

1. THE PERIODIC TABLE

1.1 Introduction

When you ask the store keeper in your Chemistry laboratory for a particular chemical, he does not search through all the shelves for it. Rather, he goes straight to a particular shelf. The same is true when you look for a book in the school library. If you know the title of the book and its subject matter, you know where to look for it. These are examples of the importance of classification.

The periodic table is a classification of the elements into vertical columns (those that react alike), and horizontal rows (those whose atoms have the same number of electronic shells). Such a classification reduces the task of studying the properties of each of more than a hundred different elements. With this classification, a prediction can be made of the behaviour of a compound yet to be prepared.

Several attempts were made by earlier chemists like Johann Wolfgang Docheriner in 1829 and John Alexander Newlands in 1863 to classify the elements on the basis of similarities in their properties. However, it was Dmitri Mendeleev and Julius Lothar Meyer who came up with a good classification.

Mendeleev and Meyer in 1869 published separate Periodic Tables which bear some resemblance to the modern Periodic Table. Mendeleev stated that when the elements were arranged in order of increasing relative atomic masses, there was a periodic reoccurrence of properties. He called this the **periodic law**.

1.2 The Modern Periodic Table

In the modern periodic table the elements are arranged in an increasing order of their atomic numbers (Table 1.1) The periodic reoccurrence of properties is even more obvious in the modern periodic table because it eliminates a number of irregularities which occur in Mendeleev's table. There are eight vertical columns, called **groups**, and seven horizontal rows, called **periods**.

The first period contains only two elements, hydrogen and helium. The second and third contain eight elements each. They are called the short periods.

The fourth and fifth periods contain eighteen elements each and are referred to as the long periods. Eight of the eighteen elements in the

long periods belong to the main groups of elements. The other ten do not, and are known as the **transition elements**. Period 6 contains thirty-two elements while period 7 contains seventeen.

1.3 Group 1 Elements: Alkali Metals (Li, Na, K, Rb, Cs)

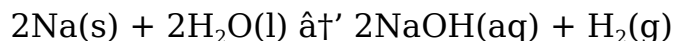
Lithium, sodium, potassium, rubidium and caesium belong to the first group of the periodic table.

Experiment 1.1: Investigating the action of the alkali metals on water.

Place a trough half-filled with water on the work bench. Drop a small cube of sodium metal into the water and observe what happens.

The sodium darts about on the surface of the water (Figure 1.1) producing a hissing sound. Some white fumes can be seen rising from the spot where the piece of sodium metal is located at any time. The sodium should be allowed to dart around the trough to avoid explosion. After some time, the piece of sodium rounds up into a tiny ball and diminishes in size, till it gradually disappears completely. Add two drops of an indicator like phenolphthalein to the solution. The solution turns purple showing that the solution is alkaline.

For a more exciting result, repeat the experiment by adding three drops of phenolphthalein indicator, with stirring. When a small cube of sodium metal is dropped into the water it darts about on the surface of the water leaving purple furrows along its path. You will also observe all the reactions that occurred in the first experiment. The purple furrows are formed as a result of phenolphthalein being turned into this colour by the alkaline solution (NaOH) produced as sodium reacts with water along its path.



The hissing sound is due to the liberation of the hydrogen, while the fume is due to the evaporation of water from the reaction which is exothermic.

Repeat the experiment with cubes of potassium and lithium metals.

PERIODIC TABLE																				
	Group 1	Group 2													Group 3	Group 4	Group 5	Group 6	Group 7	Group 8
PERIOD 1	1 H Hydrogen																			2 He Helium
PERIOD 2	3 Li Lithium	4 Be Beryllium	TRANSITION METALS												5 B Boron	6 C Carbon	7 N Nitrogen	8 O Oxygen	9 F Fluorine	10 Ne Neon
PERIOD 3	11 Na Sodium	12 Mg Magnesium													13 Al Aluminium	14 Si Silicon	15 P Phosphorus	16 S Sulphur	17 Cl Chlorine	18 Ar Argon
PERIOD 4	19 K Potassium	20 Ca Calcium	21 Sc Scandium	22 Ti Titanium	23 V Vanadium	24 Cr Chromium	25 Mn Manganese	26 Fe Iron	27 Co Cobalt	28 Ni Nickel	29 Cu Copper	30 Zn Zinc	31 Ga Gallium	32 Ge Germanium	33 As Arsenic	34 Se Selenium	35 Br Bromine	36 Kr Krypton		
PERIOD 5	37 Rb Rubidium	38 Sr Strontium	39 Y Yttrium	40 Zr Zirconium	41 Nb Niobium	42 Mo Molybdenum	43 Tc Technetium	44 Ru Ruthenium	45 Rh Rhodium	46 Pd Palladium	47 Ag Silver	48 Cd Cadmium	49 In Indium	50 Sn Tin	51 Sb Antimony	52 Te Tellurium	53 I Iodine	54 Xe Xenon		
PERIOD 6	55 Cs Caesium	56 Ba Barium	57 La Lanthanum	58 Ce Cerium	59 Pr Praseodymium	60 Nd Neodymium	61 Pm Promethium	62 Sm Samarium	63 Eu Europium	64 Gd Gadolinium	65 Tb Terbium	66 Dy Dysprosium	67 Ho Holmium	68 Er Erbium	69 Tm Thulium	70 Yb Ytterbium	71 Lu Lutetium			
PERIOD 7	87 Fr Francium	88 Ra Radium	89 Ac Actinides	90 Th Thorium	91 Pa Protactinium	92 U Uranium	93 Np Neptunium	94 Pu Plutonium	95 Am Americium	96 Cm Curium	97 Bk Berkelium	98 Cf Californium	99 Es Einsteinium	100 Fm Fermium	101 Md Mendelevium	102 No Nobelium	103 Lr Lawrencium			

Lanthanide Series	* 139 La Lanthanum	140 Ce Cerium	141 Pr Praseodymium	144 Nd Neodymium	147 Pm Promethium	150 Sm Samarium	152 Eu Europium	157 Gd Gadolinium	159 Tb Terbium	162.5 Dy Dysprosium	165 Ho Holmium	167 Er Erbium	169 Tm Thulium	173 Yb Ytterbium	175 Lu Lutetium
Actinides Series	* 227 Ac Actinium	232 Th Thorium	231 Pa Protactinium	236 U Uranium	237 Np Neptunium	242 Pu Plutonium	243 Am Americium	247 Cm Curium	247 Bk Berkelium	251 Cf Californium	254 Es Einsteinium	254 Fm Fermium	253 Md Mendelevium	256 No Nobelium	257 Lr Lawrencium

Table 1.1

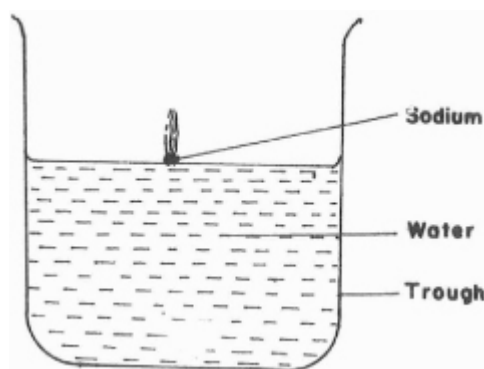


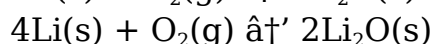
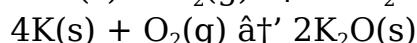
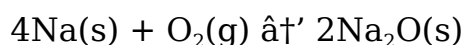
Figure 1.1 Action of alkali metals on water

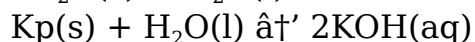
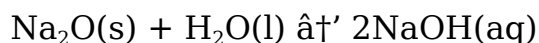
Similar reactions are observed. The rate of evolution of hydrogen and the speed of the darting metals will show that potassium reacts faster than sodium while lithium reacts more slowly.

Experiment 1.2: Investigating the action of the alkali metals on oxygen.

Collect three gas jars of dry oxygen. Melt a piece of sodium in a deflagrating spoon by heating it, then lower it into one of the gas jars (Figure .2). Repeat with potassium, and then with lithium. The metals burn with bright flames, leaving white oxide residues.

Shake each residue with water and test the resulting solutions with litmus paper. The solid oxides dissolve in water to give alkaline solutions.





Experiment 1.3: Investigating the action of alkali metals on chlorine.

Collect three gas jars of dry chlorine, then repeat Experiment 1.2 with the chlorine gas in place of oxygen. White fumes of the chlorides of each metal will be observed inside the gas jars. On shaking with water, the chlorides dissolve to form neutral solutions. Potassium again shows greater reactivity as it produces the white fumes more readily. Lithium is the least reactive.

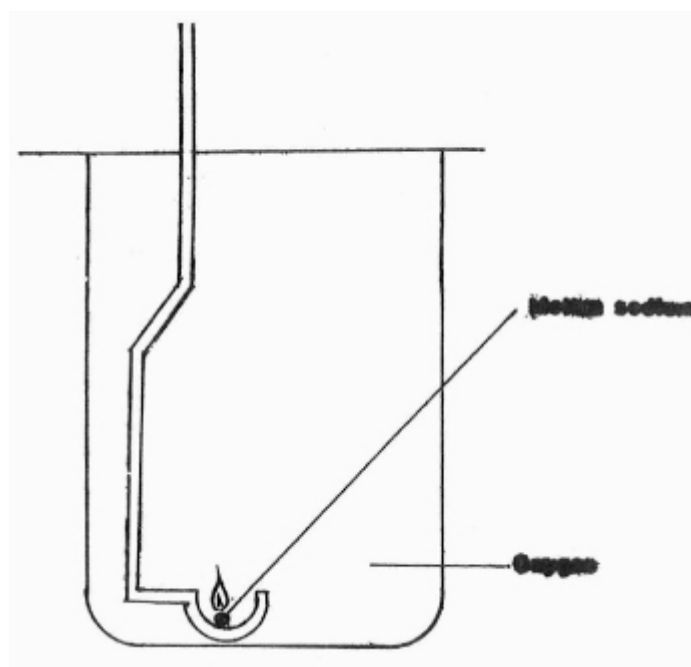
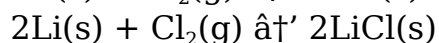
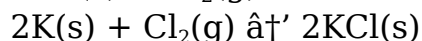
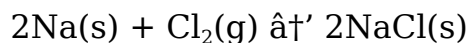
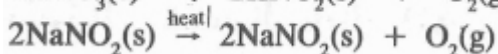
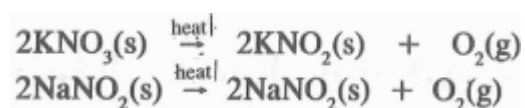


Figure 1.2 Burning sodium in oxygen

The results of Experiments 1.1 – 1.3 indicate that lithium, sodium and potassium react alike, and that potassium reacts faster than sodium, but lithium reacts more slowly. In our study of carbon and its compounds in Book 1, we saw that the trioxocarbonates(IV) of sodium and potassium are soluble in water, and are not decomposed by the heat of a Bunsen burner. The trioxonitrates(V) of these metals also show similar behaviours. They liberate oxygen when heated, and leave dioxonitrates(III) of the metals, whereas other metallic trioxonitrates(V) give off oxygen and nitrogen(IV) oxide, leaving metallic oxides.



The metals lithium, sodium, potassium, rubidium, caesium and francium are known as alkali metals because they form strongly alkaline solutions in water. Table 1.2 summarises their similarities. M in the Table represents any alkali metal.

TABLE 1.2: PROPERTIES OF ALKALI METALS

1. Action of water	All liberate hydrogen from cold water: $2M(s) + 2H_2O(l) \rightarrow 2MOH(aq) + H_2(g)$
2. Oxides	Soluble in water, yielding strong alkaline solutions: $M_2O(s) + H_2O(l) \rightarrow 2M^+OH^-(aq)$
3. Hydroxides	All form strongly alkaline solutions: $MOH(s) + aq \rightarrow M^+(aq) + OH^-(aq)$
4. Chlorides	All are soluble in water yielding neutral solutions which are electrolytes: $MCl(s) + aq \rightarrow M^+(aq) + Cl^-(aq)$
5. Trioxocarbonates(IV)	All are soluble in water and are not decomposed by heat: $M_2CO_3(s) + aq \rightarrow 2M^+(aq) + CO_3^{2-}(aq)$
6. Trioxonitrates(V)	All are decomposed to dioxonitrates(III) by heat, liberating oxygen: $2MNO_3(s) \rightarrow 2MNO_2(s) + O_2(g)$

We cannot from the results of these reactions of the alkali metals only, conclude that elements in the same group of the periodic table react alike. The alkali metals make up only one of the eight groups of the periodic table. We need to consider some other groups before making such a conclusion.

1.4 Group II Elements: Alkaline Earth Metals (Be, Mg, Ca, Sr, Ba)

Beryllium, magnesium, calcium, strontium and barium belong to the second group of the periodic table.

Experiment 1.4: Action of water on magnesium, calcium and barium.

Place about 5cm length of magnesium ribbon cleaned with sand-paper into a test-tube containing some water. Watch for 10 minutes, and if nothing happens heat the test-tube. Also place 5mm cube of calcium in a beaker of water, and a small pellet of barium into another beaker of water.

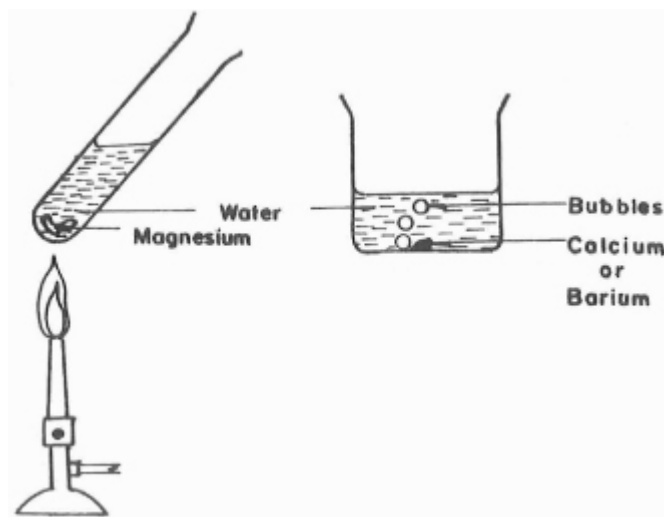
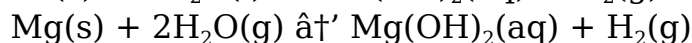
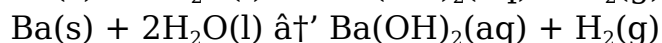
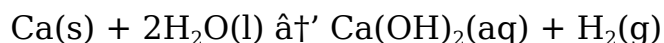


Figure 1.3 Action of water on alkaline earth metals

Barium quickly gives off bubbles of hydrogen gas. Calcium gives off hydrogen too, but less readily (Figure 1.3). Magnesium does not liberate hydrogen from cold water. Bubbles of the gas are observed only after heating the test-tube containing the magnesium. A constant flow of hydrogen will however be obtained if steam is passed through heated magnesium.



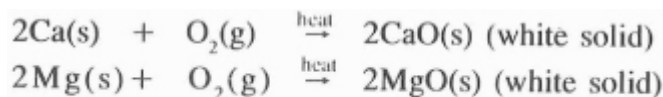
EXERCISE 1A

What is the order of reactivity of magnesium, calcium and barium? Does it agree with the order of reactivity of the alkali metals?

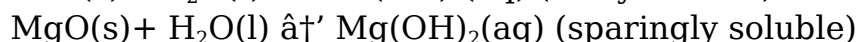
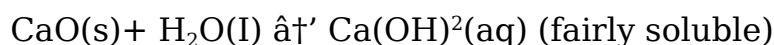
Experiment 1.5: Investigating the action of oxygen on calcium and magnesium.

Collect two gas jars of oxygen. Heat a piece of calcium metal till it is red hot, then plunge it into one of the gas jars. Light a piece of magnesium ribbon while holding it with a pair of tongs, then plunge it into the second gas jar. When the metals have stopped burning, test the oxide produced for solubility in water. Also test the resulting solutions (suspensions in water) with litmus papers.

Each of the metals burn in oxygen, forming a white oxide residue of the metal. Calcium burns with a brick-red flame.



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These oxides are less soluble in water than the oxides of the alkali metals. The magnesium oxide solution turns red litmus only faintly blue, showing that it is weakly alkaline, while the calcium oxide solution turns the red litmus blue, showing that it is a stronger alkaline solution.

Table 1.3 is a summary of the properties of the alkaline earth metals. From it and from the results of Experiments 1.4 and 1.5 we see that the group of alkaline earth metals is another family of similar elements.

TABLE 1.3 PROPERTIES OF MAGNESIUM, CALCIUM, AND BARIUM

	Magnesium	Calcium	Barium
1. Action of air	Tarnishes in air.	Tarnishes in air.	Stored in oil to avoid reaction with moisture in air.
2. Action of water or steam	Liberates H_2 from hot water.	Liberates H_2 from cold water more readily than Mg.	Liberates H_2 .
3. Chlorides (MCl_2)	Soluble in water. Forms basic chloride when solution is heated.	Less soluble in water. Hydrolysed.	Soluble in water.
4. Oxides (MO)	Very slightly soluble, giving a weak alkaline solution.	Slightly soluble, giving an alkaline solution.	Slightly soluble, giving alkaline solution.
5. Oxide + Water $MO + H_2O \rightarrow M(OH)_2$	Slight action	Heat is evolved.	More heat is evolved.
6. Hydroxide $M(OH)_2$	Sparingly soluble in water.	Slightly soluble in water.	More soluble in water.
7. Trioxocarbonates(IV) (MCO_3)	Insoluble, decomposed by heat.	Insoluble, decomposed by heat.	Insoluble, decomposed by heat.

EXERCISE 1B

Radium is an alkaline earth metal. Which of the following properties would you expect it and its compounds to show?

- Radium would liberate hydrogen from cold water.
- Radium oxide would liberate enough heat to boil water sprinkled on

it.

C. Radium hydroxide would dissolve in water forming an alkaline solution.

D. Radium trioxocarbonate(IV) would be decomposed by the bunsen flame.

All the properties should be shown by radium or its compounds. Are you in doubt? Go through Table 1.3 again.

1.5 Group VII Elements: The Halogens (F, Cl, Br, I, At)

The halogens, fluorine, chlorine, bromine, iodine and astatine, resemble each other so closely that they come readily to mind whenever we talk of families of elements. We shall therefore consider them next in our quest for evidence for the statement ‘Elements in the same group of the periodic table react alike’.

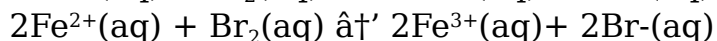
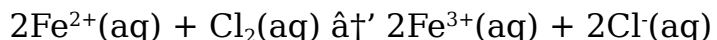
Fluorine, chlorine, bromine and iodine are coloured elements, the colour deepening with increasing atomic number. We are familiar with chlorine as the element added to our drinking or swimming pool water to kill germs. Tincture of iodine is a solution of iodine in ethanol, used for cleaning cuts to prevent them from becoming infected. These two uses already point to some similarities between the elements.

Chlorine and bromine are sufficiently soluble in water for their saturated solutions to be used when the elements are needed. Iodine is only very slightly soluble in water. It however dissolves in potassium iodide solution, and the resulting brown solution serves as iodine solution.

Experiment 1.6: Investigating the oxidising action of the halogens.

Put 2cm³ of iron(II) tetraoxosulphate(VI) solution into three separate test-tubes labelled A, B, C. To test-tube A, add 5cm³ of chlorine water, to test-tube B add 3cm³ of bromine water, and to test-tube C add 2cm³ solution of iodine in aqueous potassium iodide (Figure 1.4). Shake each test-tube after the addition.

The green solution of iron(II) tetraoxosulphate(VI) turns reddish-brown in test-tubes A and B, but remains green in test-tube C. This indicates that iron(II) has been oxidised to iron(III) by chlorine and bromine, but not by iodine.



Add two drops of ammonia solution to each of the three test-tubes. Reddish-brown precipitates are formed in test-tubes A and B. A green precipitate is formed in test tube C. This confirms the presence of Fe³⁺ in test-tubes A and B, and of Fe²⁺ in test-tube C.



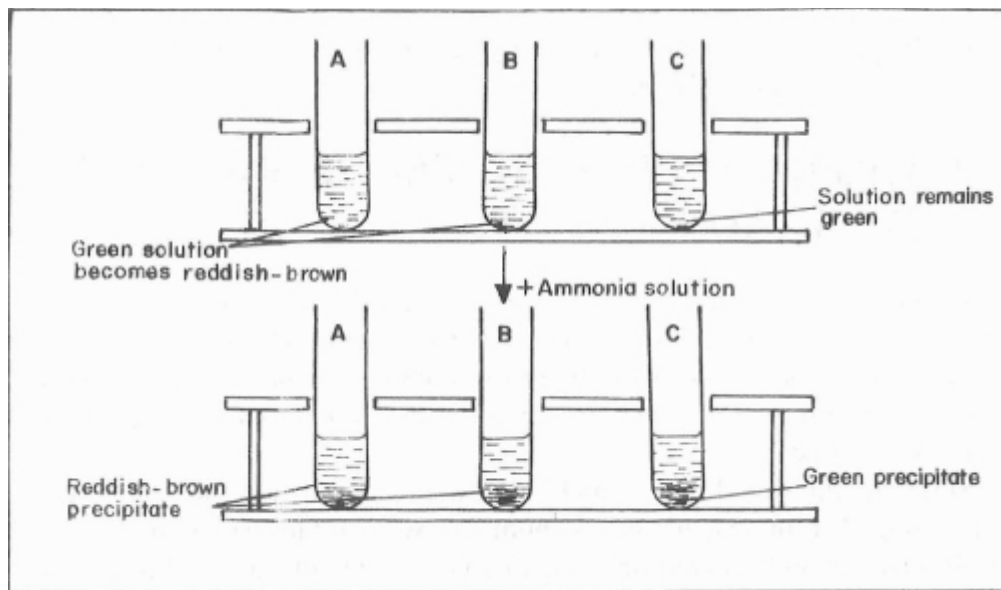
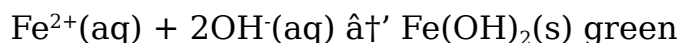


Figure 1.4 Oxidising action of halogens

Chlorine and bromine are stronger oxidizing agents than iodine. Bubble hydrogen sulphide into each of the three test-tubes containing chlorine, bromine and iodine solution respectively. A yellow deposit of sulphur is formed in each tube. Iodine, like chlorine and bromine, is able to oxidize the hydrogen sulphide.



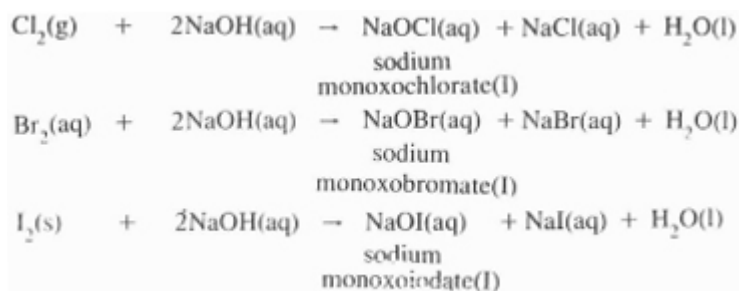
EXERCISE 1C

What is the order of oxidizing power of the halogens?

Experiment 1.7: Investigating the action of sodium hydroxide on the halogens.

Put about 2 cm³ of dilute sodium hydroxide solution into three separate test-tubes labelled X, Y, Z. Into test-tube X, pass chlorine gas for about 30 seconds. To test-tube Y, add about 10 cm³ of bromine water. To test-tube Z, add a small crystal of iodine.

In each case the colour of the halogen gradually fades away, showing that a new compound has been formed.

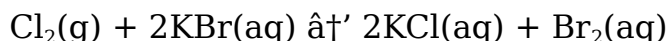


The products are mixtures of the halide and the sodium monoxohalogenate(I). These mixtures are powerful bleaching agents as the monoxohalogenates(I) give off oxygen readily to form the sodium halides.

Experiment 1.8: Investigating the relative reactivities of the halogens.

(a) Add 2 cm³ of tetrachloromethane to 2 cm³ of potassium bromide solution. The solutions separate out into two layers. Bubble chlorine gas into the mixture.

A red colour is formed in the organic layer as the bromine liberated is dissolved in it.



Add bromine water to 2 cm³ of potassium chloride solution mixed with 2 cm³ tetrachloromethane in a test-tube. The organic layer remains clear as chlorine is not liberated.

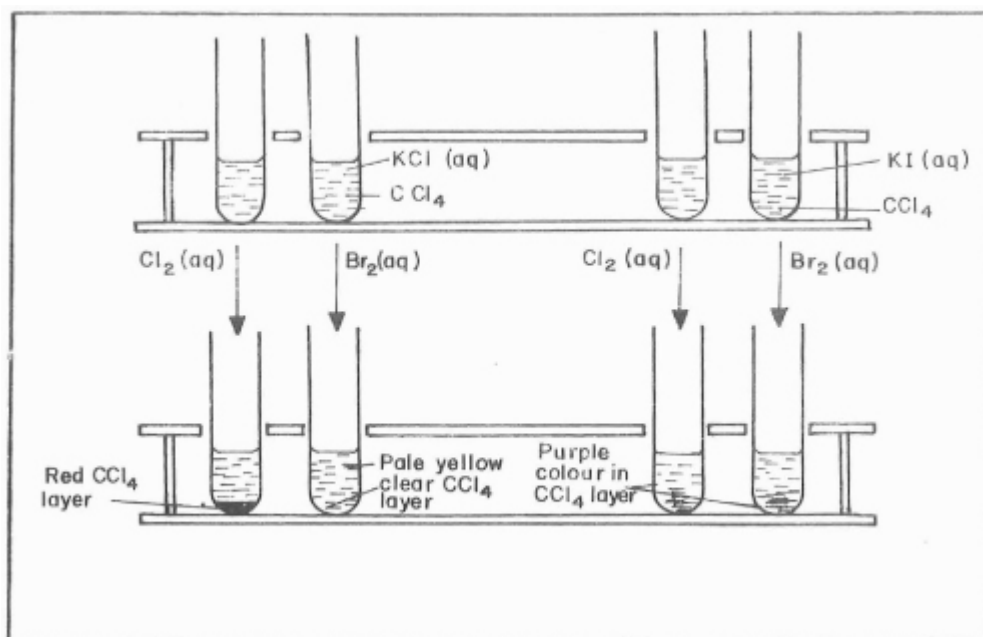
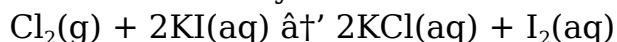
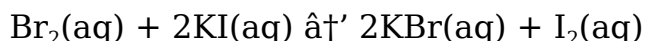


Figure 1.5 Reactivity of the halogens

(b) Add 2 cm³ of tetrachloromethane to 2 cm³ of potassium iodide solution. Bubble chlorine into the mixture. Note the formation of a purple colour in the organic layer as iodine displaced by chlorine, dissolves in this layer.



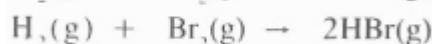
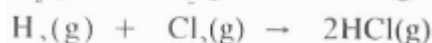
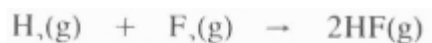
Add bromine water to another mixture of tetrachloromethane and potassium iodide solution and shake. Again, the organic layer becomes purple in colour.



In Experiment 1.8(a) we observe that chlorine displaces bromine from a bromide salt but bromine does not displace chlorine from a

chloride salt. In experiment 1.8 (b) both chlorine and bromine liberate iodine from the iodide. These two experiments show that chlorine is more reactive than bromine and iodine, while bromine is more reactive than iodine.

The reaction between the halogens and hydrogen also illustrates the similarity of the halogens as well as indicating their order of reactivity. Fluorine and hydrogen combine explosively at room temperature, even in the dark. Chlorine and hydrogen combine slowly at room temperature in daylight, but do not combine in the dark. Bromine and hydrogen combine slowly in the presence of bright and direct sunlight, while iodine and hydrogen combine only when heated and the reaction is reversible



The similarities between the halogens are summarized in Table 1.4.

TABLE 1.4 SIMILARITIES OF THE HALOGENS

Property	Fluorine	Chlorine	Bromine	Iodine
State and colour	Pale yellow gas.	Greenish-yellow gas.	Red liquid.	Grey shiny solid.
Solubility in water	Very soluble, giving acidic solution.	Very soluble, giving acidic solution.	Slightly soluble.	Very sparingly soluble.
Oxidation	Very powerful oxidizing agent.	Very powerful oxidizing agent.	Powerful oxidizing agent.	Weak oxidizing agent.
Combination with metals and non-metals	Combines with most metals and non-metals.	Combines with metals and non-metals.	Combines with metals and non-metals.	Combines with metals and a few non-metals.
Reaction with alkali	Forms fluorides.	Forms chlorate (I) with cold dilute alkali; but chlorate(V) with hot concentrated alkali.	Forms bromate(I) with cold dilute alkali; but bromate(V) with hot concentrated alkali.	Forms iodate(I) with cold dilute alkali; but iodate (V) with hot concentrated alkali.
Hydride, HX	Fuming gas; weak acid in water; undissociated at high temperature.	Fuming gas; strong acid in water; dissociates at high temperature.	Fuming gas; strong acid in water; dissociates at lower temperature.	Fuming gas; strong acid in water; readily dissociates.

EXERCISE 1D

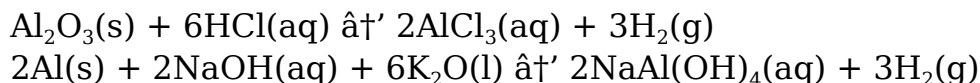
Does the order of reactivity of the halogens follow the same pattern with the alkali and alkaline earth elements?

1.6 Group III Elements (B, Al, Ga, In, Tl)

Aluminium is the only member of the group that is studied at this level of the subject. Its amphoteric nature is the most striking chemical behaviour.

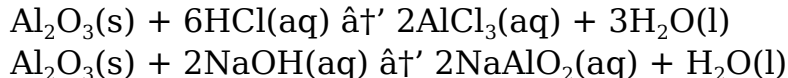
Experiment 1.9: Investigating the action of dilute hydrochloric acid and sodium hydroxide solution on aluminium.

Add dilute hydrochloric acid to about 0.5 g of aluminium in a test-tube and note that the evolution of hydrogen gas occurs only on heating. Add some sodium hydroxide solution to aluminium powder in a test-tube. Hydrogen is also evolved, showing that aluminium is able to liberate hydrogen from both acids and alkalis.



Aluminium oxide and hydroxide also react with acids and alkalis to form salt and water.

Aluminium oxide and hydroxide are said to be amphoteric.



We expect the oxides and hydroxides of the metals following aluminium in the group to behave similarly.

1.7 Group IV Elements (C, Si, Ge, Sn, Pb)

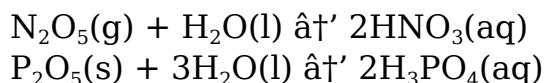
Among these elements, metallic character increases with increasing atomic number as we go down the group. Carbon(IV) oxide and silicon(IV) oxide are acidic, like oxides of other non-metals. Tin(II) oxide and lead(II) oxide are however amphoteric, indicating a tendency towards metallic character. The chlorides of carbon and silicon are covalent, but the chlorides of tin and lead are ionic. Table 1.5 is a summary of the similarities and gradation of properties among the group IV elements.

TABLE 1.5 PROPERTIES OF GROUP IV ELEMENTS

Property	Carbon	Silicon	Tin	Lead
Metallic/ non-metallic nature	Non-metallic	Non-metallic	Metallic	Metallic
Oxides	Acidic, soluble in water	Acidic but insoluble	Amphoteric	Amphoteric
Hydrides	Forms many hydrides	Forms several hydrides	Very unstable and few hydrides	Extremely unstable and few hydrides
Chlorides, XCl_4 XCl_2	Covalent and stable. None	Covalent but unstable. None	Covalent but unstable. Ionic	Covalent and very unstable. Ionic

1.8 Group V Elements (N, P, As, Sb, Bi)

The change from non-metallic to metallic character is again exhibited as atomic number increases down the group. Nitrogen and phosphorus are non-metals. Antimony is a semi-metal while bismuth is a true metal. Both nitrogen and phosphorus form fairly stable hydrides. Ammonia, a hydride of nitrogen, is alkaline, but phosphine, PH_3 is only weakly alkaline. Both nitrogen and phosphorus form acidic oxides which are acid anhydrides.



The oxides of the two elements, arsenic and antimony, which immediately follow phosphorus, are amphoteric.

1.9 Group VI Elements (O, S, Se, Te, Po)

Oxygen and sulphur appear so different in properties from each other, that we cannot appreciate their resemblance now. Both elements however, show allotropy. Normal oxygen and ozone are allotropic forms of oxygen. Rhombic and monoclinic sulphur are allotropes of sulphur. Their other similarities will become evident in due course.

1.10 Group VIII Elements: Noble Gases (He, Ne, Ar, Kr, Xe, Rn)

Helium, neon, argon, krypton, xenon and radon make up a group of elements that are similar in their unreactive natures rather than reactivities. They were regarded as inert (unreactive) until some compounds of krypton, xenon and radon were first prepared in 1962.

Despite this, reactivity among these elements is the exception rather than the rule. No compounds of helium, neon, or argon have been prepared so far, suggesting that even among these noble gases, reactivity increases with increasing atomic number.

EXERCISE 1E

Is it a general trend among all the families (groups) that reactivity increases with increasing atomic number?

1.11 Similarities in Electronic Configuration

From our investigations so far, we can make the following conclusions:

1. Groups I, II, VII and VIII contain elements that are so similar in their reactions that they are regarded as families of elements.
2. Other groups are not so closely related, but some similarities exist among their members and in the reactions of their compounds.
3. Gradation of properties within the groups is a common feature. The next stage in our study is to find out why these patterns of behaviour exist. The theories of bonding indicate that only electrons in the outermost shells are involved in chemical reactions. So an examination of the electronic arrangements of the elements may provide an explanation of the observed patterns; Table 1.6 shows this arrangement for the main group elements. We note from the table, that atoms of elements in the same group have the same number of electrons in their outermost shells. Such atoms therefore lose, gain or share the same number of electrons to form bonds with other atoms. Hence, they react alike.

TABLE 1.6 ELECTRONIC CONFIGURATION OF THE MAIN GROUP ELEMENTS

	GROUPS							
	I	II	III	IV	V	VI	VII	VIII
Electronic configuration	H 1							He 2
Electronic configuration	Li 2,1	Be 2,2	B 2,3	C 2,4	N 2,5	O 2,6	F 2,7	Ne 2,8
Electronic configuration	Na 2,8,1	Mg 2,8,2	Al 2,8,3	Si 2,8,4	P 2,8,5	S 2,8,6	Cl 2,8,7	Ar 2,8,8
Electronic configuration	K 2,8,8,1	Ca 2,8,8,2	Ga 2,8,18,3	Ge 2,8,18,4	As 2,8,18,5	Se 2,8,18,6	Br 2,8,18,7	Kr 2,8,18,8

1.12 The Horizontal Rows (Periods)

It will be interesting to compare the properties of elements in the different groups and see what relationships exist among elements in the same horizontal row (period). The third row is used as a case study. This consists of the elements sodium, magnesium, and aluminium which show metallic characters, and silicon, nitrogen, sulphur, chlorine and argon which are non-metals.

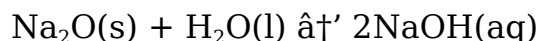
Action of water

Sodium liberates hydrogen from cold water, while magnesium does so from steam. Aluminium liberates hydrogen only from hot dilute acids. In these reactions, the metals are acting as reducing agents in that they lose electrons to the hydrogen ion which becomes a free atom. This is a typical reaction of metals. The ease with which the metals lose electrons decreases from sodium to aluminium, showing that the metallic character decreases from left to right across the period, i.e. metallic character decreases with increasing atomic number.

Silicon, phosphorus and sulphur do not have any action on water. Chlorine dissolves in water, forming the chlorate(I) acid.

Oxides

Both sodium oxide and magnesium oxide are ionic compounds. They are basic oxides. Sodium oxide forms an alkaline solution.



Aluminium oxide is amphoteric.

Oxides of silicon, phosphorus, sulphur and chlorine are acidic. This is a characteristic of non-metallic oxides. The acidity of these oxides when dissolved in water increases from SiO_2 to Cl_2O_7 .

TABLE 1.7 OXIDES OF PERIOD III ELEMENTS

Groups	I	II	III	IV	V	VI	VII
Element	Na	Mg	Al	Si	P	S	Cl
Oxide	Na_2O	MgO	Al_2O_3	SiO_2	P_4O_{10}	SO_3	Cl_2O_7

Chlorides

Sodium chloride and magnesium chloride are soluble in water, yielding electrolytes. Sodium chloride is completely ionised, but magnesium chloride is slightly hydrolysed in water, yielding basic magnesium chloride.



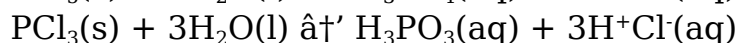
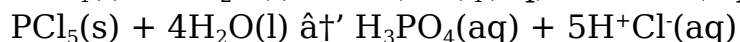
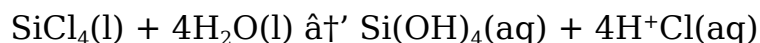
This shows the weaker metallic nature of magnesium when compared with sodium.

Aluminium chloride is much more hydrolysed.



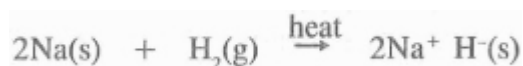
Anhydrous aluminium chloride exists as the dimer, Al_2Cl_6 , which is covalent in nature. This also shows the lower metallic character of aluminium.

The chlorides of silicon and phosphorus are readily hydrolysed in water.

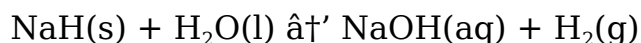


Hydrides

Sodium hydride formed when sodium is heated in dry hydrogen, is an ionic compound. Sodium is such a powerful reducing agent that it forces hydrogen to accept an electron to produce the hydride ion, H^- .

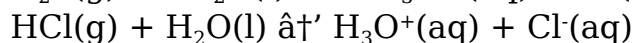
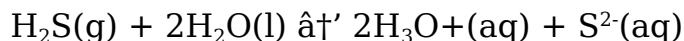


When the molten hydride is electrolysed, hydrogen gas is liberated at the anode. The hydride is decomposed by water, forming an alkaline solution and liberating hydrogen.



Magnesium hydride is unstable while aluminium hydride is essentially covalent. The hydrides of silicon, phosphorus, sulphur and chlorine, i.e. SiH_4 , PH_3 , H_2S and HCl respectively, are covalent compounds. Silane, SiH_4 , which is a solid, catches fire spontaneously in air. The others are gaseous at atmospheric conditions. Only hydrogen sulphide and hydrogen chloride have reactions with water, and this reaction is limited to ionization.

They are thus acidic in water.



1.13 Explanation of the Gradation of Properties in the Horizontal Rows

The electronic configurations of the elements (Table 1.8) reveals that sodium will form an ion, Na^+ , by loss of the single electron in its outermost shell. Magnesium forms Mg^{2+} by loss of the two electrons in its outermost shell. The loss of one electron from magnesium atom gives a positively charged ion, from which it is more difficult to lose a

second electron. Therefore Mg^{2+} is not formed as readily as Na^+ . By the same argument, it is more difficult to form Al^{3+} than Mg^{2+} . Therefore, the reducing action of the metals decreases from sodium to aluminium.

The loss of more than three electrons from a shell is practically very difficult, if not impossible. Therefore, silicon and the other elements following it do not form positive ions. They either gain electrons to form negative ions, or form covalent bonds by sharing electrons. To do this, each atom receives just enough electrons to make it attain the structure of the noble gas at the end of their period. Silicon atom receives four electrons and contributes four for sharing. The outermost shell thus has eight electrons after the sharing. Phosphorus receives three and contributes three for sharing. Two of the five electrons in its outermost shell are not shared. They constitute

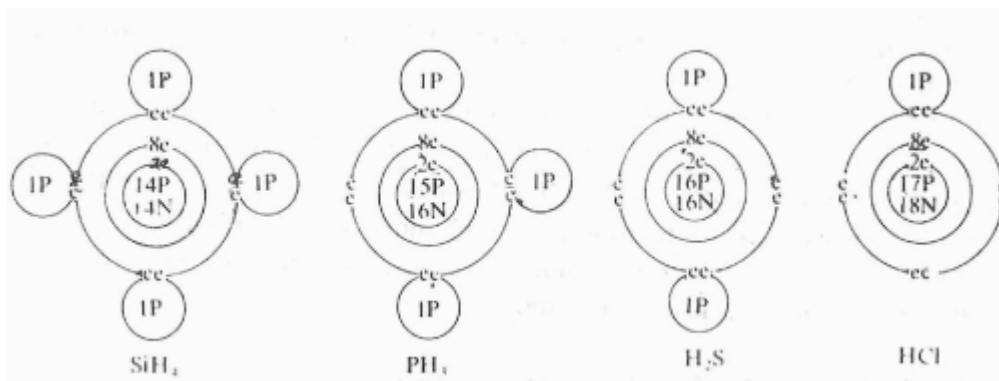


Figure 1.6 Electronic structures of the hydrides of silicon, phosphorus, sulphur and chlorine

a lone pair of electrons. Sulphur receives two and contributes two electrons for sharing. Two lone pairs of electrons are left on the valence shell. Chlorine receives only one electron and also contributes only one electron for sharing. Three lone pairs of electrons are left in its outermost shell.

Chlorine and sulphur also take part in electrovalent bond formation. They do this by accepting electrons transferred from metallic atoms. They become negatively charged ions by doing so. Each accepts enough electrons to have an outermost shell of eight electrons, thereby attaining the electronic configuration of the next noble gas. So, chlorine accepts only one electron to become Cl^- , while sulphur accepts two to become S^{2-} .

Hardness of the metals Na to Al

A noteworthy observation of the metals sodium, magnesium and aluminium is that their hardness increases from sodium to aluminium. Sodium is so soft that it can be cut with a knife. Magnesium is harder but the ribbon is so flexible that it can be wound round a pencil, and soft enough to be cut with a pair of scissors. Aluminium is light with high tensile strength. These characteristics make it suitable for use in

aircraft construction.

What is the explanation for this increasing order of hardness with increasing atomic number? To explain this we consider the force holding the atoms of the different metals together, i.e. the metallic bond. Each atom of sodium contributes only one electron to the mobile cloud of electrons which spreads over the cations formed. The sodium ions themselves carry only a single positive charge each. Each magnesium atom contributes two electrons to that metal's electron cloud, while the ion has a double positive charge, Mg^{2+} . Similarly, aluminium which has three electrons in its outermost shell contributes three electrons to its electron cloud, to form Al^{3+} . **The electrostatic forces of attraction between the cations, Na^+ , Mg^{2+} , Al^{3+} , and their respective electron clouds increase with increase in the number of electrons in the electron cloud.** This increases the binding forces between the atoms of the metals as we move from sodium to aluminium. This therefore increases their hardness from sodium to aluminium.

The increasing order of hardness for the rest of the period, from chlorine towards silicon, will be self-explanatory when we discuss molecular lattice structure. Suffice it now to say that chlorine is gaseous because the diatomic molecules exist as discrete molecules. Sulphur and phosphorus exist as S_8 and P_4 molecules which are joined together. Silicon exists as a complex network structure, and is very hard.

Ionization potential (energy) trends

The energy required to remove the most loosely-bound electron from an atom of an element in the gaseous state, is called the ionization potential (energy) of that element. Ionization potential is important for the metals because they form bonds by loss of electrons. It increases from left to right in a given period, and decreases from the top to the bottom of the groups. The reason for this trend is discussed fully in Book 3. Here, we shall only discuss the relationship between ionization potential and atomic size.

The further away an outermost electron is from its atomic nucleus, the easier it is to remove it. As we go down a group, new shells are added, and the outermost electrons get further apart from their nuclei. The ionisation potential of the elements therefore gets smaller as we go down the group.

This is why it is easier to form K^+ ion than it is to form Na^+ ion, and easier to form Na^+ than Li^+ . The reactivities of the metals thus increase down a given group of the periodic table.

The nearer an outermost electron is to the nucleus, the stronger the force binding it to the atom, hence the more difficult it is to remove. From left to right across a period, no new shells are added to the atoms, but the number of protons increases progressively. Thus, the electrons are more attracted and bound to the nuclei across the period. This leads to an increase in the ionization potential as we move

across each period.

TABLE 1.8 SUMMARY OF GRADATION OF PROPERTIES ALONG A PERIOD, ILLUSTRATED WITH PERIOD III

Groups	I	II	III	IV	V	VI	VII	VIII
1. Element	Na	Mg	Al	Si	P	S	Cl	Ar
2. 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
3. Combining power	1	2	3	4	5,3	6,2	7,1	—
4. ←—————	Increasing metallic character						←—————	
5. ←—————	Increasing ease of formation of cations.							←—————
6. ←—————	Increasing reducing power.						←—————	
7. ←—————	Increasing non-metallic character.						←—————	
8. ←—————	Increasing tendency to form anions.						←—————	
9. ←—————	Increasing oxidising power.						←—————	

1.14 Transition Elements

A question that follows from the discussions on the horizontal rows is: Does every horizontal row of the periodic table show the same gradation of properties? The answer is No. The general trend from metallic to non-metallic character and similar gradation in properties among elements in the same period are limited to only the representative elements (elements which fall in the groups I to VIII). The long periods have additional groups of ten elements which behave alike. The resemblances within each of these groups of ten elements is even more than what is found among the alkali metals or the halogens. These elements are called transition elements.

The first series of ten transition elements (scandium to zinc) occurs between calcium and gallium in the fourth period. They are all metals. Their main characteristics are:

1. They are all hard, with high densities and high tensile strengths.
2. They are ductile and malleable, and so are used in construction work.
3. Most of their compounds are coloured. Table 1.9 illustrates this point.

TABLE 1.9: COLOURED IONS OF TRANSITION METALS

Ion	Colour in solid state	Colour in aqueous solution
Cu^{2+}	Blue (SO_4^{2-} , NO_3^-) Green (Cl^- , CO_3^{2-})	Blue Blue
Fe^{2+}	Pale green	Pale green
Fe^{3+}	Brown (Cl^-) Violet (alum) White (SO_4^{2-} , NO_3^-))) Yellow)
Cr^{3+}	Green	Green
Cr^{6+}	Orange	Orange

4. Each one of them forms at least two types of cations by loss of different numbers of electrons from their atoms. Examples are shown in Table 1.10.

Metal	Oxidation State (No. of electrons lost)						
	+1	+2	+3	+4	+5	+6	+7
Copper	Cu_2O	CuO					
Iron		FeCl_2	FeCl_3				
Chromium			Cr_2O_3			$\text{K}_2\text{Cr}_2\text{O}_7$	
Manganese				MnO_2			KMnO_4

TABLE 1.10: OXIDATION STATE OF SOME TRANSITION ELEMENTS

5. They form complex compounds with ions or groups of atoms called ligands which have non-bonding lone pairs of electrons in their outermost shells, e.g $\text{Fe}(\text{CN})_6^{3-}$ and $\text{Cu}(\text{NH}_3)_4^{2+}$
6. They are often used as catalysts as illustrated in Table 1.11.

TABLE 1.11: EXAMPLES OF CATALYTIC ACTION OF TRANSITION ELEMENTS

Process	Catalyst
1. Hydrogenation $\text{-CH=CH-} + \text{H}_2 \rightarrow \text{-CH}_2\text{-CH}_2\text{-}$	Nickel, Platinum, Palladium
2. Contact process $2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{SO}_3(\text{g})$	Platinum, Vanadium(V) oxide
3. Haber process $3\text{H}_2(\text{g}) + \text{N}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$	Iron

TABLE 1.12: VALUES OF FIRST IONIZATION ENERGIES (kJ MOL) OF THE MAIN GROUP ELEMENTS

	GRP I	GRP II	GRP III	GRP IV	GRP V	GRP VI	GRP VII	GRP VIII
PRD 1	H 1312							He 2372
PRD 2	Li 520	Be 900	B 800	C 1086	N 1403	O 1314	F 1681	Ne 2081
PRD 3	Na 495	Mg 738	Al 577	Si 787	P 1060	S 1000	Cl 1255	Ar 1520
PRD 4	K 418	Ca 590	Ga 579	Ge 760	As 946	Se 941	Br 1142	Kr 1350
PRD 5	Rb 403	Sr 549	In 558	Sn 707	Sb 833	Te 869	I 1007	Xe 1170
PRD 6	Cs 374	Ba 502	Tl 81	Pb 82	Bi 83	Po —	At —	Rn —

WORKED EXAMPLE

- Study the chart below. It is a part of the Periodic Table.

1	2	3	4	5	6	7	8
H X	A		Y	B		Z	C

- (a) Classify the elements, X,Y,Z,A,B,C into two groups; metals and non-metals, giving your reasons.
- (b) Which is the most active metal, and which is the most active non-metal?
- (c) Write the formula of the compound A forms with Z, and the compound B forms with Z. Give reasons.
- (d) Which is the most unreactive?
- (e) What is the formula of the hydride of Y?
- (f) Is the hydride in (e) ionic or covalent?

SOLUTION

- (a) X and A are metals because they have one and two electrons respectively in their outermost shells. These outermost electrons are easily lost to form X^+ and A^{2+} ions. Therefore, X and A are typical reducing agents like other metals.

Y, Z, B and C are non-metals. They have four, five, seven and eight electrons respectively in their outermost shells and so cannot form cations by loss of electrons. They form bonds by sharing or gaining electrons, which is characteristic of all non-metals.

- (b) X, in group one, is the most active metal. Group one elements are very active metals because of the ease with which they lose the single electron in their outermost shell.

Z is the most active non-metal. It is a group 7 element. The group 7 elements are active because they gain electrons readily to form negatively charged ions.

- (c) (i) AZ_2

A is a group 2 element and loses two electrons easily to form A^{2+} . Z is a group 7 element and each atom of it gains one electron to form A. Two atoms of Z are therefore required to gain the two electrons lost by an atom of A.

- (ii) BZ_3

B is a group 5 element and needs three electrons to complete its octet. Z is a group 7 element and needs one electron only. One atom of Z provides three electrons for sharing, while three different atoms of Z provide one each.

- (d) C is the most unreactive element. It belongs to group 8, and therefore has a stable electronic configuration. It is a noble gas.

- (e) YH_4

Y is a group 4 element. It requires four hydrogen atoms to share its four outermost electrons.

- (f) It is covalent.

The group 4 elements form bonds by sharing electrons rather than by losing or gaining electrons.

Chapter Summary

The most important use of the Periodic Table is that it summarises a lot of chemical facts, reducing the study of about one hundred and five elements to the study of eight groups of elements. The properties of the groups are summarised below.

Group I (alkali metals)

1. Soft metals.
2. Liberate hydrogen from water.
3. Chlorides are soluble in water.
4. Oxides, hydroxides and trioxocarbonates(IV) form alkaline solutions in water.
5. Trioxocarbonates(IV) are not decomposed by heat.
6. Trioxonitrates(V) are decomposed to dioxonitrates(III) and oxygen.
7. Tetraoxosulphates(VI) are soluble in water.
8. Hydrides are salt-like.
9. Reactivity increases down the group.

Group II (alkaline earth metals)

1. Metals are harder than alkali metals.
2. First two members do not liberate hydrogen from cold water, others do.
3. Aqueous solutions of chlorides are slightly hydrolysed.
4. Oxides liberate heat in water, forming alkaline solutions.
5. Hydroxides are slightly soluble in water.
6. Trioxocarbonates(IV) are insoluble in water. They are decomposed by heat.
7. Tetraoxosulphates(VI) are insoluble in water.
8. Reactivity increases down the group.

Group III

1. Metals are sufficiently hard to be used in construction work.
2. Oxides and hydroxides are amphoteric.
3. Chlorides are hydrolysed in water.

Group IV

1. First two members (carbon and silicon) are non-metals.
2. Oxides of carbon and silicon are acidic; oxides of tin and lead are amphoteric.
3. Chlorides of carbon and silicon are covalent and liquid under

normal conditions; chlorides of others are hydrolysed in water.

4. Carbon and silicon form many stable hydrides; hydrides of others are unstable.

Group V

Not much similarity.

1. Nitrogen and phosphorus are non-metals; antimony is a semi-metal.
2. Nitrogen and phosphorus form gaseous hydrides that are alkaline.
3. The oxides of nitrogen and phosphorus are acidic.

Group VI

Not much similarity.

Group VII

Very similar elements.

1. All are oxidising agents.
2. All are very reactive non-metals.
3. All form compounds with most metals and many non-metals.

Group VIII

1. Similarity is in their unreactive natures.
2. They possess stable electronic configurations which other elements try to attain by forming compounds.

ASSESSMENT

1. Two elements, A and B, have electronic configurations
A = 2,8, 1
B = 2,8,7
 - (i) To which groups of the periodic table do the elements belong?
 - (ii) What is the nature of the hydride of each element? Would these hydrides conduct electricity or not in aqueous solution?
 - (iii) Where would the elements A and B be discharged if the molten hydrides are electrolysed?
2. The letters A to F represent the symbols of the elements whose atomic numbers are indicated after each letter:
 ${}_3\text{A}$, ${}_{10}\text{B}$, ${}_{14}\text{C}$, ${}_{14}\text{D}$, ${}_{19}\text{E}$, ${}_{35}\text{F}$.
Using the letters A to F for the elements, indicate the element(s) which:
 - (i) does/do not form common compounds.
 - (ii) liberate(s) hydrogen from cold water.

- (iii) is/are halogen.
- (iv) are in the same group.
- (v) form(s) an ion by loss of three electrons.
- (vi) is/are in the same group of the periodic table as carbon.

3. Below is a table of the electronegativities of some elements

Atomic Number	3	4	5	6	7	8	9	10	
Electronegativity	1.0	1.5	2.0	2.5	3.0	3.5	4.0		

- (i) Why is no value for the elements with atomic number 10?
 - (ii) Plot a graph of electronegativity against atomic number and comment on the shape of the graph.
 - (iii) Would the shape be the same if you had plotted atomic volume against atomic number? Explain.
4. Carbon and lead are in the same group of the periodic table yet they are widely different in properties. List three similarities which show that they belong to the same group. Why are they different?
5. Explain why argon, atomic mass 39.9 comes before potassium, atomic mass 39.1 in the periodic table.
6. Would you expect the usual gradation of properties (atomic radii, atomic volume, electronegativity, etc) from left to right among the transition elements Sc to Zn? Explain.
7. Why do the transition elements or their ions
- (i) have coloured ions,
 - (ii) show variable oxidation states,
 - (iii) act as catalysts,
 - (iv) form complexes?
8. The table below refers to a portion of the periodic table

I	II	III	IV	V	VI	VII	VIII
Lithium						P	
X						Q	
Y						R	

- (a) Write down the letters corresponding to:
 - (i) the most reactive metal,
 - (ii) the most reactive non-metal.
- (b) Name the family of the elements represented by P, Q and R.
- (c) The elements Q form compounds with lithium and carbon. Write the formula of each compound.
- (d) Which of the compounds in (c) has

- (i) the lower boiling point,
- (ii) the higher solubility in water?

(WAEC)