Step by Step: Electron Configurations and Electron Orbital Diagrams

Electron Configurations

Ex. 1) Mg: $1s^2\ 2s^2\ 2p^6\ 3s^2$

↑↑↑
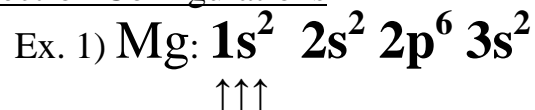
$1 = 1^{st}$ layer (row #), s = orbital type , power of 2 = the 2 electrons in the 1s orbital

**Move the Helium box next to Hydrogen (above Beryllium.) See the periodic table below.

Looking at the periodic table, we need to count each box going from Hydrogen (#1) to Magnesium (#12), including Magnesium.

H and He boxes: These boxes are in the 1st row, so **1**. These boxes are in the "s" orbital region (see below), so **S**. There are two boxes, so two electrons; the power is **2**. $1s^2$

Li and Be: These boxes are in the 2nd row, so **2**. These boxes are in the "s" orbital region (see below), so **S**. There are two boxes, so two electrons; the power is **2**. $2s^2$

B to Ne: These boxes are in the 2nd row, so **2**. These boxes are in the "p" orbital region (see below), so **p**. There are six boxes, so six electrons; the power is **6**. $2p^6$

Na to Mg: These boxes are in the 3rd row, so **3**. These boxes are in the "s" orbital region (see below), so **S**. There are two boxes, so two electrons; the power is **2**. $3s^2$
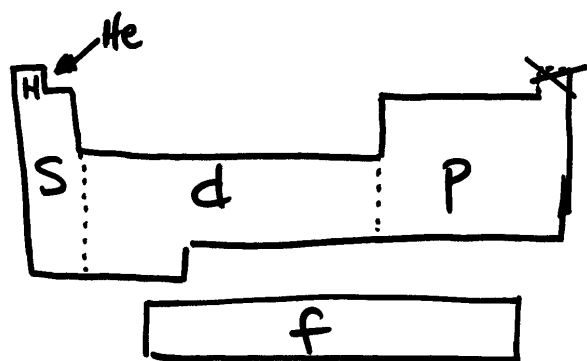
s-orbitals: column #1-2
d-orbitals: column #3-12
p-orbitals: column #13-18
f-orbitals: elements #57-71 and #89-103

2) Ba: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ 6s^2$

Looking at the periodic table, we need to count each box going from Hydrogen (#1) to Barium (#56), including Barium.

<u>H and He boxes:</u> These boxes are in the 1st row, so **1**. These boxes are in the "s" orbital region, so **S**. There are two boxes, so two electrons; the power is **2**. $\mathbf{1s^2}$

<u>Li and Be:</u> These boxes are in the 2nd row, so **2**. These boxes are in the "s" orbital region, so **S**. There are two boxes, so two electrons; the power is **2**. $\mathbf{2s^2}$

<u>B to Ne:</u> These boxes are in the 2nd row, so **2**. These boxes are in the "p" orbital region, so **p**. There are six boxes, so six electrons; the power is **6**. $\mathbf{2p^6}$

<u>Na to Mg:</u> These boxes are in the 3rd row, so **3**. These boxes are in the "s" orbital region, so **S**. There are two boxes, so two electrons; the power is **2**. $\mathbf{3s^2}$

<u>Al to Ar:</u> These boxes are in the 3rd row, so **3**. These boxes are in the "p" orbital region, so **p**. There are six boxes, so six electrons; the power is **6**. $\mathbf{3p^6}$

<u>K and Ca:</u> These boxes are in the 4th row, so **4**. These boxes are in the "s" orbital region, so **S**. There are two boxes, so two electrons; the power is **2**. $\mathbf{4s^2}$

<u>Sc to Zn:</u> <u>These boxes are in the 4th row, but the "d" orbitals are one row behind, so **3**</u>. (This is because these orbitals are closer in energy to the 3rd layer than to the 4th layer.) These boxes are in the "d" orbital region, so **d**. There are ten boxes, so ten electrons; the power is **10**. $\mathbf{3d^{10}}$

<u>Ga to Kr:</u> These boxes are in the 4th row, so **4**. These boxes are in the "p" orbital region, so **p**. There are six boxes, so six electrons; the power is **6**. $\mathbf{4p^6}$

<u>Rb and Sr:</u> These boxes are in the 5th row, so **5**. These boxes are in the "s" orbital region, so **S**. There are two boxes, so two electrons; the power is **2**. $\mathbf{5s^2}$

<u>Y to Cd</u>: These boxes are in the 5th row, but the "d" orbitals are one row behind, so **4**. (This is because these orbitals are closer in energy to the 4th layer than to the 5th layer.) These boxes are in the "d" orbital region, so **d**. There are ten boxes, so ten electrons; the power is **10**. **4d^{10}**

<u>In to Xe</u>: These boxes are in the 5th row, so **5**. These boxes are in the "p" orbital region, so **p**. There are six boxes, so six electrons; the power is **6**. **5p^{6}**

<u>Cs and Ba:</u> These boxes are in the 6th row, so **6**. These boxes are in the "s" orbital region, so **s**. There are two boxes, so two electrons; the power is **2**. **6s^{2}**

**When we do the "f" orbital, we will find they will be two rows/layers behind. So the "f" orbitals in the 6th layer will actually be 4f!
Note the "4f" orbitals in Nd below.

↓

Ex. 1a) #60 Nd: **[Xe] 6s^2 5d^1 then 4f^3**

For the short form electron configurations, we **find the noble gas before the element** and start the orbitals from there.
Count back from #60 to #58 Ce, then it jumps up to #57 La, then count back to #55 Cs, then back to the next element which is #54 Xe (the **closest noble gas that is before #60 Nd**) Xe goes into brackets: **[Xe]**
Then we start naming the boxes, like in the previous example:
Cs and Ba: are **6s^2**
#57 La is **5d^1** (remember the "d" orbitals are one layer less)
Then we jump down to Ce to Nd: **4f^3** (remember the "f" orbitals are two layers less)

**But the one electron in the "d" orbital will never stay alone. It will go down into the "f" orbital, if there is room in the "f" orbital and it is not full. If the "f" orbital is full, then it will stay in the "d" orbital. If it stays in the "d" orbital, then it can usually combine with other electrons in the "d" orbital.
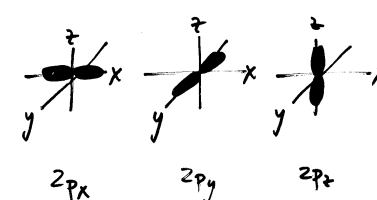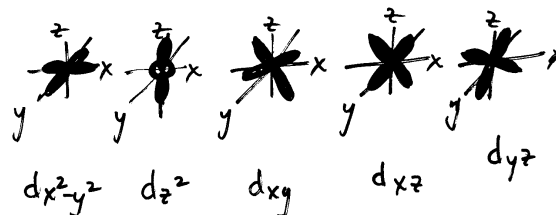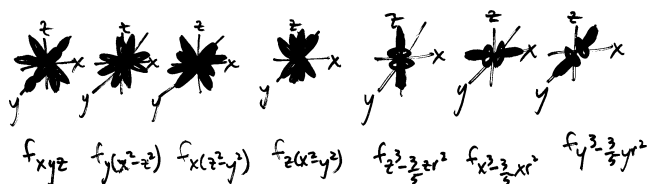
[Xe] $6s^2$ $4f^4$ $5d^0$ This is the final answer!

The 1e- alone in the "d" is lonely and moves to the "f**" with its **f**riends.

1b) #63 Eu: [Xe] $6s^2$ $5d^1$ $4f^6$ → **[Xe] $6s^2$ $4f^7$ $5d^0$**
Note the "d^1" electron moves to the "f".

1c) #77 Ir: [Xe] $6s^2$ $5d^1$ $4f^{14}$ $5d^6$ → **[Xe] $6s^2$ $4f^{14}$ $5d^7$**
Note the "d^1" electron moves to be with the other "d" electrons.
(The order of the d's and f's can be reversed. [Xe] $6s^2$ $5d^7$ $4f^{14}$ is acceptable.)

For the orbital diagrams: We will use the short form, but show the actual orbitals, where **each orbital picture is represented by a horizontal line.**

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

Bi short form = [Xe] $6s^2 4f^{14} 5d^{10} 6p^3$

f_{xyz} $f_{y(x^2-z^2)}$ $f_{x(z^2-y^2)}$ $f_{z(x^2-y^2)}$ $f_{z^3-\frac{3}{5}zr^2}$ $f_{x^3-\frac{3}{5}xr^2}$ $f_{y^3-\frac{3}{5}yr^2}$

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

$2p_x$ $2p_y$ $2p_z$

↑

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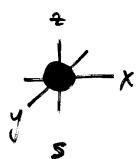
"s" orbitals have one picture, so one horizontal line.

"f" orbitals have 7 pictures, so 7 horizontal lines.

"d" orbitals have 5 pictures, so 5 horizontal lines.

"p" orbitals have 3 pictures, so 3 horizontal lines.

[Xe] __   __ __ __ __ __ __ __   __ __ __ __ __   __ __ __

6s   4f   5d   6p

Then electrons (arrows) are added one at a time into each type of orbital. **Remember the super script is the number of electrons, so $6s^2$ (2 electrons (arrows) in the s), $4f^{14}$ (14 electrons (arrows) in the f), $5d^{10}$ (10 electrons (arrows) in the d), $6p^3$ (3 electrons (arrows) in the p). These electrons (arrows) will all go up initially, one electron per horizontal line of each orbital type. When there is one in each orbital of that type, then the second electron on each horizontal line will go down. First add to the "s" orbital. When it is full, then add to the "f"orbital. When that is full, then add to the "d" orbital etc.
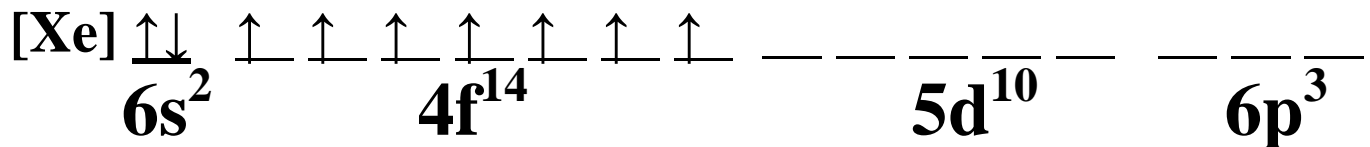
[Xe] ↑ _ _ _ _ _ _ _ _ _ _ _ _ _ _ _

$6s^2$ $4f^{14}$ $5d^{10}$ $6p^3$
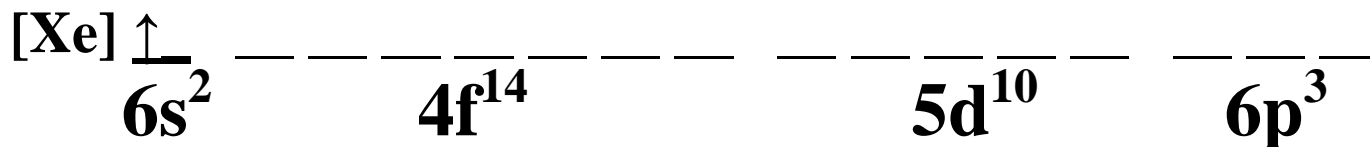
The next electron will fill the "s" and the spin direction will be opposite of the first electron.

[Xe] ↑↓ _ _ _ _ _ _ _ _ _ _ _ _ _ _ _

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

[Xe] ↑↓ ↑ _ _ _ _ _ _ _ _ _ _ _ _ _

$6s^2$ $4f^{14}$ $5d^{10}$ $6p^3$
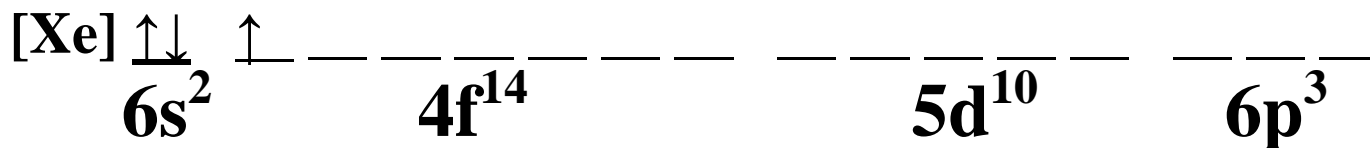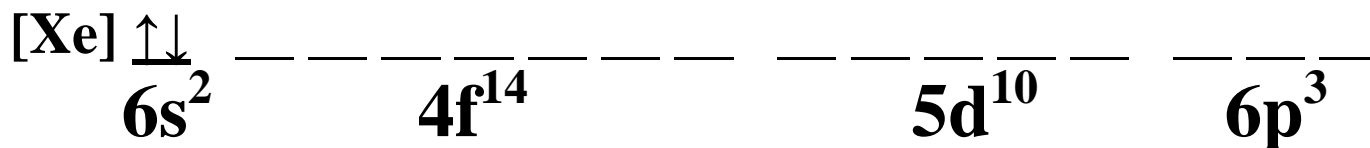
Then the next electrons go in singly and all in the up spin direction.

[Xe] ↑↓ ↑ ↑ _ _ _ _ _ _ _ _ _ _ _ _

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

[Xe] ↑↓ ↑ ↑ ↑ _ _ _ _ _ _ _ _ _ _ _

$6s^2$ $4f^{14}$ $5d^{10}$ $6p^3$
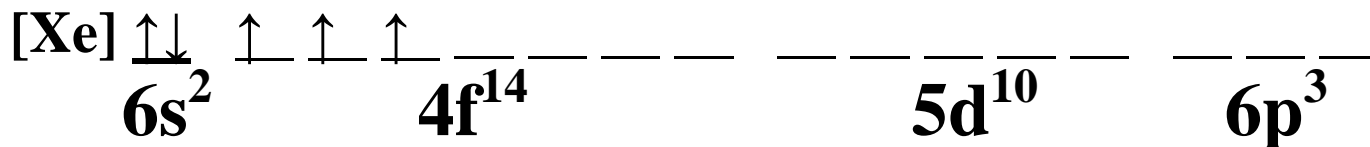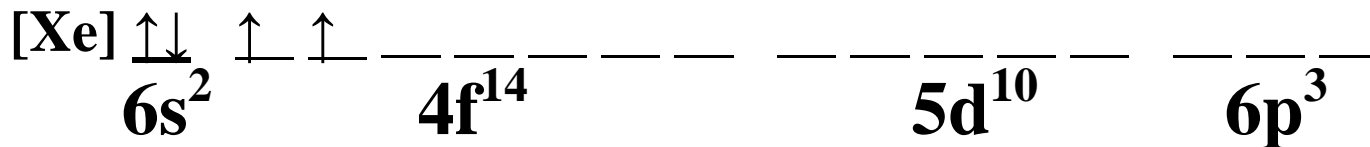
Etc. each electron will go in singly until each horizontal line has one electron.

[Xe] ↑↓ ↑ ↑ ↑ ↑ ↑ ↑ ↑ _ _ _ _ _ _ _ _

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

When there is one electron in each of the "f" orbitals, then they will start pairing with the opposite spin (down).

[Xe] ↑↓   ↑↓ ↑ ↑ ↑ ↑ ↑ ↑   — — — — — — —
$6s^2$ $4f^{14}$ $5d^{10}$ $6p^3$

[Xe] ↑↓   ↑↓ ↑↓ ↑ ↑ ↑ ↑ ↑   — — — — — — —
$6s^2$ $4f^{14}$ $5d^{10}$ $6p^3$
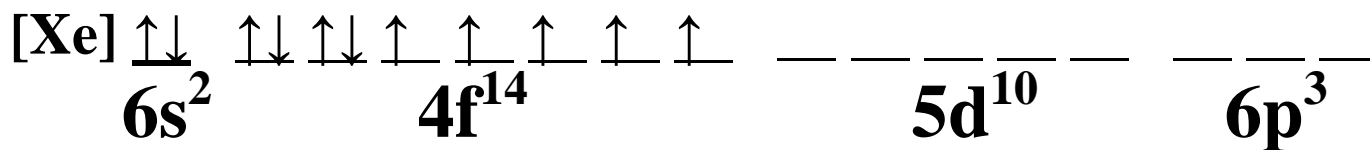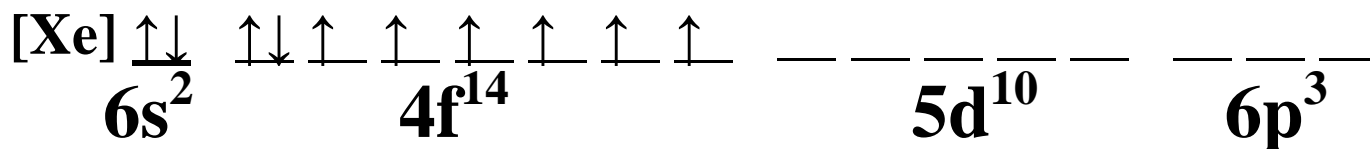
Finish filling the f's. Then keep filling for the d's and the p's. Same method: One electron per horizontal line going up, until each line has one electron. Then fill the second electron on each horizontal line down.

[Xe] ↑↓   ↑↓ ↑↓ ↑↓ ↑↓ ↑↓ ↑↓ ↑↓   ↑↓ ↑↓ ↑↓ ↑↓ ↑↓ ↑ ↑ ↑
$6s^2$ $4f^{14}$ $5d^{10}$ $6p^3$
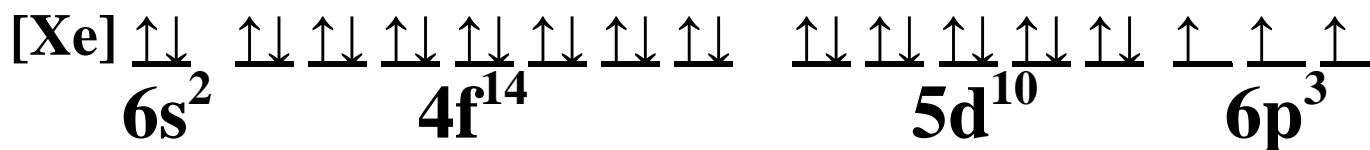
**Notice there are the 2 electrons (arrows) in the s-orbital $6s^2$, the 14 electrons (arrows) in the f-orbital $4f^{14}$, the 10 electrons (arrows) in the d-orbital $5d^{10}$, and the 3 electrons (arrows) in the p-orbital $6p^3$.

<u>*End of Notes*</u>