Chemistry 90780
Describe properties of particles and thermochemical principles

The achievement standard in a nutshell.

Properties of particles **include:**
Electron configuration of atoms and ions of the first 36 elements (using *s,p,d* notation)

Key Ideas:

Electrons are found in regions of space called **orbitals** around the nucleus.

Orbitals are put into energy levels (shells). The lowest energy level contains electrons closest to the nucleus.

The electrons increase in energy as they are found further from the nucleus.

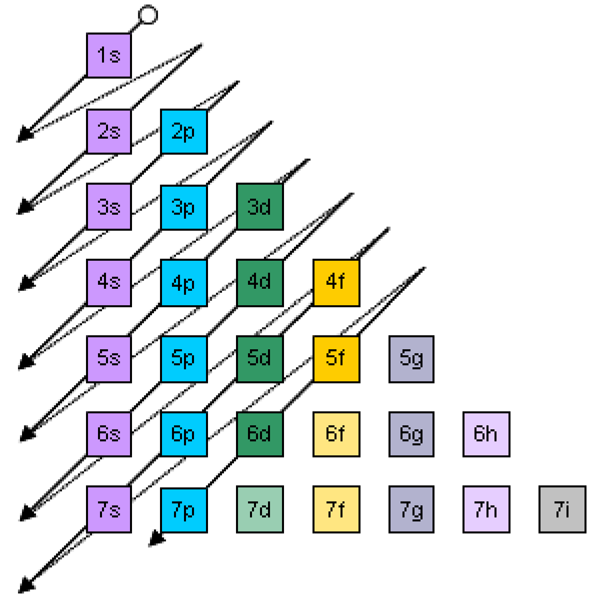
Each energy level can hold a maximum number of electrons.

Energy Level	Maximum number of electrons	Orbitals available to put them in.	Diagram
4	**32**	**s, p, d and f**	4s 4p 4d 4f
3	**18**	**s, p and d**	3s 3p 3d
2	**8**	**s and p**	2s 2p
1	**2**	**s**	1s

The outer electron shell is called the valence shell. It holds a maximum of 8 electrons.

We use spd notation to record where the electrons are found in an atom or ion. ***Always fill orbitals starting with the lowest energy first***

Full subshells and half-filled subshells are more stable than a partially filled subshell.



Notice the order

1s 2s 2p 3s 3p 4s then 3d!

then 4p (no need to go on to fill 5s orbitals at Level 3 NCEA)

The 3d orbitals have slightly higher energy than the 4s orbital and are filled after the 4s.

Key Ideas Know how to write s p d notation for atoms and ions of the first 36 elements of the Periodic Table

Some examples (use your periodic table to find the atomic number).

For Atoms – number of electrons = the atomic number

Eg Carbon – atomic number 6 so has 6 electrons

$1s^2\ 2s^2\ 2p^2$

The energy level number or shell – **start with lowest energy first**

The orbital letter containing the electrons. **Must fill according to the filling rules** (learn) Lower case s,p or d

The number of electrons. **Must be a superscript**

Always check that the number of electrons = the atomic number for that particular atom.

Eg Chlorine (atomic number 17) = 17 electrons

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

Chlorine is a **p block element**

There are **7 electrons** in the valence shell (outermost shell) with **2** electrons in the **3s** subshell and **5** electrons in the **3p** subshell.

(p subshells can hold up to 6 electrons.
d subshells can hold up to 10 electrons.
s subshells can hold up to 2 electrons).

Eg Iron (atomic number 26) = 26 electrons

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

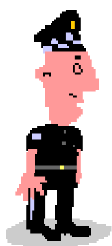
Electrons go into **4s before 3d** because 4s is at a slightly lower energy level than 3d.

Iron is a **d block element**

When using spd notation you can use a **noble gas core** to represent part of the configuration. This allows you to focus on the valence electrons.

Chlorine would become **[Ne] $3s^2\ 3p^5$**

Iron would become **[Ar] $4s^2\ 3d^6$**



Unusual electronic configurations for atoms – don't forget copper and chromium are not quite what you would expect!

Copper (Atomic number 29) **$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^1$** – (a full 3d subshell is more stable than a partially filled one).

Chromium (Atomic number 24) **$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5\ 4s^1$** (a half filled 3d subshell is more stable than a partially filled one).

What about the electronic (electron) configuration of ions?

1) Metal atoms lose electrons to form positively charged ions (Cations)

2) Non-metal atoms can gain electrons to form negatively charged ions (Anions)

3) Electrons are lost or gained so that the valence shell is filled.

Aluminium (atomic number 13) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$

Aluminium atoms **lose 3 electrons** from the valence shell (3^{rd} shell) to form an aluminium ion of +3 charge. $\mathbf{Al^{3+}}$

Aluminium ion (atomic number 13) has now got 13 protons but only 10 electrons. This accounts for the **+3** electrical charge.

s p d notation must be $1s^2\ 2s^2\ 2p^6$

Valence shell contains 8 electrons

Chlorine (atomic number 17) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$

There are 7 valence electrons. **A chlorine atom gains 1 electron** to form a chloride ion of -1 electrical charge, $\mathbf{Cl^-}$

Chloride ion (atomic number 17) has now got 17 protons but 18 electrons. This accounts for the **-1** electrical charge.

s p d notation must be $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

Valence shell contains 8 electrons

What about transition metal ions – what is the rule for s p d notations?

The 4s electrons are lost first when d-block elements form ions. (You must remember this one!)

Eg Iron (atomic number 26) = 26 electrons

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

Lose 2 electrons to form Fe^{2+}

Lose 3 electrons to form Fe^{3+}

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

The two 4s electrons are lost first.
This is the **Fe^{2+}** ion.

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

The two 4s electrons are lost as they are furthest from the nucleus- followed by an electron from the 3d subshell.
This is the **Fe^{3+}** ion

Special characteristics of transition metals (variable oxidation state, colour) related to electron configuration. Transition metals will be limited to iron, vanadium, chromium, manganese, copper and zinc.

<u>What is an oxidation state</u>? The oxidation state of an atom is a description of how many electrons it has lost or gained from its original state.

Example: Fe (oxidation state 0) means the iron atom which has its normal allocation of electrons (26) has not lost or gained any yet).

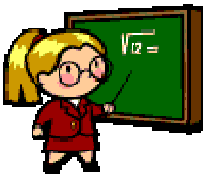
If an iron atom loses 2 electrons it forms an ion of +2 charge. Two electrons are lost. The iron is oxidised and has a **+ 2 oxidation state**. It forms Fe^{2+}

If iron loses 3 electrons to form the Fe^{3+} ion, we say the **oxidation state is +3**.

If a chlorine atom gains an electron it forms a chloride ion Cl^- with an **oxidation state of -1.**

Oxidation causes an increase in the oxidation number

Reduction causes a decrease in the oxidation number.



How to work out the oxidation state of an atom.

• The **oxidation state of an uncombined element is zero**. That's obviously so, because it hasn't been either oxidised or reduced yet! Eg , Ne or Cl_2 or S_8, or P_4

• The **sum of the oxidation states of all the atoms or ions in a neutral compound is zero.**

• The **sum of the oxidation states of all the atoms in an ion is equal to the charge on the ion.**

• The **more electronegative element in a substance is given a negative oxidation state**. The less electronegative one is given a positive oxidation state.

• Some elements almost always have the <u>same oxidation states</u> in their compounds:

Oxygen is usually -2 (-1 in peroxides).

Hydrogen is +1 (-1 in hydrides).

Transition metal atoms show variable oxidation states in their compounds. Why is this?

Transition metal atoms will form at least one ion with a partially filled 3d subshell of electrons.

(This means scandium and zinc are not transition metal atoms – and do not show variable oxidation states in their compounds).

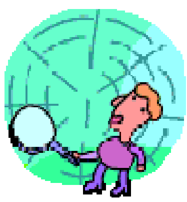
Transition metal atoms can form ions of variable oxidation states because after the 4s electrons are lost, **varying numbers of 3d electrons can also be lost.**

Iron, vanadium, chromium, manganese, copper and zinc are examples for Level 3 NCEA Chemistry

Other properties with respect to the stability of oxidation states:

- Ions in **higher oxidation** states tend to make good **oxidising agents**, whereas elements in low oxidation states become reducing agents.
- The 2+ ions across the period start as strong reducing agents and become more stable.
- The 3+ ions start stable and become more oxidising across the period.

Iron	common ions formed are	$\mathbf{Fe^{2+}}$
		$\mathbf{Fe^{3+}}$
Vanadium	common ions formed are	$\mathbf{V^{3+}}$
Manganese	common ions formed are	$\mathbf{Mn^{2+}}$
Copper	common ions formed are	$\mathbf{Cu^{+}}$
		$\mathbf{Cu^{2+}}$
Chromium	common ions formed are	$\mathbf{Cr^{3+}}$
Zinc	common ions formed is	$\mathbf{Zn^{2+}}$



Test – can you write s p d notations for all of the above ions?

Check your answers below!

Fe^{2+} [Ar] $3d^6$ Fe^{3+} [Ar] $3d^5$ V^{3+} [Ar] $3d^2$ Mn^{2+}[Ar] $3d^5$ Cu^{+} [Ar] $3d^{10}$

Cu^{2+} [Ar] $3d^9$ Cr^{3+} [Ar] $3d^3$ Zn^{2+} [Ar] $3d^{10}$

Why are transition metal compounds often coloured?

example	$Cr_2O_7^{2-}$	CrO_4^-	Fe^{2+}(aq)	Cu^{2+} (aq)	MnO_4^-(aq)
	dichromate	chromate	Iron (II)	Copper (II)	Manganate (VII)

The above solutions appear the colours they do because when white light is passed through a solution of a particular ion, **some of the energy in the light is used to promote an electron from the d orbitals into orbitals of higher energy**.

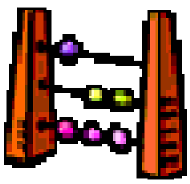
The "missing wavelengths" of light cause the eye to see a colour that is a mix of the wavelengths that pass through the solution.

Zinc and scandium cannot form coloured compounds because **they do not have partially filled 3d subshells in their ions. There are no 3d electrons to absorb energy in the light and move to a higher energy level.**

Look at the spd notations to see why …

Zn [Ar] $3d^{10}\,4s^2$

Zn^{2+} [Ar] $3d^{10}$



Sc [Ar] $3d^1 4s^2$

Sc^{3+} [Ar] $3d^0$