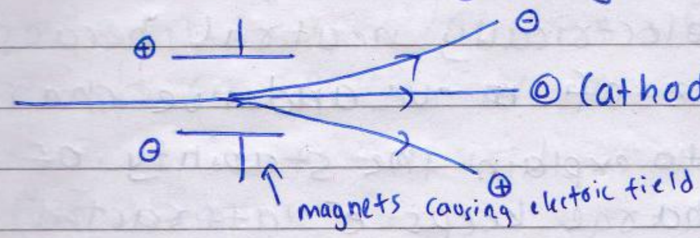
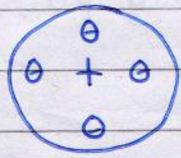
ATOMIC STRUCTURE

- Electrons have a very large charge to mass ratio.



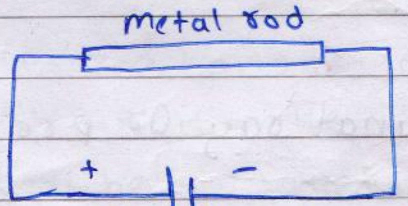
This experiment proved that the rays that deviated towards positive ~~side~~ electric field were electrons, by JJ Thompson. For this work, he was awarded with the Nobel Prize.

- Atom is a huge cloud of positive charge



This is called plum pudding model, where plums represent e^- and pudding centre represents p^+

Cathode ray



In this apparatus, the metal rod emits radiation.

- JJ Thompson discovered electrons, which he calculated must have ~~bodies~~ a very large charge to mass ratio. (e/m value)
- cathode is negative while positive is anode
- The colour of the cathode ray depends on the type of gas present in the cathode ray tube

$\frac{741}{2} = 32\frac{1}{8}$

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- Despite JJ Thompson's discovery of electrons, he only suggested the plum pudding model but was not able to explain other particles. He said that an atom was electrically neutral because of equal magnitude of -ve and +ve charges.
- It also failed to explain the stability of an atom (how positive charge keeps e^- attracted)
- However, it inspired others to work on the field of atomic structure.
- α - particles are He^{2+} (Helium nucleus)

Properties of cathode rays

- They ionize the gas through which they pass
- They travel in a straight line.
- It has energy and momentum.
- It can penetrate graphite layer or paper.
- Properties of cathode rays don't depend on cathode ~~ray~~ composition. Every cathode produces same cathode rays.

Anode ray is also called canal ray or proton ray

Proton Discovery

$$\underset{\alpha\text{-particles}}{{}^{2}_{4}He} + \underset{\text{Nitrogen Atom}}{{}^{7}_{14}N} \longrightarrow \underset{O_2\text{ atom}}{{}^{8}_{27}O} + \underset{\text{Proton}}{{}^{1}_{1}H}$$

This reaction gave out a proton and the proton was discovered.

- Proton is a ~~char~~ positively charged particle having a unit positive charge & mass equal to one H_2 atom

Radioactive material releases its mass in the form of energy spontaneously.

~~e+~~ $e^- = -(1.6 \times 10^{-19})C \approx -1$ [- is not mathematical]

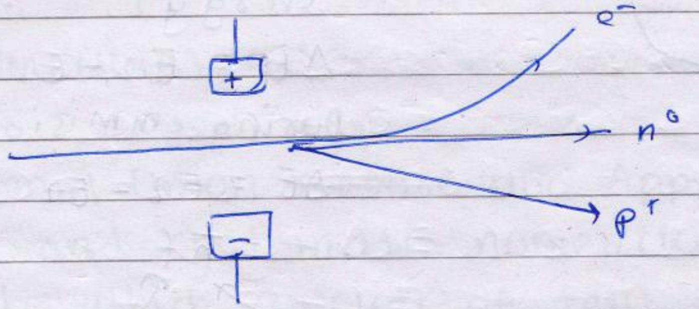
$p^+ = +(1.6 \times 10^{-19})C \approx +1$

Neutron discovery

$^{2}_{4}He + Be \rightarrow {}^{12}C + {}_{0}n^{0}$

α-particles ―――――――― neutrons

This particle had no reaction to an electric field and it was found that the particle was electrically neutral

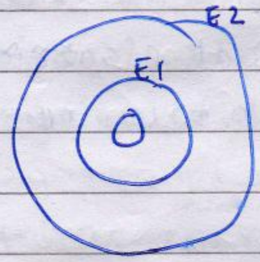


Perforated anode is anode with small, multiple holes in it. The protons (ionized H_2) in the cathode ray tube, travel towards the cathode ~~perforated anode & some of them come out of the holes.~~ (forming anode ray). Cathode is -ve, anode is +ve. (X)

Electron - ~~11~~ 1897

Bohr's Atomic Theory.

- Neils Bohr put his model of the atom in 1915. This is the most widely used at. model and is based on Planck's theory of quantization.
- Th electrons are placed in discrete orbits called stationary orbit.
- Electrons absorb higher energy and jump to quantom levels then emit lower energy.
- Energy is quantized (fixed energy), dosen't change. But naturally quantization of energy is not possible.



$E_1 < E_2$ (It jumps between quantum levels when it absorbs energy)

$\cancel{E_1 > E_2}$

$\Delta E = E_2 - E_1$

During emmision also,

$\Delta E = E_2 - E_2$

$= \hbar f$

planck's constant

frequency

$(6.626 \times 10^{-34}\ Js)$

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EXERCISE FOR ATOMIC STRUCTURE

1) How is hydrogen atom different from atoms of all other elements?

Ans. Hydrogen atom is different from other elements' atoms ~~b~~ in many ways. Firstly it has the least at. weight, it is the simplest with only one proton and one electron and a Hydrogen ion is also only one proton.

2) What is discharge tube?

Ans. Discharge tube is a glass tube which contains charged particles and a gas which is ionized by the charged particles to produce more charged particles. They are mostly used in cathode ray tubes.

3) Name the subatomic particles ~~in~~ present in an atom.

Ans. The subatomic particles present in an atom are electrons (−), protons (+) and neutrons (0)

4) What is an electron? Write its relative mass and charge?

Ans. An electron is a negatively charged sub-atomic particle which revolves around the atomic nucleus. It's relative mass is 0 and relative charge is −1.

6) Find the no. of electrons present in last shell of an atom having at. no. 15?

Ans. Proton number = 15, so Electron ~~um~~ number is also 15. Following K, L, M rule, the electronic configuration becomes 2, 8, 5, so the no. of electrons in the last shell is 5.

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7) What is the basic difference between proton and neutron?

Ans. The basic difference is that proton has a positive charge while neutron has a neutral charge.

8) What is meant by at. no of an element? Does it change when it forms ions! Give an example of diatomic and triatomic molecules.

Ans. At. no of an element means ~~th~~ the ~~p~~ number of protons present in the nucleus of an atom. It does not change when it forms ions as ions only have varying no. of electrons, not protons. A diatomic molecule would be Hydrogen (H_2) and a triatomic molecule would be Ozone (O_3).

9) Define ground state of an atom?

Ans. Ground state is the lowest energy state of an atom in which the electrons are not exicted and don't move between their quantum states or energy levels (don't undergo a quantum transistion).

10) How are cathode rays produced?

Ans. Cathode rays are produced when a cathode / metal block is provided with very high voltages which disintegrates them and releases their atoms in the form of radiation. These cathode rays have electrons, ~~protons and neutrons~~ which undergo deflection in different ways when ~~s~~ near an electric field.

//

11) what are the properties of cathode rays?

Ans. The basic properties are:

- They have electrons in them which are negatively charged.
- They deflect towards the positive terminal when near an electric field (change their trajectory)
- They travel in a straight line unless ~~o~~ near an electric field.
- They ionize certain gases around it.

12) who discovered neutron, electron & proton.

Ans. JJ Thompson discovered electron, Ernest Rutherford discovered proton and James Chadwick discovered neutron.

13) How are anode rays produced?

Ans. When cathode rays strike atoms of certain gases, they knock out their electrons from the outermost shell which ionizes them. These ions formed have a positive charge and are anode rays.

14) What was the reason behind the selection of gold foil by Rutherfood in α-particle scattering?

Ans. Gold foil was selected by Rutherford in α-particle scattering because gold ~~foil~~ is the most malleable metal and its foil can be made very thin so as to allow passage of α-rays through the foil. This proved that an atom was mostly empty space.

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ATOMIC STRUCTURE...

Nanotechnology

- It is the understanding and control of matter at the nanoscale dimensions between approx. 1 to 100 nm. where unique phenomena enable novel applications.
- Don Eigler moved the first individual atom 20 yrs ago, and shortly afterwards, he wrote IBM's name with 35 xenon atoms on a copper plate.

Principal quantum number is the shell ~~of~~ number.

$^{238}_{92}U$ (Where 92 is no. of protons & 238 is no. of protons and neutrons[nuclon number])

Atomic Radius

- The atomic radius is evaluated from the molecular form of elements.
- The at. radius decrease from lft to rt across a period and increase from top to bottom in a group.
- The inner~~most~~ shell is called penultimate shell.
- Shielding effect is when the nuclear force of attraction is less to the outer~~most~~ shells due to the electrons present in the inner shells.
- Atomic radius is larger than ionic radius. for metals which lose e^- but Atomis radius is smaller than ionic radius for non-metals which gain e^- for non-metals which gain e^-.

- Half life is the rate at which a radioactive isotope decays into smaller particles to reduce its mass/concentration to half. Half-lives for various radioisotopes can range from a few microseconds to billions of years.

$$\text{Half life } (T_{1/2}) = \frac{0.693}{\lambda}$$

ELECTRONS, ENERGY LEVELS AND ATOMIC ORBITALS

→ Heisenberg Uncertanity principle says, it is impossible to determine the position & momentum of fermions at the same time.

$\Delta m \cdot \Delta p \geq \hbar / 4\pi$

The more we are certain about one, the less we are about the other.

→ Principal quantum number is simply a name given for sub-shells. It is represented by n.

n	Sub-shells
1	1s
2	2s 2p
3	3s 3p 3d

The s-subshell is spherical.
An orbital is the sub of a subshell ~~a~~ and is a region around which the probability of finding an electron is maximum. Each orbital can only hold 2-electrons. Eg:- ~~p-orb~~ p-subshell has 6 electrons because it has three orbitals. P_x, P_y & P_z.

There are four quantum numbers for ~~partic~~ electrons.
→ Principal quantum number (n)
→ Spin quantum number (m_s)
→ ~~Arithmetical~~ Azimuthal quantum number (l)
→ Magnetic quantum number (m_l)

- Value of azimuthal quantum number (l) varies from 0 to n(principal quantum number). $l = n-1$
No. of subshells

$l = 0$ [s-subshell] $l = 1$ (p-subshell) . . . .

- ~~Number of~~ Magnetic quantum number determines value of orbitals in a subshell. Mathematically
$M_l = \cancel{2\times}\ 2l+1$
For s-subshell
$M_l = 2\times 0+1 = 1$ [1-orbital in s-subshell]

- Spin quantum number $(\pm 1/2)$

- For every electron 4 quantum numbers cannot be the same. This is pauli exclusion principal. If the first three quantum numbers are same for two electrons, they must have opposite spins. Two fermions can't exist in the same quantum states at the same time.

Hund's rule

- This rule describes the order in which electrons are filled in orbitals.

He $1s^2$
$\boxed{\uparrow\downarrow}$

$^{23}_{11}Na$ $1s^2, 2s^2, 2p^6, 3s^1$
$\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}$ $\boxed{\uparrow}$ (This unpaired electron makes Na highly reactive

1. Hydrogen

$1s^1$

$\boxed{\uparrow}$

2. Helium

$1s^2$

$\boxed{\uparrow\downarrow}$

3. Lithium

$1s^2 2s^1$

$\boxed{\uparrow\downarrow}$ $\boxed{\uparrow}$

4. Beryllium

$1s^2 2s^2$

$\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$

5. Boron

$1s^2 2s^2 2p^1$

$\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow}\boxed{}\boxed{}$

6. Carbon

$1s^2 2s^2 2p^2$

$\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow}\boxed{\uparrow}\boxed{}$

7. Nitrogen

$1s^2 2s^2 2p^3$

$\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}$

8. Oxygen

$1s^2 2s^2 2p^4$

$\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}\boxed{\uparrow}\boxed{\uparrow}$

9. Fluorine

$1s^2 2s^2 2p^5$

$\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow \mid \uparrow\downarrow \mid \uparrow}$

10. Neon

$1s^2 2s^2 2p^6$

$\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow \mid \uparrow\downarrow \mid \uparrow\downarrow}$

11. Sodium

$1s^2 2s^2 2p^6 3s^1$

$\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow \mid \uparrow\downarrow \mid \uparrow\downarrow}$ $\boxed{\uparrow}$

12. Magnesium

$1s^2 2s^2 2p^6 3s^2$

$\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow \mid \uparrow\downarrow \mid \uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$

13. Aluminium

$1s^2 2s^2 2p^6 3s^2 3p^1$

$\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow \mid \uparrow\downarrow \mid \uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow \mid \quad \mid \quad}$

14. Silicon

$1s^2 2s^2 2p^6 3s^2 3p^2$

$\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow \mid \uparrow\downarrow \mid \uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow \mid \uparrow \mid \quad}$

15. Phosphorus

$1s^2 2s^2 2p^6 3s^2 3p^3$

$\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow \mid \uparrow\downarrow \mid \uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow \mid \uparrow \mid \uparrow}$

16. Sulfur

$1s^2 2s^2 2p^6 3s^2 3p^4$

$\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow \mid \uparrow\downarrow \mid \uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow \mid \uparrow \mid \uparrow}$

17. Chlorine

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$

$\boxed{\uparrow\downarrow}\quad \boxed{\uparrow\downarrow}\quad \boxed{\uparrow\downarrow|\uparrow\downarrow|\uparrow\downarrow}\quad \boxed{\uparrow\downarrow}\quad \boxed{\uparrow\downarrow|\uparrow\downarrow|\uparrow\ }$

18. Argon

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

$\boxed{\uparrow\downarrow}\quad \boxed{\uparrow\downarrow}\quad \boxed{\uparrow\downarrow|\uparrow\downarrow|\uparrow\downarrow}\quad \boxed{\uparrow\downarrow}\quad \boxed{\uparrow\downarrow|\uparrow\downarrow|\uparrow\downarrow}$

19. Potassium

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$

$\boxed{\uparrow\downarrow}\quad \boxed{\uparrow\downarrow}\quad \boxed{\uparrow\downarrow|\uparrow\downarrow|\uparrow\downarrow}\quad \boxed{\uparrow\downarrow}\quad \boxed{\uparrow\downarrow|\uparrow\downarrow|\uparrow\downarrow}\quad \boxed{\uparrow\ }$

20. Calcium

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$

$\boxed{\uparrow\downarrow}\quad \boxed{\uparrow\downarrow}\quad \boxed{\uparrow\downarrow|\uparrow\downarrow|\uparrow\downarrow}\quad \boxed{\uparrow\downarrow}\quad \boxed{\uparrow\downarrow|\uparrow\downarrow|\uparrow\downarrow}\quad \boxed{\uparrow\downarrow}$

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IONIZATION ENERGY

- It is the quantity of energy required to ionize an atom by removing the loosely bound electrons from the atom.
- The elements must also be in gas state, as the molecules are farther apart.
- Ionization energy is represented as IE.
- First ionization energy is required to remove one-outermost electrom from the atom and this is always the least for every atom.
- If there is huge variation between two consecutive ionization energies, it indicates a change of shell.
- It is an endothermic process
- It is expressed as KJ/mol

$$E(g) + \text{ionization energy} \rightarrow E^{+}(g) + e^{-}$$

Evidence for electronic structure Ionization energy

- Energy needed to remove one electron from each atome in one mole of atoms of the element in gaseous state to form one mole of gaseous $^{1+}$ ions.
- So when we say 1st IE of Hydrogen is 1310 KJ/mol, it means that 1310 kJ is required to ionize 6.02×10^{23} atoms of hydrogen.
- Successive ionization energy refers to the successive energies provided to ionize atoms until only the nucleus is left.
- IE increases across a period and decreases down a grp- Eg: Mg has ~~more~~ less IE

Factors influencing IE

- Size of nuclear charge : As at. no increases, the nuclear charge increases, thus attraction is more. So IE increases with at. no.
- Distance of outer electrons from nucleus : If more distance, more easy to remove the outer electrons.
- Shielding effect of inner electrons : It is the cancellation of nuclear force of attraction to outer electrons due to inner electrons. If it's more, IE will be less.
- Effective nuclear charge (Z_{eff}) : ~~Z_{eff} =~~
  Z_{eff} = Total nuclear charge (Z) - screening constant (S)
  where S takes into account the screening effect of inner electrons. If no. of electrons in inner / penultimate shell is large, screening effect will be more, and attraction will be less. thus less IE
- Spin-pair repulsion : Electrons in same at. orbital in a sub-shell repel each other / which ~~makes~~ makes it easier to remove them.

Mg [↑↓] [↑↓] [↑↓|↑↓|↑↓] [↑↓]
Al [↑↓] [↑↓] [↑↓|↑↓|↑↓] [↑↓] [↑| |]

Easier to remove e^- from Al than from Mg as it is farther and is also unpaired. Thus its easier to remove e^- at high quantum energy level than at low quantum energy level

- Huge variation in IE indicates change in shell.

IONIZATION ENERGY EXERCISE

1. The first IE of sodium is 494 KJ/mol. write the eq^{n} which represents the reaction occuring during first IE of Sodium.

$$Na(g) + IE \longrightarrow Na^{+}(g) + e^{-}$$

~~where IE = 494 KJ/mol,~~

Where IE $= 8.2 \times 10^{-22}$ KJ $\left[{}^{494}/_{6.02 \times 10^{23}} \right]$ "

2. Hydrogen has an electronic structure of $1s^{1}$; helium is $1s^{2}$. Explain why helium's first IE (2370 kJ/mol) is much higher than hydrogen's (1310 KJ/mol).

→ Helium's first IE is much higher than Hydrogen's first IE because the proton number in Helium is twice that of Hydrogen's so the nuclear force of attraction is higher which makes it more difficult to remove the electron. Also the repulsion between the paired electrons is not very ~~promin~~ prominent when compared to the high nuclear force of attraction.

3. One of the factors affecting first IE is nuclear charge. Lithium has 1 more proton than Helium (and greater nuclear charge), and yet its first IE is much lower (519 KJ/mol). Explain why?

→ This is because even though lithium has more nuclear force of attraction, the valence shell is much farther than that of Helium, which makes it easier to remove ~~valence electron~~

outermost electron. Also the net nuclear attraction of outermost electron decreases due to shielding effect of electrons in the inner shells.

4. In both periods 2 and 3, although there are some fluctuations, the main trend is for the first IE to increase across both periods. Explain why?

→ It is so because more electrons are being added to the same ~~sh~~ subshell across a period. This in turn causes stronger nuclear force of attraction as the atomic radius decreases as effect of shielding effect also decreases. Thus it is more difficult to remove electrons across a period which is why the first IE increases.

5. The first IE of each element in Period 3 is less than that ~~of~~ for the corresponding element in Period 2. For eg. Mg in Grp 2 has first IE 736 KJ/mol, wheras Be, above in the same grp has first IE 900 KJ/mol. Using these as example, explain why first IEs in Period 3 are smaller than corresponding ones in Period 2.

Be $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ More IE

Mg $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ Less IE

Here Mg is in Period 3 and has less IE than that of Be which is in Period 2. By their electronic configurations we see that the outermost electrons in Be are very close to the nucleus, so the nuclear force of attraction is more, thus making it harder to remove one e^-, having high IE, while the outermost electrons in Mg are much farther which decreases nuclear force of attraction, making it easier to remove one e^- thus, decreasing IE. Thus, first IEs in period 3 are smaller than corresponding ones in period 2.

6. Explain why Al's first IE (577 KJ/mol) is less than Mg's (736 KJ/mol).

$$\text{Mg} \quad \boxed{\uparrow\downarrow} \quad \boxed{\uparrow\downarrow} \quad \boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow} \quad \boxed{\uparrow\downarrow}$$

$$\text{Al} \quad \boxed{\uparrow\downarrow} \quad \boxed{\uparrow\downarrow} \quad \boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow} \quad \boxed{\uparrow\downarrow} \quad \boxed{\uparrow}\boxed{}\boxed{}$$

It is so because the outermost electron in Al is farther than that of Mg, which makes it easier to be removed due to the lower nuclear force of attraction.

7. Explain why S's first IE (1000KJ/mol) is less than P's (1060 KJ/mol).

$$\text{S} \quad \boxed{\uparrow\downarrow} \quad \boxed{\uparrow\downarrow} \quad \boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow} \quad \boxed{\uparrow\downarrow} \quad \boxed{\uparrow\downarrow}\boxed{\uparrow}\boxed{\uparrow}$$

$$\text{P} \quad \boxed{\uparrow\downarrow} \quad \boxed{\uparrow\downarrow} \quad \boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow} \quad \boxed{\uparrow\downarrow} \quad \boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}$$

It is so because the shielding effect in Sulfur is more due to presence of one more e^- which makes it easier to remove the outermost e^- due to blocking of nulear force of attraction by inner electrons. We can verify this, because the difference in IEs is very less.

IONIZATION ENERGY EXERCISE . . .

8. In the first row of the d-block elements from Scandium to Zinc, except zinc, all have fairly similar first IEs despite no. of protons increasing accross the series. Explain why it is?

→ It is so because despite ~~first IE intrse~~ increase in proton number, the electrons are added to the same d-subshell. As there is no change of subshell or shell, the first IE is fairly similar and there is not some huge variation.

9. Explain why first IE of zinc (908 kJ/mol) is significantly higher than copper's (745 kJ/mol)?

→ It is so because in Zinc there is one extra proton which makes the nuclear force of attraction more, thus making it more difficult to remove the electron in Zinc than copper thus making the first IE more than that of copper's.

10. The second IE for Na is 4560 kJ/mol. Write the eqn which represents the reaction occuring during the second IE of Na.

$$Na^+(g) + IE \longrightarrow Na^{2+}(g) + e^-$$

where $IE = 7.72 \times 10^{-21}$ kJ $\left(4560 / 6.02 \times 10^{23} \right)$

11. The four successive IEs of beryllium are : 900, 1760, 14800, 21000. Why is there a very big gap between the 2nd and 3rd IEs.

Be $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$

→ It is so because the first two e^- are in the same sub-shell, but the third is in 1s subshell which is much closer to the nucleus so the electrons in that subshell experience much higher nuclear force of attraction. This sudden change ~~is~~ in sub-shells and shells is the reason for a very big gap between 2nd & 3rd IEs.

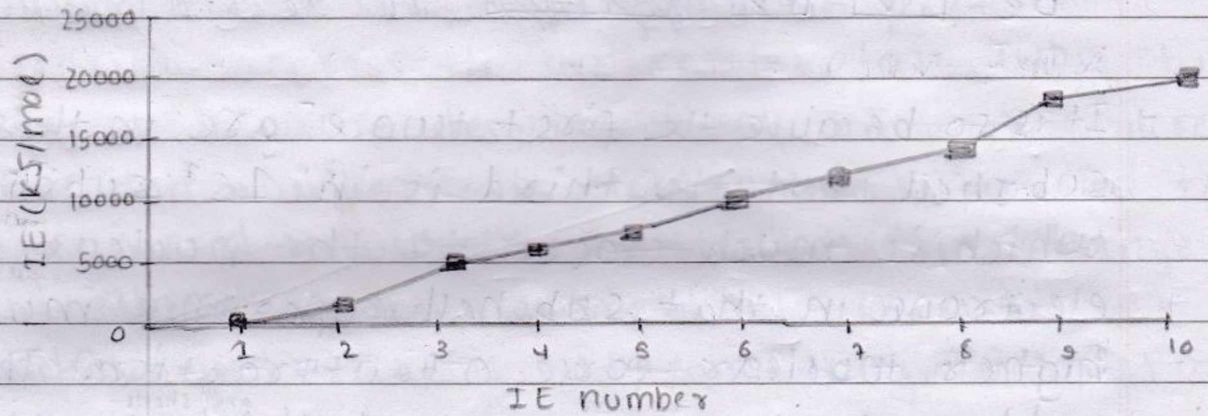
12. The following are all lists of successive IEs for different elements. In each case, decide which grp of the periodic Table the element is to be found in. None of them are from d-block.

a) 799 2420 3660 25000 32800
First three electrons are in same ~~sub~~ shell while next two are in different shell as indicated by the huge variation in IE. Since there are three electrons in one shell, the valency is three and the element lies in Group 3.

b) 736 1450 7740 10500 13600 18000 21700 25600
Similarly, this element lies in Group 2

c) 1000 2260 3390 4540 6990 8490 27100 31700
Similarly, this element lies in Group 6.

13. The graph shows the first 10 IEs for Calcium.



a) Say which electron is being removed at each of these ionizations

Ca [↑↓] [↑↓] [↑↓|↑↓|↑↓] [↑↓] [↑↓|↑↓|↑↓] [↑↓]

10 9   8 7 6 5 4 3   2 1

The numbering represents the electrons removed at the corresponding IE number.

b) Point out and explain the changes in slope in the graph.

- The 1st IE and 2nd IE are very close in values indicating the 1st and 2nd electrons are in the same shell.
- The ~~2nd IE~~ 3rd IE is somewhat larger than the 2nd IE indicating a change in shell.
- The ~~8th~~ slope of 8th IE to 3rd IE is also ~~very~~ pretty much the same indicating not a huge variation in IE which indicates 6 electrons are in the same shell.
- Similarly the slope of 10th to 9th IE also has not a huge variation

→ Then the slope of 9^{th} IE to 8^{th} IE is higher, indicating higher IE, indicating a change in shell.

→ Then the slope of 10^{th} to 9^{th} IE is pretty much the same, indicating two electrons in the same shell.

Questions

a) Define the term first ionization energy.

→ First ionization energy is the amount of energy required to remove one electron from all atoms in one mole of the element in gaseous state to form ~~g~~ one mole of gaseous 1^+ ions.

b) Explain why the first IE of elements down Grp 1 decreases even though at. no. increases

→ It is so because, the elements down group 1 have thier outermost electron in ~~up~~ outer shells. The number of shell also increases as we go down a group. Eg: Outermost e^- of Li is in second shell while that of Na is ~~is~~ in third shell. This causes shielding effect to be more prominent as we go down a group also causing nuclear force of attraction to be less. This is why first IE of elements down Grp I decreases even though at. no. increases.

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c) The eleven succesive ionization energies for Na are given below :-

496 4953 6913 9544 13352 16611 20115
24491 28934 141367 159079

(i) Explain why successive ionization energgies increase
- It is so because the electrons are in lower shells so the nuclear force of attraction is higher for successive electrons. This is why successive IE increases.

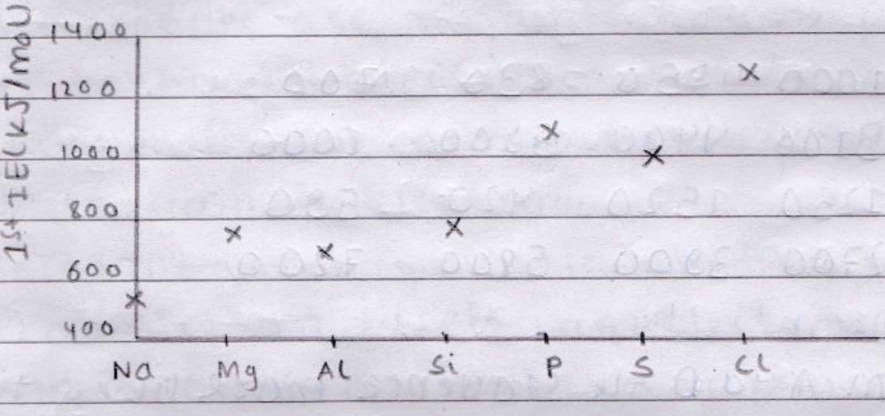
(ii) Explain how these ionization energies give evidence for the electronic structure of Na.
~~You~~
- Firstly we see there is huge variation between first IE and second IE indicating change in shell. Then we see a peculiar variation between 7th and 8th IE indicating another change of shell. Then we see a very large variation betwoen 9th and 10th IE. By this we can know that ~~8th and~~ 9th and 10th electron have to be in same shell, and one becomes unpaired when another is removed.

d) ~~(iii)~~ The first IE of aluminium is lower than Magnesium.

i) Give the electronic structures of Mg and Al in s, p and d notation.

Mg $1s^2\ 2s^2\ 2p^6\ 3s^2$
Al $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$

2. The diagram shows the first IE of some period 3 elements.



2a) Draw a cross on the diagram to show the first IE of aluminium.

b) Write an eqn to show process that occurs when first IE of aluminium is measured.

$\cancel{Al + IE}$ $Al(g) + IE \longrightarrow Al^{+}(g) + e^{-}$

c) State which of the first, second or third Ionizations of Al would produce an ion with the configuration $1s^2 2s^2 2p^6 3s^2$

$Al \; 1s^2 2s^2 2p^6 3s^2 3p^1$
$Al^{2+} \; 1s^2 2s^2 2p^6 3s^2$
$\therefore$ Second ionization of Al would produce the ion.

d) Explain why Sulfur's first IE is less than Phosphorus's first IE.
Sulfur's first IE is less because Sulfur has an extra e^- and has one electron pair in the 3p subshell. Thus the e^- repulsion and screening effect is more making it easier to remove the e^-.

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9. Question 9 is about the following IE sequences. The values are all in KJ/mol.

A	1400	1000	950	830	700
B	420	3100	4400	5900	8000
C	1000	1250	1520	420	590
D	1520	2700	3900	5800	7200

Select from A to D the sequence more likey to represent:-

a) The first IE of five consecutive members of the same group in the periodic table in order of increasing at. no.
→ A

b) The first five IEs of an s-block element
→ B

c) The first five IEs of a noble gas
→ D

Paired electrons repulse unpaired electrons in the same sub-shell. Thus, it is more easier to remove unpaired electrons than paired ones.