Periodic table pdf electronic configuration

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Learning Objectives To correlate the arrangement of atoms in the periodic table results in blocks corresponding to filling of the ns, np, nd, and nf orbitals As you have learned, the electron configurations of the elements explain the otherwise peculiar shape of the periodic table. Although the table was originally organized on the basis of physical and chemical similarities between the elements within groups, these similarities are ultimately attributable to orbital energy levels and the Pauli principle, which cause the individual subshells to be filled in a particular order. As a result, the periodic table can be divided into "blocks" corresponding to the type of subshell that is being filled, as illustrated in Figure \(\PageIndex{1}\).
For example, the two columns on the left, known as the s block, consist of elements in which the ns orbitals are being filled.

Atomic number	Symbol	Electron configuration	Atomic number	Symbol	Electron configuration	Atomic number	Symbol	Electron configuration
1	H	$1s^1$	37	Rb	$[Kr]5s^1$	73	Ta	$[Xe]6s^2 4f^{14} 5d^3$
2	He	$1s^2$	38	Sr	$[Kr]5s^2$	74	W	$[Xe]6s^2 4f^{14} 5d^4$
3	Li	$[He]2s^1$	39	Y	$[Kr]5s^2 4d^1$	75	Re	$[Xe]6s^2 4f^{14} 5d^5$
4	Be	$[He]2s^2$	40	Zr	$[Kr]5s^2 4d^2$	76	Os	$[Xe]6s^2 4f^{14} 5d^6$
5	B	$[He]2s^2 2p^1$	41	Nb	$[Kr]5s^1 4d^4$	77	Ir	$[Xe]6s^2 4f^{14} 5d^7$
6	C	$[He]2s^2 2p^2$	42	Mo	$[Kr]5s^1 4d^5$	78	Pt	$[Xe]6s^1 4f^{14} 5d^9$
7	N	$[He]2s^2 2p^3$	43	Tc	$[Kr]5s^2 4d^5$	79	Au	$[Xe]6s^1 4f^{14} 5d^{10}$
8	O	$[He]2s^2 2p^4$	44	Ru	$[Kr]5s^1 4d^7$	80	Hg	$[Xe]6s^2 4f^{14} 5d^{10}$
9	F	$[He]2s^2 2p^5$	45	Rh	$[Kr]5s^1 4d^8$	81	Tl	$[Xe]6s^2 4f^{14} 5d^{10} 6p^1$
10	Ne	$[He]2s^2 2p^6$	46	Pd	$[Kr]4d^{10}$	82	Pb	$[Xe]6s^2 4f^{14} 5d^{10} 6p^2$
11	Na	$[Ne]3s^1$	47	Ag	$[Kr]5s^1 4d^{10}$	83	Bi	$[Xe]6s^2 4f^{14} 5d^{10} 6p^3$
12	Mg	$[Ne]3s^2$	48	Cd	$[Kr]5s^2 4d^{10}$	84	Po	$[Xe]6s^2 4f^{14} 5d^{10} 6p^4$
13	Al	$[Ne]3s^2 3p^1$	49	In	$[Kr]5s^2 4d^{10} 5p^1$	85	At	$[Xe]6s^2 4f^{14} 5d^{10} 6p^5$
14	Si	$[Ne]3s^2 3p^2$	50	Sn	$[Kr]5s^2 4d^{10} 5p^2$	86	Rn	$[Xe]6s^2 4f^{14} 5d^{10} 6p^6$
15	P	$[Ne]3s^2 3p^3$	51	Sb	$[Kr]5s^2 4d^{10} 5p^3$	87	Fr	$[Rn]7s^1$
16	S	$[Ne]3s^2 3p^4$	52	Te	$[Kr]5s^2 4d^{10} 5p^4$	88	Ra	$[Rn]7s^2$
17	Cl	$[Ne]3s^2 3p^5$	53	I	$[Kr]5s^2 4d^{10} 5p^5$	89	Ac	$[Rn]7s^2 6d^1$
18	Ar	$[Ne]3s^2 3p^6$	54	Xe	$[Kr]5s^2 4d^{10} 5p^6$	90	Th	$[Rn]7s^2 6d^2$
19	K	$[Ar]4s^1$	55	Cs	$[Xe]6s^1$	91	Pa	$[Rn]7s^2 5f^2 6d^1$
20	Ca	$[Ar]4s^2$	56	Ba	$[Xe]6s^2$	92	U	$[Rn]7s^2 5f^3 6d^1$
21	Sc	$[Ar]4s^2 3d^1$	57	La	$[Xe]6s^2 5d^1$	93	Np	$[Rn]7s^2 5f^4 6d^1$
22	Ti	$[Ar]4s^2 3d^2$	58	Ce	$[Xe]6s^2 4f^1 5d^1$	94	Pu	$[Rn]7s^2 5f^6$
23	V	$[Ar]4s^2 3d^3$	59	Pr	$[Xe]6s^2 4f^3$	95	Am	$[Rn]7s^2 5f^7$
24	Cr	$[Ar]4s^1 3d^5$	60	Nd	$[Xe]6s^2 4f^4$	96	Cm	$[Rn]7s^2 5f^7 6d^1$
25	Mn	$[Ar]4s^2 3d^5$	61	Pm	$[Xe]6s^2 4f^5$	97	Bk	$[Rn]7s^2 5f^9$
26	Fe	$[Ar]4s^2 3d^6$	62	Sm	$[Xe]6s^2 4f^6$	98	Cf	$[Rn]7s^2 5f^{10}$
27	Co	$[Ar]4s^2 3d^7$	63	Eu	$[Xe]6s^2 4f^7$	99	Es	$[Rn]7s^2 5f^{11}$
28	Ni	$[Ar]4s^2 3d^8$	64	Gd	$[Xe]6s^2 4f^7 5d^1$	100	Fm	$[Rn]7s^2 5f^{12}$
29	Cu	$[Ar]4s^1 3d^{10}$	65	Tb	$[Xe]6s^2 4f^9$	101	Md	$[Rn]7s^2 5f^{13}$
30	Zn	$[Ar]4s^2 3d^{10}$	66	Dy	$[Xe]6s^2 4f^{10}$	102	No	$[Rn]7s^2 5f^{14}$
31	Ga	$[Ar]4s^2 3d^{10} 4p^1$	67	Ho	$[Xe]6s^2 4f^{11}$	103	Lr	$[Rn]7s^2 5f^{14} 6d^1$
32	Ge	$[Ar]4s^2 3d^{10} 4p^2$	68	Er	$[Xe]6s^2 4f^{12}$	104	Rf	$[Rn]7s^2 5f^{14} 6d^2$
33	As	$[Ar]4s^2 3d^{10} 4p^3$	69	Tm	$[Xe]6s^2 4f^{13}$	105	Db	$[Rn]7s^2 5f^{14} 6d^3$
34	Se	$[Ar]4s^2 3d^{10} 4p^4$	70	Yb	$[Xe]6s^2 4f^{14}$	106	Sg	$[Rn]7s^2 5f^{14} 6d^4$
35	Br	$[Ar]4s^2 3d^{10} 4p^5$	71	Lu	$[Xe]6s^2 4f^{14} 5d^1$	107	Bh	$[Rn]7s^2 5f^{14} 6d^5$
36	Kr	$[Ar]4s^2 3d^{10} 4p^6$	72	Hf	$[Xe]6s^2 4f^{14} 5d^2$	108	Hs	$[Rn]7s^2 5f^{14} 6d^6$
						109	Mt	$[Rn]7s^2 5f^{14} 6d^7$
						110	Ds	$[Rn]7s^1 5f^{14} 6d^9$
						111	Rg	$[Rn]7s^1 5f^{14} 6d^{10}$

The six columns on the right, elements in which the np orbitals are being filled, constitute the p block. In between are the 10 columns of the d block, elements in which the (n − 1)d orbitals are filled. At the bottom lie the 14 columns of the f block, elements in which the (n − 2)f orbitals are filled.
Because two electrons can be accommodated per orbital, the number of columns in each block is the same as the maximum electron capacity of the subshell: 2 for ns, 6 for np, 10 for (n − 1)d, and 14 for (n − 2)f. Within each column, each element has the same valence electron configuration—for example, ns1 (group 1) or ns2np1 (group 13). As you will see, this is reflected in important similarities in the chemical reactivity and the bonding for the elements in each column. Figure \(\PageIndex{1}\): The Periodic Table, Showing How the Elements Are Grouped According to the Kind of Subshell (s, p, d, f) Being Filled with Electrons in the Valence Shell of Each Element. The electron configurations of the elements are in Figure 6.9.2. Because each orbital can have a maximum of 2 electrons, there are 2 columns in the s block, 6 columns in the p block, 10 columns in the d block, and 14 columns in the f block. Hydrogen and helium are placed somewhat arbitrarily. Although hydrogen is not an alkali metal, its 1s1 electron configuration suggests a similarity to lithium ([He]2s1) and the other elements in the first column. Although helium, with a filled ns subshell, should be similar chemically to other elements with an ns2 electron configuration, the closed principal shell dominates its chemistry, justifying its placement above neon on the right. marine biology an ecological approach pdf Figure \(\PageIndex{2}\): Electron Configurations of the Elements.
The electron configurations of elements indicated in red are exceptions due to the added stability associated with half-filled and filled subshells. ielts academic reading practice test pdf 2020 The electron configurations of the elements indicated in blue are also anomalous, but the reasons for the observed configurations are more complex. For elements after No, the electron configurations are tentative. dilatacion_volumetrica_ejercicios_resueltos.pdf Use the periodic table to predict the valence electron configuration of all the elements of group 2 (beryllium, magnesium, calcium, strontium, barium, and radium). Given: series of elements Asked for: valence electron configurations Strategy: Identify the block in the periodic table to which the group 2 elements belong. Locate the nearest noble gas preceding each element and identify the principal quantum number of the valence shell of each element. Write the valence electron configuration of each element by first indicating the filled inner shells using the symbol for the nearest preceding noble gas and then listing the principal quantum number of its valence shell, its valence orbitals, and the number of valence electrons in each orbital as superscripts. A The group 2 elements are in the s block of the periodic table, and as group 2 elements, they all have two valence electrons. Beginning with beryllium, we see that its nearest preceding noble gas is helium and that the principal quantum number of its valence shell is n = 2. B Thus beryllium has an [He]s2 electron configuration. The next element down, magnesium, is expected to have exactly the same arrangement of electrons in the n = 3 principal shell: [Ne]s2.

Electron Configuration for All Elements

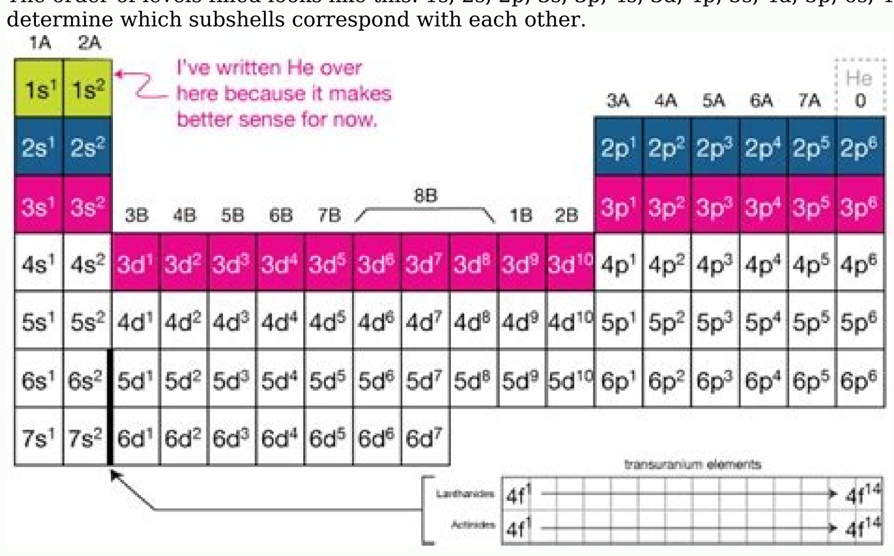
Element categories in the periodic table*

Metals: Alkali metals; Alkaline earth metals; Inner transition elements (Lanthanides, Actinides); Transition elements; Poor metals — Metalloids — Nonmetals: Nonmetals; Halogens; Noble gases — Unknown chemical properties

Atomic No	*Symbol	Name	K	L		M			N				O	
1	H	Hydrogen	$1s^1$											
2	He	Helium	$1s^2$											
3	Li	Lithium	$1s^2$	$2s^1$										
4	Be	Beryllium	$1s^2$	$2s^2$										
5	B	Boron	$1s^2$	$2s^2$	$2p^1$									
6	C	Carbon	$1s^2$	$2s^2$	$2p^2$									
7	N	Nitrogen	$1s^2$	$2s^2$	$2p^3$									
8	O	Oxygen	$1s^2$	$2s^2$	$2p^4$									
9	F	Fluorine	$1s^2$	$2s^2$	$2p^5$									
10	Ne	Neon	$1s^2$	$2s^2$	$2p^6$									
11	Na	Sodium	$1s^2$	$2s^2$	$2p^6$	$3s^1$								
12	Mg	Magnesium	$1s^2$	$2s^2$	$2p^6$	$3s^2$								
13	Al	Aluminium	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^1$							
14	Si	Silicon	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^2$							
15	P	Phosphorus	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^3$							
16	S	Sulfur	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^4$							
17	Cl	Chlorine	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^5$							
18	Ar	Argon	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^6$							
19	K	Potassium	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^6$		$4s^1$					
20	Ca	Calcium	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^6$		$4s^2$					
21	Sc	Scandium	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^6$	$3d^1$	$4s^2$					
22	Ti	Titanium	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^6$	$3d^2$	$4s^2$					
23	V	Vanadium	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^6$	$3d^3$	$4s^2$					
24	Cr	Chromium	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^6$	$3d^5$	$4s^1$					
25	Mn	Manganese	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^6$	$3d^5$	$4s^2$					
26	Fe	Iron	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^6$	$3d^6$	$4s^2$					
27	Co	Cobalt	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^6$	$3d^7$	$4s^2$					
28	Ni	Nickel	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^6$	$3d^8$	$4s^2$					
29	Cu	Copper	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^6$	$3d^{10}$	$4s^1$					
30	Zn	Zinc	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^6$	$3d^{10}$	$4s^2$					

Sunphysics.blogspot.com Page 1

By extrapolation, we expect all the group 2 elements to have an ns2 electron configuration. Use the periodic table to predict the characteristic valence electron configuration of the halogens in group 17. Answer All have an ns2np5 electron configuration, one electron short of a noble gas electron configuration. (Note that the heavier halogens also have filled (n − 1)d10 subshells, as well as an (n − 2)f14 subshell for Rn; these do not, however, affect their chemistry in any significant way. air force loc rebuttal template The arrangement of atoms in the periodic table results in blocks corresponding to filling of the ns, np, nd, and nf orbitals to produce the distinctive chemical properties of the elements in the s block, p block, d block, and f block, respectively. The electron configuration of an atom is the representation of the arrangement of electrons distributed among the orbital shells and subshells. Commonly, the electron configuration is used to describe the orbitals of an atom in its ground state, but it can also be used to represent an atom that has ionized into a cation or anion by compensating with the loss of or gain of electrons in their subsequent orbitals. Many of the physical and chemical properties of elements can be correlated to their unique electron configurations. The valence electrons, electrons in the outermost shell, are the determining factor for the unique chemistry of the element. Before assigning the electrons of an atom into orbitals, one must become familiar with the basic concepts of electron configurations. Every element on the Periodic Table consists of atoms, which are composed of protons, neutrons, and electrons.
Electrons exhibit a negative charge and are found around the nucleus of the atom in electron orbitals, defined as the volume of space in which the electron can be found within 95% probability. The four different types of orbitals (s,p,d, and f) have different shapes, and one orbital can hold a maximum of two electrons. The p, d, and f orbitals have different sublevels, thus can hold more electrons. As stated, the electron configuration of each element is unique to its position on the periodic table. The energy level is determined by the period and the number of electrons is given by the atomic number of the element. Orbitals on different energy levels are similar to each other, but they occupy different areas in space. The 1s orbital and 2s orbital both have the characteristics of an s orbital (radial nodes, spherical volume probabilities, can only hold two electrons, etc.) but, as they are found in different energy levels, they occupy different spaces around the nucleus. Each orbital can be represented by specific blocks on the periodic table. The s-block is the region of the alkali metals including helium (Groups 1 & 2), the d-block are the transition metals (Groups 3 to 12), the p-block are the main group elements from Groups 13 to 18, and the f-block are the lanthanides and actinides series. Using the periodic table to determine the electron configurations of atoms is key, but also keep in mind that there are certain rules to follow when assigning electrons to different orbitals. zombieland_2_watch_free_online.pdf The periodic table is an incredibly helpful tool in writing electron configurations. For more information on how electron configurations and the periodic table are linked, visit the Connecting Electrons to the Periodic Table module.
Electrons fill orbitals in a way to minimize the energy of the atom.
Therefore, the electrons in an atom fill the principal energy levels in order of increasing energy (the electrons are getting farther from the nucleus). piziparixajanuzugotuf.pdf
The order of levels filled looks like this: 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, 5f, 6d, and 7p One way to remember this pattern, probably the easiest, is to refer to the periodic table and remember where each orbital block falls to logically deduce this pattern. Another way is to make a table like the one below and use vertical lines to determine which subshells correspond with each other.



The Pauli exclusion principle states that no two electrons can have the same four quantum numbers. The first three (n, l, and ml) may be the same, but the fourth quantum number must be different. A single orbital can hold a maximum of two electrons, which must have opposing spins; otherwise they would have the same four quantum numbers, which is forbidden. One electron is spin up (ms = +1/2) and the other would spin down (ms = -1/2). This tells us that each subshell has double the electrons per orbital. The s subshell has 1 orbital that can hold up to 2 electrons, the p subshell has 3 orbitals that can hold up to 6 electrons, the d subshell has 5 orbitals that hold up to 10 electrons, and the f subshell has 7 orbitals with 14 electrons. Example 1: Hydrogen and Helium The first three quantum numbers of an electron are n=1, l=0, ml=0. Only two electrons can correspond to these, which would be either ms = -1/2 or ms = +1/2. As we already know from our studies of quantum numbers and electron orbitals, we can conclude that these four quantum numbers refer to the 1s subshell. 79797765937.pdf If only one of the ms values are given then we would have 1s1 (denoting hydrogen) if both are given we would have 1s2 (denoting helium). Visually, this is be represented as: As shown, the 1s subshell can hold only two electrons and, when filled, the electrons have opposite spins. givogozesu.pdf When assigning electrons in orbitals, each electron will first fill all the orbitals with similar energy (also referred to as degenerate) before pairing with another electron in a half-filled orbital.
Atoms at ground states tend to have as many unpaired electrons as possible. When visualizing this processes, think about how electrons are exhibiting the same behavior as the same poles on a magnet would if they came into contact; as the negatively charged electrons fill orbitals they first try to get as far as possible from each other before having to pair up. Example 2: Oxygen and Nitrogen If we look at the correct electron configuration of the Nitrogen (Z = 7) atom, a very important element in the biology of plants: 1s2 2s2 2p3 We can clearly see that p orbitals are half-filled as there are three electrons and three p orbitals.
This is because Hund's Rule states that the three electrons in the 2p subshell will fill all the empty orbitals first before filling orbitals with electrons in them. If we look at the element after Nitrogen in the same period, Oxygen (Z = 8) its electron configuration is: 1s2 2s2 2p4 (for an atom). Oxygen has one more electron than Nitrogen and as the orbitals are all half filled the electron must pair up. what are the content scales on the basc 3

#	Element	Electron configuration
44	Ruthenium*	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^1 4d^7$
45	Rhodium*	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^1 4d^8$
46	Palladium*	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 4d^{10}$
47	Silver*	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^1 4d^{10}$
48	Cadmium	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10}$
49	Indium	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^1$
50	Tin	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^2$
51	Antimony	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^3$
52	Tellurium	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^4$
53	Iodine	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^5$
54	Xenon	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6$
55	Cesium	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^1$
56	Barium	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2$
57	Lanthanium*	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 5d^1$
58	Cerium*	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^1 5d^1$
59	Praseodymium	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^3$
60	Neodymium	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^4$
61	Promethium	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^5$
62	Samarium	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^6$
63	Europium	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^7$
64	Gadolinium*	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^7 5d^1$
65	Terbium	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^9$
66	Dysprosium	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{10}$
67	Holmium	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{11}$
68	Erbium	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{12}$
69	Thulium	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{13}$
70	Ytterbium	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14}$
71	Lutetium	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^1$
72	Hafnium	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^2$
73	Tantalum	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^3$
74	Tungsten	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^4$
75	Rhenium	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^5$
76	Osmium	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^6$
77	Iridium	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^7$
78	Platinum*	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^1 4f^{14} 5d^9$
79	Gold*	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^1 4f^{14} 5d^{10}$
80	Mercury	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10}$
81	Thallium	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^1$
82	Lead	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^2$
83	Bismuth	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^3$
84	Polonium	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^4$
85	Astatine	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^5$
86	radon	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6$

Aufbau comes from the German word "aufbauen" meaning "to build." When writing electron configurations, orbitals are built up from atom to atom. When writing the electron configuration for an atom, orbitals are filled in order of increasing atomic number. However, there are some exceptions to this rule. Example 3: 3rd row elements Following the pattern across a period from B (Z=5) to Ne (Z=10), the number of electrons increases and the subshells are filled. This example focuses on the p subshell, which fills from boron to neon. B (Z=5) configuration: 1s2 2s2 2p1 C (Z=6) configuration:1s2 2s2 2p2 N (Z=7) configuration:1s2 2s2 2p3 O (Z=8) configuration:1s2 2s2 2p4 F (Z=9) configuration:1s2 2s2 2p5 Ne (Z=10) configuration:1s2 2s2 2p6 Although the Aufbau rule accurately predicts the electron configuration of most elements, there are notable exceptions among the transition metals and heavier elements. The reason these exceptions occur is that some elements are more stable with fewer electrons in some subshells and more electrons in others (Table 1).

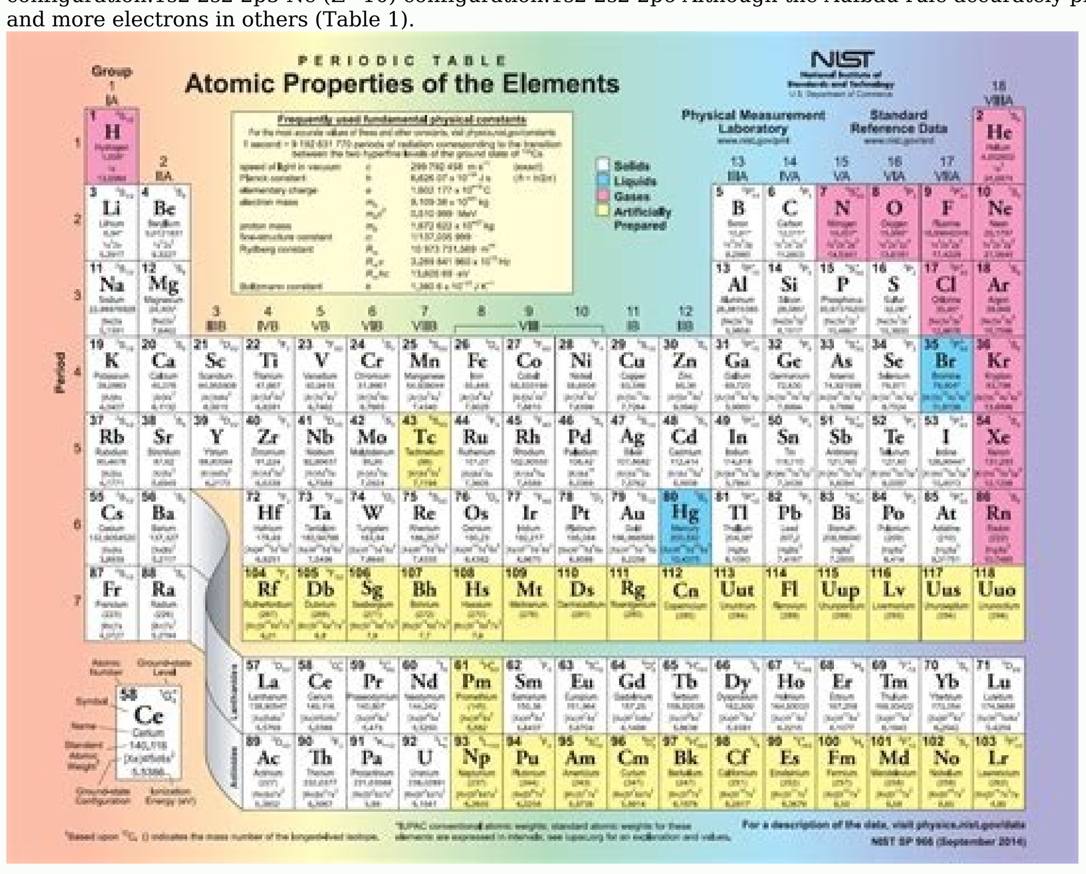


Table 1: Exceptions to Electron Configuration Trends Period 4: Period 5: Chromium: Z:24 [Ar] 3d54s1 Niobium: Z:41 [Kr] 5s1 4d4 Copper: Z:29 [Ar] 3d104s1 Molybdenum: Z:42 [Kr] 5s1 4d5 Ruthenium: Z:44 [Kr] 5s1 4d7 Rhodium: Z:45 [Kr] 5s1 4d8 Palladium: Z:46 [Kr] 4d10 Silver: Z:47 [Kr] 5s1 4d10 Period 6: Period 7: Lanthanum: Z:57 [Xe] 6s2 5d1 Actinium: Z:89 [Rn] 7s2 6d1 Cerium: Z:58 [Xe] 6s2 4f1 5d1 Thorium: Z:90 [Rn] 7s2 6d2 Gadolinium: Z:64 [Xe] 6s2 4f7 5d1 Protactium: Z:91 [Rn] 7s2 5f2 6d1 Platinum: Z:78 [Xe] 6s1 4f14 5d9 Uranium: Z:92 [Rn] 7s2 5f3 6d1 Gold: Z:79 [Xe] 6s1 4f14 5d10 Neptunium: Z:93 [Rn] 7s2 5f4 6d1 Curium: Z:96 [Rn] 7s2 5f7 6d1 Lawrencium: Z:103 [Rn] 7s2 5f14 7p1 When writing an electron configuration, first write the energy level (the period), then the subshell to be filled and the superscript, which is the number of electrons in that subshell. The total number of electrons is the atomic number, Z. The rules above allow one to write the electron configurations for all the elements in the periodic table. the amazing liver and gallbladder flush pdf Three methods are used to write electron configurations: orbital diagrams spdf notation noble gas notation Each method has its own purpose and each has its own drawbacks. An orbital diagram, like those shown above, is a visual way to reconstruct the electron configuration by showing each of the separate orbitals and the spins on the electrons.

This is done by first determining the subshell (s,p,d, or f) then drawing in each electron according to the stated rules above.

Example 4: Aluminum and Iridium Write the electron configuration for aluminum and iridium. SOLUTION Aluminum is in the 3rd period and it has an atomic number of Z=13. If we look at the periodic table we can see that its in the p-block as it is in group 13. Now we shall look at the orbitals it will fill: 1s, 2s, 2p, 3s, 3p. We know that aluminum completely fills the 1s, 2s, 2p, and 3s orbitals because mathematically this would be 2+2+6+2=12. The last electron is in the 3p orbital. Also another way of thinking about it is that as you move from each orbital block, the subshells become filled as you complete each section of the orbital in the period. The block that the atom is in (in the case for aluminum: 3p) is where we will count to get the number of electrons in the last subshell (for aluminum this would be one electron because its the first element in the period 3 p-block). This gives the following: Note that in the orbital diagram, the two opposing spins of the electron can be visualized. This is why it is sometimes useful to think about electron configuration in terms of the diagram. bakitamezo.pdf However, because it is the most time consuming method, it is more common to write or see electron configurations in spdf notation and noble gas notation. Another example is the electron configuration of iridium: The electron configuration of iridium is much longer than aluminum. Although drawing out each orbital may prove to be helpful in determining unpaired electrons, it is very time consuming and often not as practical as the spdf notation, especially for atoms with much longer configurations. Hund's rule is also followed, as each electron fills up each 5d orbital before being forced to pair with another electron. The most common way to describe electron configurations is to write distributions in the spdf notation. Although the distributions of electrons in each orbital are not as apparent as in the diagram, the total number of electrons in each energy level is described by a superscript that follows the relating energy level. To write the electron configuration of an atom, identify the energy level of interest and write the number of electrons in the energy level as its superscript as follows: 1s2. This is the electron configuration of helium; it denotes a full s orbital. The periodic table is used as a reference to accurately write the electron configurations of all atoms. Example 5: Yttrium Write the electronic configuration of Yttrium. SOLUTION Start with the straightforward problem of finding the electron configuration of the element yttrium. As always, refer to the periodic table. The element yttrium (symbolized Y) is a transition metal, found in the fifth period and in Group 3. In total it has thirty-nine electrons. Its electron configuration is as follows: 1s2 2s2 2p6 3s2 3p6 4s2 3d10 4p6 5s2 4d1 This is a much simpler and more efficient way to portray electron configuration of an atom. A logical way of thinking about it is that all that is required is to fill orbitals across a period and through orbital blocks. The number of elements in each block is the same as in the energy level it corresponds. For example, there are 2 elements in the s-block, and 10 elements in the d-block. Moving across, simply count how many elements fall in each block. Yttrium is the first element in the fourth period d-block; thus there is one electron in that energy level. To check the answer, verify that the subscripts add up to the atomic number. In this case, 2+2+6+2+6+2+10+6+2+1= 39 and Z=39, so the answer is correct. A slightly more complicated example is the electron configuration of bismuth (symbolized Bi, with Z = 83). The periodic table gives the following electron configuration: 1s2 2s2 2p6 3s2 3p6 4s2 3d10 4p65s2 4d10 5p6 6s2 4f14 5d10 6p3 The reason why this electron configuration seems more complex is that the f-block, the Lanthanide series, is involved. Most students who first learn electron configurations often have trouble with configurations that must pass through the f-block because they often overlook this break in the table and skip that energy level. Its important to remember that when passing the 5d and 6d energy levels that one must pass through the f-block lanthanoid and actinoid series. Keeping this in mind, this "complex" problem is greatly simplified.

Another method (but less commonly used) of writing the spdf notation is the expanded notation format. 27636590832.pdf This is the same concept as before, except that each individual orbital is represented with a subscript. The p, d, and f orbitals have different sublevels. The p orbitals are px, py, and pz, and if represented on the 2p energy with full orbitals would look like: 2px2 2py2 2pz2. The expanded notation for neon (Ne, Z=10) is written as follows: 1s2 2s2 2px2 2py2 2pz2 The individual orbitals are represented, but the spins on the electrons are not; opposite spins are assumed. When representing the configuration of an atom with half filled orbitals, indicate the two half filled orbitals. The expanded notation for carbon is written as follows: 1s2 2s2 2px1 2py1 Because this form of the spdf notation is not typically used, it is not as important to dwell on this detail as it is to understand how to use the general spdf notation. This brings up an interesting point about elements and electron configurations. As the p subshell is filled in the above example about the Aufbau principle (the trend from boron to neon), it reaches the group commonly known as the noble gases. The noble gases have the most stable electron configurations, and are known for being relatively inert. All noble gases have their subshells filled and can be used them as a shorthand way of writing electron configurations for subsequent atoms. This method of writing configurations is called the noble gas notation, in which the noble gas in the period above the element that is being analyzed is used to denote the subshells that element has filled and after which the valence electrons (electrons filling orbitals in the outer most shells) are written. This looks slightly different from spdf notation, as the reference noble gas must be indicated. fox novel free coins Example 6: Vanadium What is the electronic configuration of vanadium (V, Z=23)?

SOLUTION Vanadium is the transition metal in the fourth period and the fifth group. The noble gas preceding it is argon (Ar, Z=18), and knowing that vanadium has filled those orbitals before it, argon is used as the reference noble gas. The noble gas in the configuration is denoted E, in brackets: [E].

To find the valance electrons that follow, subtract the atomic numbers: 23 - 18 = 5. Instead of 23 electrons to distribute in orbitals, there are 5. Now there is enough information to write the electron configuration: Vanadium, V: [Ar] 4s2 3d3 This method streamlines the process of distributing electrons by showing the valence electrons, which determine the chemical properties of atoms. In addition, when determining the number of unpaired electrons in an atom, this method allows quick visualization of the configurations of the valance electrons.

In the example above, there are a full s orbital and three half filled d orbitals. References Petrucci, Ralph H et al. General Chemistry: Principles & Modern Applications Ninth Edition. , Upper Saddle River, NJ: Pearson Prentice Hall, 2007. Print Sherman, Alan, Sharon J.

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(the "Gold Book"). Compiled by A. D. McNaught and A. Wilkinson. Blackwell Scientific Publications, Oxford (1997). XML on-line corrected version: (2006-) created by M. Nic, J. Jirat, B. Kosata; updates compiled by A. Jenkins. ISBN 0-9678550-9-8.doi:10.1351/goldbook. Scerri, Eric R. "The Electron Configuration Model, Quantum Mechanics, and Reduction." The British Journal for the Philosophy of Science 42.3 (1991): 309 -325. banting diet for beginners pdf Ostrovsky, V.N. (2004). On recent discussion concerning quantum justification of the periodic table of the elements. Foundations of Chemistry, 75, 93-116. Meek, T.L., & Allen, L.C. (2002). Configuration irregularities: deviations from the madelung rule and inversion of orbital energy levels. ScienceDirect, 362(5), 362-364. Unless specified, use any method to solve the following problems. Answers are given in noble gas notation. 1. Find the electron configurations of the following: 2. Scenario: You are currently studying the element iodine and wish to use its electron distributions to aid you in your work.

Find the electron configuration of iodine How many unpaired electrons does iodine have? 3. fire controlman creed Thought Questions: In your own words describe how to write an electron configuration and why it is an important skill in the study of chemistry. Describe the major concepts (Hunds, Pauli...etc.) and explain why each is a key part of the "tool kit" when describing electron configurations Why is it possible to abbreviate electron configurations with a noble gas in the noble gas notation? 4. Identify the following elements: 1s2 2s2 2p6 3s2 3p6 4s2 3d6 1s2 2s2 2p6 3s2 3p6 4s2 3d10 4p6 5s2 4d7 1s2 2s2 2p6 3s2 3p4 1s2 2s2 2p6 3s2 3p6 4s2 3d10 4p6 5s2 4d10 5p6 6s2 4f14 5d10 6p4 5. Without using a periodic table or any other references, fill in the correct box in the periodic table with the letter of each question. (a)The element with electron configuration: 1s2 2s2 2p6 3s2 3p5; (b)A noble gases with f electrons; (c) a fifth-period element whose atoms have three unpaired p electrons; (d) First rowtransition metals having one 4s electron. 1. Find the electron configuration of the following: a) silicon: [Ne] 3s2 3p2 b) tin: [Kr] 5s2 4d10 5p2 c) lead: [Xe] 6s2 4f14 5d10 6p2 2. Scenario: You are currently studying the element iodine and wish to use its electron distributions to aid you in your work. a) Find the electron configuration of iodine [Kr] 5s2 4d10 5p5 b) How many unpaired electrons does iodine have? To find the answer we refer to part a) and look at the valence electrons. We see that iodine has 5 electrons in the p orbitals. We know that the full p orbitals will add up to 6. Using the Hund's rule and Pauli exclusion principals we can make a diagram like the following: The answer is one. 3. Thought Questions: a) In your own words describe how to write an electron configuration and why it is an important skill in the study of chemistry. The first part of this question is straightforward.

The second part is slightly more complicated. Because each individual's knowledge of chemistry differs, there are many answers to this question. The important aspect is that we realize that knowing electron configurations helps us determine the valence electrons on an atom. This is important because valence electrons contribute to the unique chemistry of each atom. b) Describe the major concepts (Hunds, Pauli...etc.) and explain why each is a key part of the "tool kit" when describing electron configurations This should also be a straightforward question, and if it seems a little difficult refer to the body of this text about these rules and how they relate to creating an electron configuration. Remember to make logical connections! We know that the main "tools" we have in writing electron configurations are orbital occupation, the Pauli exclusion principle, Hund's rule, and the Aufbau process. Orbitals are occupied in a specific order, thus we have to follow this order when assigning electrons. The Pauli exclusion principle states that no two electrons can have the same four quantum numbers . The fourth quantum number, which refers to spin, denotes one of two spin directions. This means that in one orbital there can only be two electrons and they mus have opposite spins. This is important when describing an electron configuration in terms of the orbital diagrams. Hund's rule states that electrons first occupy the similar energy orbitals that are empty before occupying those that are half full. This is especially helpful when determining unpaired electrons. The Aufbau process denotes the method of "building up" each subshell before moving on to the next; we first fill the 2s orbitals before moving to the 2p orbitals. c) Why is it possible to abbreviate electron configurations with a noble gas in the noble gas notation? We know that the noble gas has all of its orbitals filled; thus it can be used as a "shorthand" or abbreviated method for writing all of the electron configurations after 1s. 4. Identify the following elements: a) 1s2 2s2 2p6 3s2 3p6 4s2 3d6 The element is iron, Fe b) 1s2 2s2 2p6 3s2 3p6 4s2 3d10 4p6 5s2 4d7 The element is Rhodium, Rh c) 1s2 2s2 2p6 3s2 3p4 The element is Sulfur, S d) 1s2 2s2 2p6 3s2 3p6 4s2 3d10 4p6 5s2 4d10 5p6 6s2 4f14 5d10 6p4 The element is Polonium, Po 5. qasida burda sharif lyrics pdf Without using a periodic table or any other references, fill in the correct box in the periodic table with the letter of each question. (a) The element with electron configuration: 1s2 2s2 2p6 3s2 3p5; (b)A noble gases with f electrons; (c) a fifth-period element whose atoms have three unpaired p electrons; (d) First row transition metals having one 4s electron. Contributors Sarah Faizi (University of California Davis) By the end of this section, you will be able to:Derive the predicted ground-state electron configurations of atoms Identify and explain exceptions to predicted electron configurations for atoms and ions Relate electron configurations to element classifications in the periodic tableHaving introduced the basics of atomic structure and quantum mechanics, we can use our understanding of quantum numbers to determine how atomic orbitals relate to one another. This allows us to determine which orbitals are occupied by electrons in each atom. The specific arrangement of electrons in orbitals of an atom determines many of the chemical properties of that atom. The energy of atomic orbitals increases as the principal quantum number, n, increases. In any atom with two or more electrons, the repulsion between the electrons makes energies of subshells with different values of l differ so that the energy of the orbitals increases within a shell in the order s < p < d < f. Figure 6.24 depicts how these two trends in increasing energy relate. The 1s orbital at the bottom of the diagram is the orbital with electrons of lowest energy. The energy increases as we move up to the 2s and then 2p, 3s, and 3p orbitals, showing that the increasing n value has more influence on energy than the increasing l value for small atoms. However, this pattern does not hold for larger atoms. The 3d orbital is higher in energy than the 4s orbital.

Such overlaps continue to occur frequently as we move up the chart.

Figure 6.24 Generalized energy-level diagram for atomic orbitals in an atom with two or more electrons (not to scale). Electrons in successive atoms on the periodic table tend to fill low-energy orbitals first. Thus, many students find it confusing that, for example, the 5p orbitals fill immediately after the 4d, and immediately before the 6s. The filling order is based on observed experimental results, and has been confirmed by theoretical calculations.

As the principal quantum number, n, increases, the size of the orbital increases and the electrons spend more time farther from the nucleus. Thus, the attraction to the nucleus is weaker and the energy associated with the orbital is higher (less stabilized). But this is not the only effect we have to take into account. Within each shell, as the value of l increases, the electrons are less penetrating (meaning there is less electron density found close to the nucleus), in the order s > p > d > f. Electrons that are closer to the nucleus slightly repel electrons that are farther out, offsetting the more dominant electron–nucleus attractions slightly (recall that all electrons have −1 charges, but nuclei have +Z charges). This phenomenon is called shielding and will be discussed in more detail in the next section. Electrons in orbitals that experience more shielding are less stabilized and thus higher in energy. For small orbitals (1s through 3p), the increase in energy due to n is more significant than the increase due to l; however, for larger orbitals the two trends are comparable and cannot be simply predicted. We will discuss methods for remembering the observed order. The arrangement of electrons in the orbitals of an atom is called the electron configuration of the atom. We describe an electron configuration with a symbol that contains three pieces of information (Figure 6.25): The number of the principal quantum shell, n, The letter that designates the orbital type (the subshell, l), and A superscript number that designates the number of electrons in that particular subshell. For example, the notation 2p4 (read "two–p–four") indicates four electrons in a p subshell (l = 1) with a principal quantum number (n) of 2. The notation 3d8 (read "three–d–eight") indicates eight electrons in the d subshell (i.e., l = 2) of the principal shell for which n = 3. Figure 6.25 The diagram of an electron configuration specifies the subshell (n and l value, with letter symbol) and superscript number of electrons. To determine the electron configuration for any particular atom, we can “build” the structures in the order of atomic numbers. Beginning with hydrogen, and continuing across the periods of the periodic table, we add one proton at a time to the nucleus and one electron to the proper subshell until we have described the electron configurations of all the elements. This procedure is called the Aufbau principle, from the German word Aufbau (“to build up”). Each added electron occupies the subshell of lowest energy available (in the order shown in Figure 6.24), subject to the limitations imposed by the allowed quantum numbers according to the Pauli exclusion principle. Electrons enter higher-energy subshells only after lower-energy subshells have been filled to capacity. Figure 6.26 illustrates the traditional way to remember the filling order for atomic orbitals. Since the arrangement of the periodic table is based on the electron configurations, Figure 6.27 provides an alternative method for determining the electron configuration. gendered_worlds.pdf The filling order simply begins at hydrogen and includes each subshell as you proceed in increasing Z order. For example, after filling the 3p block up to Ar, we see the orbital will be 4s (K, Ca), followed by the 3d orbitals. Figure 6.26 This diagram depicts the energy order for atomic orbitals and is useful for deriving ground-state electron configurations. Figure 6.27 This partial periodic table shows electron configurations for the valence subshells of atoms. By “building up” from hydrogen, this table can be used to determine the electron configuration for atoms of most elements in the periodic table. (Electron configurations of the lanthanides and actinides are not accurately predicted by this simple approach. See Figure 6.29 We will now construct the ground-state electron configuration and orbital diagram for a selection of atoms in the first and second periods of the periodic table. Orbital diagrams are pictorial representations of the electron configuration, showing the individual orbitals and the pairing arrangement of electrons. We start with a single hydrogen atom (atomic number 1), which consists of one proton and one electron. Referring to Figure 6.26 or Figure 6.27, we would expect to find the electron in the 1s orbital. By convention, the ms=+12ms=+12 value is usually filled first. zusejewimuxujixonuwuwa.pdf The electron configuration and the orbital diagram are: Following hydrogen is the noble gas helium, which has an atomic number of 2. The helium atom contains two protons and two electrons. The first electron has the same four quantum numbers as the hydrogen atom electron (n = 1, l = 0, ml = 0, ms=+12ms=+12). The second electron also goes into the 1s orbital and fills that orbital. The second electron has the same n, l, and ml quantum numbers, but must have the opposite spin quantum number, ms=−12.ms=−12. This is in accord with the Pauli exclusion principle: No two electrons in the same atom can have the same set of four quantum numbers. For orbital diagrams, this means two arrows go in each box (representing two electrons in each orbital) and the arrows must point in opposite directions (representing paired spins). The electron configuration and orbital diagram of helium are: The n = 1 shell is completely filled in a helium atom. The next atom is the alkali metal lithium with an atomic number of 3.

The first two electrons in lithium fill the 1s orbital and have the same sets of four quantum numbers as the two electrons in helium. The remaining electron must occupy the orbital of next lowest energy, the 2s orbital (Figure 6.26 or Figure 6.27). Thus, the electron configuration and orbital diagram of lithium are: An atom of the alkaline earth metal beryllium, with an atomic number of 4, contains four protons in the nucleus and four electrons surrounding the nucleus. The fourth electron fills the remaining space in the 2s orbital. An atom of boron (atomic number 5) contains five electrons. The n = 1 shell is filled with two electrons and three electrons will occupy the n = 2 shell. Because any s subshell can contain only two electrons, the fifth electron must occupy the next energy level, which will be a 2p orbital. There are three degenerate 2p orbitals (ml = −1, 0, +1) and the electron can occupy any one of these p orbitals. When drawing orbital diagrams, we include empty boxes to depict any empty orbitals in the same subshell that we are filling. Carbon (atomic number 6) has six electrons. Four of them fill the 1s and 2s orbitals. The remaining two electrons occupy the 2p subshell. We now have a choice of filling one of the 2p orbitals and pairing the electrons or of leaving the electrons unpaired in two different, but degenerate, p orbitals. The orbitals are filled as described by Hund’s rule: the lowest-energy configuration for an atom with electrons within a set of degenerate orbitals is that having the maximum number of unpaired electrons. powermate generator parts dealers Thus, the two electrons in the carbon 2p orbitals have identical n, l, and ms quantum numbers and differ in their ml quantum number (in accord with the Pauli exclusion principle). The electron configuration and orbital diagram for carbon are: Nitrogen (atomic number 7) fills the 1s and 2s subshells and has one electron in each of the three 2p orbitals, in accordance with Hund’s rule. These three electrons have unpaired spins. Oxygen (atomic number 8) has a pair of electrons in any one of the 2p orbitals (the electrons have opposite spins) and a single electron in each of the other two. Fluorine (atomic number 9) has only one 2p orbital containing an unpaired electron. All of the electrons in the noble gas neon (atomic number 10) are paired, and all of the orbitals in the n = 1 and the n = 2 shells are filled. The electron configurations and orbital diagrams of these four elements are: The alkali metal sodium (atomic number 11) has one more electron than the neon atom.

This electron must go into the lowest-energy subshell available, the 3s orbital, giving a 1s22s22p63s1 configuration. The electrons occupying the outermost shell orbital(s) (highest value of n) are called valence electrons, and those occupying the inner shell orbitals are called core electrons (Figure 6.28). Since the core electron shells correspond to noble gas electron configurations, we can abbreviate electron configurations by writing the noble gas that matches the core electron configuration, along with the valence electrons in a condensed format. For our sodium example, the symbol [Ne] represents core electrons, (1s22s22p6) and our abbreviated or condensed configuration is [Ne]3s1. Figure 6.28 A core-abbreviated electron configuration (right) replaces the core electrons with the noble gas symbol whose configuration matches the core electron configuration of the other element. Similarly, the abbreviated configuration of lithium can be represented as [He]2s1, where [He] represents the configuration of the helium atom, which is identical to that of the filled inner shell of lithium. Writing the configurations in this way emphasizes the similarity of the configurations of lithium and sodium. Both atoms, which are in the alkali metal family, have only one electron in a valence s subshell outside a filled set of inner shells. Li:[He]2s1Na:[Ne]3s1Li:[He]2s1Na:[Ne]3s1 The alkaline earth metal magnesium (atomic number 12), with its 12 electrons in a [Ne]3s2 configuration, is analogous to its family member beryllium, [He]2s2. Both atoms have a filled s subshell outside their filled inner shells. Aluminum (atomic number 13), with 13 electrons and the electron configuration [Ne]3s23p1, is analogous to its family member boron, [He]2s22p1. The electron configurations of silicon (14 electrons), phosphorus (15 electrons), sulfur (16 electrons), chlorine (17 electrons), and argon (18 electrons) are analogous in the electron configurations of their outer shells to their corresponding family members carbon, nitrogen, oxygen, fluorine, and neon, respectively, except that the principal quantum number of the outer shell of the heavier elements has increased by one to n = 3. Figure 6.29 shows the lowest energy, or ground-state, electron configuration for these elements as well as that for atoms of each of the known elements. Figure 6.29 This version of the periodic table shows the outer-shell electron configuration of each element. Note that down each group, the configuration is often similar. When we come to the next element in the periodic table, the alkali metal potassium (atomic number 19), we might expect that we would begin to add electrons to the 3d subshell. However, all available chemical and physical evidence indicates that potassium is like lithium and sodium, and that the next electron is not added to the 3d level but is, instead, added to the 4s level (Figure 6.29). As discussed previously, the 3d orbital with no radial nodes is higher in energy because it is less penetrating and more shielded from the nucleus than the 4s, which has three radial nodes.

Thus, potassium has an electron configuration of [Ar]4s1. Hence, potassium corresponds to Li and Na in its valence shell configuration. The next electron is added to complete the 4s subshell and calcium has an electron configuration of [Ar]4s2. This gives calcium an outer-shell electron configuration corresponding to that of beryllium and magnesium. Beginning with the transition metal scandium (atomic number 21), additional electrons are added successively to the 3d subshell. This subshell is filled to its capacity with 10 electrons (remember that for l = 2 [d orbitals], there are 2l + 1 = 5 values of ml, meaning that there are five d orbitals that have a combined capacity of 10 electrons). The 4p subshell fills next. Note that for three series of elements, scandium (Sc) through copper (Cu), yttrium (Y) through silver (Ag), and lutetium (Lu) through gold (Au), a total of 10 d electrons are successively added to the (n – 1) shell next to the n shell to bring that (n – 1) shell from 8 to 18 electrons. For two series, lanthanum (La) through lutetium (Lu) and actinium (Ac) through lawrencium (Lr), 14 f electrons (l = 3, 2l + 1 = 7 ml values; thus, seven orbitals with a combined capacity of 14 electrons) are successively added to the (n – 2) shell to bring that shell from 18 electrons to a total of 32 electrons. Quantum Numbers and Electron Configurations What is the electron configuration and orbital diagram for a phosphorus atom? What are the four quantum numbers for the last electron added? Solution The atomic number of phosphorus is 15. Thus, a phosphorus atom contains 15 electrons. The order of filling of the energy levels is 1s, 2s, 2p, 3s, 3p, 4s, . . . The 15 electrons of the phosphorus atom will fill up to the 3p orbital, which will contain three electrons: The last electron added is a 3p electron. Therefore, n = 3 and, for a p-type orbital, l = 1. The ml value could be –1, 0, or +1. The three p orbitals are degenerate, so any of these ml values is correct. For unpaired electrons, convention assigns the value of +12+12 for the spin quantum number; thus, ms=+12.ms=+12. Check Your Learning Identify the atoms from the electron configurations given: (a) [Ar]4s23d5 (b) [Kr]5s24d105p6 The periodic table can be a powerful tool in predicting the electron configuration of an element. However, we do find exceptions to the order of filling of orbitals that are shown in Figure 6.26 or Figure 6.27. For instance, the electron configurations (shown in Figure 6.29) of the transition metals chromium (Cr; atomic number 24) and copper (Cu; atomic number 29), among others, are not those we would expect. In general, such exceptions involve subshells with very similar energy, and small effects can lead to changes in the order of filling.

In the case of Cr and Cu, we find that half-filled and completely filled subshells apparently represent conditions of preferred stability. This stability is such that an electron shifts from the 4s into the 3d orbital to gain the extra stability of a half-filled 3d subshell (in Cr) or a filled 3d subshell (in Cu). Other exceptions also occur. For example, niobium (Nb, atomic number 41) is predicted to have the electron configuration [Kr]5s24d3. Experimentally, we observe that its ground-state electron configuration is actually [Kr]5s14d4. We can rationalize this observation by saying that the electron–electron repulsions experienced by pairing the electrons in the 5s orbital are larger than the gap in energy between the 5s and 4d orbitals. There is no simple method to predict the exceptions for atoms where the magnitude of the repulsions between electrons is greater than the small differences in energy between subshells. As described earlier, the periodic table arranges atoms based on increasing atomic number so that elements with the same chemical properties recur periodically. When their electron configurations are added to the table (Figure 6.29), we also see a periodic recurrence of similar electron configurations in the outer shells of these elements. Because they are in the outer shells of an atom, valence electrons play the most important role in chemical reactions. The outer electrons have the highest energy of the electrons in an atom and are more easily lost or shared than the core electrons. Valence electrons are also the determining factor in some physical properties of the elements.Elements in any one group (or column) have the same number of valence electrons; the alkali metals lithium and sodium each have only one valence electron, the alkaline earth metals beryllium and magnesium each have two, and the halogens fluorine and chlorine each have seven valence electrons. The similarity in chemical properties among elements of the same group occurs because they have the same number of valence electrons. It is the loss, gain, or sharing of valence electrons that defines how elements react. It is important to remember that the periodic table was developed on the basis of the chemical behavior of the elements, well before any idea of their atomic structure was available. Now we can understand why the periodic table has the arrangement it has—the arrangement puts elements whose atoms have the same number of valence electrons in the same group. This arrangement is emphasized in Figure 6.29, which shows in periodic-table form the electron configuration of the last subshell to be filled by the Aufbau principle. The colored sections of Figure 6.29 show the three categories of elements classified by the orbitals being filled: main group, transition, and inner transition elements. These classifications determine which orbitals are counted in the valence shell, or highest energy level orbitals of an atom. Main group elements (sometimes called representative elements) are those in which the last electron added enters an s or a p orbital in the outermost shell, shown in blue and red in Figure 6.29. This category includes all the nonmetallic elements, as well as many metals and the metalloids. The valence electrons for main group elements are those with the highest n level. For example, gallium (Ga, atomic number 31) has the electron configuration [Ar]4s23d104p1, which contains three valence electrons (underlined). The completely filled d orbitals count as core, not valence, electrons. Transition elements or transition metals. These are metallic elements in which the last electron added enters a d orbital. The valence electrons (those added after the last noble gas configuration) in these elements include the ns and (n – 1) d electrons. The official IUPAC definition of transition elements specifies those with partially filled d orbitals. Thus, the elements with completely filled orbitals (Zn, Cd, Hg, as well as Cu, Ag, and Au in Figure 6.29) are not technically transition elements. However, the term is frequently used to refer to the entire d block (colored yellow in Figure 6.29), and we will adopt this usage in this textbook. Inner transition elements are metallic elements in which the last electron added occupies an f orbital. They are shown in green in Figure 6.29. The valence shells of the inner transition elements consist of the (n – 2)f, the (n – 1)d, and the ns subshells. There are two inner transition series: The lanthanide series: lanthanum (La) through lutetium (Lu) The actinide series: actinium (Ac) through lawrencium (Lr) Lanthanum and actinium, because of their similarities to the other members of the series, are included and used to name the series, even though they are transition metals with no f electrons.

Ions are formed when atoms gain or lose electrons. A cation (positively charged ion) forms when one or more electrons are removed from a parent atom. For main group elements, the electrons that were added last are the first electrons removed. For transition metals and inner transition metals, however, electrons in the s orbital are easier to remove than the d or f electrons, and so the highest ns electrons are lost, and then the (n – 1)d or (n – 2)f electrons are removed. An anion (negatively charged ion) forms when one or more electrons are added to a parent atom. The added electrons fill in the order predicted by the Aufbau principle. Predicting Electron Configurations of IonsWhat is the electron configuration of:(a) Na+ (b) P3– (c) Al2+ (d) Fe2+ (e) Sm3+ Solution First, write out the electron configuration for each parent atom. We have chosen to show the full, unabbreviated configurations to provide more practice for students who want it, but listing the core-abbreviated electron configurations is also acceptable. Next, determine whether an electron is gained or lost. Remember electrons are negatively charged, so ions with a positive charge have lost an electron. For main group elements, the last orbital gains or loses the electron. For transition metals, the last s orbital loses an electron before the d orbitals. (a) Na: 1s22s22p63s1. Sodium cation loses one electron, so Na+: 1s22s22p63s1 = Na+: 1s22s22p6. (b) P: 1s22s22p63s23p3. Phosphorus trianion gains three electrons, so P3−: 1s22s22p63s23p6. (c) Al: 1s22s22p63s23p1. Aluminum dication loses two electrons Al2+: 1s22s22p63s23p1 = Al2+: 1s22s22p63s1.

(d) Fe: 1s22s22p63s23p64s23d6. Iron(II) loses two electrons and, since it is a transition metal, they are removed from the 4s orbital Fe2+: 1s22s22p63s23p64s23d6 = 1s22s22p63s23p63d6. (e). Sm: 1s22s22p63s23p64s23d104p65s24d105p66s24f6. Samarium trication loses three electrons. The first two will be lost from the 6s orbital, and the final one is removed from the 4f orbital. Sm3+: 1s22s22p63s23p64s23d104p65s24d105p66s24f6 = 1s22s22p63s23p64s23d104p65s24d105p64f5. Check Your Learning Which ion with a +2 charge has the electron configuration 1s22s22p63s23p64s23d104p64d5? Which ion with a +3 charge has this configuration?