The Periodic Table

Learning Objectives

(a) Describe the Periodic Table as an arrangement of the elements in 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 structure

(c) Describe relationship between group number and ionic charge of an element

(d) Explain the similarities between elements in the same group of the Periodic Table in terms of their electronic structure

(e) Describe the change from metallic to non-metallic character from left to right across the period of the Periodic Table

(f) Describe the relationship between group number, number of valence electrons and metallic/ non-metallic character

(g) Predict the properties of elements in Group 1 and Group 17 using the Periodic Table

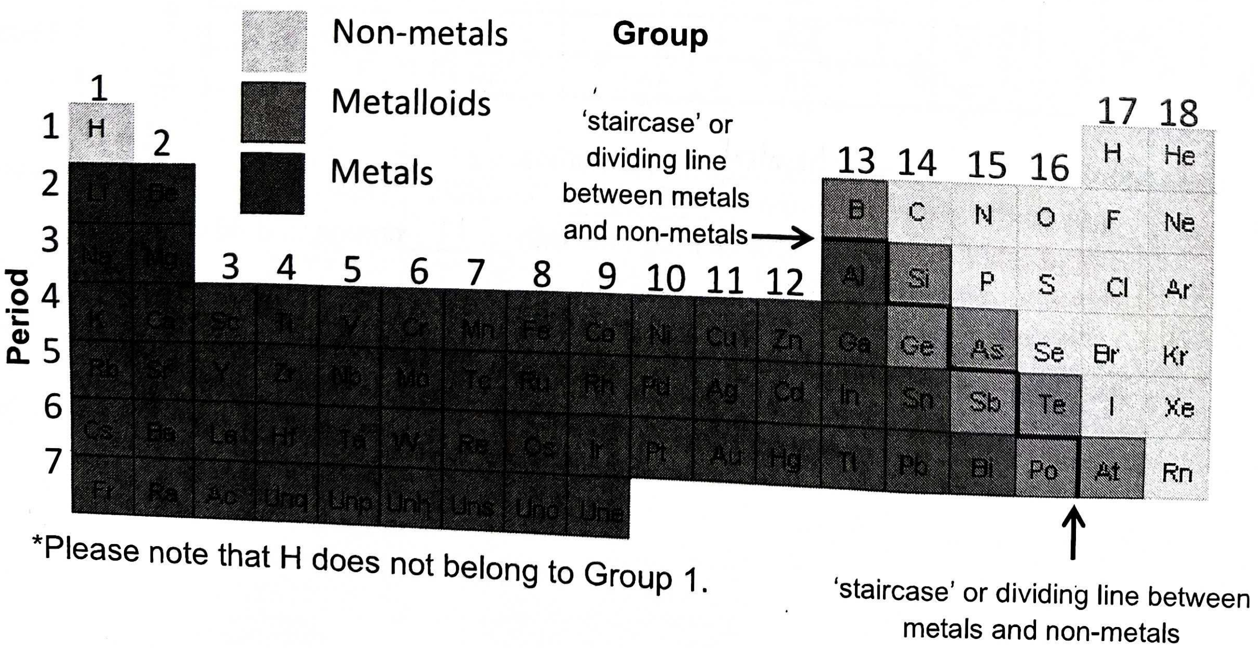
(h) Describe lithium, sodium and potassium in Group 1 as a collection of relatively soft, low density metals showing a trend in melting point and in their reaction with water

(i) Describe chlorine, bromine and iodine in Group 17 as a collection of diatomic non-metals showing a trend in colour, state and their displacement reactions

(j) Describe the element in Group 18 as a collection of monatomic elements that are chemically unreactive and hence important in providing an inert atmosphere

(k) Describe the lack of reactivity of the noble gases in terms of their electronic structures

A. Introduction to the Periodic Table



*Please note that H does not belong to Group 1.

A1. Revision on the Periodic Table

- Periodic Table is a way of arranging the elements in order of increasing atomic number.
- A Group is a vertical column of the elements. There are total of 18 groups: Group 1-18.
    - Group 1 – The Alkali metals
    - Group 2 – The Alkaline Earth Metals
    - Group 17 – The Halogens
    - Group 18 – The Noble gases
- A Period is a horizontal row of elements. There are total of 7 periods: Period 1-7.
    - Hydrogen and Helium is in Period 1.
- Elements in the Periodic Table are divided by the 'staircase' into metals and non-metals.
    - Metals: elements to the left side of the staircase.
    - Non-metals: elements to the right side of the staircase.
    - Elements which are found along the staircase e.g. B, Si, Ge, As, are metalloids. Metalloids are elements that can exhibit properties of both metals and non-metals.
- Transition metals are the elements found in Group 3 to Group 12 of the Periodic Table.

A2. Characteristics of Elements in the Same Group

1. For elements in Group 1 and 2, the elements have the same number of **valence electrons** as the group number.

   For elements from Group 13 – 18, the elements have the number of **valence electrons** equals to group number minus 10.

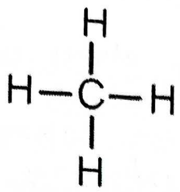
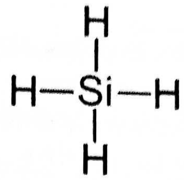
   Note: **Hydrogen does not belong to Group 1 elements.**

Group	Elements	Electronic Configuration
1	Li	2,1
	Na	2,8,1
	K	2,8,8,1
2	Mg	2,8,2
	Ca	2,8,8,2
	Sr	2,8,18,8,2
13	Al	2,8,3
	Ga	2,8,18,3
	In	2,8,18,18, 3

2. The elements have similar chemical properties.
    - They form ions of the same charges.
      Example: all Group 1 elements form ions with 1+ charge and all Group 17 elements form ions with 1- charge.
    - They form compounds with similar chemical formulae.
      Example: Formulae of the chlorides of Group 2 elements: $MgCl_2$, $CaCl_2$, $SrCl_2$

- They form the same number of bonds when they share electrons with other elements.

Example: Carbon and silicon are from Group 14.
Each atom forms four single covalent bonds with 4 hydrogen atoms.

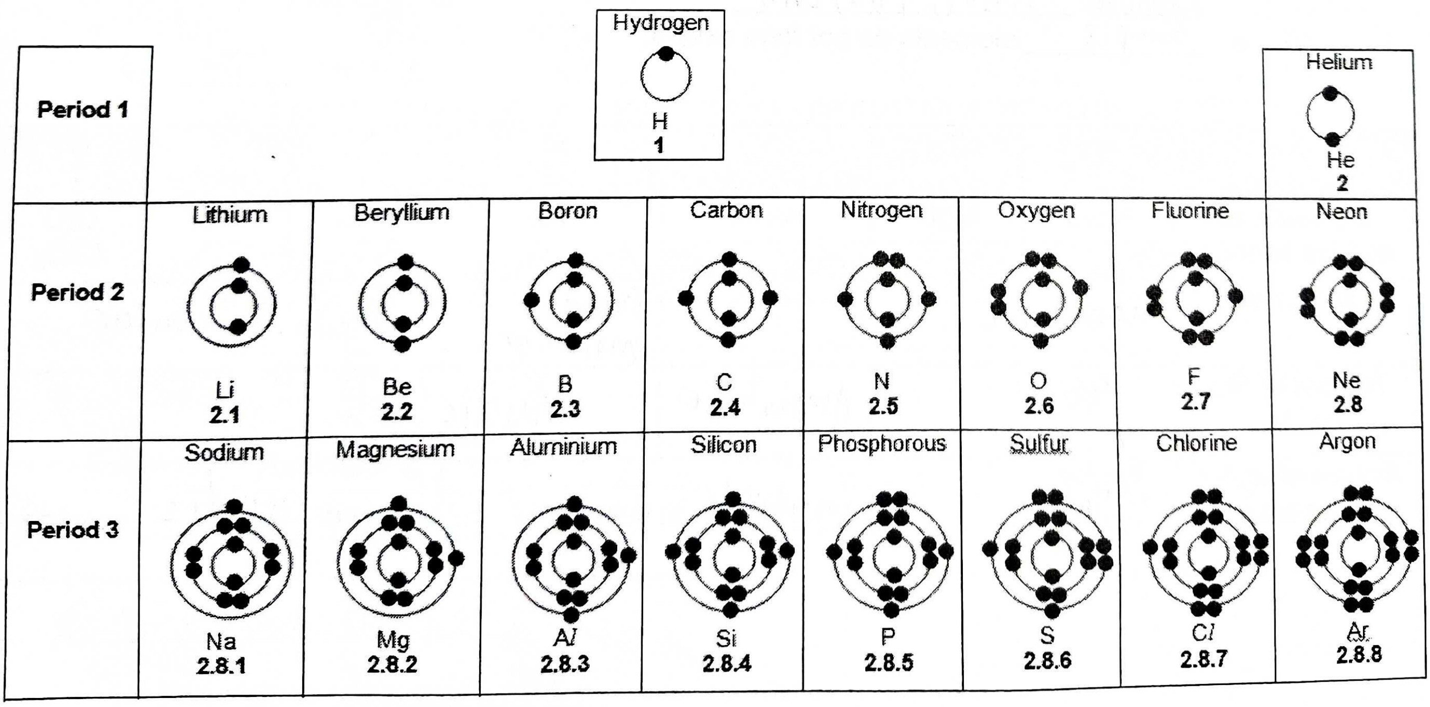
A3. Changes in the Elements Going Down a Group

1. The number of protons in an atom increases.
2. The atomic radii of the elements increases.
   - This is because the number of electron shells increases.

Group	Elements	Proton number	No. of electron shells
1	Li	3	2
	Na	11	3
	K	19	4
	Rb	37	5

A4. Characteristic of Elements in the Same Period

1. The elements have the same number of electron shells, where the number of shells is equal to the period number.



A5. Changes in the Elements Going Across a Period

1. The number of protons in an atom increases.

2. The metallic character of the elements decreases or the non-metallic character of the elements increases.
   - Metallic character is defined by the ability of the atoms to lose valence electrons (to other substance) and form positive ions.
   - Non-metallic character is defined by the ability of the atoms to gain valence electrons (either by sharing electrons or forming negative ions).

Period 3 Elements	Na	Mg	Al	Si	P	S	Cl	Ar
Proton Number	11	12	13	14	15	16	17	18
Electronic Configuration	2,8,1	2,8,2	2,8,3	2,8,4	2,8,5	2,8,6	2,8,7	2,8,8
Nature of elements	Metallic			Metalloid	Non-Metallic			
Structure	Giant metallic			Giant Molecular	Simple molecular			Free atoms
Electrical Conductivity	Good			Poor/Non-conducting				

3. The properties of the oxides of the elements change gradually across the period.
   - Oxides of metals are usually basic.
   - Oxides of non-metals are usually acidic.
   - Oxides of metalloids or the elements close to the dividing line may exhibit both basic and acidic properties, i.e. their oxides are amphoteric.
     Example: Al_2O_3, PbO, ZnO
   - Group 18 elements do not form oxides.

Period 3 Elements	Na	Mg	Al	Si	P	S	Cl
Formula of oxides formed	Na_2O	MgO	Al_2O_3	SiO_2	P_4O_6 P_4O_{10}	SO_2 SO_3	ClO ClO_2
Structure	Giant ionic			Giant molecular	Simple molecular		
Nature of the oxides	Basic		Amphoteric	Acidic			
Action of water on the oxides	Forms alkaline solution	Insoluble in water			Forms acidic solution		

B. Group 1 (The Alkali Metals)

- The elements in Group 1 are called the alkali metals because the metals react with water to produce an alkali solution (aqueous metal hydroxide), giving out hydrogen gas.
  Example: $2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$

- They have 1 valence electron, form ions with charge 1+ and have valency 1.

Note: Valency is defined as the number of electrons involved in bonding.

Element	Symbol	Proton number	Electronic configuration
Lithium	Li	3	2,1
Sodium	Na	11	2,8,1
Potassium	K	19	2,8,8,1
Rubidium	Rb	37	2,8,18,8,1
Caesium	Cs	55	2,8,18,18,8,1
Francium	Fr	87	2,8,18,32,18,8,1

Note: Hydrogen does NOT belong to Group 1.

B1. Physical Properties of Group 1 Elements

1. Good conductor of electricity.
   - The elements have giant metallic structure. There is a 'sea' of free-moving valence electrons available to conduct electricity.

2. Soft and can be easily cut with a knife.

3. Shiny, silvery solids when freshly cut.

4. Low density.
   - Lithium, sodium and potassium are less dense than water and thus, float on water.

5. Low melting points (when compared to other metals).

Element	Symbol	Density (g/cm^3)	Melting Point (°C)
Lithium	Li	0.53	180
Sodium	Na	0.97	98
Potassium	K	0.86	64
Rubidium	Rb	1.53	39
Caesium	Cs	1.88	29

B2. Trends Down Group 1

1. The melting point decreases down the Group.

2. Reactivity increases down the Group.
   y Caesium is more explosive than Sodium

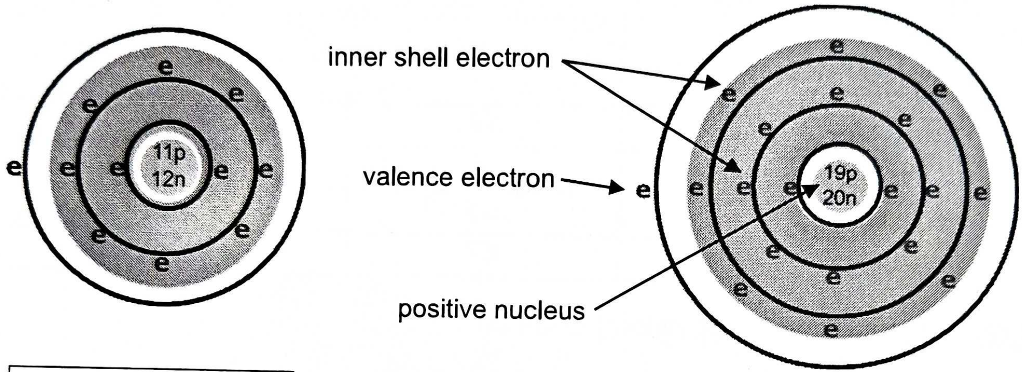
B3. Chemical Properties of Group 1 Elements

1. The metal atoms lose their one valence electron readily (and undergo oxidation*) to form ions with 1+ charge.

   Example: Na → Na^+ + e^-

2. The reactivity of Group 1 elements increases down the Group. The order of reactivity, in descending order:

   francium, caesium, rubidium, potassium, sodium, lithium



Na atom
2,8,1
2 inner shells

K atom
2,8,8,1
3 inner shells

- Down the group, the radius of the metal atom increases (due to more electron shells).
- The valence shell will be further from the nucleus.
- Thus, there will be weaker electrostatic forces of attraction between the positive nucleus and valence electrons.
- smaller amount of energy is needed to remove the valence electrons.
- Hence, valence electrons are lost more readily, i.e. the metal atoms lose electrons more readily to form positive ions.
- This makes Group 1 metals become more reactive down the Group.

3. Group 1 elements behave as powerful reducing agents*.

4. Group 1 metals react vigorously with oxygen to produce metal oxides.

Element	Chemical Equation for the Reaction with Oxygen
Lithium	$4Li(s) + O_2(g) \rightarrow 2Li_2O(s)$
Sodium	$4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$
Potassium	$4K(s) + O_2(g) \rightarrow 2K_2O(s)$
Rubidium	$4Rb(s) + O_2(g) \rightarrow 2Rb_2O(s)$

*** These terms will be taught in the topic, Redox Reactions.**

5. Group 1 metal reacts vigorously with water to produce alkali solution and hydrogen.

Element	Reaction with Water	Observations
Lithium	Reacts quickly. $2Li(s) + 2H_2O(l) \rightarrow 2LiOH(aq) + H_2(g)$	• Silvery solid darts rapidly on the surface of water. • Fizzing occurs. / Effervescence of colourless, odourless gas is seen. • Silvery solid catches fire (except lithium). • Silvery solid decreases and dissolves to form a colourless solution.
Sodium	Reacts violently. $2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$	
Potassium	Reacts very violently. $2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$	
Rubidium	Reacts very explosively. $2Rb(s) + 2H_2O(l) \rightarrow 2RbOH(aq) + H_2(g)$	

Note: Group 1 metals are stored in oil to prevent the metals from reacting with oxygen and water vapour in the air (and become tarnished).

6. Group 1 metals react vigorously and burn brightly when heated with chlorine gas to produce white solid of metal chlorides (ionic compound).

Element	Chemical Equation for the Reaction with Chlorine gas
Lithium	$2Li(s) + Cl_2(g) \rightarrow 2LiCl(s)$
Sodium	$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$
Potassium	$2K(s) + Cl_2(g) \rightarrow 2KCl(s)$
Rubidium	$2Rb(s) + Cl_2(g) \rightarrow 2RbCl(s)$

7. Oxides of Group 1 metals are basic in nature.

 - Group 1 oxides react with water to produce strong alkalis (pH 13 – 14).

 Examples: $Li_2O(s) + H_2O(l) \rightarrow 2LiOH(aq)$

 $Na_2O(s) + H_2O(l) \rightarrow 2NaOH(aq)$

 $K_2O(s) + H_2O(l) \rightarrow 2KOH(aq)$

8. Carbonates of Group 1 metals are soluble in water.

9. Carbonates of Group 1 metals are generally thermally stable.

 - Thermally stable means that the compounds do not decompose (break down) upon heating.

 Examples:

Carbonate of a Group 1 metal	Carbonate of a transition metal
$Na_2CO_3 \xrightarrow{\text{heat}}$ No reaction	$CuCO_3 \xrightarrow{\text{heat}} CuO + CO_2$

Note: Group 1 **compounds** are generally white solid and forms colourless solution.

C. Group 17 (The Halogens)

- Group 17 elements are known as halogens because they react with most metals to form salts (halogen in Greek means salt-former).
  Example: $2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$ *[a salt]*
- They have 7 valence electrons, form ions with charge 1− and have valency 1.
- They exist as diatomic molecules.
- 'Halogen': Group 17 elements e.g. F_2, Cl_2, Br_2, etc.
- 'Halide': ions formed by Group 17 elements e.g. F^-, Cl^-, Br^-, etc.

Element	Symbol	Proton number	Electronic configuration
Fluorine	F	9	2,7
Chlorine	Cl	17	2,8,7
Bromine	Br	35	2,8,18,7
Iodine	I	53	2,8,18,18,7
Astatine	At	85	2,8,18,32,18,7

C1. Physical Properties of Group 17 Elements

1. The elements are coloured.
   - The colour becomes progressively darker down the Group.

Element	Colour in various states
Fluorine	**$F_2(g)$: pale-yellow
Chlorine	**$Cl_2(g)$: greenish yellow $Cl_2(aq)$: pale yellow
Bromine	**$Br_2(l)$: reddish-brown $Br_2(aq)$: orange $Br_2(g)$: reddish-brown
Iodine	**$I_2(s)$: black $I_2(aq)$: brown $I_2(g)$: purple I_2 dissolved in organic solvent: purple

Note: (**) in the table above refers to the state at room temperature and pressure.

2. Low melting and boiling point.
   - The elements have simple molecular structure (exist as simple molecules). little energy is required to overcome weak intermolecular forces of attraction.

Element	Formula	Melting point (°C)	Boiling point (°C)	Physical state at r.t.p.
Fluorine	F_2	-220	-189	Gas
Chlorine	Cl_2	-101	-35	Gas
Bromine	Br_2	-7	59	Liquid
Iodine	I_2	114	184	Solid

3. Non-conductor of electricity.
   - The elements have simple molecular structure (exist as simple molecules). There are no free-moving ~~free~~ valence electrons available to conduct electricity.

4. Group 17 elements are generally soluble in water.

C2. Trends Down Group 17

1. The melting and boiling points increase down the Group. Physical state of elements at r.t.p. changes from gas to liquid to solid.
   - The sizes of the halogen molecules increase down the Group.
   - There are greater surface areas of interaction between the larger molecules.
   - Thus, more energy is required to overcome the stronger intermolecular forces of attraction.

2. Reactivity decreases down the Group.

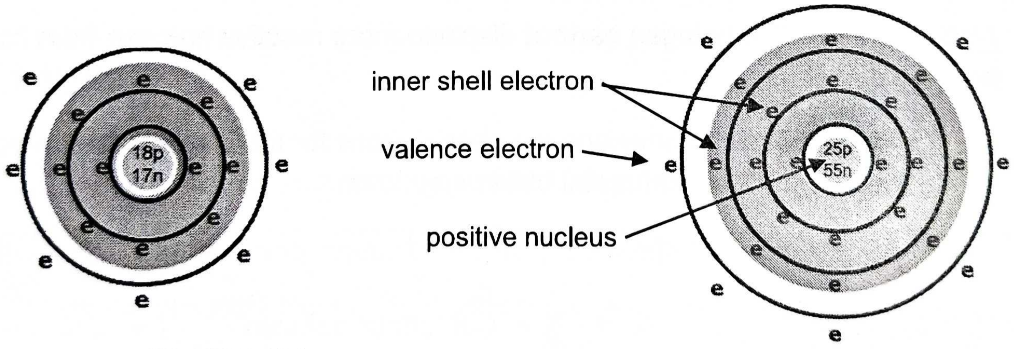
3. The colour of the element becomes progressively darker down Group 17.

C3. Chemical Properties of Group 17 Elements

1. The halogen atoms gain one electron readily (and undergo reduction*) to form ions with 1- charge.

2. The reactivity of Group 17 elements decreases down the Group. The order of reactivity, in descending order:

   fluorine, chlorine, bromine, iodine, astatine



Cl atom	Br atom
2,8,7	2,8,18,7
2 inner shells	3 inner shells

- Down the group, the sizes of the halogen atom increases.
- The valence shell will be further from the nucleus.
- Thus, it is more difficult for the positive nucleus to attract an electron into the valence shell as the electrostatic forces of attraction are weaker.
- Hence, electrons are gained less readily into the valence shell, i.e. the halogen atoms gain electrons less readily to form negative ions.
- This makes Group 17 elements become less reactive down the Group.

3. Group 17 elements behave as powerful <u>oxidising agents</u>*.

4. Group 17 elements undergo displacement reactions.
   - A more reactive halogen will displace a less reactive halogen from its aqueous halide solution.

<u>*Example 1*</u>:
When aqueous chlorine (or chlorine water) is added to sodium bromide solution, the following reaction occurs:

- Chlorine is more reactive than bromine and thus, chlorine gains electron **<u>more readily</u>** than bromine.
- Chlorine displaces bromine from colourless <u>aqueous sodium bromide</u> to form orange aqueous bromine and colourless aqueous sodium chloride
- Observation: colourless solution turns orange.
- Ionic equation: $Cl_2(aq) + 2Br^-(aq) \rightarrow 2Cl^-(aq) + Br_2(aq)$

<u>*Example 2*</u>:
When chlorine gas is bubbled through potassium iodide solution, the following reaction occurs:

- Chlorine is more reactive than iodine and thus, chlorine gains electron **<u>more readily</u>** than iodine.
- Chlorine displaces iodine from colourless <u>aqueous potassium iodide</u> to form brown aqueous iodine and colourless aqueous potassium chloride.
- Observations: colourless solution turns brown. Black Precipitate may be seen (iodine formed is partially soluble in water).
- Ionic equation: $Cl_2(g) + 2I^-(aq) \rightarrow 2Cl^-(aq) + I_2(aq)$

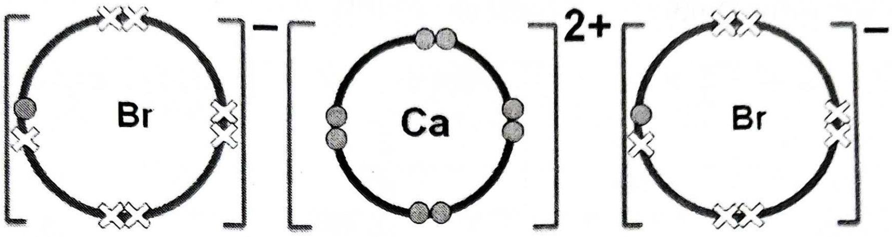
<u>*Note*</u>: Less reactive halogen **cannot** displace more reactive halogen from its aqueous halide solution.

- The table below summarises the observations for the displacement reactions of halogens with different metal halide solutions.

	Aqueous metal chloride	Aqueous metal bromide	Aqueous metal iodide
Chlorine, Cl_2		• Colourless solution turns orange. $Cl_2 + 2Br^- \rightarrow 2Cl^- + Br_2$	• Colourless solution turns brown. • Black precipitate may be seen. $Cl_2 + 2I^- \rightarrow 2Cl^- + I_2$
Bromine, Br_2	No visible reaction		• Colourless solution turns brown. • Black precipitate may be seen. $Br_2 + 2I^- \rightarrow 2Br^- + I_2$
Iodine, I_2	No visible reaction	No visible reaction	

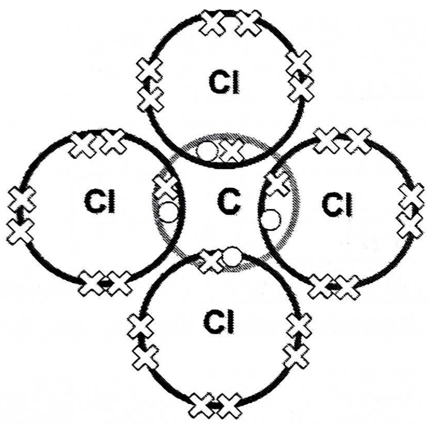
5. Group 17 elements can form both ionic and covalent compounds.
   - They react with most metals to form ionic compounds by gaining one valence electron.

     Example: $CaBr_2$

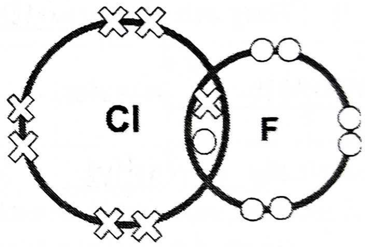


   - They react with non-metals or other halogens to form covalent compounds by sharing one of their valence electrons with another atom.

     Example 1: CCl_4



Example 2: ClF



C4. Uses of Halogens

- Disinfectant
  - Chlorine is added in swimming pools to kill germs and bacteria.

- Bleaching agent
  - Chlorine is used as bleaching agent for fabrics, papers, etc.

not reactive

D. Group 18 (The Noble Gases)

- Elements in Group 18 are called the noble gases.
- Helium has 2 valence electrons (stable duplet configuration) while the rest of the elements in the group has 8 valence electrons (stable octet configuration).
- They are monatomic (exist as atoms).

Element	Symbol	Proton number	Electronic configuration
Helium	He	2	2
Neon	Ne	10	2,8
Argon	Ar	18	2,8,8
Krypton	Kr	36	2,8,18,8
Xenon	Xe	54	2,8,18,18,8
Radon	Rn	86	2,8,18,32,18,8

D1. Properties of the Noble Gases

1. Low melting and boiling points.
   - They are all colourless gases at r.t.p.

2. Insoluble in water.

3. Chemically unreactive (or chemically inert).
   - Their valence shells are filled to the maximum number of electrons. Thus, the atoms do not gain, lose or share electrons with other atoms.

Element	Melting point (°C)	Boiling point (°C)	Density (g/cm^3)	% in atmosphere
Helium	-270	-268	0.177	0.00052
Neon	-290	-280	0.900	0.0018
Argon	-189	-186	1.78	0.93

Note: Argon is the most abundant among all the noble gases in the atmosphere.

D2. Uses of the Noble Gases

Noble Gas	Uses
Helium	• Fill weather balloons and airships because it has a low density and is not flammable.
Neon	• Making of advertisement lights.
Argon	• Fill light bulbs to provide an inert environment to protect the tungsten filament from reacting with oxygen under high temperature.

E. Transition Metals

- Transition metals are the elements in Group 3 to Group 12 of the Periodic Table.
- Most transition metals are silvery or grey in colour, except copper (pink or reddish-brown) and gold.

Note: Zinc is not considered to be a transition metal as it does not have variable oxidation states* and does not form coloured compounds.

E1. Properties of Transition Metals

1. Good conductor of electricity
   - The elements have giant metallic structure. There is a 'sea' of free-moving valence electrons available to conduct electricity.
2. High melting and boiling points.
   - A lot of energy is required to overcome the strong electrostatic forces of attraction between cations and 'sea' of free-moving valence electrons.
3. High density.
4. Exhibit variable oxidation states (form more than one type of ions) in their compounds.

Transition metals	Melting point (°C)	Boiling point (°C)	Density (g/cm^3)	Common ions
Iron	1535	3000	7.9	Fe^{2+} Fe^{3+}
Copper	1083	2595	8.9	Cu^{+} Cu^{2+}
Manganese	1240	2100	7.2	Mn^{2+}, Mn^{4+}
Chromium	1890	2482	7.2	Cr^{3+}, Cr^{6+}

Note: Zinc and silver only exhibit one valency (form one type of ion) in their compounds, Zn^{2+} and Ag^{+}. We do not include valency when we name compounds involving zinc and silver.

Examples: $Zn(NO_3)_2$ is zinc nitrate ; $AgCl$ is silver chloride.

5. Form coloured compounds.

Compound	Colour	Compound	Colour
$FeCl_2$	pale Green	$FeCl_3$	Yellow
$CuCO_3$	Green	$CuSO_4$	Blue
MnO_2	Black	$KMnO_4$	Purple

Note: The colour of transition metal compounds may be different when hydrated and when anhydrous

Examples: Copper(II) sulfate is blue when hydrated and white when anhydrous.
Cobalt(II) chloride is pink when hydrated and blue when anhydrous.

6. Transition metals and their compounds are good catalysts.

Catalyst	Industrial process
Iron, Fe	Production of ammonia from nitrogen and hydrogen (Haber process)
Nickel, Ni	Production of margarine from vegetable oil
Manganese(IV) oxide, MnO_2	Decomposition of hydrogen peroxide