

Inorganic Chemistry 1

ELECTRONIC CONFIGURATION

HOW TO WRITE ELECTRONIC CONFIGURATION FOR AN ATOM

Periodic Table of Elements

1	2																	10																	
1	H																	He																	
3	Li	4	Be																	10															
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	57	*La	72	Hf	73	Ta	74	W	75	Re	76	Os	77	Ir	78	Pt	79	Au	80	Hg	81	Tl	82	Pb	83	Bi	84	Po	85	At	86	Rn
87	Fr	88	Ra	89	+Ac	104	Rf	105	Ha	106	107	108	109	110																					

* Lanthanide Series

58	Ce	59	Pr	60	Nd	61	Pm	62	Sm	63	Eu	64	Gd	65	Tb	66	Dy	67	Ho	68	Er	69	Tm	70	Yb	71	Lu
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+ Actinide Series

90	Th	91	Pa	92	U	93	Np	94	Pu	95	Am	96	Cm	97	Bk	98	Cf	99	Es	100	Fm	101	Md	102	No	103	Lr
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Legend - click to find out more...

H - gas	Li - solid	Br - liquid	Tc - synthetic
Non-Metals	Transition Metals	Rare Earth Metals	Halogens
Alkali Metals	Alkali Earth Metals	Other Metals	Inert Elements

Step 1 : Determine the number of electrons in the atom.

- The atomic number (on the periodic table) is the number of protons of the atom, and thus equals the number of electrons in an atom with zero charge.
- If you're working with a charged atom, add or subtract electrons accordingly. For instance a sodium atom with a negative one charge would have another electron in addition to the atomic number of 11, making 12 electrons.

Step 2 : Memorize the basic list of orbitals.

Once you know the order of orbitals, you can simply fill them according to the number of electrons in the atom.

- The *s* orbital set (any number followed by an "s") contains a single orbital, and by Pauli's Exclusion Principle, a single orbital can hold a maximum of 2 electrons, so each *s* orbital set can hold 2 electrons.
- The *p* orbital set contains 3 orbitals, and thus can hold a total of 6 electrons
- The *d* orbital set contains 5 orbitals, so it can hold 10 electrons.
- The *f* orbital set contains 7 orbitals, so it can hold 14 electrons.
- Memorize the order of the orbitals. Note that orbital sets are numbered by electron shell, but ordered in terms of energy. For instance, a filled $4s^2$ is lower energy (or less potentially volatile) than a partially-filled or filled $3d^{10}$, so the 4s shell is listed first.

Step 3 : Put one electron into the highest energy orbital available starting with 1s (holds a maximum of two electrons).

Be careful!

Do not fill the orbitals in the order shown in the chart! Fill the orbitals in this order (the number following the orbital set is the maximum number of electrons it can hold):

- $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6 7s^2 5f^{14} 6d^{10} 7p^6$
- Note that the above list, if all the shells were filled, would be the electron configuration for Uuo (ununoctium), 118, the highest-numbered atom on the periodic table – so this electron configuration contains every currently known electron shell for a neutrally charged atom
- Note:
Energy level changes as you go up.
For example, when you are about to go up to the 4th energy level, it becomes 4s first, then 3d. After the fourth energy level, you'll move onto the 5th where it follows the order once again. This only happens after the 3rd energy level.

Step 4 : Fill in the orbitals according to the number of electrons in the atom

Tips For Exam

- The number following the letter is actually superscript, so do not make that mistake on a test.
- To find the atomic number of the atom when it is in electron configuration form, just add up all of the numbers that follow the letters (s, p, d, and f).
- Writing long electron configurations can be avoided by writing them in their noble gas configurations. Simply find the last symbol containing p⁶ (such as 3p⁶ or 5p⁶) and add up all of the numbers following the letters of every symbol before and including the p⁶ orbital set. Then, using the sum of the numbers, locate the element with the atomic number equal to the sum you just calculated. It should appear at the very right of the periodic table. That's called a noble gas. Now, just remove all of the symbols that you added up and put in the noble gas' symbol in brackets. So, for an antimony atom, the noble gas configuration would be [Kr] 5s² 4d¹⁰ 5p³. Notice that you don't add up the 5s² and 4d¹⁰ because they are after 4p⁶, which you do add.
- You can also write an element's electron configuration by just writing the valence configuration, which is the last s and p orbital set. So, the valence configuration of an antimony atom would be 5s² 5p³.

- When the atom is an ion, it means that the number of protons does not equal the number of electrons. The charge of the atom will then be displayed at the top right (usually) corner of the chemical symbol. So, an antimony atom with charge +2 has an electron configuration of $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^3$. Notice that the $5p^3$ changed into a $5p^1$. Be careful when the configuration of an uncharged atom ends in anything but an s and p orbital set. When you take away electrons, you can only take them away from the valence orbitals (the s and p orbitals). So, if a configuration ends in $4s^2 3d^7$, and the atom gains a charge of +2, then the configuration would change to end with $4s^0 3d^7$. Notice that $3d^7$ does not change, instead, the s orbital electrons are lost.
- There are circumstances when an electron needs to be "promoted." When an orbital set is one electron away from being half occupied or completely occupied, remove one electron from the nearest s or p orbital set and move it to the orbital set that needs the electron.
- Every atom desires to be stable, and the most stable configurations have full s and p (s^2 and p^6) orbital sets. The noble gases have this configuration, which is why they are rarely reactive and are on the right side of the periodic table. So, if a configuration ends in $3p^4$, it only needs two more electrons to become stable (losing six, including the s orbital set's electrons, takes more energy, so

losing four is easier). And if a configuration ends in $4d^3$, it only needs to lose three electrons to reach a stable state. Also, half filled shells (s^1 , p^3 , d^5 ...) are more stable than, for example p^4 or p^2 ; however, s^2 and p^6 will be even more stable.

- Ions aren't the same. They're much harder. Skip two levels and follow the same pattern depending on where you started depending on how high or how low the number of electrons is.
- There is no such thing as the "stability of a half-filled" sublevel. It is an oversimplification. Any stability relating to "half-filled" sublevels is due to the fact that each orbital is singly occupied, thus electron-electron repulsions are minimized.

Prepared by
V. Manoharan
vmano@usm.my
manovv1955@yahoo.com
04-6533888 ext 3566