

# TOPIC: 1.7 PERIODIC TRENDS

## ENDURING UNDERSTANDING:

SAP-2 The periodic table shows patterns in electronic structure and trends in atomic properties.

## LEARNING OBJECTIVE:

SAP-2.A Explain the relationship between trends in atomic properties of elements and electronic structure and periodicity.

## ESSENTIAL KNOWLEDGE:

SAP-2.A.1 The organization of the periodic table is based on the recurring properties of the elements and explained by the pattern of electron configurations and the presence of completely or partially filled shells (and subshells) of electrons in atoms.

WRITING THE ELECTRON CONFIGURATION OF ELEMENTS THAT ARE EXCEPTIONS TO THE AUFBAU PRINCIPLE WILL NOT BE ASSESSED ON THE AP EXAM.

*Rationale: The mere rote recall of the exceptions does not match the goals of the curriculum revision.*

SAP-2.A.2 Trends in atomic properties within the periodic table (periodicity) can be qualitatively understood through the position of the element in the periodic table, Coulomb's law, the shell model, and the concept of shielding/effective nuclear charge. These properties include: a. Ionization energy b. Atomic and ionic radii c. Electron affinity d. Electronegativity.

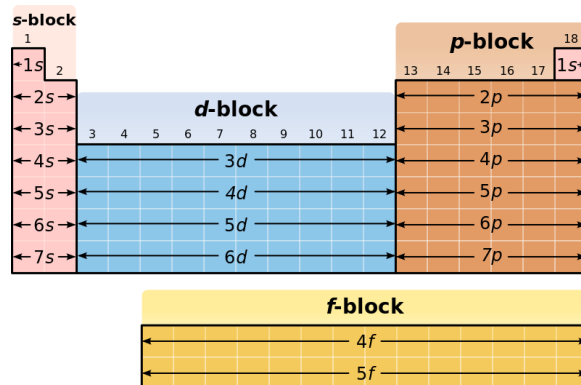
SAP-2.A.3 The periodicity (in SAP-2.A.2) is useful to predict /estimate values of properties in the absence of data.

## EQUATION(S):

N/A

## NOTES:

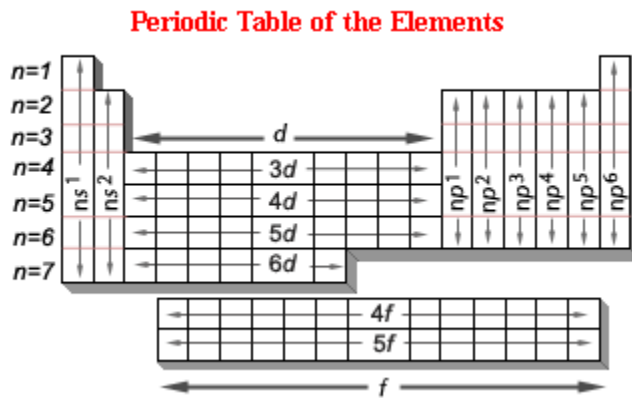
The periodic table is arranged in order from lowest atomic number (# of protons) to highest. The blocks of the periodic table correspond to the s/p/d/f groups for the electron configuration.



<https://socratic.org/questions/what-is-the-electron-configuration-for-francium>

Elements that have the same valence electron configuration tend to have similar chemical properties.

[http://nobel.scas.bcit.ca/wiki/index.php/File:Ptable\\_eco\\_nfig.gif#filelinks](http://nobel.scas.bcit.ca/wiki/index.php/File:Ptable_eco_nfig.gif#filelinks)



Most, if not all, periodic trends can be explained by the arrangement of the electrons and the number of protons in the atoms.

1 H 1s ↑							2 He 1s ↑↓
3 Li 2s ↑ 1s ↑↓	4 Be 2s ↑↓ 1s ↑↓	5 B 2p ↑ — — 2s ↑↓ 1s ↑↓	6 C 2p ↑ ↑ — 2s ↑↓ 1s ↑↓	7 N 2p ↑ ↑ ↑ 2s ↑↓ 1s ↑↓	8 O 2p ↑↓ ↑ ↑ 2s ↑↓ 1s ↑↓	9 F 2p ↑↓ ↑↓ ↑ 2s ↑↓ 1s ↑↓	10 Ne 2p ↑↓ ↑↓ ↑↓ 2s ↑↓ 1s ↑↓
11 Na 3s ↑ 2p ↑↓ ↑↓ ↑↓ 2s ↑↓ 1s ↑↓	12 Mg 3s ↑↓ 2p ↑↓ ↑↓ ↑↓ 2s ↑↓ 1s ↑↓	13 Al 3p ↑ — — 3s ↑↓ 2p ↑↓ ↑↓ ↑↓ 2s ↑↓ 1s ↑↓	14 Si 3p ↑ ↑ — 3s ↑↓ 2p ↑↓ ↑↓ ↑↓ 2s ↑↓ 1s ↑↓	15 P 3p ↑ ↑ ↑ 3s ↑↓ 2p ↑↓ ↑↓ ↑↓ 2s ↑↓ 1s ↑↓	16 S 3p ↑↓ ↑ ↑ 3s ↑↓ 2p ↑↓ ↑↓ ↑↓ 2s ↑↓ 1s ↑↓	17 Cl 3p ↑↓ ↑↓ ↑ 3s ↑↓ 2p ↑↓ ↑↓ ↑↓ 2s ↑↓ 1s ↑↓	18 Ar 3p ↑↓ ↑↓ ↑↓ 3s ↑↓ 2p ↑↓ ↑↓ ↑↓ 2s ↑↓ 1s ↑↓

[https://chem.libretexts.org/Under\\_Construction/Purgatory/Essential\\_Chemistry\\_\(Curriki\)/Unit\\_1%3A\\_Atomic\\_and\\_Molecular\\_Structure/1.4%3A\\_Electron\\_Configuration\\_and\\_Orbital\\_Diagrams](https://chem.libretexts.org/Under_Construction/Purgatory/Essential_Chemistry_(Curriki)/Unit_1%3A_Atomic_and_Molecular_Structure/1.4%3A_Electron_Configuration_and_Orbital_Diagrams)

REMEMBER: Stating a trend is not EXPLAINING a trend. Explanations of trends should never be in terms of the location of the periodic table.

Coulombic Attraction is the electrostatic attraction between two charged particles. Often when discussing periodic trends the charged particles are the nucleus (specifically the total number of protons) and the electrons. Often we are referring to the outermost electrons, the valence electrons.

Coulomb's law states that the attraction between two charged particles is proportional to the magnitude of the charge and inversely proportional to the distance between them. To make this simpler, the larger the charge, the more attractive forces between the particles. The further away the particles are from each other, the weaker the attraction.

## PERIODIC TRENDS

Key Terms:

### *COULOMBIC ATTRACTION/ ELECTROSTATIC INTERACTIONS*

The positive-negative attraction which takes place when you have two charged particles in close proximity.

- Increases with increase in charge
- Increases with decrease in distance between particles

### *EFFECTIVE NUCLEAR CHARGE AND ELECTRON SHIELDING*

The **effective nuclear charge** is the net positive **charge** experienced by valence electrons. It can be approximated by the equation:  $Z_{\text{eff}} = Z - S$ , where  $Z$  is the atomic number and  $S$  is the number of electrons in orbitals that are closer to the nucleus.

#### A) FIRST IONIZATION ENERGY

The energy required to remove the outermost (highest energy) electron from the gas from of a neutral atom in its ground state.

First Ionization energy **decreases** as you move down a group. Electrons are further from the nucleus and therefore have a lower Coulombic attraction. Additionally, the inner shells of electrons **shield** or block the protons force of attraction, so that outermost electrons do not feel as much of the nuclear force. This results in the outer electrons being even easier to remove.

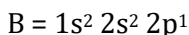
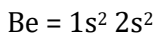
First Ionization energy **increases** as you move across a period on the periodic table, from left to right. As you move across the period the atomic radius is smaller and there is an increase in protons in the nucleus. Both factors result in greater Coulombic attraction, which in turn means that it will require more energy to remove the first electron.

<https://wps.pearsoned.com.au/ibcsl/89/22896/5861561.cw/content/index.html>

There are a few places where the ionization doesn't appear to follow a trend. You can see this on the graph between Be and B or between N and O. These are actually for two slightly different reasons.

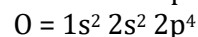
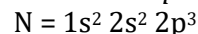
[https://useruploads.socratic.org/N5qKJ5fTLiJK3MXQAifQ\\_Ionization\\_Energy\\_Trend\\_IK.png](https://useruploads.socratic.org/N5qKJ5fTLiJK3MXQAifQ_Ionization_Energy_Trend_IK.png)

*Be and B exception ( $s^2$  to  $s^2p^1$ )*



When the first electron is removed from the boron, B, atom, the electron is being removed from the 2p orbital. Since the 2p orbital is further away from the nucleus it takes less energy to remove it even though there are more protons in the atom.

*N and O exception ( $s^2p^3$  to  $s^2p^4$ )*



When the first electron is removed from oxygen it takes less energy (despite the increase in protons) than from nitrogen because the electrons in oxygen are sharing the  $2p_x$  orbital and therefore have greater electron-electron repulsions making it easier to remove one electron.

The second ionization energy is the energy to remove a second electron from the atom and so on for each successive electron.

By examining the successive ionization energies for an element we can determine how many valence electrons there are in that element. When all of the valence electrons have been removed, you will see a large "jump" in the ionization energy values. This "jump" is due to the fact that the core electrons are closer to and less shielded from the nucleus and therefore it requires more energy to remove them.

For example:

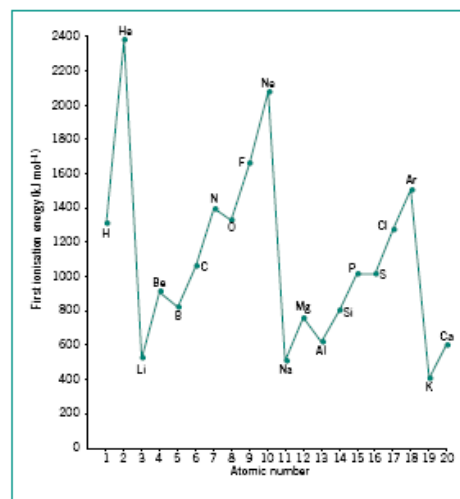
Consider magnesium, Mg, the electron configuration is  $1s^2 2s^2 2p^6 3s^2$  and we can see that it has 2 valence electrons.

<https://www.webelements.com/magnesium/atoms.html>

You can see that there is a big jump between the 2<sup>nd</sup> and 3<sup>rd</sup> ionization energies and again between the 10<sup>th</sup> and 11<sup>th</sup> ionization energies. This shows when electrons are being removed from a shell that is closer to the nucleus.

## INCREASING IONIZATION ENERGY

INCREASING IONIZATION ENERGY																	
1	2											18	19	20			
H	He											Ar	K	Ca			
3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18		
Li	Be	B	C	N	O	F	Ne	Na	Mg	Al	Si	P	S	Cl	Ar		
23	24	25	26	27	28	29	30	31	32	33	34	35	36	37	38		
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se		
39	40	41	42	43	44	45	46	47	48	49	50	51	52	53	54		
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te		
55	56	57	72	73	74	75	76	77	78	79	80	81	82	83	84		
Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po		
87	88	89	104	105	106	107	108	109	110	111	112	113	114				
Fr	Ra	Ac	Rf	Db	Sg	Bh	Hs	Mt									



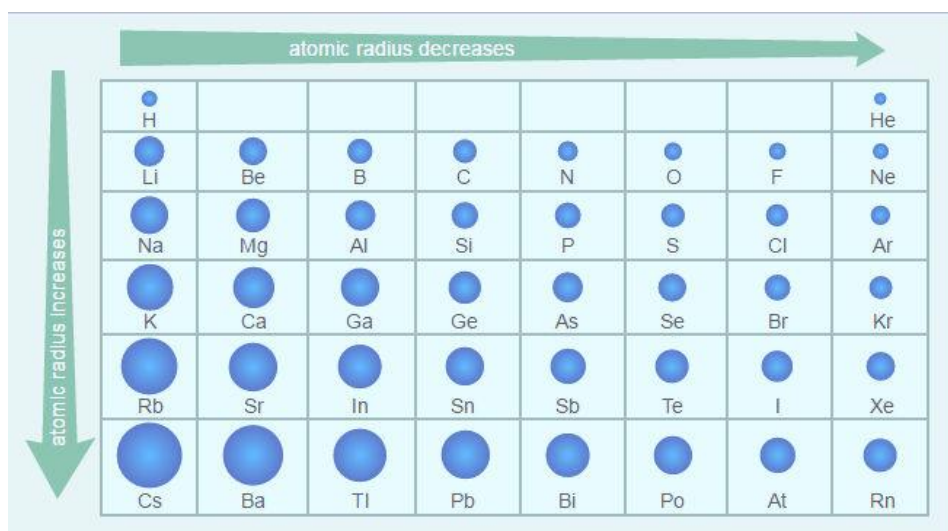
Ionization Energy Number	Enthalpy (kJ/mole)
1 <sup>st</sup>	738
2 <sup>nd</sup>	1451
3 <sup>rd</sup>	7733
4 <sup>th</sup>	10543
5 <sup>th</sup>	13636
6 <sup>th</sup>	18020
7 <sup>th</sup>	21711
8 <sup>th</sup>	25658
9 <sup>th</sup>	31646
10 <sup>th</sup>	35457
11 <sup>th</sup>	169988

## B) ATOMIC RADIUS

The **atomic radius** of a chemical element is a measure of the size of its **atoms**, usually the mean or typical distance from the center of the nucleus to the boundary of the surrounding cloud of electrons.

Atomic Radii **increases** as you move down a column as there are more electron shells.

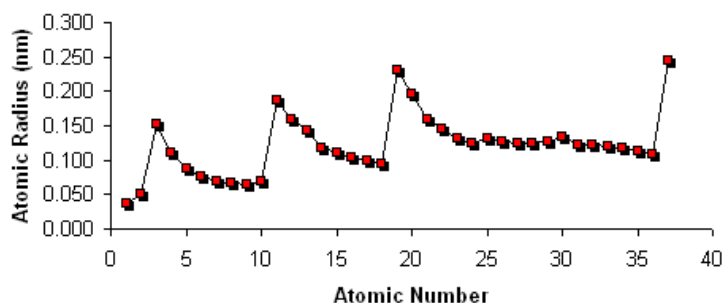
<https://byjus.com/chemistry/atomic-radius-in-periodic-table-in-basic-chemistry/>



Atomic Radii **decreases** as you move across a period on the periodic table, from left to right. Electrons are being added to the same energy level. At the same time, protons are being added to the nucleus. Increasing the number of protons gives a **higher effective nuclear charge**. In other words, there is a stronger force of attraction pulling the electrons closer to the nucleus. This results in a smaller atomic radius, as with greater numbers of protons there is more pull on the electrons.

[https://www.geocities.ws/junebug\\_sophia/atmRad.gif](https://www.geocities.ws/junebug_sophia/atmRad.gif)

Atomic Radius vs Atomic Number

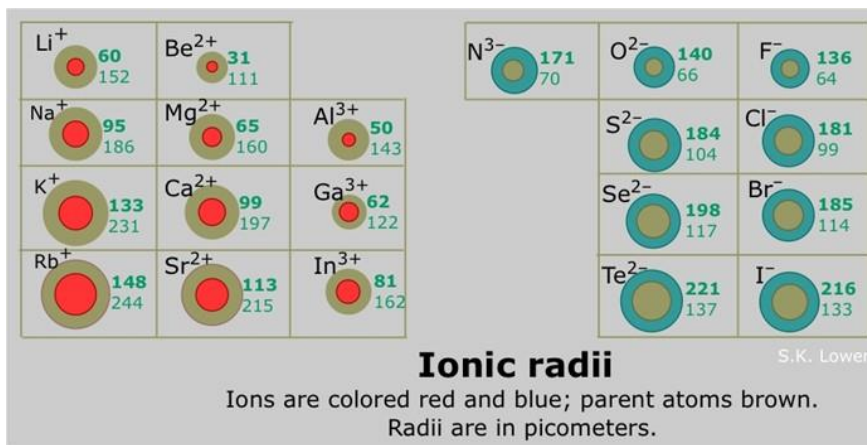


## IONIC RADIUS

The trends for ionic radii are similar to those of atomic radii, except that cations and anions are different from each other.

Cations are always smaller than the parent atoms, because they have lost their valence shell. This causes them to be smaller. They also decrease in size because the nuclear attraction is now acting on fewer electrons so they are drawn in toward the nucleus due to the greater attraction. Additionally there are fewer electron-electron repulsions.

Anions, on the other hand, are always larger than the parent atom. Electrons are added to the same valence shell; however, there are greater electron-electron repulsions so the ion increases in size.



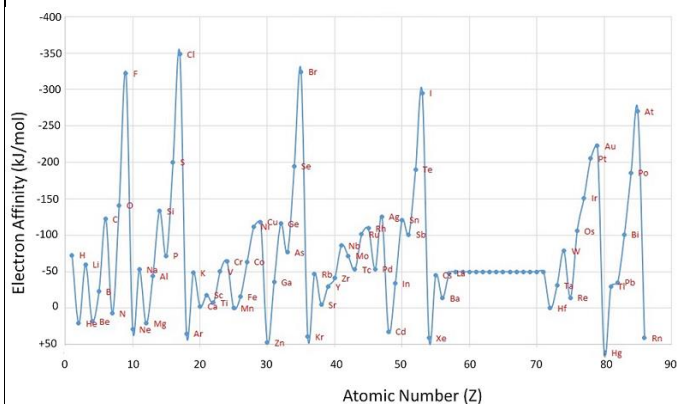
<https://slideplayer.com/slide/8861824/>

## C) ELECTRON AFFINITY

Electron affinity is the amount of energy involved when an electron is accepted by a gaseous atom to form a negative ion. In other words, the neutral atom's likelihood of gaining an electron. The values tend to be negative to show the energy is released as electrons are added to the atoms.

In general, the electron affinity increases from left to right on the periodic table. This is caused by the filling of the valence shell of the atom; a Group 17 atom releases more energy than a Group 1 atom when it gains an electron, this indicates that it is more stable.

[https://chem.libretexts.org/Bookshelves/Physical\\_and\\_Theoretical\\_Chemistry\\_Textbook\\_Maps/Supplemental\\_Modules\\_\(Physical\\_and\\_Theoretical\\_Chemistry\)/Physical\\_Properties\\_of\\_Matter/Atomic\\_and\\_Molecular\\_Properties/Electron\\_Affinity](https://chem.libretexts.org/Bookshelves/Physical_and_Theoretical_Chemistry_Textbook_Maps/Supplemental_Modules_(Physical_and_Theoretical_Chemistry)/Physical_Properties_of_Matter/Atomic_and_Molecular_Properties/Electron_Affinity)



## D) ELECTRONEGATIVITY

Electronegativity is a measure of the ability of an atom (or group of atoms) to attract shared electrons.

Electronegativity **decreases** as you move down a column as there is a greater distance from the nucleus and because there is also more electron shielding.

Electronegativity **increases** as you move across a period on the periodic table, from left to right. This is because the atomic radius is decreasing while the number of protons (and effective nuclear charge) is increasing.

Fluorine is the most electronegative element.

### INCREASING ELECTRON AFFINITY

1 H Hydrogen 1.00794	2 He Helium 4.002602																	17 Cl Chlorine 35.453	18 Ar Argon 39.948
3 Li Lithium 6.941	4 Be Beryllium 9.012182	5 B Boron 10.811	6 C Carbon 12.011	7 N Nitrogen 14.00643	8 O Oxygen 15.999	9 F Fluorine 18.998403	10 Ne Neon 20.1797	11 Na Sodium 22.989769	12 Mg Magnesium 24.30508	13 Al Aluminum 26.981538	14 Si Silicon 28.0855	15 P Phosphorus 30.973762	16 S Sulfur 32.065	17 Cl Chlorine 35.453	18 Ar Argon 39.948				
19 K Potassium 39.0983	20 Ca Calcium 40.078	21 Sc Scandium 44.955912	22 Ti Titanium 47.867	23 V Vanadium 50.9415	24 Cr Chromium 51.9961	25 Mn Manganese 54.938044	26 Fe Iron 55.845	27 Co Cobalt 58.933200	28 Ni Nickel 58.6934	29 Cu Copper 63.546	30 Zn Zinc 65.39	31 Ga Gallium 69.723	32 Ge Germanium 72.63	33 As Arsenic 74.92160	34 Se Selenium 78.96	35 Br Bromine 79.904	36 Kr Krypton 83.80		
37 Rb Rubidium 85.4678	38 Sr Strontium 87.62	39 Y Yttrium 88.90584	40 Zr Zirconium 91.224	41 Nb Niobium 92.90638	42 Mo Molybdenum 95.94	43 Tc Technetium 98.90625	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.90550	46 Pd Palladium 106.42	47 Ag Silver 107.8682	48 Cd Cadmium 112.411	49 In Indium 114.818	50 Sn Tin 118.710	51 Sb Antimony 121.757	52 Te Tellurium 127.60	53 I Iodine 126.90545	54 Xe Xenon 131.29		
55 Cs Cesium 132.90545	56 Ba Barium 137.327	57 La Lanthanum 138.9055	72 Zn Zinc 65.39	73 Ga Gallium 69.723	74 Ge Germanium 72.63	75 As Arsenic 74.92160	76 Se Selenium 78.96	77 Br Bromine 79.904	78 Kr Krypton 83.80	79 Rb Rubidium 85.4678	80 Sr Strontium 87.62	81 Y Yttrium 88.90584	82 Zr Zirconium 91.224	83 Nb Niobium 92.90638	84 Mo Molybdenum 95.94	85 Tc Technetium 98.90625	86 Ru Ruthenium 101.07		
87 Fr Francium 223	88 Ra Radium 226	89 Ac Actinium 227	104 Rf Rutherfordium 261	105 Db Dubnium 262	106 Sg Seaborgium 263	107 Bh Bohrium 264	108 Hs Hassium 265	109 Mt Meitnerium 266	110 Ds Darmstadtium 269	111 Rg Roentgenium 272	112 Cn Copernicium 277	113 Nh Nihonium 284	114 Fl Flerovium 289	115 Mc Moscovium 288	116 Lv Livermorium 293	117 Ts Tennessine 289	118 Og Oganesson 294		

A trend of decreasing electron affinity when moving down the groups in the periodic table might be expected. The additional electron will be entering an orbital farther away from the nucleus. Since this electron is farther from the nucleus it is less attracted to the nucleus and would release less energy when added. However, a clear counterexample to this trend can be found in Group 2, and inspecting the entire periodic table, it turns out that the proposed trend only applies to Group 1 atoms.

### INCREASING ELECTRONEGATIVITY

1 H Hydrogen 1.00794	2 He Helium 4.002602																	17 Cl Chlorine 35.453	18 Ar Argon 39.948
3 Li Lithium 6.941	4 Be Beryllium 9.012182	5 B Boron 10.811	6 C Carbon 12.011	7 N Nitrogen 14.00643	8 O Oxygen 15.999	9 F Fluorine 18.998403	10 Ne Neon 20.1797	11 Na Sodium 22.989769	12 Mg Magnesium 24.30508	13 Al Aluminum 26.981538	14 Si Silicon 28.0855	15 P Phosphorus 30.973762	16 S Sulfur 32.065	17 Cl Chlorine 35.453	18 Ar Argon 39.948				
19 K Potassium 39.0983	20 Ca Calcium 40.078	21 Sc Scandium 44.955912	22 Ti Titanium 47.867	23 V Vanadium 50.9415	24 Cr Chromium 51.9961	25 Mn Manganese 54.938044	26 Fe Iron 55.845	27 Co Cobalt 58.933200	28 Ni Nickel 58.6934	29 Cu Copper 63.546	30 Zn Zinc 65.39	31 Ga Gallium 69.723	32 Ge Germanium 72.63	33 As Arsenic 74.92160	34 Se Selenium 78.96	35 Br Bromine 79.904	36 Kr Krypton 83.80		
37 Rb Rubidium 85.4678	38 Sr Strontium 87.62	39 Y Yttrium 88.90584	40 Zr Zirconium 91.224	41 Nb Niobium 92.90638	42 Mo Molybdenum 95.94	43 Tc Technetium 98.90625	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.90550	46 Pd Palladium 106.42	47 Ag Silver 107.8682	48 Cd Cadmium 112.411	49 In Indium 114.818	50 Sn Tin 118.710	51 Sb Antimony 121.757	52 Te Tellurium 127.60	53 I Iodine 126.90545	54 Xe Xenon 131.29		
55 Cs Cesium 132.90545	56 Ba Barium 137.327	57 La Lanthanum 138.9055	72 Zn Zinc 65.39	73 Ga Gallium 69.723	74 Ge Germanium 72.63	75 As Arsenic 74.92160	76 Se Selenium 78.96	77 Br Bromine 79.904	78 Kr Krypton 83.80	79 Rb Rubidium 85.4678	80 Sr Strontium 87.62	81 Y Yttrium 88.90584	82 Zr Zirconium 91.224	83 Nb Niobium 92.90638	84 Mo Molybdenum 95.94	85 Tc Technetium 98.90625	86 Ru Ruthenium 101.07		
87 Fr Francium 223	88 Ra Radium 226	89 Ac Actinium 227	104 Rf Rutherfordium 261	105 Db Dubnium 262	106 Sg Seaborgium 263	107 Bh Bohrium 264	108 Hs Hassium 265	109 Mt Meitnerium 266	110 Ds Darmstadtium 269	111 Rg Roentgenium 272	112 Cn Copernicium 277	113 Nh Nihonium 284	114 Fl Flerovium 289	115 Mc Moscovium 288	116 Lv Livermorium 293	117 Ts Tennessine 289	118 Og Oganesson 294		

Example	Difference in electronegativity	Type of Bond
H-H	No difference – Electrons are shared equally	Nonpolar covalent bond
H-Br	Slight difference in values – Electrons are shared unequally	Polar covalent bond
NaCl	Large difference in values – Electrons are not shared, they are transferred	Ionic Bond

**IDO:**

For each of the following pairs of elements

Na or Li      F and O

Choose the atom with:

- a) Higher first ionization energy

*Li and F*

- b) Larger atomic radius

*Na and O*

- c) Higher electronegativity

*Li and F*

1																	2
H																	He
1.00794																	4.002602
3	4											5	6	7	8	9	10
Li	Be											B	C	N	O	F	Ne
6.941	9.012182											10.811	12.0107	14.00674	15.9994	18.9984032	20.1797
11	12											13	14	15	16	17	18
Na	Mg											Al	Si	P	S	Cl	Ar
22.989770	24.3050											26.981538	28.0855	30.973761	32.066	35.4527	39.948
19	20	21	22	23	24	25	26	27	28	29	30	31	32	33	34	35	36
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
39.0983	40.078	44.955910	47.867	50.9415	51.9961	54.938049	55.845	58.933200	58.6334	63.545	65.39	69.723	72.61	74.92160	78.96	79.904	83.80
37	38	39	40	41	42	43	44	45	46	47	48	49	50	51	52	53	54
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
85.4678	87.62	88.90585	91.224	92.90638	95.94	(98)	101.07	102.90550	106.42	106.90555	112.411	114.818	118.710	121.760	127.60	126.90447	131.29
55	56	57	72	73	74	75	76	77	78	79	80	81	82	83	84	85	86
Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
132.90545	137.327	138.9055	178.49	180.9479	183.84	186.207	190.23	192.217	195.078	196.96655	200.59	204.3833	207.2	208.98038	(209)	(210)	(222)
87	88	89	104	105	106	107	108	109	110	111	112				114	116	118
Fr	Ra	Ac	Rf	Db	Sg	Bh	Hs	Mt							(289)	(287)	(293)
(223)	(226)	(227)	(261)	(262)	(263)	(262)	(265)	(266)	(269)	(272)	(277)				(289)	(287)	(293)

58	59	60	61	62	63	64	65	66	67	68	69	70	71
Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu
140.116	140.50765	144.24	(145)	150.36	151.964	157.25	158.92534	162.50	164.93032	167.26	168.93421	173.04	174.967
90	91	92	93	94	95	96	97	98	99	100	101	102	103
Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr
232.0381	231.03688	238.0289	(237)	(244)	(243)	(247)	(247)	(251)	(252)	(257)	(258)	(259)	(262)

**WE DO:**

Rank the following from smallest to largest atomic/ionic radius.

- a) Na<sup>+</sup>, Na, Na<sup>-</sup>  
 b) C, N, O  
 c) Cl, Ar, K  
 d) Be, Mg, Ca

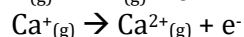
1																	2
H																	He
1.00794																	4.002602
3	4											5	6	7	8	9	10
Li	Be											B	C	N	O	F	Ne
6.941	9.012182											10.811	12.0107	14.00674	15.9994	18.9984032	20.1797
11	12											13	14	15	16	17	18
Na	Mg											Al	Si	P	S	Cl	Ar
22.989770	24.3050											26.981538	28.0855	30.973761	32.066	35.4527	39.948
19	20	21	22	23	24	25	26	27	28	29	30	31	32	33	34	35	36
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
39.0983	40.078	44.955910	47.867	50.9415	51.9961	54.938049	55.845	58.933200	58.6334	63.545	65.39	69.723	72.61	74.92160	78.96	79.904	83.80
37	38	39	40	41	42	43	44	45	46	47	48	49	50	51	52	53	54
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
85.4678	87.62	88.90585	91.224	92.90638	95.94	(98)	101.07	102.90550	106.42	106.90555	112.411	114.818	118.710	121.760	127.60	126.90447	131.29
55	56	57	72	73	74	75	76	77	78	79	80	81	82	83	84	85	86
Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
132.90545	137.327	138.9055	178.49	180.9479	183.84	186.207	190.23	192.217	195.078	196.96655	200.59	204.3833	207.2	208.98038	(209)	(210)	(222)
87	88	89	104	105	106	107	108	109	110	111	112				114	116	118
Fr	Ra	Ac	Rf	Db	Sg	Bh	Hs	Mt							(289)	(287)	(293)
(223)	(226)	(227)	(261)	(262)	(263)	(262)	(265)	(266)	(269)	(272)	(277)				(289)	(287)	(293)

**YOU DO:**

- 1) On the basis of their position on the periodic table determine which element in the pair would have a larger atomic radius
- a) P or S
- b) Cl or Br
- c) Sr or Sc
- 2) Based on the successive ionization energies for the following element "X", predict the formula that would be formed when "X" reacts with chlorine, Cl.

Ionization Energy Number	Enthalpy (kJ/mole)
1 <sup>st</sup>	577
2 <sup>nd</sup>	1820
3 <sup>rd</sup>	2740
4 <sup>th</sup>	11600
5 <sup>th</sup>	14841

- 3) The first ionization energy for potassium, K, is 419 kJ/mol and the second ionization energy for calcium, Ca, is 1145 kJ/mol. Using concepts from this unit explain why they are different even though they are isoelectric (have the same number of electrons).



- 4) Element X has an electron configuration of  $1s^2 2s^2 2p^6 3s^1$ , while element Z has an electron configuration of  $1s^2 2s^2 2p^5$ .
- Which element would have greater first ionization energy?
  - Which element would have a larger radius?
  - Which element would have higher electronegativity?
  - Which element would form an ion that has a larger radius?
  - Which element would release more energy when it gains an electron?

- 5) Predict two elements that would have properties similar to:

- Chlorine
- Sodium
- Calcium

- 6) Nitrogen is in column 5A of the periodic table, which is called the pnictogens. When nitrogen reacts with iodine it forms nitrogen triiodide,  $\text{NI}_3$ , which is a contact explosive that explodes with a snap releasing clouds of purple iodine vapor. Select another pnictogen and predict the formula of the compound that would be formed with a reaction with bromine.

- 7) Based on the given electron configurations, group together the elements that would have similar chemical properties.

- $1s^2 2s^2 2p^6 3s^1$
- $1s^2 2s^2 2p^6 3s^2 3p^4 4s^2$
- $1s^2 2s^2 2p^5$
- $1s^2 2s^2 2p^6 3s^2 3p^4 4s^2 4p^6 5s^2$
- $1s^2 2s^1$
- $1s^2 2s^2 2p^6 3s^2 3p^5$