Quantum Mechanics 2

Robert C. Roleda
Physics Department

The Periodic Table

De La Salle University

Periodic Table of the Elements

1A	2A	3B	4B	5B	6B	7B	8B			1B	2B	3A	4A	5A	6A	7A	8A
1 **H** 1.00794 Hydrogen																	2 **He** 4.002602 Helium
3 **Li** 6.941 Lithium	4 **Be** 9.012182 Beryllium											5 **B** 10.811 Boron	6 **C** 12.0107 Carbon	7 **N** 14.0067 Nitrogen	8 **O** 15.9994 Oxygen	9 **F** 18.9984032 Fluorine	10 **Ne** 20.1797 Neon
11 **Na** 22.989769 Sodium	12 **Mg** 24.3050 Magnesium											13 **Al** 26.9815386 Aluminum	14 **Si** 28.0855 Silicon	15 **P** 30.973762 Phosphorus	16 **S** 32.065 Sulfur	17 **Cl** 35.453 Chlorine	18 **Ar** 39.948 Argon
19 **K** 39.0983 Potassium	20 **Ca** 40.078 Calcium	21 **Sc** 44.955912 Scandium	22 **Ti** 47.867 Titanium	23 **V** 50.9415 Vanadium	24 **Cr** 51.9961 Chromium	25 **Mn** 54.938045 Manganese	26 **Fe** 55.845 Iron	27 **Co** 58.933195 Cobalt	28 **Ni** 58.6934 Nickel	29 **Cu** 63.546 Copper	30 **Zn** 65.38 Zinc	31 **Ga** 69.723 Gallium	32 **Ge** 72.64 Germanium	33 **As** 74.92160 Arsenic	34 **Se** 78.96 Selenium	35 **Br** 79.904 Bromine	36 **Kr** 83.798 Krypton
37 **Rb** 85.4678 Rubidium	38 **Sr** 87.62 Strontium	39 **Y** 88.90585 Yttrium	40 **Zr** 91.224 Zirconium	41 **Nb** 92.90638 Niobium	42 **Mo** 95.96 Molybdenum	43 **Tc** [98] Technetium	44 **Ru** 101.07 Ruthenium	45 **Rh** 102.90550 Rhodium	46 **Pd** 106.42 Palladium	47 **Ag** 107.8682 Silver	48 **Cd** 112.411 Cadmium	49 **In** 114.818 Indium	50 **Sn** 118.710 Tin	51 **Sb** 121.760 Antimony	52 **Te** 127.60 Tellurium	53 **I** 126.90447 Iodine	54 **Xe** 131.293 Xenon
55 **Cs** 132.9054519 Cesium	56 **Ba** 137.327 Barium	57-71 Lanthanides	72 **Hf** 178.49 Hafnium	73 **Ta** 180.94788 Tantalum	74 **W** 183.84 Tungsten	75 **Re** 186.207 Rhenium	76 **Os** 190.23 Osmium	77 **Ir** 192.217 Iridium	78 **Pt** 195.084 Platinum	79 **Au** 196.966569 Gold	80 **Hg** 200.59 Mercury	81 **Tl** 204.3833 Thallium	82 **Pb** 207.2 Lead	83 **Bi** 208.98040 Bismuth	84 **Po** [209] Polonium	85 **At** [210] Astatine	86 **Rn** [222] Radon
87 **Fr** [223] Francium	88 **Ra** [226] Radium	89-103 Actinides	104 **Rf** [267] Rutherfordium	105 **Db** [268] Dubnium	106 **Sg** [271] Seaborgium	107 **Bh** [272] Bohrium	108 **Hs** [270] Hassium	109 **Mt** [276] Meitnerium	110 **Ds** [281] Darmstadtium	111 **Rg** [280] Roentgenium	112 **Cn** [285] Copernicium	113 **Uut** [284] Ununtrium	114 **Uuq** [289] Ununquadium	115 **Uup** [288] Ununpentium	116 **Uuh** [293] Ununhexium	117 **Uus** [294] Ununseptium	118 **Uuo** [294] Ununoctium

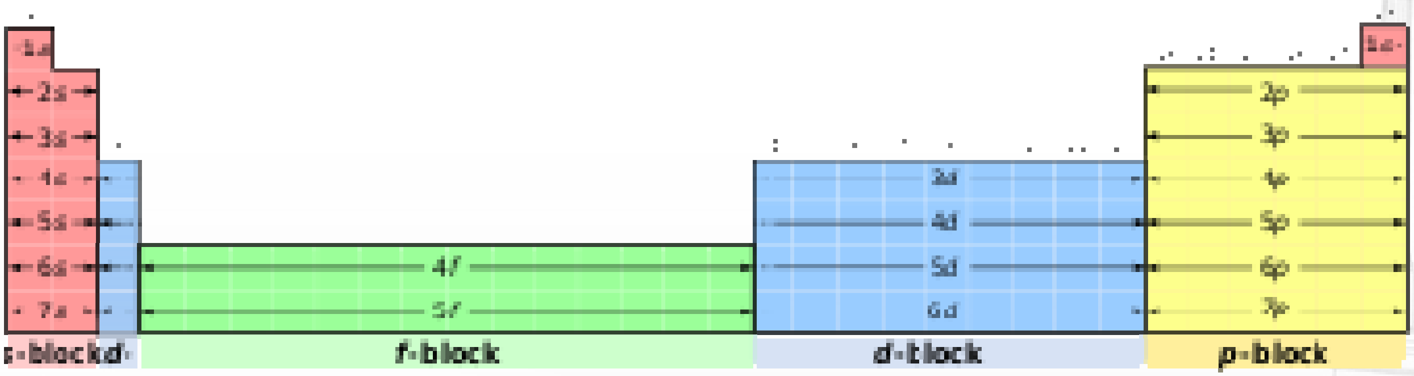
Lanthanides	57 **La** 138.90547 Lanthanum	58 **Ce** 140.116 Cerium	59 **Pr** 140.90765 Praseodymium	60 **Nd** 144.242 Neodymium	61 **Pm** [145] Promethium	62 **Sm** 150.36 Samarium	63 **Eu** 151.964 Europium	64 **Gd** 157.25 Gadolinium	65 **Tb** 158.92535 Terbium	66 **Dy** 162.500 Dysprosium	67 **Ho** 164.93032 Holmium	68 **Er** 167.259 Erbium	69 **Tm** 168.93421 Thulium	70 **Yb** 173.054 Ytterbium	71 **Lu** 174.9668 Lutetium
Actinides	89 **Ac** [227] Actinium	90 **Th** 232.03806 Thorium	91 **Pa** 231.03588 Protactinium	92 **U** 238.02891 Uranium	93 **Np** [237] Neptunium	94 **Pu** [244] Plutonium	95 **Am** [243] Americium	96 **Cm** [247] Curium	97 **Bk** [247] Berkelium	98 **Cf** [251] Californium	99 **Es** [252] Einsteinium	100 **Fm** [257] Fermium	101 **Md** [258] Mendelevium	102 **No** [259] Nobelium	103 **Lr** [262] Lawrencium

Alkali Metals	Alkaline Earth	Basic Metal	Halogen	Noble Gas	Non Metal	Rare Earth	Semi Metal	Transition Metal

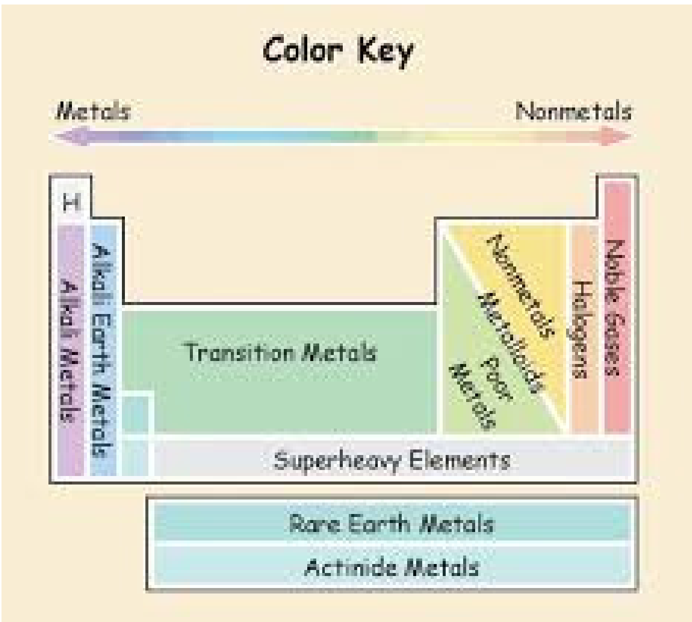
The Periodic Table

The periodic table that Dmitri Mendeleev first published in 1869 was based on the organization of elements according to atomic weight. In 1871, he developed it in a new form with groups of elements arranged in columns. The columns numbered I to VIII corresponds with the element's oxidation state.

The modern periodic table was developed in 1923 by Horace Groves Deming with elements now arranged according to atomic number, which is equal to the number of electrons. With the advent of quantum mechanics, it then became apparent that each period in the table corresponds to the filling of shells of electrons.



Groups



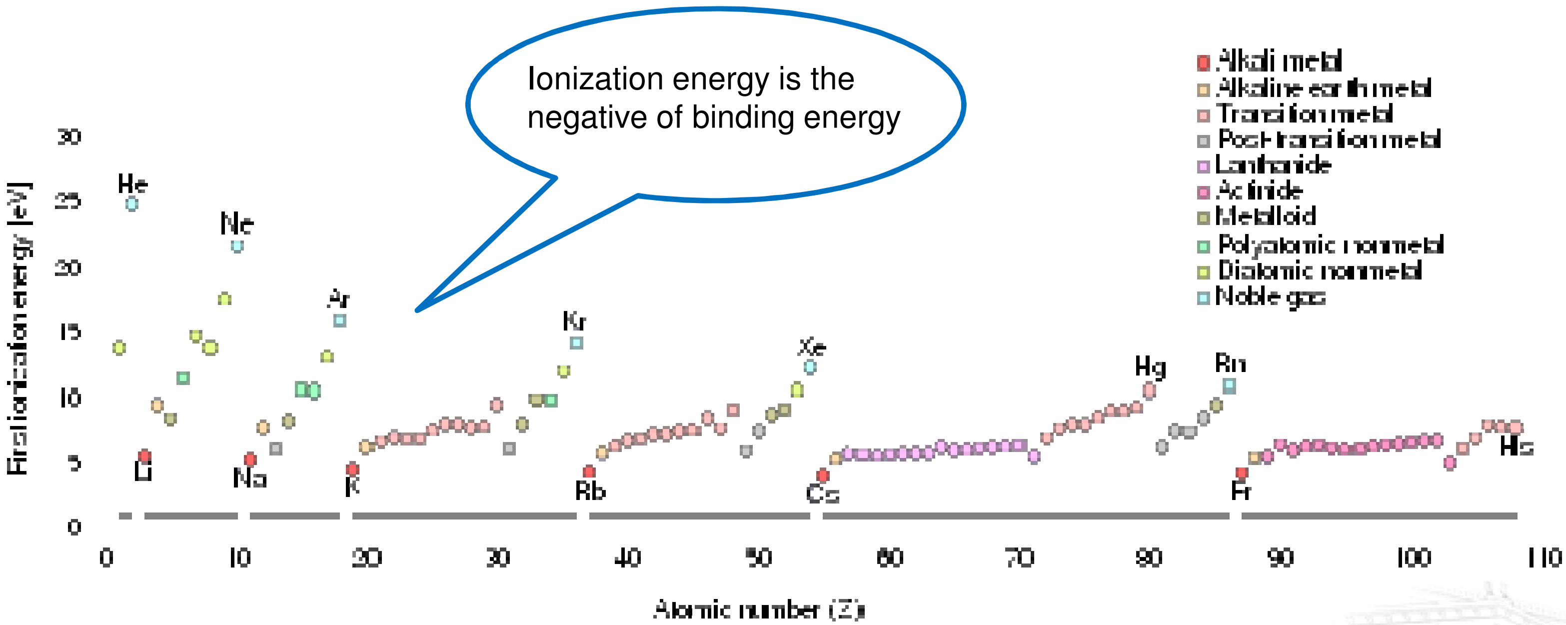
Elements of the same group exhibit similar chemical characteristics because they have the same number of electrons in their outermost shell.

Alkali metals (1A) for example all have one electron on the outermost s – subshell. Alkaline Earth elements (2A) all have two electrons on the outermost s – subshell. Noble gases* (8A) have p – subshells filled with 6 electrons. Halogens (7A) have outermost p – subshells filled with 5 electrons.

The outermost subshell of transition metals is a d – subshells (hence the 10 columns). The 14 elements in each rare earth period have outermost f – subshells.

* Except Helium

Noble Gases

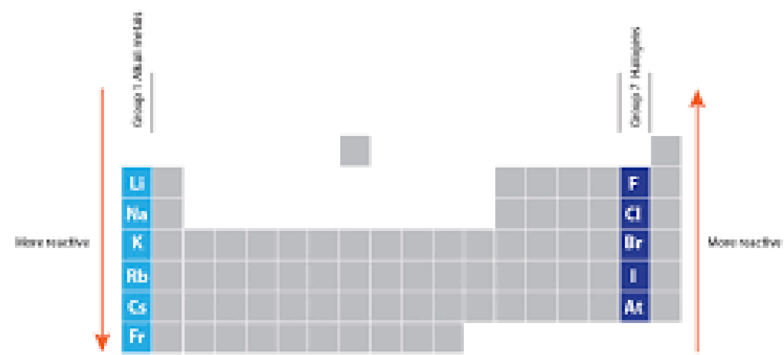


Elements in Group 8A are called noble gases because they are inert, that is non-reactive. The reason for this is that charge distribution in a subshell that is filled out with electrons become spherical and therefore more compact. The electrical interaction between the nucleus and the outermost electrons becomes stronger. The atom will therefore end up in a higher energy state if it either accepts another electron or give up one. Helium has a completely filled s – subshell while the other noble elements have completely filled p – subshells.

The Alkali and the Halogens

Group 1A elements are called Alkali metals, and Group VIIA are called Halogens.

Both group are one step away from the inert gases, Alkali metals having one electrons more than closed p – subshells, while Halogens are short of one electron to form a closed p – subshell.



By giving away its extra electron, Alkali metals would dip to a much lower energy level (or higher ionization energy). Likewise, Halogens goes to a lower energy state if they gain an extra electron. It is for this reason that both groups are the most reactive elements on the Table.

By shedding an electron, Alkali metals become positive ions. By gaining an extra electron, Halogens become negative ions. It is therefore quite common to see Alkali metals bonding with Halogens to form neutral Alkali Halide molecules. The bonds between two ions are called Ionic Bonds.

Atomic Radius

In [hydrogen 5], we have shown that the average distance of an electron from the nucleus is

$$\langle r \rangle = \frac{3n^2 - l(l+1)}{2Z} a_0$$

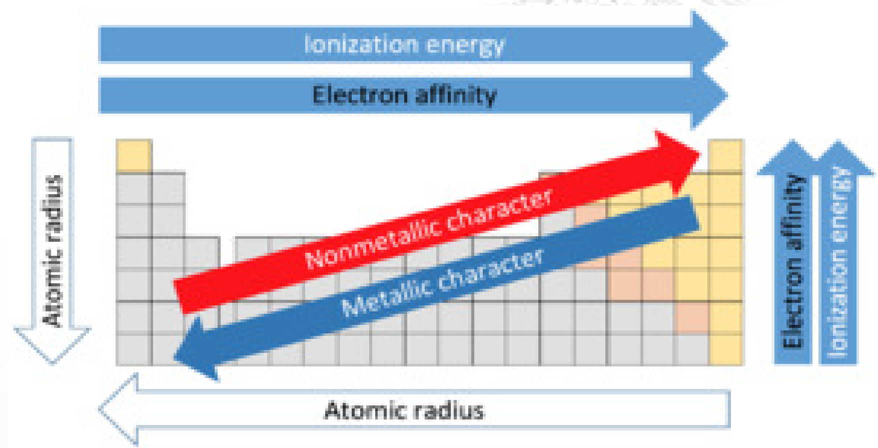
minus

denominator

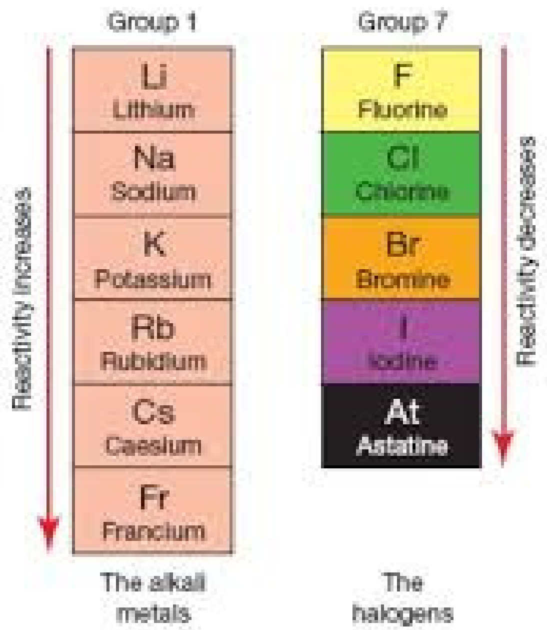
Thus, within one period (row, with fixed n), the atomic radius decreases as we move right since the electrons occupy higher l states as we move to the right, and the atomic number Z likewise go higher as we move to the right.

Within one group (column), atomic radius increases as we go down from one period to the next.

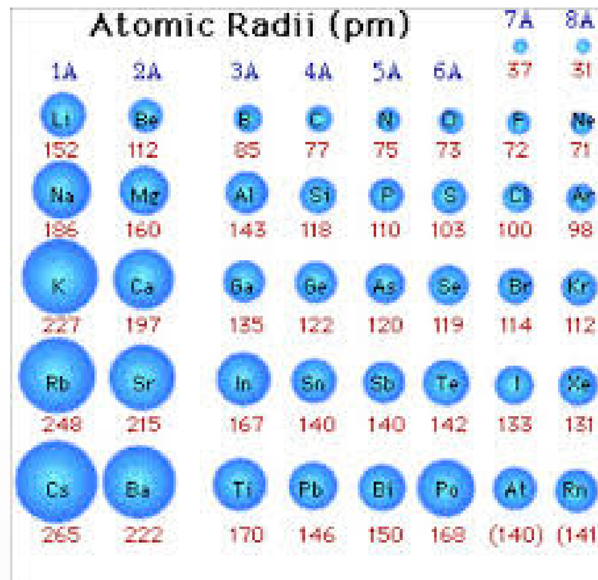
Electron affinity and ionization energy are inversely related to atomic size since the electrical interaction between the nucleus and an outermost electron is weaker with larger distances.



Reactivity



As one goes down the periodic table, the atomic radius increases. For Alkali metals, this means that the lone electron in the outermost shell has a weaker bond with the nucleus. As such, chemical elements become more reactive as we go down the periods. While Sodium seethes and boils when in contact with water, Potassium ignites, while Rubidium and Cesium explode.

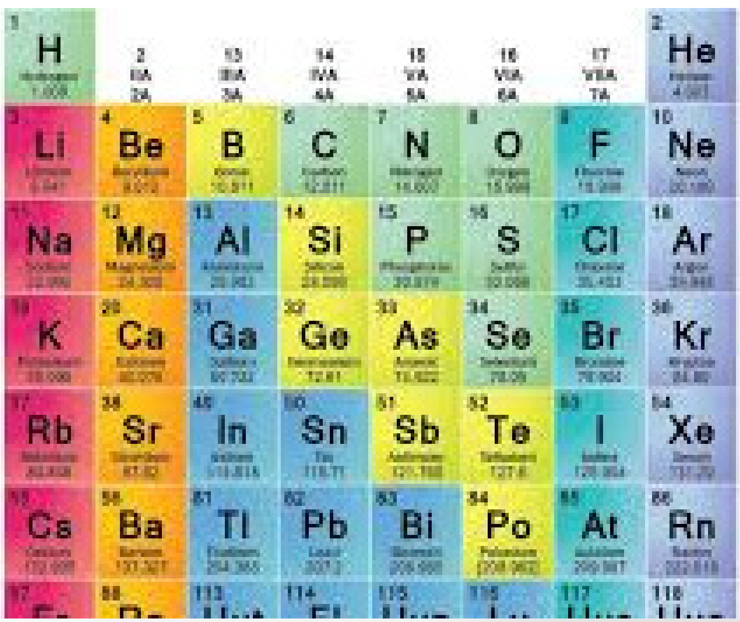


Halogens on the other hand become more reactive as we go up the periods. This is because the atom "gains more" binding energy (binding energy becomes more negative) when a smaller atom takes in the additional electron.

The Main Group Elements

The Alkaline Earth elements of Group 2A are less reactive than the Group 1A Alakali metals. Not only is it because the Alkalines are smaller, but it is also because they have filled s-shells.

Electron-electron interaction becomes more important as the number of electrons increases. This interaction consists of two parts – the electrical (Coulomb) interaction which is repulsive, and an attractive component due to spins. As a rule, electrons with the same spin orientation attract each other more than those with opposite orientation

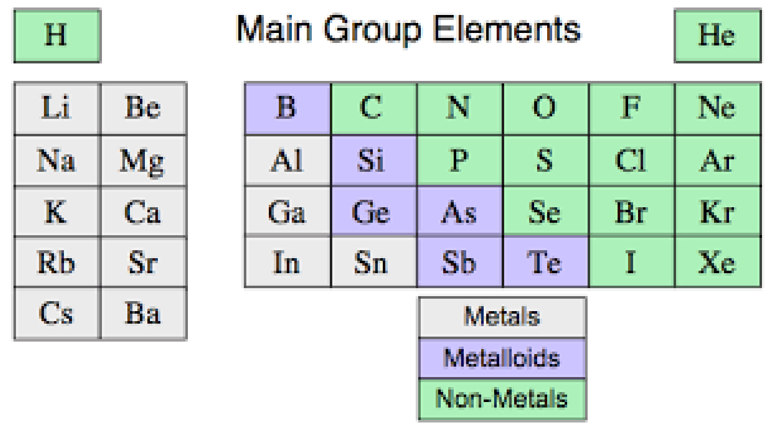


Because there are only two electrons in the s-shells, electron-electron binding is quite weak, so Alkaline Earths are still quite reactive even with filled s-shells.

There is therefore an impetus for electrons in the s-shells, to interact quite strongly with electrons in the p-shells, forming octaves (8 states)

The *p* Block

Among the orbital blocks of the periodic table, the p-block is the most diverse in its characteristics, and this may be explained by the way the p-states are filled up. The block is divided into three sections through a diagonal band. Elements above this band are non-metals, and those below are metals. Those in the middle are metalloids.



The chemical behavior is governed by two things – the atomic radius and the electrons in the p-states. Elements below the metalloids have large atomic radii and low number of p –electrons. The electrons are thus lightly bound to the nucleus, and these elements share similar characteristics as the metals to their left.

The elements above the metalloids have smaller atomic radii and are therefore more reactive. The p-states however lack more than one electron. As more electrons are needed, the chance of getting them is smaller.

Covalent Bonds

The solution to the "dilemma" is that instead of acquiring, these atoms tend to share electrons with other atoms that are similarly situated. The bond formed between such atoms is called covalent bond. In this way, Carbon, Nitrogen, Oxygen, Phosphorous and Sulfur easily bonds with other atoms, and together with Hydrogen, form molecules that are the building blocks of life

	13	14	15	16	17
2	B Boron	C Carbon	N Nitrogen	O Oxygen	F Fluorine
3	Al Aluminium	Si Silicon	P Phosphorus	S Sulfur	Cl Chlorine
4	Ga Gallium	Ge Germanium	As Arsenic	Se Selenium	Br Bromine
5	In Indium	Sn Tin	Sb Antimony	Te Tellurium	I Iodine
6	Tl Thallium	Pb Lead	Bi Bismuth	Po Polonium	At Astatine

Melting points (K)

covalent polyatomic molecule or network

diatomic molecules

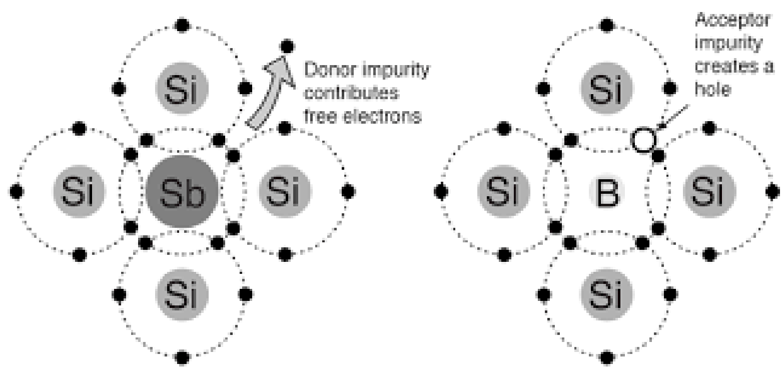
monoatomic

B 2349	C 3800	N 63	O 54	F 53.5	He 0.95
Al 933	Si 1687	P 317	S 388	Cl 172	Ne 24.6
Ga 303	Ge 1211	As 1090	Se 494	Br 266	Ar 83.8
In 430	Sn 505	Sb 904	Te 723	I 387	Kr 115.8
Tl 577	Pb 601	Bi 544	Po 527	At ?	Xe 161.4
					Rn 202

metallic

Elements of Group IVA are special because they have 4 electrons that can fill the 8 sp – states. As this is exactly half of the states, the covalent bonds they form with multiple atoms are very strong.

Metalloids



In its "pristine" state, that is a solid formed by atoms of the same element, the material is non-metallic. But once a small amount of atoms to the left or right of Group IVA is added, a process called doping, free charges are created.

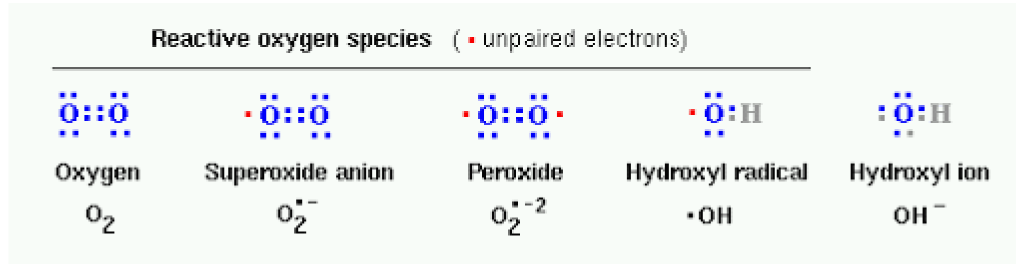
Since Group IIIA elements have only 3 *sp*-electrons, it is one less than what is needed to form closed shells, an empty space is made available for an electron to freely move into. This is called a hole, and is treated as a "positive charge" that can move from one atom to the next as an electron from the next atom moves in to fill the hole.

	13	14	15	16	17
2	B Boron	C Carbon	N Nitrogen	O Oxygen	F Fluorine
3	Al Aluminium	Si Silicon	P Phosphorus	S Sulfur	Cl Chlorine
4	Ga Gallium	Ge Germanium	As Arsenic	Se Selenium	Br Bromine
5	In Indium	Sn Tin	Sb Antimony	Te Tellurium	I Iodine
6	Tl Thallium	Pb Lead	Bi Bismuth	Po Polonium	At Astatine

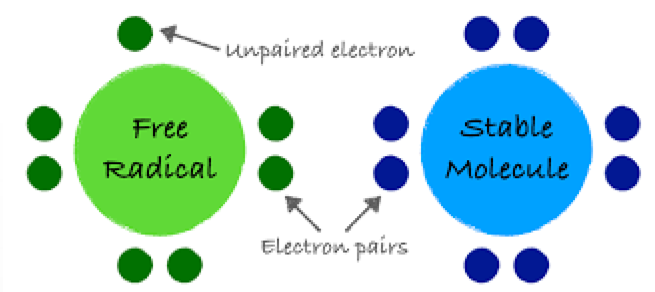
On the other a Group VA element has one electron more than what is needed for a closed shell, and this excess electron would be free to move around. Under these conditions, the substances behave like metals.

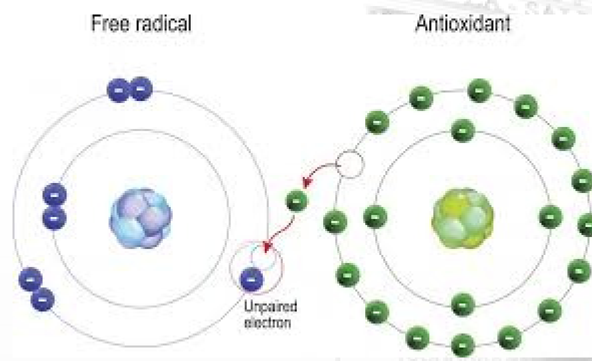
Free Radicals

Because Oxygen has 4 $p-$ electrons, it requires two more to form a closed shell. But two electrons are hard to come by at a given time.



Thus, even as they form molecules through covalent bonding, the molecules have empty spaces for electrons. And when there is an unpaired valence electron, the molecule is highly reactive. An atom, molecule, or ion that has an unpaired valence electron is called a free radical. It is for this reason that oxygen is a highly corrosive and flammable material. While radicals are necessary for life, they also induce damages of cells.





The *d* Block

Elements to the left of the $d-$block are the highly reactive Alkali and Alkaline metals. To the right are weakly reacting metals of the $p-$block. As there are 10 $d-$states, these transition metals show a variety of characteristics depending on the number of filled states.

Transition Elements

d - block

3	4	5	6	7	8	9	10	11	12
Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn
Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd
	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg
	Rf	Db	Sg	Bh	Hs	Mt			

Elements on the tenth column (Group 2B), Zinc, Cadmium and Mercury, have completely-filled $d-$shells, and are therefore non-reactive. Elements on the ninth column (Group 1B) are the coinage metals Copper, Silver, Gold as they are highly resistant to corrosion. Elements on the eighth column (Nickel, Palladium, Platinum) are excellent catalysts. Elements in the middle of the block (Manganese, Iron, Cobalt)are ferromagnetic materials. Elements on the left (Titanium, Vanadium, Chromium, Molybdenum) form durable steel when combined with Iron.

The *d* Block

With 9 electrons filling up 10 slots, elements in the ninth column like copper and gold have plenty of electrons that can move from one atom to the next, filling up the remaining empty spot. They therefore make very good conductors.

In the middle of the block, the $d-$ shells are half-occupied. As energy is lower when spins are aligned, elements like Iron have strong magnetic moments.

Elements on the left side of the block have few electrons. But when combined with the two $s-$ electrons, Titanium for example has 4 electrons that can bind with the 6 electrons of Iron, to form a very strong bond of 10 electrons.