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

**The table**

<table>
<thead>
<tr>
<th>S-block</th>
<th>P-Block</th>
<th>D-Block</th>
</tr>
</thead>
</table>

**Properties of Metals**
- Shiny + high m.p/l.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 increase.
   - As a result, the atomic radius decreases across a period because the attractive forces holding 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 bond 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 a period as the number of protons increases, hence the attractive force 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.
   - Note: Sulfur has lower FIE due to presence of 3p powered electron.

[Graph of ionisation energy (kJ/mol)]
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 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) M₇P and B₃P

- Metals

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

Non-Metals (affected by m.p size of elemental state)

- Lower m.p because they are simple covalent structures.
- Higher at S as it exists as S₂.
- P exists as P₄, hence the highest.
- The larger the molecule, the stronger the Van der Waals forces holding the molecule together.

Periodic Trends

Group 1

<table>
<thead>
<tr>
<th>No.</th>
<th>Name</th>
<th>Reactivity</th>
<th>Ionisation Energy</th>
<th>Reactivity increases as you descend because they react by losing their outer electrons to form their ion. The outer electron is in successively higher energy shells as you descend, where force of attraction holding them together are weaker. Hence, less energy is required to remove it.</th>
</tr>
</thead>
<tbody>
<tr>
<td>Li</td>
<td></td>
<td>1</td>
<td>Lower</td>
<td></td>
</tr>
<tr>
<td>Na</td>
<td></td>
<td>2</td>
<td>Lower</td>
<td></td>
</tr>
<tr>
<td>K</td>
<td></td>
<td>3</td>
<td>Lower</td>
<td></td>
</tr>
<tr>
<td>Rb</td>
<td></td>
<td>4</td>
<td>Lower</td>
<td></td>
</tr>
<tr>
<td>Cs</td>
<td></td>
<td>5</td>
<td>Lower</td>
<td></td>
</tr>
</tbody>
</table>
Physical Properties
- good conductors of heat and electricity
- low densities

Chemical Properties
- reactive
- form ionic compounds with non-metals

Reactions of P₃ elements and Group I elements

<table>
<thead>
<tr>
<th>Group</th>
<th>Reaction</th>
<th>Product</th>
</tr>
</thead>
<tbody>
<tr>
<td>Li(s) + H₂O(l)</td>
<td>2LiOH(aq) + H₂(g)</td>
<td>P₃ Oxides with H₂O</td>
</tr>
<tr>
<td>Na(s) + Cl₂(g)</td>
<td>2NaCl(aq)</td>
<td>P₃ elements with H₂O</td>
</tr>
</tbody>
</table>

Mg + H₂O(g) → MgO + H₂

Across the table: alkali oxide → amphoteric → acidic

Al₂O₃ + 6HCl → 2AlCl₃ + 3H₂O

Halogenes

<table>
<thead>
<tr>
<th>F₂</th>
<th>Cl₂</th>
<th>Br₂</th>
<th>I₂</th>
</tr>
</thead>
<tbody>
<tr>
<td>↑</td>
<td>↑</td>
<td>↑</td>
<td>↑</td>
</tr>
<tr>
<td>reactivity</td>
<td>h.p.</td>
<td>darkening color</td>
<td></td>
</tr>
</tbody>
</table>

Reactivity decreases as you ascend.

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

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

e.g. 1) Cl₂ (aq) + 2Br⁻ → 2Cl⁻ + Br₂

Phys Properties
- coloured compounds

Tests

<table>
<thead>
<tr>
<th>Halide</th>
<th>+ AgNO₃</th>
<th>+ dilute HCl</th>
<th>+ conc HCl</th>
<th>+ Pb(NO₃)₂</th>
</tr>
</thead>
<tbody>
<tr>
<td>F⁻</td>
<td>colourless</td>
<td>no N₂L</td>
<td>ppt reaction</td>
<td>white</td>
</tr>
<tr>
<td>Cl⁻</td>
<td>white</td>
<td>ppt chlorine</td>
<td>ppt solution</td>
<td>white</td>
</tr>
<tr>
<td>Br⁻</td>
<td>cream</td>
<td>none</td>
<td>ppt solution</td>
<td>yellow</td>
</tr>
<tr>
<td>I⁻</td>
<td>yellow</td>
<td>ppt red solution</td>
<td>none</td>
<td>yellow</td>
</tr>
</tbody>
</table>

F very soluble A
Noble Gases
- Colourless gases
- Monoatomic
- Very unreactive [because they cannot gain/lose electrons]
  Do not form cations as they have the highest first ionisation energy.
  Do not form anions as extra electron would have to be added to an empty extra 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

Chemical Properties
- Varying oxidation numbers
- Coloured compounds
- Catalyst \(_{\text{elemental}}\) compounds

Atomic Radii
- There 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 change is largely offset by the addition of one electron.
- The similarity in atomic radii results in an ease for transition metals to form ternary alloys.

Why Zn is not a transition metal
- Zn only forms the 2+ oxidation state in its compound.
- 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

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 at 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 a electron 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+ become 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

Coordination 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 ligands, 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) \[ \text{M}^{2+}+\text{L} \rightarrow \text{M}^{2+}\text{L} \text{ (dative)} \] repre...ing an arrow from trigonal Lewis base (a scalar 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

<table>
<thead>
<tr>
<th>Complex</th>
<th>Ligand</th>
<th>CN</th>
<th>Shape</th>
</tr>
</thead>
<tbody>
<tr>
<td>$\text{Fe(OH}_2\text{)}_6^{2+}$</td>
<td>H$_2$O</td>
<td>6</td>
<td>octahedral</td>
</tr>
<tr>
<td>$\text{[CuCl}_4\text{]}^{2-}$</td>
<td>Cl$^-$</td>
<td>4</td>
<td>tetrahedral</td>
</tr>
<tr>
<td>$\text{Fe(C_2\text{H}_2\text{O}_2}_2\text{)}^2-$</td>
<td>O$^-$ and NH$_3$</td>
<td>4</td>
<td>tetrahedral</td>
</tr>
<tr>
<td>MnO$_4^{-}$</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Ligands can be replaced by other ligands.

For example: $\text{Ag(CuCl}_4\text{)}^{2-} + \text{Cl}^- \rightarrow \text{CuCl}_4^{2-} + \text{AgCl} \downarrow$

$\text{CuCl}_4^{2-} + \text{H}_2\text{O} \rightarrow \text{Cu(H}_2\text{O}_4\text{)}^{2+} + \text{HCl}$

$\text{Cu(H}_2\text{O}_4\text{)}^{2+} + \text{NH}_3 \rightarrow \text{Cu(NH}_3\text{)}_4^{2+} + \text{H}_2\text{O}$

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

Polydentate Ligands: ligands that can utilize 2 or more pairs of unbound electrons to form a single coordinate bond. Examples include oxalate ions ($\text{C}_2\text{O}_4^{2-}$) and EDTA and EDTA$^{4-}$ (EthyleneDiamineTetraAcetic acid). EDTA forms a hexadentate ligand as there are 6 unpaired pairs of electrons available in a single molecule of EDTA. For example, the coordination number of $\text{[Cu(EDTA)}^{2-}$ is 6 as 6 bonds are made to the central metal ion.

Magnetism

- Every single spinning electron can behave as a small magnet
- Electrons with opposite spins (paired electrons) have opposing orientation and hence cancel the net magnetic effect.

Types of Magnetism

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

- Paramagnetism: paramagnetic metals contain unpaired electrons, each of which create a small magnetic field and will line up in domains when an electric or magnetic field is applied. This makes the complex weakly magnetic or reinforcing the external field. Increase in number of unpaired electrons, increase paramagnetic effect.
Diamagnetism: When all electrons in a complex are paired.

Iron ligands: \( \text{Fe} \)
- Can be dia or paramagnetic
- d orbitals are split by ligands
- Up position of ligand on series, \( \uparrow \Delta E \) (splitting)
- If low ligands like \( \text{H}_2\text{O} \) are used, \( \frac{1}{2} \) electrons can occupy all d orbitals, leaving 4 unpaired electrons, a.k.a. 4 units of paramagnetism
  \[
  \downarrow \quad \frac{1}{2} \quad \frac{1}{2} \quad \frac{1}{2} \quad \frac{1}{2} \quad \uparrow \Delta E
  \]
  \[
  [\text{Fe} \left( \text{H}_2\text{O} \right)_6]^{3+}
  \]
- If higher ligands like \( \text{CN}^- \) 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 metals can be related to the presence of partially filled d orbitals. The ion \( \text{Sc}^{3+} \) is colourless as a complex ion because its 3d block is empty. \( \text{Zn}^{2+} \) is colourless because its 3d shell is full.

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

The d orbitals in an 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 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 electron will repel the d\( \text{xy} \) and d\( \text{yz} \) 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 (\( \Delta E \)) 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 or transmitted.

The frequencies of light absorbed by e\textsuperscript{-} or those moves from lower to higher are complementary to absorbed.
Factors affecting the colour of a complex ion

1) Nuclear charge
   - $\uparrow$ protons, $\uparrow$ effective nuclear charge
   - This increases the electrostatic attraction between the donated pairs of electrons and the nucleus
   - Hence, $\Delta E$ is higher, resulting in a higher wavelength of light being emitted/demitted

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 of repulsion within orbitals
   - The higher the split, $\uparrow \Delta E$

3) Geometry/Stereochemistry

4) Oxidation states of the metalion
   - $\uparrow$ oxidation state, $\uparrow$ charge, $\uparrow$ number of electrons
   - $\uparrow$ electron repulsion between the ligand and d electrons
   - hence $\uparrow \Delta E$

Alkali Metals with water

Li $\rightarrow$ floats, no flame, effervescence
Na $\rightarrow$ floats, flame colored (orange flame), a lower k, m.p. \{ all form
K $\rightarrow$ Faster than Na, lite flame
Rb $\rightarrow$ dense, s int