

3.1 THE PERIODIC TABLE

arrangement

- four blocks - s, p, d, f, in increasing atomic number
 - groups - vertical columns periods - horizontal rows.
 - GP 1 : alkali metals
 - 17 : halogens
 - 18 : noble gases
- } no of valence electrons
- ↳ no of shells (principle quantum no.)
- d block = transition metals
 - f block = lanthanoids (starting at La)
= actinoids (starting @ Ac)
 - metals vs non-metals divided by zig-zag line
↳ close to the diagonal lines are METALLOIDS
they have both metal + non-metal properties

3.2 PERIODIC TRENDS

physical properties include atomic radius, ionic radius, ionisation energy, electron affinity, electronegativity.
due to changes in **effective nuclear charge**.

outer electrons attracted by nucleus
outer electrons repelled by inner electrons.

effective nuclear charge is NET attraction factors:

- ① DISTANCE from nucleus to shells (no. of shells)
- ② SHIELDING from inner shells (↑ shielding ↓ attraction)
- ③ NUCLEAR CHARGE from protons (if shells + shielding = same)
- ④ SUB-SHELLS & ELECTRON PAIRING

→ **atomic radius** (measured as HALF distance between 2 bonded atoms)

↓ group ↑ atomic radius

because outer electrons = new energy level.

→ period ↓ atomic radius

because more protons but same shielding effect.

→ **ionic radius**

metal cations, non-metal anions increase in size ↓ group

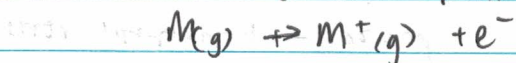
metal cations = smaller than metals' atoms = bc lost electron.

∴ reduction of shielding ∴ greater attraction

non-metal anions = larger than atoms = bc gained electron

∴ increase shielding ∴ increased repulsion.

→ **ionisation energy** is the minimum amount of energy required to remove one mole of electron from one mole of gaseous atoms



down a group, 1st IE ↓ (more shielding, further, more shells)

across a period 1st IE ↑ (increase in proton number, same shell)

→ **electronegativity** is the ability of an atom to attract a pair of electrons towards itself within a covalent bond

↓ group ↓ electronegativity ∴ electrons further, more shielded from nucleus.

→ period ↑ electronegativity ∴ proton no, same E level.

most electronegative = F, O, N

least electronegative = Fr

→ **electron affinity** is the energy change when one mole of electrons is added to one mole of gaseous atom



the first electron affinity is exothermic because the nucleus can still attract the electron ∴ it releases energy

∴ ΔH is negative (the ion loses energy)

when you **descend a group**, EA becomes **LESS NEGATIVE**

because the energy is needed to keep the electron there

∴ less energy released ∴ ΔH less negative.

→ more shielded, further away, more shells.

when you **go across a period**, EA becomes **MORE NEGATIVE**

it's easier to keep that electron there bc ↑ proton ↑ nuclear pull

∴ more energy released ∴ ΔH more negative

→ **metallic & non-metallic character**

metals tend to **LOSE** electrons : low IE levels

non-metals **GAIN** electrons : high IE levels

or **SHARE**

largely exothermic (v. negative ΔH)

Fluorine = most?

high electro-negativity

high electron affinity.

most metallic : Fr

increase down a group

decrease across period.

chemical properties of Alkali metals, Halogens & oxides

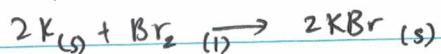
the Alkali metals

- group 1 metals react with WATER to produce hydroxide + Hydrogen



↓ group ↑ reactivity (IE is lower, easier to remove)

- they also react with HALOGENS to produce ionic salts



the Halogens

- get less reactive ↓ group
- after reaction = less attraction between nucleus + valence electrons

∴ ↑ distance ↑ shielding

- electron affinity & negativity decrease down a group



↑ more reactive displace less reactive.

↳ halogen is reduced (gain electron) ∴ is an OXIDISING agent

↳ halide is oxidized (lose electron)

REDUCING agent

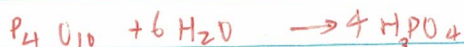
↳ Br is best

F is best

metal and non-metal oxides

metals - form alkalis

non-metals - form acidic



Al₂O₃ and SiO₂ are INSOLUBLE

↓
amphoteric

↓
still considered acidic (bc it will react w/ alkalis)

Na₂O + Al₂O₃ are ionic = high mpt, ✓ conduct electricity ✓ strong attraction

SiO₂ is giant covalent = ✓ strong, high mpt X conduct

P₄O₁₀, SO₂, CO₂, NO₂ = simple covalent X conduct ↓ melting pt.

1. The first step in the process of...

is to identify the...

2. The second step is to...

analyze the data...

3. The third step is to...

interpret the results...

4. The final step is to...

5. The next step is to...

evaluate the findings...

6. The final step is to...

draw conclusions...

7. The next step is to...

8. The final step is to...

9. The next step is to...

10. The final step is to...

11. The next step is to...

12. The final step is to...

13. The next step is to...

14. The final step is to...

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21.

22. The next step is to...

23. The final step is to...

24. The next step is to...

25. The final step is to...

13.1 FIRST ROW D-BLOCK ELEMENTS

transition metals have certain properties: variable ox. states, catalytic & magnetic properties, form complex ions (ligands) & have coloured compounds.

↳ they have the same outermost electron subshells.

↳ all of them will have ions that make partially filled d-shells
Sc and Zn are not transition metals.

variable oxidation states

• all lose $4s^2 \rightarrow$ charge of $2+$ (0.5 of $+2$)

• it is also easy to lose d-shell electrons because $4s$ and $3d$ overlap and have a similar energy level.

successive IE (how much energy to remove an electron) is small

e.g. manganese



easy to remove 1, 2, 3, 4, 5, 6, 7 of d-shell

catalytic properties

heterogeneous: catalyst in diff state from reactants

↳ T-metals allow reactants to ADSORB \therefore more contact

Fe in Haber, Ni in hydrogenation, Pd/Pt in catalytic converters

MnO₂ in H₂O₂ decomposition

homogeneous: catalyst in same state of reactants.

↳ T-metals: > 1 stable O.S. \therefore redox reactions \checkmark

: coordinate bonds

Fe²⁺ in heme

Co³⁺ in B₁₂ vitam.
B₁₂ vitam.

magnetic properties

- every spinning electron can be a magnet
- if there are $\uparrow\downarrow$ (opp direction) = no magnetic effect
- but some T-metals have unpaired electrons \rightarrow can align \rightarrow magnet.

\hookrightarrow para/dia/ferromagnetic

diamagnetism weak magnet ($\uparrow\downarrow$ electrons)

paramagnetic ok magnet (an unpaired electron $\uparrow\downarrow$ 1)

ferromagnetic strong magnet ($\uparrow\uparrow\uparrow$ a lot unpaired)

e.g. Fe $\uparrow\downarrow$ \uparrow \uparrow \uparrow is ferromagnetic (Ni + Co also)

Cu $\uparrow\downarrow$ $\uparrow\downarrow$ $\uparrow\downarrow$ $\uparrow\downarrow$ is diamagnetic

Cr $\uparrow\downarrow$ \uparrow \uparrow \uparrow is paramagnetic

$\uparrow\uparrow\uparrow\uparrow$

$\uparrow\downarrow$ $\uparrow\downarrow$ $\uparrow\uparrow\uparrow$

complex formation

T-metal ions are SMALL so they can attract LIGANDS

ligands: non-bonded pair of electrons, and donate this to an empty orbital of a T-metal ion

forming a COORDINATE BOND

a number of ligands \Rightarrow 1 metal ion = complex ion
 $[H_2O, NH_3, Cl^-, CN^-]$ \hookrightarrow many coordinate bonds.

no. of ligands \Rightarrow to metal = coordination number.

depends on ligand size

and empty orbital availability.

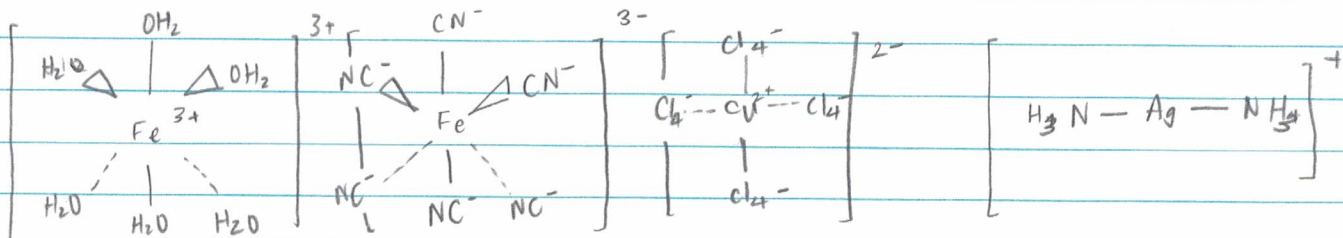
e.g. Cl^- is > bigger than H_2O

only 4 Cl^- fit into metal ion > 6 H_2O or NH_3

monodentate ligands: form ONE COOD bond w/ metal

bidentate ligands: form TWO COOD bonds

polydentate ligands: form THREE + (two + atoms with a lone pair)

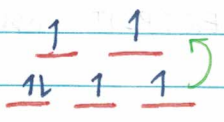


change, metal, ligand, size of ligand.

COLOURED COMPLEXES

in a d orbital, all the electrons have the same energy - are in the same energy level [degenerate] when the atom is ISOLATED

but... when you INTRODUCE another atom to form a "complex," the sub-shell SPLITS into 2 different energy levels



- ① shine light through
- ② electrons **absorb** light (visible light) and **jump** energy levels
- ③ a different wavelength/frequency is produced (different colour).

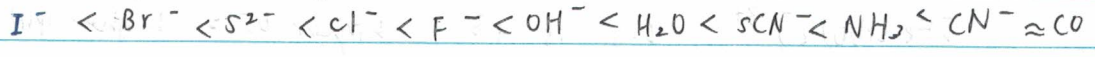
- the greater the splitting \downarrow
- the more energy required to promote an electron
- larger frequency (shorter wavelength)

* zinc has a full d shell cannot promote electron

CONCLUSION

- ① d-orbital splits
- ② particular frequency of light is absorbed by electron which moves from lower to higher energy levels
- ③ complementary frequency of light TRANSMITTED.

Different ligands affect colours differently



according to the energy difference they produce between 2 sets of d-orbitals

\uparrow
lower frequency of light

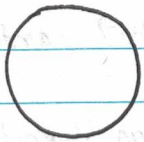
THE COLOUR DEPENDS ON... (amt of wavelength absorbed)

- A. Nuclear charge & identity of metal ion
more $[Cu(H_2O)_6]^{2+}$ = blue $[Co(H_2O)_6]^{2+}$ = pink less.
- B. Oxidation state of metal ion (changes repulsion between lig and t ion)
more $[Fe(H_2O)_6]^{2+}$ = green $[Fe(H_2O)_6]^{3+}$ = yellow less
- C. geometry of complex (coord. no = changes repulsion)
- D. identity of ligand (spectrochemical series)

the colours depend on the split of electron shells

larger split, larger jump, larger wavelength absorbed ROYGBIV

e.g. $[\text{Mn}(\text{H}_2\text{O})_6]^{2+}$ vs $[\text{Fe}(\text{H}_2\text{O})_6]^{2+}$ ↑ largest
25 protons vs 26 protons
↳ smaller split = absorbs RED ↳ BIGGER SPLIT absorbs BLUE



identity of the ligands

$\text{I}^- < \text{Br}^- < \text{S}^{2-} < \text{Cl}^- < \text{F}^- < \text{OH}^- < \text{H}_2\text{O} < \text{SCN}^- < \text{NH}_3 < \text{CN}^- \approx \text{CO}$

$[\text{Cu}(\text{H}_2\text{O})_6]^{2+}$ absorbs orange absorbed

$[\text{Cu}(\text{H}_2\text{O})_5\text{NH}_3]^{2+}$ MORE +

$[\text{Cu}(\text{H}_2\text{O})_4(\text{NH}_3)_2]^{2+}$ MORE as NH_3 is higher yellow absorbs
↳ bigger split

oxidation state

$[\text{Fe}(\text{H}_2\text{O})_6]^{2+}$ O.S is +2 lower OS = lower split absorbs green

$[\text{Fe}(\text{H}_2\text{O})_6]^{3+}$ O.S is +3 higher OS = higher split absorbs violet