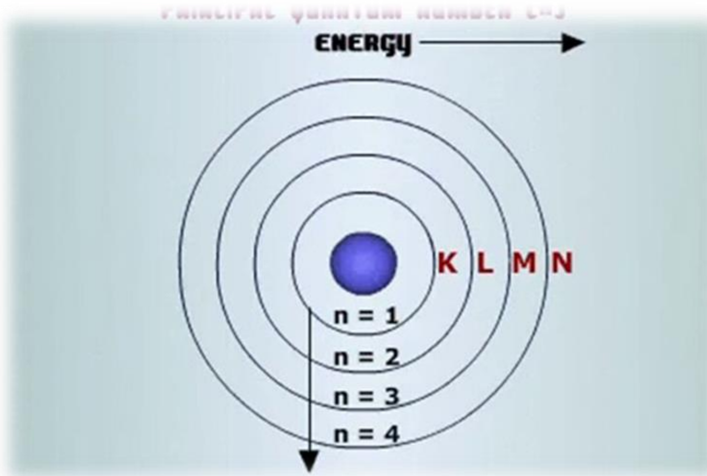
Quantum Mechanical Model of an Atom

Quantum Numbers

- In quantum mechanics, three ***quantum numbers are required to describe the distribution*** *of electrons in hydrogen and other atoms.*
- They are-
  - The principal quantum number (n),
  - The angular momentum quantum number (l),
  - The magnetic quantum number (m_l) & The magnetic spin quantum number(m_s).

The Principal Quantum Number (*n)*



- The principal quantum number (*n) can have integral values 1, 2, 3,…*
- the value of *n determines the energy of an orbital*
- It also relates to the average distance of the electron from the nucleus in a particular orbital.
- The larger the *n is, the* greater the average distance of an electron in the orbital from the nucleus and therefore the larger the orbital.

The Angular Momentum Quantum Number (l)

- Tells the "shape" of the orbitals.
- The values of l depend on the value of the principal quantum number, *n* .
- For a given value of *n* , *l has possible integral values from* *0 to (n - 1)*.
- If n = 1, l = (1-1) = 0, So, l has one value "0".
- If *n = 2, there are* two values of l, given by 0 and 1.

ℓ	0	1	2	3	4	5
Name of orbital	*s*	*p*	*d*	*f*	*g*	*h*

The Magnetic Quantum Number (m_l)

- It tells about the orientation of the e
- m depends on l. it will have (2l +1) number of integral values.

-l , (-l +1), ……,0,……(+l-1), l

Ex: l= 1, m= -1, 0, +1

There is another "**Magnetic spin quantum number(m_s)**. It tells spinning motion of electron if one is clockwise the other is counter-clockwise.

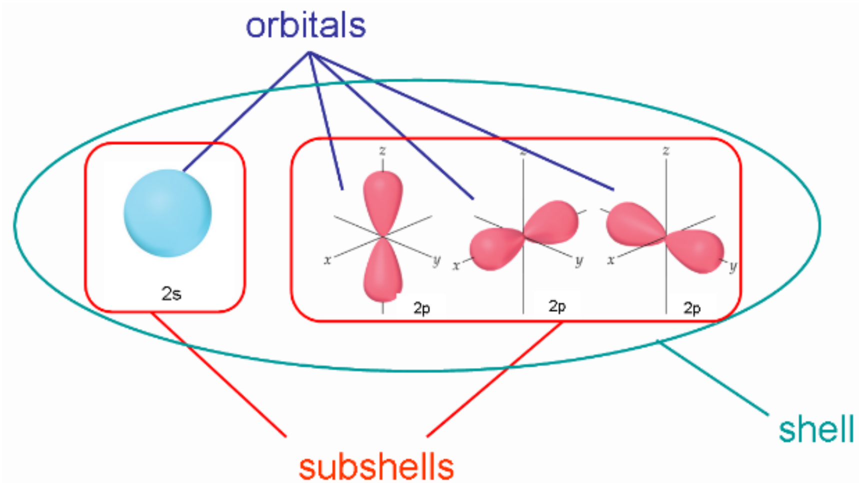
It has 2 values: +1/2 and -1/2.

Quantum Numbers

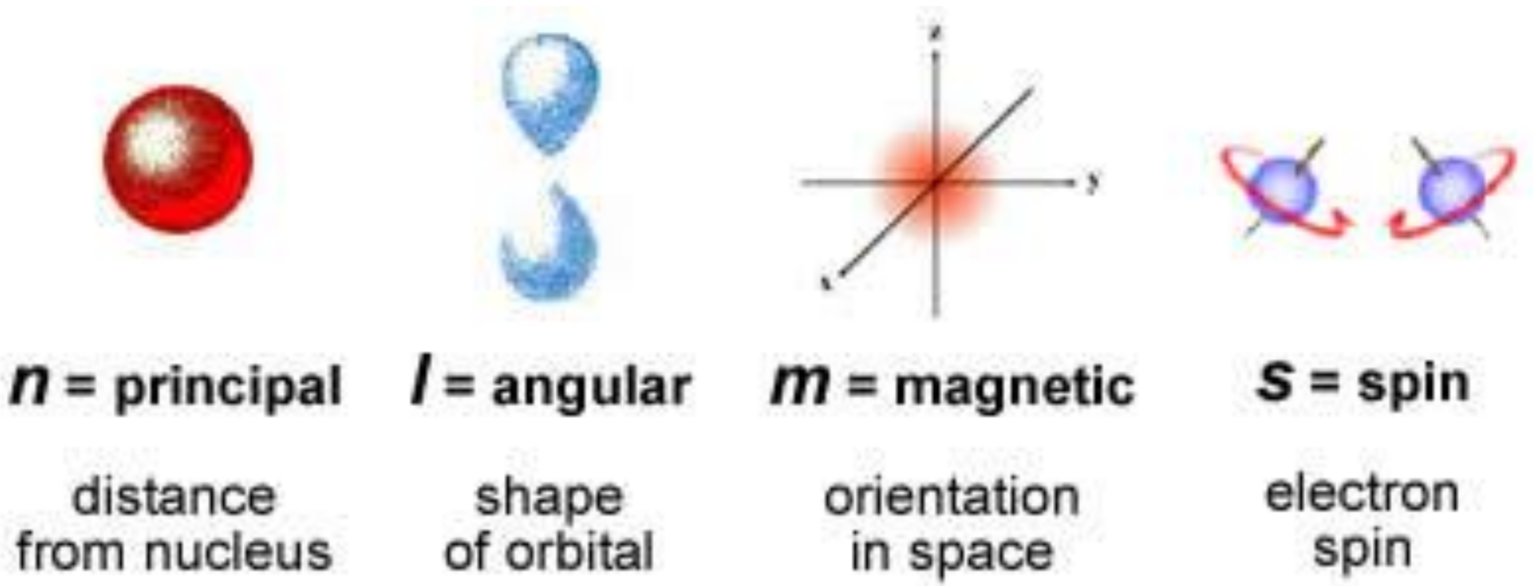
• All electrons that have the same value for n (the principle quantum number) are in the same shell

• Within a shell (same n), all electrons that share the same l (orbital shape) are in the same sub-shell

• When electrons share the same n, l and ml, we say they are in the same orbital (they have the same energy level, shape, and orientation)



Atomic Orbital



Atomic Orbital

- **A 3 dimensional space around a nucleus in which electrons are most likely to be found**
- **Shape represents electron density (*not a path the electron follows*)**
- **Each orbital can hold up to 2 electrons.**
- **Each of the major energy levels (n=1, 2, 3...) contain one or more sublevels (s, p, d and f) in which there are one or more orbitals.**

Atomic Orbital (s, p, d, f)

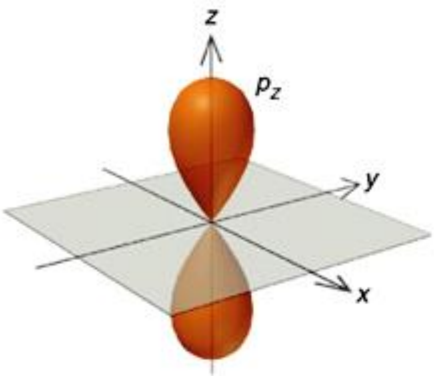
TABLE 7.2 Relation Between Quantum Numbers and Atomic Orbitals

n	ℓ	m_ℓ	Number of Orbitals	Atomic Orbital Designations
1	0	0	1	$1s$
2	0	0	1	$2s$
	1	$-1, 0, 1$	3	$2p_x, 2p_y, 2p_z$
3	0	0	1	$3s$
	1	$-1, 0, 1$	3	$3p_x, 3p_y, 3p_z$
	2	$-2, -1, 0, 1, 2$	5	$3d_{xy}, 3d_{yz}, 3d_{xz}, 3d_{x^2-y^2}, 3d_{z^2}$
⋮	⋮	⋮	⋮	⋮

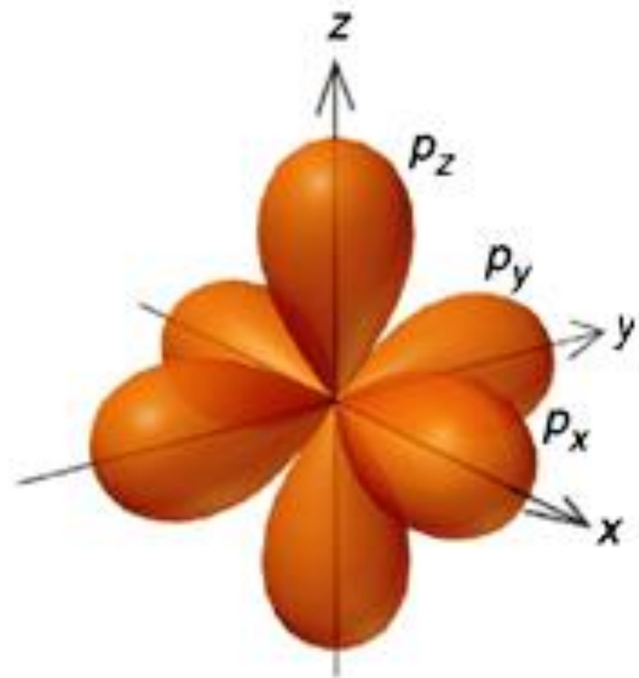
The *s* orbital

1s orbital 2s orbital 3s orbital

The three *p* orbitals



Stylized probability contour



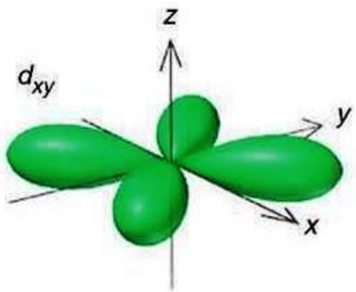
The three *p* orbitals

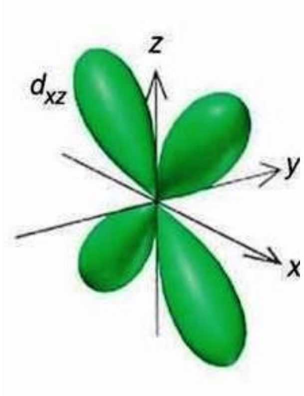
p orbitals, $\ell = 1$.

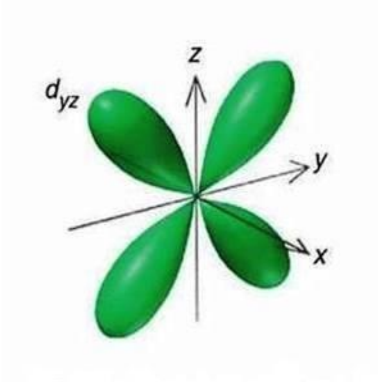
$m_\ell = -1, 0, +1$

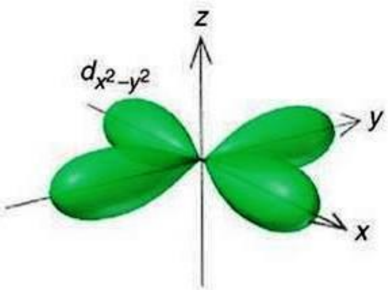
p_x $\quad$ p_z $\quad$ p_y

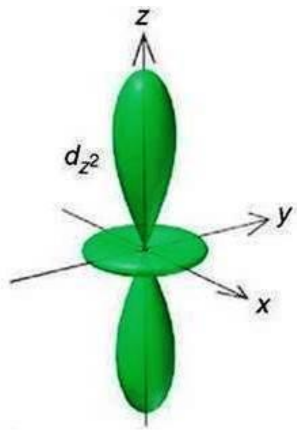
The five *d* orbitals

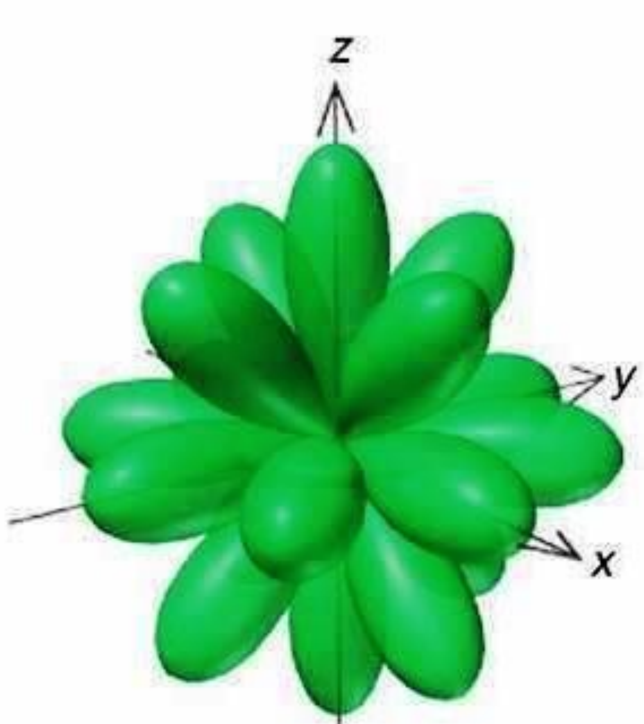












The five *d* orbitals

d orbitals, $l = 2$

$m_l = -2, -1, 0, +1, +2$

d_{xy} d_{xz} d_{z^2} d_{yz} $d_{x^2-y^2}$

Quantum Numbers and Atomic Orbitals

$n = 1, \ell = 0, m_\ell = 0$

$n = 2, \ell = 0, m_\ell = 0$
$\ell = 1, m_\ell = -1, 0, +1$

$n = 3, \ell = 0, m_\ell = 0$
$\ell = 1, m_\ell = -1, 0, +1$
$\ell = 2, m_\ell = -2, -1, 0, +1, +2$

m_ℓ values $= 2\ell + 1$

$\ell = 0, 1, 2, 3, 4, 5, 6, \ldots$

s p d f g h i

Table 7.2 The Hierarchy of Quantum Numbers for Atomic Orbitals

Name, Symbol (Property)	Allowed Values	Quantum Numbers
Principal, *n* **(size)**	**Positive integer (1, 2, 3, ...)**	1 2 3
Angular momentum, *l* **(shape)**	**0 to *n* – 1**	0 0 1 0 1 2
Magnetic, m_l **(orientation)**	**$-l,\ldots,0,\ldots,+l$**	0 0 -1 0 +1 0 -1 0 +1 -2 -1 0 +1 +2

Quantum Numbers for Electrons

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	+½ or –½	direction of e– spin

- Each electron in any atom is described completely by a set of ***four*** quantum numbers.

➢ The **first three** quantum numbers describe the **orbital**, while the **fourth** quantum number describes **electron spin**.

Sample Problem 7.6

What values of the angular momentum (l) and magnetic (m_l) quantum numbers are allowed for a principal quantum number (n) of 3? How many orbitals are allowed for n = 3?

$n = 3$, $l = 0$, $m_l = 0$

$l = 1$, $m_l = -1, 0, +1$

$l = 2$, $m_l = -2, -1, 0, +1, +2$

Total # orbitals for $\underline{n=3}$ is $\underline{9}$

For any n, # of orbitals is n^2

For any n, # of electrons is $2n^2$

Sample Problem 7.7

Give the name, magnetic quantum numbers, and number of orbitals for each sublevel with the following quantum numbers:

(a) $n = 3$, $l = 2$ **(b)** $n = 2$, $l = 0$ **(c)** $n = 5$, $l = 1$ **(d)** $n = 4$, $l = 3$

	name	m_l	# orbitals
(a) $n = 3$, $l = 2$	3d	−2, −1, 0, +1, +2	5
(b) $n = 2$, $l = 0$	2s	0	1
(c) $n = 5$, $l = 1$	5p	−1, 0, +1	3
(d) $n = 4$, $l = 3$	4f	−3, −2, −1, 0, +1, +2, +3	7

Sample Problem 7.8

What is wrong with each of the following quantum numbers designations and/or sublevel names?

	n	l	m_l	Name
(a)	1	$\cancel{1}$	0	1*p*
(b)	4	3	+1	~~4*d*~~ 4f
(c)	3	1	$\cancel{-2}$	3*p*

a l can't be 1.

b 4f.

c m_l can't be −2

CLASS ACTIVITY

Problem 7.50

How many orbitals in an atom can have each of the following designations:

(a) $5f$

(b) $4p$

(c) $5d$

(d) $n = 2$?

Problem 7.52

Give all possible m_l values for orbitals that have each of the following:

(a) $l = 3$

(b) $n = 2$

(c) $n = 6$, $l = 1$.

Problem 7.72

The following combinations are not allowed. If n and m_l are correct, change the l value to create an allowable combination:

(a) $n = 3; l = 0; m_l = -1$

(b) $n = 3; l = 3; m_l = +1$

(c) $n = 7; l = 2; m_l = +3$

(d) $n = 4; l = 1; m_l = -2$

Electronic Configuration of Atoms with Many Electrons

Energies of Orbital

- **Total energy of a subshell =**
  energy of the main shell + the subshell
- **The 4s energy < 3d energy**

4p	___	
3d	___	**(finishes the n=3 shell)**
4s	___	**(starts the n=4 shell)**
3p	___	
3s	___	
2p	___	
2s	___	
1s	___	

Aufbau (or *Building Up*) Principle

- The **Aufbau Principle** states that *electrons are always placed in the lowest energy level first, and then placed in higher levels in order.*

Period 1:

H (Z = 1) 1s^1 $\boxed{\uparrow}$ 1s

He (Z = 2) 1s^2 $\boxed{\uparrow\downarrow}$ 1s

Period 2:

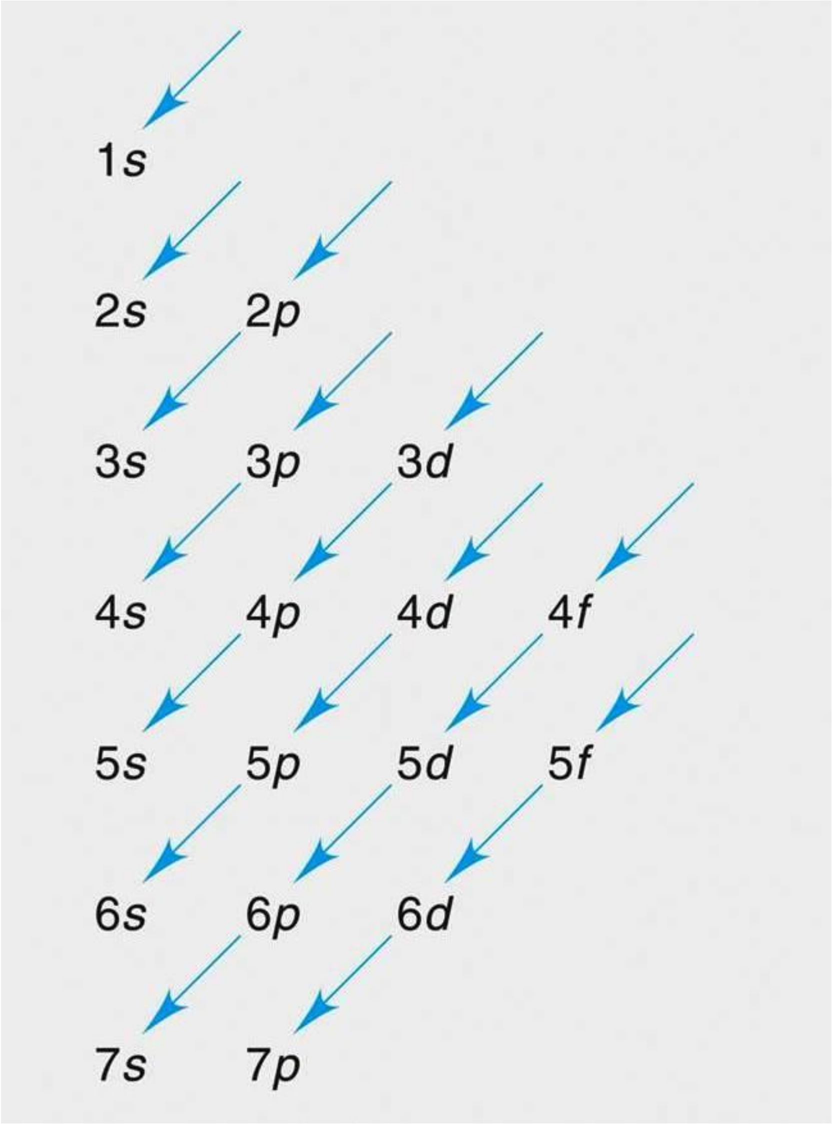
$_3\text{Li} \rightarrow 1s^2\,2s^1$ $\quad \underset{}{\overset{1s^2}{\boxed{\uparrow\downarrow}}}\ \overset{2s^1}{\boxed{\uparrow}}$

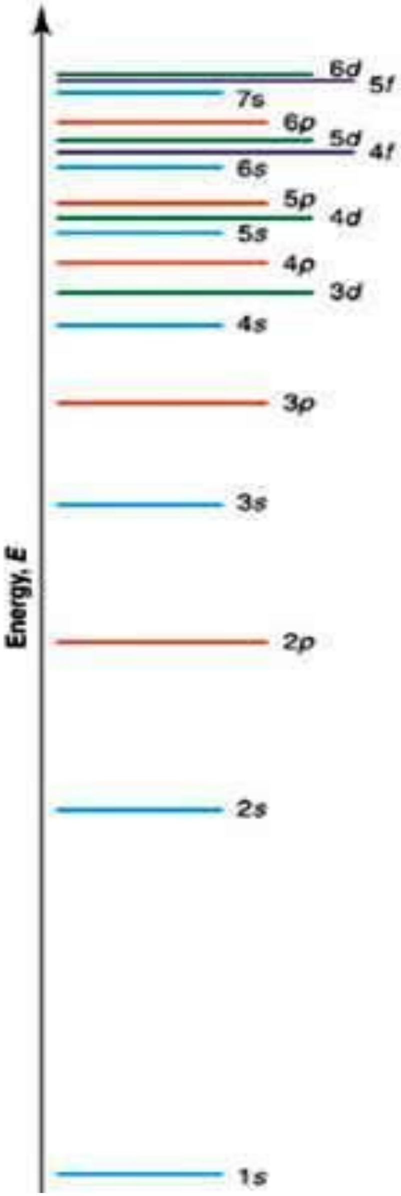
$_4\text{Be} \rightarrow 1s^2\,2s^2$ $\quad \overset{1s^2}{\boxed{\uparrow\downarrow}}\ \overset{2s^2}{\boxed{\uparrow\downarrow}}$

$_5\text{B} \rightarrow 1s^2\,2s^2\,2p^1$ $\quad \overset{1s^2}{\boxed{\uparrow\downarrow}}\ \overset{2s^2}{\boxed{\uparrow\downarrow}}\ \overset{2p^1}{\boxed{\uparrow}\boxed{}\boxed{}}$

$-1 \quad 0 \quad +1$

Aid to memorizing sub-shell filling order





Energies of Orbital

The Pauli Exclusion Principle

- *"No two electrons in an atom can have the same set of four quantum numbers."*
- If two electrons in an atom have the same n , l, *and* m_l *values,* then they must have different values of m_s

Hund’s Rule

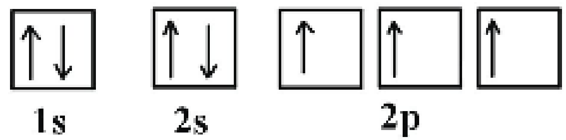
- Hund's rule states that:

1. Every orbital in a sublevel is singly occupied before any orbital is doubly occupied.
2. All of the electrons in singly occupied orbitals have the same spin (to maximize total spin).

Hund’s Rule

Consider the correct electron configuration of the nitrogen (Z = 7) atom: $1s^2\ 2s^2\ 2p^3$

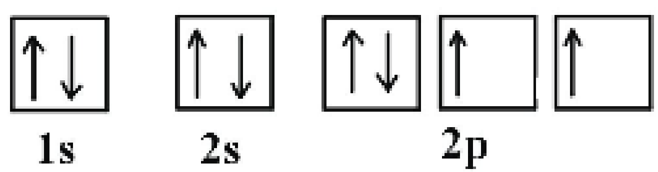
Nitrogen



The p orbitals are half-filled; there are three electrons and three p orbitals. This is because the three electrons in the 2p subshell will fill all the empty orbitals first before pairing with electrons in them.

Next, consider oxygen (Z = 8) atom, the element after nitrogen in the same period; its electron configuration is: $1s^2\ 2s^2\ 2p^4$

Oxygen



General Rules for Assigning Electrons to Atomic Orbitals

Based on the preceding examples we can formulate some general rules for determining the maximum number of electrons that can be assigned to the various subshells and orbitals for a given value of *n :*

1. Each shell or principal level of quantum number *n contains n subshells*. *For*

example, if *n = 2, then there are two subshells (two values of l) of angular* momentum quantum numbers 0 and 1.

2. Each subshell of quantum number l contains (2l+1) orbitals. For example, if l = 1, then there are three *p orbitals*.

3. No more than two electrons can be placed in each orbital. Therefore, the maximum number of electrons is simply twice the number of orbitals that are employed.

4. A quick way to determine the maximum number of electrons that an atom can have in a principal level *n is to use the formula* $2n^2$.

Electron Configuration

- List of subshells containing electrons
- Written in order of increasing energy
- Superscripts give the number of electrons

Example: Electron configuration of neon

number of electrons

$1s^2$ $2s^2$ $2p^6$

main shell subshell

Learning Check

Write the complete electronic configuration for each:

A. 99Es

B. 77Ir

Learning Check

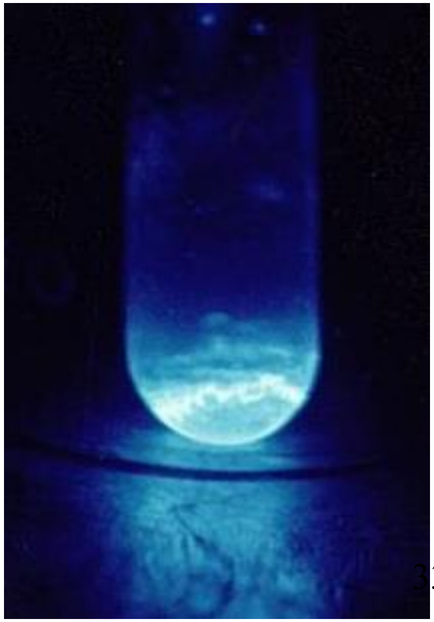
Write the complete electronic configuration

A. 99Es

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6 7s^2 5f^{11}$

[Rn] $5f^{11}$ $7s^2$

- Einsteinium is a synthetic element with symbol Es . Einsteinium was discovered as a component of the debris of the first hydrogen bomb explosion in 1952, and named after Albert Einstein. Its most common isotope einsteinium-253 (half life 20.47 days) is produced artificially.



Learning Check

Write the complete electronic configuration

A. 77Ir

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^7$

[Xe] $4f^{14}\ 5d^7\ 6s^2$

- A very hard, brittle, silvery-white transition metal
- Iridium was discovered in 1803 among insoluble impurities in natural platinum

TABLE 7.3 The Ground-State Electron Configurations of the Elements*

Atomic Number	Symbol	Electron Configuration	Atomic Number	Symbol	Electron Configuration	Atomic Number	Symbol	Electron Configuration
14	Si	$[Ne]3s^23p^2$	51	Sb	$[Kr]5s^24d^{10}5p^3$	88	Ra	$[Rn]7s^2$
15	P	$[Ne]3s^23p^3$	52	Te	$[Kr]5s^24d^{10}5p^4$	89	Ac	$[Rn]7s^26d^1$
16	S	$[Ne]3s^23p^4$	53	I	$[Kr]5s^24d^{10}5p^5$	90	Th	$[Rn]7s^26d^2$
17	Cl	$[Ne]3s^23p^5$	54	Xe	$[Kr]5s^24d^{10}5p^6$	91	Pa	$[Rn]7s^25f^26d^1$
18	Ar	$[Ne]3s^23p^6$	55	Cs	$[Xe]6s^1$	92	U	$[Rn]7s^25f^36d^1$
19	K	$[Ar]4s^1$	56	Ba	$[Xe]6s^2$	93	Np	$[Rn]7s^25f^46d^1$
20	Ca	$[Ar]4s^2$	57	La	$[Xe]6s^25d^1$	94	Pu	$[Rn]7s^25f^6$

Exceptions

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. A list of the exceptions to the Aufbau process is given below.

Table 1: Exceptions to Electron Configuration Trends

Period 4:	Period 5:
Chromium: Z:24 [Ar] $3d^5 4s^1$	Niobium: Z:41 [Kr] $5s^1\ 4d^4$
Copper: Z:29 [Ar] $3d^{10} 4s^1$	Molybdenum: Z:42 [Kr] $5s^1\ 4d^5$
	Ruthenium: Z:44 [Kr] $5s^1\ 4d^7$
	Rhodium: Z:45 [Kr] $5s^1\ 4d^8$
	Palladium: Z:46 [Kr] $4d^{10}$
	Silver: Z:47 [Kr] $5s^1\ 4d^{10}$

Table 1: Exceptions to Electron Configuration Trends

Period 6:	Period 7:
Lanthanum: Z:57 [Xe] $6s^2 5d^1$	Actinium: Z:89 [Rn] $7s^2 6d^1$
Cerium: Z:58 [Xe] $6s^2 4f^1 5d^1$	Thorium: Z:90 [Rn] $7s^2 6d^2$
Gadolinium: Z:64 [Xe] $6s^2 4f^7 5d^1$	Protactium: Z:91 [Rn] $7s^2 5f^2 6d^1$
Platinum: Z:78 [Xe] $6s^1 4f^{14} 5d^9$	Uranium: Z:92 [Rn] $7s^2 5f^3 6d^1$
Gold: Z:79 [Xe] $6s^1 4f^{14} 5d^{10}$	Neptunium: Z:93 [Rn] $7s^2 5f^4 6d^1$
	Curium: Z:96 [Rn] $7s^2 5f^7 6d^1$
	Lawrencium: Z:103 [Rn] $7s^2 5f^{14} 7p^1$

Anomalies of $_{24}Cr$ and $_{29}Cu$

❖ ***Half-filled*** ***and fulfilled d-orbitals tend to be more stable than partially filled ones***

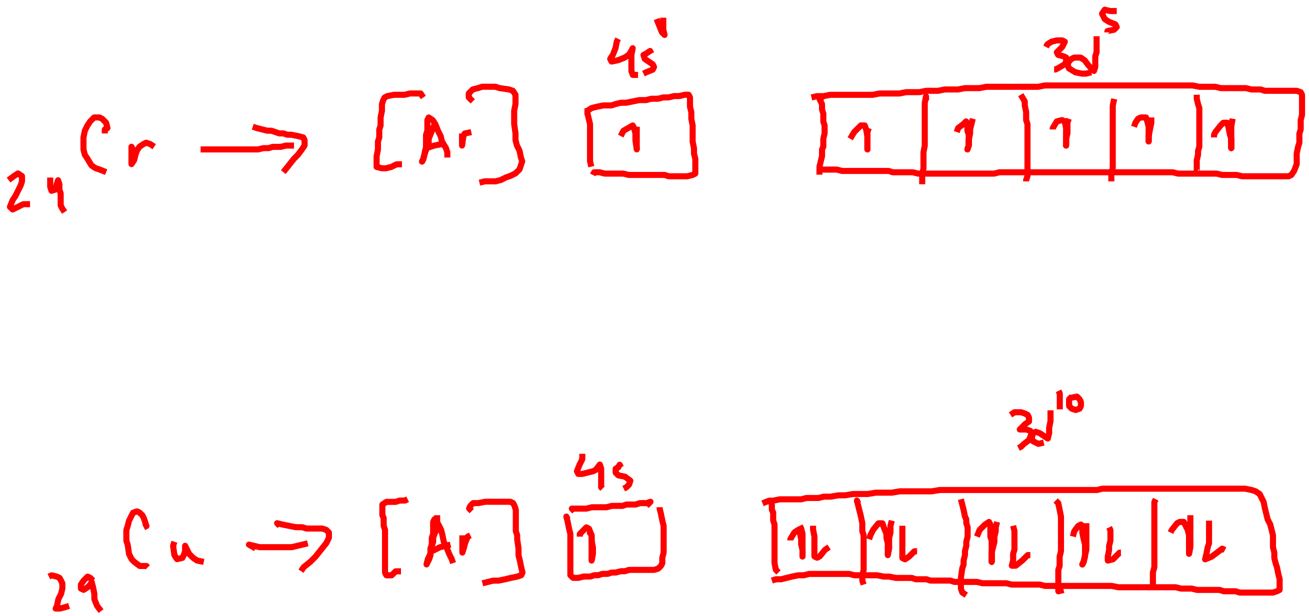


Table 8.4 Partial Orbital Diagrams and Electron Configurations* for the Elements in Period 4

Atomic Number	Element	Partial Orbital Diagram (4*s*, 3*d*, and 4*p* Sublevels Only)									Full Electron Configuration	Condensed Electron Configuration
		4*s*	3*d*					4*p*				
19	K	↑									$1s^22s^22p^63s^23p^64s^1$	$[Ar]\ 4s^1$
20	Ca	↑↓									$1s^22s^22p^63s^23p^64s^2$	$[Ar]\ 4s^2$
21	Sc	↑↓	↑								$1s^22s^22p^63s^23p^64s^23d^1$	$[Ar]\ 4s^23d^1$
22	Ti	↑↓	↑	↑							$1s^22s^22p^63s^23p^64s^23d^2$	$[Ar]\ 4s^23d^2$
23	V	↑↓	↑	↑	↑						$1s^22s^22p^63s^23p^64s^23d^3$	$[Ar]\ 4s^23d^3$
24	Cr	↑	↑	↑	↑	↑	↑				$1s^22s^22p^63s^23p^64s^13d^5$	$[Ar]\ 4s^13d^5$
25	Mn	↑↓	↑	↑	↑	↑	↑				$1s^22s^22p^63s^23p^64s^23d^5$	$[Ar]\ 4s^23d^5$
26	Fe	↑↓	↑↓	↑	↑	↑	↑				$1s^22s^22p^63s^23p^64s^23d^6$	$[Ar]\ 4s^23d^6$
27	Co	↑↓	↑↓	↑↓	↑	↑	↑				$1s^22s^22p^63s^23p^64s^23d^7$	$[Ar]\ 4s^23d^7$
28	Ni	↑↓	↑↓	↑↓	↑↓	↑	↑				$1s^22s^22p^63s^23p^64s^23d^8$	$[Ar]\ 4s^23d^8$
29	Cu	↑	↑↓	↑↓	↑↓	↑↓	↑↓				$1s^22s^22p^63s^23p^64s^13d^{10}$	$[Ar]\ 4s^13d^{10}$
30	Zn	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓				$1s^22s^22p^63s^23p^64s^23d^{10}$	$[Ar]\ 4s^23d^{10}$
31	Ga	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑			$1s^22s^22p^63s^23p^64s^23d^{10}4p^1$	$[Ar]\ 4s^23d^{10}4p^1$
32	Ge	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑		$1s^22s^22p^63s^23p^64s^23d^{10}4p^2$	$[Ar]\ 4s^23d^{10}4p^2$
33	As	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑	↑	$1s^22s^22p^63s^23p^64s^23d^{10}4p^3$	$[Ar]\ 4s^23d^{10}4p^3$
34	Se	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑	$1s^22s^22p^63s^23p^64s^23d^{10}4p^4$	$[Ar]\ 4s^23d^{10}4p^4$
35	Br	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	$1s^22s^22p^63s^23p^64s^23d^{10}4p^5$	$[Ar]\ 4s^23d^{10}4p^5$
36	Kr	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	$1s^22s^22p^63s^23p^64s^23d^{10}4p^6$	$[Ar]\ 4s^23d^{10}4p^6$

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

Sample Problem 8.1

Write a set of quantum numbers for the ***third*** electron and a set for the ***eighth*** electron of the F atom.

SOLUTION

$_9F \Rightarrow 1s^2\, 2s^2\, 2p^5$

$1s^2$ | ↑↓ | $2s^2$ | ↑↓ | $2p^5$ | ↑↓ | ↑↓ | ↑ |

$-1 \quad 0 \quad +1$

$n = 2$
$l = 0$
$m_l = 0$
$m_s = +\frac{1}{2}$

$n = 2$
$l = 1$
$m_l = -1$
$m_s = -\frac{1}{2}$

Main-Group Elements (*s* block) | Transition Elements (*d* block) | Main-Group Elements (*p* block)

Period number: highest occupied energy level

Period	1A (1) ns^1	2A (2) ns^2	3B (3)	4B (4)	5B (5)	6B (6)	7B (7)	8B (8)	8B (9)	8B (10)	1B (11)	2B (12)	3A (13) ns^2np^1	4A (14) ns^2np^2	5A (15) ns^2np^3	6A (16) ns^2np^4	7A (17) ns^2np^5	8A (18) ns^2np^6
1	1 **H** $1s^1$																	2 **He** $1s^2$
2	3 **Li** $2s^1$	4 **Be** $2s^2$											5 **B** $2s^22p^1$	6 **C** $2s^22p^2$	7 **N** $2s^22p^3$	8 **O** $2s^22p^4$	9 **F** $2s^22p^5$	10 **Ne** $2s^22p^6$
3	11 **Na** $3s^1$	12 **Mg** $3s^2$											13 **Al** $3s^23p^1$	14 **Si** $3s^23p^2$	15 **P** $3s^23p^3$	16 **S** $3s^23p^4$	17 **Cl** $3s^23p^5$	18 **Ar** $3s^23p^6$
4	19 **K** $4s^1$	20 **Ca** $4s^2$	21 **Sc** $4s^23d^1$	22 **Ti** $4s^23d^2$	23 **V** $4s^23d^3$	24 **Cr** $4s^13d^5$	25 **Mn** $4s^23d^5$	26 **Fe** $4s^23d^6$	27 **Co** $4s^23d^7$	28 **Ni** $4s^23d^8$	29 **Cu** $4s^13d^{10}$	30 **Zn** $4s^23d^{10}$	31 **Ga** $4s^24p^1$	32 **Ge** $4s^24p^2$	33 **As** $4s^24p^3$	34 **Se** $4s^24p^4$	35 **Br** $4s^24p^5$	36 **Kr** $4s^24p^6$
5	37 **Rb** $5s^1$	38 **Sr** $5s^2$	39 **Y** $5s^24d^1$	40 **Zr** $5s^24d^2$	41 **Nb** $5s^14d^4$	42 **Mo** $5s^14d^5$	43 **Tc** $5s^24d^5$	44 **Ru** $5s^14d^7$	45 **Rh** $5s^14d^8$	46 **Pd** $4d^{10}$	47 **Ag** $5s^14d^{10}$	48 **Cd** $5s^24d^{10}$	49 **In** $5s^25p^1$	50 **Sn** $5s^25p^2$	51 **Sb** $5s^25p^3$	52 **Te** $5s^25p^4$	53 **I** $5s^25p^5$	54 **Xe** $5s^25p^6$
6	55 **Cs** $6s^1$	56 **Ba** $6s^2$	57 **La*** $6s^25d^1$	72 **Hf** $6s^25d^2$	73 **Ta** $6s^25d^3$	74 **W** $6s^25d^4$	75 **Re** $6s^25d^5$	76 **Os** $6s^25d^6$	77 **Ir** $6s^25d^7$	78 **Pt** $6s^15d^9$	79 **Au** $6s^15d^{10}$	80 **Hg** $6s^25d^{10}$	81 **Tl** $6s^26p^1$	82 **Pb** $6s^26p^2$	83 **Bi** $6s^26p^3$	84 **Po** $6s^26p^4$	85 **At** $6s^26p^5$	86 **Rn** $6s^26p^6$
7	87 **Fr** $7s^1$	88 **Ra** $7s^2$	89 **Ac**** $7s^26d^1$	104 **Rf** $7s^26d^2$	105 **Db** $7s^26d^3$	106 **Sg** $7s^26d^4$	107 **Bh** $7s^26d^5$	108 **Hs** $7s^26d^6$	109 **Mt** $7s^26d^7$	110 **Ds** $7s^26d^8$	111 **Rg** $7s^26d^9$	112 **Cn** $7s^26d^{10}$	113 $7s^27p^1$	114 **Fl** $7s^27p^2$	115 $7s^27p^3$	116 **Lv** $7s^27p^4$	117 $7s^27p^5$	118 $7s^27p^6$

Inner Transition Elements (*f* block)

Period	Series														
6	*Lanthanides	58 **Ce** $6s^24f^15d^1$	59 **Pr** $6s^24f^3$	60 **Nd** $6s^24f^4$	61 **Pm** $6s^24f^5$	62 **Sm** $6s^24f^6$	63 **Eu** $6s^24f^7$	64 **Gd** $6s^24f^75d^1$	65 **Tb** $6s^24f^9$	66 **Dy** $6s^24f^{10}$	67 **Ho** $6s^24f^{11}$	68 **Er** $6s^24f^{12}$	69 **Tm** $6s^24f^{13}$	70 **Yb** $6s^24f^{14}$	71 **Lu** $6s^24f^{14}5d^1$
7	**Actinides	90 **Th** $7s^26d^2$	91 **Pa** $7s^25f^26d^1$	92 **U** $7s^25f^36d^1$	93 **Np** $7s^25f^46d^1$	94 **Pu** $7s^25f^6$	95 **Am** $7s^25f^7$	96 **Cm** $7s^25f^76d^1$	97 **Bk** $7s^25f^9$	98 **Cf** $7s^25f^{10}$	99 **Es** $7s^25f^{11}$	100 **Fm** $7s^25f^{12}$	101 **Md** $7s^25f^{13}$	102 **No** $7s^25f^{14}$	103 **Lr** $7s^25f^{14}6d^1$

Electronic Configuration of Period 3: $_{11}$Na, $_{12}$Mg, ..., $_{15}$P...., $_{18}$Ar (using Ne configuration)

$_{10}Ne \rightarrow 1s^2\ 2s^2\ 2p^6$

$1s^2$ ↑↓ $2s^2$ ↑↓ $2p^6$ ↑↓ ↑↓ ↑↓

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

Electronic Configuration of Period 4: $_{19}$K, ... $_{23}$V, ..., $_{28}$Ni, ..., $_{36}$Kr (using Ar configuration)

$_{18}Ar \rightarrow 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6 \rightarrow 18e^-$

$_{19}K \rightarrow [Ar]\ 4s^1 \rightarrow [Ar]\ \underset{4s^1}{\boxed{\uparrow}}$

$_{23}V \rightarrow [Ar]\ 4s^2\ 3d^3 \rightarrow [Ar]\ \underset{4s^2}{\boxed{\uparrow\downarrow}}\ \underset{3d^3}{\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}\boxed{}\boxed{}}$

$_{28}Ni \rightarrow [Ar]\ 4s^2\ 3d^8 \rightarrow [Ar]\ \underset{4s^2}{\boxed{\uparrow\downarrow}}\ \underset{3d^8}{\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow}\boxed{\uparrow}}$

$_{36}Kr \rightarrow [Ar]\ 4s^2\ 3d^{10}\ 4p^6 \rightarrow [Ar]\ \underset{4s^2}{\boxed{\uparrow\downarrow}}\ \underset{3d^{10}}{\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}}$

$\underset{4p^6}{\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}}$

Categories of Electrons

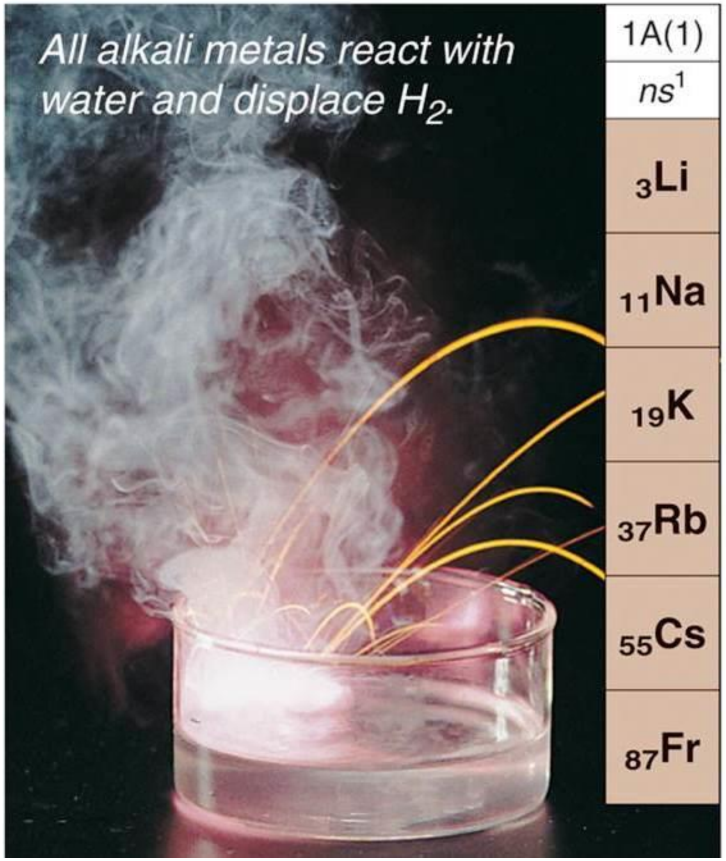
- **Inner (core) electrons** are those an atom has in common with the previous noble gas and any ***completed*** transition series.

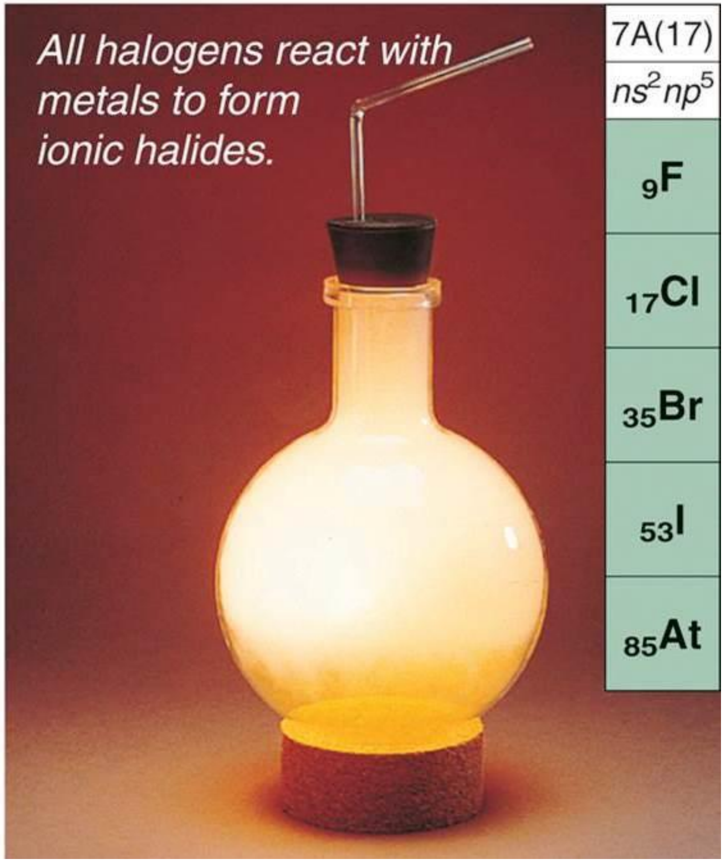
- **Outer electrons** are those in the ***highest*** energy level (highest *n* value).

- **Valence electrons** are those involved in forming compounds.
  - For **main group** elements, the valence electrons ***are*** the outer electrons.
  - For **transition elements**, the valence electrons include the outer electrons and any $(n-1)d$ electrons.

Electron Configuration and Group

- Elements in the same group of the periodic table have the same outer electron configuration, and thus they exhibit similar chemical behavior.





Main-Group Elements (s block)

Main-Group Elements (p block)

Transition Elements (d block)

Period number: highest occupied energy level

Period	1A (1)	2A (2)	3B (3)	4B (4)	5B (5)	6B (6)	7B (7)	8B (8)	8B (9)	8B (10)	1B (11)	2B (12)	3A (13)	4A (14)	5A (15)	6A (16)	7A (17)	8A (18)
	ns^1	ns^2											ns^2np^1	ns^2np^2	ns^2np^3	ns^2np^4	ns^2np^5	ns^2np^6
1	1 **H** $1s^1$																	2 **He** $1s^2$
2	3 **Li** $2s^1$	4 **Be** $2s^2$											5 **B** $2s^22p^1$	6 **C** $2s^22p^2$	7 **N** $2s^22p^3$	8 **O** $2s^22p^4$	9 **F** $2s^22p^5$	10 **Ne** $2s^22p^6$
3	11 **Na** $3s^1$	12 **Mg** $3s^2$											13 **Al** $3s^23p^1$	14 **Si** $3s^23p^2$	15 **P** $3s^23p^3$	16 **S** $3s^23p^4$	17 **Cl** $3s^23p^5$	18 **Ar** $3s^23p^6$
4	19 **K** $4s^1$	20 **Ca** $4s^2$	21 **Sc** $4s^23d^1$	22 **Ti** $4s^23d^2$	23 **V** $4s^23d^3$	24 **Cr** $4s^13d^5$	25 **Mn** $4s^23d^5$	26 **Fe** $4s^23d^6$	27 **Co** $4s^23d^7$	28 **Ni** $4s^23d^8$	29 **Cu** $4s^13d^{10}$	30 **Zn** $4s^23d^{10}$	31 **Ga** $4s^24p^1$	32 **Ge** $4s^24p^2$	33 **As** $4s^24p^3$	34 **Se** $4s^24p^4$	35 **Br** $4s^24p^5$	36 **Kr** $4s^24p^6$
5	37 **Rb** $5s^1$	38 **Sr** $5s^2$	39 **Y** $5s^24d^1$	40 **Zr** $5s^24d^2$	41 **Nb** $5s^14d^4$	42 **Mo** $5s^14d^5$	43 **Tc** $5s^24d^5$	44 **Ru** $5s^14d^7$	45 **Rh** $5s^14d^8$	46 **Pd** $4d^{10}$	47 **Ag** $5s^14d^{10}$	48 **Cd** $5s^24d^{10}$	49 **In** $5s^25p^1$	50 **Sn** $5s^25p^2$	51 **Sb** $5s^25p^3$	52 **Te** $5s^25p^4$	53 **I** $5s^25p^5$	54 **Xe** $5s^25p^6$
6	55 **Cs** $6s^1$	56 **Ba** $6s^2$	57 **La*** $6s^25d^1$	72 **Hf** $6s^25d^2$	73 **Ta** $6s^25d^3$	74 **W** $6s^25d^4$	75 **Re** $6s^25d^5$	76 **Os** $6s^25d^6$	77 **Ir** $6s^25d^7$	78 **Pt** $6s^15d^9$	79 **Au** $6s^15d^{10}$	80 **Hg** $6s^25d^{10}$	81 **Tl** $6s^26p^1$	82 **Pb** $6s^26p^2$	83 **Bi** $6s^26p^3$	84 **Po** $6s^26p^4$	85 **At** $6s^26p^5$	86 **Rn** $6s^26p^6$
7	87 **Fr** $7s^1$	88 **Ra** $7s^2$	89 **Ac**** $7s^26d^1$	104 **Rf** $7s^26d^2$	105 **Db** $7s^26d^3$	106 **Sg** $7s^26d^4$	107 **Bh** $7s^26d^5$	108 **Hs** $7s^26d^6$	109 **Mt** $7s^26d^7$	110 **Ds** $7s^26d^8$	111 **Rg** $7s^26d^9$	112 **Cn** $7s^26d^{10}$	113 $7s^27p^1$	114 **Fl** $7s^27p^2$	115 $7s^27p^3$	116 **Lv** $7s^27p^4$	117 $7s^27p^5$	118 $7s^27p^6$

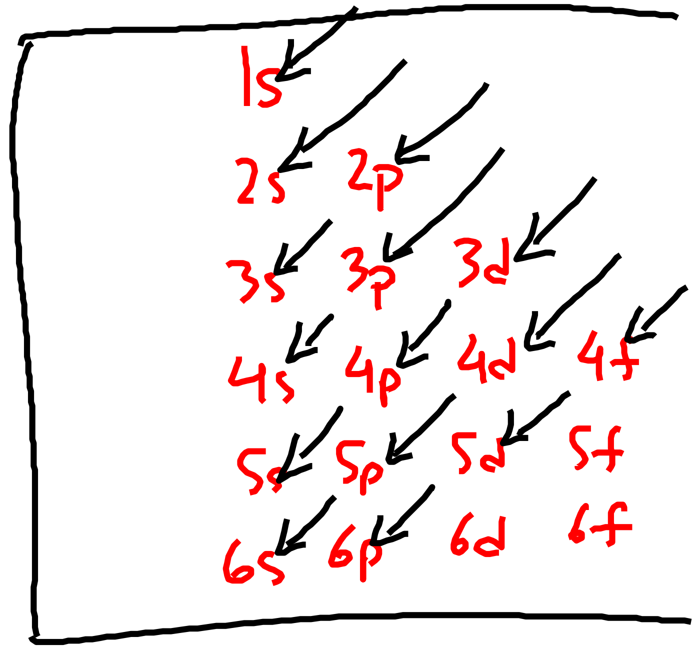
Inner Transition Elements (f block)

Period	Series														
6	*Lanthanides	58 **Ce** $6s^24f^15d^1$	59 **Pr** $6s^24f^3$	60 **Nd** $6s^24f^4$	61 **Pm** $6s^24f^5$	62 **Sm** $6s^24f^6$	63 **Eu** $6s^24f^7$	64 **Gd** $6s^24f^75d^1$	65 **Tb** $6s^24f^9$	66 **Dy** $6s^24f^{10}$	67 **Ho** $6s^24f^{11}$	68 **Er** $6s^24f^{12}$	69 **Tm** $6s^24f^{13}$	70 **Yb** $6s^24f^{14}$	71 **Lu** $6s^24f^{14}5d^1$
7	**Actinides	90 **Th** $7s^26d^2$	91 **Pa** $7s^25f^26d^1$	92 **U** $7s^25f^36d^1$	93 **Np** $7s^25f^46d^1$	94 **Pu** $7s^25f^6$	95 **Am** $7s^25f^7$	96 **Cm** $7s^25f^76d^1$	97 **Bk** $7s^25f^9$	98 **Cf** $7s^25f^{10}$	99 **Es** $7s^25f^{11}$	100 **Fm** $7s^25f^{12}$	101 **Md** $7s^25f^{13}$	102 **No** $7s^25f^{14}$	103 **Lr** $7s^25f^{14}6d^1$

Sample Problem 8.2

Give the **condensed** and **full** electron configurations for the following elements:

(a) Technetium (Tc; Z = 43)   **(b)** Lead (Pb; Z = 82)



SOLUTION

(a) For Tc (Z = 43)

Condensed: [Kr]$5s^2 4d^5$

Full: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^5$

(b) For Pb (Z = 82)

Condensed: [Xe] $6s^2 4f^{14} 5d^{10} 6p^2$

Full: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^2$


Thank you!