The Electronic Configurations and Oxidation states

The following table represents the electronic configuration of the elements of the first transition series and their common oxidation states. **(For illustration).**

The Element	The Group	The Electronic Configuration	Oxidation States	Some of the compounds
$_{21}Sc$	IIIB	[Ar], $4s^2$, $3d^1$	3	Sc_2O_3
$_{22}Ti$	IVB	[Ar], $4s^2$, $3d^2$	2,3,4	TiO_2, Ti_2O_3, TiO
$_{23}V$	VB	[Ar], $4s^2$, $3d^3$	2,3,4,5	V_2O_5, VO_2, V_2O_3, VO
$_{24}Cr$	VIB	[Ar], $4s^2$, $3d^4$ expected ✗ [Ar], $4s^1$, $3d^5$ actual ✓	2,3,6	CrO_3, Cr_2O_3, CrO
Highest oxidation $_{25}Mn$	VIIB	[Ar], $4s^2$, $3d^5$	2,3,4,6,7	MnO_2, Mn_2O_3, MnO $KMnO_4$, K_2MnO_4
$_{26}Fe$	VIII	[Ar], $4s^2$, $3d^6$	2,3,6	Fe_2O_3, FeO
$_{27}Co$	VIII	[Ar], $4s^2$, $3d^7$	2,3,4	$CoCl_3$, $CoCl_2$
$_{28}Ni$	VIII	[Ar], $4s^2$, $3d^8$	2,3,4	NiO_2, Ni_2O_3, NiO
Lowest oxidation $_{29}Cu$	IB	[Ar], $4s^2$, $3d^9$ expected ✗ [Ar], $4s^1$, $3d^{10}$ actual ✓	1,2,3	CuO, Cu_2O
$_{30}Zn$	IIB	[Ar], $4s^2$, $3d^{10}$	2	ZnO

Configuration		4s	3d				
$_{21}Sc$	[Ar], $4s^2$, $3d^1$	↑↓	↑				
$_{22}Ti$	[Ar], $4s^2$, $3d^2$	↑↓	↑	↑			
$_{23}V$	[Ar], $4s^2$, $3d^3$	↑↓	↑	↑	↑		
$_{24}Cr$	[Ar], $4s^1$, $3d^5$	↑	↑	↑	↑	↑	↑
$_{25}Mn$	[Ar], $4s^2$, $3d^5$	↑↓	↑	↑	↑	↑	↑
$_{26}Fe$	[Ar], $4s^2$, $3d^6$	↑↓	↑↓	↑	↑	↑	↑
$_{27}Co$	[Ar], $4s^2$, $3d^7$	↑↓	↑↓	↑↓	↑	↑	↑
$_{28}Ni$	[Ar], $4s^2$, $3d^8$	↑↓	↑↓	↑↓	↑↓	↑	↑
$_{29}Cu$	[Ar], $4s^1$, $3d^{10}$	↑	↑↓	↑↓	↑↓	↑↓	↑↓
$_{30}Zn$	[Ar], $4s^2$, $3d^{10}$	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓

1. After $_{20}Ca$, there is a gradual filling of the five orbitals of (3d) sublevel by single electron in each orbital in sequence till manganese ($3d^5$).
After manganese, pairing of electrons takes place in each orbital till zinc ($3d^{10}$) According to **Hund's rule.**

2. **Chromium and copper** are **abnormal** from the expected electronic configuration as their configurations are

$_{24}Cr$ (Ar) $3d^5\ 4s^1$ and $_{29}Cu$ (Ar) $3d^{10}\ 4s^1$.

Because their outer most sublevels 4s and 3d in **chromium** atom are half filled while in **copper** 4s is half filled but 3d is completely filled,
lowering their energy, and more stable.

Note on stability of atoms

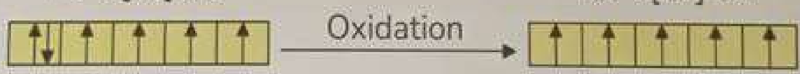
The atom become more stable when their outmost sublevel is:

1.Completely filled. 2. Half filled.

1.iron (II) is easily converted to iron (III)

The electronic configuration of an iron atom is: **$_{26}Fe$: [Ar]$4s^2 3d^6$**

Fe^{2+}: [Ar] $3d^6$ Fe^{3+}: [Ar] $3d^5$



Less stable half-filled (d^5) more stable

Iron (III) ion is **more stable** as the 3d sublevel is **half-filled** $3d^5$ so the reaction goes toward the formation of the more stable compound.

2. Mn (II) is not easily converted to Mn (III)

The electronic configuration of an iron atom is: **$_{25}Mn$: [Ar]$4s^2 3d^5$**

Mn^{2+}:$3d^5$ Mn^{3+}:$3d^4$

half-filled (d^5) more stable Less stable

We find that the 3d sublevel is **half-filled** $3d^5$ in Mn^{+2} ion, so it is **more stable** than Mn (III) ion, then Mn^{+2} ion is not easily (readily) oxidized to Mn^{+3} ion.

3. All elements of the first transition series have oxidation state **(+2)** after losing the electrons of (4s) sublevel at first, **except scandium,** it gives oxidation state **(+3) only.**

Because the energy needed to lose two electrons from 4s is nearly equal the energy needed to lose one electron from 3d
and to be more stable when outer sublevel is completely filled (noble gas)

4. The oxidation states **increase** from scandium to manganese which has the highest oxidation state **(+7)** in group VIIB, after that the oxidation states **decreases** gradually to be **(+2)** in zinc in group IIB, from the previous , we find that the maximum oxidation state doesn't exceed its group number
except group IB that contains **copper, silver and gold (Coinage metals).**

5. The main transition elements are characterized by several oxidation states (while representative metals have one oxidation state).

Because the sublevels 4s and 3d are close in energy so the atom of the main transition elements is oxidized by losing electrons from 4s then 3d sublevel in sequence.

when a transition metal loses electron → the ionization energy increase

while representative metals have one oxidation state
(like sodium Na, magnesium Mg and aluminum Al)

Because it is **difficult to break a complete energy level**, so it is difficult to obtain Na^{+2} or Mg^{+3} or Al^{+4} during the chemical reaction under normal conditions

The transition elements

It is the element in which the orbitals of (d or f) sublevels are occupied with electrons but not completely filled either in its atomic state " free state" or in one of its oxidation states (compounds).

Give Reason

Coinage metals (group IB) which are Cu ($3d^{10}$ $4s^1$), Ag ($4d^{10}5s^1$) and Au ($5d^{10}$ $6s^1$) can be considered as transition elements?

Because their (d) sublevel is completely filled (d^{10}) in their atomic state, but in the oxidation state (+2) or (+3) the sublevel (d) will be partially filled (d^9) or (d^8), **so they are transition elements.**

$$_{29}Cu\ [Ar]\ 3d^{10}, 4s^1 \xrightarrow{\text{Lose } 2e^-} {}_{29}Cu^{+2}\ [Ar]\ 3d^9$$

Give Reason

Group IIB Zn ($4s^2$, $3d^{10}$), Cd ($5s^2$, $4d^{10}$), Hg ($6s^2$, $5d^{10}$) can't be considered as transition elements?

Because their (d) sublevel is completely filled (d^{10}) in both atomic state and oxidation state (+2 ion) **so they are not transition elements.**

$$_{30}Zn\ [Ar]\ 4s^2, 3d^{10} \xrightarrow{\text{Lose } 2e^-} {}_{30}Zn^{+2}\ [Ar]\ 3d^{10}$$