NOTE: Kindly continue checking the school website during this holiday for more content including practical work items.

ATOMIC STRUCTURE

An atom is the smallest particle of an element that can take part in chemical reactions.

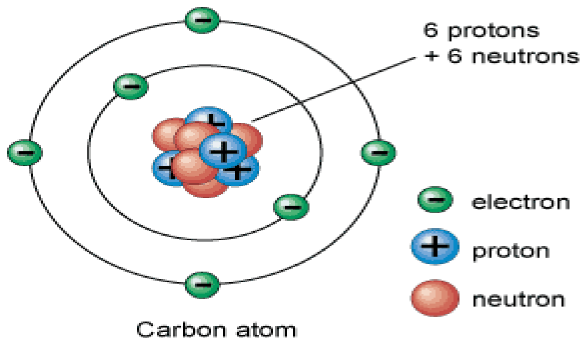
A typical atom is composed of protons, neutrons and electrons.

Protons and neutrons together occupy the core of the atom called the nucleus and they are collectively called nucleons.

Electrons on the other hand, occupy proximal space but in orbits called shells or energy levels.

Only a hudrogen atom lacks neutrons.

THE STRUCTURE OF CARBON-12 ISOTOPE



A proton has a relative mass of 1 and a charge of +1

An electron has a relative mass of approximately zero and a charge of -1.

A neutron has a relative mass of nearly 1(normally considered to be 1) and a charge of zero making it neutral. That's why it is called a neutron.

SUB ATOMIC PARTICLE	RELATIVE MASS	CHARGE
Proton(P)	1	+1
Neutron(n)	1	0
Electron (e)	0	-1

This fact makes the mass of an atom be determined entirely by the nucleus since only the nucleons (protons and neutrons have mass).

In any atom, the number of protons is equal to the number of electrons, making the atom electrically neutral.

ATOMIC NUMBER (Z)

This refers to the number of protons in the nucleus of an atom. It is denoted by Z.

MASS NUMBER (A)

This is the number of protons and neutrons in the nucleus of an atom. It is denoted by A. Mass number is sometimes called **Atomic Mass**.

An atom represented by a symbol X can be fully represented as $_{Z}^{A}X$. eg $_{20}^{40}Ca$, $_{6}^{12}C$, $_{1}^{1}H$.

The number of neutrons can be calculated if atomic number and mass number of an element are known i.e. Mass number A = atomic number(no.of protons) + number of neutrons

i.e A = Z+n

Hence n = A-Z

EXAMPLES

ATOM	ATOMIC NUMBER	NUMBER OF PROTONS	NUMBER OF ELECTRONS	NUMBER OF NEUTRONS	MASS NO
Oxygen, $_{8}^{16}O$,	8	8	8	16-8=8	16
Magnesium, $_{12}^{24}Mg$	12	12	12	24-12=12	24
Sulphur, $_{16}^{32}S$					
Chlorine, $_{17}^{35}Cl$	17	17	17	35-17=18	35
Sodium, $_{11}^{23}Na$					

Complete the rest of the table above.

ELECTRONIC CONFIGURATION

This is the arrangement of electrons in energy levels round the nucleus. Electronic configuration is also called electron structure of an atom.

GUIDELINES TO WRITING ELECTRONIC CONFIGURATION

- The innermost energy level accommodates a maximum of two (2) electrons, while the rest of the energy levels accommodate a maximum of eight (8) electrons.
- Before electrons occupy outer energy levels, inner energy levels must first get filled to their capacity.
- Energy levels are separated using a comma(,) or the symbol)

EXAMPLES.

Write the electronic configuration of the following:

a) Hydrogen (atomic number 1)

Ans. 1, since the hydrogen atom has 1 electron.

b) Sodium (Atomic number 11)

Ans: 2,8,1 or 2)8)1. I encourage you to use the former.

c) Calcium (Atomic number 20)

Ans: 2,8,8,2

d) Chlorine (Atomic number 17)

Ans: 2,8,7

Activity: Draw structures to show electron arrangement for these atoms above.

ION FORMATION

An ion is any particle that carries a charge.

The charge of an ion may be negative in which case this ion has more electrons than protons, and it is called an **ANION.**

The charge may be positive implying that the ion has more protons than electrons, and it is called a **CATION.**

All metals form cations by losing all the electrons in the outermost energy level thus attaining a stable electron structure of a noble/inert gas.

All nonmetal atoms form anions by gaining electrons to fill the outermost energy level to its capacity, attaining an electron structure of a noble/inert gas. Noble gases are inert (unreactive) because they have fully filled energy levels.

Atoms of metals lose electrons which in number are equal to their valency while atoms of non-metals gain electrons which in number are equal to their valencies.

ATOM	**No of protons in the atom**	**No of electrons in the atom**	**Electronic configuration of the atom**	**No. of Protons in the ion**	**No of electrons in the ion**	**Electronic configuration of the ion**	**Formula of the ion**
$^{16}_{8}O$	8	8	2,6	8	10(since 2 are gained)	2, 8	O^{2-}
$^{23}_{11}Na$	11	11	2,8,1	11	10(since 1 is lost)	2,8	Na^{+}
$^{14}_{7}N$	7	7	2,5	7	10(since 3 are gained)	2,8	N^{3-}
$^{27}_{13}Al$	13	13	2,8,3	13	10(since 3 are lost)	2,8	Al^{3+}
$^{39}_{19}K$	19	19	2,8,8,1	19	18(since 1 is lost)	2,8,8	K^{+}

ISOTOPY

Isotopy is the occurance/existence of two or more atoms of the same element having the same atomic number but with different mass numbers resulting from having different numbers of neutrons.

Isotopes therefore, are different atoms of the same element which have the same atomic number but with different mass numbers.

EXAMPLES OF ISOTOPES

ELEMENT	ISOTOPES
Hydrogen	${}^{1}_{1}\boldsymbol{H}$ (Hydrogen), ${}^{2}_{1}\boldsymbol{H}$ ***or*** ${}^{2}_{1}\boldsymbol{D}$ (***Deuterium***), ***and*** ${}^{3}_{1}\boldsymbol{H}$ ***or*** ${}^{3}_{1}\boldsymbol{T}$(***Tritium***). Hydrogen-1 is the mo abundant.
Oxygen	${}^{16}_{8}\boldsymbol{O}$, ${}^{17}_{8}\boldsymbol{O}$ and ${}^{18}_{8}\boldsymbol{O}$, Oxygen-16 is the most abundant
Carbon	${}^{12}_{6}\boldsymbol{C}$, ${}^{13}_{6}\boldsymbol{C}$, ${}^{14}_{6}\boldsymbol{C}$. ${}^{14}_{6}\boldsymbol{C}$ ***is the radioactive isotope of carbon while*** ${}^{12}_{6}\boldsymbol{C}$ ***is the most abundant isotop***
Chlorine	${}^{35}_{17}\boldsymbol{Cl}$, ${}^{36}_{17}\boldsymbol{Cl}$, ${}^{37}_{17}\boldsymbol{Cl}$. Chlorine-35 isotope is the most abundant

Activity: In the remaining space in the table, list other isotopic elements and their isotopes.

THE PERIODIC TABLE

The Periodic Table is the arrangement of elements in order of increasing atomic numbers. The Periodic Table is made up of rows and columns. The horizontal are called **periods** and are assigned numbers according to the number of electron shells/energy levels that are occupied by electrons. All elements in the same period have the same number of electron shells /energy levels. The vertical columns arereferred to as **groups** and the group number is written in **Roman numeral** on top of the group (column).
The transition metals (a strip between groups II and III) belong to no group.

Kindly correct the error in the Periodic Table below as follows:

i. Remove the numbers 3,4,5,….,12 on top of the columns of transition metals ie from the column containing Scandium, Sc to the column containing zinc, Zn.

ii. Change the numbers in Arabic numerals i.e 1,2,13,14,15,16,17,18 on top of the columns to Roman numerals I, II, III, IV,V,VI,VII and VIII respectively. These Roman numerals represent the correct groups of the elements in these columns.

iii. **<u>Note:</u>** Elements in a particular group resemble each other chemically because they have the **same number of electrons** in the outer most shell/energy level.

Periodic table of the elements

- Alkali metals
- Alkaline-earth metals
- Transition metals
- Other metals
- Other nonmetals
- Halogens
- Noble gases
- Rare-earth elements (21, 39, 57–71) and lanthanoid elements (57–71 only)
- Actinoid elements

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

lanthanoid series 6	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
actinoid series 7	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

*Numbering system adopted by the International Union of Pure and Applied Chemistry (IUPAC).

<u>PART OF THE PERIODIC TABLE SHOWING THE FIRST 20 ELEMENTS</u>

Periods \ Groups	I	II	III	IV	V	VI	VII	VIII
1	$^{1}_{1}H$							$^{4}_{2}He$
2	$^{7}_{3}Li$	$^{9}_{4}Be$	$^{11}_{5}B$	$^{12}_{6}C$	$^{14}_{7}N$	$^{16}_{8}O$	$^{19}_{9}F$	$^{20}_{10}Ne$
3	$^{23}_{11}Na$	$^{24}_{12}Mg$	$^{27}_{13}Al$	$^{28}_{14}Si$	$^{31}_{15}P$	$^{32}_{16}S$	$^{35.5}_{17}Cl$	$^{36}_{18}Ar$
4	$^{39}_{19}K$	$^{40}_{20}Ca$						

EXAMPLES

Group I elements	**Atomic Number**	**Electronic configuration**
Lithium	3	2,1
Sodium	11	2,8,1
Potassium	19	2,8,8,1

These metals are ***very reactive*** and are kept under oil in the laboratory where theyhave no contact with water and air. They include **lithium (Li), sodium (Na)** and **potassium (K)**. Atoms of all these elements each have a single electron on the outer most shell, which is easily lost during chemical reactions leaving a single positively charged ion.

All elements in the **same period** have the same number of energy levels/electronshells. The periods are written in **roman numbers** as 1, 2, 3, etc.

Also from one element to the next, across the period, the number of electrons in theouter most energy level increases by one.

Period 3 elements	**Na**	**Mg**	**Al**	**Si**	**P**	**S**	**Cl**	**Ar**
Atomic 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

The elements in the periodic table are classified as metals, metalloid, non-metalsand noble/inert/rare gases.

i) **Metals:** these are on the left of the periodic table in group I, II and III. Metalsreact by loss of electrons.

ii) **Semi metals (metalloids):** These are found in the middle of the periodic tablein group IV.

iii) **Non-metals:** they are on the right of the periodic table in groups V, VI and VII.Non-metals react by gain of electrons.

iv) **Noble/rare or inert gases:** They are at the extreme right of the periodic tablein group VIII (O). Non- metals are not reactive because they are very stable.

Transition elements: these are found between group II and III and towards thebottom of the periodic table. Transition elements form colored compounds andhave variable valencies. Examples of transition elements include copper, iron, chromium, manganese, lead, cobalt etc.

General Trends in the periodic table

1. Elements in the same group react in similar ways because they have the samenumber of electrons in their outer most energy level/ shell.
2. Reactivity increases as you go **down the group** of metals because of **increase** in **atomic size (atomic radius)**. As the size of atoms increase, the outer most electrons become loosely held because they are no longer strongly attracted by the nucleus hence such electrons are easily lost. Because metals react by loss ofelectrons, the more easily a metal loses electrons, the more reactive that metal is.

3. As you go down the group of non-metals, reactivity decreases due to increase in**atomic size (atomic radius)**. As the atomic sizes increase, the forces of attraction of the nucleus for the electrons decrease hence electrons cannot be gained readily. So, the more easily a non-metal gains electrons, the more reactive that non-metal is.

4. **Across the period** of elements from **left to right**, reactivity **decreases** from group I to group IV and **then increases** up to group VII and falls sharply in group VIII. This is because, as you go across the period, the effective force of the nucleusincreases hence the outer most electrons become more strongly attracted.

TRENDS

ATOMIC RADIUS

The atomic radius of an element is the measure of the size of an atom. It is definedas the distance between the center of the nucleus of an atom and its outermost shell.

Variation in atomic radius for elements in the same period

Consider **period 2** elements; **Lithium** has the **least nuclear charge**. Thus its outermost electrons experience the **least nuclear attraction** which results in the**greatest** distance between its nucleus and outermost electrons.

Oxygen has the **highest nuclear charge**, so its outermost electrons experience thegreatest nuclear attraction. This results in the smallest distance between its nucleusand outermost electrons.

Conclusion

Therefore generally, atomic radius decreases across the period from left to right.

Variation in atomic radius down the group

Consider group I elements, Lithium has the smallest atomic radius because itsouter most shell are the least shielded from nuclear attraction.

Potassium has the largest atomic radius because its outermost electrons are themost shielded from the nucleus.

In addition, down the group, an extra shell of electrons is added and these eight electrons increase the repulsive force on outer electrons causing them to be pushed outwards.This causes an increase in atomic size.

Variation in metallic character

Metallic character depends on how readily atoms of elements lose their electrons toform cations **(positively charged ions)**.

Increase in atomic size increases the tendency of metals to lose electrons, hence increasingthe metallic character. This is because when the atomic radius or size is bigger outermost electrons are farther away from the nucleus and therefore the nuclear attraction on these outermost electrons is lower making it easier to remove these outermost electrons, viceversa.

Decrease in atomic size decreases the tendency to lose electrons hence decreasing the metallic character. Metallic character is the tendency of metals (or elements) to lose outermost electrons.

Activity:Complete the table below by stating whether the properties indicated increase ordecrease.

Property	**Trend down the group**	**Trend across the period from left to right**
Atomic size		
Tendency to lose electrons		
Metallic character		

Variation in melting and boiling points of period 3 elements

Some substances require more heat than others to change their physical state(melting). **Why is this so?**

Melting points and boiling points depend on the strengths of forces which existbetween the particles which make up a substance.

Trend

Melting and boiling points increase among metals across the period and generallydecrease among non-metals from left to right.

Consider the table below;

Period 3 elements	**Na**	**Mg**	**Al**	**Si**	**P**	**S**	**Cl**	**Ar**
Atomic number	11	12	13	14	15	16	17	18
Melting point (^{0}C)	97.8	650	660.3	1410	44.1	112.8	-101.5	-189.4
Boiling point (^{0}C)	882.8	1091	2470	2355	280.5	444.6	-34.04	-185.8

Question: On the same axes, draw a graph of melting and boiling point for theelements from Na to Ar against atomic number

Factors which contribute to the above trend in the melting and boiling point

- Number of electrons in the outermost shell (for metals)
- Atomic radius
- Molecular size for non-metals

Explanation

The melting points and boiling points **increase across the three metals (Sodium,Magnesium, and Aluminium)** because of **increasing strength** of the **metallic bonds.**

Silicon adopts a 3-dimensional structure **(giant covalent structure)** with strong interlinking covalent bonds that give it extra stability. These require high energy tobreak hence the exceptionally high melting and boiling points.

Phosphorus, sulphur and chlorine exist as simple molecules held by **Vander Waal's forces of attraction**, whose strength increases with increase in molecularsize. Therefore their melting points are low because their bonds require little energy for melting and boiling to take place.

Density

All group I elements have very low densities. Density **increases** as we go down thegroup because the **increase in atomic mass** is **more than** the **increase in atomic size.** Potassium is an exception to this trend as it has a lower density than sodium.

The density of period 3 elements increases among metals and generally decreasesamong non-metals.

Increase in densities from sodium to Aluminium is due to decrease in atomic radius, which favours formation of stronger metallic bonds, hence closer packingof the atoms.

<u>Summary of the physical properties of Group I elements</u>

1. The metals are soft and can be cut with a knife. The softness decreases downthe group. **(Hardness)**

2. Their **melting and boiling points** are low. The melting point decreases downthe group.

3. They have **low densities** and can float on water.

4. They have shinny surfaces when freshly cut. The shiny surfaces soon tarnish dueto reaction with oxygen to form the oxides.

5. They are good conductors of electricity and heat.

<u>Chemical properties of Group I Metals</u>

All group I elements have one electron in the outer most shell therefore undergosimilar chemical reactions.

1. Reaction with water

Potassium reacts very vigorously, darts on water and burns with a bright purple(lilac) flame producing fumes of hydrogen gas and an alkaline solution of potassium hydroxide.

$$2K\,(s) + 2H_2O \longrightarrow 2KOH\,(aq) + H_2\,(g)$$

Sodium melts into a silvery ball, darts on water reacting vigorously with it producing fumes of hydrogen gas with a hissing sound and an alkaline solution ofsodium hydroxide. The reducing in size as the reaction proceeds until when it finally disappears.

$$2Na_{(s)} + 2H_2O \longrightarrow 2NaOH_{(aq)} + H_{2\,(g)}$$

Lithium reacts slowly with cold water liberating hydrogen gas.

$$2Li_{(s)} + 2H_2O \longrightarrow 2LiOH_{(aq)} + H_{2\,(g)}$$

2. Reaction with air (oxygen)

Potassium burns with a bright purple flame producing white solid of potassium oxide.

$$4K\,(s) + O_2\,(g) \longrightarrow 2\,K_2O$$

Sodium burns in air with a bright yellow flame forming yellow solids of sodiumperoxide in plenty of air.

$$2Na\ (s) + O_2\ (g) \longrightarrow Na_2O_2\ (s)$$

In limited in air, sodium forms white solids of sodium oxide.

$$4Na\ (s) + O2\ (g) \longrightarrow 2\ Na2O\ (s)$$

Lithium burns with a red flame forming white residue of lithium oxide.

$$4Li\ (s) + O_2\ (g) \longrightarrow 2\ Li_2O\ (s)$$

3. Reaction with chlorine

The alkaline metals react with chlorine to form metal chloride salts e.g. Sodium continues to burn in chlorine to form white fumes of sodium chloride.

$2Na_{(s)} + Cl_{2\ (g)} \longrightarrow 2NaCl_{(s)}$ Potassium

burns in chlorine to form fumes of potassium chloride.

$$2K(s) + Cl_2\ (g) \longrightarrow 2KCl(s)$$

<u>Chemical properties of period 3 elements</u>

1. Reaction with water

Sodium vigorously reacts with cold water to produce sodium hydroxide solutionand hydrogen gas.

$$Na(s) + 2H2O\ (l) \longrightarrow 2NaOH(aq) + H_2(g)$$

Magnesium has a very slight reaction with cold water but burns in steam with awhite flame and produces magnesium oxide and hydrogen gas

$$Mg(s) + H_2O\ (g) \longrightarrow MgO(s) + H_2(g)$$

Aluminium also reacts with steam to produce Aluminium oxide and hydrogen gas.

$$2Al(s) + 3H_2O\ (g) \longrightarrow Al_2O_3(s) + 3H_2(g)$$

1. Reaction with oxygen

Note: All elements in Period 3 except chlorine and argon combine directly withoxygen to form oxides.

Sodium burns in oxygen with an orange flame to produce sodium oxide.

$$Na(s) + O_2(g) \longrightarrow 2Na_2O(s)$$

In excess oxygen, sodium burns to produce sodium peroxide.

$$2Na(s) + O_2(g) \longrightarrow Na_2O_2(s)$$

<u>Magnesium</u> burns in oxygen with an intense white flame to form a white solid ofmagnesium oxide.

$$Mg(s) + O_2(g) \longrightarrow 2MgO(s)$$

Powdered **<u>aluminium</u>** burns when sprinkled in a bunsen flame to produce whitesolid of aluminium oxide.

$$Al(s) + 3O_2(g) \longrightarrow 2Al_2O_3(s)$$

<u>Silicon</u> burns in oxygen when heated strongly to form silicon dioxide or silicon(IV) oxide

$$Si(s) + O_2(g) \longrightarrow 2SiO_2(s)$$

<u>Phosphorus</u> burns in oxygen with a white flame to produce phosphorus trioxide (phosphorus (III) oxide)

$$P_4(s) + 3O_2(g) \longrightarrow 2P_2O_3(s)$$

In excess oxygen, **<u>phosphorus</u>** burns to produce phosphorus pentoxide(phosphorus (V) oxide)

$$P_4(s) + 5O_2(g) \longrightarrow 2P_2O_5(s)$$

NOTE:

- **Phosphorus occurs as a tetratomic molecule, P_4**
- **Oxides of phosphorus normally occur as dimers ie P_4O_6 for P_2O_3 and P_4O_{10} for P_2O_5. These dimers are still named as phosphorus trioxide and phosphorus pentoxide respectively.**

<u>Sulphur</u> burns in oxygen on gentle heating with a blue flame to form sulphur dioxide gas

$$S(s) + O_2(g) \longrightarrow 2SO_2(g)$$

<u>Chlorine</u> does not react directly with oxygen. **<u>Argon</u>** does not either.

Note: Sodium and magnesium form **<u>basic oxides</u>**, aluminium forms an **<u>amphoteric oxide</u>** whereas silicon, phosphorus and sulphur form acidic oxides.

END