Periodic table

Created	@June 8, 2025 1:19 AM
Tags	

Halogens

State and explain the relative thermal stability of hydrogen halides

- Thermal stability decreases down the group 17
- As H-X bond energy decreases as less effective orbital overlapping between orbitals of H atom and halogen atom as size of halogen atom increases resulting in longer and weaker H-X bond
- Hence less energy to break the H-X bond decreases
- HC*l* does not decompose even on strong heating. **[0.5]**
- HBr yields reddish-brown (accept “brown”) fumes of Br2 under strong heating. **[0.5]**
- HI gives violet fumes of I2 when red-hot rod is plunged into jar of HI. **[0.5]**

Trend in volatility down the group

- they have simple molecular structures with weak I’d-I’d between molecules
- Down the group, electron cloud size of halogen molecules increases and becomes more easily polarised, leading to stronger London dispersion forces, more energy to overcome the stronger I’d-I’d, bp increases and volatility decreases

Transition metal

Definitions:

- Transition metals are d block elements which form one or more stable ions with a partially filled d subshell
- Ligand: an ion or molecule that contains at least one atom bearing a lone pair of electrons which can be donated into a low-lying vacant orbital of a central metal cation, forming a dative covalent bond
- A complex contains a central metal cation, bonded ligands by dative covalent bonds

Why is copper/ any metal in d block element transition metals?

- it is a d-block element which forms one or more stable ions with a partially filled d-subshell

Questions regarding properties of transition metals

Physical properties:

Property	Explanation
Higher mp/mp than non-transition metals	transition metal can contribute the3d and 4s electrons for metallic bonding while non-transition metal can only contribute the 4s electrons for metallic bonding. The metallic bond is stronger for transition metal, and hence it has a higher melting and boiling point
Higher density	Density = mass/volume. Transition metals has a heavier relative atomic mass than non-tm, tm have a smaller atomic radius than non-tm .Hence, there are more tm atoms per unit volume than in non-tm and hence

Property	Explanation
	densities of tm is significantly greater than that of non-tm
Similar atomic radii across the period	Across the transition elements series, nuclear charge increases and electrons are addedto the inner principal quantum shell (3d subshell). Shielding effect for valence shell 4selectrons increases due to the increase in no. of inner shell (3d) electrons. Nuclearattraction for the valence electrons is almost constant, hence the atomic radii are relativelysimilar.

Chemical properties:

Properties	Explanation
Transition metals less reactive	Less likely to undergo oxidation unlike group 1 and 2 metals, compare E values
Exhibit variable oxidation states	Transition elements exhibit *variable oxidation states* in their compounds due to the *close similarity in energy* of the 3d and 4s orbitals. Hence, both 3d and 4s electrons can be removed to form stable ions of different oxidation states.

why is aqueous transition metal ions such as Cu2+ are coloured while zn2+ is colurless/ why certain transition metal ions produce colour less solutions

- In the presence of H2O ligands, repulsion between the lone pair electrons of ligands and the d orbital electrons causes the degenerate d-orbitals of the Cu2+ ion are split into two different energy levels with a small energy gap, ΔE.
- As 3d Su shell is partially filled, Electrons in the lower energy d-orbitals can absorb light of a certain wavelength with energy corresponding to the energy gap, ΔE, and be promoted to the higher energy d-orbitals (d-d transition).
- The colour of the complex is the complementary colour of the wavelength absorbed.
- For Zn2+(aq), the 3d orbitals are fully filled with 3d10 configuration. The is no d-d transition and all the light are transmitted and Zn2+(aq) appears colourless
- Look at the electronic configuration:
    - Has a empty d sub shell/ fully filled d subshell, no d-d transition, no light absorbed and appears colourless

Note: when question says aqueous ions it means that the ligand is H2O

Why do different complex ions have different colours?

Different ligands:

- The ligands are different in the two compounds, resulting in a different degree of replusion between the lone pair electrons of the ligands and d orbital electrons of transition metal ion
- The degenerate 3d orbitals are split into 2 different energy levels with an energy gap and the energy gap is is different for the two compounds.
- In each case, electrons in the lower energy 3d orbitals absorb visible light of a certain wavelength with energy correspond to energy gap as they get promoted to the higher energy d orbitals.
- Different wavelength of light is absorbed and hence we observed a different complemantary colour.

Different overall charge:

- As overall charges of complex ions are different, The number of d electrons are different, resulting in a different degree of replusion between the lone pair electrons of the ligands and d orbital electrons
- The degenerate 3d orbitals are split into 2 different energy levels with an energy gap, and energy gap is different for the two compounds.
- In each case, electrons in the lower energy 3d orbitals absorb visible light of a certain wavelength with energy correspond to energy gap as they get promoted to the higher energy d orbitals.
- Different wavelength of light is absorbed and hence we observed a different complementary colour.

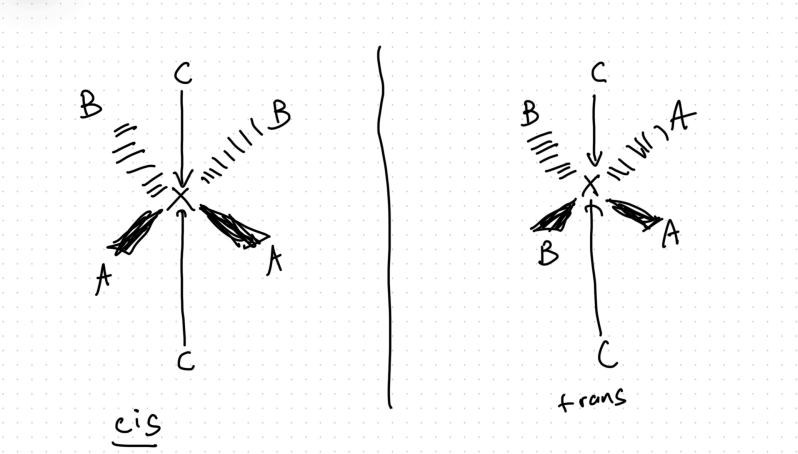
Why is carbon monoxide poisonous?

- Binds more strongly than oxygen with iron ion in haemoglobin to form very stable complexes, cutting down the supply of oxygen to the body

Why are transition metal ions able to form variable oxidation states unlike group 1 and 2 metals

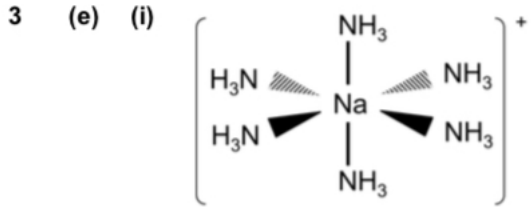
- Due to the close similarity in energy of the 3d and 4s orbitals ,different numbers of) 3d and 4s electrons can be removed to form stable ions of different oxidation states.
- However, once the valence electrons from s orbitals of group 1 and 2 are removed, the subsequent electrons removed must come from the next inner quantum shell which required large additional amount of energy and hence, does not have variable oxidation states.

Cis trans isomerism in complexes



For cis isomerism, the same atoms must face the same direction of the plane

Drawing 3D diagram for complex



- Need to show the 3d lines

Why is there different number of ligands surrounding the same central atom in different complexes

- Electron cloud of the ligand larger than the other ligand, less of that ligand can be placed around central atom or ion in a complex ion

Reaction of Cu complex with reagents

- Dilute sulfuric acid
  - Observation: blue solid dissolves in sulfuric acid to give a pale blue solution containing [Cu(H2O)6]2+
  - Type of reaction: acid base reaction
  - Equation: Cu(H2O)4(OH)2 + H2SO4 → [Cu(H2O)6]SO4 or Cu(H2O)4(OH)2 + H2SO4 → CuSO4 + 6H2O (depends on context of question)

- Excess ammonia
    - Observation: blue solid dissolves in excess to give a deep blue solution containing $[Cu(NH_3)_4(H_2O)_2]^{2+}$ or $[Cu(NH_3)_4]^{2+}$
    - Type of reaction: Ligand exchange reaction
    - Equation: $Cu(H_2O)_4(OH)_2 + 4NH_3 \rightarrow [Cu(NH_3)_4(H_2O)_2]^{2+} + 2OH^- + 2H_2O$ Or $Cu(H_2O)_4(OH)_2 + 4NH_3 \rightarrow [Cu(NH_3)_4]^{2+} + 2OH^- + 4H_2O$
- Concentrated HCl
- Observation: Blue solid dissolves in excess to give a yellow solution of $[CuCl_4]^{2-}$
- Type of reaction: Ligand exchange reaction
- Equation: $Cu(H_2O)_4(OH)_2 + 4Cl^- \rightarrow [CuCl_4]^{2-} + 2OH^- + 4H_2O$

Explain how some transition metal can act as catalyst

- have ions of variable oxidation states, allowing it to function as a homogenous catalyst
- Have partially filled d subshell allowing it to accept electrons from reactants and act as a heterogeneous catalyst

How are compounds containing copper (II) ions different and similar to Mg ions in terms of chemical behaviour

Differences

- Formation of complexes
- Formation of coloured compounds vs white for non-transition metals
- Transition metal able to catalyse reactions

Similar:

- Both compounds can react with sodium hydroxide to form ppt that is insoluble in excess (refer to QA section in data booklet)
- Can conduct electricity in molten state but not solid state

- Both form ppt with carbonate

Period 3 elements

Reactivity of chlorides with water:

Element	resulting pH	Explanation with equations
Na	7	NaCl + aq → Na+ + Cl *NaCl dissolves readily in water, no reaction occurs.*
Mg		
Al	pH<7	AlCl3(s) + 6 H2O(l) → [Al(H2O)6]3+(aq) + 3Cl–(aq) [Al(H2O)6]3+(aq) + H2O(l) ⇌ [Al(H2O)5(OH)]2+(aq) + H3O+(aq Since Al3+ ions have a high charge density, it has a **high polarising power** to polarise the electron cloud from the datively bonded water molecules, **weakening the O–H bond to a large extent oand producing H3O+ in water**
Si	1	SiCl4 reacts with water to form an acidic solution of pH 1 (complete dissociation of HCl). SiCl4 + 2H2O → SiO2 + 4HCL *HCl*→ H+ + Cl–
P		
S		

Behaviours of oxidising with acid/base:

Elements	Equations with explanations (if any)
Na2O (s)	It is a basic oxide and hence react with acid 3Na2O + 2H3PO4 —→ 2Na3PO4 + 3H2O
Mg2O	
Al2O3	It is an amphoteric oxide and hence can react with both acid and base 1. Rxn with acid: Al2O3 + 2H3PO4 —→ 2AlPO4 + 3H2O

Elements	Equations with explanations (if any)
	2. Reaction with base: Al2O3 + 2NaOH + 3H2O —→ 2NaAl(OH)4
SiO2	It is an acidic oxide and hence only react with base SiO2 + 2NaOH (conc) —→ Na2SiO3 + H2O Note: SiO2 only reacts with hot, concentrated NaOH due to strong covalent bonds

Properties of oxide:

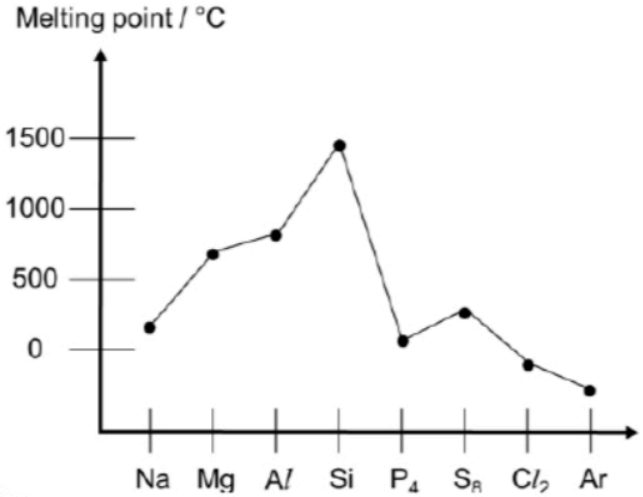
Elements	Properties	Explanations/ equations
Na2O	1. Soluble in water	1. Na2O(s) + H2O(l) → 2NaOH(aq), gives pH 14 as NaOH is a strong base that dissociates completely
MgO		
Al2O3	1. Insoluble in water 2. Amphoteric oxide	1.due to high lattice energy caused by strong ionic bonds, insufficient energy to overcome the bonds, pH 7 2. Reaction with base: Al2O3(s) + 2NaOH(aq) + 3H2O(l) → 2NaAl(OH)4(aq)
SiO2	1. Insoluble 2. Acidic oxide	
P4O10	1. Acidic oxide	1.P4O10 + 6H2O —→ 4H3PO4 (pH=2) as acid formed is a weak acid and dissociates partially to produce some H+: H3PO4(aq) H+(aq) ><+ H2PO4−(aq)
SO3	1.acidic oxide	1.SO3 + H2O —→ H2SO4 (pH < 7)

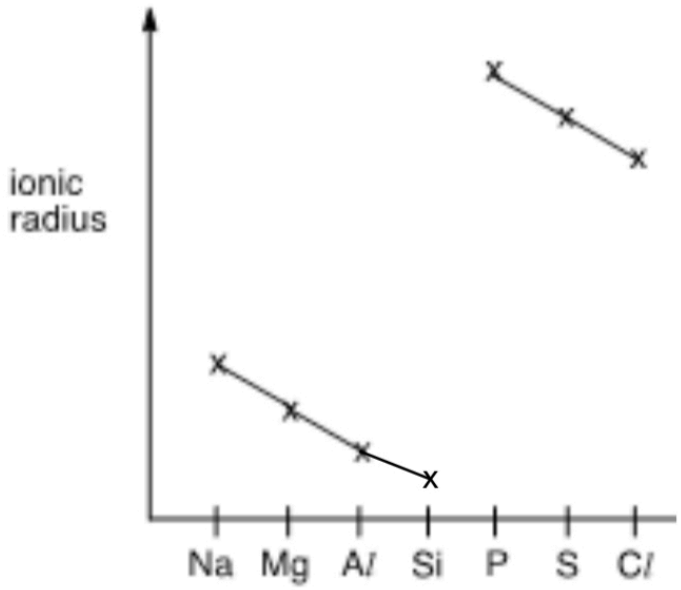
Physical properties of period 3 elements

Properties	Trend and explanation
Melting point	1.Na, Mg and A*l* exist as giant metallic lattice structure. Melting point increases from Na to Al because of ***stronger metallic bond strength*** between cations and sea of delocalized electrons as the charge of cations increases and number of delocalised electrons increase. Note: Metallic bonding strength depends on number of delocalised electrons and inversely proportional to size ionic radius(similar to lattice energy) 2.Si exist as giant covalent lattice structure. ***Much higher energy*** is required to break the strong covalent bonds between Si atoms. Hence, it has the highest melting point. 3. P4 to Ar exist as simple covalent molecules. ***Little energy*** is required to overcome the weaker instantaneous dipole – induced dipole (id-id) interaction between molecules.*As strength of id-id* $\propto$ *no. of electrons in a molecule/atom*,Melting point decrease in order: **S8 > P4 > C*l*2 > Ar (ease of melting)**
Ionic radius	Na^{+}, Mg^{2+}, Al^{3+} and Si^{4+} are isoelectronic (or have the same number of electrons) and P^{3-}, $S^{2-}$and Cl^{-} are isoelectronic, hence **shielding effect** is the **same**. Across each series, number of protons and thus **nuclear charge** increases**. [1]**Thus, **effective nuclear charge** increases, attraction of the nucleus on the **valence/outermost**electrons become stronger (*or* outermost electrons pulled closer to the nucleus), and ionicradius

Properties	Trend and explanation
	decreases.There is a sharp increase in radius from Si^{4+} to P^{3-}, as anions have one more**quantum/electron shell** than cations. **[1]**

Graph of properties:





Group 2

Trend	Explanation
Ionic radii increases down the group	Number of filled electronic shells increases and shielding effect increases causing outermost electrons are further away From nucleus. This outweighs the increase in nuclear charge as there is weaker nuclear attraction for outermost electrons and ionic radius increases
Ease of decomposition of carbonates or nitrates decreases down the group/ thermal stability increases down the group	Down the group, ionic radii increases, charge quantity remains the same, polarising power decreases as charge to size ratio of cation decreases, electron cloud of carbonate/nitrate is weakened to a lower extent, decreasing the extent of weakening of bond, more energy to break the C-O or N-O bonds in the anions, ease of

Trend	Explanation
	decomposition decreases and thermal stability increases

General periodicity

Explaining the trend of successive ionisation energy for same element

- There is an increasing trend for the successive ionization energy for the element
- With the removal of each electron, there are less electrons to be attracted by the same no. of proton
- Hence, as the number of electrons decreases, there is an increase in nuclear attraction for remaining electrons which are removed to form an increasingly positively charged ion
- the subsequent electrons require more energy to be removed, resulting in higher successive ionisation energy.
- The rapid increase in IE corresponds to a removal of the most loosely held electron from an **inner principal quantum shell**.
- The electrons from the inner principal quantum shell are closer to the nucleus and experience stronger nuclear attraction.
- Hence much more energy is required to remove the 8th electron as compared to the 7th electron.

Explaining trend of first IE down a group

- Down the group, nuclear charge increases. However, the number of filled electronic shells also increases.
- The most loosely held electron is further away from the nucleus and experiencesgreater shielding effect which outweighs the increasing nuclear

charge.

- Nuclear attraction for the most loosely held electron decreases. Less energy is required to remove the electron and 1st IE decreases.

Explaining trend for electronegativity down the group

- Down the group, valence electron further away from nucleus due to more filled principal quantum shell and increased shielding effect, outweighing increase in nuclear charge, weaker nuclear attraction for bonding pairs of electron in a covalent bond and thus decreased electronegativity

Explaining the trend in ionic radii down the group

- Number of filled quantum shells increases down the group
- Valence electrons further away from nucleus and experiences more shielding effect from inner shell electron
- Valence electrons thus experience weaker nuclear attraction that outweighs increasing nuclear charge
- Ionic radii increases down the group

Explaining trend in density down the group

- Quote Ar of elements and ionic sizes
- Ar increases down the group more than twice/thrice/by a factor of __ but ionic radii does not increase at the same rate
- For same number of ions, increase in mass is more significant than increase in volume ocupied by element, increasing density

Explaining the trend of first ionisation across a period

- **First ionisation energy generally increases** across Period due to **increasing nuclear attraction** as a result of **increasing nuclear charge** (due to larger number of protons) and **approximately constant shielding effect** (due to same number of inner principal quantum shell electrons).
- First ionisation energy of atom in group 16 is lower than that of group 15 as the **paired electron in one of the 3p orbital of sulfur experiences inter electronic**

repulsion, hence **less energy is required to remove the most loosely held electron of sulfur** than expected.