Chemistry

Periodic Table
Notes

Focus

Periodic Trends
(a) describe the Periodic Table as an arrangement of the elements in the order of increasing proton (atomic) number
(b) describe how the position of an element in the Periodic Table is related to proton number and electronic configuration
(c) describe the relationship between number of outer (valence) electrons and the ionic charge of an ion for the first twenty elements
(d) explain the similarities between the elements in the same group of the Periodic Table in terms of their electronic configuration
(e) describe the change from metallic to non-metallic character from left to right across a period of the Periodic Table
(f) describe the relationship between number of outer (valence) electrons and metallic/nonmetallic character
(g) predict the properties of elements in Group 1 and Group 17 using the Periodic Table

Group Properties
(a) describe lithium, sodium and potassium in Group 1 (the alkali metals) as a collection of relatively soft, low density metals showing a trend in melting point and in their reaction with water
(b) describe chlorine, bromine and iodine in Group 17 (the halogens) as a collection of diatomic non-metals showing a trend in colour, state and their displacement reactions with solutions of other halide ions
(c) describe the elements in Group 18 (the noble gases) as a collection of monoatomic elements that are chemically unreactive and hence important in providing an inert environment, e.g. argon and neon in light bulbs; helium in balloons; argon in the manufacture of steel
(d) describe the lack of reactivity of the noble gases in terms of their electronic configurations

Transition Elements
(a) describe typical transition elements as metals having high melting point, high density, variable oxidation state and forming coloured compounds
(b) state that the elements and/or their compounds are often able to act as catalysts.

Past Periodic Table

1. After numerous elements were discovered by the mid 1800s, chemists were looking hard for a way of organizing the elements based on what they found out regarding their physical and chemical properties.

- In 1866, the Periodic Table was proposed to be arranged in accordance to the relative atomic mass of element. The idea was put forth by a British chemist named John Newlands. The elements were arranged into columns according to the properties of the elements known at that time. Non-metal gases were placed in the 1st column, while metals that are highly reactive were placed in the 2nd column, resembling part of the modern Periodic Table that we know of.

H	Li	Be	B	C	N	O
F	Na	Mg	Al	Si	P	S
Cl	K	Ca	Cr	Ti	Mn	Fe

- Subsequently, a Russian scientist named Dmitri Mendeleev developed John Newlands' idea, which was not supported then. The Periodic Table was still being arranged in accordance to the relative atomic mass of element. He introduced the term '**period**' to describe elements that are arranged in a row and '**group**' to describe the elements arranged in a column.

	Group 1	Group 2	Group 3	Group 4	Group 5	Group 6	Group 7	Group 8
Row 1	H							
Row 2	Li	Be	B	C	N	O	F	
Row 3	Na	Mg	Al	Si	P	S	Cl	
Row 4	K	Ca	-	-	-	-	-	Ti, V, Cr, Mn, Fe, Co, Ni
Row 5	Cu	Zn	-	-	As	Se	Br	

- The developments he made towards the modern Periodic Table was as follows:
    - He introduced the concept of period and group,
    - He left gaps, allowing chemists to focus their effort in the search of these elements that are yet to be discovered, and;
    - He separated the metals from the non-metals clearly: metals were arranged on the left and non-metals on the right.

- The proposal was still far different from the modern Periodic Table in the following aspects:
    - The elements were arranged according to increasing relative atomic mass.
    - Hydrogen, a non-metal, was in group 1.
    - There were no noble gases (because they were not discovered yet)
    - The transition metal block consisting of Sc, Ti, V, Cr, Mn, Fe, Co, Ni, Cu and Zn was missing (because they were not discovered yet).

<u>**Intro - Period and Group**</u>

2. The modern Periodic Table is arranged in order of ***increasing number of protons/atomic number***, and the elements are arranged according to their physical and chemical properties.

   <u>Take Note:</u> The Periodic Table is a list of elements arranged in order of increasing proton (atomic) numbers.

- **Group number of an element indicates the number of valence electrons or outermost electrons of an atom for elements in Group 1 and 2, For elements in Group 13 to 18 (except Helium), the number of valence electrons is the group number minus 10.** A group is a vertical column of elements that runs from top to bottom. There are a total of eighteen groups, numbered from 1 to 18. The last group consisting of noble gases is labelled as Group 18.

➔ The electronic arrangement of elements in the Periodic Table (Groups):

Element	Proton number	Electron arrangement	Number of valence electron	Group
Lithium, Li	3	2,1	1	1
Sodium, Na	11	2,8,1	1	1
Beryllium, Be	4	2,2	2	2
Magnesium, Mg	12	2,8,2	2	2
Aluminium	13	2,8,3	13-10 = 3	13
Fluorine	9	2,7	17-10 = 7	17

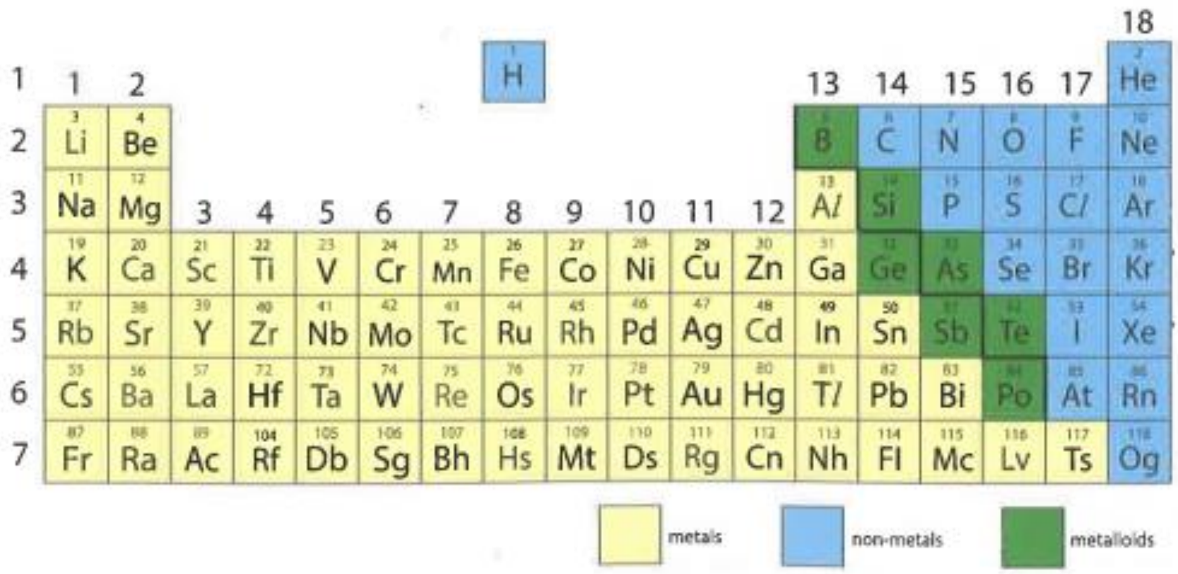
- **Period number of an element indicates the number of electron shells in an atom.** A period is a horizontal row of elements that runs from left to right. There are a total of seven periods, numbered from 1 to 7.

➔ The electronic arrangement of elements in The Periodic Table (Periods):

Element	Proton number	Electron arrangement	Number of Shells filled with electrons	Period
Boron, B	5	2,3	2	2
Magnesium, Mg	12	2,8,2	3	3
Potassium, K	19	2,8,8,1	4	4

- There is an imaginary line in the Periodic Table, which separates the metals from the non-metals. This **Metal – Non Metal Line** is shown in bold in the figure below. To the right of the line are the elements: B-Si-As-Te-At. Elements that are found next to this line are called metalloids. Metalloids have the properties of both metals and non-metals

- All Group 1, 2 and transition metal block elements are metals, while all Group 17 & 18 elements are non-metals. Group 13 to 16 elements consist of both metals and non-metals.



3. **Elements from the Same Group** display similar chemical properties since they have the same number of valence electrons. During chemical reactions they can lose/gain/share the same number of electrons to attain a noble gas electronic configuration. Elements do share some similar physical properties.

4. **Elements from the Same Period** have the same number of electron shell/s. Elements do not share the same physical and chemical properties.

Down a Group

Going down a group, there is an increase in metallic properties and a decrease in non-metallic properties.
This is because the size of the atom increases down the group.
Hence the valence electrons of the element will be further away from the attractive force of the nucleus.

Tendency to lose valence electron/s (metallic characteristic) increases; while tendency to gain electron/s (non-metallic characteristic) decreases.

Across the Periodic Table

Period 3

5. Across Period 2 and 3, the following are the changes in structures, bonding, physical and chemical properties of the elements.

Period 3	**Na**	**Mg**	**Al**	**Si**	**P**	**S**	**Cl**	**Ar**
Group	1	2	13	14	15	16	17	18
Valence electrons	1	2	3	4	5	6	7	8
Electron arrangement	2,8,1	2,8,2	2,8,3	2,8,4	2,8,5	2,8,6	2,8,7	2,8,8
Size of atom	big --→ small							
Element is a	metal			metalloid	non-metal			
Structure	• giant metallic • regular lattice arrangement of atoms			• giant molecular (covalent) or macromolecular • extensive covalent network	• simple molecular (covalent)			
Bonding	strong forces of attraction between the positive cations and the sea of negatively charged delocalised electrons			strong covalent bonds between Si atoms, forming a 3-dimensional giant network	weak intermolecular forces between covalent molecules			
Melting point	high			high	low			
Electrical conductivity	good electrical conductor			fair electrical conductor	poor electrical conductor			
Nature of oxides	Basic	Amphoteric		Acidic				none

- Across the period, the elements become less metallic / more non-metallic. This is because due to increase in number of protons, the valence electrons are attracted more strongly to the nucleus. Hence tendency to lose electron/s (metallic characteristic) decreases and tendency to gain electron/s (non-metallic characteristic) increases.
- Across the period, the size of the atom decreases.
  ➔ The proton number increases by one. The positive charge of the nucleus increases. The force of attraction of nucleus on the electrons increases. Electrons are pulled closer to the nucleus. Hence, the size of the atom decreases.

Down the Periodic Table
Group 1 – Alkali Metals

6. Physical Properties of Alkali Metals:

- 1 electron in outermost electron shell.
  → Lose 1 electron to become a cation.
  → Highly unstable and high in the reactivity series which explains its vigorous reaction with water to form an alkaline solution and hydrogen gas
- Soft silvery solids that can be cut with a knife.
- Have low melting and boiling points when compared to normal metals.
- Have low densities [Li, Na and K are less dense than water (float on water), whereas from K to Rb, they are higher in density due to a sharp increase in mass number] when compared to other metals.
- Conduct electricity since they are metals with free-moving delocalised electrons.
- All alkali salts are soluble in water.
- Going down the group, the density increases, the melting point decreases and element becomes more reactive.

7. Properties of Alkali Metals down the group:

Question 1 What explains the increase in density, increase in reactivity and decrease in melting point going down the group?

Density
Down group,
- **Relative Molecular Mass increases**
- **Density = mass/volume**
- **Density increases**

Reactivity
Down the group,
- **Atomic radius of Group I element increases**
- **Easier for elements to lose electron**
- **Therefore element is more reactive**

Melting Point
Melting point = strength of metallic bond
Down group,
- **Larger Atomic Radius**
- **Weaker bond**
- **Same number of delocalised electrons**

8. Chemical Properties of Alkali Metals:

(a) Alkali metals are highly reactive as they have an outermost electron which is lost when they take part in chemical reactions. They become cations.

(b) **Down the group, alkali metals become more reactive** since atomic size increases, valence electron is further from the nucleus and the attractive forces between nucleus and valence electron weakens and it becomes easier for the metal atom to lose its outermost electron to form a cation. Potassium (K) is more reactive than sodium (Na), but less reactive than rubidium (Rb).

(c) Alkali metals are **strong reducing agents** as they chemically react to form cations by losing an electron and undergoing oxidation. For example: $Li \rightarrow Li^{+} + e^{-}$

Reactions that show alkali metals behave as reducing agents:

➔ Alkali metals burn readily with oxygen in air to produce white metal oxide. (Hence, alkali metals are normally submerged in oil to prevent it from coming into direct contact with oxygen in air). The white metal oxide is dissolves in water to form an alkaline (metal hydroxide) solution.

$4\ M + O_2 \rightarrow 2\ M_2O$	
$4\ Li(s) + O_2(g) \rightarrow 2\ Li_2O(s)$	Lithium burns with a red flame.
$4\ Na(s) + O_2(g) \rightarrow 2\ Na_2O(s)$	Sodium burns quickly with a yellow flame.
$4\ K(s) + O_2(g) \rightarrow 2\ K_2O(s)$	Potassium burns violently with a lilac flame.

➔ Alkali metals react with halogens (e.g. chlorine gas) to produce white metal halides (metal chlorides) which are ionic. The halide compounds have similar chemical formulae.

$2\ M + Cl_2 \rightarrow 2\ MCl$	
$2\ Li(s) + Cl_2(g) \rightarrow 2\ LiCl(s)$	Lithium burns with a bright flame.
$2\ Na(s) + Cl_2(g) \rightarrow 2\ NaCl(s)$	Sodium burns quickly with a bright flame.
$2\ K(s) + Cl_2(g) \rightarrow 2\ KCl(s)$	Potassium burns violently with a bright flame.

➔ All alkali metals when dissolved in water form alkaline (metal hydroxide) solutions. In this chemical reaction, hydrogen gas is produced. The respective reactions between the 3 alkali metals: Na, K and Rb and water have different observations since their densities and reactivities are different.

Question 2 Pen down the equations for the following reactions:

- Water and Li
- Water and Na
- Water and Rb
- General Equation for the reaction between Water and A (where A is an alkali metal)

$2Li + 2H_2O \rightarrow 2LiOH + H_2$
$2Na + 2H_2O \rightarrow 2NaOH + H_2$
$2Rb + 2H_2O \rightarrow 2RbOH + H_2$

$2A + 2H_2O \rightarrow 2AOH + H_2$

💣 **Question 3** Describe the chemical reactions between each of the Group 1 metals: Na, K and Rb and water. State the observations.

- **Rate of Reaction**
- **Down the Group, reactivity increases**
- **Rb reacts explosively**
- **K react very violently**
- **Na react violently**

- **Na, K – density less than water**
- **K to Rb – significant jump in Ar → Rb more dense than water**
- **Rb sinks in water**
- **K skids on water**
- **Na skids on water**

- **Rb – no flame**
- **K – burns with lilac flame**
- **Na – occasional yellow flame observed**

- **Rb – sinks in water → no burning of Rb**
- **Rb – no flame is seen**
- **K, Na produces great amount of heat that cause K and Na to burn**
- **K being more reactive enables K to burn, while Na being less reactive only burn occasionally**

💣 **Question 4** Going down Group 2, will the density of the elements increase? Will the melting point decrease? Will the metals become more reactive? Are Group I elements more reactive than Group II elements?

Yes. Group 2 and Group 1 have the same trends.

Yes. Group 1 elements are more reactive than Group 2 generally as they only need to lose one valence electron to achieve a noble gas structure while Group 2 elements need to lose two.

Group 17 – Halogens

9. **Group 17 (Halogens)** is a group of five chemically active non-metallic elements—fluorine (F), chlorine (Cl), bromine (Br), iodine (I), and astatine (At). Halogens are the most reactive non-metals in the Periodic Table. The name halogen refers to the property that each of the halogens has in forming a salt similar to common salt (sodium chloride) when reacted with sodium. Each member of the group has a valency of 1 and combines with metals to form halides, as well as reacts with metals and non-metals to form complex ions.

10. Properties of Halogens:

    - 7 electrons in outermost electron shell.
      → Gains 1 electron to become an anion.
      → Shares 1 electron to attain electronic structure of a noble gas. This is why the elements exist as diatomic molecules.

- The states of Group 17 elements at r.t.p., and their colours:

Element at r.t.p.	Colour at r.t.p.	Colour in an aqueous solution
F_2 (gas)	Pale yellow	pale yellowish/colourless
Cl_2 (gas)	Yellow-green	Light yellow
Br_2 (liquid)	red-brown	Orange/ red-brown
I_2 (solid)	Purple-black	Brown
At_2 (solid)	Black	--

Colour of Group 17 elements darkens down the Group.

- Most silver and lead halides [halides are formed from halogens after they have achieved a noble gas electronic configuration, e.g. chloride (Cl^-), bromide (Br^-), iodide (I^-)] are not soluble in water.

- Do not conduct electricity since the molecules are covalent compounds and there are no free-moving electrons.

- Have low melting and boiling points because of weak intermolecular forces of attraction between the covalent molecules that can be easily overcame with little energy.

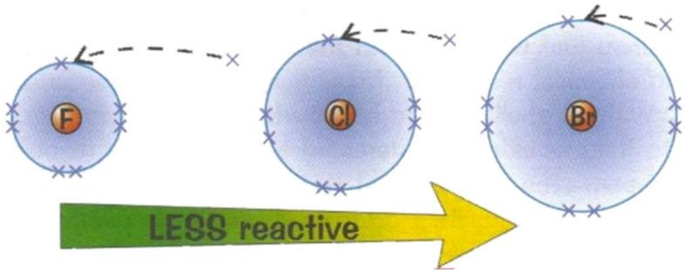
- Going down the group, density increases, melting and boiling points increase (from gas to solid) and elements become less reactive.

11. Properties of Halogens down the Group:

Question 5 What explains the increase in density, decrease in reactivity and increase in melting and boiling points?

<u>Density</u>
Down the group,
- **Relative Molecular Mass increases**
- **Density = mass/volume**
- **Density increases**

<u>Reactivity</u>
Down the group,
- **Atomic radius of Group 17 elements increases**
- **More difficult for elements to take in or share electron**
- **Therefore, element is less reactive**



<u>Melting and boiling points</u>
Melting point / boiling point = strength of the intermolecular forces between covalent molecules
Larger the covalent molecules, Stronger the intermolecular force, higher the mp / bp

12. Chemical Properties of Halogens:

(a) Halogens are **highly reactive** as they have 7 outermost electrons. So, they obtain an additional electron to achieve a noble gas electronic configuration by gaining/sharing an electron in a chemical reaction.

(b) **Down the group, halogens become less reactive** since their atomic sizes increase, valence electron is further from the nucleus and the attractive forces between nucleus and valence electron weakens. It is more difficult for the atom to attract an electron to its outermost shell. Chlorine (Cl) is less reactive than fluorine (F) but more reactive than bromine (Br).

(c) Halogens are **strong oxidising agents** as they gain an electron and undergo reduction. In the process, halides are formed. For example: $Cl + e^- \rightarrow Cl^-$

Reactions that show halogens behave as oxidising agents:

➔ Halogens react with most metals to form salts called metal halides. Fluoride ions (F^-), chloride ions (Cl^-), bromide ions (Br^-) and iodide ions (I^-) are halide ions. For example, sodium reacts with chlorine gas to produce sodium chloride.
$2Na\ (s) + Cl_2\ (g) \rightarrow 2NaCl\ (s)$

➔ Halogens undergo displacement reactions with halide solutions. A displacement reaction is a reaction in which one element takes the place of another element in a compound. A more reactive halogen will displace a less reactive halogen from its halide solution.
If X and Y are halogens, halogen X has to be more reactive than Y for a displacement reaction to take place: $X_2 + 2Y^- \rightarrow 2X^- + Y_2$

- X_2 oxidises Y^- and acts as the oxidising agent. X_2 gains electrons to form X^- and is reduced.
- Oxidation state of X decreases from 0 in X_2 to –1 in X^-. X_2 is reduced and is an oxidising agent.

Question 6 Pen down the equations for the reactions (if any) between the following substances and explain the reaction:

- $KBr(aq)$ and $Cl_2(g)$ (What would you expect to observe in terms of colour change?)

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

Observation: Colourless solution turns red-brown.
Explanation: Chlorine is more reactive than bromine. Hence, chlorine displaces bromine from colourless aqueous potassium bromide to form colourless aqueous potassium chloride and reddish-brown aqueous bromine.

- $KI(aq)$ and $Cl_2(g)$ (What would you expect to observe in terms of colour change?)

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

Observation: Colourless solution turns brown.
Explanation: Chlorine is more reactive than iodine. Hence, chlorine displaces iodine from colourless aqueous potassium iodide to form colourless aqueous potassium chloride and brown aqueous iodine.

- $KCl(aq)$ and $Br_2(g)$ (What would you expect to observe in terms of colour change?)

Observation: No visible change.
Explanation: Bromine is less reactive than chlorine. Hence, bromine cannot displace chlorine from colourless aqueous potassium chloride.

Question 7 Lab tests for chloride (Cl^-) and iodide (I^-).

Add dilute nitric acid followed by aqueous silver nitrate solution.
For chloride, white precipitate (of silver chloride) is formed.
$Cl^-\,(aq) + Ag^+\,(aq) \rightarrow AgCl\,(s)$
For iodide, yellow precipitate (of silver iodide) is formed.
$I^-\,(aq) + Ag^+\,(aq) \rightarrow AgI\,(s)$

Why silver nitrate used must be added with dilute nitric acid first?
To remove carbonate ions which could also form white precipitate of silver carbonate with aqueous silver nitrate.
$CO_3^{2-}\,(aq) + 2Ag^+\,(aq) \rightarrow Ag_2CO_3\,(s)$

The carbonate ions react with the acid to form carbon dioxide which is released into the air.
$CO_3^{2-}\,(aq) + 2H^+(aq) \rightarrow CO_2\,(s) + H_2O\,(l)$

→ Halogens chemically react with hydrogen to form hydrogen halides (HX). Hydrogen halides form strong acids when dissolved in water.

$$X_2 + H_2 \rightarrow 2HX$$

	Hydrogen Chloride	**Hydrogen Bromide**	**Hydrogen Iodide**
Formation	$Cl_2 + H_2 \rightarrow 2HCl$	$Br_2 + H_2 \rightarrow 2HBr$	$I_2 + H_2 \rightarrow 2HI$
Soluble in water to form acid	HCl dissolves in water to form hydrochloric acid. (strong acid)	HBr dissolves in water to form hydrobromic acid. (stronger acid)	HI dissolves in water to form hydroiodic acid. (strongest acid)

<u>Group 18 – Noble Gases</u>

13. **Group 18 (Noble Gases/Inert Gases)** - They are all unreactive gases as they have **fully filled valence shell**. They do not lose, gain or share electrons. Hence, they rarely react to form compounds.

14. Physical Properties of Noble Gases:
    - Colourless at r.t.p.
    - Have low melting and boiling points since their molecules are held together by weak intermolecular forces.
    - Do not conduct electricity since they exist as monoatomic, neutral covalent molecules.
    - Insoluble in water

15. Physical Properties of Noble Gases down the Group:
    - Going down the group, density increases as the relative atomic mass of the elements increase.
    - Going down the group, melting and boiling points increases due to stronger intermolecular forces of attraction. Molecules have more electrons and are larger in size.

16. Commercially, these gases are obtained from liquid air by fractional distillation.

17. Uses of noble gases:
    - Helium is used for filling balloons as they have low density.
    - Argon is used to fill light bulbs because it is unreactive and protects the white-hot filament and is used to provide an inert atmosphere. Argon is also mixed with oxygen and blown through molten steel to remove excess carbon during making of steel.
      Argon is cheap because it is of abundance.
    - Neon is used in advertisement boards and in landing light bulbs for planes.

Transition Elements

18. Transition elements are metals found in Groups 3 to 12 of the Periodic Table. Transition elements are also called transition metals. Examples of common transition metals include chromium (Cr), manganese (Mn), iron (Fe), copper (Cu).

19. Properties of Transition Metals:

- Transition elements have **higher melting point and boiling point than normal metals** since they form stronger metallic bonds. They have a larger number of delocalized electrons as compared to alkali metals which only has 1 valence electron.

- Transition elements have **high densities** due to their large relative atomic mass and relatively smaller atomic size as compared to alkali metals.

	Group I metal	Transition metal			
Element	Potassium (K)	Chromium (Cr)	Manganese (Mn)	Iron (Fe)	Copper (Cu)
Melting point/°C	63	1857	1244	1538	1084
Boiling point/°C	774	2672	1962	2750	2567
Density/(g/cm^3)	0.86	7.19	7.21	7.86	8.92

- Atoms of transition metals are capable of delocalising more than 2 valence electrons. Thus, they are capable of **displaying variable oxidation states** and forming compounds with **varied colours**.

Name	**Formula**	**Oxidation state**	**Colour**
Manganate (VII) anion	MnO_4^-	+7	purple
Iron (II) cation	Fe^{2+}	+2	green
Iron (III) cation	Fe^{3+}	+3	yellow
Copper (II) cation	Cu^{2+}	+2	blue
Copper (II) cation in excess NH_3 (aq)	$[Cu(NH_3)_4]^{2+}$	+2	blue

Take note: When naming the compounds of a transition element, specify the oxidation state of the transition element in the compound. E.g. iron (II) chloride ($FeCl_2$) or iron (III) chloride ($FeCl_3$). Do not simply write 'iron chloride'.

- Because of the ability of transition metals to **display variable oxidation states**, transition metals and their compounds can act as **catalysts** for certain chemical reactions.

Reaction	Catalyst
N_2 (g) + $3H_2$ (g) → $2NH_3$ (g) (Haber Process for the manufacture of ammonia)	iron
C_2H_4 (g) + H_2 (g) → C_2H_6 (g) Vegetable oil + hydrogen → margarine (Manufacture of margarine from vegetable oil)	nickel
$2H_2O_2$ (aq) → O_2 (g) + $2H_2O$(l) (Decomposition of hydrogen peroxide)	manganese (IV) oxide

Take note: A catalyst is a substance that increases the speed of a chemical reaction and remains chemically unchanged at the end of the reaction.