

IB Periodic Table 1

1 Describe how the elements in the Periodic Table arranged

In order of atomic number and in four groups associated with the four sub-levels (s,p,d,f)

2 State the names of the following elements

- The element in Period 3 and Group 14
- The element in Period 5 and Group 8
- The element in Period 1 and Group 18

- The element in Period 3 and Group 14 - Silicon
- The element in Period 5 and Group 8 - Ruthenium
- The element in Period 1 and Group 18 - Helium

3 State whether each of the following elements is in the s, p, d or f block

Mg	Mn	Sm	As	Ne	U	Ti	Cs
s	d	f	p	p	f	d	s

4 State the significance of the period number (n)

The number of the outer main energy level (shell) occupied by electrons/
The number of occupied shells (main energy levels) of electrons

5 For each of the following atoms state the number of the principal energy level and the number of valence electrons

Atom	principal energy level	number of valence electrons
Ca	4	2
P	3	5
Br	4	7
Si	3	4
Rb	5	1

6 State the condensed electron configuration of each of the following

Atom	condensed electron configuration
Be	[He] 2s ²
S	[Ne] 3s ² 3p ⁴
Ti	[Ar] 4s ² 3d ²
Cu	[Ar] 4s ¹ 3d ¹⁰
As	[Ar] 4s ² 3d ¹⁰ 4p ³

7 Classify each of the following elements by putting ticks in boxes

	alkali metal	halogen	noble gas	transition metal	lanthanoid	actinoid
Br		✓				
Gd					✓	
Xe			✓			
Mo				✓		
Cm						✓
Os				✓		
K	✓					

- 8 Classify each of the following elements based on their position in the Periodic Table by putting ticks in boxes

	metal	non-metal	metalloid
Nb	✓		
Se		✓	
Ge			✓
Pt	✓		
Ho	✓		
Si			✓

- 9 State and explain the trends in atomic radius down a group

Atomic radius increases down the group
Atoms have more shells of electrons

Values are given in the data booklet (Table 9)

- 10 State and explain the trends in atomic radius across a period.

Atomic radius decreases from left to right across a period
From left to right there is an increase in nuclear charge (number of protons in the nucleus)
But as all atoms have the same number of shells there is no significant change in shielding
Outer electrons pulled in more strongly

- 11 Explain why Ca^+ is smaller than Ca

Same nuclear charge (therefore same attraction for electrons)
 Ca^+ has 1 fewer electron, therefore less $e^- - e^-$ repulsion
Electrons are pulled in more closely together

- 12 Explain why Cl^- is larger than Cl

Same nuclear charge (therefore same attraction for electrons)
 Cl^- has 1 more electron, therefore more $e^- - e^-$ repulsion
Electrons cannot be pulled in as closely together

- 13 State and explain which is the larger species in each of the following pairs:

Ar and Cl^- Ca^{2+} and S^{2-} Ar and K Ar and K^+ Cl^- and Br Na^+ and Cl^-

Ar and Cl^- Cl^- larger	isoelectronic (both have 18 e^-), therefore $e^- - e^-$ repulsion very similar Ar has higher nuclear charge so electrons are pulled in more closely together
Ca^{2+} and S^{2-} S^{2-} larger	isoelectronic (both have 18 e^-), therefore $e^- - e^-$ repulsion very similar Ar has higher nuclear charge so electrons are pulled in more closely together
Ar and K K larger	K has one extra shell of electrons (4 th main energy level occupied but the highest occupied energy level for Ar is 3)
Ar and K^+ Ar larger	isoelectronic (both have 18 e^-), therefore $e^- - e^-$ repulsion very similar K has higher nuclear charge so electrons are pulled in more closely together
Cl^- and Br Br larger	Br has one extra shell of electrons (4 th main energy level occupied but the highest occupied energy level for Cl^- is 3)
Na^+ and Cl^- Cl^- larger	Cl^- has one extra shell of electrons (3 rd main energy level occupied but the highest occupied energy level for Na^+ is 2)

- 14 Write an equation for the first ionisation energy of calcium



- 15 State and explain the trend in first ionisation energy down a group

Ionisation energy decreases down the group
 Atoms get larger down the group
 Outer electron further from the nucleus, therefore less strongly attracted and easier to remove

Down a group the trends are all to do with the increase in size of the atoms
 Across a period the trends are to do with an increase in nuclear charge with no significant change in shielding.

- 16 State and explain the *general* trend in first ionisation energy across a period.

Ionisation generally increases from left to right across a period
 From left to right there is an increase in nuclear charge (number of protons in the nucleus)
 But as all atoms have the same number of shells there is no significant change in shielding
 Outer electron attracted more strongly

- 17 Explain why Mg has a higher first ionisation energy than Al



Electron removed from p subshell in Al but from s subshell in Mg

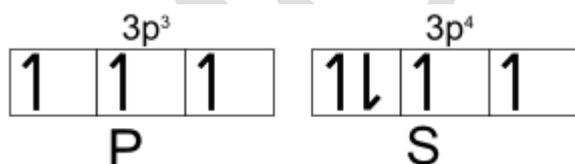
p subshell in Al higher in energy than s subshell in Mg, therefore less energy required to remove electron in A;

OR: p electron in Al slightly more shielded than s electron in Mg as the $3s^2$ electrons shield the p electron to a certain extent – more shielding, electron not as strongly attracted – easier to remove.

- 18 Explain why S has a lower first ionisation energy than P

S – 2 electrons paired in the same p orbital – greater repulsion – easier to remove an electron.

P – all p electrons are unpaired in separate p orbitals – less repulsion.



- 19 Write an equation for the first electron affinity of sulfur



- 20 State and explain the trend in first electron affinity down a group

First electron affinity becomes less exothermic down the group

Atoms get larger down the group
 Outer shell further from the nucleus,
 less energy released when an electron is brought into this shell as it is less strongly attracted by the nucleus.

Better to use this phrase rather than 'larger' or 'smaller' because the values are all negative and it can get confusing.

F has a less exothermic first electron affinity than Cl – the F atom is very small and greater e⁻e⁻ repulsion in the smaller atom outweighs the effect of the electron being brought into a shell closer to the nucleus – exceptions not mentioned on the syllabus

- 21 State and explain the general trend in first electron affinity across a period

First electron affinity generally becomes more exothermic from left to right across a period
 From left to right there is an increase in nuclear charge (number of protons in the nucleus)
 But as electrons added to the same shell there is no significant change in shielding
 Added electron attracted more strongly

Exceptions not mentioned on the syllabus – P has less exothermic EA as electron is paired up in a p orbital already containing an electron – greater repulsion

- 22 State and explain the trend in electronegativity down a group

Electronegativity decreases down the group
 Atoms get larger down the group
 Shared pair of electrons further from the nucleus, therefore less strongly attracted

- 23 State and explain the trend in electronegativity across a period

Electronegativity increases from left to right across a period
 From left to right there is an increase in nuclear charge (number of protons in the nucleus)
 But as all atoms have the same number of shells there is no significant change in shielding
 Shared pair of electrons attracted more strongly

- 24 Explain why metallic character decreases across a period

For an element to be metallic it must fairly readily form a positive ion
 Across a period ionisation energy increases therefore it becomes more difficult to form a positive ion and therefore metallic character decreases.

- 25 Explain the trend in metallic character down a group using Group 14 as an example

Down group 14 elements go from carbon, which is a non-metal to silicon and germanium to tin and lead, which are metals – therefore metallic character increases down the group.
 For an element to be metallic it must fairly readily form a positive ion
 Down a group ionisation energy decreases therefore it becomes easier to form a positive ion and therefore metallic character increases.

- 26 Classify each of the oxides below as acidic or basic or amphoteric.

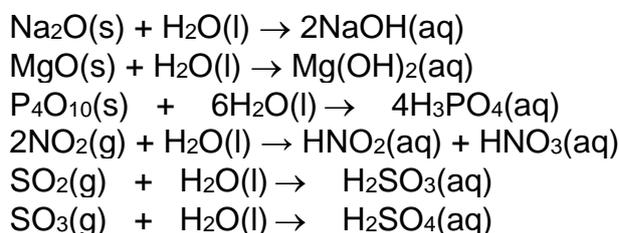
Na₂O MgO Al₂O₃ SiO₂ P₄O₆ P₄O₁₀ SO₂ SO₃ Cl₂O Cl₂O₇

basic		amphoteric	acidic						
Na ₂ O	MgO	Al ₂ O ₃	SiO ₂	P ₄ O ₆	P ₄ O ₁₀	SO ₂	SO ₃	Cl ₂ O	Cl ₂ O ₇

In general, metallic oxides are basic and non-metallic oxides are acidic

- 27 Write equations for the reactions of the following with water:

Na₂O MgO P₄O₁₀ NO₂ SO₂ SO₃



28 State whether each of the solutions formed in 27 are acidic, alkaline or neutral.

$\text{Na}_2\text{O}(\text{s}) + \text{H}_2\text{O}(\text{l}) \rightarrow 2\text{NaOH}(\text{aq})$	alkaline
$\text{MgO}(\text{s}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{Mg}(\text{OH})_2(\text{aq})$	alkaline
$\text{P}_4\text{O}_{10}(\text{s}) + 6\text{H}_2\text{O}(\text{l}) \rightarrow 4\text{H}_3\text{PO}_4(\text{aq})$	acidic
$2\text{NO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{HNO}_2(\text{aq}) + \text{HNO}_3(\text{aq})$	acidic
$\text{SO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2\text{SO}_3(\text{aq})$	acidic
$\text{SO}_3(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2\text{SO}_4(\text{aq})$	acidic

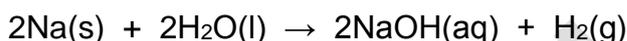
29 Explain two reasons why the elements Li and Na are placed in the same group in the Periodic Table

They have the same number of electrons in the outer shell (1)
Therefore have the same chemical properties – react in the same way

30 Explain the decrease in melting point of the alkali metals from Li to Cs.

The atoms get bigger down the group
Weaker attraction between the positive ion and the delocalised electrons
(the delocalised electrons are further from the nucleus of the positive ion)

31 Write an equation for the reactions of sodium with water.



32 State two observations for the reaction between sodium and water

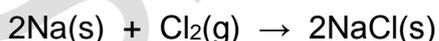
- melts into a ball,
- fizzes
- moves around on the surface of the water
- disappears/gets smaller

For lithium: you need the idea that the reaction is less vigorous, e.g. does not fizz as rapidly or melt into a ball
For potassium the reaction is more vigorous – bursts into flames.

33 Explain why potassium is more reactive than sodium

Reactions involve the formation of a positive ion (M^+)
Potassium atom is bigger
Outer electron is further from the nucleus, therefore less strongly attracted by the nucleus
Outer electron lost more easily/ionisation energy lower.

34 Write an equation for the reaction between sodium and chlorine.



35 Explain why elements in the same group have similar chemical properties

Reactions depend on the number of electrons in the outer shell of an atom
Atoms in the same group have the same number of electrons in the outer shell

36 Explain the increase in melting point and boiling point down the halogen group

Covalent molecular with London forces between molecules
Down the group the number of electrons increases/relative molecular mass increases
London forces get stronger, therefore more energy required to overcome them.

37 State and explain the colour changes that are observed when the following are mixed:

- chlorine solution and potassium bromide solution
- bromine solution and potassium iodide solution
- bromine solution and potassium chloride solution

- chlorine solution and potassium bromide solution

reaction mixture turns yellow/orange

chlorine displaces bromine

chlorine is a stronger oxidising agent than bromine, therefore oxidises bromide ions to bromine

- bromine solution and potassium iodide solution

reaction mixture turns orange/brown

bromine displaces iodine

bromine is a stronger oxidising agent than iodine, therefore oxidises iodide ions to iodine

- bromine solution and potassium chloride solution

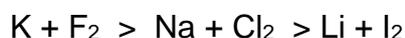
No reaction

chlorine is a stronger oxidising agent than bromine, therefore bromine cannot displace chlorine/oxidise chloride ions

38 Write an ionic equation for the reaction that occurs when chlorine solution is added to potassium bromide solution



39 Predict which of the following reactions will be most vigorous:



The order of reactivity of the alkali metals is $\text{K} > \text{Na} > \text{Li}$

The order of reactivity of the halogens is $\text{F}_2 > \text{Cl}_2 > \text{I}_2$

40 State the characteristic properties of transition elements

- They can exhibit more than one oxidation state in compounds/complexes.
- They form complex ions.
- They usually form coloured compounds/complexes.
- The elements and their compounds/complexes can act as catalysts in many reactions.

41 Write the electron configurations for the following:

Ti Fe Fe^{2+} Fe^{3+} Ni^{2+} Cr Cr^{3+} Cu Cu^+ Zn^{2+}

Ti	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$
Fe	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$
Fe^{2+}	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^6$
Fe^{3+}	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^5$
Ni^{2+}	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^8$
Cr	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$
Cr^{3+}	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^3$
Cu	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$
Cu^+	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10}$
Zn^{2+}	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10}$

42 Explain why Zn is not considered as a transition element

Transition element forms at least one stable oxidation state (other than 0) with a partially-filled d subshell

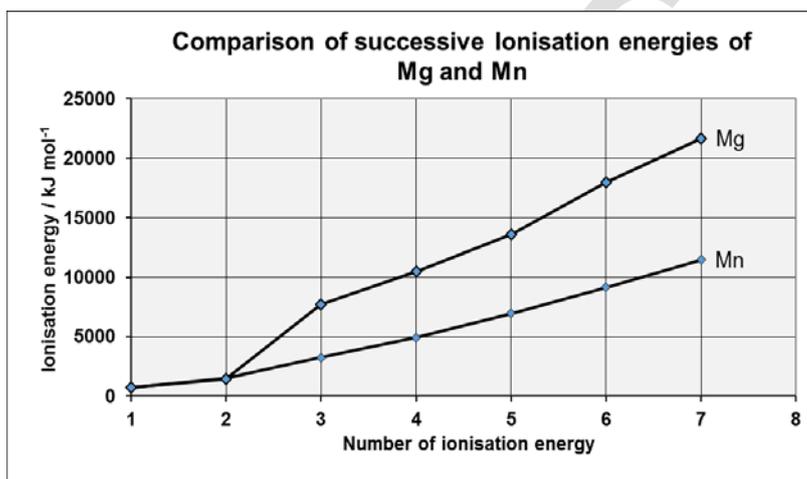
Zn only forms the Zn^{2+} ion – this has a full 3d subshell

43 Explain why all transition metals exhibit the +2 oxidation state

They can lose the $4s^2$ electrons

Cr/Cu on have $4s^1$ but the 4s and 3d are very close in energy so can also lose a second electron

44 Use the graph below to explain why magnesium only exhibits oxidation state +2 but manganese exhibits a range of oxidation states.



Mg has 2 electrons in the outer shell which can be lost relatively easily but to remove a third electron requires a lot more energy because the third electron is removed from a shell closer to the nucleus – this does not occur.

Mn – the 4s and 3d subshells are close in energy, therefore there are no big jumps in ionisation energies – it changes gradually. Therefore a range of ions can be formed.

45 Deduce whether each of the following ions is diamagnetic or paramagnetic

Fe^{2+} Fe^{3+} Ni^{2+} Cr^{3+} Cu^+ Zn^{2+}

paramagnetism is caused by unpaired electrons – paramagnetic substances are attracted by a magnetic field.

	3d subshell					
Fe^{2+}	$\uparrow\downarrow$	\uparrow	\uparrow	\uparrow	\uparrow	unpaired electrons, therefore paramagnetic
Fe^{3+}	\uparrow	\uparrow	\uparrow	\uparrow	\uparrow	unpaired electrons, therefore paramagnetic
Ni^{2+}	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$	\uparrow	\uparrow	unpaired electrons, therefore paramagnetic
Cr^{3+}	\uparrow	\uparrow	\uparrow			unpaired electrons, therefore paramagnetic
Cu^+	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$	all electrons paired, therefore diamagnetic
Zn^{2+}	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$	all electrons paired, therefore diamagnetic

The paramagnetic effect is much larger, so any unpaired electrons means the compound is paramagnetic

Diamagnetism is caused by paired electrons – diamagnetic substances are repelled slightly by a magnetic field.

- 46 Explain what a *ligand* is and state whether it is a Lewis acid or Lewis base.

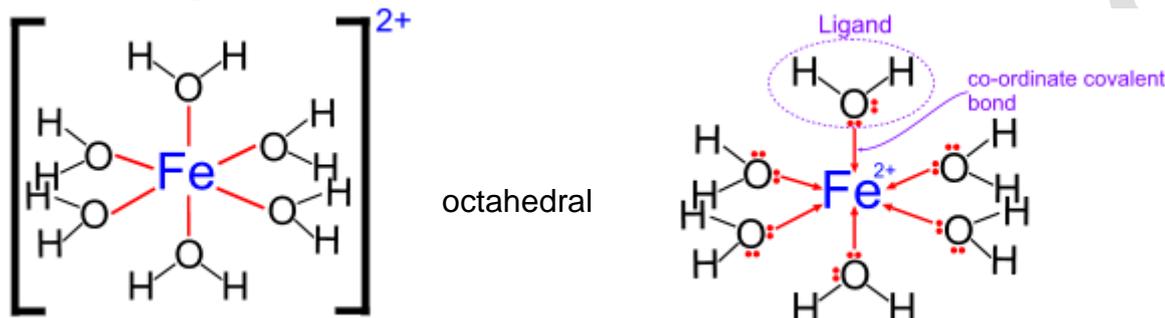
Negative ions or neutral molecules which use lone pairs of electrons to bond to a transition metal ion to form a complex ion. Co-ordinate covalent bonds are formed between the ligand and the transition metal ion.

A ligand is a Lewis base – it can donate a lone pair of electrons

- 47 Explain how a ligand bonds to a transition metal ion

Each ligand uses a lone pair to form a co-ordinate covalent bond

- 48 Draw diagrams of the complex ion $[\text{Fe}(\text{H}_2\text{O})_6]^{2+}$



- 49 Deduce the total charge on each of the following ions:

$[\text{Fe}(\text{CN})_6]^{?}$ contains iron(III)

$[\text{CuCl}_4]^{?}$ contains copper(II)

$[\text{Ag}(\text{NH}_3)_2]^{?}$ contains silver(I)

$[\text{Ti}(\text{OH})(\text{H}_2\text{O})_5]^{?}$ contains titanium(III)

$[\text{Fe}(\text{CN})_6]^{3-}$ - CN^- ligand has a 1- charge: $3+ + 6 \times 1- = 3-$

$[\text{CuCl}_4]^{2-}$ - Cl^- ligand has a 1- charge: $2+ + 4 \times 1- = 2-$

$[\text{Ag}(\text{NH}_3)_2]^+$ - NH_3 ligand has 0 charge: $1+ + 0 = 1+$

$[\text{Ti}(\text{OH})(\text{H}_2\text{O})_5]^{2+}$ - H_2O ligand has 0 charge, OH^- ligand has 1- charge: $3+ + 1- + 0 = 2+$

- 50 Deduce the oxidation state of the transition metal in each of the following:

$[\text{TiCl}_2(\text{NH}_3)_4]\text{Cl}$ $[\text{Co}(\text{NH}_3)_5\text{Cl}]\text{Cl}_2$ K_3NiF_6 $\text{K}_2[\text{Ni}(\text{CO})_2(\text{CN})_2]$

$[\text{TiCl}_2(\text{NH}_3)_4]^+ \text{Cl}^-$ NH_3 0 charge, Cl^- , therefore Ti +3 oxidation state

$[\text{Co}(\text{NH}_3)_5\text{Cl}]^{2+} (\text{Cl}^-)_2$ NH_3 0 charge, Cl^- , therefore Co +3 oxidation state

$(\text{K}^+)_3 [\text{NiF}_6]^{3-}$ F^- , therefore Ni +3 oxidation state

$(\text{K}^+)_2 [\text{Ni}(\text{CO})_2(\text{CN})_2]^{2-}$ CO 0 charge, CN^- , therefore Ni 0 oxidation state

- 51 Explain why some transition metal ions are coloured.

- d orbitals split into 2 groups at different energies
- certain frequencies of **visible** light are absorbed
- to promote an electron from the lower to the higher set of d orbitals
- the complementary colour is reflected/transmitted
- a partially-filled d subshell is required

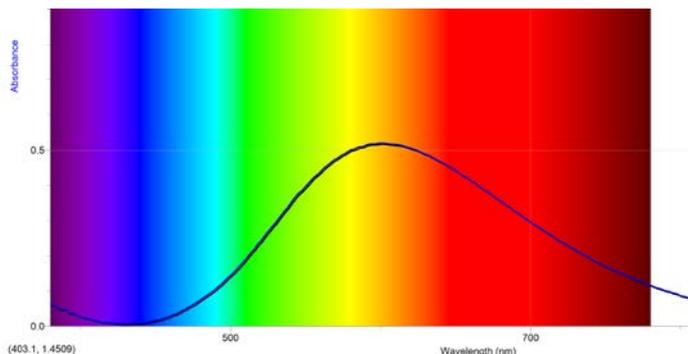
- 52 Explain why complexes of Ti^{4+} and Zn^{2+} are colourless.

Ti^{4+} has no d electrons, therefore there are no electrons to be promoted

Zn^{2+} has $3d^{10}$, so the d subshell is full – no space in the highest set of orbitals to promote an electron to.

53 In each case below explain what colour the transition metal complex ion is:

- It absorbs orange light
- It absorbs green light
- It absorbs light in the range 424 nm to 491 nm
- The complex ion with the spectrum shown



- It absorbs orange light – using the colour wheel (in data booklet) – the complementary colour to orange is blue, therefore the complex ion is blue
 - It absorbs green light– using the colour wheel (in data booklet) – the complementary colour to green is red, therefore the complex ion is red
 - It absorbs blue light– using the colour wheel (in data booklet) – the complementary colour to blue is orange, therefore the complex ion is orange
 - The complex ion with the spectrum shown – absorbs orange, yellow, green, red. Some red/blue/violet transmitted, therefore likely to be purple/violet
- 54 Explain why $[\text{Fe}(\text{H}_2\text{O})_6]^{2+}$ and $[\text{Co}(\text{H}_2\text{O})_6]^{3+}$ are different colours

Fe^{2+} and Co^{3+} ions present, both $3d^6$ – same number of electrons
 Co^{3+} has higher nuclear charge – therefore ligands pulled in more closely so greater splitting of d orbitals.
 Therefore different frequencies of visible light absorbed.

55 Explain why $[\text{Co}(\text{H}_2\text{O})_6]^{2+}$ and $[\text{Co}(\text{H}_2\text{O})_6]^{3+}$ are different colours

Co^{2+} and Co^{3+} have different numbers of electrons and therefore different amounts of electron-electron repulsion.
 $3+$ ions pulls the ligands in more closely, therefore greater splitting of d orbitals
 Therefore different frequencies of visible light absorbed.

56 Explain why $[\text{Co}(\text{H}_2\text{O})_6]^{2+}$ and $[\text{Ni}(\text{H}_2\text{O})_6]^{2+}$ are different colours

Co^{2+} and Ni^{2+} have different numbers of electrons and therefore different amounts of electron-electron repulsion.
 d orbital splitting is different in each
 Therefore different frequencies of visible light absorbed.

57 Explain why $[\text{Co}(\text{H}_2\text{O})_6]^{2+}$ and $[\text{Co}(\text{NH}_3)_6]^{2+}$ are different colours

NH_3 causes greater splitting of d orbitals (from spectrochemical series in data booklet)
 Bigger difference in energy between lower and higher set of d orbitals in $[\text{Co}(\text{NH}_3)_6]^{2+}$
 Therefore shorter wavelength/higher frequency energy absorbed

58 Explain whether $[\text{Cu}(\text{H}_2\text{O})_6]^{2+}$ or $[\text{Cu}(\text{NH}_3)_4(\text{H}_2\text{O})_2]^{2+}$ absorbs light of a longer wavelength

NH_3 causes greater splitting of d orbitals (from spectrochemical series in data booklet)
 Bigger difference in energy between lower and higher set of d orbitals in $[\text{Cu}(\text{NH}_3)_4(\text{H}_2\text{O})_2]^{2+}$
 Therefore shorter wavelength/higher frequency energy absorbed in $[\text{Cu}(\text{NH}_3)_4(\text{H}_2\text{O})_2]^{2+}$