General Chemistry I

NAC 1101

Lecture 4

Electronic structure of Atoms

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Outline

- Wave Behavior of Matter (de Broglie)
- Heisenberg’s Uncertainty Principle
- Schrödinger’s wave equation
- Orbitals
- Quantum Numbers
- Many-Electron Atoms
- Pauli exclusion principle
- Electron Configurations
- Abnormal Electron Configurations

Wave Behavior of Matter

Louis de Broglie (1892–1987): matter waves

If radiant energy could, under appropriate conditions, behave as though it were a stream of particles (photons), could matter, under appropriate conditions, possibly show the properties of a wave?

- suggested that an electron moving about the nucleus of an atom behaves like a wave and therefore has a λ that depends on its mass, m, and on its velocity, v:

$$\lambda = \frac{h}{mv}$$

The quantity mv for any object is called its **momentum**

Exercise

- What is the wavelength of an electron moving with a speed of $5.97 \times 10^{6}\, m\, s^{-1}$? The mass of the electron is $9.11 \times 10^{-31} kg$.

Answer

$$\lambda = \frac{h}{mv}$$

$$= \frac{(6.626 \times 10^{-34}\, J.s)}{(9.11 \times 10^{-31}\, kg)(5.97 \times 10^{6}\, m\, s^{-1})} \times \left(\frac{1\, kg\, m^{2} s^{-2}}{1\, J}\right)$$

$$= 1.22 \times 10^{-10}\, m = 0.122\, nm = 122\, pm$$

Close to X-rays

Heisenberg's Uncertainty Principle

It is impossible for us to know simultaneously both the exact momentum of an electron and its exact location in space.

Math. he related the uncertainty in position, Δx, and the uncertainty in momentum, $\Delta(mv)$, to a quantity involving the Planck constant:

$$\Delta x \cdot \Delta(mv) \geq \frac{h}{4\pi}$$

The result is a model that precisely describes the energy of the electron while describing its location not precisely but rather in terms of probabilities.

Exercise

- If the electron has a mass of $9.11 \times 10^{-31} kg$ and moves at an average speed of about $5 \times 10^{6}\, m\, s^{-1}$ with an uncertainty of 1% in a hydrogen atom. Calculate the uncertainty in the position of this electron.

Answer $\Delta v = 0.01 \times 5 \times 10^{6}\, m\, s^{-1} = 5 \times 10^{4}\, m\, s^{-1}$

Assume that this is the only important source of uncertainty in the momentum $\Rightarrow \Delta(mv) = m\,\Delta v$

$$\Delta x \geq \frac{h}{4\pi \cdot m\, \Delta v} = \frac{(6.626\times 10^{-34}\ J.s)}{4\pi (9.11\times 10^{-31}\ kg)(5\times 10^{4}\ m\, s^{-1})} \times \left(\frac{1\ kg\, m^{2} s^{-2}}{1\ J}\right)$$

$$\Delta x \geq 1 \times 10^{-9}\ m$$

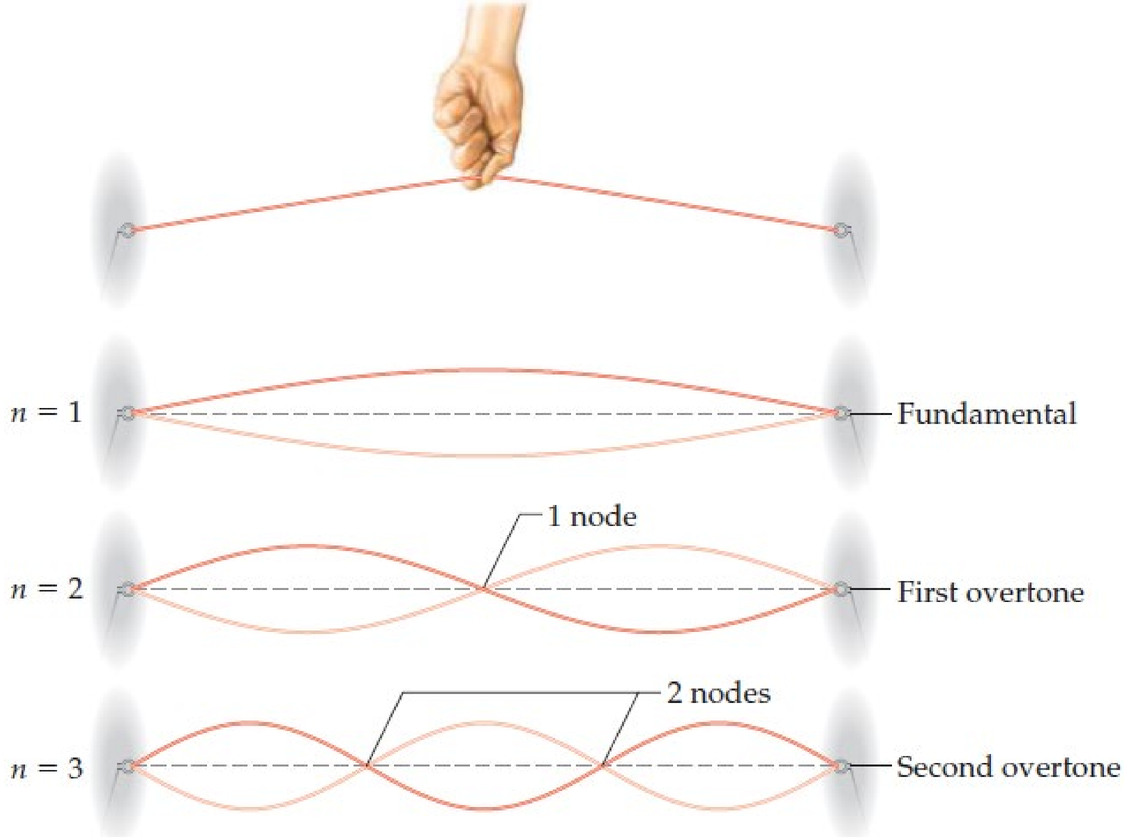
Because the diameter of a H atom is about $1 \times 10^{-10}\ m$, the uncertainty in the position of the electron in the atom is an order of magnitude greater than the size of the atom. Thus, we have essentially no idea where the electron is in the atom.

Schrödinger’s wave equation

- incorporates both the wave-like and particle-like behaviors of the electron.
- opened an approach (quantum mechanics or wave mechanics) to dealing with subatomic particles.
- gives a powerful way to view electronic structure.
- treated the electron in a hydrogen atom like a standing wave (SW: do not travel in space) as the wave on a plucked guitar string.
- SWs (can be either mechanical or electromagnetic) form when two waves of the same frequency and amplitude interfere with each other, either by traveling in opposite directions or by being reflected from a boundary.

Standing Waves

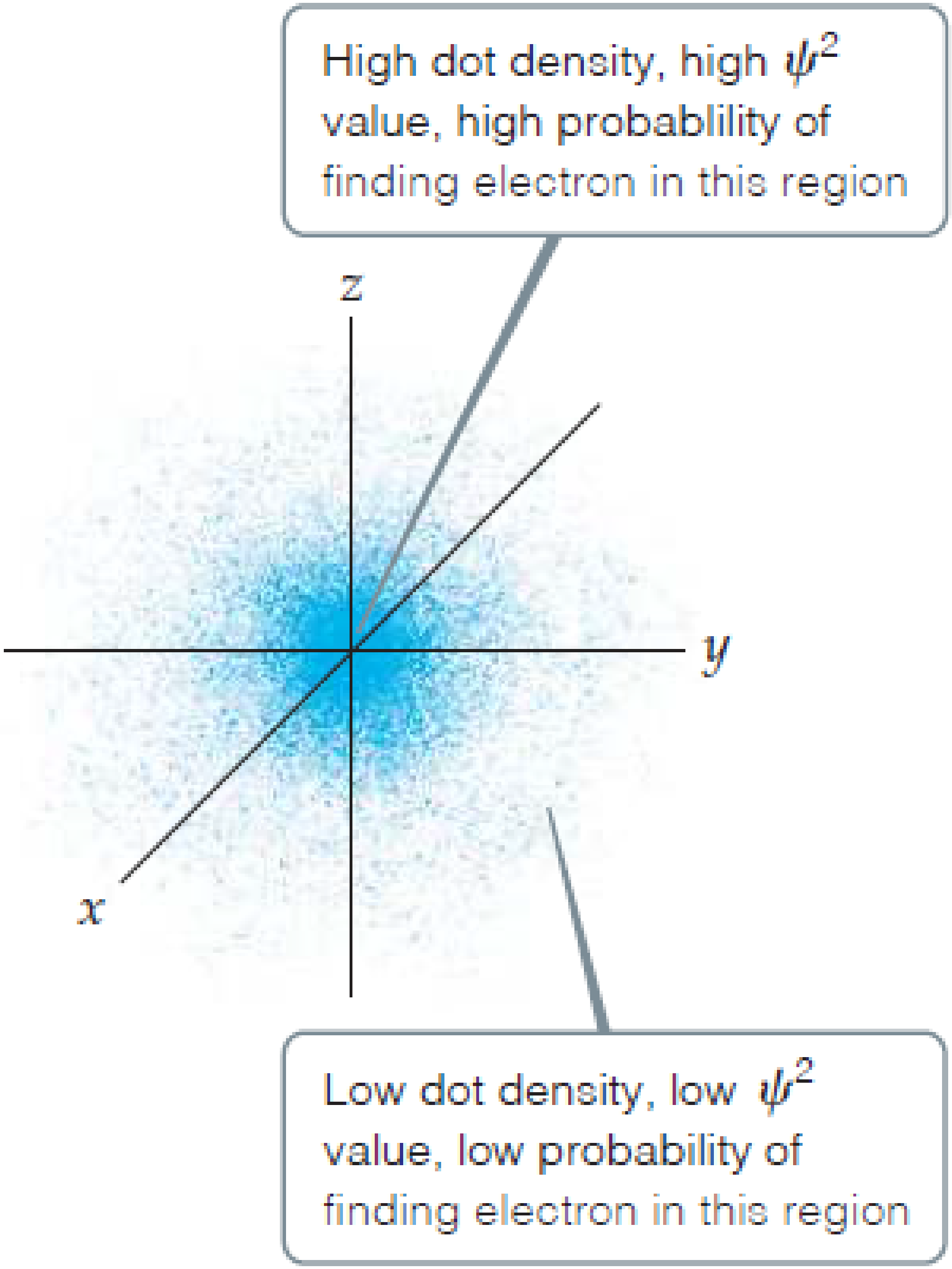
https://www.walter-fendt.de/html5/phen/standingwavereflection_en.htm



- As standing wave that has a fundamental frequency and higher overtones (harmonics), the electron exhibits a lowest-energy standing wave and higher-energy ones.
- As the overtones of the guitar string have nodes, *points where the magnitude of the wave is zero*, so do the waves characteristic of the electron.

wave Functions, ψ

- Solving Schrödinger's equation for the H atom leads to a series of mathematical functions called wave functions (ψ , psi).
- ψ has no direct physical meaning, yet ψ^2 inform about the electron's location when it is in an allowed (as predicted by Bohr) energy state.
- ψ^2 (called the probability density or the electron density) represents the probability that the electron will be found at a certain location of space at a given instant.
- Bohr model assumed that the electron is in a **circular** orbit of some particular radius about the nucleus. In contrast, the quantum-mechanical model assumed the electron's location cannot be described so simply.



High dot density, high ψ^2 value, high probablility of finding electron in this region

Low dot density, low ψ^2 value, low probability of finding electron in this region

An atom consists of a nucleus surrounded by an electron cloud.

Orbitals

- The solution to Schrödinger’s equation for H atom yields a set of ψ called orbitals. Each orbital has a characteristic shape and energy.
- For example, the lowest-energy orbital in the H atom has the spherical shape and an energy of -2.18×10^{-18} J.
- Note that an orbital that is visualized as “electron clouds” according to the quantum- mechanical model describes electrons in terms of probabilities because the motion of the electron in an atom cannot be precisely determined (Heisenberg’s uncertainty principle).
- This is not the same as an orbit (as in Bohr model), which visualizes the electron moving in a physical orbit, like a planet around a star.

Quantum Numbers

Principal quantum number, n

- n can have positive integral values 1, 2, 3, ... ∞
- As n increases, the orbital becomes larger, and the electron spends more time farther from the nucleus.
- An increase in n also means that the electron has a higher energy and is therefore less tightly bound to the nucleus.

- For the hydrogen atom, as in the Bohr model:

$$E_n = (-2.18 \times 10^{-18}) \left(\frac{1}{n^2}\right)$$

Angular momentum quantum number, l

- l can have integral values from 0 to $(n-1)$ for each value of n. l defines the shape of the orbital.
- The value of l for a particular orbital is generally designated by the letters s (sharp), p (principal), d (diffuse), and f (fundamental), corresponding to l values of 0, 1, 2, and 3, respectively.

Value of l	0	1	2	3
Letter used	s	p	d	f

Magnetic quantum number, m_l

- can have integral values between $-l$ and l, including **zero**.
- m_l describes the orientation of the orbital in space.

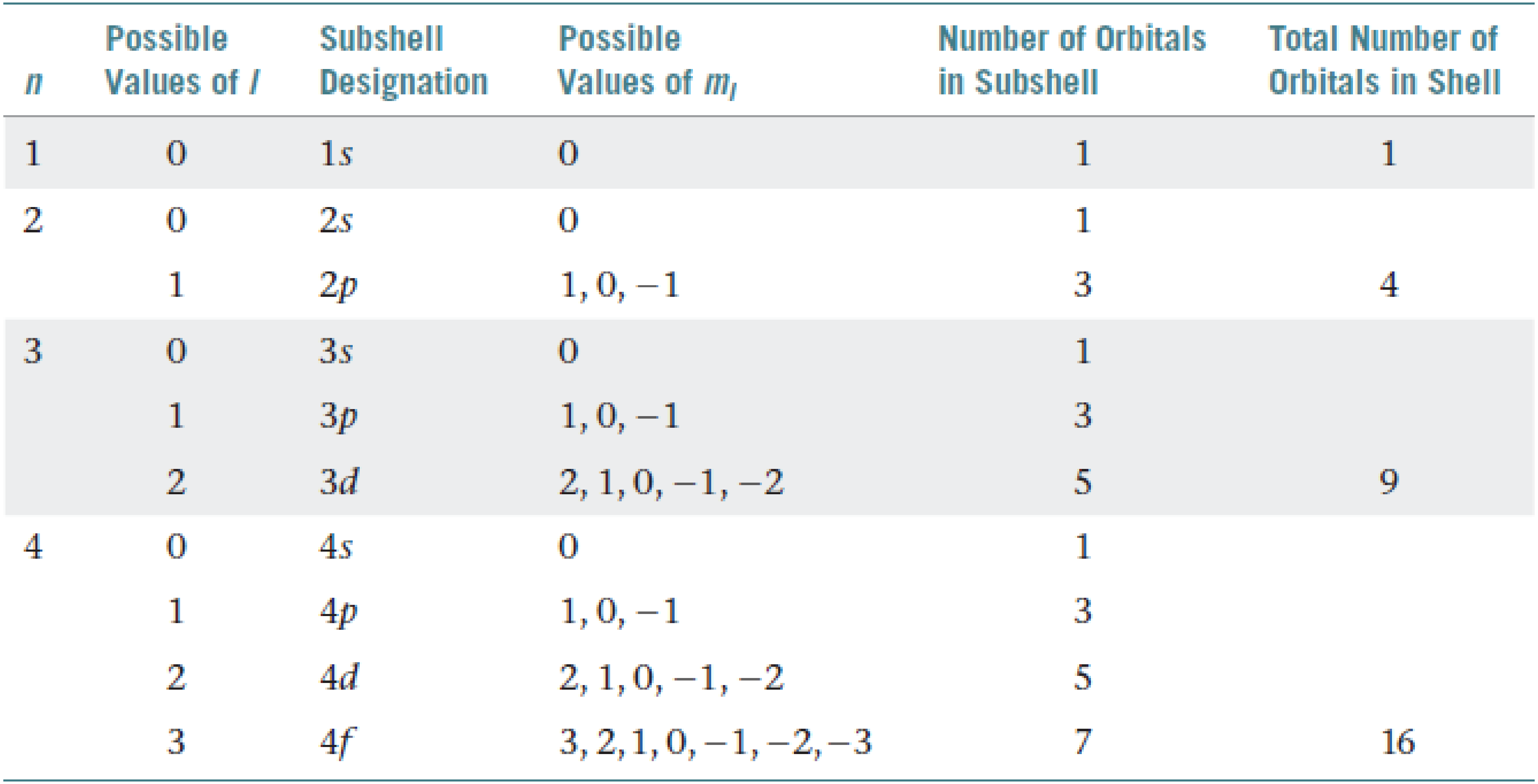
Shells and subshells

- The collection of orbitals with the same value of n is called an electron shell. All the orbitals that have n = 3, for example, are said to be in the third shell.

- The set of orbitals that have the same n and l values is called a subshell. Each subshell is designated by a number (the value of n) and a letter (s, p, d, or f, corresponding to the value of l).

- For example, the orbitals that have n = 3 and l = 2 are called 3d orbitals and are in the 3d subshell.

Relationship among Values of n, l, and m_l



n	Possible Values of l	Subshell Designation	Possible Values of m_l	Number of Orbitals in Subshell	Total Number of Orbitals in Shell
1	0	1*s*	0	1	1
2	0	2*s*	0	1	
	1	2*p*	1, 0, −1	3	4
3	0	3*s*	0	1	
	1	3*p*	1, 0, −1	3	
	2	3*d*	2, 1, 0, −1, −2	5	9
4	0	4*s*	0	1	
	1	4*p*	1, 0, −1	3	
	2	4*d*	2, 1, 0, −1, −2	5	
	3	4*f*	3, 2, 1, 0, −1, −2, −3	7	16

Important Observations

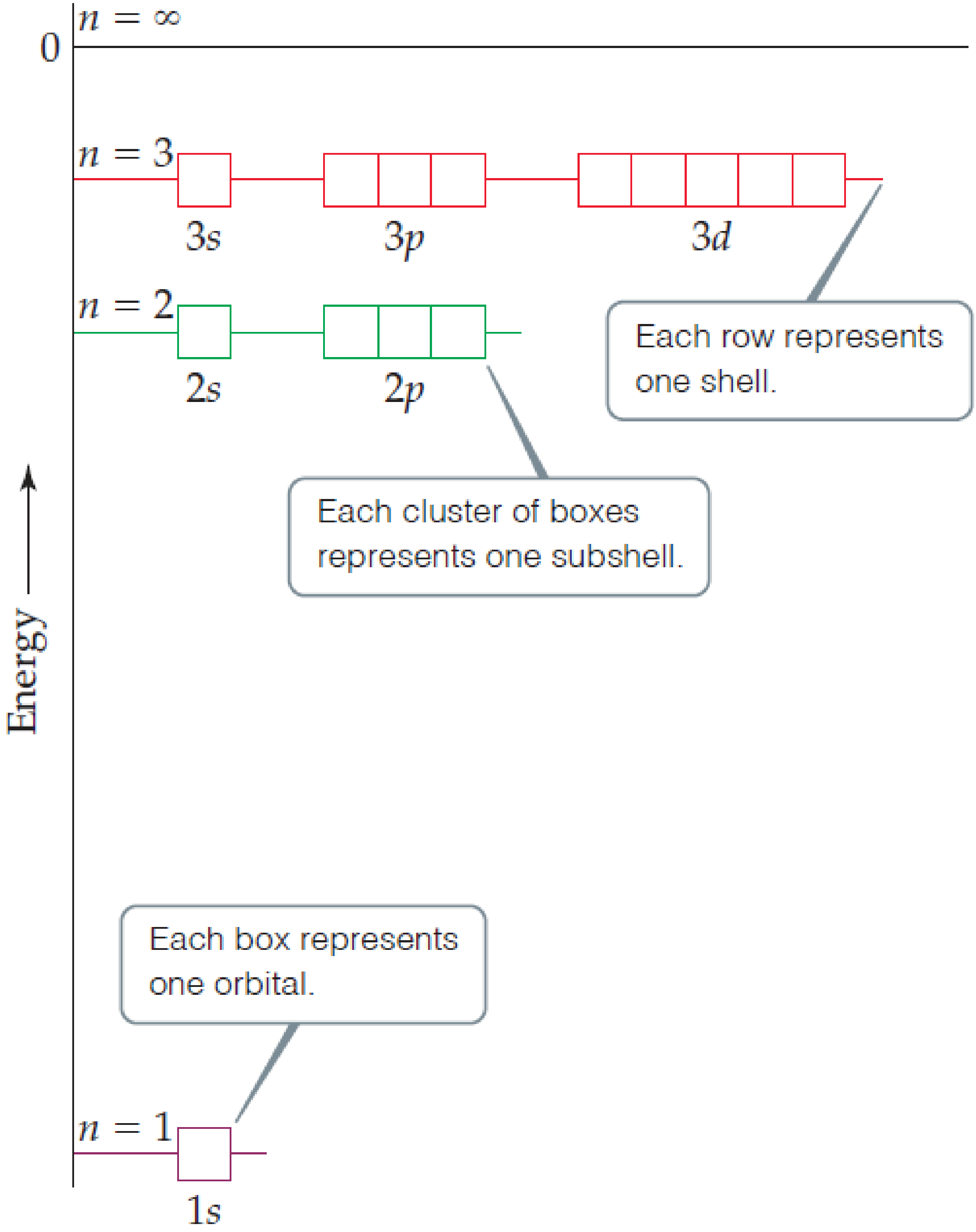
- The shell with n consists of exactly n subshells. Each subshell corresponds to a different allowed value of l from 0 to $(n - 1)$.
    - The first shell ($n = 1$) consists of only one subshell, the $1s$ ($l = 0$).
    - The second shell ($n = 2$) consists of two subshells, the $2s$ ($l = 0$) and $2p$ ($l = 1$).
    - The third shell consists of three subshells, $3s$, $3p$, and $3d$, and so forth.

- Each subshell consists of a specific number of orbitals. Each orbital corresponds to a different allowed value of m_l.
- For a given value of l, there are $(2l+1)$ allowed values of m_l, ranging from $-l$ to $+l$. Thus, each s $(l=0)$ subshell consists of **one** orbital; each p $(l=1)$ subshell consists of **three** orbitals; each d $(l=2)$ subshell consists of **five** orbitals, and so forth.

- The total number of orbitals in a shell is n^2. The resulting number of orbitals for the shells—1, 4, 9, 16—is related to a pattern seen in the periodic table: We see that the number of elements in the rows of the periodic table—2, 8, 18, and 32—equals twice these numbers.

- When e occupies the lowest-energy orbital (1*s*, normally at ordinary T) the H atom is said to be in its ground state.

- When e occupies any other orbital, the atom is in an excited state.



Each row represents one shell.

Each cluster of boxes represents one subshell.

Each box represents one orbital.

Exercise

- Which set of quantum numbers is a valid description of an orbital?

(a) n = 2; l = 2; m_l= 0 **(b)** n = 0; l = 2; m_l = 2

(c) n = 3; l = 2; m_l = 1 **(d)** n = 2; l = 0; m_l = 1

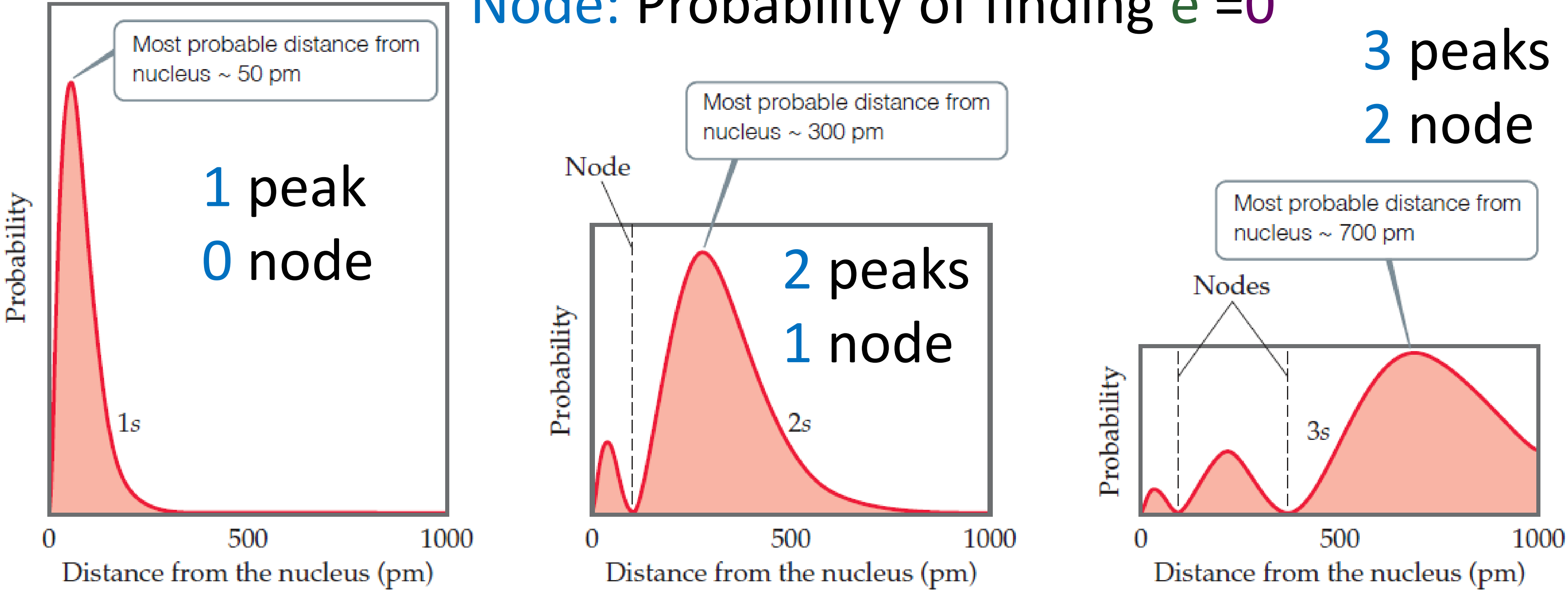
Answer

(c) n = 3; l = 2; m_l = 1

s orbital shape

- All *s* orbitals are spherically symmetric and centered on the nucleus.

Node: Probability of finding e =0



1 peak
0 node

2 peaks
1 node

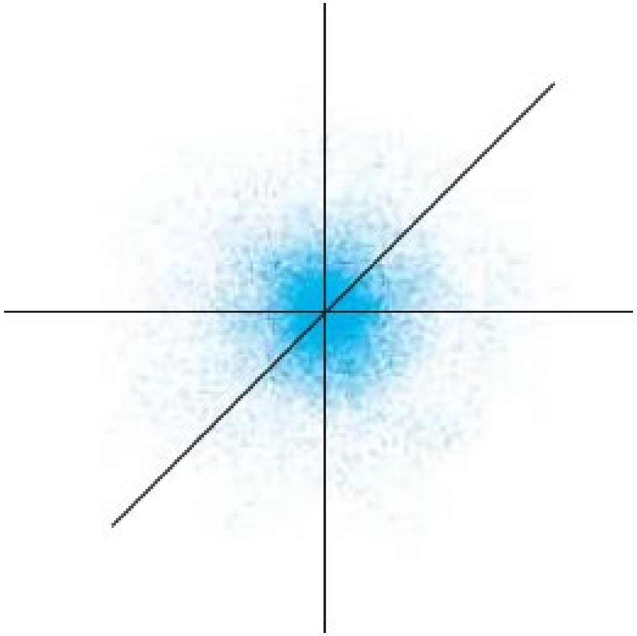
3 peaks
2 node

Radial probability functions for the 1*s*, 2*s*, and 3*s* orbitals of H atom. As *n* increases, the most likely distance at which to find the electron (highest peak) moves farther from nucleus.

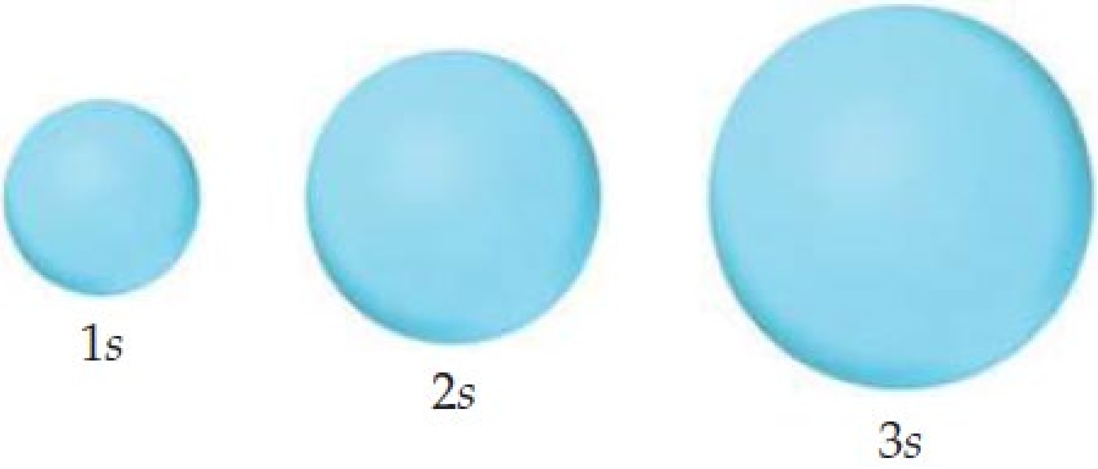
- For an ns orbital, the number of peaks is equal to n, with the outermost peak being larger than inner ones.
- For an ns orbital, the number of nodes is equal to $n-1$.
- As n increases, the electron density becomes more spread out; that is, there is a greater probability of finding the electron further from the nucleus.

Contour representation

A widely used method of representing orbital shape by drawing a boundary surface that encloses some substantial portion, say 90%, of the electron density for the orbital.



(a) An electron density model

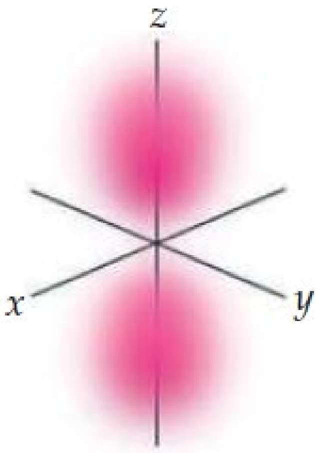


(b) Contour models

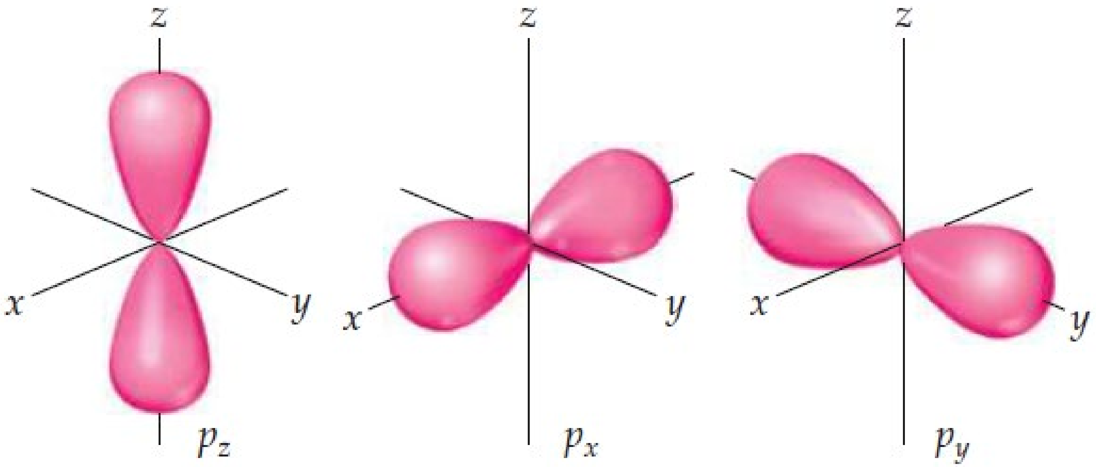
All the orbitals have the same shape, but they differ in size, becoming larger as n increases, reflecting the fact that the electron density becomes more spread out as n increases.

p orbitals shapes

- Each p subshell has 3 orbitals, corresponding to the three allowed values of m_l: −1, 0, and 1.
- The distribution of electron density for a $2p$ orbital is not spherically, but the density is concentrated in two regions on either side of the nucleus, separated by a node at the nucleus.
- We say that this dumbbell-shaped orbital has two lobes.



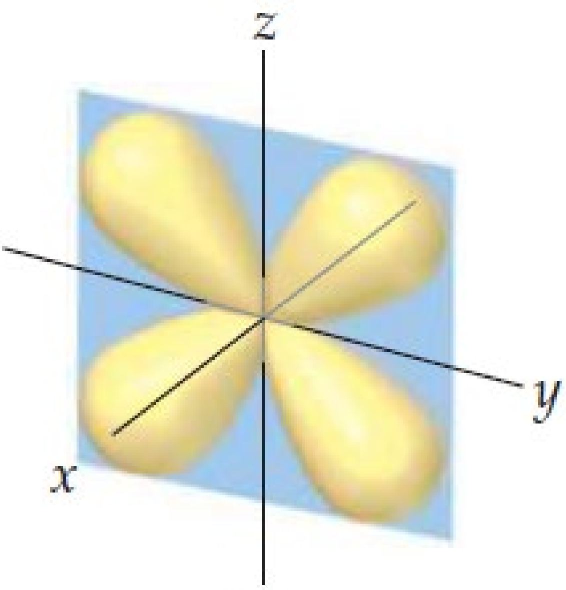
(a) e density model



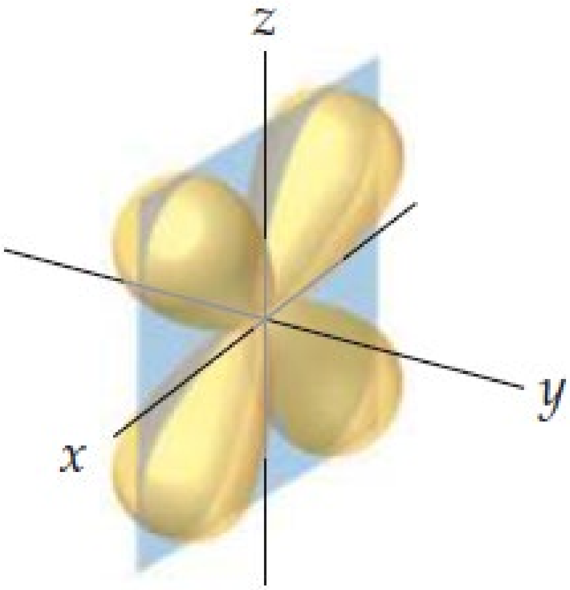
(b) Contour representations

d orbitals shapes

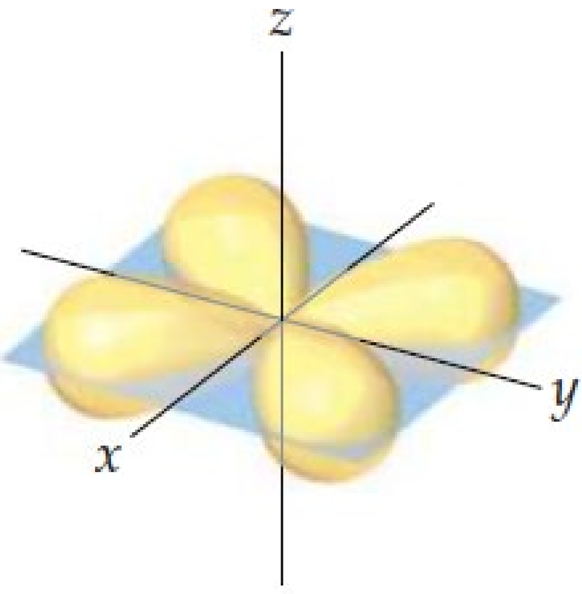
- When n is 3 or greater, d orbitals appear (for which l = 2). There are **five** $3d$ orbitals, **five** $4d$ orbitals, and so forth, because in each shell there are **five** possible values for the m_l quantum number: –2, –1, 0, 1, and 2.
- The different d orbitals in a given shell have different shapes and orientations in space.
- 4 of the d-orbital contour representations have a “four-leaf clover” shape, with four lobes, and each lies primarily in a plane.



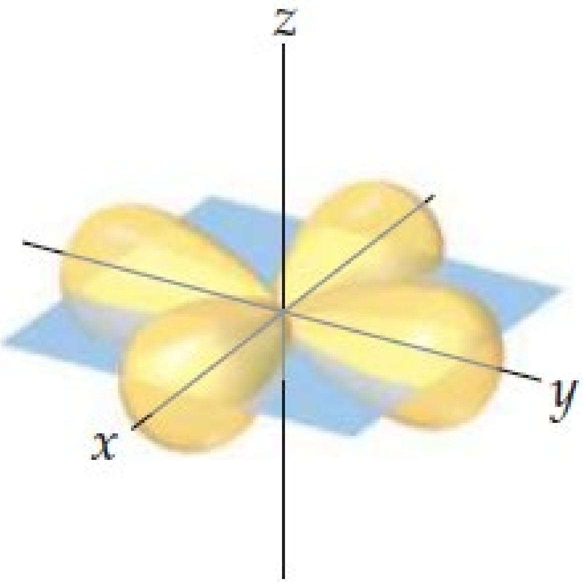
d_{yz}



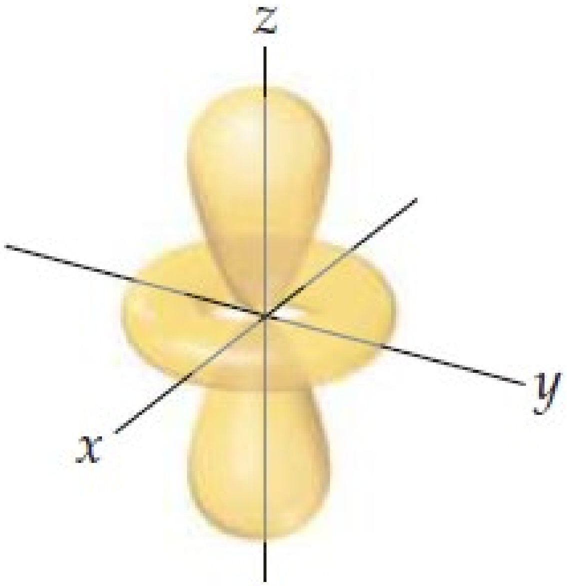
d_{xz}



d_{xy}



$d_{x^2-y^2}$

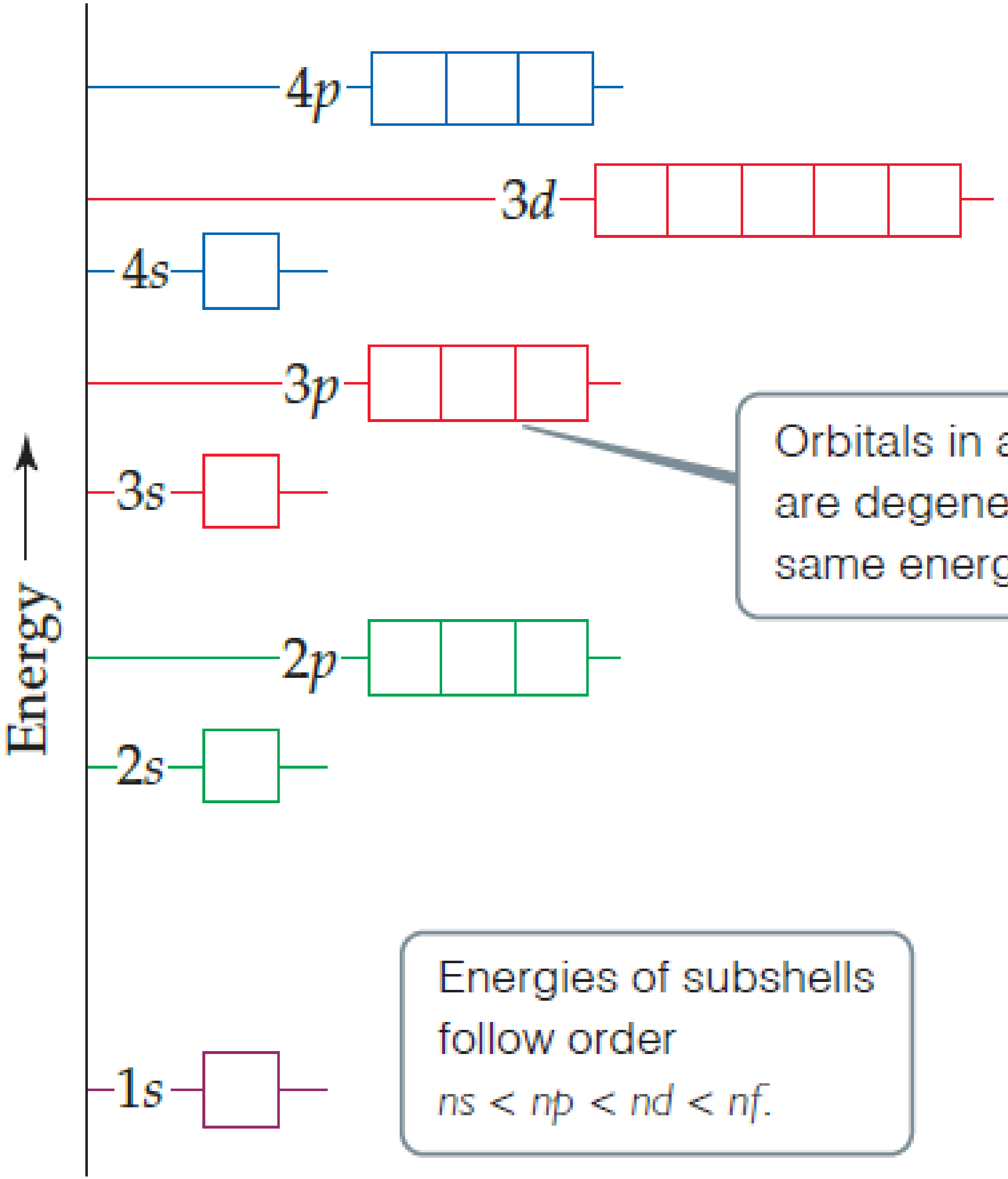


d_{z^2}

- The d_{xy}, d_{xz}, and d_{yz} orbitals lie in the xy, xz, and yz planes, respectively, with the lobes oriented between the axes.
- The lobes of the $d_{x^2-y^2}$ orbital also lie in the xy plane, but the lobes lie **along** the x and y axes.
- The d_{z^2} orbital looks very different from the other four: It has two lobes along the z axis and a “doughnut” in the xy plane.
- Even though the d_{z^2} orbital looks different from the other d orbitals, it has the same energy as the other four d orbitals.

Many-Electron Atoms

- The shapes of the orbitals are the same as those for hydrogen, but their energies are different due to electron–electron repulsions.
    - In hydrogen, the energy of an orbital depends only on its n (3s, 3p, and 3d subshells all have the same energy, **degenerate**).
- For a given value of n, the energy of an orbital increases with increasing value of l.
- All orbitals of a given subshell (as the 5 $3d$ orbitals) have the same energy, just as they do in the hydrogen atom (**degenerate**).



Orbitals in any subshell are degenerate (have same energy).

Energies of subshells follow order *ns* < *np* < *nd* < *nf*.

Electron Spin

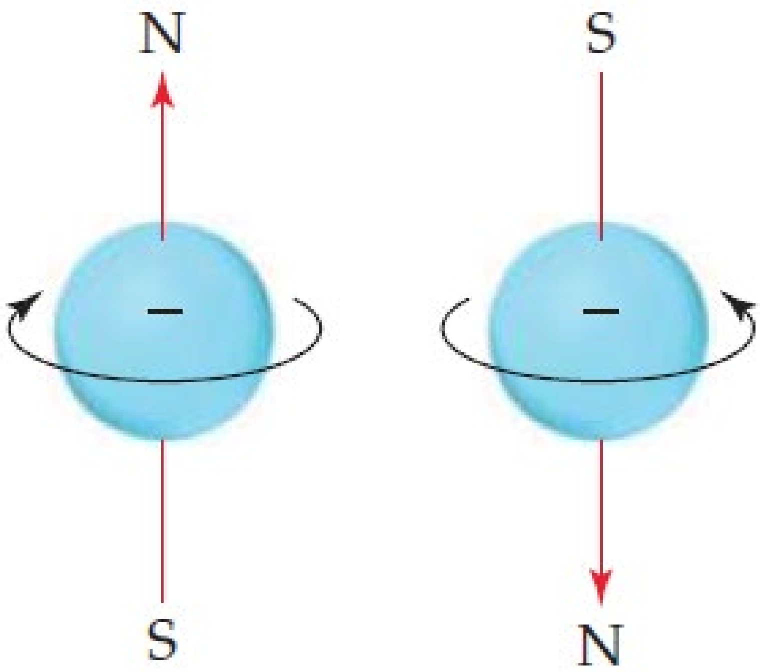
- Spectral lines of many electron atoms that were originally thought to be single were actually closely spaced pairs. $\Rightarrow$ There were twice as many energy levels as there were “supposed” to be.

- In 1925, It is postulated that electrons have an intrinsic property, called electron spin, that causes each electron to behave as if it were a tiny sphere spinning on its own axis.

- By now, it is learnt that electron spin is **quantized**, that led to the assignment of the spin magnetic quantum number, $m_s = {}^{1}/_{2}$ or ${}^{-1}/_{2}$ (indicating the two opposite directions in which the electron can spin).

Electron Spin

- A spinning charge produces a magnetic field. The two opposite directions of spin therefore produce two oppositely directed magnetic fields that lead to the splitting of spectral lines.



Pauli exclusion principle

No two electrons in an atom can have the same set of four quantum numbers n, l, m_l, and m_s.

- If for a given orbital where n, l and m_l are fixed, and to put more than one electron in an orbital and satisfy the Pauli exclusion principle, they should take different our only choice is to assign different m_s values.
- Because there are only two m_s values, an orbital can hold a **maximum** of two electrons, and they must have opposite spins.

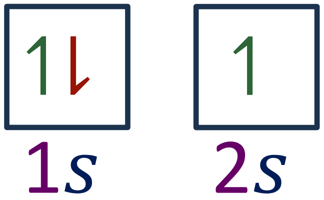
Electron Configurations

- The way electrons are distributed among the various orbitals of an atom.

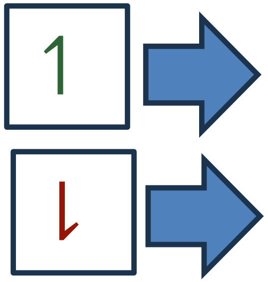
The most stable electron configuration—the **ground state**—is that in which the electrons are in the lowest possible energy states.

Orbitals are filled in order of **increasing** energy, with no more than two electrons per orbital.

$_{3}$**Li**



Orbital diagram



$m_s = {}^{1}/_{2}$ spin-up

$m_s = -{}^{1}/_{2}$ spin-down

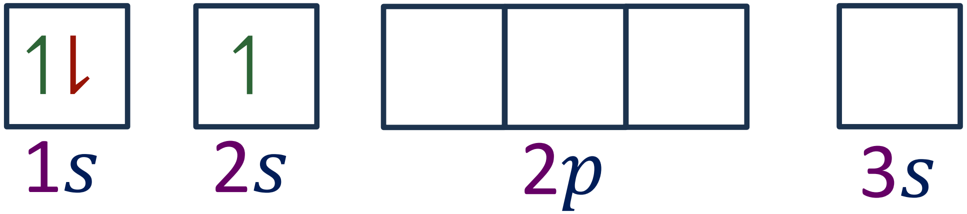
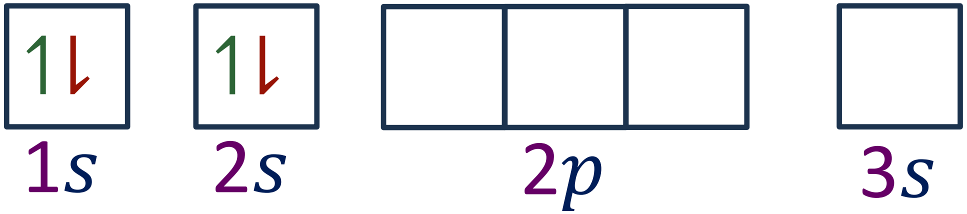
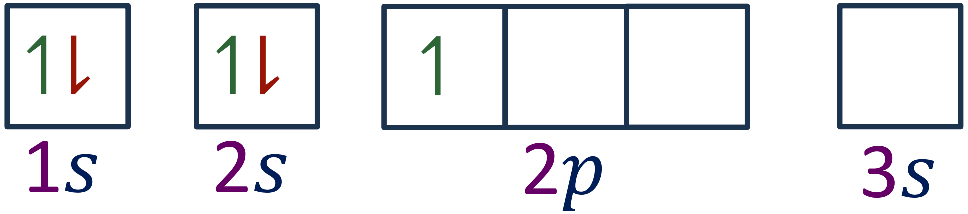
Unpaired electrons

$_{3}$**Li:** $1s^{2}\ 2s^{1}$

Paired electrons

Hund's Rule

- When filling degenerate orbitals, the lowest energy is attained when the number of electrons having the same spin is maximized.

Element	Orbital diagram	e's configuration
$_{3}\mathrm{Li}$		$1s^2\ 2s^1$
$_{4}\mathrm{Be}$		$1s^2\ 2s^2$
$_{5}\mathrm{B}$		$1s^2\ 2s^2\ 2p^1$

Element	Orbital diagram						e's configuration
	$1s$	$2s$	$2p$			$3s$	
$_{6}\mathrm{C}$	↑↓	↑↓	↑	↑			$1s^2\ 2s^2 2p^2$
$_{7}\mathrm{N}$	↑↓	↑↓	↑	↑	↑		$1s^2\ 2s^2 2p^3$
$_{8}\mathrm{O}$	↑↓	↑↓	↑↓	↑	↑		$1s^2\ 2s^2\ 2p^4$
$_{10}\mathrm{Ne}$	↑↓	↑↓	↑↓	↑↓	↑↓		$1s^2\ 2s^2\ 2p^6$
$_{11}\mathrm{Na}$	↑↓	↑↓	↑↓	↑↓	↑↓	↑	$1s^2\ 2s^2\ 2p^6\ 3s^1$

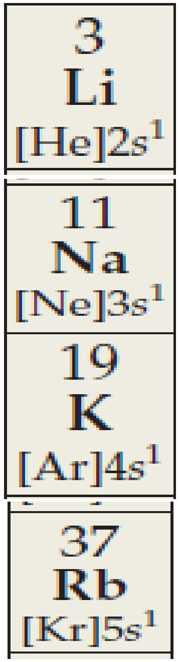
Condensed Electron Configurations

$_{3}\text{Li}: 1s^2\ 2s^1 \Rightarrow {}_{3}\text{Li}: [\text{He}]\ 2s^1$

$_{11}\text{Na}: 1s^2\ 2s^2\ 2p^6\ 3s^1 \Rightarrow {}_{11}\text{Na}: [\text{Ne}]\ 3s^1$

Inner-shell Core electrons

Outer-shell Valence electrons



3
Li
[He]$2s^1$

11
Na
[Ne]$3s^1$

19
K
[Ar]$4s^1$

37
Rb
[Kr]$5s^1$

Elements having the same type of electron configuration in the outermost occupied shell have **similar** chemical properties.

Transition Metals

Element	Orbital diagram						e’s configuration
	$4s$	$3d$					
$_{25}$Mn: [Ar]	↿⇂	↿	↿	↿	↿	↿	[Ar] $4s^2\ 3d^5$
$_{30}$Zn: [Ar]	↿⇂	↿⇂	↿⇂	↿⇂	↿⇂	↿⇂	[Ar] $4s^2\ 3d^{10}$

Lanthanides

- The $4f$ subshell starts to filling in the sixth row of the periodic table after $_{55}$Cs: [Xe] $6s^1$ and $_{56}$Ba: [Xe] $6\,s^2$ with seven degenerate $4\,f$ orbitals corresponding to the seven allowed m_l values (from 3 to -3).

- The 14 elements corresponding to the filling of the $4f$ orbitals are known as either the lanthanide elements or the rare earth elements.

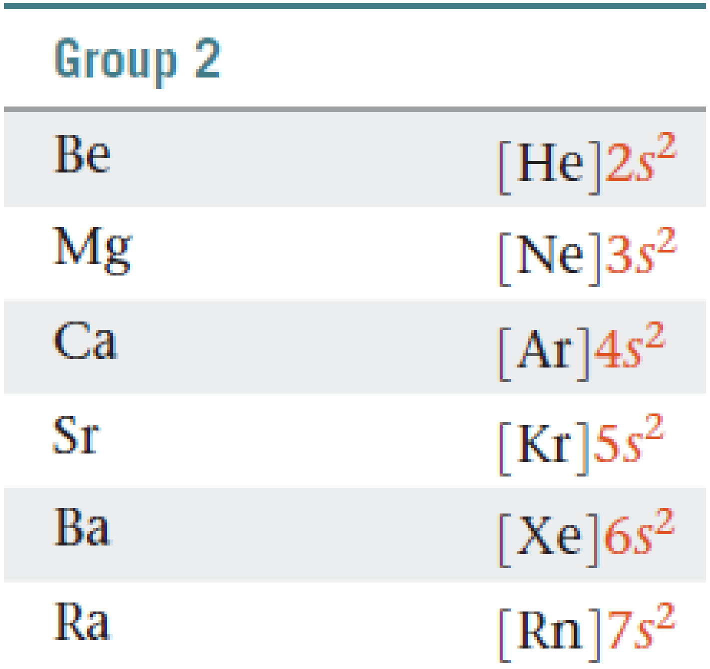
[Xe] $6s^2\ 5d^1$	[Xe] $6s^2\ 5d^1\ 4f^1$	[Xe] $6s^2\ 4f^3$
$_{57}$**La**	$_{58}$**Ce**	$_{59}$**Pr**
Lanthanum	**Cerium**	**Praseodymium**

Actinides

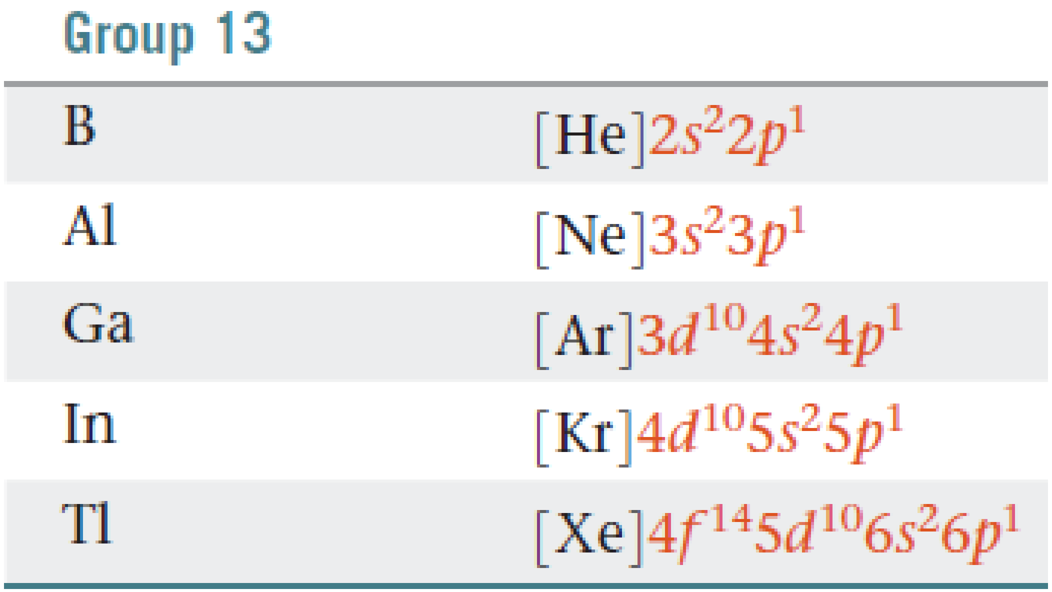
- After the lanthanide series, the third transition element series is completed by the filling of the $5d$ orbitals, followed by the filling of the $6p$ orbitals. This brings us to radon ($_{86}$Rn), the heaviest naturally occurring noble-gas element.
- The final row of the periodic table begins by filling the $7s$ orbitals. The actinide ($_{89}$Ac) elements, of which uranium (U, element 92) and plutonium (Pu, element 94) are the best known, are then built up by completing the $5f$ orbitals.
- All of the actinide elements are radioactive, and most of them are not found in nature.

Electron Configurations & Periodic Table

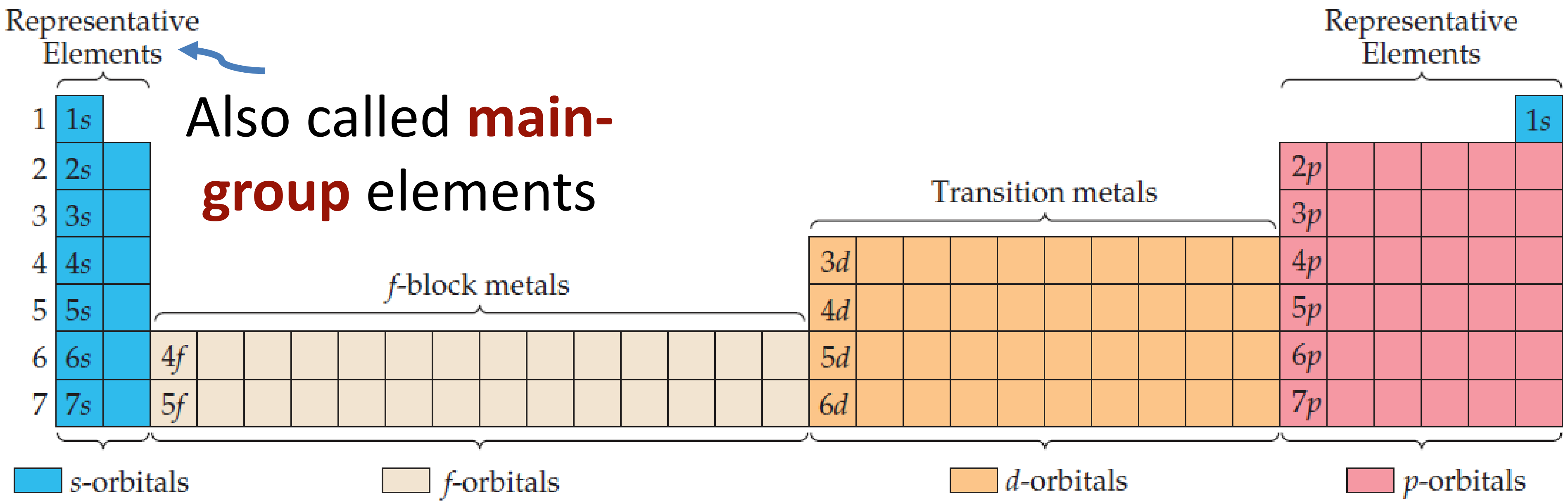
- Elements in the same column of the table have related outer-shell (valence) electron configurations.



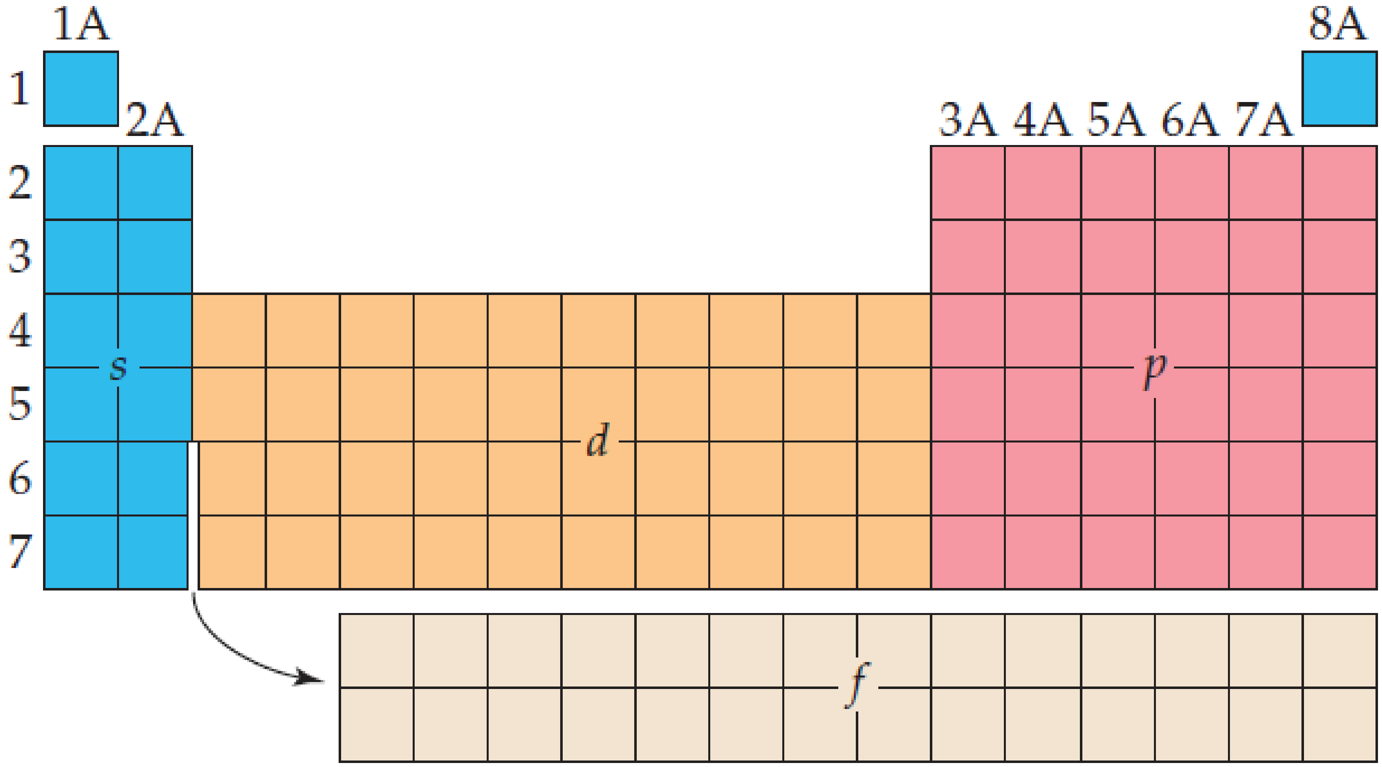
Group 2	
Be	$[\mathrm{He}]2s^2$
Mg	$[\mathrm{Ne}]3s^2$
Ca	$[\mathrm{Ar}]4s^2$
Sr	$[\mathrm{Kr}]5s^2$
Ba	$[\mathrm{Xe}]6s^2$
Ra	$[\mathrm{Rn}]7s^2$



Group 13	
B	$[\mathrm{He}]2s^2 2p^1$
Al	$[\mathrm{Ne}]3s^2 3p^1$
Ga	$[\mathrm{Ar}]3d^{10} 4s^2 4p^1$
In	$[\mathrm{Kr}]4d^{10} 5s^2 5p^1$
Tl	$[\mathrm{Xe}]4f^{14} 5d^{10} 6s^2 6p^1$



Also called **main-group** elements

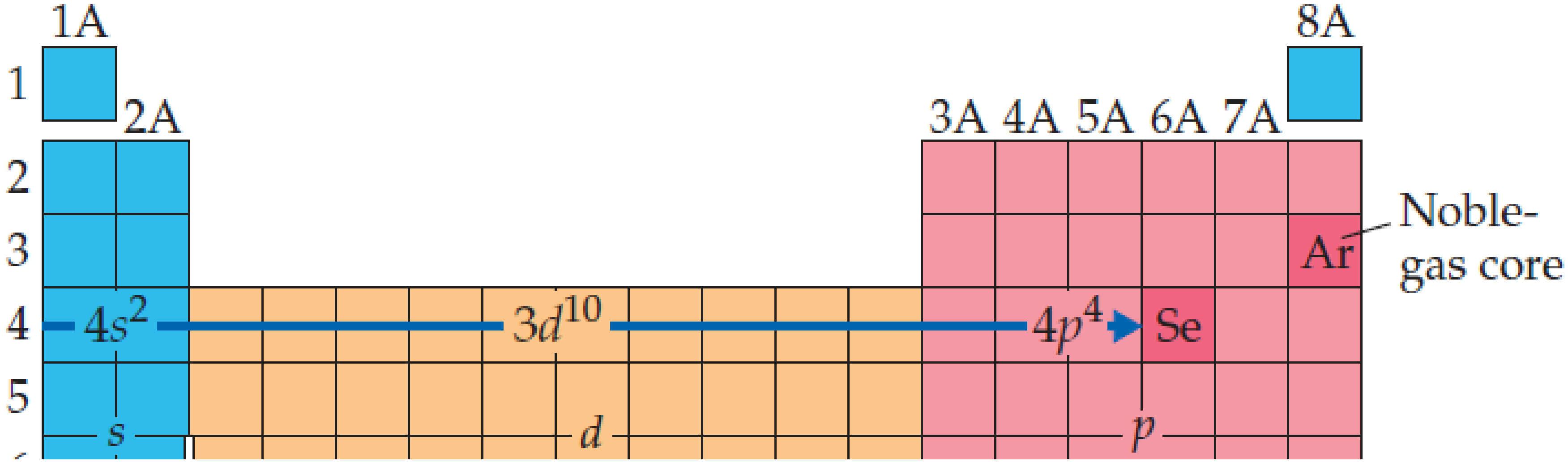


The electron configuration of an element can be known from its position in PT

Exercise

- Use the periodic table to write the electron configuration of selenium (Se, element 34).

Answer

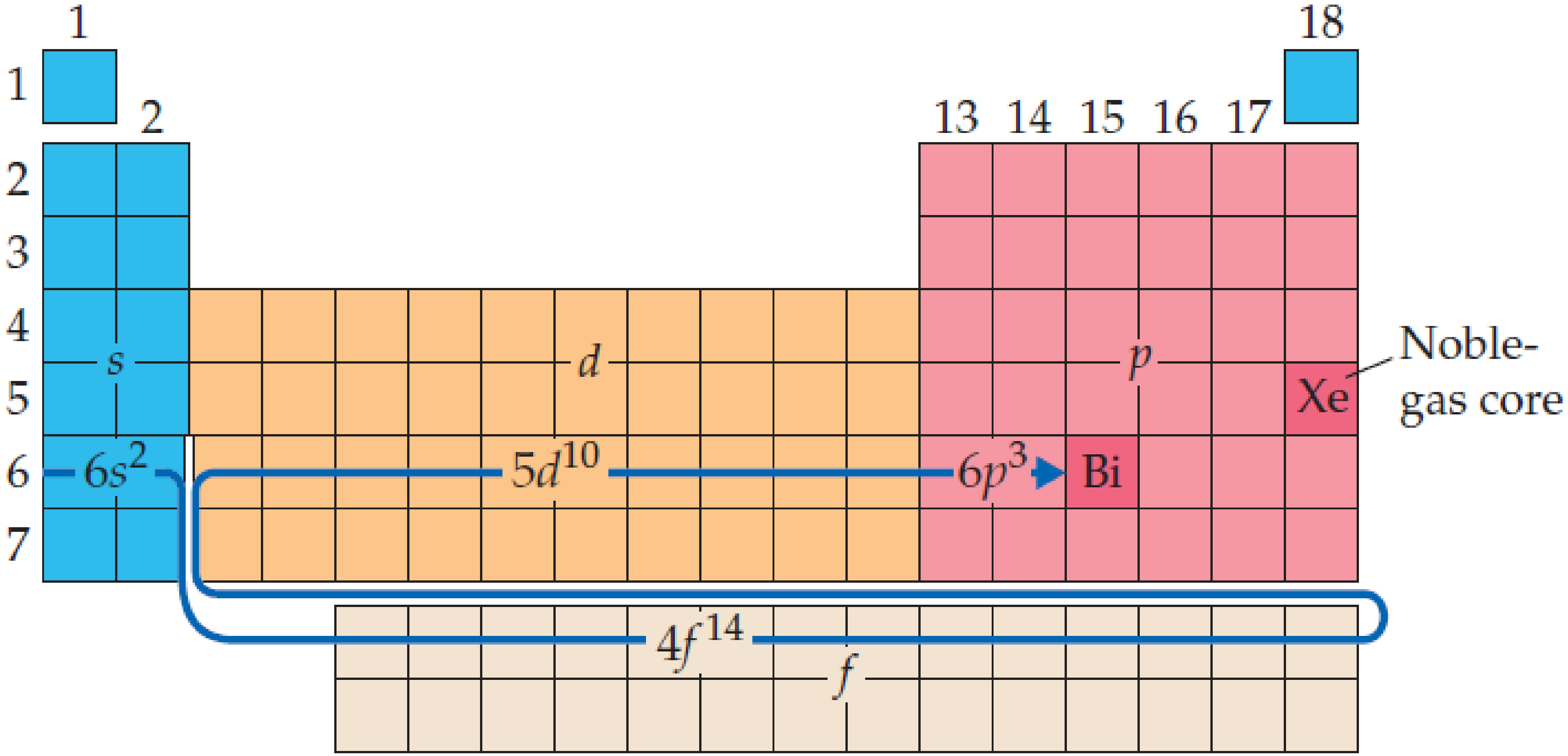


$$_{34}\text{Se: [Ar] } 4s^2\ 3d^{10}\ 4p^4$$

Exercise $_{83}\mathrm{Bi} : [_{54}\mathrm{Xe}]\ 4f^{14}\ 5d^{10}\ 6s^{2}\ 6p^{3}$

- Use the periodic table to write the electron configuration of Bithmus (Bi, element 83).

Answer



$$_{83}\mathrm{Bi} : [_{54}\mathrm{Xe}]\ 6s^{2}4f^{14}\ 5d^{10}\ 6p^{3}$$

For representative elements we do not consider the electrons in completely filled d or f subshells to be valence electrons, and for transition elements we do not consider the electrons in a completely filled f subshell to be valence electrons.

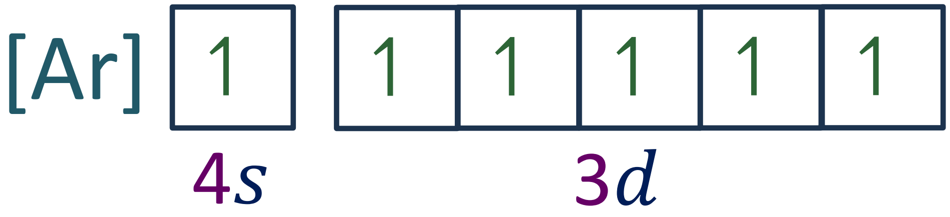
$${}_{17}\mathrm{Cl{:}[Ne]}\ \underbrace{3s^2\ 3p^5} \qquad {}_{35}\mathrm{Br{:}[Ar]}\ 3d^{10}\ \underbrace{4s^2\ 4p^5}$$

7 Valence electrons

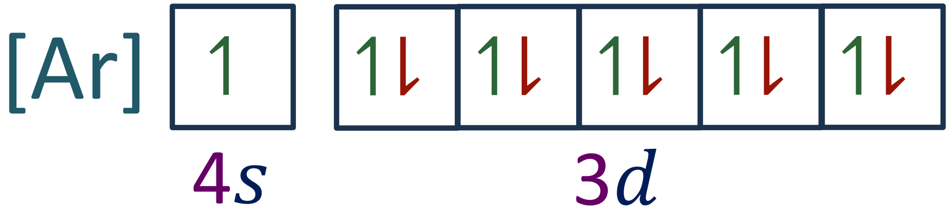
Abnormal Electron Configurations

- Elements in the same column of the table have related outer-shell (valence) electron configurations.

$_{24}$Cr:[Ar] $4s^1\ 3d^5$

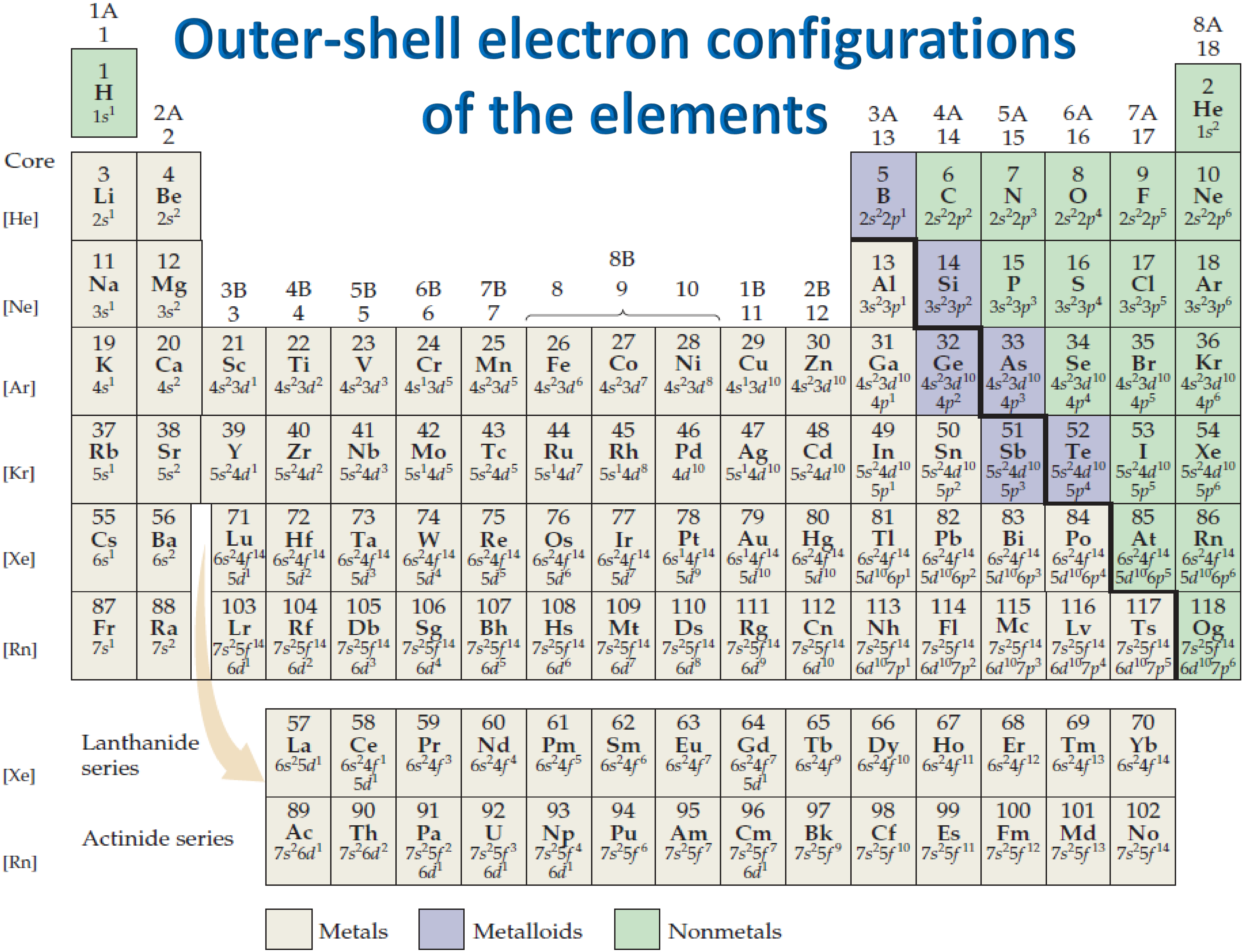


$_{29}$Cu:[Ar] $4s^1\ 3d^{10}$



This is largely a consequence of the closeness of the 3*d* and 4*s* orbital energies. It frequently occurs when there are enough electrons to form precisely half-filled sets of degenerate orbitals (as in chromium) or a completely-filled *d* subshell (as in copper).

Outer-shell electron configurations of the elements



Core	1A	2A	3B	4B	5B	6B	7B	8B			1B	2B	3A	4A	5A	6A	7A	8A
	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
	1 **H** $1s^1$																	2 **He** $1s^2$
[He]	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** $2s^22p^6$
[Ne]	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** $3s^23p^6$
[Ar]	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** $4s^23d^{10}4p^1$	32 **Ge** $4s^23d^{10}4p^2$	33 **As** $4s^23d^{10}4p^3$	34 **Se** $4s^23d^{10}4p^4$	35 **Br** $4s^23d^{10}4p^5$	36 **Kr** $4s^23d^{10}4p^6$
[Kr]	37 **Rb** $5s^1$	38 **Sr** $5s^2$	39 **Y** $5s^24d^1$	40 **Zr** $5s^24d^2$	41 **Nb** $5s^24d^3$	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** $5s^24d^{10}5p^1$	50 **Sn** $5s^24d^{10}5p^2$	51 **Sb** $5s^24d^{10}5p^3$	52 **Te** $5s^24d^{10}5p^4$	53 **I** $5s^24d^{10}5p^5$	54 **Xe** $5s^24d^{10}5p^6$
[Xe]	55 **Cs** $6s^1$	56 **Ba** $6s^2$	71 **Lu** $6s^24f^{14}5d^1$	72 **Hf** $6s^24f^{14}5d^2$	73 **Ta** $6s^24f^{14}5d^3$	74 **W** $6s^24f^{14}5d^4$	75 **Re** $6s^24f^{14}5d^5$	76 **Os** $6s^24f^{14}5d^6$	77 **Ir** $6s^24f^{14}5d^7$	78 **Pt** $6s^14f^{14}5d^9$	79 **Au** $6s^14f^{14}5d^{10}$	80 **Hg** $6s^24f^{14}5d^{10}$	81 **Tl** $6s^24f^{14}5d^{10}6p^1$	82 **Pb** $6s^24f^{14}5d^{10}6p^2$	83 **Bi** $6s^24f^{14}5d^{10}6p^3$	84 **Po** $6s^24f^{14}5d^{10}6p^4$	85 **At** $6s^24f^{14}5d^{10}6p^5$	86 **Rn** $6s^24f^{14}5d^{10}6p^6$
[Rn]	87 **Fr** $7s^1$	88 **Ra** $7s^2$	103 **Lr** $7s^25f^{14}6d^1$	104 **Rf** $7s^25f^{14}6d^2$	105 **Db** $7s^25f^{14}6d^3$	106 **Sg** $7s^25f^{14}6d^4$	107 **Bh** $7s^25f^{14}6d^5$	108 **Hs** $7s^25f^{14}6d^6$	109 **Mt** $7s^25f^{14}6d^7$	110 **Ds** $7s^25f^{14}6d^8$	111 **Rg** $7s^25f^{14}6d^9$	112 **Cn** $7s^25f^{14}6d^{10}$	113 **Nh** $7s^25f^{14}6d^{10}7p^1$	114 **Fl** $7s^25f^{14}6d^{10}7p^2$	115 **Mc** $7s^25f^{14}6d^{10}7p^3$	116 **Lv** $7s^25f^{14}6d^{10}7p^4$	117 **Ts** $7s^25f^{14}6d^{10}7p^5$	118 **Og** $7s^25f^{14}6d^{10}7p^6$

[Xe]	Lanthanide series	57 **La** $6s^25d^1$	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** $6s^24f^{14}$
[Rn]	Actinide series	89 **Ac** $7s^26d^1$	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** $7s^25f^{14}$

⬜ Metals 🟪 Metalloids 🟩 Nonmetals

Thank You