

PROSPERITY ACADEMY

AS CHEMISTRY 9701

Crash Course

RUHAB IQBAL

ATOMIC STRUCTURE

COMPLETE NOTES



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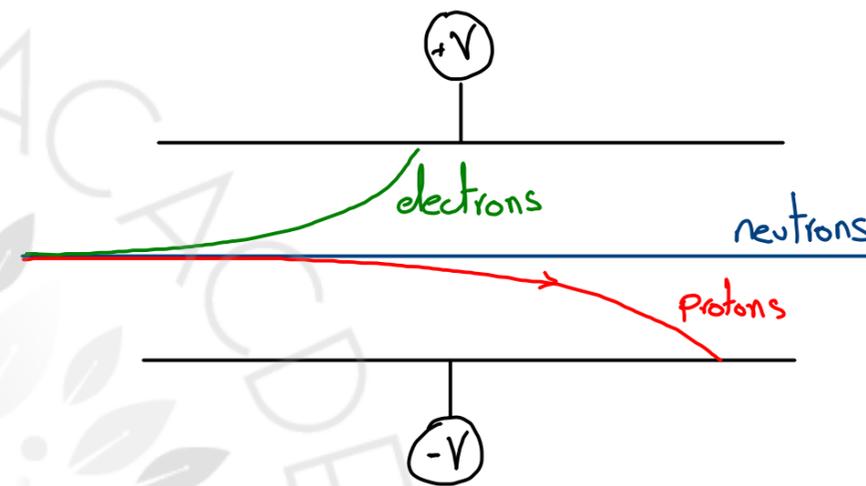


Atomic Structure

Subatomic particles:-

	Relative charge	Relative mass
protons	+1	1
electrons	0	1
neutrons	-1	$\frac{1}{1840} \approx 0$

* Relative to a proton



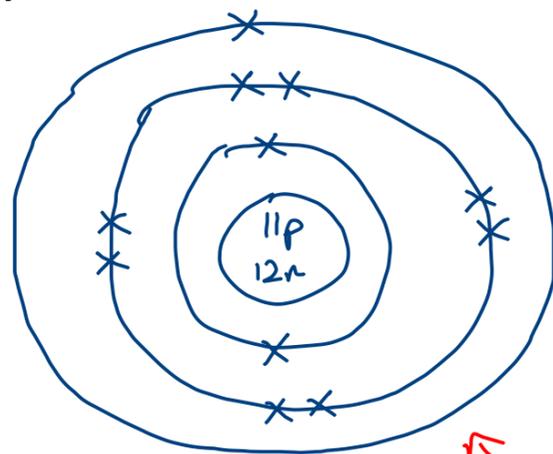
The greater the charge/mass ratio, the more the deflection

- electrons deflected most
- protons deflected less
- neutrons undeflected

Electronic Structure:-

In O-levels:-

Sodium:-



Bohr model

This does not explain trend in ionization energies (discussed later) and atomic emission and absorption spectrum (Physics A2 - Quantum Physics)

This is a simplistic model, but we needed a better understanding, hence - the Quantum model.

Quantum model:-

- Electrons are arranged in shells, $n=1, 2, 3$ and so on. (n is known as principal quantum number)
- The shells increase in energy with increasing number (with some exceptions)
- Each shell has subshells equal to its principal quantum number
- The shells are named s, p, d, f
- Each shell has one or more orbitals, s has 1 orbital, p has 3, d has 5, f has 7 (Goes up in odd number)

Shell	No. of Subshells	Names of subshells	Orbitals present
$n=1$	1	$1s$	$1s$
$n=2$	2	$2s, 2p$	$2s, 2p_x, 2p_y, 2p_z$
$n=3$	3	$3s, 3p, 3d$	$3s, 3p_x, 3p_y, 3p_z, (5 \times 3d)$
$n=4$	4	$4s, 4p, 4d, 4f$	$4s, 4p_x, 4p_y, 4p_z, (5 \times 4d)$ and $(7 \times 4f)$

- Electrons have a property known as spin:- Spin up: \uparrow or \uparrow / Spin down: \downarrow or \downarrow
- Electrons fill in order of increasing energies (Work out through periodic table)
- An orbital can only have maximum 2 electrons. If 2 electrons are present they must be opposite spin $\uparrow\downarrow$
- If orbitals of equal energy are available, the orbitals will each fill up with 1 electron before they start filling up with 2.

The Periodic Table of Elements

9 The Periodic Table of Elements

		Group										13	14	15	16	17	18	
1	2											13	14	15	16	17	18	
s block 3 Li lithium 6.9 4 Be beryllium 9.0 11 Na sodium 23.0 12 Mg magnesium 24.3		Key atomic number atomic symbol name relative atomic mass										1 H hydrogen 1.0 → (Also in s block) ←	p-block					2 He helium 4.0
		d-block (takes the number of previous period)										5 B boron 10.8 13 Al aluminium 27.0	6 C carbon 12.0 14 Si silicon 28.1	7 N nitrogen 14.0 15 P phosphorus 31.0	8 O oxygen 16.0 16 S sulfur 32.1	9 F fluorine 19.0 17 Cl chlorine 35.5	10 Ne neon 20.2 18 Ar argon 39.9	
19 K potassium 39.1	20 Ca calcium 40.1	21 Sc scandium 45.0	22 Ti titanium 47.9	23 V vanadium 50.9	24 Cr chromium 52.0	25 Mn manganese 54.9	26 Fe iron 55.8	27 Co cobalt 58.9	28 Ni nickel 58.7	29 Cu copper 63.5	30 Zn zinc 65.4	31 Ga gallium 69.7	32 Ge germanium 72.6	33 As arsenic 74.9	34 Se selenium 79.0	35 Br bromine 79.9	36 Kr krypton 83.8	
37 Rb rubidium 85.5	38 Sr strontium 87.6	39 Y yttrium 88.9	40 Zr zirconium 91.2	41 Nb niobium 92.9	42 Mo molybdenum 95.9	43 Tc technetium -	44 Ru ruthenium 101.1	45 Rh rhodium 102.9	46 Pd palladium 106.4	47 Ag silver 107.9	48 Cd cadmium 112.4	49 In indium 114.8	50 Sn tin 118.7	51 Sb antimony 121.8	52 Te tellurium 127.6	53 I iodine 126.9	54 Xe xenon 131.3	
55 Cs caesium 132.9	56 Ba barium 137.3	57-71 lanthanoids	72 Hf hafnium 178.5	73 Ta tantalum 180.9	74 W tungsten 183.8	75 Re rhenium 186.2	76 Os osmium 190.2	77 Ir iridium 192.2	78 Pt platinum 195.1	79 Au gold 197.0	80 Hg mercury 200.6	81 Tl thallium 204.4	82 Pb lead 207.2	83 Bi bismuth 209.0	84 Po polonium -	85 At astatine -	86 Rn radon -	
87 Fr francium -	88 Ra radium -	89-103 actinoids	104 Rf rutherfordium -	105 Db dubnium -	106 Sg seaborgium -	107 Bh bohrium -	108 Hs hassium -	109 Mt meitnerium -	110 Ds darmstadtium -	111 Rg roentgenium -	112 Cn copernicium -	113 Nh nihonium -	114 Fl flerovium -	115 Mc moscovium -	116 Lv livermorium -	117 Ts tennessine -	118 Og oganesson -	

1
2
3
4
5
6
7

lanthanoids

57 La lanthanum 138.9	58 Ce cerium 140.1	59 Pr praseodymium 140.9	60 Nd neodymium 144.4	61 Pm promethium -	62 Sm samarium 150.4	63 Eu europium 152.0	64 Gd gadolinium 157.3	65 Tb terbium 158.9	66 Dy dysprosium 162.5	67 Ho holmium 164.9	68 Er erbium 167.3	69 Tm thulium 168.9	70 Yb ytterbium 173.1	71 Lu lutetium 175.0
89 Ac actinium -	90 Th thorium 232.0	91 Pa protactinium 231.0	92 U uranium 238.0	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 -

actinoids

f-block (not important)
(takes the number of 2 periods before)

Electronic Configuration of elements upto atomic number 30.

Element	Electronic Configuration	Arrangement
1) H ₁	1s ¹	$\begin{array}{c} 1s \\ \boxed{1} \end{array}$
2) He ₂	1s ²	$\begin{array}{c} 1s \\ \boxed{1\downarrow} \end{array}$
3) Li ₃	1s ² 2s ¹	$\begin{array}{cc} 1s & 2s \\ \boxed{1\downarrow} & \boxed{1} \end{array}$
4) Be ₄	1s ² 2s ²	$\begin{array}{cc} 1s & 2s \\ \boxed{1\downarrow} & \boxed{1\downarrow} \end{array}$
5) B ₅	1s ² 2s ² 2p ¹	$\begin{array}{ccc} 1s & 2s & 2p \\ \boxed{1\downarrow} & \boxed{1\downarrow} & \boxed{1} \quad \boxed{} \quad \boxed{} \end{array}$
6) C ₆	1s ² 2s ² 2p ²	$\begin{array}{ccc} 1s & 2s & 2p \\ \boxed{1\downarrow} & \boxed{1\downarrow} & \boxed{1} \quad \boxed{1} \quad \boxed{} \end{array}$
7) N ₇	1s ² 2s ² 2p ³	$\begin{array}{ccc} 1s & 2s & 2p \\ \boxed{1\downarrow} & \boxed{1\downarrow} & \boxed{1} \quad \boxed{1} \quad \boxed{1} \end{array}$
8) O ₈	1s ² 2s ² 2p ⁴	$\begin{array}{ccc} 1s & 2s & 2p \\ \boxed{1\downarrow} & \boxed{1\downarrow} & \boxed{1\downarrow} \quad \boxed{1} \quad \boxed{1} \end{array}$
9) F ₉	1s ² 2s ² 2p ⁵	$\begin{array}{ccc} 1s & 2s & 2p \\ \boxed{1\downarrow} & \boxed{1\downarrow} & \boxed{1\downarrow} \quad \boxed{1\downarrow} \quad \boxed{1} \end{array}$
10) Ne ₁₀	1s ² 2s ² 2p ⁶	$\begin{array}{ccc} 1s & 2s & 2p \\ \boxed{1\downarrow} & \boxed{1\downarrow} & \boxed{1\downarrow} \quad \boxed{1\downarrow} \quad \boxed{1\downarrow} \end{array}$

Electronic Configuration of elements upto atomic number 30.

Element

Electronic Configuration

Arrangement

Element	Electronic Configuration	1s	2s	2p	3s	3p	4s
11) Na ₁₁	1s ² 2s ² 2p ⁶ 3s ¹	↑↓	↑↓	↑↓ ↑↓ ↑↓	↑		
12) Mg ₁₂	1s ² 2s ² 2p ⁶ 3s ²	↑↓	↑↓	↑↓ ↑↓ ↑↓	↑↓		
13) Al ₁₃	1s ² 2s ² 2p ⁶ 3s ² 3p ¹	↑↓	↑↓	↑↓ ↑↓ ↑↓	↑↓	↑	
14) Si ₁₄	1s ² 2s ² 2p ⁶ 3s ² 3p ²	↑↓	↑↓	↑↓ ↑↓ ↑↓	↑↓	↑ ↑	
15) P ₁₅	1s ² 2s ² 2p ⁶ 3s ² 3p ³	↑↓	↑↓	↑↓ ↑↓ ↑↓	↑↓	↑ ↑ ↑	
16) S ₁₆	1s ² 2s ² 2p ⁶ 3s ² 3p ⁴	↑↓	↑↓	↑↓ ↑↓ ↑↓	↑↓	↑↓ ↑ ↑	
17) Cl ₁₇	1s ² 2s ² 2p ⁶ 3s ² 3p ⁵	↑↓	↑↓	↑↓ ↑↓ ↑↓	↑↓	↑↓ ↑↓ ↑	
18) Ar ₁₈	1s ² 2s ² 2p ⁶ 3s ² 3p ⁶	↑↓	↑↓	↑↓ ↑↓ ↑↓	↑↓	↑↓ ↑↓ ↑↓	
19) K ₁₉	1s ² 2s ² 2p ⁶ 3s ² 3p ⁶ 4s ¹	↑↓	↑↓	↑↓ ↑↓ ↑↓	↑↓	↑↓ ↑↓ ↑↓	↑
20) Ca ₂₀	1s ² 2s ² 2p ⁶ 3s ² 3p ⁶ 4s ²	↑↓	↑↓	↑↓ ↑↓ ↑↓	↑↓	↑↓ ↑↓ ↑↓	↑↓

Electronic Configuration of elements upto atomic number 30.

Element	Electronic Configuration	Arrangement						
21) Sc ₂₁	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^1$	1s 1↓	2s 1↓	2p 1↓ 1↓ 1↓	3s 1↓	3p 1↓ 1↓ 1↓	4s 1↓	3d 1
22) Ti ₂₂	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$	1s 1↓	2s 1↓	2p 1↓ 1↓ 1↓	3s 1↓	3p 1↓ 1↓ 1↓	4s 1↓	3d 1 1
23) V ₂₃	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$	1s 1↓	2s 1↓	2p 1↓ 1↓ 1↓	3s 1↓	3p 1↓ 1↓ 1↓	4s 1↓	3d 1 1 1
24) Cr ₂₄ *	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$	1s 1↓	2s 1↓	2p 1↓ 1↓ 1↓	3s 1↓	3p 1↓ 1↓ 1↓	4s 1	3d 1 1 1 1 1
25) Mn ₂₅	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^5$	1s 1↓	2s 1↓	2p 1↓ 1↓ 1↓	3s 1↓	3p 1↓ 1↓ 1↓	4s 1↓	3d 1 1 1 1 1
26) Fe ₂₆	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$	1s 1↓	2s 1↓	2p 1↓ 1↓ 1↓	3s 1↓	3p 1↓ 1↓ 1↓	4s 1↓	3d 1↓ 1 1 1 1
27) Co ₂₇	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^7$	1s 1↓	2s 1↓	2p 1↓ 1↓ 1↓	3s 1↓	3p 1↓ 1↓ 1↓	4s 1↓	3d 1↓ 1↓ 1 1 1
28) Ni ₂₈	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^8$	1s 1↓	2s 1↓	2p 1↓ 1↓ 1↓	3s 1↓	3p 1↓ 1↓ 1↓	4s 1↓	3d 1↓ 1↓ 1↓ 1 1
29) Cu ₂₉ *	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$	1s 1↓	2s 1↓	2p 1↓ 1↓ 1↓	3s 1↓	3p 1↓ 1↓ 1↓	4s 1	3d 1↓ 1↓ 1↓ 1↓ 1↓
30) Zn ₃₀	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$	1s 1↓	2s 1↓	2p 1↓ 1↓ 1↓	3s 1↓	3p 1↓ 1↓ 1↓	4s 1↓	3d 1↓ 1↓ 1↓ 1↓ 1↓

→ Very important

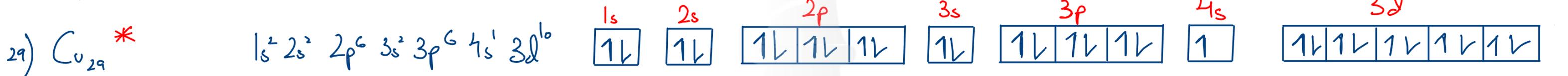
→ Very important

Shorthand Configuration:-

Instead of writing the whole configuration, you can quote the noble gas before the element and start from there.

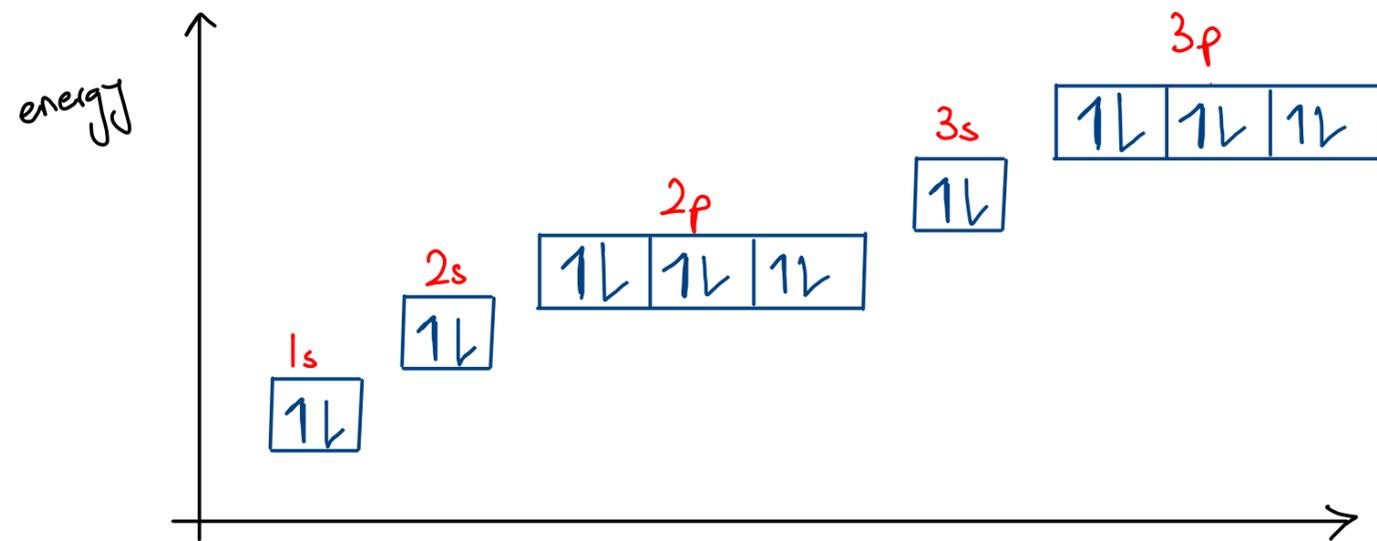
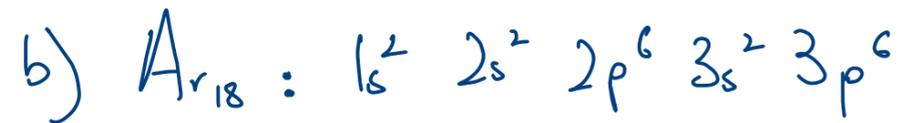
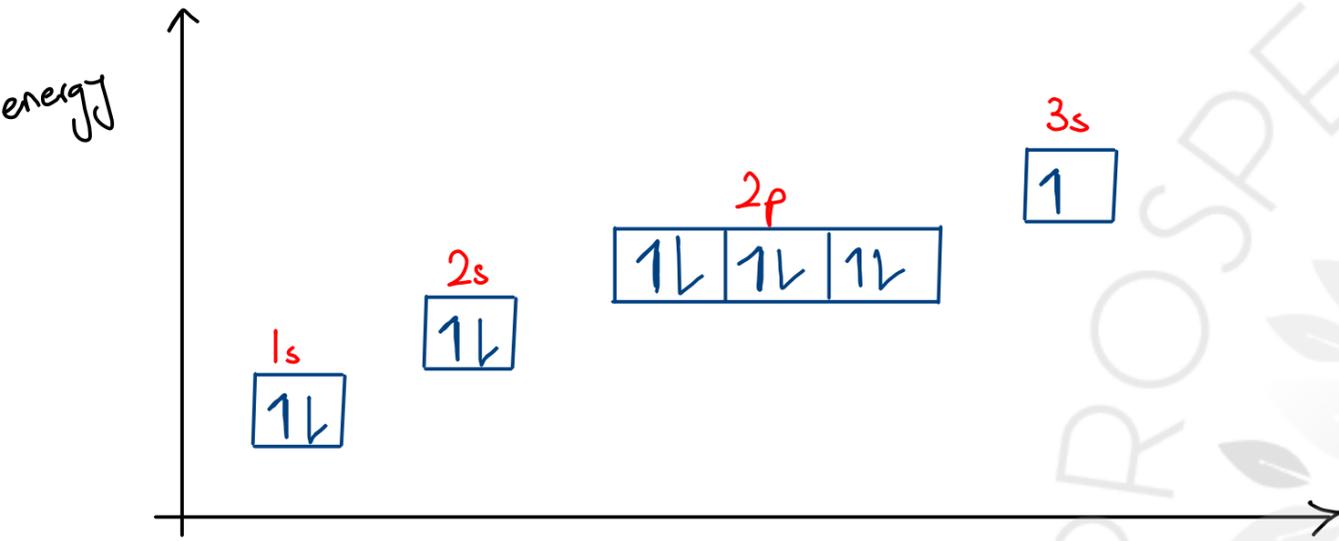


* Anomalous Configurations of Cr (Chromium) and Cu (Copper):-



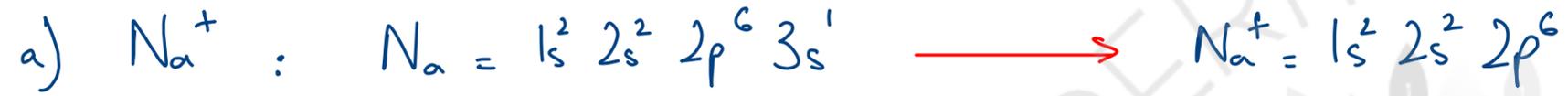
A half or completely filled d orbital is more stable and so an electron is transferred from 4s to 3d.

Electronic Configuration on a graph:-



Electronic configuration of ions:-

Remove or add electrons as shown by charge.



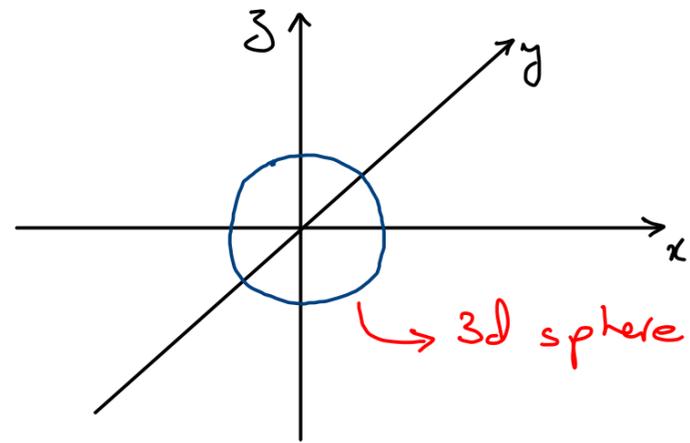
* For elements in d-block / transition metals, remove electrons from 4s first and then 3d.

Using Short hand Configuration.



Shape of s orbital:-

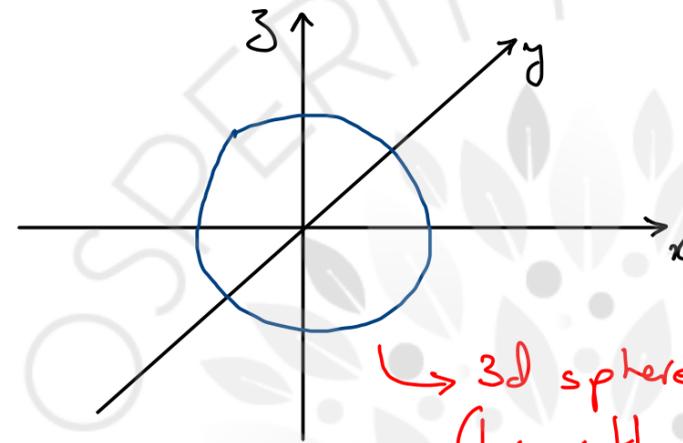
1s orbital:-



↳ 3d sphere

higher in energy

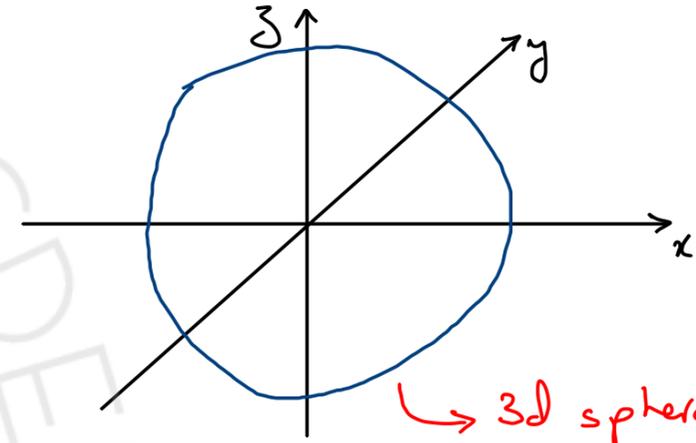
2s orbital:-



↳ 3d sphere (larger than 1s)

higher in energy

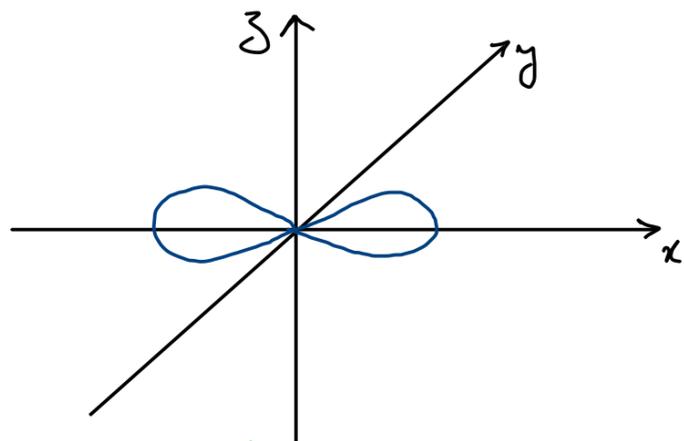
3s orbital:-



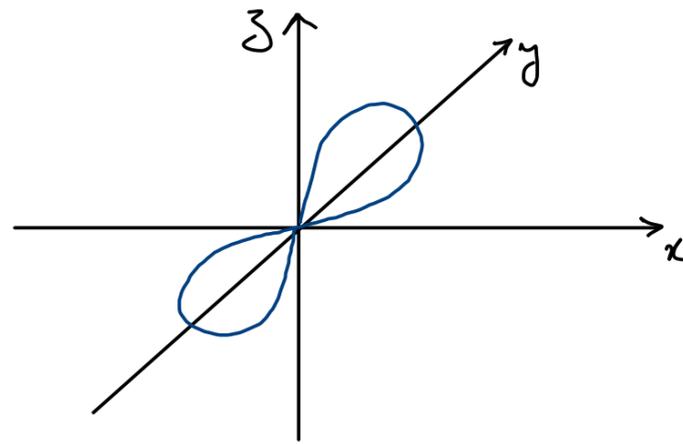
↳ 3d sphere (larger than 2s)

Shape of p orbital:-

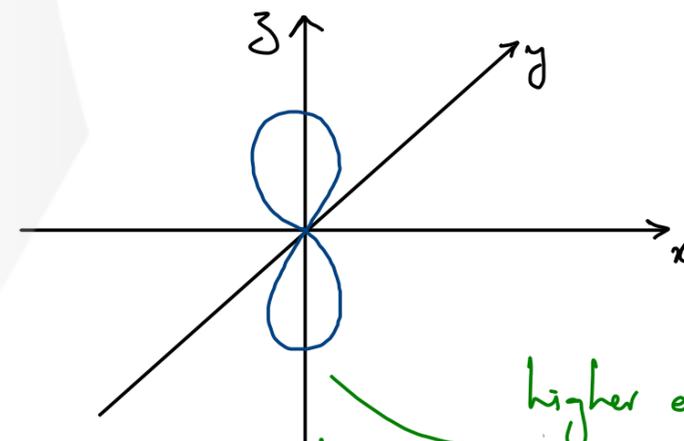
2p_x orbital:-



2p_y orbital:-

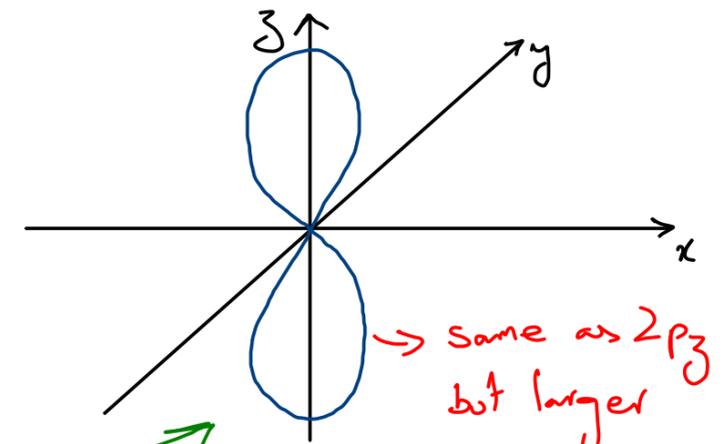


2p_z orbital:-



higher energy

3p_z orbital:-

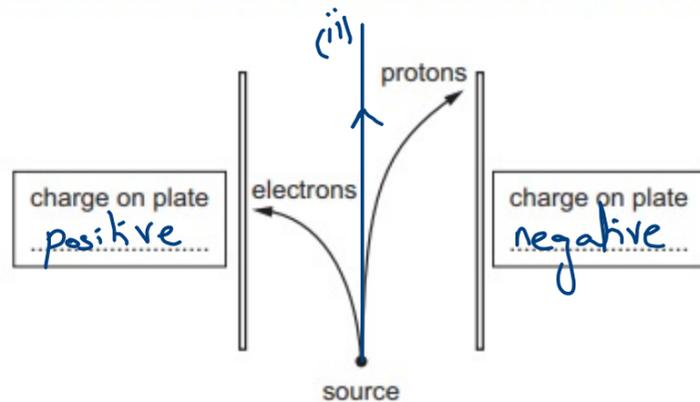


↳ same as 2p_z but larger

equal energy

1 Atoms contain the subatomic particles electrons, protons and neutrons. Protons and electrons were discovered by observations of their behaviours in electric fields.

(a) The diagram shows the behaviour of separate beams of electrons and protons in an electric field.



(i) Complete the diagram with the relative charge of each of the electrically charged plates. [1]

(ii) On the diagram, draw a line to show how a separate beam of neutrons from the same source behaves in the same electric field. [1]

(b) Electrons in atoms up to $_{36}\text{Kr}$ are distributed in s, p and d orbitals.

(i) State the number of occupied orbitals in an isolated atom of $_{36}\text{Kr}$. : $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$

type of orbital	s	p	d
number of orbitals	4	9	5

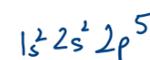
[3]

(ii) Complete the diagram to show the number and relative energies of the electrons in an isolated atom of $_{14}\text{Si}$.



[2]

(iii) The diagram shows a type of orbital.



State the total number of electrons that exist in all orbitals of this type in an atom of $_{9}\text{F}$.

5 electrons

[1]

(iv) The first ionisation energies of elements in the first row of the d block ($_{21}\text{Sc}$ to $_{29}\text{Cu}$) are very similar. For all these elements, it is a 4s electron that is lost during the first ionisation.

Suggest why the first ionisation energies of these elements are very similar.

The nuclear charge increases and the shielding also increases as electrons increase in 3d subshell across the period and therefore cancel each other out. The atomic radius therefore remains roughly constant. [3]

(c) Hydron is a general term used to represent the ions $^1\text{H}^+$, $^2\text{H}^+$ and $^3\text{H}^+$.

State, in terms of subatomic particles in the nucleus, what is the same about each of these ions and what is different.

same proton number

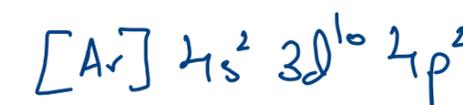
different neutron number / mass number. [1]

[Total: 12]

11 Germanium has the electronic configuration $[\text{Ar}] 3d^{10} 4s^2 4p^2$, where $[\text{Ar}]$ represents the configuration of argon.

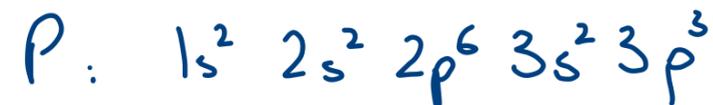
In which order are the electrons lost in forming the Ge^{4+} ion?

	1st	2nd	3rd	4th
A	4p ✓	4p ✓	4s ✓	4s ✓
B	4p	4p	3d	3d
C	4s	4s	4p	4p
D	3d	4s	4p	3d



3 Which atomic orbitals are occupied in an atom of phosphorus?

- A $1p2s2p$ ~~X~~ B $2s2p2d$ ~~X~~ C $2s2p3s$ ✓ D $2p3s3d$ ~~X~~



31 The symbol for a phosphorus ion is ${}^{33}_{15}\text{P}^{3-}$.

$33 - 15 = 18$

The symbol for a potassium ion is ${}^{39}_{19}\text{K}^+$.

$37 - 19 = 18$

What do these two ions have in common?

- 1 the same number of electrons ✓
 2 the same number of neutrons ✓
 3 the same number of protons ~~X~~

2 The electronic configuration of an atom of sulfur is $1s^2 2s^2 2p^6 3s^2 3p^4$.

How many valence shell and unpaired electrons are present in one sulfur atom?

	valence shell electrons	unpaired electrons
A	2	1
B	4	2
C	6 ✓	0
D	6 ✓	2 ✓



1 Which particle has equal numbers of protons and neutrons and an electronic structure of $1s^2 2s^2 2p^6 3s^2 3p^6$?

- A ${}^{39}_{18}\text{Ar}$ ~~X~~ B ${}^{40}_{20}\text{Ca}^{2+}$ ✓ C ${}^{16}_8\text{O}^{2-}$ ~~X~~ D ${}^{32}_{16}\text{S}$ ~~X~~

8 What is the electronic configuration of an isolated Ni^{2+} ion?

- A $1s^2 2s^2 2p^6 3s^2 3p^6 3d^6 4s^2$
 B $1s^2 2s^2 2p^6 3s^2 3p^6 3d^7 4s^1$
 C $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2$
 D $1s^2 2s^2 2p^6 3s^2 3p^6 3d^8$

2 In which pair of species do both species have only one unpaired p electron?

- A Ar^+ and C^- ~~X~~ B B and Ti^+ C F and Ga D Se^- and Si^-



Ionization Energies

First ionization energy:- The minimum amount of energy required to remove 1 mole of electrons from 1 mole of gaseous atoms to form 1 mole of $1+$ charged gaseous ions.

e.g.



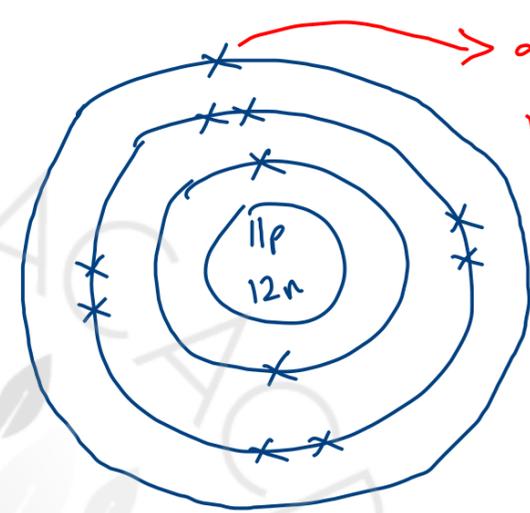
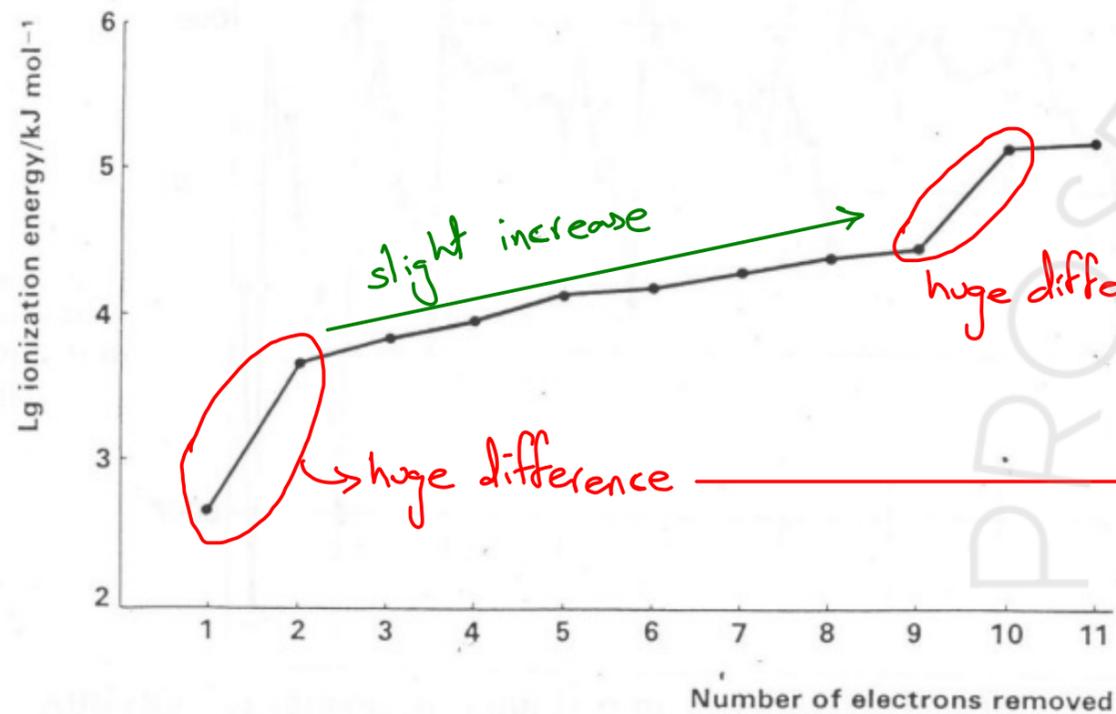
Second ionization energy:- The minimum amount of energy required to remove 1 mole of electrons from 1 mole of $1+$ charged gaseous ions to form 1 mole of $2+$ charged gaseous ions.



and so on for other ionization energies.....

Successive ionization energies of a single element:-

Ionization energies of Sodium:-



after this electron is lost, you start removing electrons from inner shell, which is closer to nucleus, so more energy required

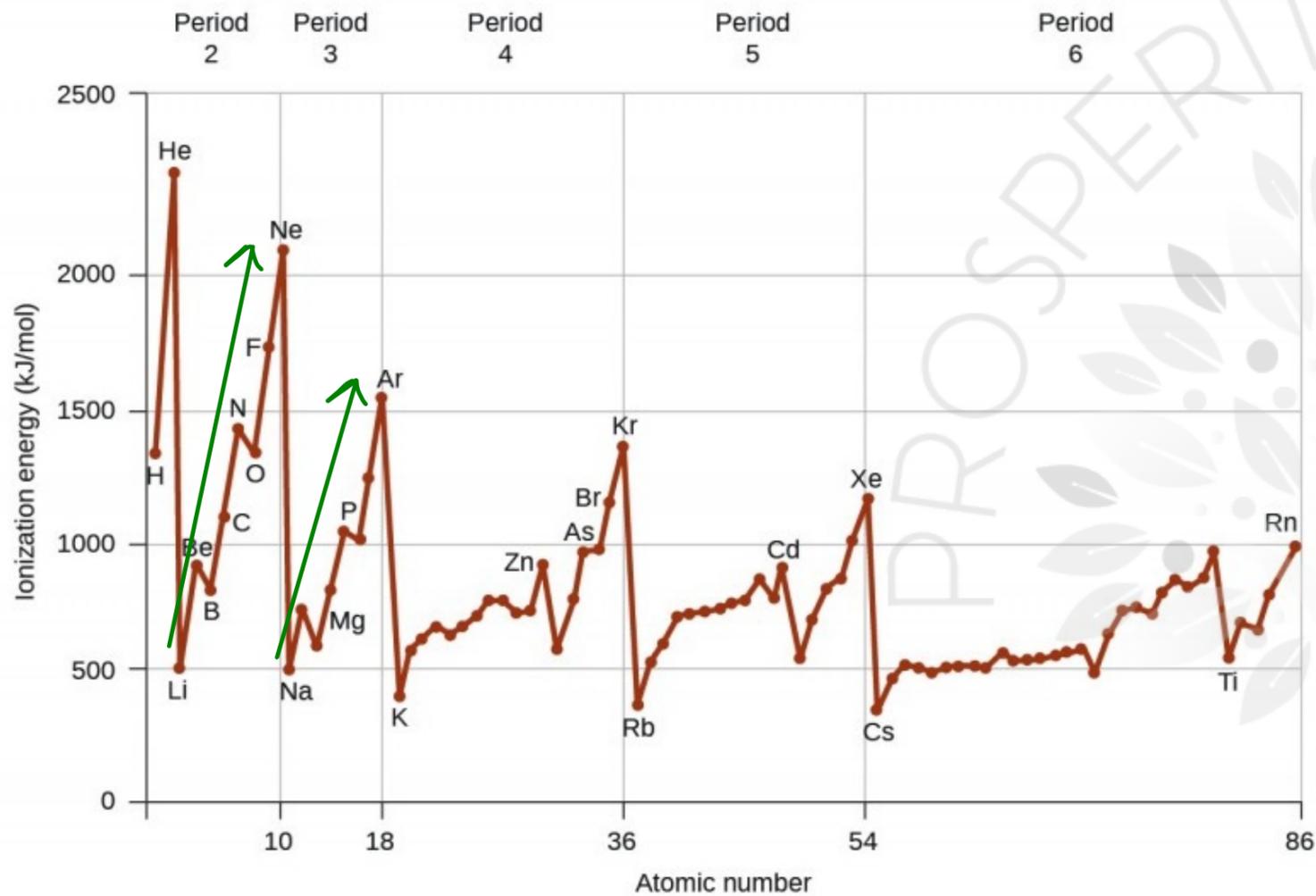
Q. Why is there a slight increase in successive ionization energies for the same shell?

- 1) As we remove electrons, there are now more protons than electrons hence stronger attraction
- 2) The repulsion between electrons decreases and so they move closer resulting in smaller radius

Hence more energy is required to remove next electron.

(This does not provide evidence for quantum model but provides evidence for Bohr's model [0 levels])

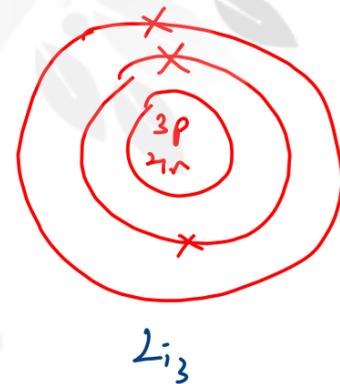
Trends of First ionization energies across the periodic table:-



Across a period, the first ionization energy generally increases

To explain this, you need to understand:

Shielding:- The electrons in the inner shells tend to shield the valence electrons from experiencing the full nuclear charge. Each inner shell electron approximately cancels out the effect of 1 proton.

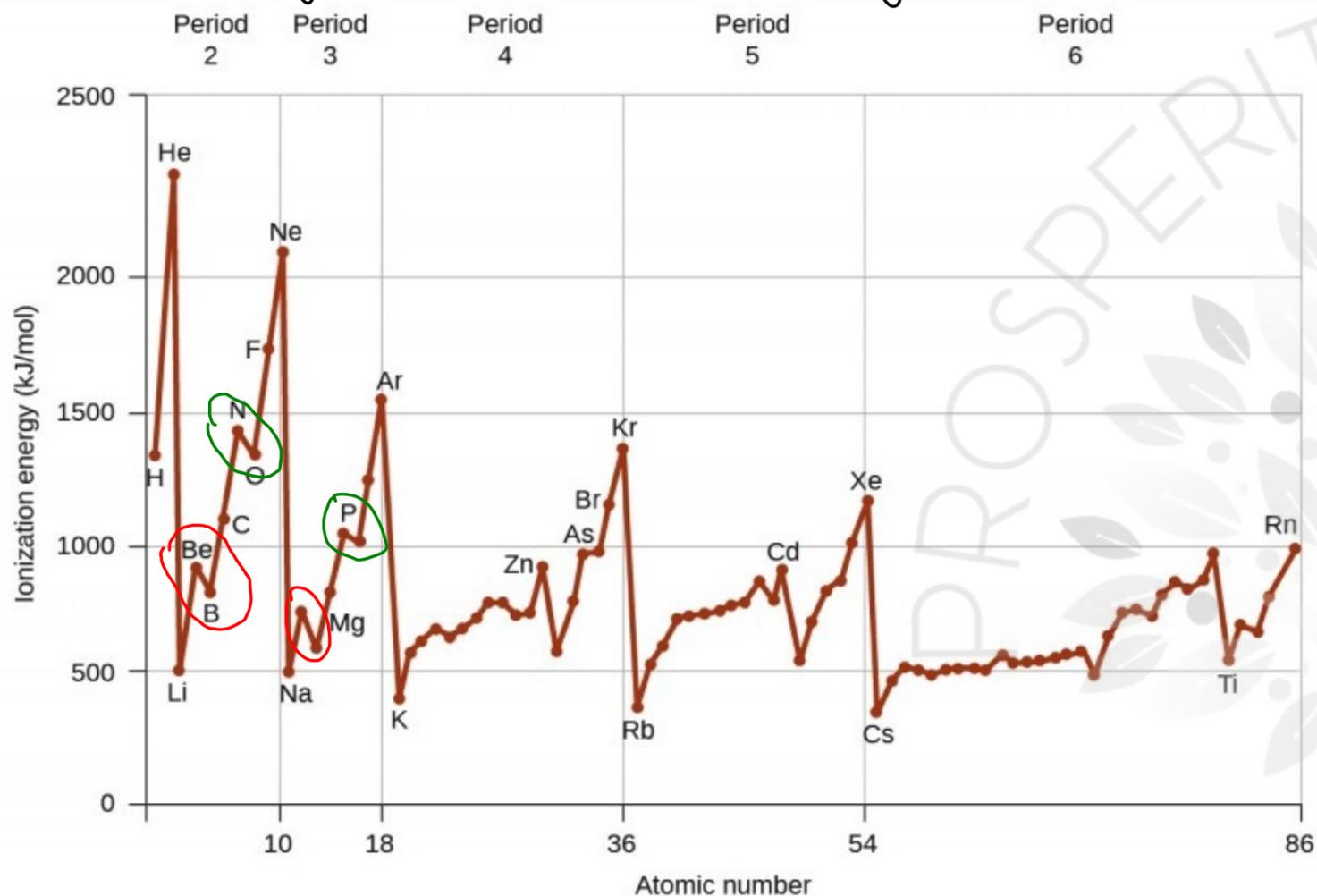


Inner shell electrons: $2e^-$
 protons: 3
 charge experienced by valence electron: +1

The first ionization energy increases across a period as:

- 1) The nuclear charge increases across a period
- 2) The shielding by inner shell electrons remains constant
- 3) As a result, the atomic radius decreases and so the force of attraction between the nucleus and valence electrons increases and so more energy is required to remove an electron.

Trends of First ionization energies across the periodic table:-

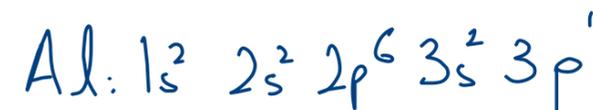


The decrease between Groups 2 and 13:-

The first ionization energy decreases from Be to B or Mg to Al



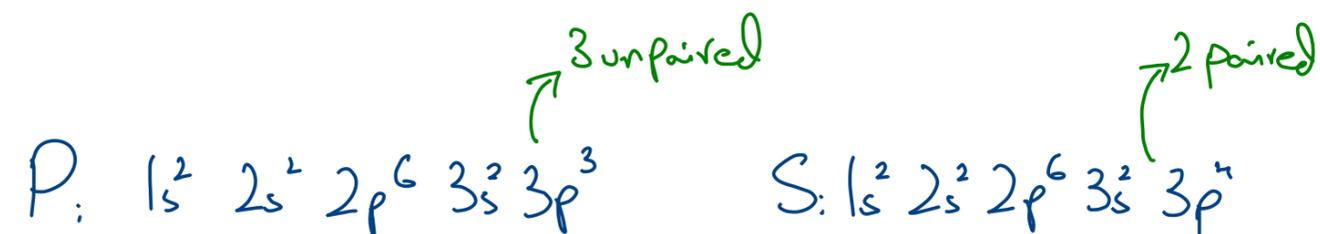
In Be the electron is removed from 2s and in B from 2p. 2s is a lower energy orbital than 2p and so removing an electron from it is harder.



Once again, 3s is a lower energy orbital than 3p so removing an electron from it is harder.

The decrease between Groups 15 and 16:

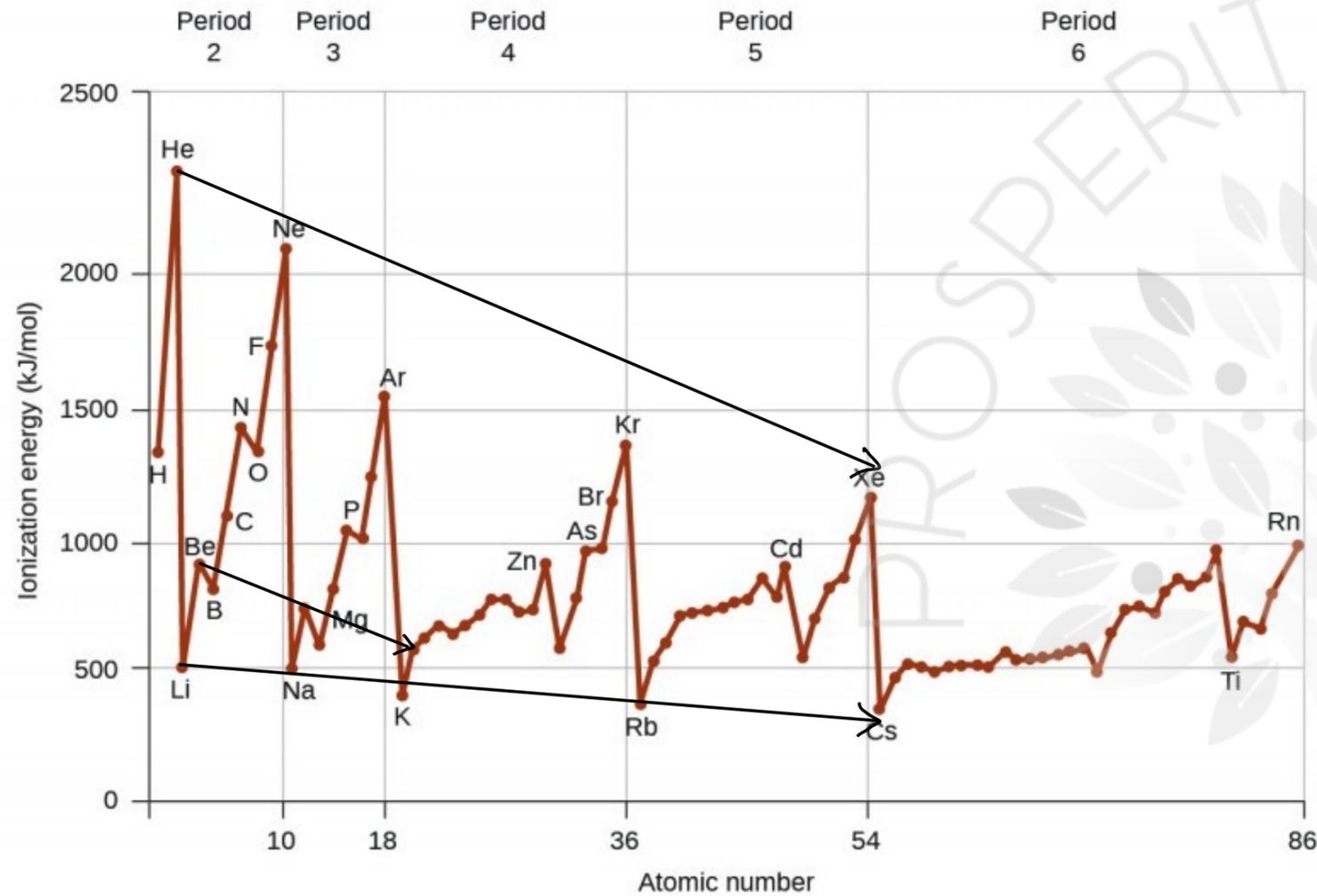
The first ionization energy decreases from N to O or P to S



It is easier to remove a paired electron than an unpaired electron as in paired electrons, they are already repelling each other

Same explanation applies here

Trends of First ionization energies across the periodic table:-



Down a group, first ionization energy decreases:-

- 1) Nuclear charge increases → cancel each other out
- 2) Shielding by inner electrons also increases (as more inner shells now)
- 3) The atomic radius therefore increases due to increased number of shells, hence the force of attraction between nucleus and valence electrons decreases so first ionization energy decreases.

Decrease from Noble gas to Group 1:-

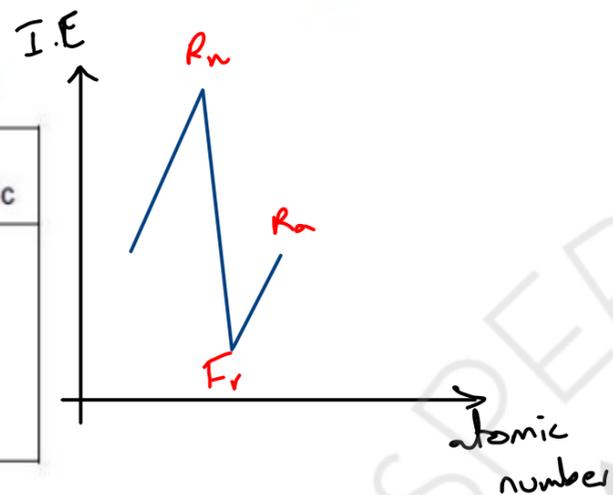
- 1) Nuclear charge increases → cancel each other
- 2) The shielding also increases → cancel each other
- 3) The atomic radius increases due to more number of shells hence it becomes easier to remove an electron.

The variation of ionization energies across the periodic table provide evidence for the Quantum model.

The elements radon (Rn), francium (Fr) and radium (Ra) have consecutive proton numbers in the Periodic Table.

What is the order of their first ionisation energies?

	least endothermic	→	most endothermic
A	Fr		Rn
B	Fr		Ra
C	Ra		Rn
D	Rn		Fr



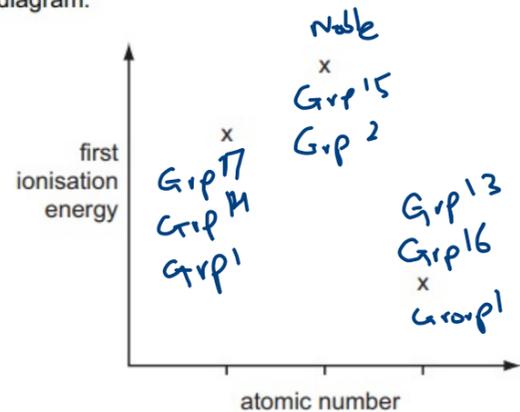
Which equation represents the second ionisation energy of an element X?

- A** $X(g) \rightarrow X^{2+}(g) + 2e^{-}$
- B** $X^{+}(g) \rightarrow X^{2+}(g) + e^{-}$
- C** $X(g) + 2e^{-} \rightarrow X^{2-}(g)$
- D** $X^{-}(g) + e^{-} \rightarrow X^{2-}(g)$

From which particle is the removal of an electron the most difficult?

- A** $Cl^{-}(g)$ X
- B** $F(g)$ X
- C** $K^{+}(g)$ X
- D** $Na^{+}(g)$

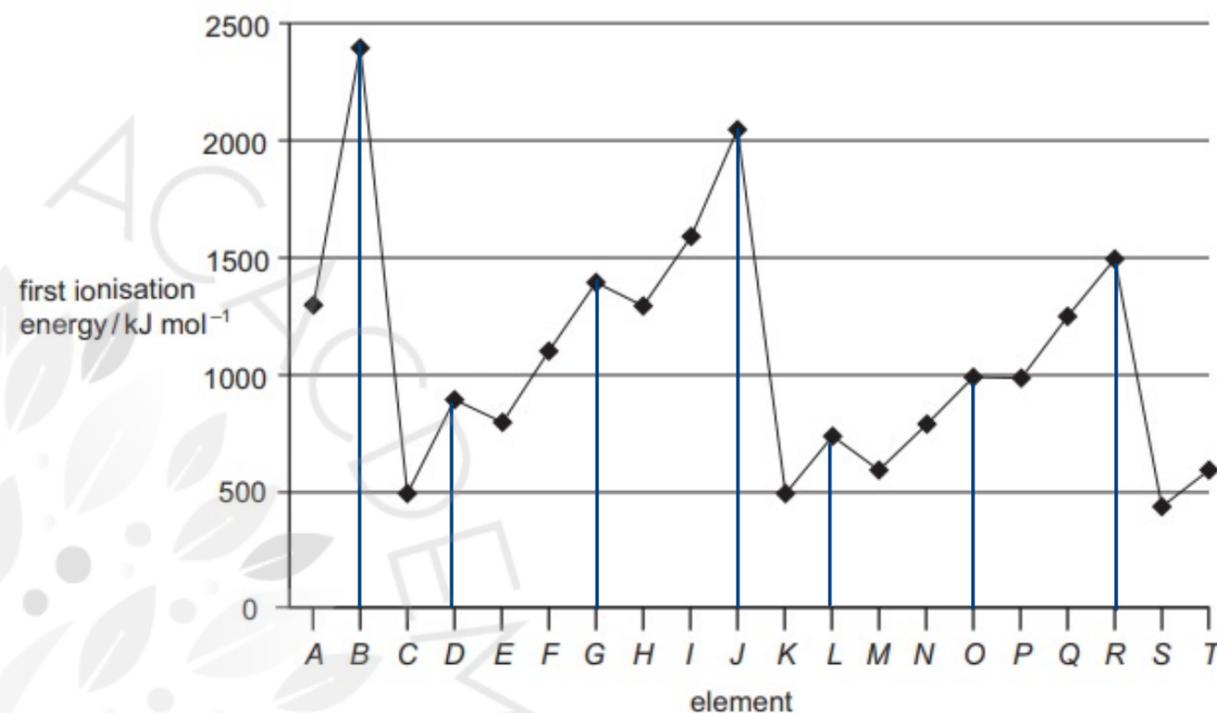
2 Three successive elements in the Periodic Table have first ionisation energies which have the pattern shown in the diagram.



What could be the first element of this sequence?

- A** C X
- B** N X
- C** F ✓
- D** Na X

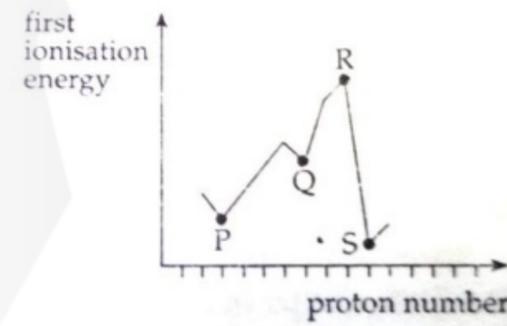
34 The first ionisation energies of successive elements in the Periodic Table are represented in the graph.



Which of these statements about this graph are correct?

- 1 Elements B, J and R are in Group 0 of the Periodic Table. ✓
- 2 Atoms of elements D and L contain 2 electrons in their outer shells. ✓
- 3 Atoms of elements G and O contain half-filled p orbitals. ✓

13 Which one of the elements marked on the graph is an alkali metal?



- A** Element P
- B** Element Q
- C** Element R
- D** Element S

(b) Fig. 1.2 shows the relative first ionisation energies of six successive elements in the Periodic Table.

The letters are **not** the symbols of the elements.

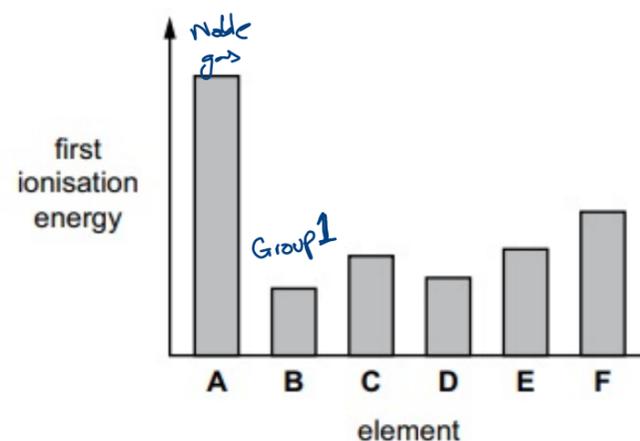


Fig. 1.2

(i) Define first ionisation energy.

The minimum amount of energy required to remove 1 mole of electrons from 1 mole of gaseous atoms to form 1 mole of 1+ charged gaseous atoms.

[2]

(ii) Suggest why the first ionisation energy of B is much less than that of A in Fig 1.2.

The electron from B is removed from a higher energy level than that of A. From B to A, the nuclear charge increases and so does the shielding so they cancel each other out. Therefore, the atomic radius increases and hence less energy is required to remove an electron.

[3]

(c) (i) On Fig. 1.3, sketch a graph to show the trend in the atomic radius of successive elements in Period 3.

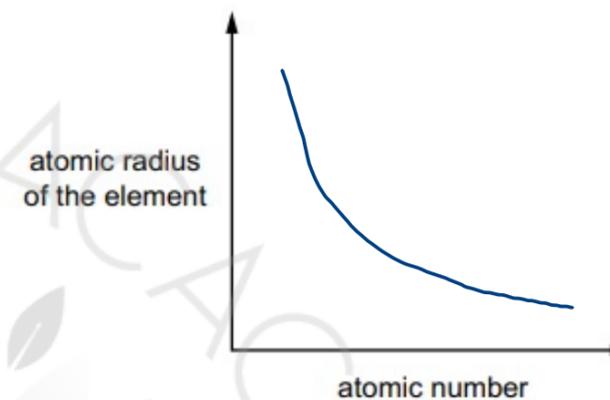


Fig. 1.3

[1]

(ii) Explain your answer to (c)(i).

The nuclear charge increases while the shielding by inner electrons remains constant across the period. Hence, the valence electrons experience a greater effective charge and the atomic radius decreases.

[3]

[Total: 12]

(a) Define the term **first ionization energy**.

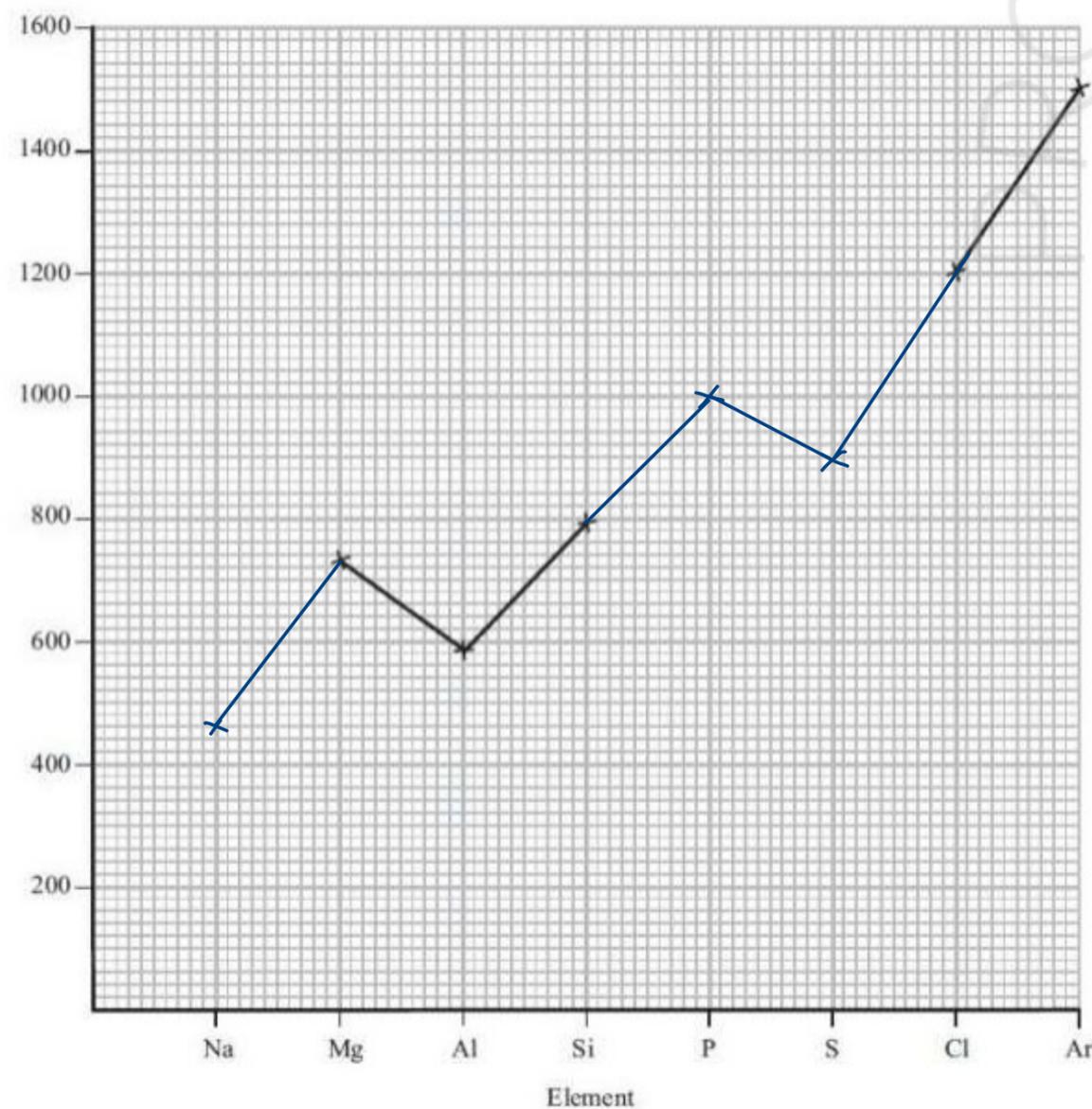
The minimum amount of energy required to remove 1 mole of electrons from 1 mole of gaseous atoms to form 1 mole of 1+ charged gaseous atoms.

(b) Write an equation, with state symbols, to illustrate the process occurring when the **second** ionization energy of sodium is measured.



(c) The graph below shows the variation in the **first** ionization energies of some of the elements in Period 3.

First ionization energy / kJ mol^{-1}



(i) On the graph, use crosses to show the approximate values of the first ionization energies for the elements Na, P and S.

Join the crosses to complete your graph.

*(ii) Explain why the first ionization energies generally increase across the period sodium to argon (Na to Ar).

Nuclear charge increases but shielding by inner electrons remains constant across a period. Therefore the atomic radius decreases and more energy is required to remove an electron.

*(iii) Explain why the first ionization energy of aluminium is less than that of magnesium. Mg: $1s^2 2s^2 2p^6 3s^2$ Al: $1s^2 2s^2 2p^6 3s^2 3p^1$

The electron in magnesium is removed from a 3s orbital which is lower in energy and closer to the nucleus than a 3p orbital from which the electron of aluminium is removed. Therefore more energy is required to remove an electron from Mg.

(d) Place the following species

S^+ S S^-

in order of increasing first ionization energy, starting with the lowest.

Lowest first ionization energy

Highest first ionization energy

S^-

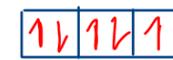
S

S^+



radius

paired - easy



radius \uparrow - easy

paired - easy



radius \downarrow - hard

unpaired - hard