

# ELECTRON CONFIGURATION OF ATOMS

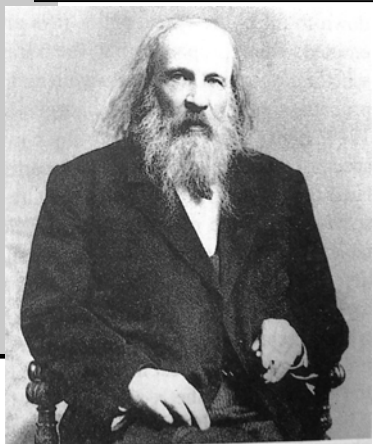
**Figure 3.11** A periodic table of partial ground-state electron configurations. These ground-state electron configurations show the electrons beyond the previous noble gas in the sublevel block being filled (excluding filled inner sublevels). For main-group elements, the group heading identifies the general outer configuration. Anomalous electron configurations occur often among the *d*-block and *f*-block elements, with the first two appearing for Cr (*Z* = 24) and Cu (*Z* = 29). Helium is colored as an *s*-block element but placed with the other members of Group 8A(18). Configurations for elements 112, 114, and 116 have not yet been confirmed.

Main-Group Elements ( <i>s</i> block)		Transition Elements ( <i>d</i> block)																Main-Group Elements ( <i>p</i> block)									
1A (1)	2A (2)	3B (3)	4B (4)	5B (5)	6B (6)	7B (7)	8 (8)	9 (9)	10 (10)	1B (11)	2B (12)	3A (13)	4A (14)	5A (15)	6A (16)	7A (17)	8A (18)										
1 H $1s^1$	2 He $1s^2$	Transition Elements ( <i>d</i> block)																3 Li $2s^1$	4 Be $2s^2$	5 B $2s^2 2p^1$	6 C $2s^2 2p^2$	7 N $2s^2 2p^3$	8 O $2s^2 2p^4$	9 F $2s^2 2p^5$	10 Ne $2s^2 2p^6$		
3 Na $3s^1$	4 Mg $3s^2$	5 Sc $3d^1 4s^2$	6 Ti $3d^2 4s^2$	7 V $3d^3 4s^2$	8 Cr $3d^5 4s^1$	9 Mn $3d^5 4s^2$	10 Fe $3d^6 4s^2$	11 Co $3d^7 4s^2$	12 Ni $3d^8 4s^2$	13 Cu $3d^{10} 4s^1$	14 Zn $3d^{10} 4s^2$	15 Ga $4s^2 4p^1$	16 Ge $4s^2 4p^2$	17 As $4s^2 4p^3$	18 Se $4s^2 4p^4$	19 Br $4s^2 4p^5$	20 Kr $4s^2 4p^6$										
5 Rb $5s^1$	6 Sr $5s^2$	7 Y $4d^1 5s^2$	8 Zr $4d^2 5s^2$	9 Nb $4d^4 5s^1$	10 Mo $4d^5 5s^1$	11 Tc $4d^5 5s^2$	12 Ru $4d^7 5s^1$	13 Rh $4d^8 5s^1$	14 Pd $4d^{10}$	15 Ag $4d^{10} 5s^1$	16 Cd $4d^{10} 5s^2$	17 In $5s^2 5p^1$	18 Sn $5s^2 5p^2$	19 Sb $5s^2 5p^3$	20 Te $5s^2 5p^4$	21 I $5s^2 5p^5$	22 Xe $5s^2 5p^6$										
7 Fr $7s^1$	8 Ra $7s^2$	9 Ac $6d^1 7s^2$	10 Th $6d^2 7s^2$	11 Pa $5f^2 6d^1 7s^2$	12 U $5f^3 6d^1 7s^2$	13 Np $5f^4 6d^1 7s^2$	14 Pu $5f^6 7s^2$	15 Am $5f^7 7s^2$	16 Cm $5f^7 7s^2$	17 Bk $6d^1 7s^2$	18 Cf $6d^2 7s^2$	19 Es $6d^3 7s^2$	20 Fm $6d^4 7s^2$	21 Md $6d^5 7s^2$	22 No $6d^6 7s^2$	23 Lr $6d^7 7s^2$	24 Lu $4f^14 5d^1 6s^2$										
		Inner Transition Elements ( <i>f</i> block)																									
		*Lanthanides																									
		58 Ce $4f^1 5d^1 6s^2$	59 Pr $4f^3 6s^2$	60 Nd $4f^4 6s^2$	61 Pm $4f^5 6s^2$	62 Sm $4f^6 6s^2$	63 Eu $4f^7 6s^2$	64 Gd $4f^7 5d^1 6s^2$	65 Tb $4f^9 6s^2$	66 Dy $4f^10 6s^2$	67 Ho $4f^11 6s^2$	68 Er $4f^12 6s^2$	69 Tm $4f^13 6s^2$	70 Yb $4f^14 6s^2$	71 Lu $4f^14 5d^1 6s^2$												
		**Actinides																									
		90 Th $6d^2 7s^2$	91 Pa $5f^2 6d^1 7s^2$	92 U $5f^3 6d^1 7s^2$	93 Np $5f^4 6d^1 7s^2$	94 Pu $5f^6 7s^2$	95 Am $5f^7 7s^2$	96 Cm $5f^7 7s^2$	97 Bk $6d^1 7s^2$	98 Cf $6d^2 7s^2$	99 Es $6d^3 7s^2$	100 Fm $6d^4 7s^2$	101 Md $6d^5 7s^2$	102 No $6d^6 7s^2$	103 Lr $6d^7 7s^2$												

## Electron Configuration ?

- is the distribution of electrons within the orbitals of its atoms, in relation with chemical and physical properties
- Objectives:**
  - to show how the organization of the table,
  - condensed from countless hour of laboratory works,
  - which was explained by quantum-mechanical atomic model

# Development of The Periodic Table



Mendeleev's Great Contribution

- Earliest organizing attempt: Johann Döbereiner (mid 19<sup>th</sup> century)
- At 1870, Dmitri Mendeleev ==> 65 elements ==> periodic table

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**Table 8.1** Mendeleev's Predicted Properties of Germanium ("eka Silicon") and Its Actual Properties

Property	Predicted Properties of eka Silicon (E)	Actual Properties of Germanium (Ge)
Atomic mass	72 amu	72.61 amu
Appearance	Gray metal	Gray metal
Density	5.5 g/cm <sup>3</sup>	5.32 g/cm <sup>3</sup>
Molar volume	13 cm <sup>3</sup> /mol	13.65 cm <sup>3</sup> /mol
Specific heat capacity	0.31 J/g·K	0.32 J/g·K
Oxide formula	EO <sub>2</sub>	GeO <sub>2</sub>
Oxide density	4.7 g/cm <sup>3</sup>	4.23 g/cm <sup>3</sup>
Sulfide formula and solubility	ES <sub>2</sub> ; insoluble in H <sub>2</sub> O; soluble in aqueous (NH <sub>4</sub> ) <sub>2</sub> S	GeS <sub>2</sub> ; insoluble in H <sub>2</sub> O; soluble in aqueous (NH <sub>4</sub> ) <sub>2</sub> S
Chloride formula (boiling point)	ECl <sub>4</sub> (<100°C)	GeCl <sub>4</sub> (84°C)
Chloride density	1.9 g/cm <sup>3</sup>	1.844 g/cm <sup>3</sup>
Element preparation	Reduction of K <sub>2</sub> EF <sub>6</sub> with sodium	Reduction of K <sub>2</sub> GeF <sub>6</sub> with sodium

## Characteristics of Many-Electron Atoms

- The Schrödinger equation does not give exact solutions for many-electron atoms
- Unlike the Bohr model, the Schrödinger equation gives very good *approximate solutions*.
- that, the atomic orbitals of many-electron atoms are hydrogen-like
- ==> we can use the same quantum numbers for the H atom to describe the orbitals of other atoms.
- Three features that were not relevant with the case of H atom:
  - a) the need for a fourth quantum number
  - b) a limit on the number of electrons allowed in a given orbital
  - c) a more complex set of orbital energy levels.

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## The Electron-Spin Quantum Number

- An additional quantum number is needed to describe a property of the electrons itself, called **SPIN**, which is not a property of the orbital.
- Electron spin becomes important when more than one electron is present
- Like its charge, spin is an intrinsic property of the electron, and the **spin quantum number** ( $m_s$ ) has values of either  $+\frac{1}{2}$  or  $-\frac{1}{2}$ .
- Therefore: *each electron in an atom is described completely by a set of four quantum numbers; the first three describe its orbital, and the fourth describes its spin.*
- Look at the following table:

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**Table 8.2 Summary of Quantum Numbers of Electrons in Atoms**

Name	Symbol	Permitted Values	Property
Principal	$n$	Positive integers (1, 2, 3, ...)	Orbital energy (size)
Angular momentum	$l$	Integers from 0 to $n - 1$	Orbital shape (The $l$ values 0, 1, 2, and 3 correspond to $s$ , $p$ , $d$ , and $f$ orbitals, respectively.)
Magnetic	$m_l$	Integers from $-l$ to 0 to $+l$	Orbital orientation
Spin	$m_s$	$+\frac{1}{2}$ or $-\frac{1}{2}$	Direction of $e^-$ spin

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■ Example:

- Hydrogen atom (H,  $Z=1$ )  $\implies n=1, l=0, m_l=0$ , and  $m_s=+\frac{1}{2}$
- (By convention,  $+\frac{1}{2}$  should be assigned as the first electron in an orbital rather than  $-\frac{1}{2}$ )

■ **The Exclusion Principle (Pauli's Principle):**

- Helium (He,  $Z=2$ ): the first electron in the He ground state has the same set of quantum numbers as that in the H atom, but the second He electron does not
- Wolfgang Pauli  $\implies$  observe the excited states of atoms
- Exclusion Principle: "*no two electrons in the same atom can have the same four quantum numbers*"
- That is, each electron must have a unique identity
- $\implies$  the second He electron occupies the same orbital as the first but has an opposite spin  $\implies n=1, l=0, m_l=0, m_s=-\frac{1}{2}$

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- Spin quantum number ( $m_s$ ) can have only two values ( $+\frac{1}{2}$  or  $-\frac{1}{2}$ ), therefore: “an atomic orbital can hold a maximum of two electrons and they must have opposing spins”
- or otherwise: the 1s orbital in He is filled, and that the electrons have paired spins.
- In the quantum – mechanical model:



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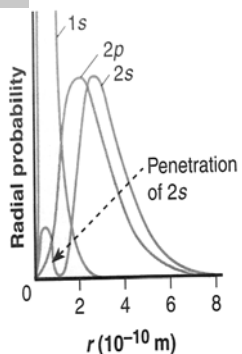
## Electrostatic Effects and Energy-Level Splitting

- Electrostatic effects play a major role on the energy states of many-electron atoms
- The energy states of many-electron atoms arise not only from nucleus-electron attractions, but also electron-electron repulsions.
- The electrons of an atom in its ground state occupy the orbitals of lowest energy
- Therefore, the first two electrons in the ground state of Li fill its 1s orbital. Then, the third Li electron must go into the  $n=2$  level. But, this level has 2s and 2p sublevels
- Which does Li's third electron enter?
- You'll see that the 2s is lower in energy than the 2p
- Why? Why?



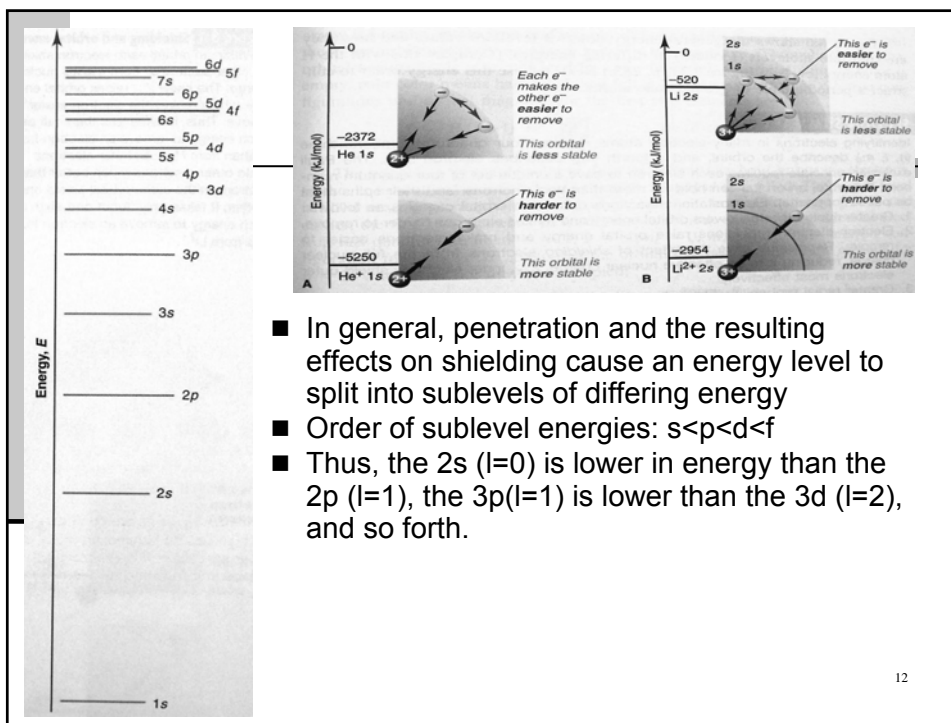
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## Why the 2s orbital is occupied rather than 2p in Li atom?



- The electron would enter the 2p orbital (orange) because it is slightly closer to the nucleus, on average, than the major portion of the 2s orbital (blue).
- But note that a minor portion of the 2s radial probability distribution appears within the 1s region.
- As a result, an electron in the 2s orbital spends part of its time “penetrating” very close to the nucleus.
- Penetration by the 2s electron increases its overall attraction to the nucleus relative to that for 2p electron
- At the same time, penetration into the 1s region decreases the shielding of the 2s electron by the 1s electron
- Indeed, the 2s orbital of Li is lower in energy than the 2p orbital
- Because it takes more energy to remove a 2s electron (520 kJ/mol) than a 2p (341 kJ/mol)

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- In general, penetration and the resulting effects on shielding cause an energy level to split into sublevels of differing energy
- Order of sublevel energies:  $s < p < d < f$
- Thus, the 2s ( $l=0$ ) is lower in energy than the 2p ( $l=1$ ), the 3p ( $l=1$ ) is lower than the 3d ( $l=2$ ), and so forth.

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## Summaries of Electrostatic Interactions

- Greater nuclear charge lowers orbital energy and makes electrons harder to remove
- Electron-electron repulsions raise orbital energy and makes electrons easier to remove
- Repulsions have the effect of shielding electrons from the full nuclear charge.
- Greater radial probability distribution near the nucleus (greater penetration) makes an electron harder to remove because it is attracted more strongly and shielded less effectively
- As a result, an energy level is split into sublevels with the energy order  $s < p < d < f$

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## The Quantum-Mechanical Model and The Periodic Table

- Quantum mechanics provides the theoretical foundation for the experimentally based periodic table.
- AUFBAU PRINCIPLE: *one electron per element to the lowest energy orbital available.*
- it results in ground-state electron configuration.
- **For H atom:**
  - $Z=1$ ;  $n=1$ ,  $l=0$ ,  $m_l=0$ ,  $m_s=+\frac{1}{2}$
- For He atom:
  - (the first electron in He has the same set as the electron in H, but the second He electron has opposing spin):
    - $Z=2$ ;  $n=1$ ,  $l=0$ ,  $m_l=0$ ,  $m_s=-\frac{1}{2}$

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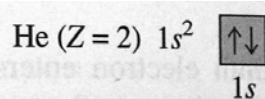
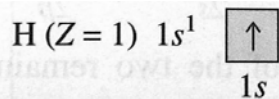
■ **The electron Configuration:**

- designate as:  $nl^{\#}$
- $n$ =energy level;  $l$ =sub level; and  $\#$ =number of electrons

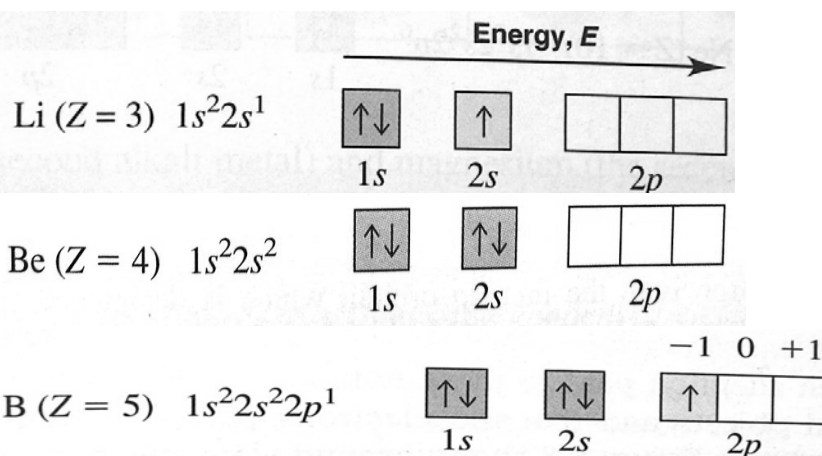
■ **The orbital diagram:**

- A box for each orbital in a given energy level, grouped by sublevel, with an arrow indicating an electron and its spin.
  - an orbital can hold only two electrons
  - The energy of the sublevels increases from left to right
  - Hund's rule: *orbitals of equal energy become half-filled, with electron spins parallel, before any pairing occurs.*
  - Pauli's exclusion: *no two electrons can have the same set of quantum numbers*
- The following are ground state atom's electron configuration

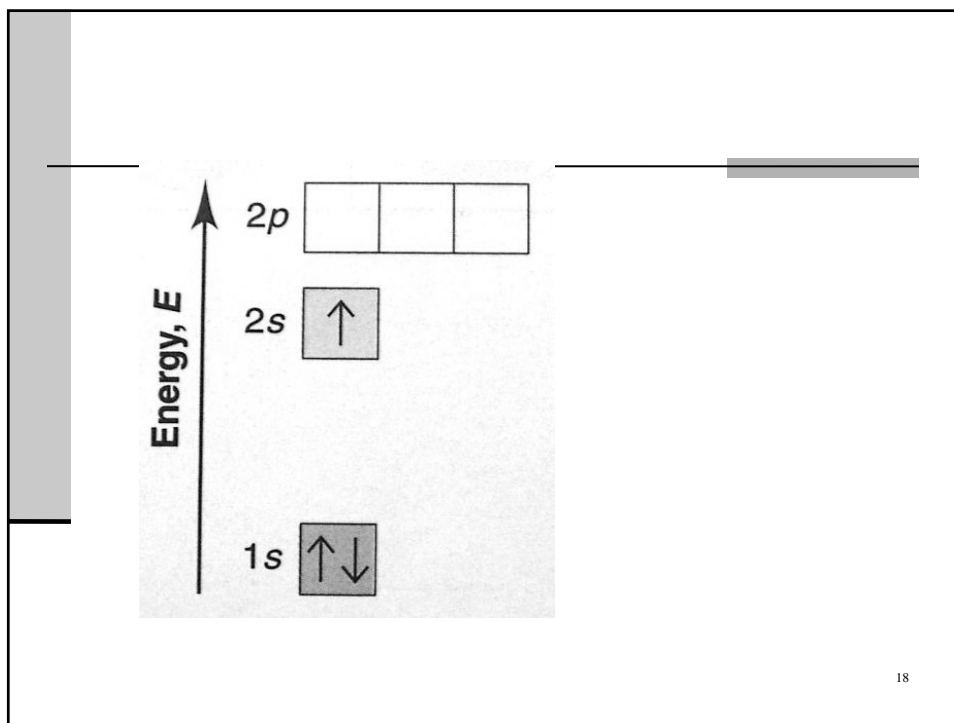
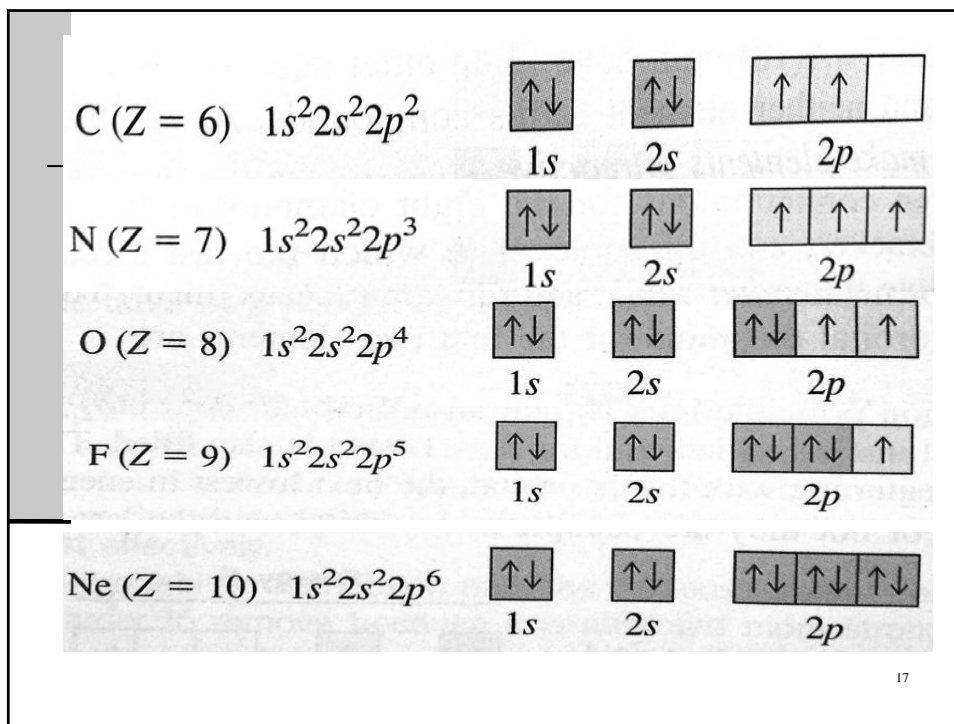
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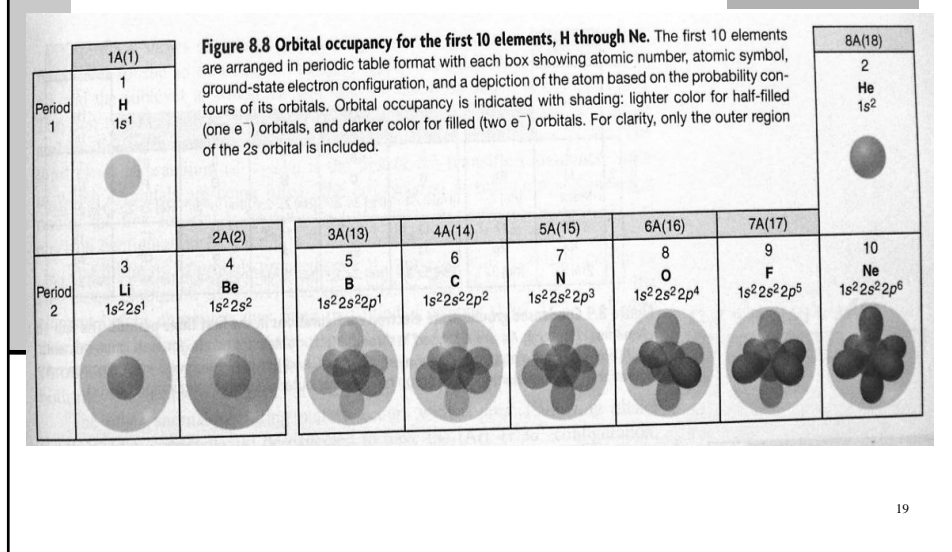
Period 2 (Li .. Ne)







## Building up Period 3 (Na .. Ar)



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**Table 8.3 Partial Orbital Diagrams and Electron Configurations\* for the Elements in Period 3**

Atomic Number	Element	Partial Orbital Diagram (3s and 3p Sublevels Only)	Full Electron Configuration	Condensed Electron Configuration
11	Na	$3s$ $\uparrow$ $3p$ $\square$ $\square$ $\square$	$[1s^2 2s^2 2p^6] 3s^1$	$[\text{Ne}] 3s^1$
12	Mg	$3s$ $\uparrow\downarrow$ $3p$ $\square$ $\square$ $\square$	$[1s^2 2s^2 2p^6] 3s^2$	$[\text{Ne}] 3s^2$
13	Al	$3s$ $\uparrow\downarrow$ $3p$ $\uparrow$ $\square$ $\square$	$[1s^2 2s^2 2p^6] 3s^2 3p^1$	$[\text{Ne}] 3s^2 3p^1$
14	Si	$3s$ $\uparrow\downarrow$ $3p$ $\uparrow$ $\uparrow$ $\square$	$[1s^2 2s^2 2p^6] 3s^2 3p^2$	$[\text{Ne}] 3s^2 3p^2$
15	P	$3s$ $\uparrow\downarrow$ $3p$ $\uparrow$ $\uparrow$ $\uparrow$	$[1s^2 2s^2 2p^6] 3s^2 3p^3$	$[\text{Ne}] 3s^2 3p^3$
16	S	$3s$ $\uparrow\downarrow$ $3p$ $\uparrow\downarrow$ $\uparrow$ $\uparrow$	$[1s^2 2s^2 2p^6] 3s^2 3p^4$	$[\text{Ne}] 3s^2 3p^4$
17	Cl	$3s$ $\uparrow\downarrow$ $3p$ $\uparrow\downarrow$ $\uparrow\downarrow$ $\uparrow$	$[1s^2 2s^2 2p^6] 3s^2 3p^5$	$[\text{Ne}] 3s^2 3p^5$
18	Ar	$3s$ $\uparrow\downarrow$ $3p$ $\uparrow\downarrow$ $\uparrow\downarrow$ $\uparrow\downarrow$	$[1s^2 2s^2 2p^6] 3s^2 3p^6$	$[\text{Ne}] 3s^2 3p^6$

\*Colored type indicates the sublevel to which the last electron is added.

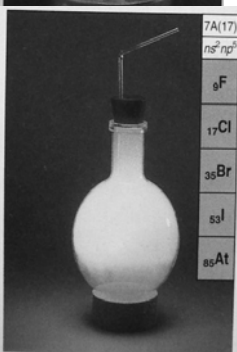
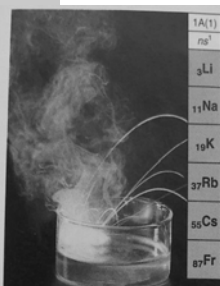
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## Condensed ground state electron configurations in the first three periods

Period	1A (1)							8A (18)
	1 H $1s^1$	2A (2)	3A (13)	4A (14)	5A (15)	6A (16)	7A (17)	2 He $1s^2$
2	3 Li $[\text{He}] 2s^1$	4 Be $[\text{He}] 2s^2$	5 B $[\text{He}] 2s^2 2p^1$	6 C $[\text{He}] 2s^2 2p^2$	7 N $[\text{He}] 2s^2 2p^3$	8 O $[\text{He}] 2s^2 2p^4$	9 F $[\text{He}] 2s^2 2p^5$	10 Ne $[\text{He}] 2s^2 2p^6$
3	11 Na $[\text{Ne}] 3s^1$	12 Mg $[\text{Ne}] 3s^2$	13 Al $[\text{Ne}] 3s^2 3p^1$	14 Si $[\text{Ne}] 3s^2 3p^2$	15 P $[\text{Ne}] 3s^2 3p^3$	16 S $[\text{Ne}] 3s^2 3p^4$	17 Cl $[\text{Ne}] 3s^2 3p^5$	18 Ar $[\text{Ne}] 3s^2 3p^6$

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## Electron Configuration Within Groups



- Similar outer electron configurations correlate with similar chemical behavior
- In Group 1A(1): Li, Na, and alkali metals (K, Rb, Cs, Fr)
- All are highly reactive metals that form ionic compounds with nonmetals with formulas such as  $\text{MCl}$ ,  $\text{M}_2\text{O}$ , and  $\text{M}_2\text{S}$
- In Group 7A(17): F, Cl
- All form ionic compound with metals ( $\text{KX}$ ,  $\text{MgX}_2$ )
- Connection between quantum mechanic with chemical periodicity: "*orbitals are filled in order of increasing energy, which leads to outer electron configurations that recur periodically, which leads to chemical properties that recur periodically*".

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## The First *d*-orbital Transition States: Building Up Period 4

- The 3*d* orbitals are filled in Period 4, however, the 4*s* orbital is filled before the 3*d*.
- This switch in filling order is due to the shielding and penetration effects
- The radial probability distribution of the 3*d* orbital is greater outside the filled, inner *n*=1 and *n*=2, so a 3*d* electron is shielded very effectively from the nuclear charge
- In contrast, penetration by the 4*s* electron means that it spend a significant part of its time near the nucleus and feels a greater nuclear attraction
- Thus, the 4*s* orbital is slightly lower in energy than the 3*d*, and so fill first
- In general, ***the ns sublevel fills before the (n-1)d sublevel.***
- The transition elements have electron configuration with filling *d* orbitals

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**Table 8.4 Partial Orbital Diagrams and Electron Configurations\* for the Elements in Period 4**

Atomic Number	Element	Partial Orbital Diagram (4 <i>s</i> , 3 <i>d</i> , and 4 <i>p</i> Sublevels Only)			Full Electron Configuration	Condensed Electron Configuration
19	K	↑			$[1s^2 2s^2 2p^6 3s^2 3p^6] 4s^1$	[Ar] 4 <i>s</i> <sup>1</sup>
20	Ca	↑↓			$[1s^2 2s^2 2p^6 3s^2 3p^6] 4s^2$	[Ar] 4 <i>s</i> <sup>2</sup>
21	Sc	↑↓	↑		$[1s^2 2s^2 2p^6 3s^2 3p^6] 4s^2 3d^1$	[Ar] 4 <i>s</i> <sup>2</sup> 3 <i>d</i> <sup>1</sup>
22	Ti	↑↓	↑ ↑		$[1s^2 2s^2 2p^6 3s^2 3p^6] 4s^2 3d^2$	[Ar] 4 <i>s</i> <sup>2</sup> 3 <i>d</i> <sup>2</sup>
23	V	↑↓	↑ ↑ ↑		$[1s^2 2s^2 2p^6 3s^2 3p^6] 4s^2 3d^3$	[Ar] 4 <i>s</i> <sup>2</sup> 3 <i>d</i> <sup>3</sup>
24	Cr	↑	↑ ↑ ↑ ↑ ↑		$[1s^2 2s^2 2p^6 3s^2 3p^6] 4s^1 3d^5$	[Ar] 4 <i>s</i> <sup>1</sup> 3 <i>d</i> <sup>5</sup>
25	Mn	↑↓	↑ ↑ ↑ ↑ ↑		$[1s^2 2s^2 2p^6 3s^2 3p^6] 4s^2 3d^5$	[Ar] 4 <i>s</i> <sup>2</sup> 3 <i>d</i> <sup>5</sup>
26	Fe	↑↓	↑↓ ↑ ↑ ↑ ↑		$[1s^2 2s^2 2p^6 3s^2 3p^6] 4s^2 3d^6$	[Ar] 4 <i>s</i> <sup>2</sup> 3 <i>d</i> <sup>6</sup>
27	Co	↑↓	↑↓ ↑↓ ↑ ↑ ↑		$[1s^2 2s^2 2p^6 3s^2 3p^6] 4s^2 3d^7$	[Ar] 4 <i>s</i> <sup>2</sup> 3 <i>d</i> <sup>7</sup>
28	Ni	↑↓	↑↓ ↑↓ ↑↓ ↑ ↑		$[1s^2 2s^2 2p^6 3s^2 3p^6] 4s^2 3d^8$	[Ar] 4 <i>s</i> <sup>2</sup> 3 <i>d</i> <sup>8</sup>
29	Cu	↑	↑↓ ↑↓ ↑↓ ↑↓ ↑↓		$[1s^2 2s^2 2p^6 3s^2 3p^6] 4s^1 3d^{10}$	[Ar] 4 <i>s</i> <sup>1</sup> 3 <i>d</i> <sup>10</sup>
30	Zn	↑↓	↑↓ ↑↓ ↑↓ ↑↓ ↑↓		$[1s^2 2s^2 2p^6 3s^2 3p^6] 4s^2 3d^{10}$	[Ar] 4 <i>s</i> <sup>2</sup> 3 <i>d</i> <sup>10</sup>
31	Ga	↑↓	↑↓ ↑↓ ↑↓ ↑↓ ↑↓	↑	$[1s^2 2s^2 2p^6 3s^2 3p^6] 4s^2 3d^{10} 4p^1$	[Ar] 4 <i>s</i> <sup>2</sup> 3 <i>d</i> <sup>10</sup> 4 <i>p</i> <sup>1</sup>
32	Ge	↑↓	↑↓ ↑↓ ↑↓ ↑↓ ↑↓	↑ ↑	$[1s^2 2s^2 2p^6 3s^2 3p^6] 4s^2 3d^{10} 4p^2$	[Ar] 4 <i>s</i> <sup>2</sup> 3 <i>d</i> <sup>10</sup> 4 <i>p</i> <sup>2</sup>
33	As	↑↓	↑↓ ↑↓ ↑↓ ↑↓ ↑↓	↑ ↑ ↑	$[1s^2 2s^2 2p^6 3s^2 3p^6] 4s^2 3d^{10} 4p^3$	[Ar] 4 <i>s</i> <sup>2</sup> 3 <i>d</i> <sup>10</sup> 4 <i>p</i> <sup>3</sup>
34	Se	↑↓	↑↓ ↑↓ ↑↓ ↑↓ ↑↓	↑↓ ↑ ↑	$[1s^2 2s^2 2p^6 3s^2 3p^6] 4s^2 3d^{10} 4p^4$	[Ar] 4 <i>s</i> <sup>2</sup> 3 <i>d</i> <sup>10</sup> 4 <i>p</i> <sup>4</sup>
35	Br	↑↓	↑↓ ↑↓ ↑↓ ↑↓ ↑↓	↑↓ ↑↓ ↑	$[1s^2 2s^2 2p^6 3s^2 3p^6] 4s^2 3d^{10} 4p^5$	[Ar] 4 <i>s</i> <sup>2</sup> 3 <i>d</i> <sup>10</sup> 4 <i>p</i> <sup>5</sup>
36	Kr	↑↓	↑↓ ↑↓ ↑↓ ↑↓ ↑↓	↑↓ ↑↓ ↑↓	$[1s^2 2s^2 2p^6 3s^2 3p^6] 4s^2 3d^{10} 4p^6$	[Ar] 4 <i>s</i> <sup>2</sup> 3 <i>d</i> <sup>10</sup> 4 <i>p</i> <sup>6</sup>

\*Colored type indicates sublevel(s) whose occupancy changes when the last electron is added.

# Anomalous Filling Pattern: Cr & Cu

- The filling of 3d orbitals proceeds one at a time, as with p orbitals, except in 2 cases: chromium (Cr, Z=24) and copper (Cu, Z=29)
- Vanadium (Z=23), has three half-filled d orbitals ([Ar]4s<sup>2</sup>3d<sup>3</sup>).
- Rather than having its last electron enter a fourth empty d orbital to give [Ar]4s<sup>2</sup>3d<sup>4</sup>, Cr has one electron in the 4s sublevel and five in the 3d sublevel. Thus, both the 4s and 3d sublevels are half-filled.
- Copper would be expected to have the [Ar]4s<sup>2</sup>3d<sup>9</sup> configuration. Instead, the 4s orbital of copper is half-filled (1 e), and the 3d orbitals are filled with 10 electrons
- The anomalous filling patterns in Cr and Cu lead to conclusion that half-filled and filled sublevels are unexpectedly stable.
- In Zn, both 4s and 3d sublevels are completely filled, and the transition series ends.

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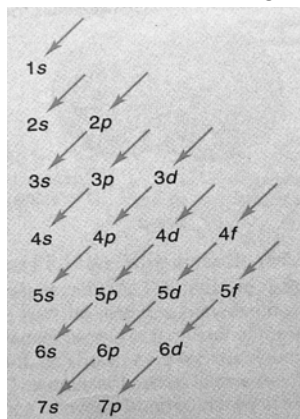
# General Principles of Electron Configuration

**Figure 8.11** A periodic table of partial ground-state electron configurations. These ground-state electron configurations show the electrons beyond the previous noble gas in the sublevel block being filled (excluding filled inner sublevels). For main-group elements, the group heading identifies the general outer configuration. Anomalous electron configurations occur often among the d-block and f-block elements, with the first two appearing for Cr (Z = 24) and Cu (Z = 29). Helium is colored as an s-block element but placed with the other members of Group 8A(18). Configurations for elements 112, 114, and 116 have not yet been confirmed.

Main-Group Elements (s block)		Transition Elements (d block)										Main-Group Elements (p block)									
1A (1)	2A (2)											3A (13)	4A (14)	5A (15)	6A (16)	7A (17)	8A (18)				
ns <sup>1</sup>	ns <sup>2</sup>											ns <sup>2</sup> np <sup>1</sup>	ns <sup>2</sup> np <sup>2</sup>	ns <sup>2</sup> np <sup>3</sup>	ns <sup>2</sup> np <sup>4</sup>	ns <sup>2</sup> np <sup>5</sup>	ns <sup>2</sup> np <sup>6</sup>				
1 H	2 He											3 B	4 C	5 N	6 O	7 F	8 Ne				
1s <sup>1</sup>	1s <sup>2</sup>											2s <sup>2</sup> 2p <sup>1</sup>	2s <sup>2</sup> 2p <sup>2</sup>	2s <sup>2</sup> 2p <sup>3</sup>	2s <sup>2</sup> 2p <sup>4</sup>	2s <sup>2</sup> 2p <sup>5</sup>	2s <sup>2</sup> 2p <sup>6</sup>				
2 Li	3 Be											3 B	4 C	5 N	6 O	7 F	8 Ne				
2s <sup>1</sup>	2s <sup>2</sup>											3s <sup>2</sup> 3p <sup>1</sup>	3s <sup>2</sup> 3p <sup>2</sup>	3s <sup>2</sup> 3p <sup>3</sup>	3s <sup>2</sup> 3p <sup>4</sup>	3s <sup>2</sup> 3p <sup>5</sup>	3s <sup>2</sup> 3p <sup>6</sup>				
3 Na	4 Mg	3B (3)	4B (4)	5B (5)	6B (6)	7B (7)	8B (8)	9B (9)	10B (10)	11B (11)	12B (12)	13 Al	14 Si	15 P	16 S	17 Cl	18 Ar				
3s <sup>1</sup>	3s <sup>2</sup>											3s <sup>2</sup> 3p <sup>1</sup>	3s <sup>2</sup> 3p <sup>2</sup>	3s <sup>2</sup> 3p <sup>3</sup>	3s <sup>2</sup> 3p <sup>4</sup>	3s <sup>2</sup> 3p <sup>5</sup>	3s <sup>2</sup> 3p <sup>6</sup>				
4 K	5 Ca	3B (3)	4B (4)	5B (5)	6B (6)	7B (7)	8B (8)	9B (9)	10B (10)	11B (11)	12B (12)	13 Al	14 Si	15 P	16 S	17 Cl	18 Ar				
4s <sup>1</sup>	4s <sup>2</sup>	4s <sup>2</sup> 3d <sup>1</sup>	4s <sup>2</sup> 3d <sup>2</sup>	4s <sup>2</sup> 3d <sup>3</sup>	4s <sup>2</sup> 3d <sup>4</sup>	4s <sup>2</sup> 3d <sup>5</sup>	4s <sup>2</sup> 3d <sup>6</sup>	4s <sup>2</sup> 3d <sup>7</sup>	4s <sup>2</sup> 3d <sup>8</sup>	4s <sup>2</sup> 3d <sup>9</sup>	4s <sup>2</sup> 3d <sup>10</sup>	4s <sup>2</sup> 4p <sup>1</sup>	4s <sup>2</sup> 4p <sup>2</sup>	4s <sup>2</sup> 4p <sup>3</sup>	4s <sup>2</sup> 4p <sup>4</sup>	4s <sup>2</sup> 4p <sup>5</sup>	4s <sup>2</sup> 4p <sup>6</sup>				
5 Rb	6 Sr	3B (3)	4B (4)	5B (5)	6B (6)	7B (7)	8B (8)	9B (9)	10B (10)	11B (11)	12B (12)	13 Al	14 Si	15 P	16 S	17 Cl	18 Ar				
5s <sup>1</sup>	5s <sup>2</sup>	5s <sup>2</sup> 4d <sup>1</sup>	5s <sup>2</sup> 4d <sup>2</sup>	5s <sup>2</sup> 4d <sup>3</sup>	5s <sup>2</sup> 4d <sup>4</sup>	5s <sup>2</sup> 4d <sup>5</sup>	5s <sup>2</sup> 4d <sup>6</sup>	5s <sup>2</sup> 4d <sup>7</sup>	5s <sup>2</sup> 4d <sup>8</sup>	5s <sup>2</sup> 4d <sup>9</sup>	5s <sup>2</sup> 4d <sup>10</sup>	5s <sup>2</sup> 5p <sup>1</sup>	5s <sup>2</sup> 5p <sup>2</sup>	5s <sup>2</sup> 5p <sup>3</sup>	5s <sup>2</sup> 5p <sup>4</sup>	5s <sup>2</sup> 5p <sup>5</sup>	5s <sup>2</sup> 5p <sup>6</sup>				
6 Cs	8 Ba	3B (3)	4B (4)	5B (5)	6B (6)	7B (7)	8B (8)	9B (9)	10B (10)	11B (11)	12B (12)	13 Al	14 Si	15 P	16 S	17 Cl	18 Ar				
6s <sup>1</sup>	6s <sup>2</sup>	6s <sup>2</sup> 4f <sup>1</sup>	6s <sup>2</sup> 4f <sup>2</sup>	6s <sup>2</sup> 4f <sup>3</sup>	6s <sup>2</sup> 4f <sup>4</sup>	6s <sup>2</sup> 4f <sup>5</sup>	6s <sup>2</sup> 4f <sup>6</sup>	6s <sup>2</sup> 4f <sup>7</sup>	6s <sup>2</sup> 4f <sup>8</sup>	6s <sup>2</sup> 4f <sup>9</sup>	6s <sup>2</sup> 4f <sup>10</sup>	6s <sup>2</sup> 5p <sup>1</sup>	6s <sup>2</sup> 5p <sup>2</sup>	6s <sup>2</sup> 5p <sup>3</sup>	6s <sup>2</sup> 5p <sup>4</sup>	6s <sup>2</sup> 5p <sup>5</sup>	6s <sup>2</sup> 5p <sup>6</sup>				
7 Fr	8 Ra	3B (3)	4B (4)	5B (5)	6B (6)	7B (7)	8B (8)	9B (9)	10B (10)	11B (11)	12B (12)	13 Al	14 Si	15 P	16 S	17 Cl	18 Ar				
7s <sup>1</sup>	7s <sup>2</sup>	7s <sup>2</sup> 5f <sup>1</sup>	7s <sup>2</sup> 5f <sup>2</sup>	7s <sup>2</sup> 5f <sup>3</sup>	7s <sup>2</sup> 5f <sup>4</sup>	7s <sup>2</sup> 5f <sup>5</sup>	7s <sup>2</sup> 5f <sup>6</sup>	7s <sup>2</sup> 5f <sup>7</sup>	7s <sup>2</sup> 5f <sup>8</sup>	7s <sup>2</sup> 5f <sup>9</sup>	7s <sup>2</sup> 5f <sup>10</sup>	7s <sup>2</sup> 6p <sup>1</sup>	7s <sup>2</sup> 6p <sup>2</sup>	7s <sup>2</sup> 6p <sup>3</sup>	7s <sup>2</sup> 6p <sup>4</sup>	7s <sup>2</sup> 6p <sup>5</sup>	7s <sup>2</sup> 6p <sup>6</sup>				
		Inner Transition Elements (f block)																			
		6	7	8	9	10	11	12	13	14	15	16	17	18	19	20	21	22			
		*Lanthanides	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu					
		5f <sup>0</sup> 6s <sup>2</sup>	5f <sup>1</sup> 6s <sup>2</sup>	5f <sup>2</sup> 6s <sup>2</sup>	5f <sup>3</sup> 6s <sup>2</sup>	5f <sup>4</sup> 6s <sup>2</sup>	5f <sup>5</sup> 6s <sup>2</sup>	5f <sup>6</sup> 6s <sup>2</sup>	5f <sup>7</sup> 6s <sup>2</sup>	5f <sup>7</sup> 6s <sup>2</sup>	5f <sup>7</sup> 6s <sup>2</sup>	5f <sup>7</sup> 6s <sup>2</sup>	5f <sup>7</sup> 6s <sup>2</sup>	5f <sup>7</sup> 6s <sup>2</sup>	5f <sup>7</sup> 6s <sup>2</sup>	5f <sup>7</sup> 6s <sup>2</sup>					
		**Actinides	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr					
		5f <sup>0</sup> 7s <sup>2</sup>	5f <sup>2</sup> 7s <sup>2</sup>	5f <sup>3</sup> 7s <sup>2</sup>	5f <sup>4</sup> 7s <sup>2</sup>	5f <sup>4</sup> 7s <sup>2</sup>	5f <sup>4</sup> 7s <sup>2</sup>	5f <sup>4</sup> 7s <sup>2</sup>	5f <sup>4</sup> 7s <sup>2</sup>	5f <sup>4</sup> 7s <sup>2</sup>	5f <sup>4</sup> 7s <sup>2</sup>	5f <sup>4</sup> 7s <sup>2</sup>	5f <sup>4</sup> 7s <sup>2</sup>	5f <sup>4</sup> 7s <sup>2</sup>	5f <sup>4</sup> 7s <sup>2</sup>	5f <sup>4</sup> 7s <sup>2</sup>					

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- Elements in a group have similar chemical properties because they have similar outer electron configurations.
- To memorize the Orbital filling order:



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## Categories of Electrons

- **Inner (Core) electrons:** are those seen in the previous noble gas and any completed transition series. They fill all the lower energy levels of an atom
- **Outer electrons:** are those in the highest energy level (highest  $n$  value). They spend most of their time farthest from the nucleus
- **Valence electrons:** are those involved in forming compounds. Among the main-group elements, the valence electrons are the outer electrons.
- Among the transition elements, the  $(n-1)d$  electrons are counted among the valence electrons because some or all of them are often involved in bonding

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## Group and Period Number

- Among the main-group elements (A group), *the group number equals the number of outer electrons (those with the highest  $n$ )*
- *The period number is the  $n$  value of the highest energy level.*
- The  $n$  value squared ( $n^2$ ) gives the total number of orbitals in that energy level. Because an orbital can hold no more than two electrons (exclusion principle),  $2n^2$  gives the maximum number of electrons in the energy level.
  - $n=3 \implies$  the number of orbitals is  $n^2=9$ : one  $3s$ , three  $3p$ , and five  $3d$ . The number of electrons is  $2n^2 (=18)$ : two  $3s$ , six  $3p$ , and ten  $3d$  electrons
- The period 6 inner transition series fills the  $4f$  orbitals and consists of the lanthanides (or rare earth), so called because they occur after and are similar to lanthanum (La).
- The other inner transition series holds the actinides, which fills  $5f$  orbitals that appear in Period 7 after actinium (Ac).

## Summary

- **Aufbau method:** one electron is added to an atom of each successive element in accord with:
  - *Pauli's exclusion principles* (no two electrons can have the same set of quantum numbers)
  - and *Hund's rule* (orbitals of equal energy become half-filled, with electron spins parallel, before any pairing occurs)
- The element of a group have similar outer electron configurations and similar chemical behavior
- For the main group elements, valence electrons (those involved in reactions) are in the outer (highest energy) level only.
- For transition elements,  $(n-1)d$  electrons are also involved in reactions. In general,  $(n-1)d$  orbitals fill after  $ns$  and before  $np$  orbitals.
- In Periods 6 and 7,  $(n-2)f$  orbitals fill between the first and second  $(n-1)d$  orbitals.

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