## 2.1 THE NUCLEAR ATOM

- mass of an atom is in its minute, +ve nucleus
- nature of science → plum pudding model ⇒ gold leaf experiment (Rutherford)
  - → subatomic particle theory was a paradigm shift in 1800s
- bohr model of an atom: a small solar system with electrons orbiting nucleus
  - neutrons needed to stabilize nuclei (+ve repel = fall apart)

a nucleus is composed of protons + neutrons [this is the mass number], (Z)
↳ electrons occupy space around the nucleus (outside)

- atomic number = number of protons. (A)

$^{A}_{Z}X^{n}$ → charge

e.g. $^{7}_{3}Li$ has 3 protons and 4 neutrons, 3 electrons.

isotopes are atoms of the same element with different numbers of neutrons.
↳ different MASS numbers. (diff physical properties – mpt, bpt, diff. rate)

relative atomic mass average mass of ALL isotopes of an element compared to an atom of C-12

e.g. $\frac{(\% \times mass) + (\% \times mass)}{100}$ or if you're trying to find indiv mass (100−x)

chemical properties of isotopes are the same as they have same electronic configuration and electrons in outer shell. [USE MASS SPECTROMETER 4 R.A. MASS]

a radioisotope has an unstable nucleus and undergoes radioactive decay
↳ emits gamma / alpha / beta particle
↳ used in medicine = diagnosis / treatment / research
e.g. tracers e.g. radiocarbon dating
↳ PET scans = 3D images of tracer concentration in the body (cancer)
e.g. Iodine131 (Positron emission tomography) injected into blood, absorbed by good tissue, scanner detects photons of unabsorbed.

- RELATIVE ATOMIC MASS found via a mass spectrometer – determines abundance of isotope in an element to work out RAM

## 2.2 ELECTRON ARRANGEMENT

- electrons are arranged in energy levels of increasing energy and distance from the nucleus (first energy level, n = 1, is closest to nucleus)
  - ↳ it has the lowest energy
  - ↳ can contain max 2 electrons. (n=2, max 8, n=3, max 18)
- electrons in the outermost energy level = valence electrons
  - ↳ they define the group number of the element.

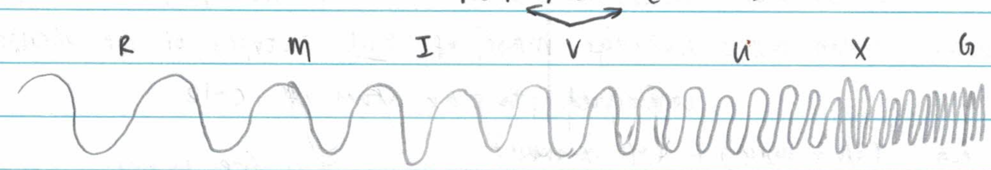
evidence for energy levels come from the LINE EMISSION SPECTRUM

1. when electrons are excited = move from ground state to HIGHER energy level.
2. when it drops back into a lower shell = emit energy in form of light (photon

↳ what could it release?

anything on the electromagnetic spectrum.

ROY G BIV (colours)




low frequency
long wavelength

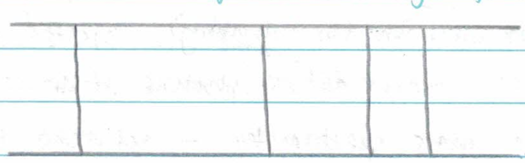
[ energy ∝ frequency
frequence ∝ 1/wavelength ]

high frequency
short wavelength

continuous spectrum: all possible wavelengths / frequencies

2 (a). when energy is supplied / released, the element will emit a spectrum of CERTAIN lines at CERTAIN wavelengths.
the line spectrum is a characteristic spectrum of discrete lines.

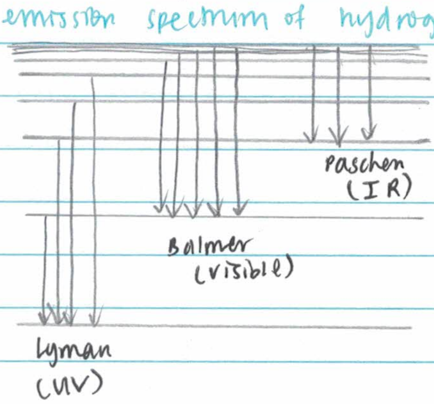
line emission spectrum of hydrogen



lower energy higher energy

- this is evidence of electrons in discrete energy levels which CONVERGE at higher energies.

emission spectrum of hydrogen




- the jump back to n=1 is the largest drop in energy. High energy, high frequency, UV.
- when electron jumps to n=2 level, this is seen in (VISIBLE) region
- each element has different energy level distances = diff. line spectrum.

## 2.2 ELECTRON ARRANGEMENT (CONT'D)

- each energy level is n and has a number (1st is n = 1 etc)
- each energy level can hold $2n^2$ number of electrons.
- within each energy level, there are subshells.
- there are FOUR types of subshells: s, p, d, f.
- each sub-shell is made of electron orbitals which hold 2 electrons.
  - (orbital) this is where an electron is likely to be.

s subshells have lowest energy, then p, then d, then f.
- s → 1 orbital
- p → 3 orbitals
- d → 5 orbitals

in each energy level ... 1st (n=1) has 1 subshell (s)
(n=2) has 2 subshells (s, p)
n = 3 has 3 subshells (s, p, d)

each orbital has a different shape
↳ s = spherical (O)
↳ $p_{x,z,y}$ = dumbell (8)   x = ∞   y = 8   z = 8

- electron configuration
  ① aufbau principle
  - orbitals of lowest energy = filled first
  ② hund's rule
  - each orbital must have ONE electron b4 pairing.

1s
2s 2p
3s 3p 3d
4s 4p 4d 4f

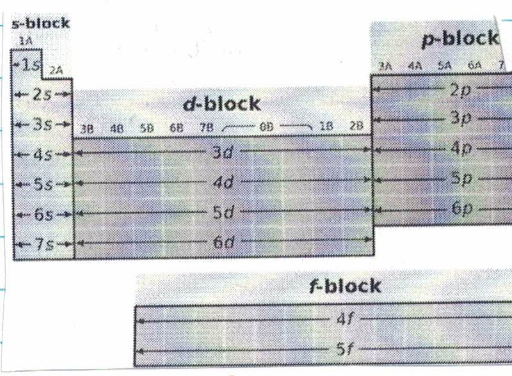
e.g He = $1s^2$
1 = energy level
s = subshell
2 = no. of $e^-$

e.g. Li = $1s^2 2s^1$
N = $1s^2 2s^2 2p^3$
or 'electrons in boxes' diagrams: 1s □ 2s □ □□□ 2p

because 3d has more energy that 4s, 3d AFTER 4s.
Ne = $1s^2 2s^2 2p^6 3s^2 3p^6$

shorthand = nearest noble gas configuration.

e.g Zr is [Kr] $5s^2 4d^2$
Cs is [Xe] $6s^1$



exceptions

- Cu is $4s^1 3d^{10}$ [ $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$ ]
- Cr is $4s^1 3d^5$ [ $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$ ]

→ because $4s^1 3d^5$ is more stable than $4s^2 3d^4$

4s [↑] 3d [↑][↑][↑][↑][↑]

vs

[↑↓] [↑][↑][↑][↑][ ]

Electronic configuration of IONS

- electrons **furthest** from the nucleus are lost first
- 4s are lost before 3d (4s gains energy when occupied)
- electrons are removed in numerical order (4 before 3, 3 before 2, etc)
- isoelectronic: same electronic configuration.

## 12.1 Electrons in Atoms.

evidence for the existence of sub-shells & shells can be seen from studies in trends of ionisation energies within the periodic table.

first ionisation energy: energy absorbed when 1 MOLE of electrons is removed from 1 MOLE of an atoms in GASEOUS state, giving 1 mole of positive ions.

$$M_{(g)} \longrightarrow M^{+}_{(g)} + e^{-}$$

ionisation energy (1st) is different, depending on...

① DISTANCE : → greater the distance between nucleus + valence electrons
→ less attractive force
→ decrease in IE

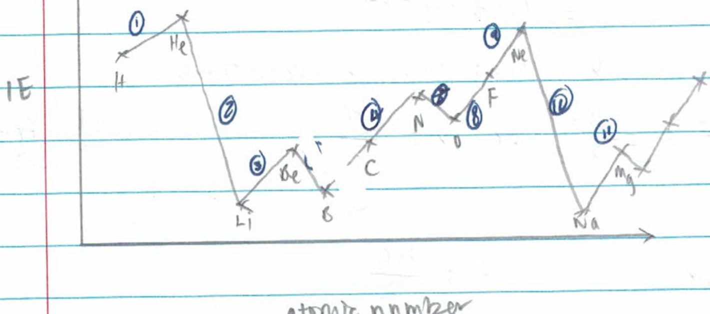
② SHIELDING : → repulsion filled by inner shells / subshells makes the electrons in the outer shell easier to remove
→ decrease in IE.

③ NUCLEAR CHARGE : → nucleus is +ve charge
→ more protons, more charge, more pull (force of attraction)
→ increase in IE

④ ELECTRON PAIRING : → paired $e^-$ is easier to remove than unpaired $e^-$
→ electrons in shells that are half full are hardest to remove
e.g. $3p^3$ | 1 | 1 | 1 |

successive ionisation energies : evidence for SHELLS.
↳ after removing FIRST electron, can remove other electrons.

graph of 1ST IONISATION energies : evidence for SUBSHELLS.




1. H → He : bigger nucleus (+ve charge), increase
2. He → Li : distance from nucleus (shell) ↓ IE
   - increase in shielding   Be → B : new shell ↓
3. Li → Be : increased nuclear charge ↑ IE ↙
4. B → C → N : increased nuclear charge ↑ IE
5. N → O : in N ($2p^3$), all unpaired, in $2p^4$ increased repulsion between paired electrons = ↓
6. O → F → Ne : increased nuclear charge (since all $e^-$ removed same shell)

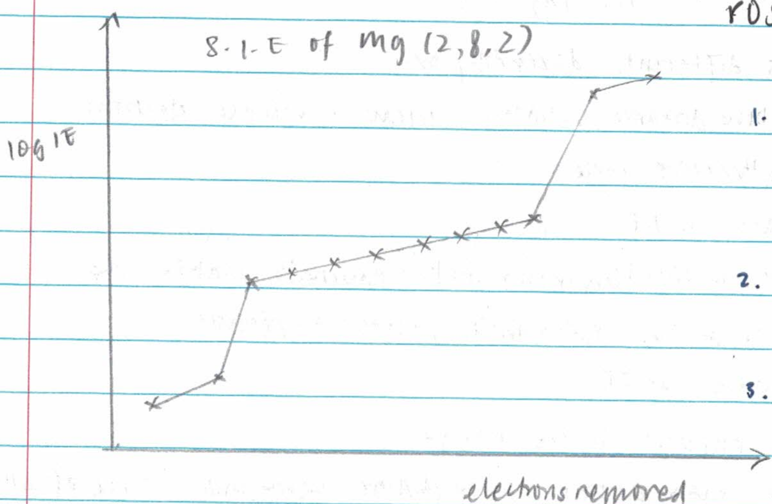
EVIDENCE for SHELLS:
→ large decrease in IE from group 18 to group 1 due to an increased distance of the outer electron from the nucleus when a new shell is added (He → Li, Ne → Na)

## SUCCESSIVE IONISATION ENERGIES. (evidence for shells!)

- you can continue to remove electrons (2nd, 3rd... ionisation energies)

e.g.
$Mg_{(g)} \rightarrow Mg^{+}_{(g)} + e^-$
$Mg^{+}_{(g)} \rightarrow Mg^{2+}_{(g)} + e^-$
$Mg^{2+}_{(g)} \rightarrow Mg^{3+}_{(g)} + e^-$
etc etc etc

there is a general increase
① because ratio of protons : electrons is bigger → greater attraction
② removing electrons forms a GREATER POSITIVE CHARGE.




1. large increase – 2nd electron removed from 3rd shell, 3rd one is removed from 2nd
   ↳ closer to nucleus.
2. general increase = higher ratio of +ve : negative
3. large increase – 10th electron from 2nd shell, 11th from 1st = even closer to nucleus (closer than 2nd)

- the FIRST large INCREASE indicates no of electrons in the outer shell.

## QUANTIZATION OF ENERGY

- precise lines in line emission spectrum have a specific wavelength which corresponds to a discrete unit of energy (= energy difference between energy levels)
  ↳ ∴ it is foundation of quantization: ~~radio~~ electromagnetic radiation comes in parcels / discrete.
  ↳ photon is a quantum of radiation.

[ frequency, energy, related by..]

$E = h\nu$ $\quad h = 6.63 \times 10^{-34}$

$\nu$ = frequency of radiation.

[frequency + wavelength related by...] $c = \nu\lambda$ or $\nu = \frac{c}{\lambda}$

!! frequency is $\frac{c}{\lambda}$

therefore $E = h\nu$ is $E = \frac{hc}{\lambda}$

$E \propto \frac{1}{\lambda}$ (greater energy, smaller wavelength, higher frequency)

# EXAM QUESTIONS & EXAMPLES

examples: photon of red light, 4 s.f — $\lambda = 450^{nm}$ $h = 6.626\times10^{-34}$, $c = 2.998\times10^8$

find the energy. convert!! $450\times10^{-9}$

$$E = \frac{hc}{\lambda} \Rightarrow \frac{6.626\times10^{-34} \times 2.998\times10^8}{450\times10^{-9}} = 3.056\times10^{-19}$$

★ wavelength must always be in m.

first ionisation energy in $kJ\ mol^{-1}$ — $\lambda = 91.16$ nm, $h = 6.626\times10^{-34}$, $c = 2.998\times10^8$

$N_A = 6.022\times10^{23}$

$$E = \frac{hc}{\lambda} \quad \frac{6.626\times10^{-34}\times 2.998\times10^8}{91.16\times10^{-9}}$$

$= 2.179\times10^{-18}$ per photon

$= 2.179\times10^{-18} \times 6.022\times10^{23}$ per mole

$= 2.179\times10^{-18}\times 6.022\times10^{23}\times10^{-3}$ per kJ/mol

first ionisation energy in $kJ\ mol^{-1}$, frequency is 3.3 PHz,

★ $c = f\lambda$ so $f = \frac{\lambda}{c}$ and $\frac{c}{f} = \lambda$

$$\frac{2.998\times10^8}{3.3\times10^{15}} = \lambda$$ ... and work out the rest.

state a radioactive isotope used in medicine.

- iodine - 131.

bromine exists as the isotope Br-79 and Br-81. what is % of $Br^{79}$ with 79.9 ma

$79x + 81(100-x) = 7990$

$-2x = -110$, $x = 55\%$

the value of the relative molecular mass of HBr is 80.91. why doesn't any sample of HBr have this Mr value? (2)

① it is an average of all HBr values

② each HBr has a different mass depending on the isotopes of H and Br.

describe and explain how the chemical / physical properties of 2 isotopes compare.

① similar chemical properties ② same valence electrons

③ different physical properties ④ different masses.

explain how the formation of lines in a visible line spectrum indicates the presence of energy levels. (1)

- shows electrons transition from high to low energy levels, and so they emit energy in the form of light describe the emission spectrum of hydrogen; explain how this is related to energy levels.

① · emission spectrum is a series of lines that converge @ higher frequencies

② · electrons transfer from high to low energy levels and light is emitted.

③ as they go down the energy levels (releasing energy)

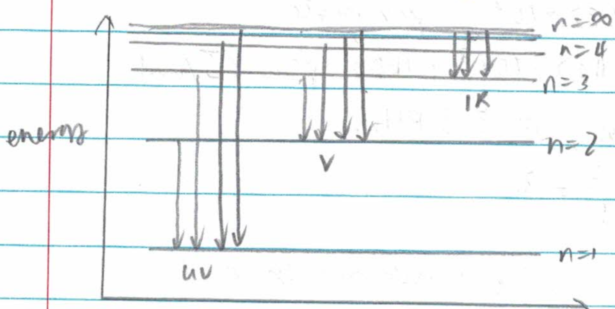
④ n=1 is ground = UV, n=2, visible light, n=3, IR. → transitions to n=1,2,3.

(n=1: lyman; n=2: Balmer; n=3: paschen.)

the line which represents the FIRST IONISATION ENERGY of hydrogen is from

n=1 to n=∞ (to remove / ionise, electron must be taken OUT)

energy level diagram for hydrogen




what is meant by the term "first ionisation energy"?

· the amount of energy absorbed when one mole of electrons is removed from one mole of ~~atd~~ a gaseous atom.

what is "periodicity"?

· repeating patterns / trends across periods.

explain how the graph of first ionisation energies provides evidence for the existence of main energy levels and sub-levels

· existence of shells - huge drop between He and Li (different shells)

· existence of subshells - general increase across a period. (electrons added to same shell) → greater attraction from higher +ve : -ve ratio.

furthermore = highest value for noble gases, lowest for alkaline → value drops drastically after Gp 8 as a new shell added = further away from nucleus.

· drop in IE from Be to B ($s^2$ to $s^2p^1$) and

drop in IE from N to O ($p^3$ to $p^4$)

why is the difference between Germanium's 4th and 5th ionisation energies is much greater than other successive values? $3d^{10}4p^2$

· instead of n=4 (taking the electron) taken from n=3 (taken from energy level closer to the nucleus)

↳ attraction of protons is greater

EXAM QUESTIONS

define the term isotope.

· atoms of the same element but with different NUMBER OF NEUTRONS (not mass)

state how a line spectra differs from a continuous spectrum.

· line spectrum = discrete LINES at certain frequencies.

· continuous spectrum = all possible wavelengths/frequences

successive ionisation energies:

1st - 799. 2nd - 2420 3rd - 3660 4th - 25000 5th - 32800

↳ why is there a large increase between 3 → 4?

· electron taken from n = 1 level instead of n = 2 = closer to nucleus

· more attraction = more energy needed to ionise.

↳ why is the increase between 1st → 2nd > 2nd → 3rd?

· electron in 1→2 taken away from a different shell; closer to nucleus (2p to 2s) ⇒ more attraction ↑ IE.

↳ state with reason how the value for the second IE of carbon would compare with the second IE of boron.

$B = 1s^2 2s^2 2p^1$
$C = 1s^2 2s^2 2p^2$

carbon = less IE because taking away from $2p^2$, further away than $2s^2$ → less attraction = less IE. whereas B taken from 2s level = closer = ↑ IE.

★ provide 2 sides!!

However, nuclear charge of C is greater. ↑IE due to ↑ protons than Boron (+ve : -ve ratio higher).

technitium is used as a tracer in medicine. suggest why it could be dangerous

· radioactive isotope which could ionise body's cells = cause cancer.

what is relative atomic mass?

· the weighted mean mass of all atoms and isotopes (naturally occurring)

· relative to a C-12 atom (1/12 th)

why is there a general increase in the value of 1st IE Li → Ne

· electrons taken ^ all in the same shell, relatively the same distance from the nucleus (n=2 energy level), hence a general increase merely based on nuclear attraction (+ve : -ve ratio) as proton no. increase as atomic no. increases.

[ATOMIC RADIUS DECREASES]

first ionisation energy of Hydrogen is 1312 $kJmol^{-1}$. determine frequency and wavelength.

$E = h\nu$ ← frequency.
↑ planck's constant

$\frac{c}{\nu} = \lambda$ / $\frac{c}{\lambda} = \nu$

explain how the line spectrum is related to the energy level in a hydrogen atom.

↳ electrons jump from excited state to the n=2 level

↳ converge @ higher frequencies

↳ frequencies of the light = directly proportional to the difference in energylevels.