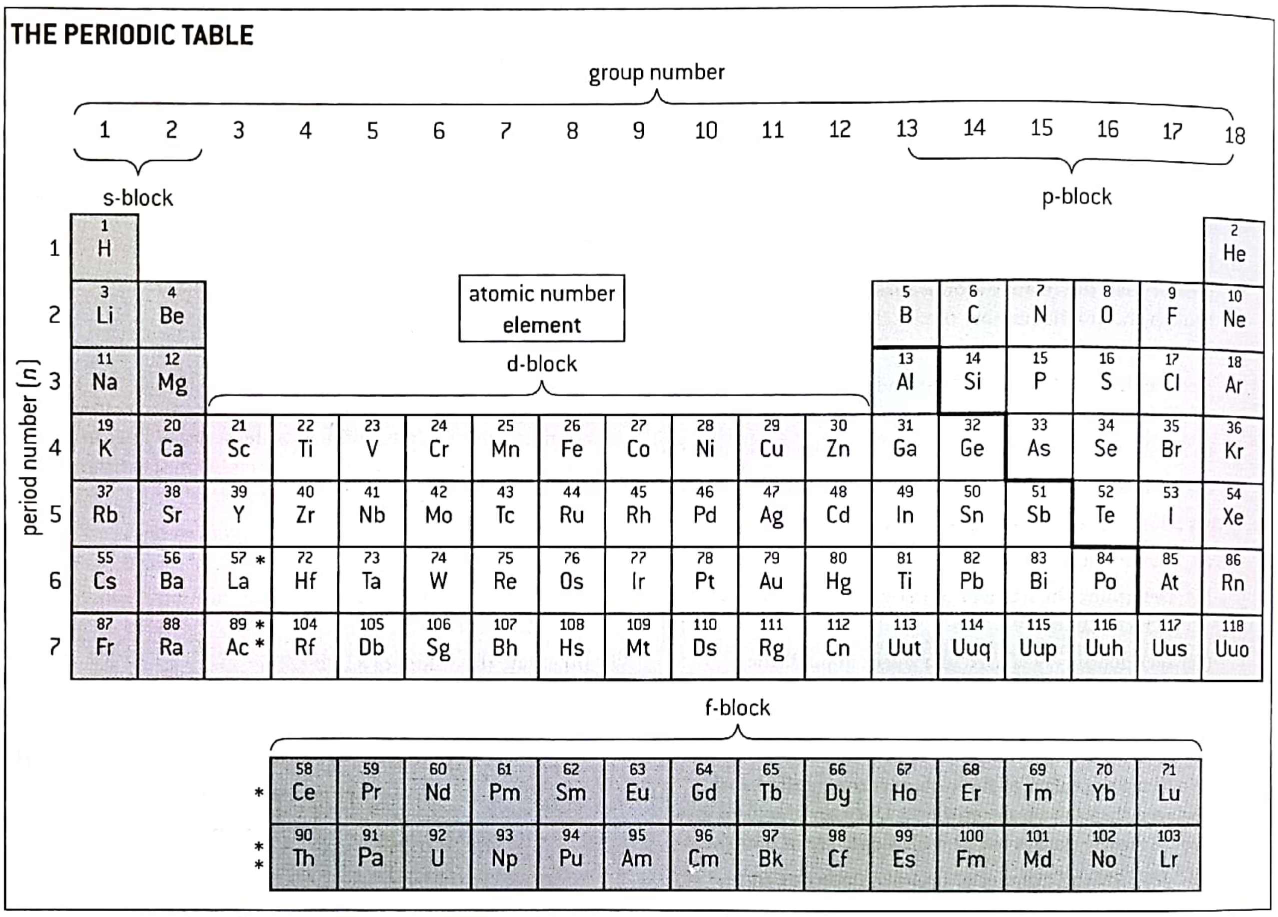
The periodic table

THE PERIODIC TABLE



period number (n)	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba	57 * La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Ti	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra	89 * Ac *	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Uut	114 Uuq	115 Uup	116 Uuh	117 Uus	118 Uuo

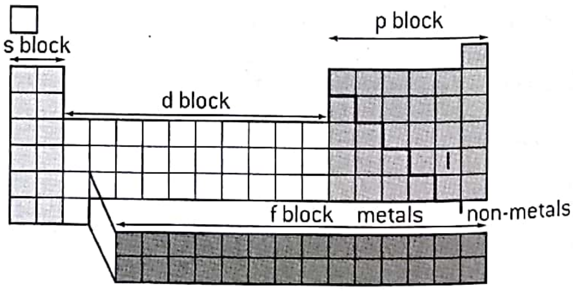
f-block

*	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
**	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

FEATURES OF THE PERIODIC TABLE

- In the periodic table elements are placed in order of increasing atomic number.
- Elements are arranged into four blocks associated with the four sub-levels – s, p, d and f.
- The period number (*n*) is the outer energy level that is occupied by electrons.
- Elements in the same vertical group contain the same number of electrons in the outer energy level.
- The number of the principal energy level and the number of the valence electrons in an atom can be deduced from its position on the periodic table.
- The periodic table shows the positions of metals, non-metals and metalloids.
- Certain groups have their own name. For example, group 1 is known as the alkali metals, group 17 as the halogens and group 18 as the noble gases.
- The d-block elements (groups 3 to 12) are known as the transition metals and the f-block elements form two distinct groups – the lanthanoids (from La to Lu) and the actinoids (from Ac to Lr).
- Metals are on the left and in the centre of the table and non-metals are on the right (distinguished by the thick line). Metalloids, such as boron, silicon and germanium have properties intermediate between those of a metal and a non-metal.

ELECTRON CONFIGURATION AND THE PERIODIC TABLE



An element's position in the periodic table is related to its valence electrons so the electronic configuration of any element can be deduced from the table, e.g. iodine ($Z = 53$) is a p block element. It is in group 17 so its configuration will contain ns^2np^5. If one takes H and He as being the first period then iodine is in the fifth period so $n = 5$. The full configuration for iodine will therefore be:

$1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^5$ or $[Kr]\ 5s^24d^{10}5p^5$

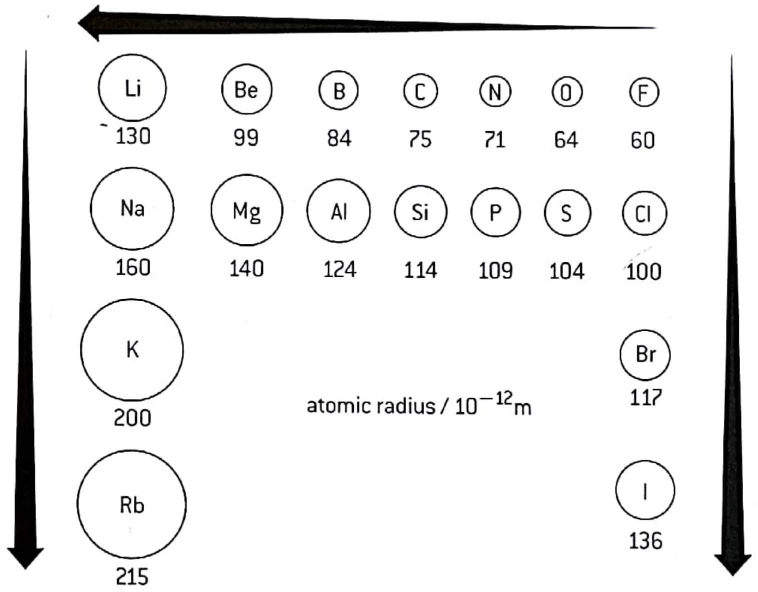
Periodic trends (1)

PERIODICITY

Elements in the same group tend to have similar chemical and physical properties. There is a change in chemical and physical properties across a period. The repeating pattern of physical and chemical properties shown by the different periods is known as **periodicity**.

These periodic trends can clearly be seen in atomic radii, ionic radii, ionization energies, electronegativities, electron affinities and melting points.

ATOMIC RADIUS



The atomic radius is the distance from the nucleus to the outermost electron. Since the position of the outermost electron can never be known precisely, the atomic radius is usually defined as half the distance between the nuclei of two bonded atoms of the same element.

As a group is descended, the outermost electron is in a higher energy level, which is further from the nucleus, so the radius increases.

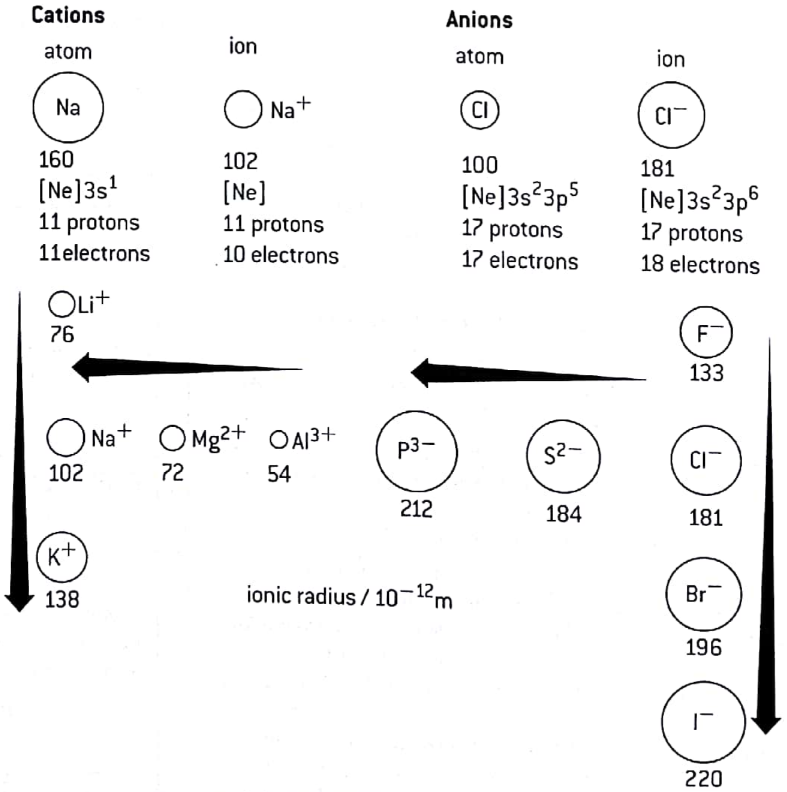
Across a period electrons are being added to the same energy level, but the number of protons in the nucleus increases. This attracts the energy level closer to the nucleus and the atomic radius decreases across a period.

IONIC RADIUS

It is important to distinguish between positive ions (**cations**) and negative ions (**anions**). Both cations and anions increase in size down a group as the outer level gets further from the nucleus.

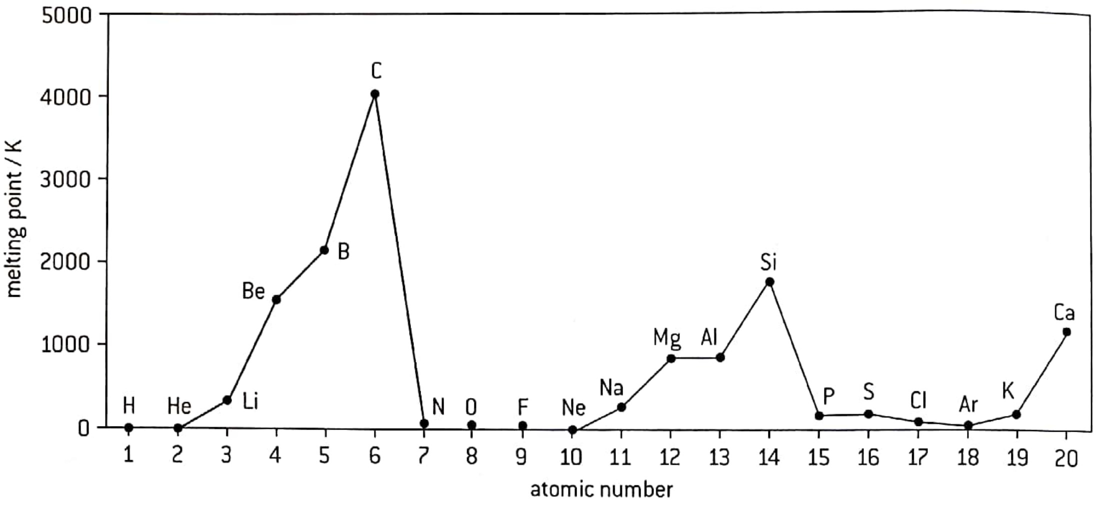
Cations contain fewer electrons than protons so the electrostatic attraction between the nucleus and the outermost electron is greater and the ion is smaller than the parent atom. It is also smaller because the number of electron shells has decreased by one. Across the period the ions contain the same number of electrons (**isoelectronic**), but an increasing number of protons, so the ionic radius decreases.

Anions contain more electrons than protons so are larger than the parent atom. Across a period the size decreases because the number of electrons remains the same but the number of protons increases.



Periodic trends (2)

MELTING POINTS



Melting points depend both on the structure of the element and on the type of attractive forces holding the atoms together. Using period 3 as an example:

- At the left of the period elements exhibit metallic bonding (Na, Mg, Al), which increases in strength as the number of valence electrons increases.
- Silicon, the metalloid, in the middle of the period has a macromolecular covalent structure with very strong bonds resulting in a very high melting point.
- Elements in groups 15, 16, and 17 (P_4, S_8, and Cl_2) show simple molecular structures with weak intermolecular forces of attraction between the molecules.
- The noble gases (Ar) exist as **monatomic molecules** (single atoms) with extremely weak forces of attraction between the atoms.

Within groups there are also clear trends:

- In group 1 the melting point decreases down the group as the atoms become larger and the strength of the metallic bond decreases.

	Li	Na	K	Rb	Cs
M. pt / K	454	371	336	312	302

- In group 17 the intermolecular attractive forces between the diatomic molecules increase down the group so the melting points increase.

	F_2	Cl_2	Br_2	I_2
M. pt / K	53	172	266	387

FIRST IONIZATION ENERGY

The definition of first ionization energy and a graph showing the values for the first 20 elements has already been given on page 13. The values decrease down each group as the outer electron is further from the nucleus and therefore less energy is required to remove it, e.g. for the group 1 elements, Li, Na and K.

Element:	Li	Na	K
Electron configuration	$1s^22s^1$	$[Ne]3s^1$	$[Ar]4s^1$
First ionization energy (kJ mol^{-1})	520	496	419

Generally the values increase across a period. This is because the extra electrons are filling the same energy level and the extra protons in the nucleus attract this energy level closer making it harder to remove an electron.

The values do not increase regularly across a period because new sub-levels are being filled. The p sub-level is higher in energy than the s sub-level. This explains why the value for B ($1s^22s^22p^1$) is slightly lower than the value for Be (Is^22s^2) and the value for Al ($[Ne]3s^23p^1$) is slightly lower than Mg ($[Ne]3s^2$). There is also a drop in value between N ($1s^22s^22p^3$) and O ($1s^22s^22p^4$) and between P ($[Ne]3s^23p^3$) and S ($[Ne]3s^23p^4$). This is because when electrons pair up in an orbital there is increased repulsion so the paired electron is easier to remove compared with when the three electrons are all unpaired, one each in the three separate p orbitals.

ELECTRONEGATIVITY

Electronegativity is a relative measure of the attraction that an atom has for a shared pair of electrons when it is covalently bonded to another atom. As the size of the atom decreases the electronegativity increases, so the value increases across a period and decreases down a group. The three most electronegative elements are F, N, and O.

ELECTRON AFFINITY

The electron affinity is the energy change when an electron is added to an isolated atom in the gaseous state, i.e.

$$X(g) + e^- \rightarrow X^-(g)$$

Atoms 'want' an extra electron so electron affinity values are negative for the addition of the first electron. However, when oxygen forms the O^{2-} ion the overall process is endothermic.

	$O(g) + e^- \rightarrow O^-(g)$	$\Delta H^\ominus = -141$ kJ mol^{-1}
	$O^-(g) + e^- \rightarrow O^{2-}(g)$	$\Delta H^\ominus = +753$ kJ mol^{-1}
overall	$O(g) + 2e^- \rightarrow O^{2-}(g)$	$\Delta H^\ominus = +612$ kJ mol^{-1}

Periodic trends (3)

CHEMICAL PROPERTIES OF ELEMENTS IN THE SAME GROUP

Group 1 – the alkali metals

Li
Na
K
Rb
Cs

↓ increasing reactivity

Lithium, sodium, and potassium all contain one electron in their outer shell. They are all reactive metals and are stored under liquid paraffin to prevent them reacting with air. They react by losing their outer electron to form the metal ion. Because they can readily lose an electron they are good reducing agents. The reactivity increases down the group as the outer electron is in successively higher energy levels and less energy is required to remove it.

They are called alkali metals because they all react with water to form an alkali solution of the metal hydroxide and hydrogen gas. Lithium floats and reacts quietly, sodium melts into a ball which darts around on the surface, and the heat generated from the reaction with potassium ignites the hydrogen.

$$2Li(s) + 2H_2O(l) \rightarrow 2Li^+(aq) + 2OH^-(aq) + H_2(g)$$

$$2Na(s) + 2H_2O(l) \rightarrow 2Na^+(aq) + 2OH^-(aq) + H_2(g)$$

$$2K(s) + 2H_2O(l) \rightarrow 2K^+(aq) + 2OH^-(aq) + H_2(g)$$

They all also react readily with chlorine, bromine and iodine to form ionic salts, e.g.

$$2Na(s) + Cl_2(g) \rightarrow 2Na^+Cl^-(s)$$

$$2K(s) + Br_2(l) \rightarrow 2K^+Br^-(s)$$

$$2Li(s) + I_2(g) \rightarrow 2Li^+I^-(s)$$

Group 17 – the halogens

F_2
Cl_2
Br_2
I_2

↑ increasing reactivity

The halogens react by gaining one more electron to form halide ions. They are good oxidizing agents. The reactivity decreases down the group as the outer shell is increasingly at higher energy levels and further from the nucleus. This, together with the fact that there are more electrons between the nucleus and the outer shell, decreases the attraction for an extra electron.

Chlorine is a stronger oxidizing agent than bromine, so can remove the electron from bromide ions in solution to form chloride ions and bromine. Similarly both chlorine and bromine can oxidize iodide ions to form iodine.

$$Cl_2(aq) + 2Br^-(aq) \rightarrow 2Cl^-(aq) + Br_2(aq)$$

$$Cl_2(aq) + 2I^-(aq) \rightarrow 2Cl^-(aq) + I_2(aq)$$

$$Br_2(aq) + 2I^-(aq) \rightarrow 2Br^-(aq) + I_2(aq)$$

Test for halide ions

The presence of halide ions in solution can be detected by adding silver nitrate solution. The silver ions react with the halide ions to form a precipitate of the silver halide. The silver halides can be distinguished by their colour. These silver halides react with light to form silver metal. This is the basis of old-fashioned film photography.

$$\begin{array}{c} Ag^+(aq) + X^-(aq) \rightarrow AgX(s) \\ \text{light}\ \big\downarrow \\ Ag(s) + \tfrac{1}{2}X_2 \end{array} \qquad \text{where } X = Cl, Br, \text{ or } I$$

AgCl white
AgBr cream
AgI yellow

CHANGE FROM METALLIC TO NON-METALLIC NATURE OF THE ELEMENTS ACROSS PERIOD 3

Metals tend to be shiny and are good conductors of heat and electricity. Sodium, magnesium and aluminium all conduct electricity well. Silicon is a semiconductor and is called a **metalloid** as it possesses some of the properties of a metal and some of a non-metal. Phosphorus, sulfur, chlorine and argon are non-metals and do not conduct electricity. Metals can also be distinguished from non-metals by their chemical properties. Metal oxides tend to be basic, whereas non-metal oxides tend to be acidic.

Sodium oxide and magnesium oxide are both basic and react with water to form hydroxides,

e.g. $Na_2O(s) + H_2O(l) \rightarrow 2NaOH(aq)$ $\quad MgO(s) + H_2O(l) \rightarrow Mg(OH)_2$

Aluminium is a metal but its oxide is amphoteric, that is, it can be either basic or acidic depending on whether it is reacting with an acid or a base.

The remaining elements in period 3 have acidic oxides. For example, sulfur trioxide reacts with water to form sulfuric acid, phosphorus pentoxide reacts with water to form phosphoric(V) acid and nitrogen(IV)oxide reacts with water to form nitric acid.

$$SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq) \qquad P_4O_{10}(s) + 6H_2O(l) \rightarrow 4H_3PO_4(aq) \qquad 3NO_2(g) + H_2O(l) \rightarrow 2HNO_3(aq) + NO(g)$$

HL The transition metals

THE FIRST ROW TRANSITION ELEMENTS

Element	(Sc)	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	(Zn)
Electron configuration [Ar]	$4s^23d^1$	$4s^23d^2$	$4s^23d^3$	$4s^13d^5$	$4s^23d^5$	$4s^23d^6$	$4s^23d^7$	$4s^23d^8$	$4s^13d^{10}$	$4s^23d^{10}$

A transition element is defined as an element that possesses an incomplete d sub-level in one or more of its oxidation states. Scandium is not a typical transition metal as its common ion Sc^{3+} has no d electrons. Zinc is not a transition metal as it contains a full d sub-level in all its oxidation states. (Note: for Cr and Cu it is more energetically favourable to half-fill and completely fill the d sub-level respectively so they contain only one 4s electron).

Transition elements have variable oxidation states, form complex ions with ligands, have coloured compounds, and display catalytic and magnetic properties.

VARIABLE OXIDATION STATES

The 3d and 4s sub-levels are very similar in energy. When transition metals lose electrons they lose the 4s electrons first. All transition metals can show an oxidation state of +2. Some of the transition metals can form the +3 or +4 ion (e.g. Fe^{3+}, Mn^{4+}) as the ionization energies are such that up to two d electrons can also be lost. The M^{4+} ion is rare and in the higher oxidation states the element is usually found not as the free metal ion but either covalently bonded or as the oxyanion, such as MnO_4^-. Some common examples of variable oxidation states in addition to +2 are:

Cr(+3)	$CrCl_3$	chromium(III) chloride
Cr(+6)	$Cr_2O_7^{2-}$	dichromate(VI) ion
Mn(+4)	MnO_2	manganese(IV) oxide
Mn(+7)	MnO_4^-	manganate(VII) ion
Fe(+3)	Fe_2O_3	iron(III) oxide
Cu(+1)	Cu_2O	copper(I) oxide

A full list of the common oxidation states can be found in Section 14 of the IB data booklet.

CATALYTIC BEHAVIOUR

Many transition elements and their compounds are very efficient catalysts, that is, they increase the rate of chemical reactions. This helps to make industrial processes, such as the production of ammonia and sulfuric acid, more efficient and economic. Platinum and palladium are used in catalytic converters fitted to cars. In the body, iron is found in haem and cobalt is found in vitamin B_{12}. Other common examples include:

Iron in the Haber process	$3H_2(g) + N_2 \overset{Fe(s)}{\rightleftharpoons} 2NH_3(g)$
Vanadium(V) oxide in the Contact process	$2SO_2(g) + O_2(g) \overset{V_2O_5(s)}{\rightleftharpoons} 2SO_3(g)$
Nickel in hydrogenation reactions	$C_2H_4(g) + H_2(g) \xrightarrow{Ni(s)} C_2H_6(g)$
Manganese(IV) oxide with hydrogen peroxide	$2H_2O_2(aq) \xrightarrow{MnO_2(s)} 2H_2O(l) + O_2(g)$

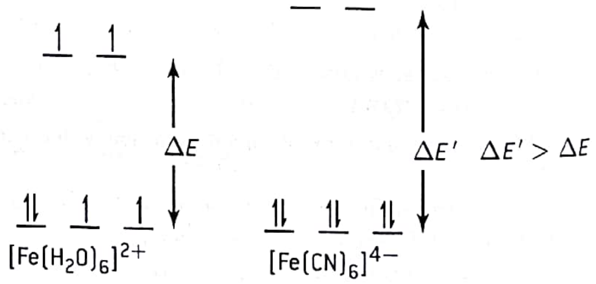
MAGNETIC PROPERTIES

Transition metals and their complexes that contain unpaired electrons can exhibit magnetism. Iron metal and some other metals (e.g. nickel and cobalt) show ferromagnetism. This is a permanent type of magnetism. In this type of magnetism the unpaired electrons align parallel to each other in domains irrespective of whether an external magnetic or electric field is present. This property of iron has been utilized for centuries to make compasses, which align with the Earth's magnetic field to point north.

Many complexes of transition metals contain unpaired electrons. Unlike paired electrons, where the spins cancel each other out, the spinning unpaired electrons create a small magnetic field and will line up in an applied electric or magnetic field to make the transition metal complex weakly magnetic when the field is applied, i.e. they reinforce the external magnetic field. This type of magnetism is known as paramagnetism. The more unpaired electrons there are in the complex the more paramagnetic the complex will be.

When all the electrons in a transition metal complex are paired up the complex is said to be diamagnetic.

Iron(II) complex ions ([Ar]$3d^6$) can be paramagnetic or diamagnetic. The five d orbitals are split by the ligands according to the spectrochemical series (see page 22). If the ligands are low in the spectrochemical series (e.g. H_2O) and split the d orbitals by only a small amount, the electrons can occupy all the d orbitals giving four units of paramagnetism as there will be four unpaired electrons. This is known as weak ligand field splitting. Ligands high in the spectrochemical series (e.g. CN^-) will cause a larger splitting. Only the lower d orbitals will be occupied and the complex will be diamagnetic as there are no unpaired electrons.



Weak field splitting
4 unpaired electrons
paramagnetic

Strong ligand field splitting
no unpaired electrons
diamagnetic

HL Transition metal complex ions

FORMATION OF COMPLEX IONS

Because of their small size d-block ions act as Lewis acids and attract species that are rich in electrons. Such species are known as **ligands**. Ligands are neutral molecules or anions which contain a non-bonding pair of electrons. These electron pairs can form co-ordinate covalent bonds with the metal ion to form **complex ions**.

A common ligand is water and most (but not all) transition metal ions exist as hexahydrated complex ions in aqueous solution, e.g. $[Fe(H_2O)_6]^{3+}$. Ligands can be replaced by other ligands. A typical example is the addition of ammonia to an aqueous solution of copper(II) sulfate to give the deep blue colour of the tetraaminecopper(II) ion. Similarly if concentrated hydrochloric acid is added to a solution of $Cu^{2+}(aq)$ the yellow tetrachlorocopper(II) anion is formed.

Note: in this ion the overall charge on the ion is −2 as the four ligands each have a charge of −1.

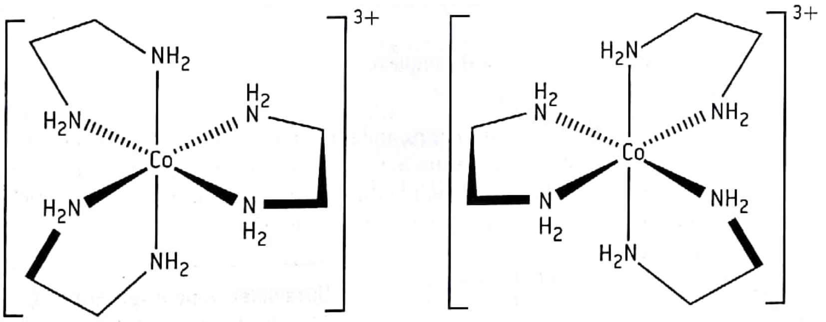
$$[CuCl_4]^{2-} \underset{H_2O}{\overset{Cl^-}{\rightleftharpoons}} [Cu(H_2O)_4]^{2+} \underset{H_2O}{\overset{NH_3}{\rightleftharpoons}} [Cu(NH_3)_4]^{2+}$$

The number of lone pairs bonded to the metal ion is known as the **coordination number**. Compounds with a coordination number of six are octahedral in shape, those with a coordination number of four are tetrahedral or square planar, whereas those with a coordination number of two are usually linear.

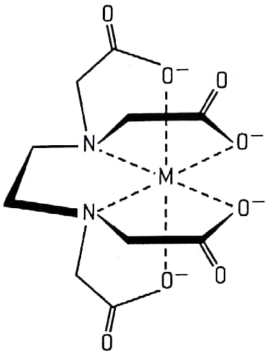
Coordination number	6	4	2
Examples	$[Fe(CN)_6]^{3-}$ $[Fe(OH)_3(H_2O)_3]$	$[CuCl_4]^{2-}$ $[Cu(NH_3)_4]^{2+}$	$[Ag(NH_3)_2]^+$

POLYDENTATE LIGANDS

Ligands such as water and cyanide ions are known as monodentate ligands as they utilize just one non-bonding pair to form a coordinate covalent bond to the metal ion. Some ligands contain more than one non-bonding pair and can form two or more coordinate bonds to the metal ion. Three common examples are ethylenediamine, $H_2NCH_2CH_2NH_2$, oxalate ions $(COO^-)_2$, both of which can use two non-bonding pairs to form bidentate ligands, and EDTA (EthyleneDiamineTetraAcetic acid) or its ion, $EDTA^{4-}$, which can act as hexadentate ligands.



The Co^{3+} ion can form two different stereoisomers (mirror images) with the three separate ethylenediamine molecules acting as bidentate ligands. The coordination number is six as although each metal ion is surrounded by only three ligands the metal ion forms a total of six coordinate bonds with the ligands.



$EDTA^{4-}$ acting as a hexadentate ligand with a transition metal, M to form a complex such as $[Cu(EDTA)]^{2-}$. Note that the coordination number is still six even though only one ligand surrounds the metal ion.

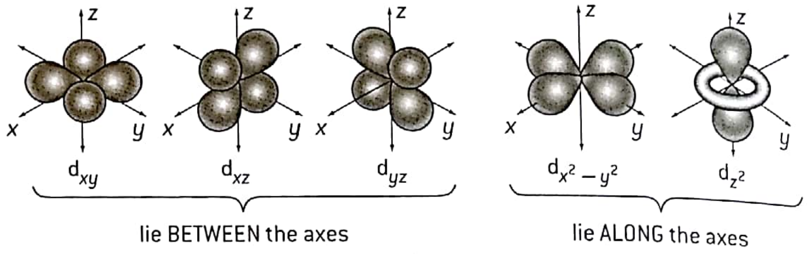
FACTORS AFFECTING THE COLOUR OF TRANSITION METAL COMPLEXES

Transition metals are defined as elements having an incomplete d sub-level in one or more of their oxidation states. Compounds of Sc^{3+} which have no d electrons and of Cu^+ and Zn^{2+} which both have complete d sub-shells are colourless. This strongly suggests that the colour of transition metal complexes is related to an incomplete d level. The actual colour is determined by four different factors.

1. The nature of the transition element. For example $Mn^{2+}(aq)$ and $Fe^{3+}(aq)$ both have the configuration $[Ar]3d^5$. $Mn^{2+}(aq)$ is pink whereas $Fe^{3+}(aq)$ is yellow.
2. The oxidation state. $Fe^{2+}(aq)$ is green whereas $Fe^{3+}(aq)$ is yellow.
3. The identity of the ligand. $[Cu(H_2O)_6]^{2+}$ (sometimes shown as $[Cu(H_2O)_4]^{2+}$), is blue, $[Cu(NH_3)_4(H_2O)_2]^{2+}$ (sometimes shown as $[Cu(NH_3)_4]^{2+}$), is blue/violet whereas $[CuCl_4]^{2-}$ is yellow (green in aqueous solution).
4. The stereochemistry of the complex. The colour is also affected by the shape of the molecule or ion. In the above example $[Cu(H_2O)_6]^{2+}$ is octahedral whereas $[CuCl_4]^{2-}$ is tetrahedral. However for the IB only octahedral complexes in aqueous solution will be considered.

HL Colour of transition metal complexes

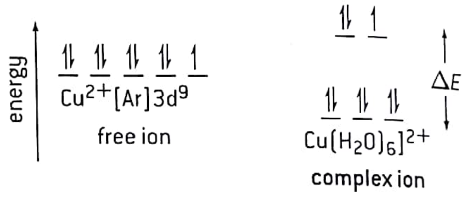
SPLITTING OF THE d ORBITALS



In the free ion the five d orbitals are degenerate. That is they are all of equal energy.

Note that three of the orbitals (d_{xy}, d_{yz}, and d_{zx}) lie *between* the axes whereas the other two ($d_{x^2-y^2}$ and d_{z^2}) lie *along* the axes. Ligands act as Lewis bases and donate a non-bonding pair of electrons to form a co-ordinate bond. As the ligands approach the metal along the axes to form an octahedral complex the non-bonding pairs of electrons on the ligands will repel the $d_{x^2-y^2}$ and d_{z^2} orbitals causing the five d orbitals to split, three to lower energy and two to higher energy. The difference in energy between the two levels corresponds to the wavelength of visible light.

When white light falls on the aqueous solution of the complex the colour corresponding to ΔE is absorbed as an electron is promoted and the transmitted light will be the complementary colour. For example, $[Cu(H_2O)_6]^{2+}$ absorbs red light so the compound appears blue. The amount that the d orbitals are split will determine the exact colour. Changing the transition metal changes the number of protons in the nucleus which will affect the levels. Similarly changing the oxidation state will affect the splitting as the number of electrons in the level is different. Different ligands will also cause different amounts of splitting depending on their electron density.

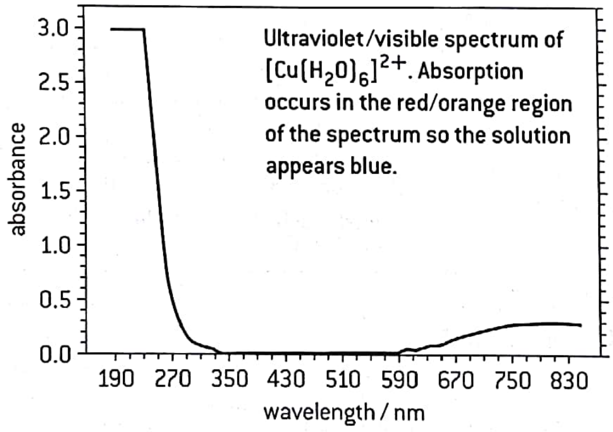


THE SPECTROCHEMICAL SERIES

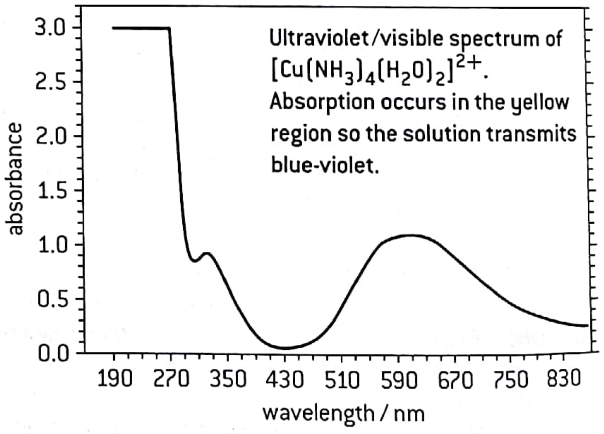
Ligands can be arranged in order of their ability to split the d orbitals in octahedral complexes.

$$I^- < Br^- < S^{2-} < Cl^- < F^- < OH^- < H_2O < SCN^- < NH_3 < CN^- < CO$$

This order is known as the spectrochemical series. Iodide ions cause the smallest splitting and the carbonyl group, CO, the largest splitting. The energy of light absorbed increases when ammonia is substituted for water in Cu^{2+} complexes as the splitting increases, i.e. in going from $[Cu(H_2O)_6]^{2+}$ to $[Cu(NH_3)_4(H_2O)_2]^{2+}$. This means that the wavelength of the light absorbed decreases and this is observed in the colour of the transmitted light which changes from blue to a blue-violet (purple) colour.

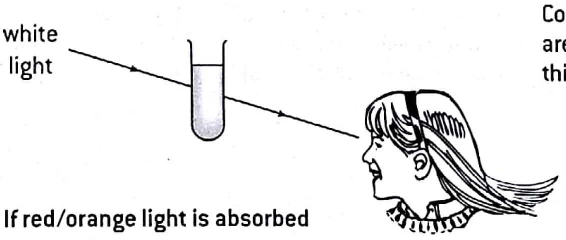


Ultraviolet/visible spectrum of $[Cu(H_2O)_6]^{2+}$. Absorption occurs in the red/orange region of the spectrum so the solution appears blue.



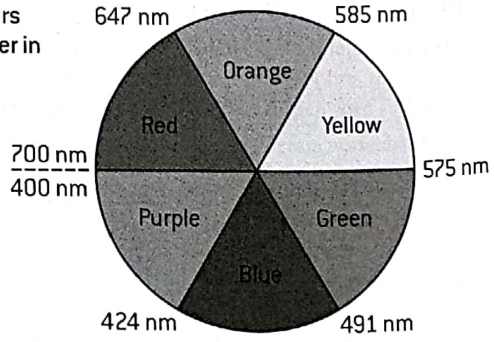
Ultraviolet/visible spectrum of $[Cu(NH_3)_4(H_2O)_2]^{2+}$. Absorption occurs in the yellow region so the solution transmits blue-violet.

COMPLEMENTARY COLOURS



If red/orange light is absorbed the solution appears blue-green as that is the complementary colour which is transmitted.

Complementary colours are opposite each other in this 'colour wheel'.



MULTIPLE CHOICE QUESTIONS – PERIODICITY

1. Where in the periodic table would the element with the electronic structure $1s^22s^22p^63s^23p^64s^23d^{10}4p^2$ be located?

   A. s block
   B. p block
   C. d block
   D. lanthanoids

2. Which property decreases with increasing atomic number in group 17 (the halogens)?

   A. Melting point
   B. Electronegativity
   C. Atomic radius
   D. Ionic radius of the negative ion

3. Which property increases with increasing atomic number for both the halogens and the alkali metals?

   A. Reactivity with water
   B. Electronegativity
   C. Electron affinity
   D. Atomic radius

4. Which one of the following series is arranged in order of increasing size?

   A. P^{3-}, S^{2-}, Cl^-
   B. Cl, Ar, K
   C. Na^+, Mg^{2+}, Al^{3+}
   D. H^+, H, H^-

5. Which element has the lowest electronegativity value?

   A. potassium
   B. fluorine
   C. iodine
   D. hydrogen

6. Which statements about the periodic table are correct?

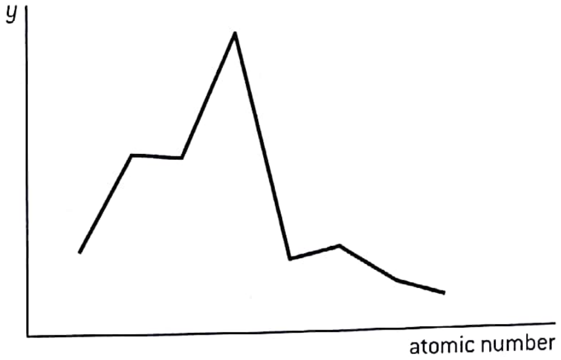
   I. The elements Mg, Ca and Sr have similar chemical properties.

   II. Elements in the same period have the same number of main energy levels.

   III. The oxides of Na, Mg and P are basic.

   A. I and II only
   B. I and III only
   C. II and III only
   D. I, II and III

7. The *x*-axis of the graph below represents the atomic number of the elements in period 3.



   Which variable could represent the *y*-axis?

   A. Melting point
   B. Electronegativity
   C. Ionic radius
   D. Atomic radius

8. Which is the correct definition for the electron affinity of element X?

   A. $X(s) + e^- \rightarrow X^-(s)$
   B. $X(s) \rightarrow X^+(s) + e^-$
   C. $X(g) + e^- \rightarrow X^-(g)$
   D. $X(g) \rightarrow X^+(g) + e^-$

9. Which oxide will react with water to give a solution with a pH greater than 7?

   A. NO_2
   B. P_4O_{10}
   C. MgO
   D. SiO_2

10. Which reaction does **not** occur readily?

    A. $Cl_2(aq) + 2Br^-(aq) \rightarrow Br_2(aq) + 2Cl^-(aq)$
    B. $2Na(s) + H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$
    C. $Na_2O(s) + H_2O(l) \rightarrow 2NaOH(aq)$
    D. $I_2(aq) + 2Br^-(aq) \rightarrow Br_2(aq) + 2I^-(aq)$

HL

11. Which complex ion is colourless in aqueous solution?

    A. $[Mn(H_2O)_6]^{2+}$
    B. $[Zn(H_2O)_4]^{2+}$
    C. $[Cu(NH_3)_4]^{2+}$
    D. $[CuCl_4]^{2-}$

12. A certain element has the electron configuration $1s^22s^22p^63s^23p^64s^23d^3$. What oxidation state(s) would this element most likely show?

    A. +2 only
    B. +3 only
    C. +2 and +5 only
    D. +2, +3, +4 and +5

13. What is the overall charge on the complex ion formed by Fe(III) and six CN^- ligands?

    A. +3
    B. −6
    C. −3
    D. +6

14. Which statements about transition metal complexes are correct?

    I. The colour of the complex is due to light being emitted when an electron falls from a higher split d sub-level to a lower split d sub-level.

    II. The difference in energy between the split d sub-levels depends on the nature of the surrounding ligands.

    III. The colour of the complex is influenced by the oxidation state of the metal.

    A. I and II only
    B. I and III only
    C. II and III only
    D. I, II and III

15. What is the electron configuration of the Mn^{2+} ion?

    A. $1s^22s^22p^63s^23p^63d^5$
    B. $1s^22s^22p^63s^23p^64s^23d^5$
    C. $1s^22s^22p^63s^23p^64s^23d^3$
    D. $1s^22s^22p^63s^23p^64s^13d^4$

16. Which statement is true for all polydentate ligands?

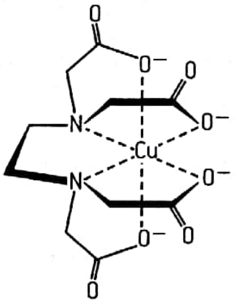
    A. They contain more than one pair of non-bonding electrons.
    B. They can only form one coordinate bond with a transition metal ion.
    C. They do not affect the size of the splitting of the d sub-level.
    D. They can only form complexes with transition metal ions.

SHORT ANSWER QUESTIONS – PERIODICITY

1. Carbon, silicon and tin belong to the same group in the periodic table.
   a) Distinguish between the terms *group* and *period* in terms of electron configuration. [2]
   b) State in which block of the periodic table (s, p, d or f) these three elements are located. [1]
   c) Explain why the first ionization energy of silicon is lower than that of carbon. [2]
   d) Describe a simple experiment to show the change in non-metallic to metallic properties as the atomic number of the elements in the group increases. [3]
2. a) The maximum number of electrons in each energy level is determined by the expression $2n^2$ where n is the number of the level. Explain why the third level is sometimes said to contain eight electrons and sometimes said to contain eighteen electrons. [2]
   b) In terms of electron configuration describe the essential difference between the lanthanoids and the actinoids. [2]
3. a) Describe how the acid–base nature of the oxides changes across period 3 (Na→Cl). [3]
   b) State the equation for the reaction of water with
      (i) sodium oxide, Na_2O [1]
      (ii) phosphorus(V) oxide, P_4O_{10} [1]
      (iii) sulfur(VI) oxide, SO_3 [1]
   c) Explain why oxides of nitrogen and sulfur cause damage to many types of buildings. Include relevant equations in your explanation. [4]
4. a) State the electron configuration for the P^{3-} ion. [1]
   b) Explain why the ionic radius of the S^{2-} ion is smaller than that of the P^{3-} ion. [2]
   c) Potassium and bromine are in the same period. Explain why the atomic radius of bromine is considerably smaller than the atomic radius of potassium. [2]
   d) Suggest a reason why no value is given for the atomic radius of neon in the IB data booklet. [1]
5. The periodic table shows the relationship between electron configuration and the properties of elements and is a valuable tool for making predictions in chemistry.
   a) (i) Identify the property used to arrange elements in the periodic table. [1]
      (ii) Outline two reasons why electronegativity increases across period 3 in the periodic table and one reason why noble gases are not assigned electronegativity values. [3]
   b) (i) Define the term first ionization energy of an atom. [2]
      (ii) Explain the general increasing trend in the first ionization energies of the period 2 elements, Li to Ne. [2]
      (iii) Explain why the first ionization energy of boron is lower than that of beryllium. [2]
   c) Explain why sodium conducts electricity but phosphorus does not. [2]
6. In many cities around the world, public transport vehicles use diesel, a liquid hydrocarbon fuel, which often contains sulfur impurities and undergoes incomplete combustion. All public transport vehicles in New Delhi, India, have been converted to use compressed natural gas (CNG) as fuel. Suggest two ways in which this improves air quality, giving a reason for your answer. [3]

HL

7. Transition elements, such as iron, have many characteristic properties and uses.
   a) State and explain in terms of **acid–base properties** the type of reaction that occurs between Fe^{2+} ions and water to form $[Fe(H_2O)_6]^{2+}$. [2]
   b) Explain why the colour of $[Fe(H_2O)_6]^{2+}$ is different to that of $[Fe(H_2O)_6]^{3+}$. [2]
   c) Explain why iron metal is magnetic and why some of its complexes such as $[Fe(H_2O)_6]^{2+}$ are paramagnetic whereas others such as $[Fe(CN)_6]^{4-}$ are diamagnetic. [4]
   d) The Haber process to form ammonia is exothermic. Explore the economic significance of using an iron catalyst. [3]
8. The diagram shows the $[Cu(EDTA)]^{4-}$ ion.



   a) Deduce the formula mass of this complex ion. [2]
   b) State the coordination number and the shape of the ion. [2]
   c) Explain how the EDTA ion is able to bond to the Cu^{2+} ion. [3]
9. Ligands can be listed in terms of their ability to split the d orbitals. This is known as the spectrochemical series.
   a) Suggest a reason why different ligands cause the d orbitals to be split by different amounts. [2]
   b) Ammonia molecules are higher in the spectrochemical series (cause greater splitting) than water molecules. Explain why the light blue colour of $[Cu(H_2O)_6]^{2+}$ changes to a more purple colour when excess ammonia is added to copper(II) sulfate solution to form the $[Cu(NH_3)_4(H_2O)_2]^{2+}$ ion. [3]
   c) Explain why aqueous solutions of scandium(III) compounds are not coloured. [2]