CHEM 3.4 (AS91390)

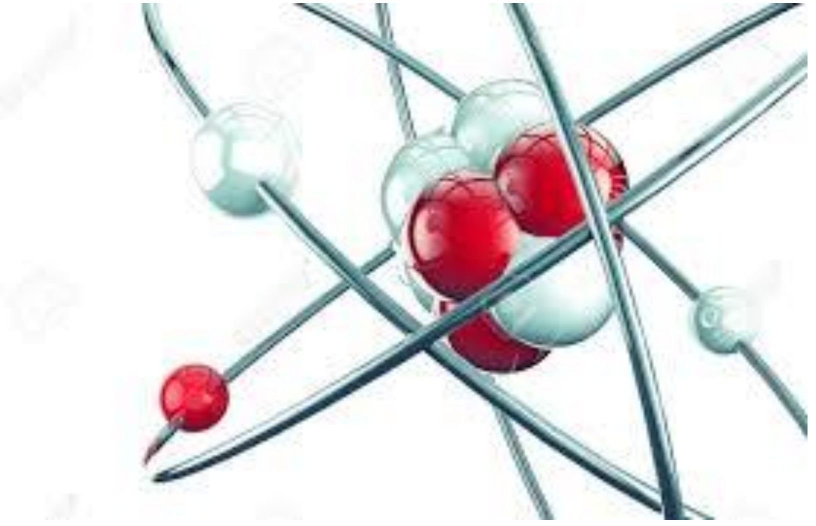
5 credits

Demonstrate understanding of thermochemical principles and the properties of particles and substances

Do now:

Describe in your own words the structure of an atom.

Describe what information you can get from the periodic table about the structure of an atom.



Electron configuration

Atoms have the same number of protons and electrons. We use the atomic number to tell us how many protons an atom has.

Up until now we have written the electron configuration of atoms (and ions) in terms of the number of electrons in each shell. For example:

Write the electron configuration of

2nd energy level

1st energy level

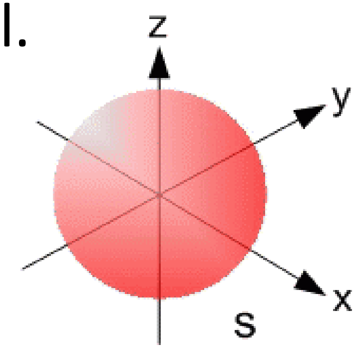
3rd energy level

Mg	**2, 8, 2**	S	**2, 8, 6**	Al	**2, 8, 3**
Na^+	**2, 8**	F^-	**2, 8**	Ca^{2+}	**2, 8, 8**

Electron configuration

In level 3 we learn to write electron configurations differently, in terms of sub orbitals.

Each energy level has sub-orbitals. Only 2 electrons can fit in each sub-orbital.
The 1st energy level can have up to 2 electrons so it has only 1 sub-orbital – we call this an s orbital.



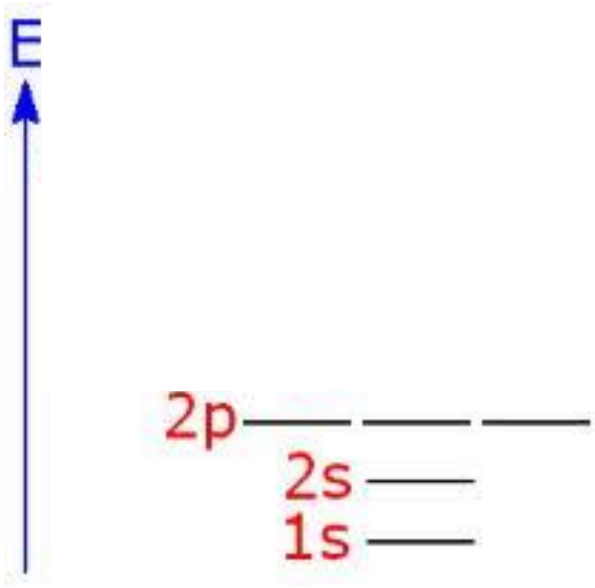
H has only 1 electron. It goes into the 1st energy level in the s orbital
We would write the electron configuration as $1s^1$

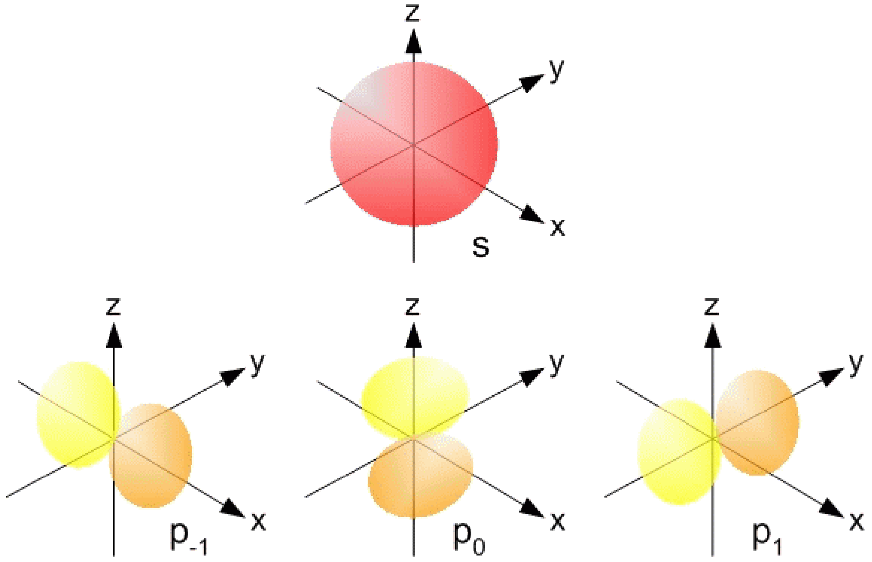
energy level orbital number of electrons

Electron configuration

The 2nd energy level can have up to 8 electrons so it has 4 sub-orbitals – 1 s orbital and 3 p orbitals.

The s orbital is lower in energy so electrons fill this orbital first. The p orbitals are called **degenerate orbitals** and they are all the same energy.

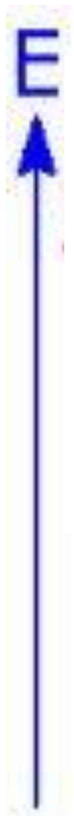




Electron configuration

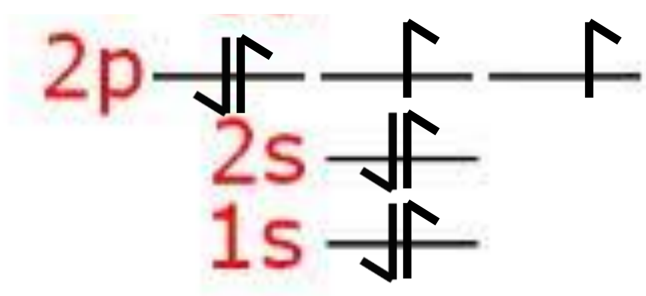
For example:

O has 8 electrons: 2, 6



Orbitals fill up so that one electron goes in each orbital first, then they pair up

We now write the electron configuration of O as: $1s^2 2s^2 2p^4$



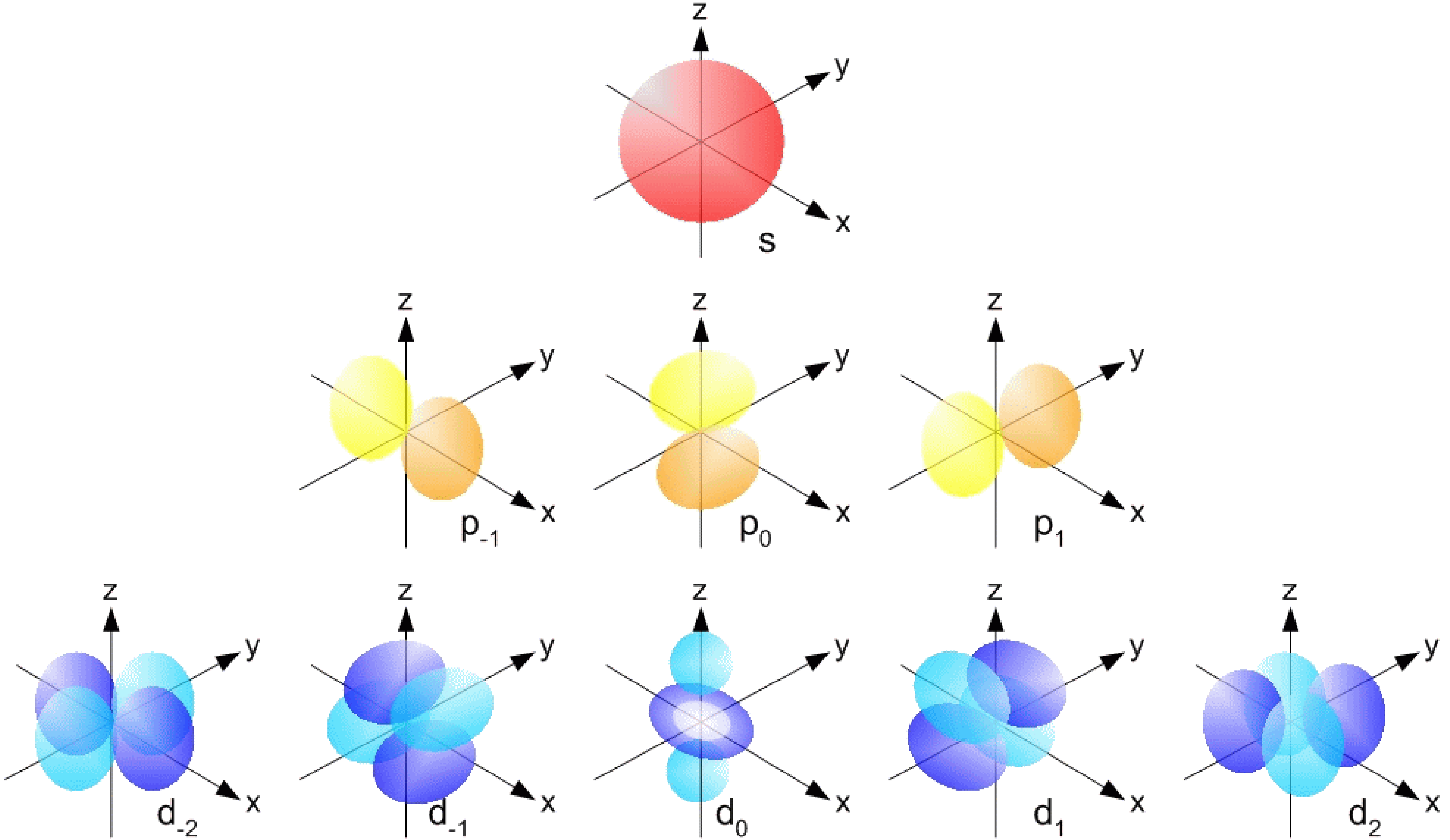
Write the electron configuration of

B **$1s^2 2s^2 2p^1$** F **$1s^2 2s^2 2p^5$**

Li **$1s^2 2s^1$**

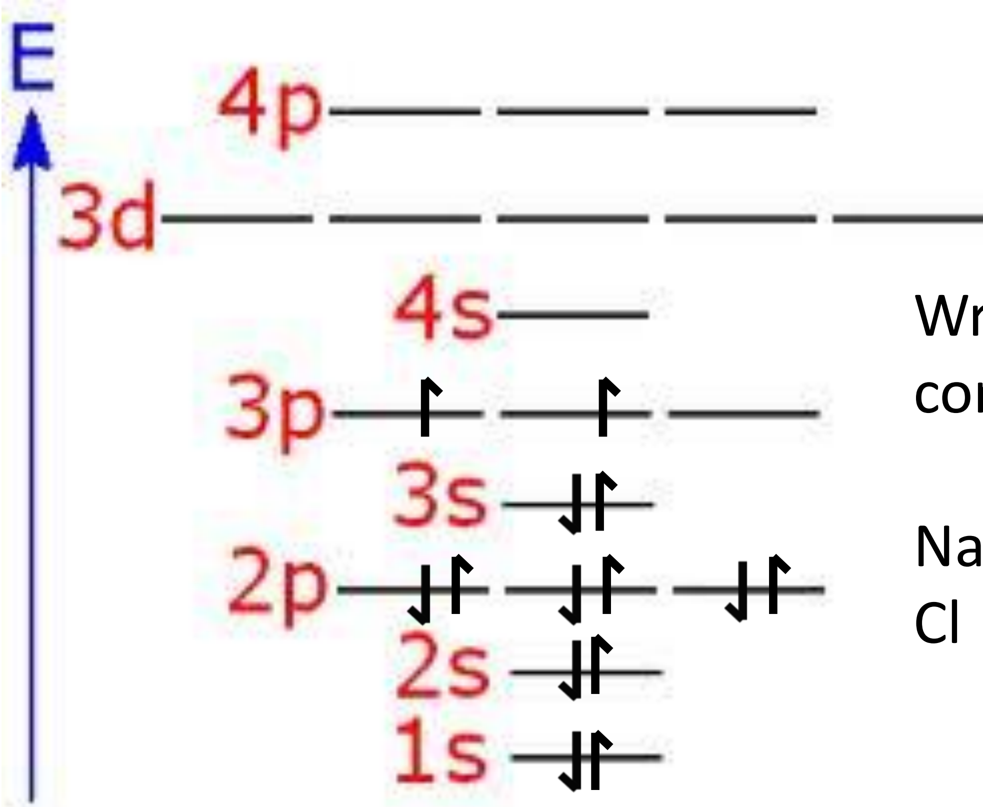
Electron configuration

The 3rd energy level can have up to 18 electrons so it has 9 sub-orbitals – 1 s orbital and 3 p orbitals and 5 d orbitals.



Electron configuration

For example:
Si has 14 electrons



We write this as
$1s^2 2s^2 2p^6 3s^2 3p^2$

Write the electron configuration of

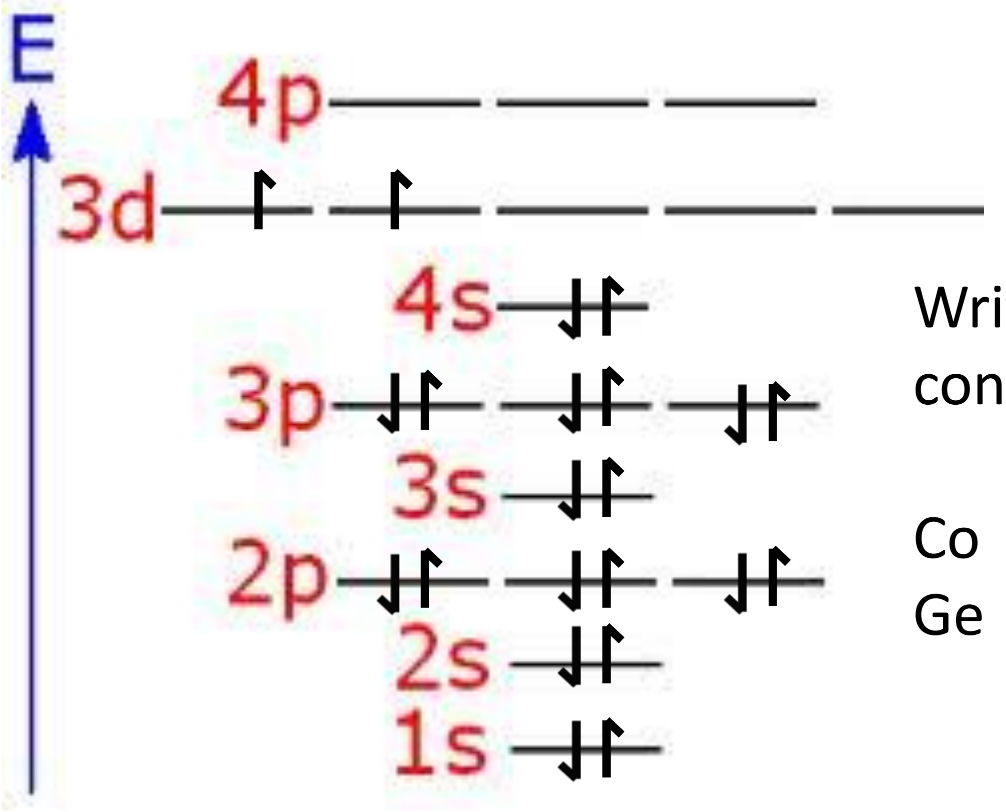
Na **$1s^2 2s^2 2p^6 3s^1$**
Cl **$1s^2 2s^2 2p^6 3s^2 3p^5$**

Note that the 3d orbital is higher in energy than the 4s orbital. This means that it gets filled after the 4s orbital.

Electron configuration

For example:

Ti has 22 electrons



We write this as

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$

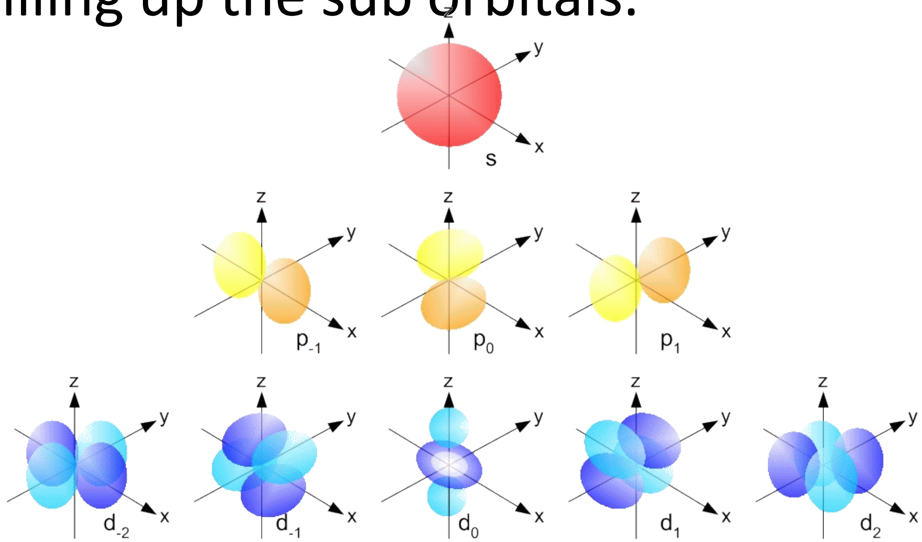
Write the electron configuration of

Co $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^7$

Ge $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^7 4p^2$

Do now:

Write down what you remember about s, p, d electron configuration and the order for filling up the sub orbitals.



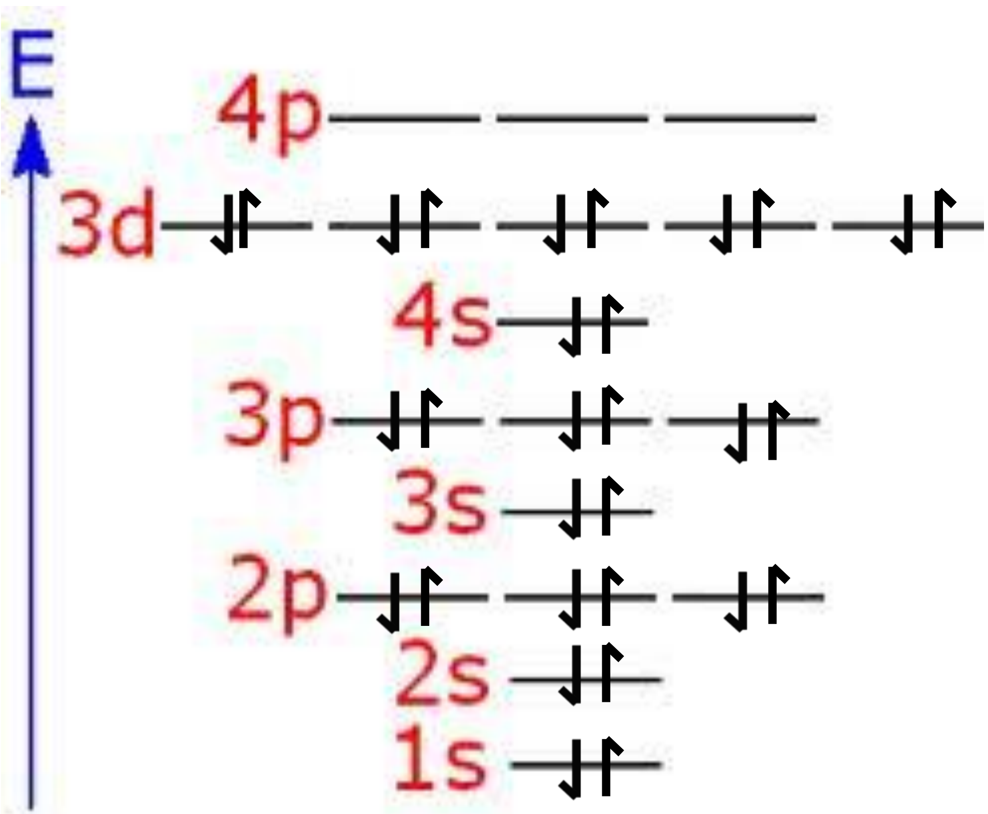
Write down the electron configuration of the following atoms using s, p, d notation

Nitrogen	Magnesium	Vanadium
$1s^2 2s^2 2p^3$	**$1s^2 2s^2 2p^6 3s^2$**	**$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$**

Electron configuration

Exceptions:

Half filled orbital levels and complete orbital levels are more favourable than partially filled orbital levels. This applies to **chromium** and **copper**.



Chromium – 24 electrons

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$

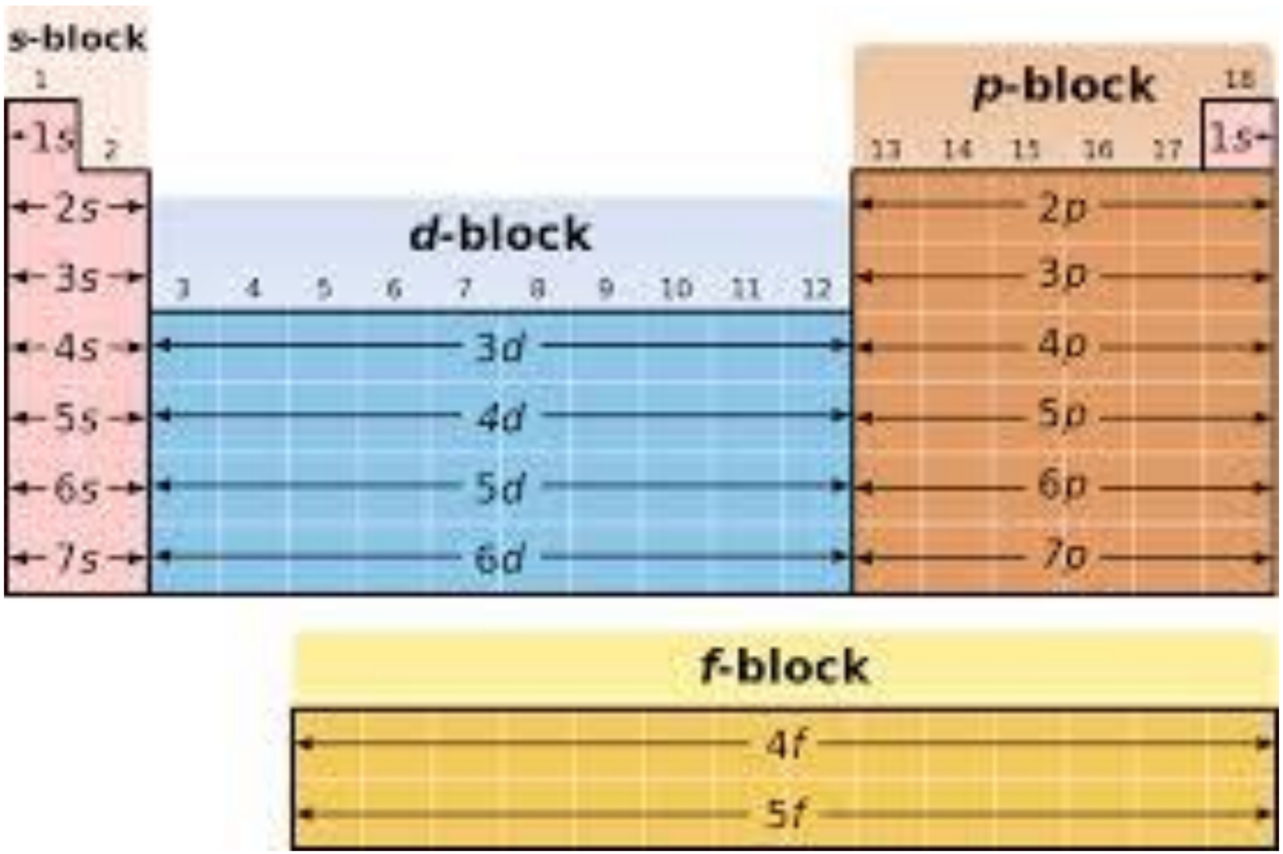
Copper – 29 electrons

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

Exercises in book pg 9 – left hand column (1a, c, e, g, i, k, m)

Electron configuration

Look at the electron configuration of the atoms you have written and look at their position on the periodic table. What do you notice?



Electron configuration

Writing out electron configuration for elements can get tedious!
We use a shorthand to make this process easier.

Write the electron configuration of:

Co **$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^7$** Ar = $1s^2 2s^2 2p^6 3s^2 3p^6$ Co: [Ar] $4s^2 3d^7$

Na **$1s^2 2s^2 2p^6 3s^1$** Ne: $1s^2 2s^2 2p^6$ Na: [Ne] $3s^1$

We can use the previous group 18 element to represent most of the electron configuration.

Ar: $1s^2 2s^2 2p^6 3s^2 3p^6$
Ne: $1s^2 2s^2 2p^6$

Electron configuration of ions

Atoms gain or lose electron to form ions. We will mostly focus on losing electrons to form cations.

Write the electron configuration for these ions:

Cl^-	**$1s^22s^22p^63s^23p^6$**	Mg^{2+}	**$1s^22s^22p^6$**
O^{2-}	**$1s^22s^22p^6$**	K^+	**$1s^22s^22p^63s^23p^6$**

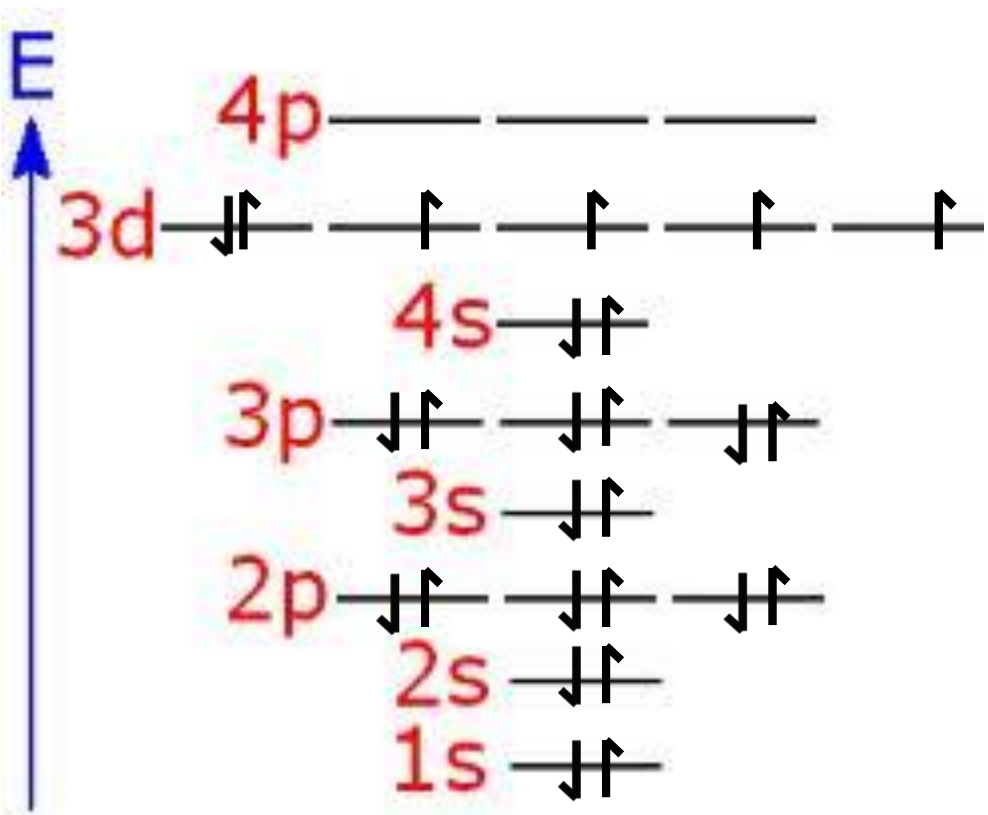
When electrons are lost to form ions they are lost from the outer most shell. This means the 4s electrons are usually lost before the 3d electrons.

Using s, p, d notation for electron configurations helps to explain why transition metals have various oxidation states.

Electron configuration of ions

For example:

Fe^{2+} and Fe^{3+} Fe as an atom has 26 electrons



Fe^{2+} - lose the 4s electrons

$1s^2 2s^2 2p^6 3s^2 3p^6 3d^6$

Fe^{3+} - lose 1 3d electron as well to have a half full energy level

$1s^2 2s^2 2p^6 3s^2 3p^6 3d^5$

Exercises in book pg 10

Exam Question 2013

QUESTION ONE

(a) Complete the following table.

Symbol	Electron configuration
Se	$[Ar]3d^{10}4s^24p^4$ or $4s^23d^{10}4p^4$
V	$[Ar]3d^34s^2$ or $4s^23d^3$
V^{3+}	$[Ar]3d^2$

2 rows correct required for Achieved