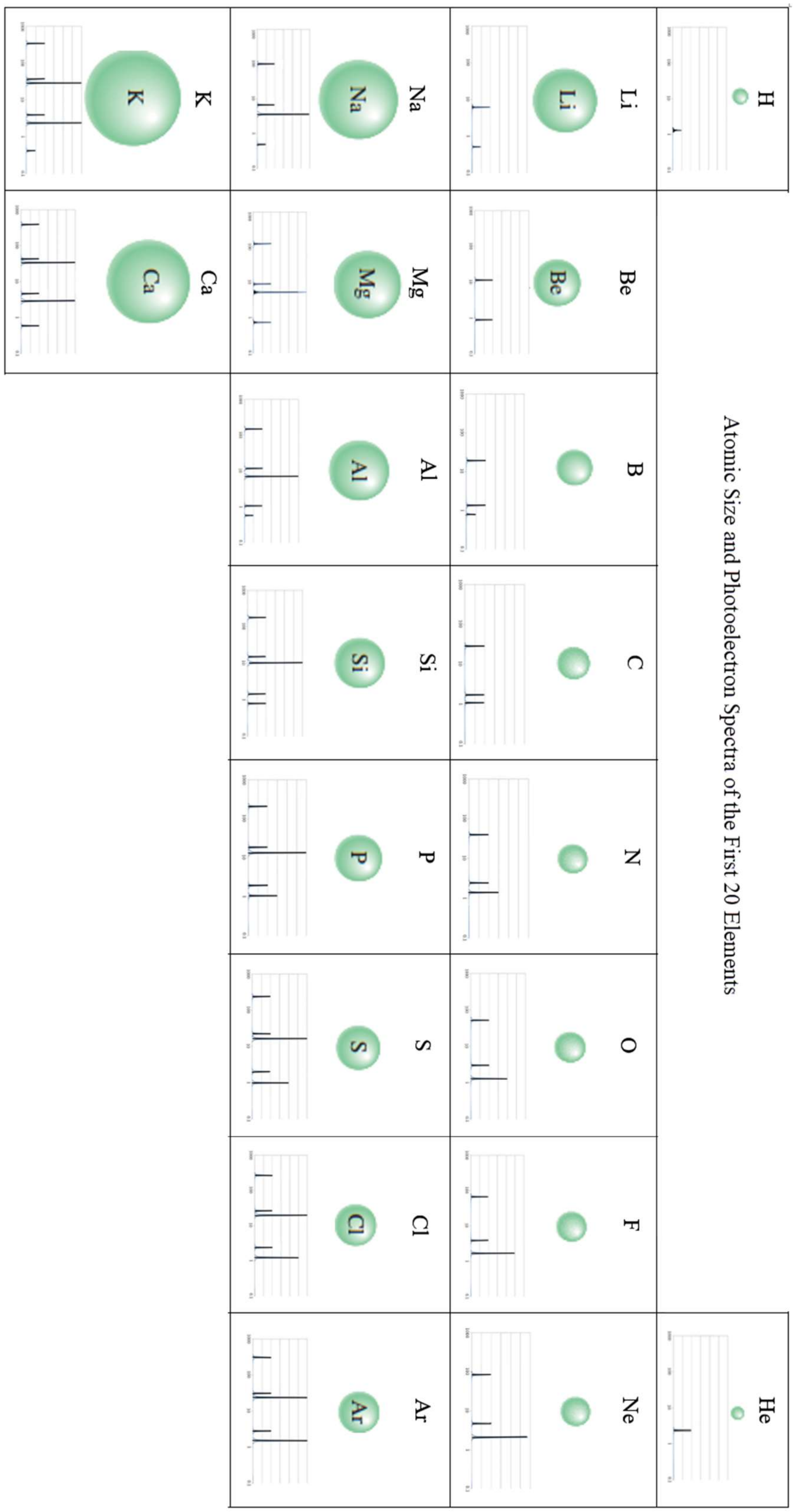


S 16 protons 

atomic radius 103 pm

$1s^2$	$2s^1$	$2p^6$	$3s^2$	$3p^4$
239			2.1	1.0

Again, the effect of the increased number of protons shifts the peaks to the left, with increased the Binding Energies that makes the atom smaller.



Atomic Size and Photoelectron Spectra of the First 20 Elements