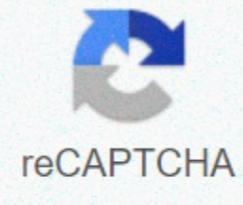




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## Electron configuration table

The commonly used long form of the periodic table is designed to emphasize electron configurations. Since it is the outermost (valens) electrons that are primarily involved in chemical interactions between atoms, the last electron added to an atom in the build-up process is of far greater interest to a chemist than the first. This last electron is called the characteristic electron because it separates an atom from that located immediately before it in the periodic table. The type of substation (s, p, d, f) in which the characteristic electron is located is very closely related to the chemical behaviour of an element and gives rise to the classification shown by the colour coding on the periodic table seen here. The representative elements are those where the characteristic electron enters or p subshell. Most of the elements whose chemistry and behavior we have discussed so far fall into this category. Many of the chemical properties of the representative elements can be explained on the basis of Lewis charts. That is, the valences of the representative elements can be predicted based on the number of valence electrons they have, or on the basis of the number of electrons that had to be added to achieve the same electron configuration as an atom of a noble gas. For representative elements, the number of valence electrons is the same as the periodic group number, and the number needed to match the next noble gas configuration is 8 minus the group number. This is consistent with valence rules derived from the periodic table, and results in formulas for chlorides of the first dozen elements, showing the periodic variation of valence. Element Atomic Weight Hydrogen Compounds Oxygen Compounds Chlorine Compounds Hydrogen 1.01 H<sub>2</sub> H<sub>2</sub>O, H<sub>2</sub>O<sub>2</sub> HCl Helium 4.00 No formed No formed Lithium 6.94 LiH Li<sub>2</sub>O, Li<sub>2</sub>O<sub>2</sub> LiCl Beryllium 9.01 BeH<sub>2</sub> BeO BeCl<sub>2</sub> Boron 10.81 B<sub>2</sub>H<sub>6</sub> B<sub>2</sub>O<sub>3</sub> BCl<sub>3</sub> Carbon 12.01 CH<sub>4</sub>, C<sub>2</sub>H<sub>6</sub>, C<sub>3</sub>H<sub>8</sub> CO<sub>2</sub>, CO, C<sub>2</sub>O<sub>3</sub> CCl<sub>4</sub>, C<sub>2</sub>Cl<sub>6</sub> Nitrogen 14.01 NH<sub>3</sub>, N<sub>2</sub>H<sub>4</sub>, HN<sub>3</sub> N<sub>2</sub>O, NO, NO<sub>2</sub>, N<sub>2</sub>O<sub>5</sub> NCl<sub>3</sub> Oxygen 16.00 H<sub>2</sub>O, H<sub>2</sub>O<sub>2</sub> O<sub>2</sub>, O<sub>3</sub> &lt;Cl<sub>2</sub>O, ClO<sub>2</sub>, Cl<sub>2</sub>O<sub>7</sub> Fluoro 19.00 HF OF<sub>2</sub>, O<sub>2</sub>F<sub>2</sub> ClF, ClF<sub>3</sub>, ClF<sub>5</sub> Neon 20.18 No formed None formed No formed Sodium 22.99 NaH Na<sub>2</sub>O, Na<sub>2</sub>O<sub>2</sub> NaCl Magnesium 24.31 MgH<sub>2</sub> MgO MgCl<sub>2</sub> The first three horizontal rows or periods of the modern periodic system consist entirely of representative elements. In the first period the characteristic electrons to H and He are in the 1s subshell. Across the second period Li and Be have distinguish electrons in 2s subshell, and electrons are being added to the 2p subshell in atoms from B to Ne. In the third period the subs to be filled for Na and Mg and therefore Al, Si, P, S, Cl and Ar. As a general rule, the characteristic electron associated with the representative elements will be in a ns or np subshell. The value of n, the most important can be quickly determined by counting down from the top of the periodic table. For example, iodine is a representative element in the fifth period. Therefore, the characteristic electron must occupy either the 5s or 5p subshell. Since I'm on the right side of the table, 5p is the right choice. When the primary quantum number is three or more, it is also possible to obtain subshells of the d-type. The transition elements or transition metals are those whose characteristic electron is found in a d orbital. The first examples of transitional metals (Sc, Ti, V, Cr, Mn, Fe, Co, Ni, Cu, Zn) are found in the fourth period, although the characteristic electron in each case is a 3d electron and belongs to the third shell. This interruption results, which we have already seen because 4s are lower in energy than 3d. 4s orbital thus begins to fill up, beginning the fourth period before any of the 3d orbitals can be occupied. Figure \(\PageIndex{1}\)) compares the probability distributions of a 4 and a 3d electron in a V atom. Although 4s electron cloud is farther from the nucleus on average than does 3d cloud, a small portion of 4s electron density is found very close to the nucleus, where it is barely shielded from the total nuclear charge of +23. It is the very strong attractive force of this small fraction of the total 4s electron density that lowers the energy of the 4s electron below 3d. Figure \(\PageIndex{1}\)) Comparison of 3d (crying) and 4s (red) electron clouds for a vanadium atom. The fact that 4s electron cloud is more extensive than 3d has an important influence on the chemistry of the transition elements. When an atom like V (Figure \(\PageIndex{1}\)) interacts with another atom, it is 4s electrons that stretch furthest away from the nucleus that first contacts the second atom. Thus, 4s electrons are often more significant than 3d in determining the valence and formulas of connections. 3d electrons are buried under the surface of atoms of transition metals. Adding a more 3d electron has significantly less effect on their chemical properties than adding a more 3s or 3p electron did in the case of the representative elements. Therefore, there is a slow but steady transition in properties from one transition element to another. Note, for example, that except sc, all the transitional metals form chlorides, MCl<sub>2</sub>, where the metal has a valence of 2; examples are TiCl<sub>2</sub>, VCl<sub>2</sub>, CrCl<sub>2</sub>, etc. This can be seen in the table at the top of this page. Valence of 2 corresponds to the two 4s roller electrons. Each of the transition metals also exhibits other valences, with one or more of the 3d electrons also involved. For example, in some compounds V (vanadium) has a valence of 2 (VO, VCl<sub>2</sub>) in others have a valence of 3 (V<sub>2</sub>O<sub>3</sub>, VCl<sub>3</sub>), in still others have a valence of 4 (VO<sub>2</sub>, and in at least one case (V<sub>2</sub>O<sub>5</sub>) it has a valence of 5. The chemistry of transition metals is more complicated and a wider range of formulas for transition-metal connections is possible due to this variable valence. In some cases electrons in d subshells act as valence electrons, while in other cases they do not. Although 3d electron clouds do not extend further from the nucleus than 3s and 3p (and thus do not form another shell as 4s electrons do), they are thoroughly shielded from the nuclear charge and thus often act as valens electrons. This Jekyll and Hyde behavior of 3d electrons makes life more complicated (and often much more interesting) for chemists who study transition elements. Table 1: Atomic Electron Configurations Z Element Configuration 1 H 1s 1 2 He 1s 2 3 Li [He] 2s 1 4 Be [He] 2s 2 5 B [He] 2s 2 2p 1 6 C [He] 2s 2 2p 2 7 N [He] 2s 2 2p 3 8 O [He] 2s 2 2p 4 9 F [He] 2s 2 2p 5 10 Ne [He] 2s 2 2p 6 11 Na [Ne] 3s 1 12 Mg [Ne] 3s 2 13 Al [Ne] 3s 2 3p 1 14 Si [Ne] 3s 2 3p 2 15 P [Ne] 3s 2 3p 3 16 S [Ne] 3s 2 3p 4 17 Cl [Ne] 3s 2 3p 5 18 Ar [Ne] 3s 2 3p 6 19 K [Ar] 4s 1 20 Ca [Ar] 4s 2 21 Sc [Ar] 3d 1 4s 2 22 Ti [Sc] 3d 2 4s 2 23 V [Ar] 3d 3 4s 2 24 Cr [Ar] 3d 5 4s 1 25 Mn [Sc] 3d 5 4s 2 26 Fe [Ar] 3d 6 4s 2 27 Co [Ar] 3d 7 4s 2 28 Ni [Ar] 3d 8 4s 2 29 Cu [Ar] 3d 10 4s 1 30 Zn [Ar] 3d 10 4s 2 31 Ga [Ar] 3d 10 4s 2 4p 1 32 Ge [Ar] 3d 10 4s 2 4p 2 33 As [Sc] 3d 10 4s 2 4p 3 34 Se [Ar] 3d 10 4s 2 4p 4 35 Br [Ar] 3d 10 4s 2 4p 5 36 Kr [Ar] 3d 10 4s 2 4p 6 37 Rb [Kr] 5s 1 38 Sr [Kr] 5s 2 39 Y [Kr] 4d 1 5s 2 40 Zr [Kr] 4d 2 5s 2 41 Nb [Kr] 4d 4 5s 1 42 Mo [Kr] 4d 5 5s 1 43 Tc [Kr] 4d 5 5s 2 44 Ru [Kr] 4d 7 5s 1 45 Rh [Kr] 4d 8 5s 1 46 Pd [Kr] 4d 10 47 Ag [Kr] 4d 10 48 Cd [Kr] 4d 10 49 In [Kr] 4d 10 50 Sn [Kr] 4d 10 5s 2 5p 2 51 Sb [Kr] 4d 10 5s 2 5p 3 52 Te [Kr] 4d 10 5s 2 5p 4 53 I [Kr] 4d 10 5s 2 5p 5 54 Xe [Kr] 4d 10 5s 2 5p 6 55 Cs [Xe] 6s 1 56 Ba [Xe] 6s 2 57 La [Xe] 5d 1 6s 2 58 Ce [Xe] 4f 1 5d 1 6s 2 59 Pr [Xe] 4f 3 6s 2 60 Nd [Xe] 4f 4 6s 2 61 Pm [Xe] 4f 5 6s 2 62 Sm [Xe] 4f 6 6s 2 63 Eu [Xe] 4f 7 6s 2 64 Gd [Xe] 4f 7 5d 1 6s 2 65 Tb [Xe] 4f 9 6s 2 66 Dy [Xe] 4f 10 6s 2 67 Ho [Xe] 4f 11 6s 2 68 Er [Xe] 4f 12 6s 2 69 Tm [Xe] 4f 13 6s 2 70 Yb [Xe] 4f 14 6s 2 71 Lu [Xe] 4f 14 5d 1 6s 2 72 Hf [Xe] 4f 14 5d 2 6s 2 73 Ta [Xe] 4f 14 5d 3 6s 2 74 W [Xe] 4f 14 5d 4 6s 2 75 Re [Xe] 4f 14 5d 5 6s 2 76 Os [Xe] 4f 14 5d 6 6s 2 77 Ir [Xe] 4f 14 5d 7 6s 2 78 Pt [Xe] 4f 14 5d 9 6s 1 79 Au [Xe] 4f 14 5d 10 6s 1 80 Hg [Xe] 4f 14 5d 10 6s 2 81 Tl [Xe] 4f 14 5d 10 6s 2 82 Pb [Xe] 4f 14 5d 10 6s 2 83 Bi [Xe] 4f 14 5d 10 6s 2 84 Po [Xe] 4f 14 5d 10 6s 2 85 At [Xe] 4f 14 5d 10 6s 2 86 Rn [Xe] 4f 14 5d 10 6s 2 87 Fr [Rn] 7s 1 88 Ra [Rn] 7s 2 89 Ac [Rn] 6d 1 7s 2 90 Th [Rn] 6d 2 7s 2 91 Pa [Rn] 5f 2 6d 1 7s 2 92 U [Rn] 5f 3 6d 1 7s 2 93 Np [Rn] 5f 4 6d 1 7s 2 94 Pu [Rn] 5f 6 7s 2 95 Am [Rn] 5f 7s 2 96 Cm [Rn] 5f 7 6d 1s 2 97 Bk [Rn] 5f 9s 2 98 Cf [Rn] 5f 10 s 2 99 Es [Rn] 5f 11 s 2 100 Fm [Rn] 5f 12 s 2 101 Md [Rn] 5f 13 s 2 102 No. [Rn] 5f 14 s 2 103 Lr [Rn] 5f 14 6d 1 s 2 104 Rf [Rn] 5f 14 6d 2 s 2 The third major category of items occurs when distinguishing electron occupies a f subshell. The first example occurs in the case of lanthanoids (elements with atomic numbers between 57 and 71). The lanthanoids have the general electron configuration [Kr]4d104f i5s25p65d0 or 16s2 where i am a number between 0 and 14. Thus, in the construction process for lanthanoids, electrons are added to a subshell (4f), whose main quantum number is two smaller than the outer orbital (6s). The addition of another electron to an inner shell buried as deep as 4f has little or no effect on the chemical properties of these elements. All are very similar to lanthan (La) and can fit into the exact same room in the periodic table as La. The lanthanoid elements are so similar that special techniques are needed to separate them. As a result, even some pure samples of most of them were first produced in the 1870s. After the element actinium (Ac) is a series of atoms in which 5f subshell is filling. The actinoids are somewhat less similar to Ac than lanthanoids are to La, because some exceptions to the usual order filling orbitals occur in the case of Th, Pa, and U (Table \(\PageIndex{1}\)) ). Figure \(\PageIndex{2}\)): Periodic system showing the subshells that distinguish between areas of the periodic table. Note: although it is found in the p-subshell area of the periodic table, the box for He is purple, not green, since the s-subshell is filled in He, not the p-subshell. Because lanthanoids and most of the actinoids behave chemically as if they were to fit into group IIIB of the periodic table (where Lu and Lr exist), both groups are separated from the rest of the table and placed together in a block below. Together, lanthanoids and actinoids are called internal transition elements because f undershells that are filled lie so deep inside the remaining electronic structure of their atoms. Figure \(\PageIndex{2}\)) summarizes the type of subshell in which the characteristic electron exists for atoms of elements in different areas of the periodic table. This summary information makes it relatively simple to use the periodic table to obtain electron configurations, as the following example shows. Example \(\PageIndex{1}\)): Electron Configuration Get electron configuration for (a) Nb; (b) Per Solution (a) Nb, element number 41, is found in the fifth period and in an area of the periodic table where a d-subshell is filling (the second transition series). Moving back (towards lower atomic numbers) through the periodic table, the nearest noble gas is Kr, and we use Kr kernel: Nb [Kr] \_\_\_\_\_ 5s2 for Nb configuration. We are now moving further straight into the 4d subshell region of the periodic table and counting over three spaces (Y, Zr, Nb) to reach Nb. Thus, the total electron configuration is Nb [Kr]4d35s2 (Note that the primary quantum number for the d-subshell is 4 – one less than the number of the period. Also, if you look at the table of electron configurations, it should be noted that Nb is an exception to the typical orbital filling rules) b) A similar procedure is followed for Pr, element number 59. Moving backwards through the table, the nearest noble gas is Xe, and then we use the Xe core. Counting forward again, Cs and Ba equate to 6s2. So La, Ce, and Pr correspond to three more electrons in the 4f subshell. The configuration is thus Pr ... [Xe]4f36s2 Another point must be emphasized about the relationship between electron configuration and the periodic table. The atoms of elements in the same vertical column in the table have similar electron configurations. Consider, for example, the following: Using our rules for emitting electron configurations (Example 1) we have Element Electron Configuration Lewis Diagram Be [He] 2s2 Be: Mg [Ne]3s2 Mg: Ca [Ar] 4s2 Ca: Sr [Kr]5s2 Sr: Ba [Xe]6s2 Ba: Ra [Rn]7s2 Ra: Thus similarities of chemical behavior and valence noted earlier for these elements correlate with the similarities between their outer electron clouds. Such similarities account for the success of Mendeleev's predictions about the properties of undiscovered elements. Contributors and Attributes Attributes Attributions