

ELECTRON CONFIGURATIONS

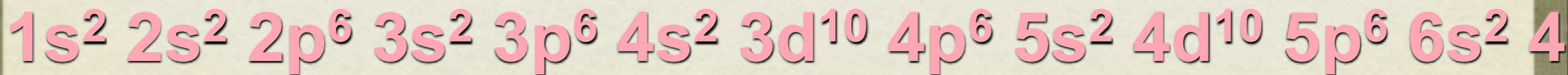
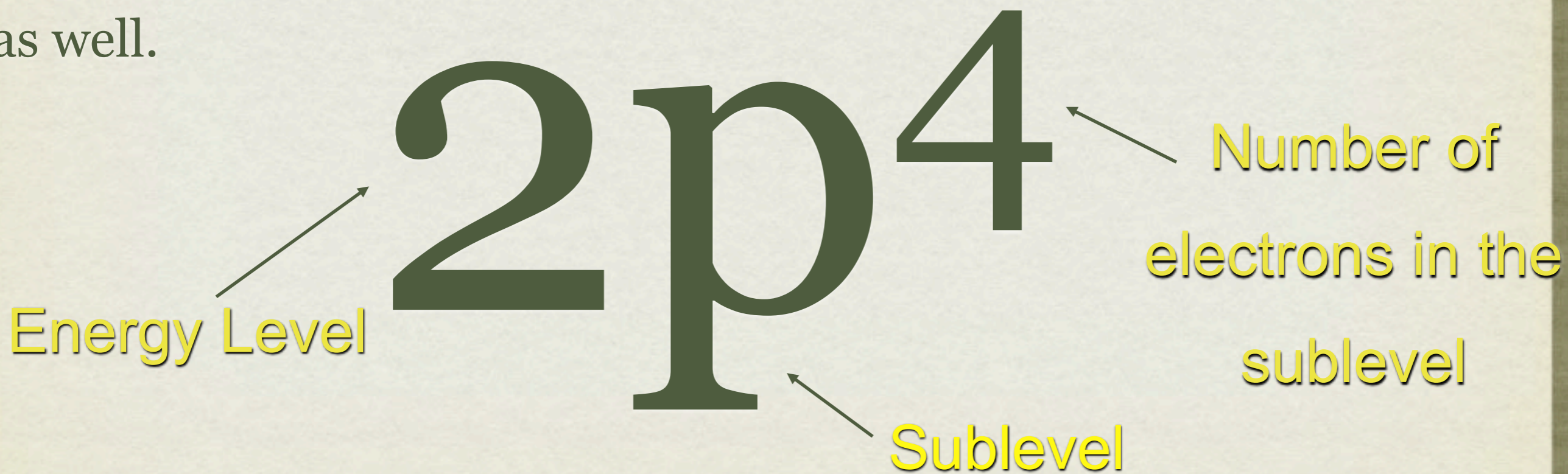
- ELECTRON CONFIGURATIONS,
ORBITAL DIAGRAMS, AUFBAU
PRINCIPLE, HUND'S RULE

REPRESENTING ELECTRONS ...

- Now that you know what an orbital is, you need to be able to use that to describe the electronic nature of an element
- Two ways:
 - Electron configuration is a **concise** way to describe where the electrons are with respect to energy level and sublevel
 - Orbital diagrams are a **visual way** to describe where the electrons are with respect to energy level and sublevel

ELECTRON CONFIGURATIONS

- With BR diagrams, you could say how many electrons are in each shell
- With electron configurations, you can now say not just what **shell** the electrons are in, but also what **sublevel (s, p, d, f)** they are in as well.



AUFBAU PRINCIPLE

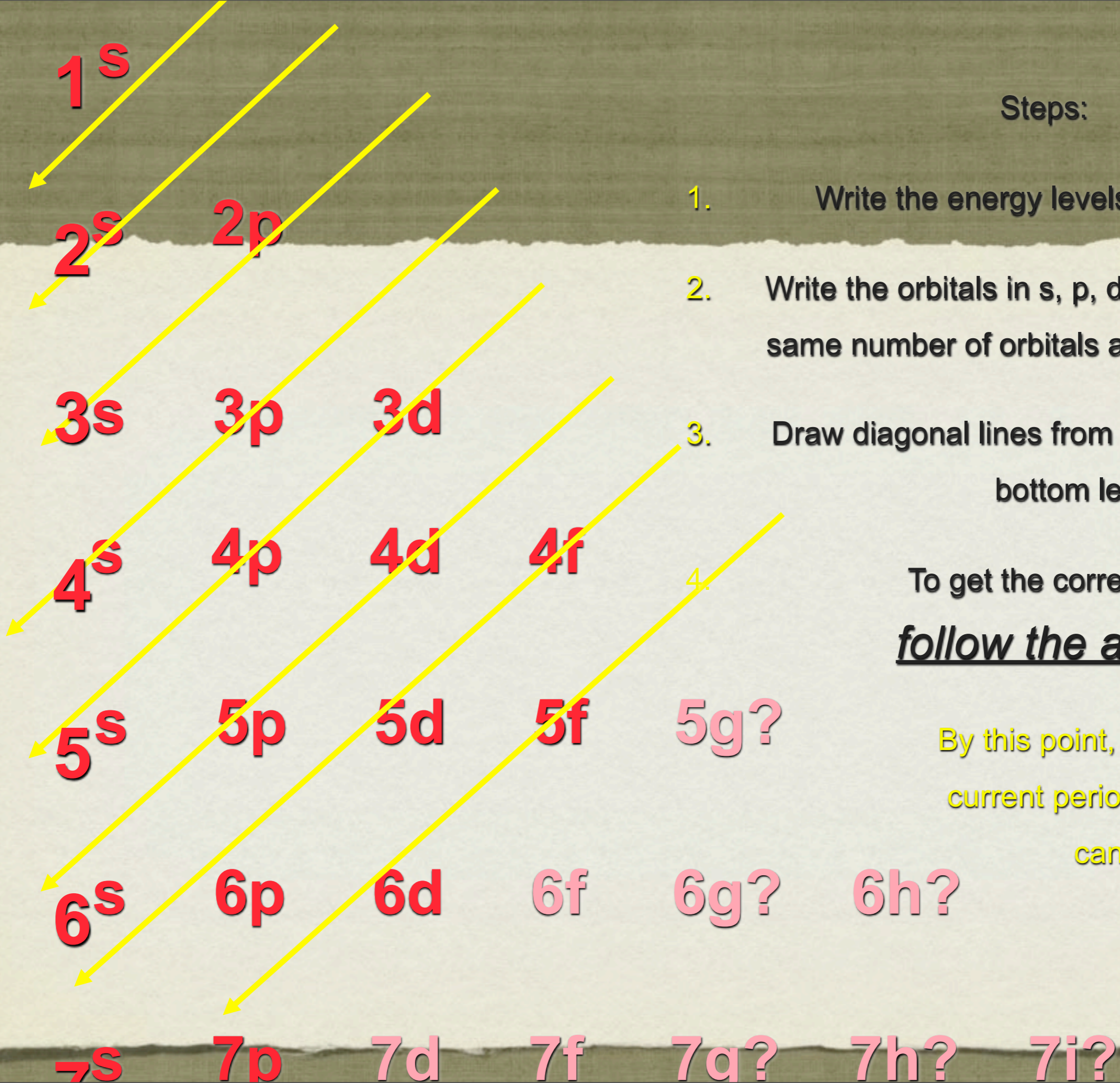
- Aufbau described the filling order of the orbitals (ie what order do the electrons go in)
- He said they must be filled from **lowest energy** to **highest energy**
- Like in BR diagrams, you filled the first shell before you moved on to the second shell...but it gets more complicated!
- Sadly, they don't go numerically: 4S is lower in energy than 3D!

LINK TO QUANTUM

- How do we know how many **sublevels** there are in a shell?
- For example, what **sublevels** exist for first shell ($n=1$)?
 - when $n=1$, $\ell=0$ ONLY so **1s** is the only sublevel
- What about when $n=2$?
 - $\ell=0$ or 1 2s and 2p exist
- What about when $n=3$?
 - $\ell=0$ or 1 or 2 3s and 3p and 3d exist

DIAGONAL RULE

- The diagonal rule is a *memory device* that helps you remember the order of the filling of the orbitals from lowest energy to highest energy
- You won't be given this on your exam!



Steps:

1. Write the energy levels top to bottom.
2. Write the orbitals in s, p, d, f order. Write the same number of orbitals as the energy level.
3. Draw diagonal lines from the top right to the bottom left.

To get the correct order, **follow the arrows!**

By this point, we are past the current periodic table so we can stop.

ORBITALS & THE PERIODIC TABLE

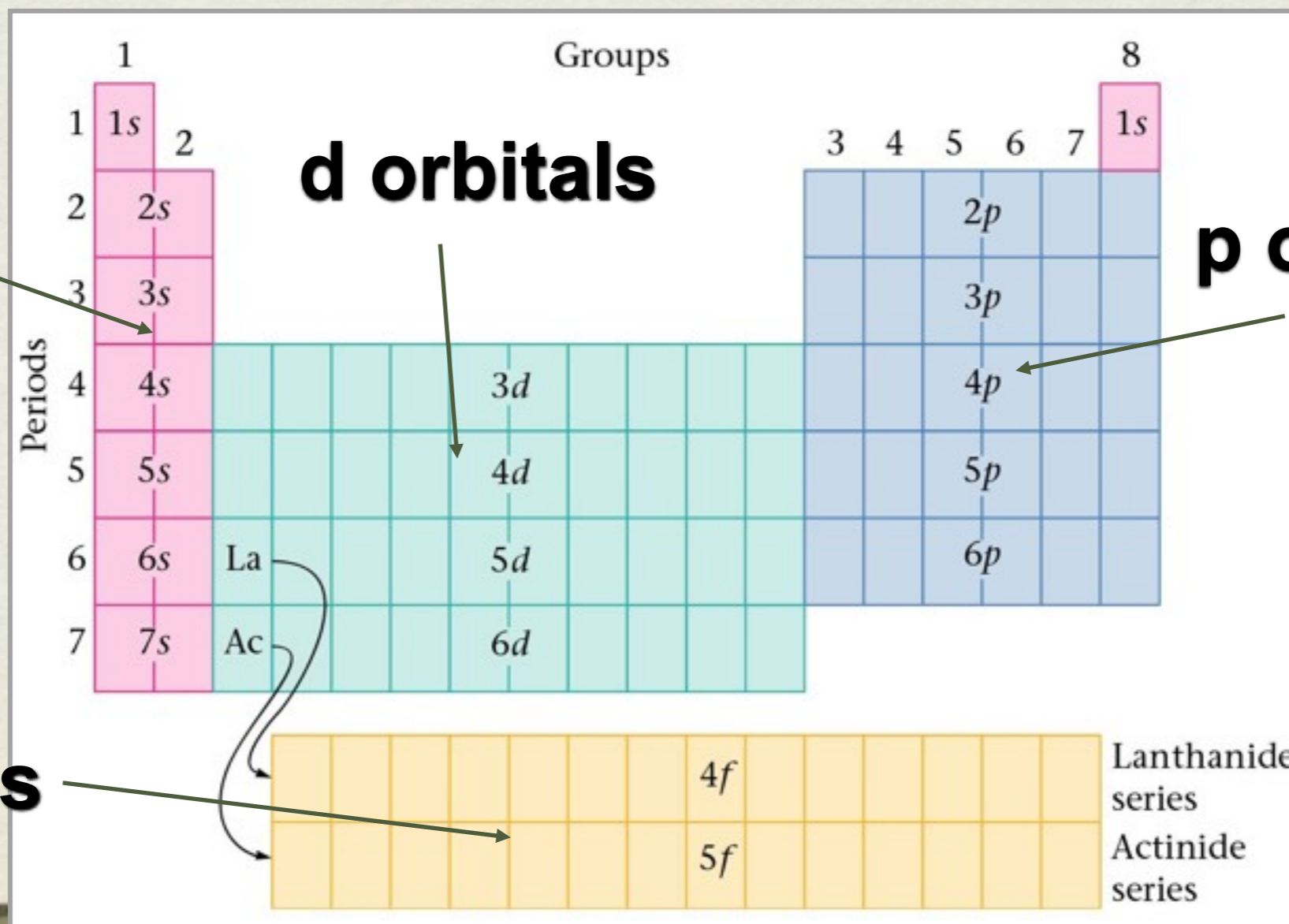
- Orbitals grouped in s, p, d, and f orbitals
- Knowing the blocks is up to you!

s orbitals

d orbitals

p orbitals

f orbitals



ELECTRON CONFIGURATIONS

Guidelines for “Filling” Orbitals

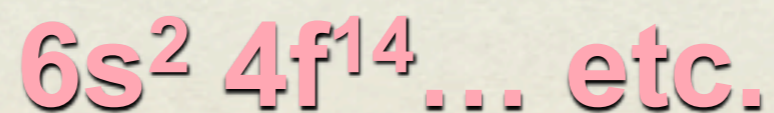
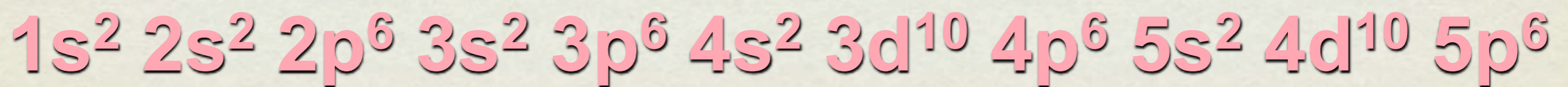
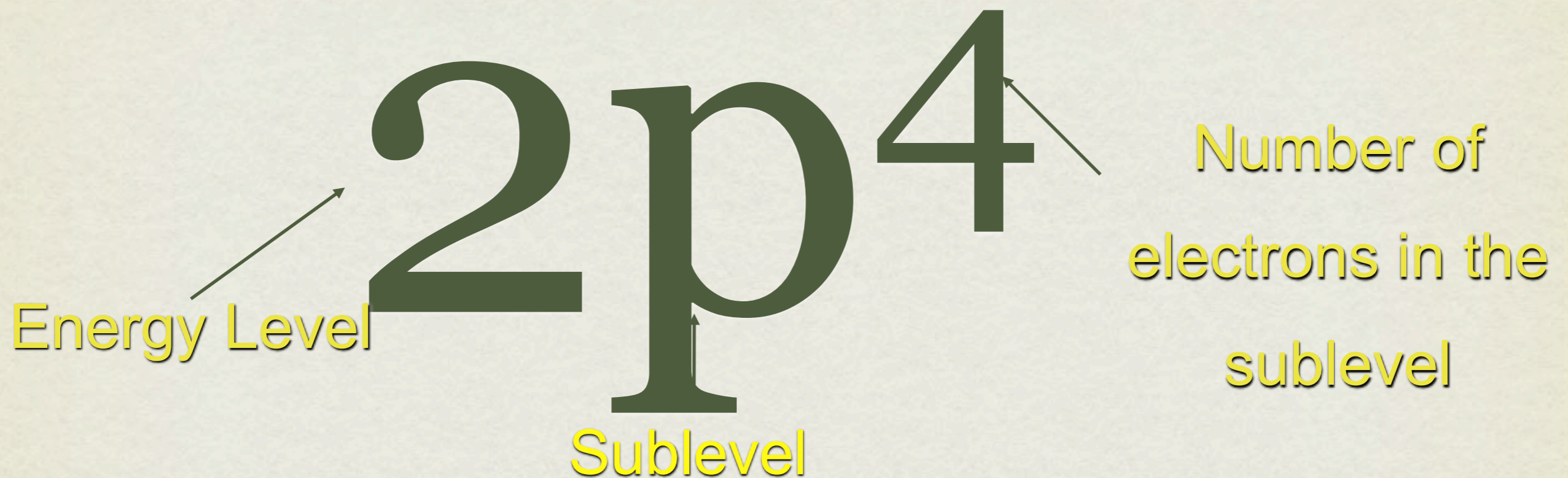
1. Place electrons into the orbitals in order of increasing energy level.
2. Each set of orbitals of the same energy level must be completely filled before proceeding to the next orbital or series of orbitals.
3. Whenever electrons are added to orbitals of the same energy sub-level, each orbital receives one electron before any pairing occurs.
4. When electrons are added singly to separate orbitals of the same energy, the electrons must all have the same spin.

How many electrons can be in a sublevel?

Remember: A maximum of two electrons can be placed in an orbital.

	s orbitals	p orbitals	d orbitals	f orbitals
Number of orbitals	1	3	5	7
Number of electrons	2	6	10	14

ELECTRON CONFIGURATIONS



LET'S TRY IT!

- Write the electron configuration for the following elements:

H

F

Li

Cr

N

Ne

Periodic Table of the Elements

MAIN-GROUP ELEMENTS

MAIN-GROUP ELEMENTS

1	1 1.01 1.01 0.00 0.01 3.00	2 6.94 1.57 800 1500 2744
	3 6.94 0.88 520 437 15.5	4 9.01 1.57 800 1500 2744
2	11 22.99 0.93 405 371 1155	12 24.31 1.31 738 1013 1363
	19 39.10 0.82 49 365.7 1032	20 40.08 1.00 500 1115 1757
3	37 85.47 0.82 403 3125 941.2	38 87.62 0.95 540 1050 1655
	55 132.91 0.79 376 3117 944	56 137.33 0.89 510 1101 2170
4	87 (223) 0.7 375 301.2	88 (226) 0.9 500 93.2
	104 (261) -	105 (262) -

Atomic number	6	12.01	Average atomic mass*
Electronegativity	2.5	4	Common ion charge
First ionization energy (kJ/mol)	1085	1085	Other ion charges
Melting point (K)	476	476	
Boiling point (K)	476	476	

C
carbon

- metals (main group)
- metals (transition)
- metals (inner transition)
- metalloids
- nonmetals
- Gases
- Liquids
- Synthetics

TRANSITION ELEMENTS

3 (IIIB)	4 (IVB)	5 (VB)	6 (VIB)	7 (VIIB)	8 (VIII)	9 (VIII)	10 (VIII)	11 (IB)	12 (IIB)	13 (IIIA)	14 (IVA)	15 (VA)	16 (VIA)	17 (VIIA)	18 (VIIIA)
21 44.96 1.36 309	22 47.87 1.54 350	23 50.94 1.50 368	24 52.00 1.51 368	25 54.94 1.55 368	26 55.85 1.83 368	27 58.93 1.88 368	28 58.69 1.91 368	29 63.55 1.90 368	30 65.39 1.95 368	31 69.72 1.81 368	32 72.61 2.01 368	33 74.92 2.19 368	34 78.96 2.55 368	35 79.90 3.16 368	36 83.80 -
39 88.91 1.32 309	40 91.22 1.33 309	41 92.91 1.5 309	42 95.94 1.56 309	43 (98) -	44 101.07 2.2 309	45 102.91 2.28 309	46 106.42 2.2 309	47 107.87 1.93 309	48 112.41 1.69 309	49 114.82 1.78 309	50 118.71 1.96 309	51 121.76 2.05 309	52 127.60 2.1 309	53 126.90 2.66 309	54 131.29 -
57 138.91 1.10 309	72 178.49 1.3 309	73 180.95 1.5 309	74 183.84 1.7 309	75 186.21 1.9 309	76 190.23 2.2 309	77 192.22 2.2 309	78 195.08 2.2 309	79 196.97 2.4 309	80 200.59 1.9 309	81 204.38 1.8 309	82 207.20 2.15 309	83 208.98 2.0 309	84 (209) -	85 (210) -	86 (222) -
89 (227) -	104 (261) -	105 (262) -	106 (263) -	107 (264) -	108 (269) -	109 (268) -	110 (271) -	111 (272) -	112 (277) -	113 -	114 (285) -	115 -	116 (289) -	117 -	118 -

INNER TRANSITION ELEMENTS

6	Lanthanoids	58 140.12 1.12 527 1071 3716	59 140.91 1.13 528 1071 3716	60 144.24 1.14 528 1071 3716	61 (145) -	62 150.36 1.17 540 1317 2057	63 151.96 1.17 540 1005 1802	64 157.25 1.21 540 1896 3546	65 158.93 1.22 572 1629 3503	66 162.50 1.23 581 1685 2840	67 164.93 1.23 581 1747 2873	68 167.26 1.24 581 1802 341	69 168.93 1.25 581 1888 2223	70 173.04 1.3 603 1002 160	71 174.97 1.3 524 1035 3635
7	Actinoids	90 232.04 1.3 507 3773 901	91 231.04 1.5 508 1045 -	92 238.03 1.7 508 1045 404	93 237.05 1.3 508 117 311	94 (244) -	95 (243) -	96 (247) -	97 (247) -	98 (251) -	99 (252) -	100 (257) -	101 (258) -	102 (259) -	103 (262) -

*Average atomic mass data in brackets indicate atomic masses of most stable isotope of the element.

SHORTHAND NOTATION

- A way of abbreviating long electron configurations
- Since we are only concerned about the outermost electrons, we can skip to places we know are completely full, i.e. the noble gases , and then finish the configuration

SHORTHAND NOTATION

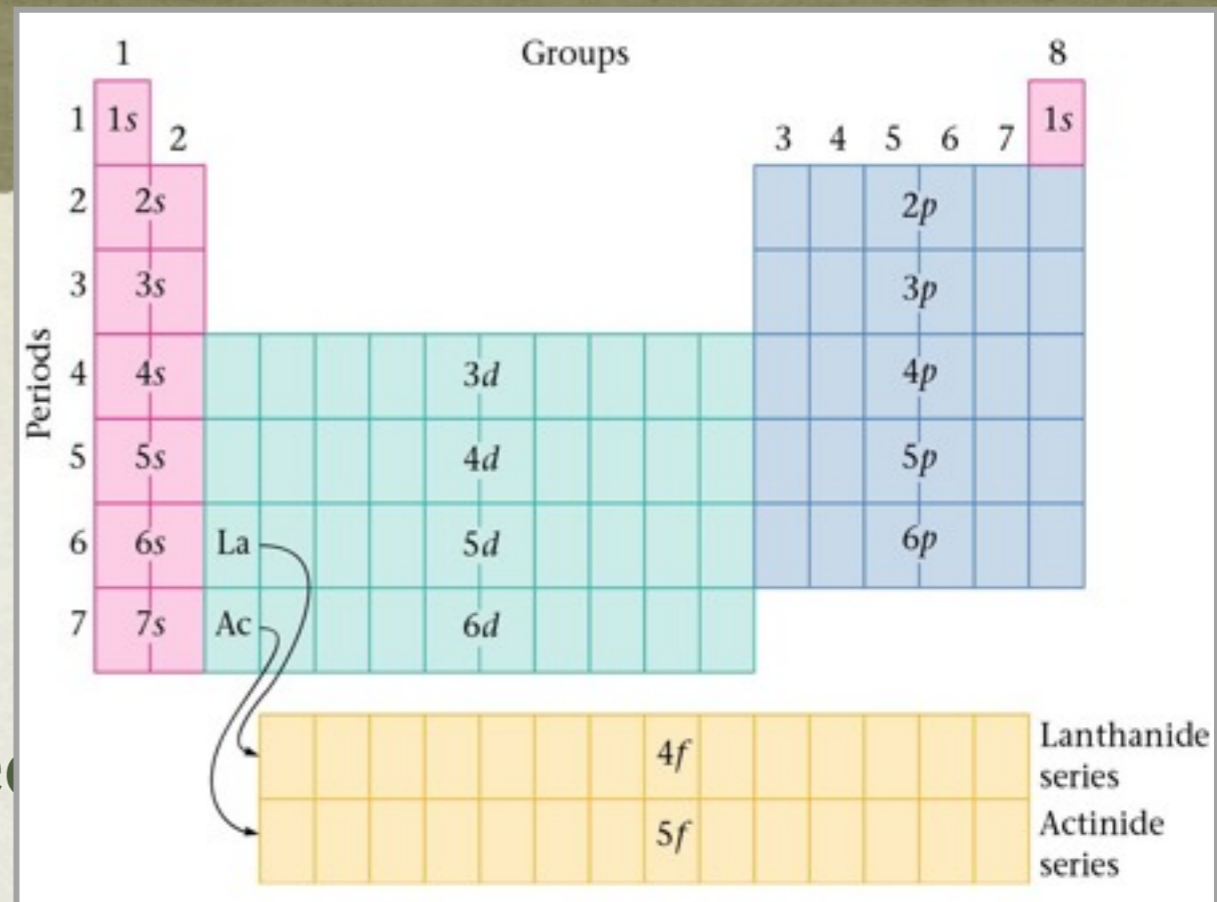
- **Step 1:** Find the closest noble gas to the atom (or ion), **WITHOUT GOING OVER** the number of electrons in the atom (or ion). It will be at the end of the period above the element that you are working with. Write the noble gas in brackets [].
- **Step 2:** Find where to resume by finding the next energy level.
- **Step 3:** Resume the configuration until it's finished.

SHORTHAND NOTATION

- Chlorine

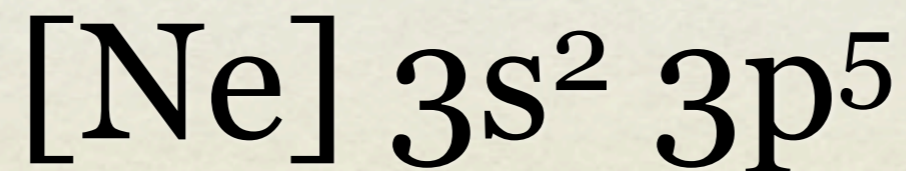
- Longhand is $1s^2 2s^2 2p^6 3s^2 3p^5$

You can abbreviate the first 10 electrons with a noble gas, Ne
[Ne] replaces $1s^2 2s^2 2p^6$



The next energy level after Neon is 3

So you start at level 3 on the diagonal rule (all levels start with s) and finish the configuration by adding 7 more electrons to bring the total to 17



F BLOCK

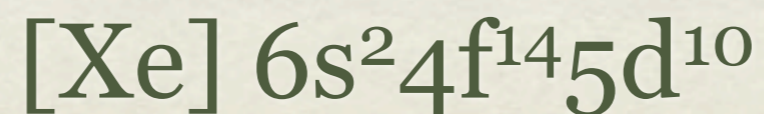
Periodic table diagram showing the f-block elements (Lanthanide and Actinide series) highlighted in yellow. The diagram includes labels for Periods (1-7) and Groups (1-8). The f-block is labeled as the Lanthanide series (4f) and Actinide series (5f). A red star is placed in the 6th period, 8th group area.

Periods	1	2	3	4	5	6	7	8
1	1s							1s
2	2s							
3	3s							
4	4s			3d				
5	5s			4d				
6	6s	La		5d				
7	7s	Ac		6d				

Lanthanide series
Actinide series

Any element past #57 will have f block electrons

Let's try Hg (#80)



PRACTICE SHORTHAND NOTATION

- Write the shorthand notation for the following atoms:

S

K

Ca

I

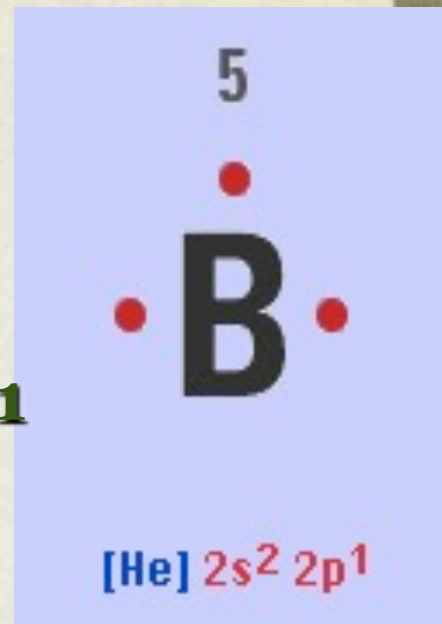
Bi

VALENCE ELECTRONS

Electrons are divided between **core** & valence electrons

Write the electron configuration of Boron and determine how many valence electrons Boron has

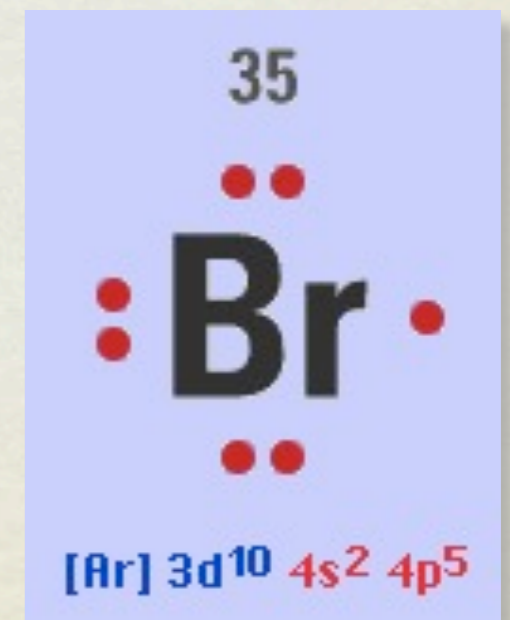
B $1s^2 2s^2 2p^1$ Core = [He] Valence = $2s^2 2p^1$



Write the electron configuration of Bromine and determine how many valence electrons Br has

Br [Ar] $4s^2 3d^{10} 4p^5$

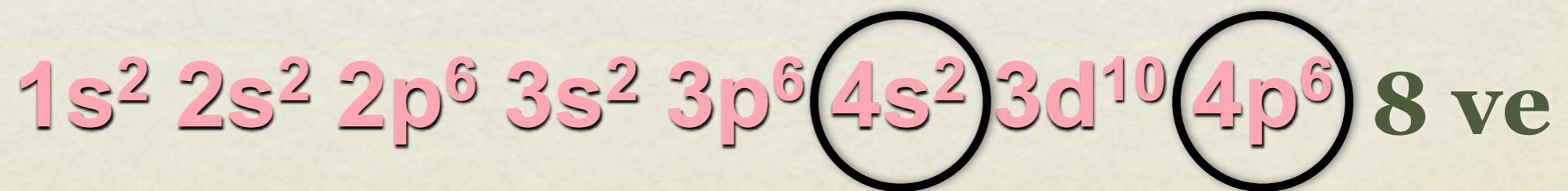
Core = [Ar] $3d^{10}$, valence = $4s^2 4p^5$



VALENCE ELECTRONS

- Valence electrons are always found in the **OUTERMOST** shell.
- It doesn't matter in the order they are written in!
- Always the shell with the highest principal number

How many valence electrons?



KEEP AN EYE ON THOSE IONS!

- Electrons can be lost or gained by atoms to form ions
- negative ions have gained electrons, positive ions have lost electrons
- The electrons that are lost or gained should be added/removed from the **highest energy level** (not the highest orbital in energy!)

FORMING IONS!

- Write electron configurations: Sn, Sn²⁺, Sn⁴⁺

Atom: [Kr] 5s² 4d¹⁰ 5p²

Sn²⁺ ion: [Kr] 5s² 4d¹⁰

Sn⁴⁺ ion: [Kr] 4d¹⁰

Note that the electrons came out of the highest energy level, not the highest energy orbital!

FORMING IONS

- Bromine

Atom: $[\text{Ar}] 4s^2 3d^{10} 4p^5$

Br⁻ ion: $[\text{Ar}] 4s^2 3d^{10} 4p^6$

Note that the electrons went into the highest energy level, not the highest energy orbital!

TRY SOME IONS!

- Write the longhand notation for these:



Write the shorthand notation for these:



HOMework

- Complete all of p.40 in your workbook

EXCEPTIONS TO THE AUFBAU PRINCIPLE

- Remember d and f orbitals require **LARGE** amounts of energy
- If we can't fill these sublevels, then the next best thing is to be **HALF** full (**one electron in each orbital in the sublevel**)
- There are many exceptions, but the most common ones are

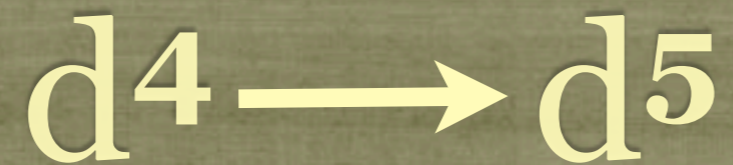
d^4 and d^9



Is the glass half full or half empty ?

Technically, the glass is almost empty.





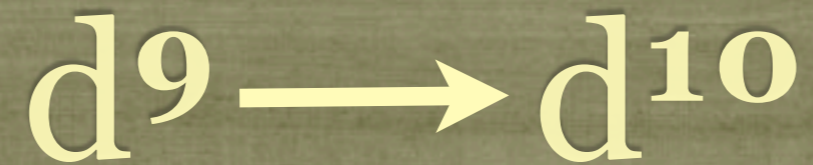
- d orbitals can hold up to 10 electrons so ...d⁴ is one electron short of being HALF full (d⁵)
- In order to become more stable (require less energy), one of the **closest** s electrons will actually go into the d, making it d⁵ instead of d⁴.

Write electron configuration of Cr



Procedure: Find the closest s orbital. Steal one electron from it, and add it to the d.

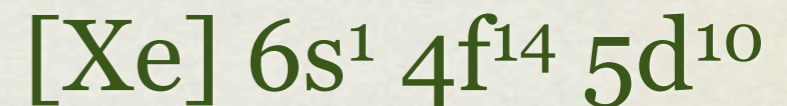




- d^9 is one electron short of being full
- Just like d^4 , one of the **closest s** electrons will go into the d, this time making it d^{10} instead of d^9 .

Write electron configuration of Au $[\text{Xe}] 6s^2 4f^{14} 5d^9$

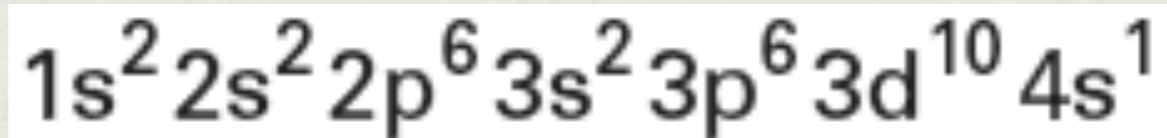
Procedure: Find the closest s orbital. Steal one electron from it, and add it to the d.



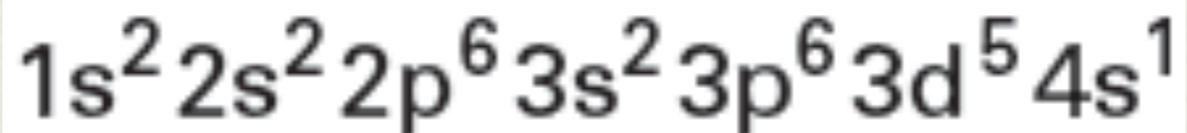
TRY THESE!

- Write the longhand notation for:

Cu



Cr

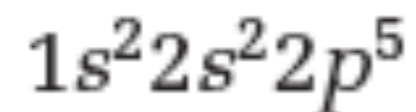
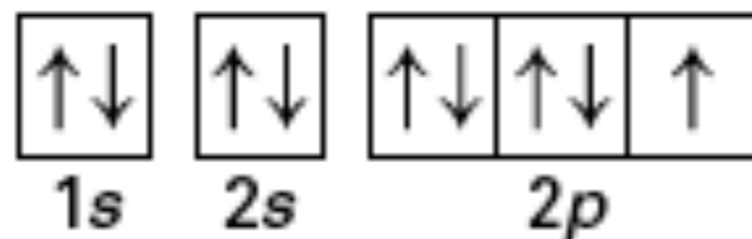


Mo



ORBITAL DIAGRAMS

- Graphical representation of an electron configuration
- One arrow represents one electron
- Shows spin and which orbital within a sublevel
- Same rules as before (Aufbau principle, d^4 and d^9 exceptions, two electrons in each orbital, etc. etc.)

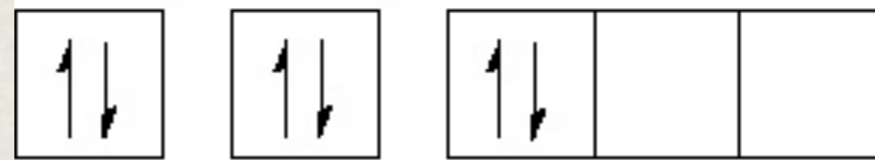


HUND'S RULE

- In orbitals of EQUAL ENERGY (p, d, and f), place one electron in each orbital before making any pairs
- All single electrons must spin the same way
- Think of this rule as the “Monopoly Rule”
- In Monopoly, you have to build houses EVENLY. You can not put 2 houses on a property until all the properties has at least 1 house.



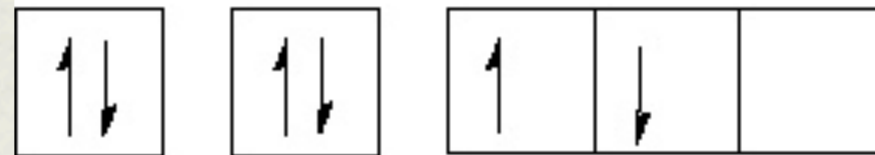
CARBON



1s

2s

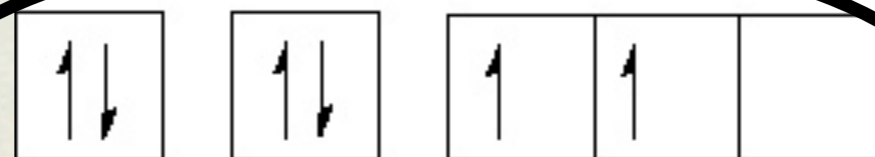
2p



1s

2s

2p



1s

2s

2p

Three possibilities,
which is correct?

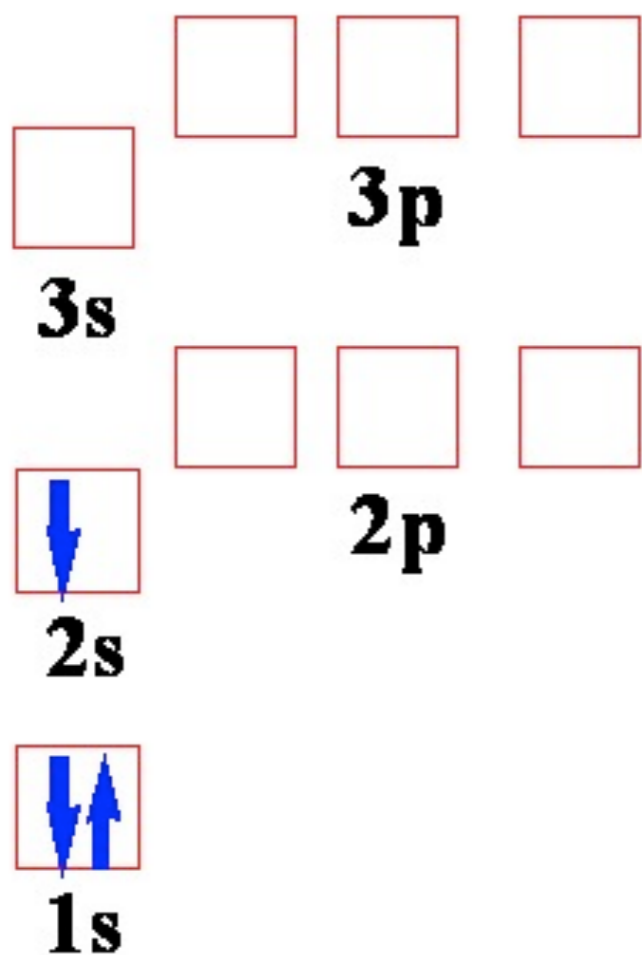
Need to use
Hund's Rule

LITHIUM

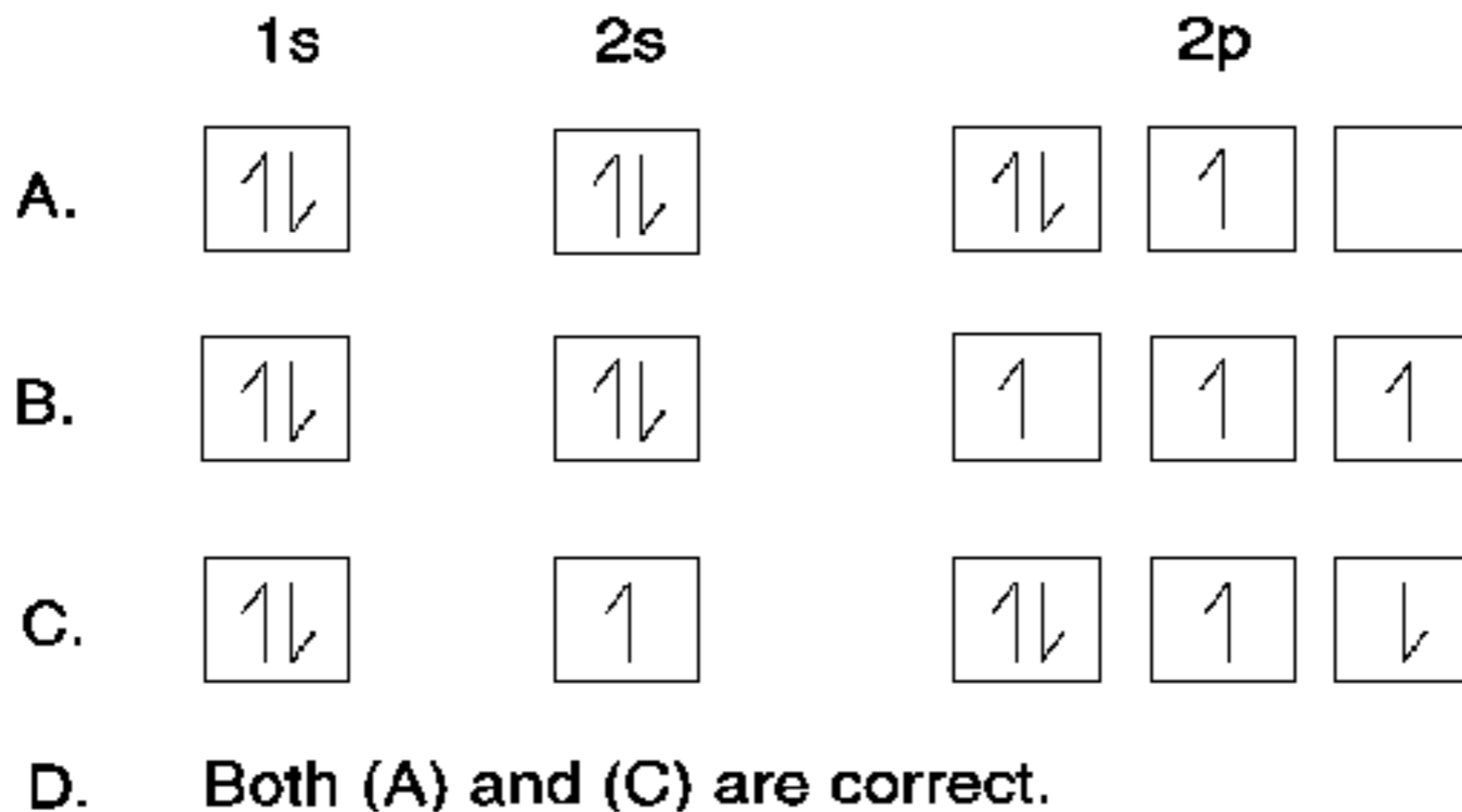


Electron configuration:

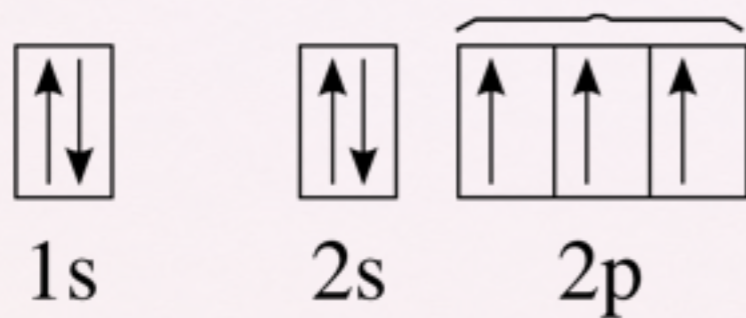
$1s^2 2s^1$ ---> 3 total electrons



NITROGEN



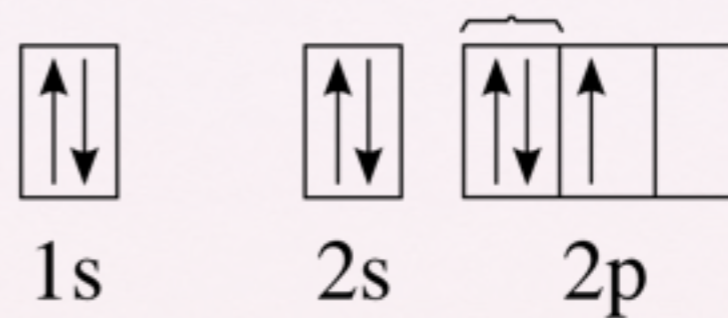
no electron-electron repulsion
equals lower energy



correct

or

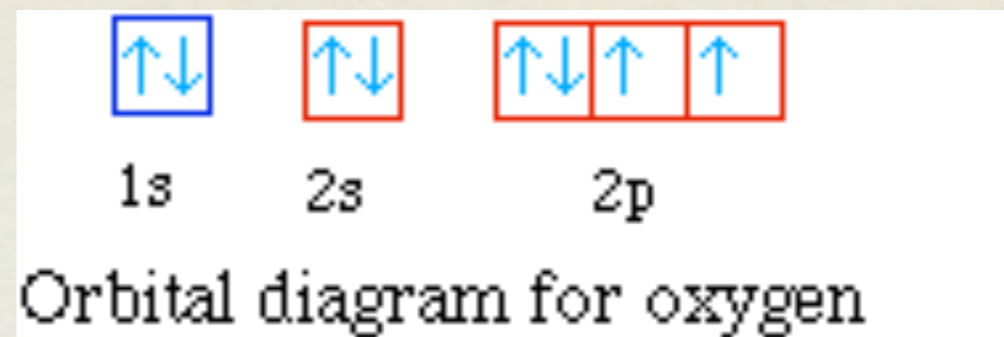
electron-electron repulsion
equals higher energy



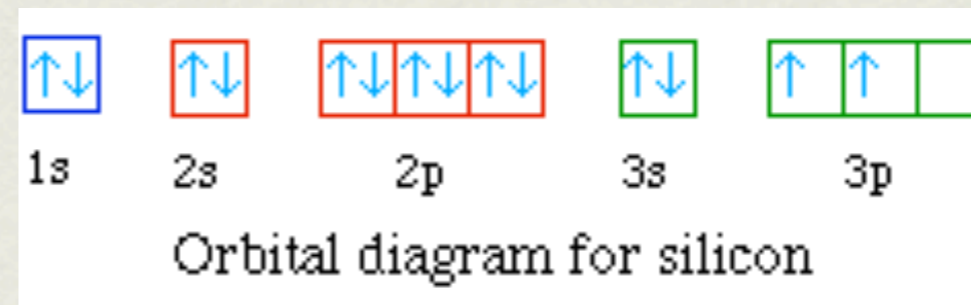
incorrect

DRAW THESE ORBITAL DIAGRAMS!

- Oxygen (O)

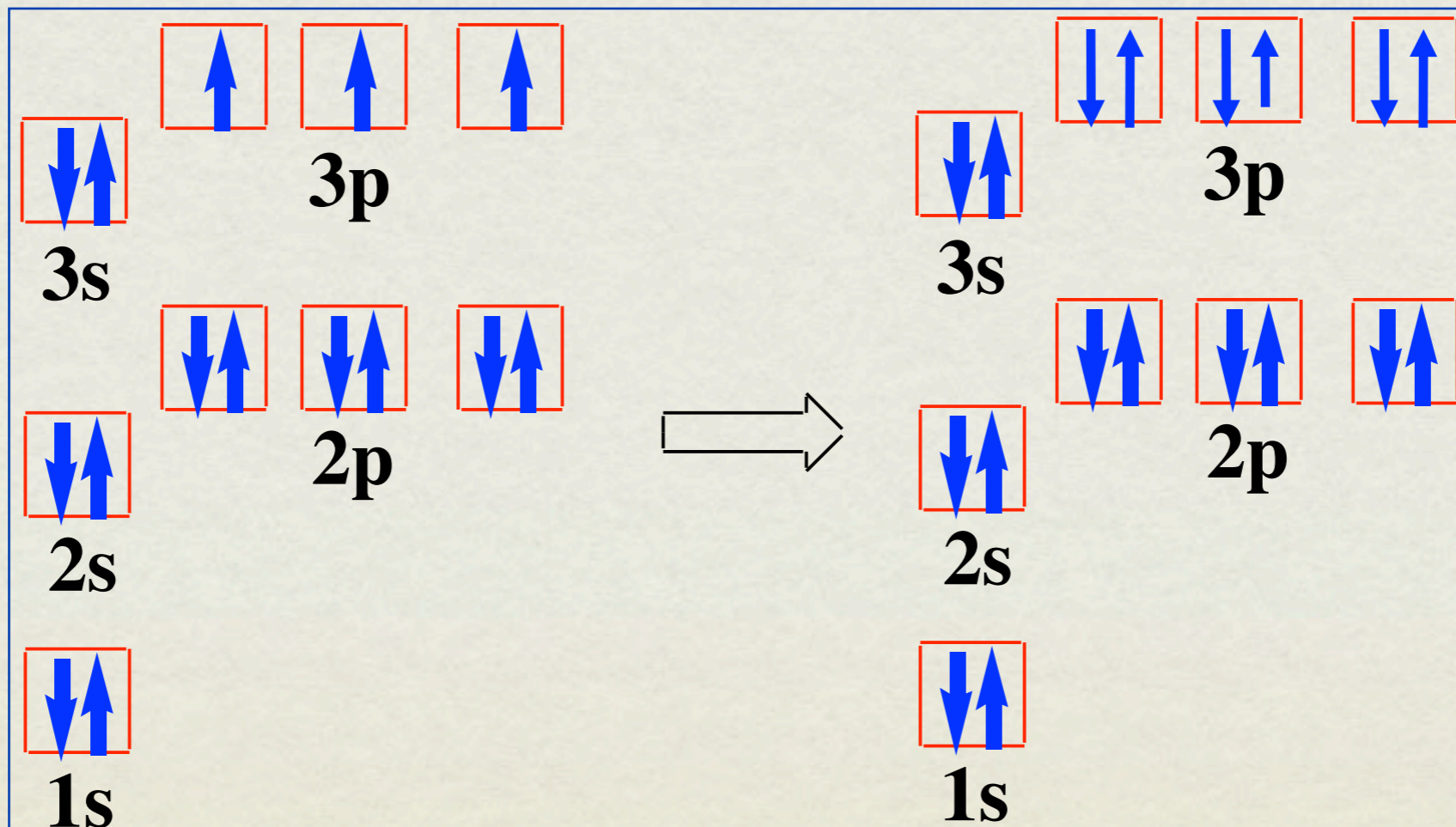


- Silicon (Si)



IONS!

To form anions from elements, add 1 or more e⁻ from the highest sublevel.



TRY IT!

Element	Total number of electrons	Orbital diagram				Full Electron configuration
		1s	2s	2p	3s	
Helium	2					
Lithium	3					
Boron	5					
Nitrogen	7					
Fluorine	9					
Sodium	11				s ¹	

TRY IT!



P.191 #3,4

P.194 # 6,8,9,10