

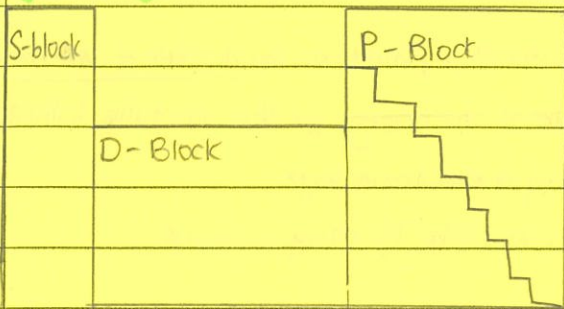
3

Periodicity → recurring trends in element properties, with increasing proton number

The table

Date

No.



Properties of Metals

- Shiny + high m.p./b.p
- Sonorous
- Malleable
- ductile
- Good conductors of heat and electricity

Periodic Trends [Across Period 3]

1) Atomic Radius

- The number of protons and valence electrons increase across the period
- Hence, the attractive forces between the protons and valence electrons increases
- As a result, the atomic radius decreases across a period because the attractive forces holding pull the shells towards the nucleus

2) Electronegativity (make reference to how shielding is similar)

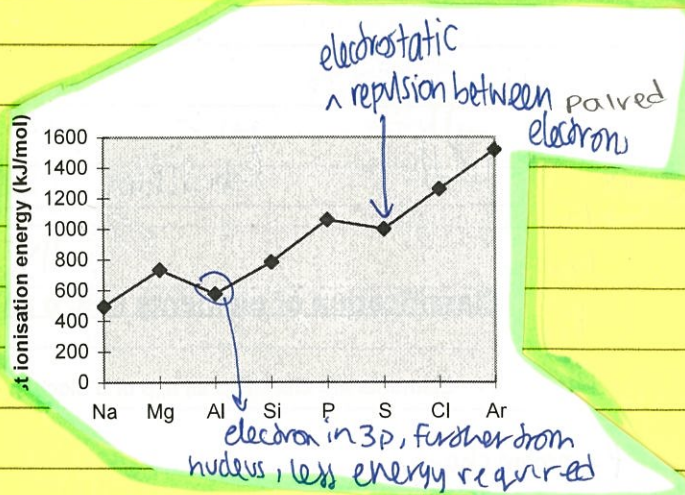
- definition: ability of an atom to attract a pair of electrons in a covalent to itself
- Electronegativity increases across a period, as nuclear charge increases, due to a rise in proton number, hence so does the attractive forces between the element and the shared pair of electrons in the covalent pair
- However, it decreases down the group due to decreasing effective nuclear charge [↑ shielding, atomic radius]

3) 1st ionisation energy

- generally increases across period as the number of protons increases, hence the attractive forces holding the electrons in place increases
- When the 3p shell starts to get filled, the electron is further from the nucleus, hence less energy is required to remove it
- There is a dip at sulfur because the p shell fills up for a second time, hence repulsive forces between electrons increase, decreasing FIE.

FIE.

Note: Sulfur has lower FIE due to presence of paired electron



4) Electrical Conductivity

- electrical conductivity of metals increases as you go across a period, as the number of valence electrons increases, hence there are more delocalised electrons to carry in the metallic lattice, hence there are more charged particles available to carry an electrical charge
- in non-metals, there are no charged particles available to carry a charge
- Si is slightly above O as it is a metalloid

5) M.P and B.P

Metals

- More protons as you go across, ↑ attractive forces between delocalised electrons and protons in the nucleus
- There will also be an increasing charge on the cations in the metal lattice, hence the number of delocalised electrons increases and the attractive forces holding the lattice together increases
- There is a jump in Si because it exists in a macromolecular structure, hence more bonds need to be broken.

Non-Metals (affected by size of elemental state)

- lower m.p because they are simple covalent structures
- higher at S as it exists as S_8
- P exists as P_4 , hence T_m higher
- The larger the molecule, the stronger the Van der Waals forces holding the molecule together

⚡ Periodic Trends

Group 1

No flame → Li			
Na	↓ b.p+m.p ↑	↓ softer as you descend	↓ reactivity ↑ as you descend
K			
Rb			
Cs			

* Reactivity increases as you descend because they react by losing their outer electrons to form their ions. The outer electron is in successively higher energy shells as you descend, where forces of attraction holding them together are weaker. Hence, less energy is required to remove it

Physical Properties

- good conductors of heat and electricity
- low densities

chem Properties

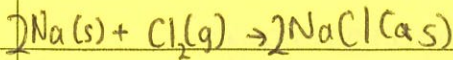
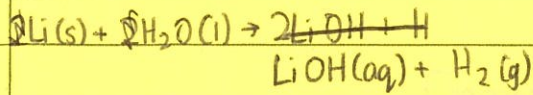
- reactive
- form ionic compounds with non-metals

Reactions of P3 elements and Group I elements

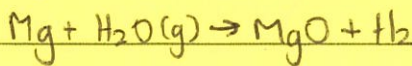
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Group I

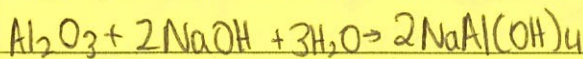
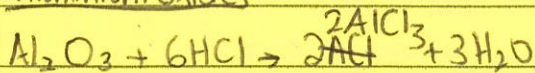


Magnesium with steam

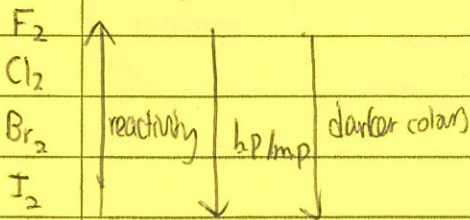


Across Table → alkali oxides → amphoteric → acidic

Aluminium oxides



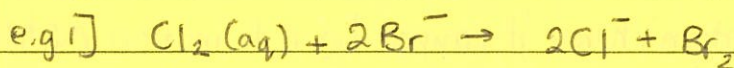
Halogens



* Reactivity decreases as you ascend

the outer shell is increasingly at higher energy levels and further from the nucleus. This decreases the attractive forces that pull an electron into the valence shell of a halogen, decreasing its reactivity.

The halogens can displace less reactive halogens as the halogens that are more reactive are better oxidising agents, hence they gain electrons from the other halogen that they oxidise.



Phys Properties

→ coloured compounds

Tests

	Halide	+ AgNO ₃	+ dilute NH ₃	+ conc NH ₃	+ Pb(NO ₃) ₂
F ⁻ very soluble ↑	F ⁻	colourless	no NIL ppt redissolves	NIL ppt redissolves	white
	Cl ⁻	white	ppt dissolves none	ppt redissolves	white
	Br ⁻	cream	none	ppt redissolves none	yellow
	I ⁻	yellow	ppt redissolves none	none	yellow

Noble Gases

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- colourless gases
- Monatomic
- very unreactive [because they cannot gain/lose electrons]

Do not form cations as they have the highest first ionisation energies

Do not form anions as extra electrons would have to be added to an empty outer energy level, where they would experience a negligible effective nuclear pull up.

Transition Metals

Definition: A metal that forms at least 1 stable ion with a partially full d-shell of electrons

Physical Properties

- high electrical and heat conductivity
- high m.p
- malleable / ductile
- high tensile strength

Chem Properties

- varying oxidation numbers
- coloured compounds
- catalyst in elements / compounds

Atomic Radii

- There ~~is~~ is a relative decrease (small) in atomic radii across d-block due to a small increase in effective nuclear charge across the d block. The increase is small because the charge is largely offset by the addition of one electron.
- The ~~similarity~~ similarity in atomic radii results in an ease for transition metals to form ~~the~~ alloys

Why Zn is not a transition metal

- Zn only forms the $2+$ oxidation state in its compounds
- d shell is full in both ion and atom, hence it cannot be a transition metal as a transition metal ion has an incomplete d shell.

Why Transition Metals have Varying oxidation states

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- 1) When there is a large jump in ionisation energy, the ion that is formed after that jump is energetically unstable, hence it rarely exists.
- 2) The increase in ionisation energies are gradual for transition metals as the 4s and 3d shells are close in energy. However, when the inner 3p electron is removed, the ion is energetically unstable due to a large jump in ionisation energy. Hence as electrons are being removed from 4s and 3d, stable ions are formed.

Notes on transition metals

- all transition metals show 2+ and 3+ oxidation states but 3+ becomes less common as you go across a period because ↑ nuclear charge hence ↑ 3rd ionisation energy
- Oxidation states above +3 show covalent character. Higher charge ions can polarise negative ions and increase the covalent character of the compound
- ↑ oxidation, ↑ ability as a oxidising agent.

Examples of Metal Catalysts

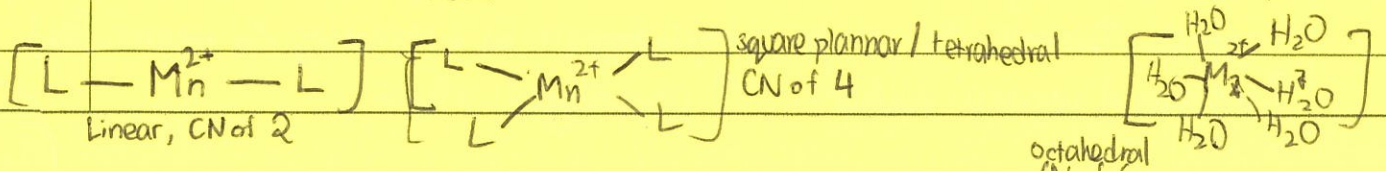
- V₂O₅ in contact process
- Fe in Haber Process

Coordinate bonds and ligands

Transition metal ions in solution have a high charge density and as a result, they act as Lewis acids and attract species rich in electrons. These species are known as ligand, neutral molecules or anions that contain one or more non-bonding pairs of electrons. The ligands form covalent bonds with the central transition metal ion to form a complex ion, (dative) → represented by an arrow from Lewis base (a.k.a. ligand)

A complex ion has a central metal ion at its centre with a number of other molecules surrounding it

The number of coordinate bonds to one central ion is the coordination number



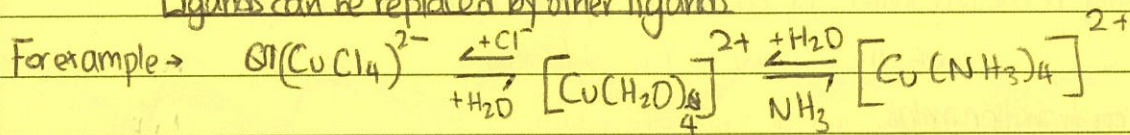
Examples of Ligand complexes

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Complex	Ligand	CN	Shape
$[\text{Fe}(\text{H}_2\text{O})_6]^{2+}$	H_2O	6	octahedral
$[\text{CuCl}_4]^{2-}$	Cl^-	4	tetrahedral
cisplatin) $[\text{Pt}(\text{Cl}_2)(\text{NH}_3)_2]$	Cl^- and NH_3		
MnO_4^-	O^{2-}	4	tetrahedral

Ligands can be replaced by other ligands



If one ligand is higher on the spectrochemical series than another, it can displace that ligand.

Polydentate Ligands \rightarrow ligands that can utilise 2 or more pairs of unbonded electrons to form a ~~separate~~ coordinate bond. Examples include oxalate ions $(\text{COO}^-)_2$ and EDTA and EDTA^{4-} (Ethylenediamine Tetra Acetic acid). EDTA forms a hexadentate ligand as there are 6 ~~un~~ unbonded pairs of electrons available in a single molecule of EDTA. For example, the coordination number of $[\text{Cu}(\text{EDTA})]^{2-}$ is 6 as 6 bonds are made to the central metal ion.

Magnetism

- \rightarrow every single spinning electron can behave as a small magnet
- \rightarrow Electrons with opposite spins (paired electrons) have opposing orientation and hence ~~cancel the~~ ^{have no} net magnetic effect

Types of Magnetism

Ferromagnetism \rightarrow metals like Fe, Co and Ni are ferromagnetic. Their unpaired electrons line up with parallel spins in regions called domains, irrespective of whether an external electric or magnetic field is present. Ferromagnetism is permanent magnetism

Paramagnetism \rightarrow paramagnetic metals contain unpaired electrons, each of which create a small magnetic field and will line up in domains when an electric or magnetic is applied. This makes the complex weakly magnetic, reinforcing the external field.

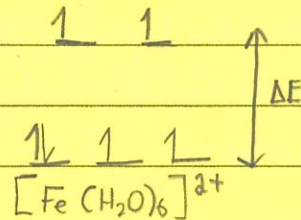
\uparrow number of unpaired electrons, \uparrow paramagnetic effect

Diamagnetism \rightarrow when all electrons in a complex are paired

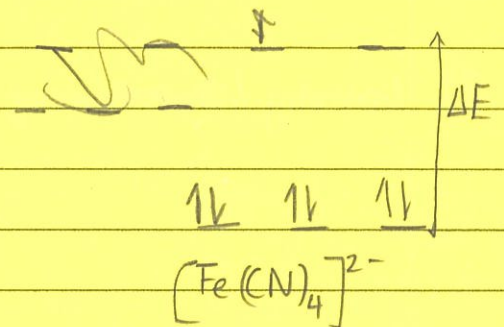
Iron ligands.



- \rightarrow can be dia or paramagnetic
- \rightarrow d orbitals are split by ligands
- \rightarrow \uparrow position of ligand on series, \uparrow ΔE (\uparrow splitting)
- \rightarrow if low ligands like H_2O are used (fig 2.1), electrons can occupy all d orbitals, leaving 4 unpaired electrons, a.k.a 4 units of paramagnetism



- \rightarrow if higher ligands like CN^- are used, higher splitting is observed, hence only lower d orbitals will be occupied with no unpaired electrons, making it diamagnetic



Coloured complexes

The colour of transition metal ions can be related to the presence of partially filled d orbitals. The ion Sc^{3+} is colourless as a complex ion because its 3d block is empty. Zn^{2+} is colourless because the 3d shell is full.

Ligands have the ability to split the d orbitals into 2 sub levels

The d orbitals in a isolated transition metal atom or ion, the d-orbitals are degenerate as they contain the same amount of energy. When a complex ion is formed with a ligand, the ligands act as Lewis ~~bases~~ ^{acids/bases} and donate a pair of non-bonding electrons to form a coordinate bond. As the ligands approach the d orbitals along the axis, the ligand electrons will repel the $d_{x^2-y^2}$ and d_{z^2} orbitals as they lie on the axis. As a result, the d orbitals are split, 2 to higher energy, 3 to lower energy. This difference in energy (ΔE) corresponds to the wavelength of light absorbed. ($\Delta E = h\nu$) ($c = \nu\lambda$). Using the colour wheel, we can see what wavelength of light is emitted/transmitted.

The frequencies of light absorbed by e^- as they move from lower to higher are complementary to absorbed.

Factors affecting the colour of a complex ion

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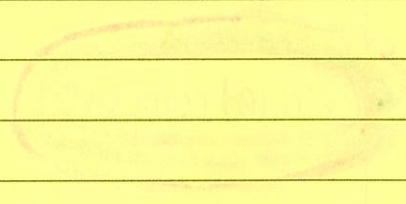
1) Nuclear charge

- ↑ protons, ↑ effective nuclear charge
- This increases the electrostatic attraction between the donated pairs of electrons and the nucleus
- Hence, ΔE is higher, resulting in a higher wavelength of light being emitted (transmitted)

2) Ligand identity

- The higher the ligand on the spectrochemical series, the higher the charge density
- The higher the charge density, the higher the split in the d orbitals, because ↑ repulsion with orbitals
- The higher the split, ↑ ΔE

3) Geometry / Stereochemistry



4) Oxidation states of the metal ion

- ↑ oxidation state, ↑ charge, ↓ number of electrons
- ↑ electron repulsion between the ligand and d electrons
- hence ↑ ΔE

Alkali Metals with water

- Li → floats, no flame, effervescent
 - Na → floats, flame produced (orange flame), w. lower b.m.p
 - K → Faster than Na, lilac flame
 - Rb → dense = sinks
- } all form alkaline solutions