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100 SOLVED QUESTIONS SERIES

VOLUME 3: THE ATOM

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Part one
QUESTIONS

Question 1

(a) In the isotope symbol of each atom, there is a superscripted number. This number is called **mass number**.
  (i) How is the mass number determined?
  (ii) Why is this number called a 'mass' number?

(b) Magnesium has three stable isotopes of masses 23.985, 24.985, 25.982 a.m.u. The relative abundances of the three isotopes are 39.35, 5.065, and 5.585 respectively. Calculate the average atomic mass.

Question 2

(a) In a neutral atom, how does the number of electrons compare to the number of protons?

(b) Why the relationship stated in (a) above is important in making a neutral atom?

(c) A sample of pure unknown element **X** was analysed and the data is given in the table below. Calculate the relative atomic mass of **X** strictly to two decimal places.

Symbol	**Mass (a.m.u)**	**Natural abundance (%)**
$^{20}_{10}X$	19.992	90.22
$^{21}_{10}X$	20.994	0.257
$^{22}_{10}X$	21.991	8.82

Question 3

(a) Today atomic model is the result of contributions from different scientists lived at different times. With the help of diagram, give the name of atomic model contributed by each of the following scientists.
  (i) John Dalton
  (ii) J.J. Thomson
  (iii) Ernest Rutherford

(iv) Niels Bohr
(v) Erwin Schrödinger

(b) If ^{69}Ga and ^{71}Ga occur in proportions 60: 40, calculate the average atomic mass of Ga.

Question 4

The mass spectrum of an element enables the relative abundance of each isotope of the element to be determined. Data relating to the mass spectrum of an element **X** whose atomic number is 35 appear as follows;

Mass number of isotopes	Relative Abundance
79	50.5%
81	49.5%

(i) Define the term **isotope**.
(ii) Write down the conventional symbols for the two isotopes of **X**.
(iii) Calculate the relative atomic of **X** to the three significant figures.

Question 5

(a) The symbol, $^{35}_{17}Cl$, can simply be written as ^{35}Cl and **not** $_{17}Cl$. Explain.

(b) Silver consists of two isotopes $^{107}_{47}Ag$ and $^{109}_{47}Ag$ of atomic masses 106.91g/mol and 108.91g/mol respectively. The relative abundances of these isotopes are 51.88% for $^{107}_{47}Ag$ and 48.12% for $^{109}_{47}Ag$. Calculate the average atomic mass of Ag.

Question 6

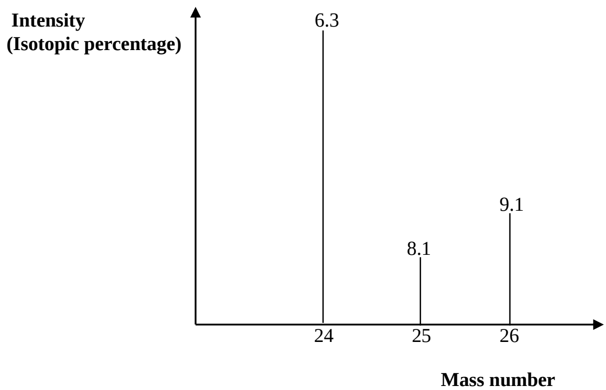
(a) In Rutherford's experiment, generally the thin foil of heavy atoms, like gold and platinum have been used to be bombarded by the α-particles. If the thin foil of light atoms like aluminium is used, what difference would be observed from the Rutherford's experimental results?

(b) Copper has two naturally occurring isotopes. $Cu - 63$ Have an atomic mass of 62.9296 amu and an abundance of 69.15%. What is the atomic mass of the second isotope? Given that average atomic mass of copper is 63.546 amu.

Question 7

(a) Give two reasons which makes gold to be useful in Rutherford's α-particle scattering experiment.

(b) The diagram below shows the mass spectrum of magnesium. The heights of the three peaks and the mass number of isotopes are shown. Calculate the relative atomic mass of magnesium.



Question 8

(a) With an example, give the meaning of the term **isodiapheres**.

(b) In a sample of 400 lithium atoms, it is found that 30 atoms are lithium-6 (6.015g/mol) and 370 atoms are lithium -7 (7.016g/mol).Calculate the average atomic mass of lithium.

Question 9

(a) Among the following pairs of orbitals which orbital will experience the larger effective nuclear charge? Give reason to justify your choices.

(i) 2s and 3s,

(ii) 4d and 4f,

(iii) 3d and 3p

(b) Natural rubidium has the average mass 85.4678amu and is compared of isotope $\text{Rb}-85(\text{mass}=84.9117\text{amu})$ and $\text{Rb}-87$. The ratio of atoms $\frac{\text{Rb}-85}{\text{Rb}-87}$ in natural rubidium is 2.591. Calculate the mass of $\text{Rb}-87$.

Question 10

(a) What does it mean to say that the energy of the electrons in an atom is quantized?

(b) The mass spectrum of Cl_2 shows peaks at masses 70, 72 and 74 a.m.u. The heights of the peaks are in the ratio of 9:6:1 respectively. What is the relative abundance of ^{35}Cl ad ^{37}Cl? Hence calculate the average mass.

Question 11

(a) The quantum numbers of six electrons are given below. Arrange them in order of increasing energies.

A. n = 4, l = 2, m = –2, s = –1/2

B. n = 3, l = 2, m = 1, s = +1/2

C. n = 4, l = 1, m = 0, s = +1/2

D. n = 3, l = 2, m = –2, s = –1/2

E. n = 3, l = 1, m= –1, s= +1/2

F. n = 4, l = 1, m = 0, s = +1/2

G. n = 4, l = 0, m = 0, s = +1/2

(b) Calculate the wavelength of the first line in Balmer series. ($R_H = 1.097\times10^7\text{m}^{-1}$)

Question 12

(a) A sample of argon (Ar) known to contain isotopes of mass numbers 36, 39 and 40 is introduced in a mass spectrometer. The sample is bombarded with electrons to form positively charge ions.

(i) Which ion ${}^{36}_{18}Ar^{+}$ or ${}^{40}_{18}Ar^{+}$, is likely to be deflected most in the magnetic field? Give a reason for your answer.

(ii) Some atoms lose two electrons in the ionisation chamber. Which ion ${}^{36}_{18}Ar^{+}$ or ${}^{36}_{18}Ar^{2+}$, is likely to be deflected most in the magnetic field? Give a reason for your answer.

(b) Calculate the energy of a photon of radiations whose wavelength is 5.89×10^{-5}cm emitted from sodium atoms when heated.

Given that c = 3×10^{8}m/s and h = 6.62×10^{-34}Js

Question 13

(a) The neutron was not discovered until more than 30 years after the discovery of the proton and the electron. Why was the neutron more difficult to detect?

(b) Using Plank's equation, calculate the energy of photons of light for radiation of wavelength 242.4nm, the longest wavelength that will bring about the photo dissociation of O_2. What is the energy of:

(i) one photon.

(ii) a mole of photons of this light?

($h = 6.626 \times 10^{-34}$Js and c = 2.998×10^{8}m/s)

Question 14

(a) Explain why the relative atomic mass of argon is greater than that of potassium although potassium has larger atomic number.

(b) If the wavelength of the first member of Balmer series in hydrogen spectra is 6563A°. Calculate the wavelength of the first member of the Lyman series.

Question 15

(a) The average atomic mass of calcium is 40.08amu. What is probably its most abundant isotope?

(b) A diode laser emits at wavelength of 987nm. All of the radiation it emits is absorbed in a detector which measures a total energy of 0.52J over a period of 32 sec. How many photons per second are being emitted by the laser?

Question 16

(a) From the following nuclei select the isotopes and isobars

$^{238}_{92}U$, $^{234}_{90}Th$, $^{234}_{92}U$ $^{238}_{93}Np$, $^{234}_{91}Pa$

(b) Calculate the wavelength of the line in Balmer series associated with energy transitions: $E_4 \rightarrow E_2$, where:

$$E_4 = -1.362 \times 10^{-19} J$$

$$E_2 = -5.498 \times 10^{-19} J$$

$$h = 6.63 \times 10^{-34} Js$$

$$c = 3 \times 10^8 m/s$$

Question 17

An electromagnetic radiation was emitted in the Balmer series as result of electron transition between n = 2 and n = 5. Calculate:

(a) The energy of radiation in kJ/mol
(b) The frequency of radiation
(c) The wavelength of radiation in metres

Where;

$h = 6.63 \times 10^{-34}$Js R= 1.097×10^{7}/m and c = 3×10^{8}m/s

Question 18

(a) What is meant by the expressions "Carbon-14" and "Silver-108"?
(b) Calculate the ionisation energy of hydrogen.

Given that R_H = $1.097 \times 10^{7} m^{-1}$, h = 6.63×10^{-34}Js and c = $3 \times 10^{8} m/s$

Question 19

(a) Natural neon consists of a mixture of three isotopes: 90.92% neon $- 20$, atomic mass

$19.9924amu: 0.257\% neon - 21$, atomic mass $20.9930amu$ and 8.82% neon $- 22$, atomic mass $21.9914amu$. Without doing any calculations, estimate the approximate average atomic mass of neon.

(b) An electron in potassium metal was excited from its ground energy level to a higher energy level after being heated. On returning to its ground state, a violet light of wavelength 4.34×10^{-7}m was emitted. Calculate the energy difference between the ground level and the higher energy level reached by the electron (h = 6.626×10^{-34}Js and $c = 3 \times 10^{8} m/s$)

Question 20

(a) The RAM listed in the Periodic Table for Magnesium is 24.305. A form five student claims "there isn't a single magnesium atom in the universe with a mass of 24.305." Is this statement correct? Explain your answer.

(b) The uncertainty in the momentum of a particle is 2.5×10^{-14}gcmsec^{-1}

Find its approximate position (h $= 6.626 \times 10^{-34}$ and π 3.14)

Question 21

(a) Explain where the different colours of light come from in the bright line spectrum of an element.

(b) What is the de Broglie's wavelength for a neutron travelling with a velocity 5% of the speed of light (speed of light = 3×10^{8}m/s, h = 6.626×10^{-34} and mass of neutron = 1.61×10^{-27}kg)

Question 22

(a) In what way does the photoelectric effect support the particles theory of light?

(b) What is the minimum uncertainty in position as imposed by the uncertainty principle on 100g ball thrown at $42 \pm 1 m/s$? (Hint: $\pm m/s$ means maximum error in the measurement of speed of the ball is $1 m/s$)

Question 23

(a) Under what circumstances can an atom emit a photon? Give two possibilities.

(b) A car weighing 3×10^3kg is moving on a highway. Its speed can be measured with an accuracy of ± 0.0025 mile/hour and its position with an accuracy of ± 0.01mile. Is the Heisenberg uncertainty principle valid? (1mile = 1.6×10^3m)

Question 24

(a) Why Bohr's orbits called stationary states?

(b) By using quantum numbers, deduce maximum number of electrons which can be contained in third shell.

Question 25

(a) Compare and contrast the models of the atom that were proposed by Rutherford and Thomson.

(b) What type of orbital may be specified by the following quantum numbers?

(i) $n = 1, l = 0$

(ii) $n = 2, l = 1, m_l = 0$

(iii) $n = 2, l = 1$

(iv) $n = 3, l = 1, m_l = +2$

Question 26

(a) A beam of electrons and beam of protons are moving with the same speed. Which has longer De Broglie wavelength?

(b) Assign quantum numbers for the following orbitals:

(i) $3s -$ orbital

(ii) $3P_z -$ orbital

(iii) $2P_x -$ orbital

(iv) $3d_{x^2-y^2} -$ orbital

Question 27

(a) What does the 'photoelectric effect' say about nature of light?

(b) Write electronic configuration of elements with the following atomic numbers: (i) 13 (ii) 9 (iii) 27 (iv) 21

Question 28

(a) If matter has wave nature why is this wave- like character not observed in our daily experiences?

(b) By using boxes and arrows draw electronic configuration of the following elements:

(i) Chromium (Cr)

(ii) Copper (Cu)

Question 29

(a) Which of the following sets of quantum numbers are allowed and which are not? For those which are not allowed state why?

(i) $n = 2, l = 1, m_l = 0$

(ii) $n = 2, l = 2, m_l = 2$

(iii) $n = 2, l = 1, m_l = -1$

(iv) $n = 2, l = 1, m_l = -2$

(b) Assign quantum numbers for the following:

(i) 4^{th} electron in nitrogen

(ii) Last electron in oxygen

(iii) Valence electron in sodium

Question 30

(a) Give the differences between orbit and orbital.

(b) Write electronic configuration to show excited state of:

(i) Aluminium (ii) Silicon (iii) Phosphorous

Question 31

(a) Which of the following orbitals are not possible?

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

Give reason for your choices.

(b) Write electronic configuration of sulphur at excited state to show:

(i) Its covalency of 4

(ii) Its covalency of 6

Question 32

(a) Give the values for the quantum numbers n, l, m_l for each orbital in the 4d subshell.

(b) Write electronic configuration of chlorine at excited state to show:

(i) Its covalency of 3
(ii) Its covalency of 5
(iii) Its covalency of 7

Question 33

(a) By referring to the rules of writing electronic configuration justify the following statements:

(i) Electron are lazy
(ii) Electron are unfriendly
(iii) Quantum numbers are fingerprints for electrons

(b) Draw electronic configuration of the following ions:

(i) Fe^{2+}
(ii) zn^{2+}
(iii) Fe^{3+}

Question 34

(a) How electronic configuration of chromium is said to violate Aufbau principle?

(b) For each of the following electronic configuration state whether it is valid. Give reason for invalid ones.

(i) $1s^2\ 2s^2 2p^6 3s^2 3p^6 4s^2 4d^{10} 4p^5$
(ii) $1s^2\ 2s^2\ 2p^6\ 3s^3\ 3p^5$
(iii) $[Ra]7s^2\ 5f^5$
(iv) $[Kr]5s^2 4s^{10} 5p^5$
(v) [Xe]

Question 35

(a) Arrange the following orbitals in order of increasing their energies for:

(i) Hydrogen atom

(ii) Copper atom

$1s, 2s, 2p, 3d, 4s, 4p, 3p, 3s$

(b) A sample of naturally occurring silicon consists $Si-28(27.9769\ amu), Si-29(28.9765\ amu)$ and $Si-30(29.9738\ amu)$. If the atomic mass of silicon is 28.0855 and the natural abundance of $Si-29$ is 4.67%, what are the natural abundances of $Si-28$ and $Si-30$?

Question 36

(a) With reference to xenon at ground state, how many electrons have the following set of quantum number?

(i) $n = 4$
(ii) $n = 4, l = 2$
(iii) $l = 0$
(iv) $n = 2, l = 2, m_l = -1 , s = {}^1/_2$
(v) $n = 4 , l = 3$

(b) Pb has an average atomic mass of 207.19amu.The three major isotopes of Pb are $Pb-206\ (205.98amu); Pb-207(206.98amu)$; and $Pb-208(207.98amu)$. If the isotopes of $Pb-207$ and $Pb-208$ are present in equal amounts, calculate the percentage abundance of $Pb-206, Pb-207, Pb-208$.

Question 37

(a) With reference to calcium at ground state, how many electrons have the following quantum numbers?

(i) $m_l = 0$
(ii) $l = 1$

(b) The atomic weight of naturally occurring neon is 20.18 amu.

Naturally occurring neon is composed of two isotopes:

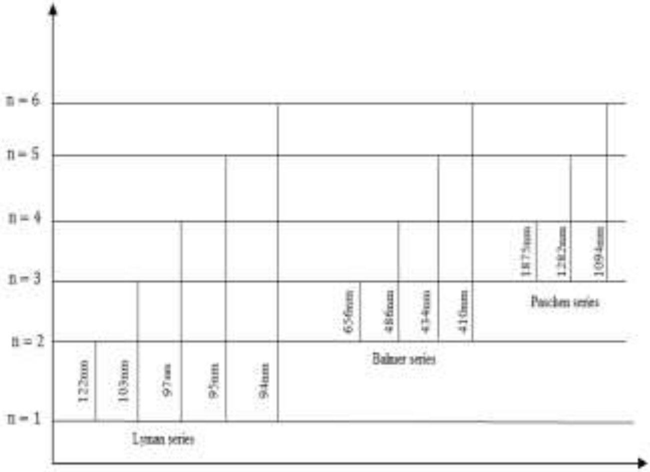
$Ne-20: 19.99amu$

$Ne-22: 21.99amu$

Calculate the number of $Ne-22$ atoms in $12.55g$ sample of naturally occurring neon.

Question 38

Consider the following diagram:



(a) The above diagram illustrates the Bohr model of the hydrogen atom

- i) Explain what horizontal and vertical lines represent.
- ii) Explain how this model explains emission line spectral and absorption line spectra.

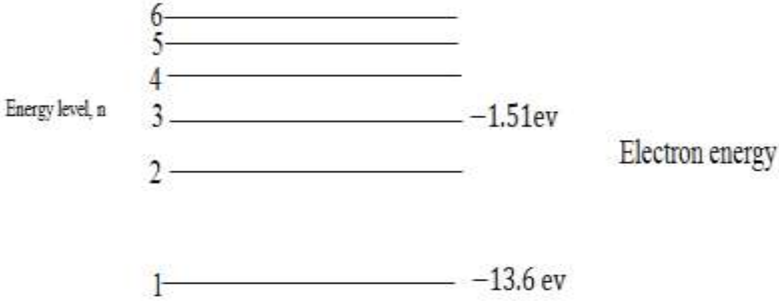
(b) The line between n = 5 and n = 2 is labeled 434nm. Show that this is correct for a hydrogen atom

(c) Calculate the energy of a hydrogen atom in its ground state. Express your answer in ev.

Question 39

The electron in the hydrogen atom emits or absorbs electromagnetic radiation when it moves between energy levels.

The visible part of the spectrum emitted by hydrogen can be seen in the laboratory by applying a high voltage to a hydrogen discharge tube. The diagram below represents some of the electron energy levels in the hydrogen atom.



(a) To which energy level does the electron drop when it emits visible light?

(b) Absorption spectrum for hydrogen gas consists of a series of dark lines within the full spectrum of colours. Explain clearly how the dark line in the red part of the spectrum is produced.

(c) Calculate the frequency of the photon produced when an electron drops from the second excited state to the ground state.

An electron in energy level 4 jumps to a higher energy level, and then drops down to the ground state, releasing a photon of frequency 3.2×10^{5}Hz.

(d) Calculate the frequency of photon required for the first jump.

Question 40

Nuclear reactions in the sun produced light. The main element in the sun is hydrogen. The spectrum of hydrogen can be observed in the laboratory with a hydrogen discharged tube. The visible lines in the hydrogen spectrum are called the Balmer series and are described by the formula: $\frac{1}{\lambda} = \mathrm{R}\left(\frac{1}{\mathrm{S}^2} - \frac{1}{\mathrm{L}^2}\right)$

Where $\mathrm{s} = 2$

(a) Calculate the wavelength of the lowest frequency line in the Balmer series.

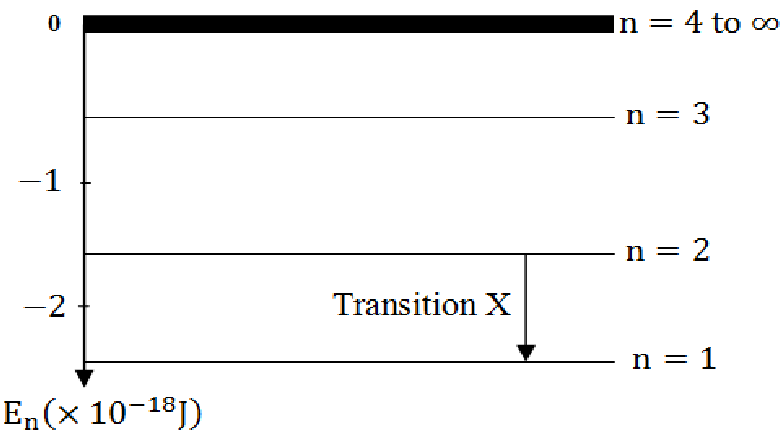
(b) Explain how light of this particular frequency is produced in the hydrogen atom

(c) An electron in the 6^{th} excited state (L=7) returns to the ground state in two jumps. It releases one photon with a wavelength of 2.165×10^{-6}m. what is the wavelength of the second photon?

Question 41

Light from stars is photons of electromagnetic radiation created by electron transitions between energy states. These photons produce spectra that identify the atoms that are producing light. A common element that produces light from star is hydrogen.

The possible energy states (levels) of the hydrogen atom electrons are shown in the next page.



(a) In which part of this electromagnetic spectrum is the radiation emitted by transition X?

(b) Calculate the wavelength of the photons emitted by the transition.

(c) Explain which transition produces the red line in the visible part of the hydrogen atom.

(d) In order for an electron in a hydrogen atom to move from the third energy level to the fifth energy level, a photon of electromagnetic radiation must be absorbed. Calculate the energy of this photon.

Question 42

Three experiments or pieces of scientific equipment were key to our present understanding of atomic structure. For each of the following, outline major discovery that it contributed to:

(a) Crookes tube

(b) Rutherford's gold foil experiment

(c) The bright line spectrum for hydrogen atom

Question 43

(a) Write the electronic configuration of the following by the orbital method:

(i) N^{3-} (ii) Ar (iii) Sc^{3+}.

(b) The human eye can detect 3.15×10^{-17}J of 510nm radiation. How many photons does this correspond?

Question 44

(a) Write atomic number of an atom whose dispositive ion has electronic configuration of $1s^2 2s^2 2p^6$.

(b) What is the maximum number of electrons in an atom that can have the following quantum numbers?

(i) $n = 4, l = 3, m_l = -3$

(ii) $n = 5, l = 3$

(iii) $n = 2, m_s = -{}^1/_2$

Question 45

(a) An atom X may gain electrons to form anion; X^{3-} with the electronic configuration of $1s^2\ 2s^2\ 2p^6$. What is the atomic number of X.

(b) Most of cell phones use electromagnetic radiation in the radio-frequency range, i.e.100KHz to 1GHz; calculate the maximum energy that a photon in this range can have. Convert it to a per mole basic and compare to the energy of covalent bond ($>$ 100kJ/mol). Based on this comparison, how likely is cell phone radiation to cause breaking of chemical bonds in the body?

Question 46

A sample of chlorine containing 75% of the isotope chlorine-35 and 25% of the isotope chlorine-37 was analysed by mass spectrometry. Three corresponding to Cl_2^+ were recorded.

(a) Why were three peaks recorded?

(b) What are the relative masses of the three peaks

(c) What are the relative amounts in the three peaks?

Question 47

What maximum number of electrons may have the following quantum numbers?

(i) $n = 2, m_l = 0$
(ii) $n = 3, l = 1$
(iii) $n = 2, l = 1$
(iv) $n = 3, l = 1, m_l = -1$
(v) $n = 3, l = 1, m_l = 0$ $s = {}^{+1}/_{2}$

Question 48

(a) Why the following set of quantum numbers is invalid?

(i) $n = 3, \quad l = 2, m_l = +1, m_s = +1$
(ii) $n = 4, \quad l = 3, \; m_l = -4, m_s = +{}^{1}/_{2}$

(b) An electron and proton have the same kinetic energy and are moving at non-relativistic speeds. Determine the ratio of the De-Broglie wavelength of the electron to that of the proton.

Give that:

Mass of an electron$= 9.11 \times 10^{-31}$kg

Mass of proton $= 1.67 \times 10^{-27}$kg

Question 49

An electron and proton both moving at non-relativistic speeds have the same De-Broglie wavelength. Which one of the two particles has greater of the following?

(i) Momentum
(ii) Speed
(iii) Kinetic energy
(iv) Frequency

Question 50

(a) What designation is given to the orbitals having:
(i) $n = 2, \; l = 0$
(ii) $n = 1, \; l = 0$

(iii) $n = 3, \;\; l = 2$

(iv) $n = 4, \;\; l = 1, \quad m_l = -1$

(b) An electron moves in a straight line with a constant speed $V = 1 \times 10^6 \mathrm{m/s}$ which has been measured to a precision of 0.2%.What is the maximum precision with which its position could be simultaneously measured? Mass of an electron$= 9.11 \times 10^{-31}$kg

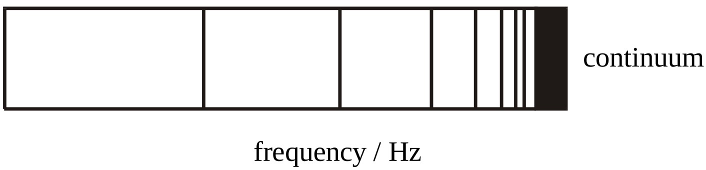
Question 51

(a) Give an account for an experiment which led Rutherford to conclude that every atom has heavy positively charged nucleus which occupies small volume.

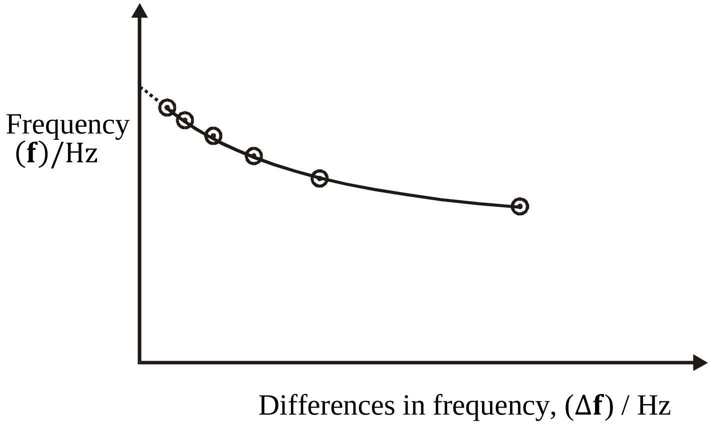
(b) A measurement established the position of a proton with an uncertainty of 1.5×10^{-11}m; find the minimum uncertainty in the proton's position in 2 seconds later (Mass of a proton$= 1.67 \times 10^{-27}$kg)

Question 52

The emission spectrum of an element is seen as a series of bright coloured lines on a dark background.



Within a series the intervals between the frequencies of each line decrease until the lines are so close together that they converge to form a **continuous spectrum** or **continuum** as shown in the diagram. A graphical method can be used to find the start of the continuum. A plot of **f** against Δ**f** can be extrapolated back to find where Δ**f** is 0. This is the start of the continuum.



(a) What causes a line in an emission spectrum?

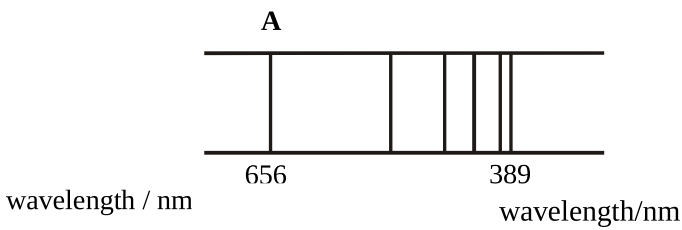
(b) Why do the lines converge as they reach the continuum?

(c) (i) Calculate the energy, in kJ/mol, of the emission line at the start of the continuum if the curve $\Delta \mathbf{f}$ intersects the *y*-axis ($\mathbf{f}$) at 1.26×10^{15}Hz.

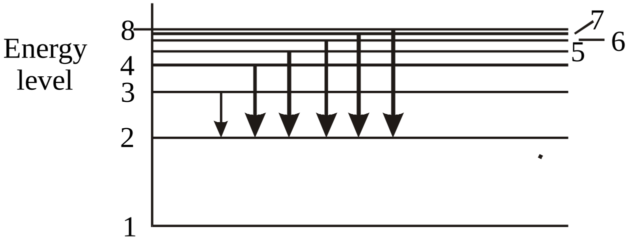
(ii) What does this energy represent?

Question 53

Below is a simplified diagram of the Balmer series in the emission spectrum of atomic hydrogen.



Spectral lines arise as a result of electronic transitions in atoms. The Balmer series is produced by the transitions shown in the following diagram.

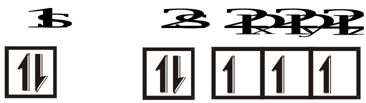


(a) What transition corresponds to a line A in the spectrum? Explain your answer.

(b) Calculate the energy difference, in $kJmol^{-1}$, that gives rise to line A, with wavelength 656 nm.

Question 54

The electron configuration for nitrogen is:



(a) What do the symbols ↑ and ↓ represent?

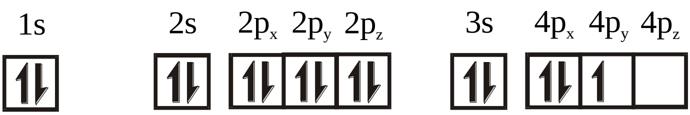
(b) What is the significance of *x*, *y* and *z* in the 2p sublevel?

(c) (i) Describe the shape of the s and p orbitals.

(ii) Describe the position of the p orbitals relative to each other.

(d) Why is the $2p_z$ electron for nitrogen not placed in the $2p_x$ or $2p_y$ orbital?

(e) Phosphorus is in the same group as nitrogen but has 15 electrons. **Clyfar**, a form five student wrote the following configuration for phosphorus:



Explain the two mistakes in the **Clyfar's** answer**.**

Question 55

There are four statements that you have come across in your study of electrons and atomic orbitals. These statements are:

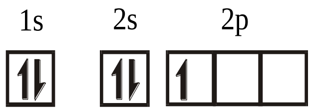
(1) The Aufbau principle

(2) Heisenberg's uncertainty principle

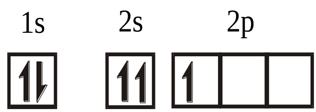
(3) The Pauli Exclusion Principle

(4) Hund's rule of maximum multiplicity

(a) The electronic configuration for boron is given by (i) and not (ii).



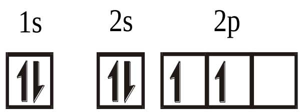
(i)



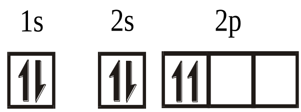
(ii)

Explain why (ii) is wrong and identify which of the above statements justifies your choice.

(b) The electronic configuration for carbon is given by (iii) and not (iv).



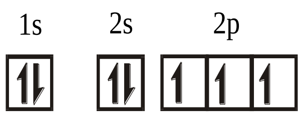
(iii)



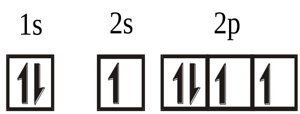
(iv)

Explain why (iv) is wrong and identify which of the above statements justifies your choice.

(c) The electronic configuration for nitrogen is given by (v) and not (vi).



(v)



(vi)

Explain why (vi) is wrong and identify which of the above statements justifies your choice.

Question 56

(a) Draw diagrams, including axes, to represent a 2s orbital and the three 2p orbitals.

(b) What does an orbital diagram represent?

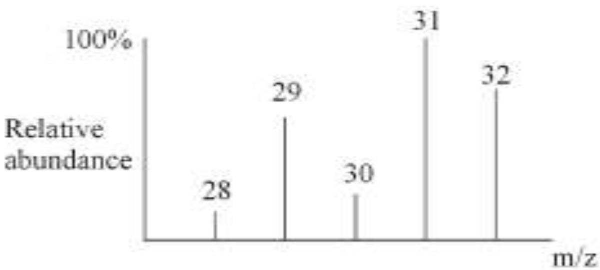
(c) What is the significance of the number 2 in the terms 2s and 2p?

(d) The three 2p orbitals are often degenerate. What does the term ‘degenerate’ mean in this context?

(e) Draw an energy level box diagram to represent the relative energies of the 1s, 2s and 2p orbitals in an isolated atom.

Question 57

The mass spectrum of a molecule of empirical formula CH_4O is shown below.



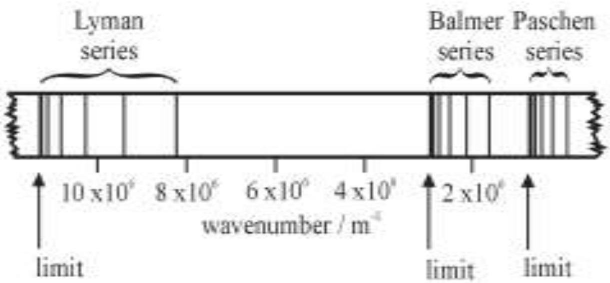
(a) Suggest a formula for an ion for each peak in the above spectrum. Identify the parent ion.

(b) What is measured by the peak height in the above spectrum?

(c) Occasionally two electrons can be removed from each fragment produced. Where will these peaks appear in the mass spectrum?

Question 58

The diagram below represents a section of the line emission spectrum for hydrogen.



(a) If the spectrum of hydrogen is viewed through a spectroscope only one set of lines is seen.

(i) Why is this?

(ii) Which series is seen?

(b) Explain how any particular line in this spectrum is produced.

(c) The last line or convergence limit of the Lyman series has a wavenumber of approximately 11×10^6 m^{-1}. Calculate the energy equivalent of this in $kJmol^{-1}$.

(d) What does the energy referred to in (c) correspond to?

Question 59

Imagine a universe in which the value of the magnetic spin quantum number, m_s, can be + ½, 0, or – ½. Assuming that all of the other quantum numbers can take only the values possible in our world and the Pauli Exclusion Principle applies, give the following:

(a) The new electronic configuration of phosphorus.

(b) The atomic number of the element with a complete n=3 shell.

(c) The number of unpaired electrons in aluminum.

Question 60

(a) State:

(i) Aufbau principle

(ii) Hund's rule of maximum multiplicity

(iii) The uncertainty principle

(b) The wavelength of lines in the balmier series of hydrogen spectrum is given by the expression:

$$\frac{1}{\lambda} = R_H\left(\frac{1}{2^2} - \frac{1}{n^2}\right)$$

Where n is an integer greater than 2.

(i) Draw an energy level diagram to show the origin of second and fourth lines in the Balmer series.

(ii) Calculate the frequency of the first line in the Balmier series. $(R_H = 1.09678 \times 10^7\ m^{-1})$

Question 61

(a) Explain why ground state electronic configuration of Cr and Cu are different from what might be expected.

(b) What is the total number of electrons that can be held in all orbitals having the same principle quantum number n?

Question 62

The following value is the only allowed energy levels of hypothetical one-electron atom.

$$E_6 = -2 \times 10^{-19}J \quad E_5 = -7 \times 10^{-19}J$$

$$E_4 = -11 \times 10^{-19}J \quad E_3 = -15 \times 10^{-19}J$$

$$E_2 = -17 \times 10^{-19}J \quad E_1 = -20 \times 10^{-19}J$$

(a) If the electron was in the n = 3 level, what would be the highest frequency and minimum wavelength of radiation that could be emitted?

(b) What is the ionisation energy (in kJ/mol) of the electron in its ground state?

(c) If the electron was in the $n = 4$ level what would be the shortest wavelength of radiation that could be absorbed without causing ionisation.

Question 63

The energy of the electron in hydrogen atom in the ground state is given by;

$$E_1 = \frac{-2.178\times10^{-18}}{n_1^2}\text{Joules}$$

The energy of the same electron if it occupies a higher level n_2 is given by:

$$E_2 = \frac{-2.178\times10^{-18}}{n_2^2}\text{Joules}$$

(i) Why is the energy negative?

(ii) Calculate the energy in joules and the wavelength in metres of the light which must be absorbed by atom to excite its electron from $n = 1$ to $n = 2$.

Question 64

(a) With reference to calcium at ground state, how many electrons have the following quantum numbers?

(i) $n = 3$

(ii) $m_l = 0$

(iii) $l = 1$

(b) Soap bubbles pick up colour because they reflect light with wavelength equal to the thickness of the walls of the bubble .What frequency of light will be reflected by a soap bubble of 6 nanometers thick?

Question 65

(a) Write the electronic configuration of

(i) Cu

(ii) Mg^{2+}

(iii) Cr

(iv) Li^{+}

(b) State the postulates of Bohr's atomic theory, briefly state the shortcomings of the theory

Question 66

(a) Define the following terms
(i) Quantization of angular momentum
(ii) Azimuthal quantum number
(iii) Wave particle duality of matter

b) Briefly explain the physical significance of the state $n = \infty$ and $E = 0$ for hydrogen atom.

Question 67

(a) The nucleus of certain element is presented as $^{15}_{7}\mathbf{X}$. Clearly deduce the number of each of the fundamental particles in atom **X**.
(b) An electromagnetic radiation of wavelength $2420A°$ is sufficient to ionize the sodium atom. Calculate the ionisation energy of sodium atom in kJ/mol.

Question 68

(a) List down four information that can be obtained from principal quantum number.
(b) Calculate the wavenumber for the longest wavelength transition in Balmier series of atomic hydrogen.

Question 69

When an electron jumps from certain higher energy level, E_2, to its ground state E_1, green light in the Ballmer series is emitted. If the energy released during this transition is 4.071×10^{-19}J, Determine the:-

(i) Wavelength of the green light
(ii) Higher energy level E_2, from which the electron jumps to the ground energy level, E_1.

Question 70

The atomic spectrum of hydrogen in the visible region is given by the following relationship;

$$\frac{1}{\lambda} = R_H\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

(a) What do symbol λ, R_H, n_1 and n_2 represent
(b) Calculate the frequency of the third line of visible spectrum.

Question 71

(a) Samples of pure silicon obtained from natural silicon ores, mined in different parts of the world, have slightly different relative atomic masses. Explain.
(b) A hydrogen emission spectral line in a infra-red region at 1875nm corresponds to a transition from a higher level to n=3 level. Calculate the value of n for the higher energy level.

Question 72

(a) What is the maximum number of orbitals with:
(i) $n = 2, l = 1$
(ii) $n = 2, l = 2$
(iii) $n = 3, l = 2$
(iv) $n = 5, l = 1, m_l = -1$

(b) The first ionisation energy of sodium atom is 145.684kJ/mol. Calculate the wavelength of electromagnetic radiation in angstrom which is sufficient to ionize the atom.

Question 73

(a) From Heisenberg equation; show that Heisenberg's uncertainty principle is true.
(b) Alpha particles emitted from Barium have energy of 4.8Mev. Given that a mass of alpha particles is 6.6×10^{-27}kg and 1Mev $= 10^6$eV; calculate De-Broglie's wavelength.

Question 74

(a) X is the hypothetical element with two isotopes; ^{m}X and ^{n}X. If the element, X exists as triatomic molecule, X_3, explain how many peaks the element will show in its mass spectrum.
(b) Calculate the energy emitted when electrons of 1g atoms of hydrogen undergo transition giving line of lowest energy in the visible region of its atomic spectrum.

Question 75

(a) The mass number of two atoms, **X** and **Y** with the same atomic number are 35 and 37 respectively. If **X** contains 18 neutrons in its nucleus, find the number of neutrons and electrons in **Y**.

(b) State whether each of the following sets of quantum number is permissible for an electron in an atom. If a set is not permissible, explain why?

(i) $n = 1,\ l = 1,\ m_l = 0,\ m_s = +{}^1/_2$

(ii) $n = 3,\ l = 1,\ m_l = -2,\ m_s = -{}^1/_2$

(iii) $n = 2\ \ l = 1,\ m_l = 0,\ m_s = +{}^1/_2$

(iv) $n = 2\ ,\ l = 0\ \ m_l = 0\ \ m_s = -1$

Question 76

In order to obtain a mass spectrum of an element, a gaseous sample of the element is first ionised.

(a) Describe how ionisation is achieved in mass spectrometer.
(b) Give three reasons why ionisation is necessary.

Question 77

Some data obtained from the mass spectrum of a sample of carbon are given below.

Ion	$^{12}C^+$	$^{13}C^+$
Absolute mass of one ion/g	1.993×10^{-23}	2.158×10^{-23}
Relative abundance %	98.9	1.1

Use these data to calculate a value for:

(i) the mass of one neutron,
(ii) the relative atomic mass of ^{13}C and
(iii) the relative atomic mass of carbon in the sample.

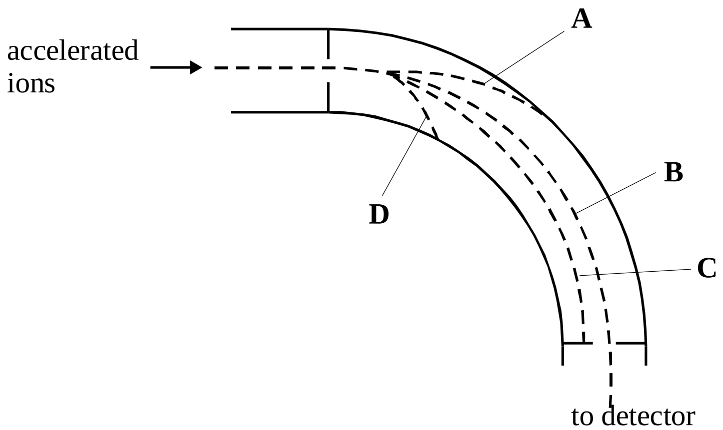
You may neglect the mass of an electron.

Question 78

The diagram below shows a section of mass spectrometer between the acceleration stage and the detection stage. The accelerated ions are from a sample of krypton which has been ionised as follows:

$$Kr(g) \longrightarrow Kr^{+}(g) + e$$

The ions are deflected in four distinct paths, **A, B, C** and **D.** Ions are detected and a mass spectrum is then produced.



(i) What accelerates the Kr^{+} ions before being deflected?

(ii) What deflects the moving ions around a curved path?

(iii) Why do the Kr^{+} ions from this sample of krypton separate into four paths?

(iv) What adjustment could be made to the operating conditions of the mass spectrometer in order to direct the ions following path C onto the detector?

(v) For each type of ion what two measurements can be made from the mass spectrum?

(vi) What factors, other than the mass to charge ratio of an ionised particle, determines how much that particle is deflected in a magnetic field of a given strength.

Question 79

The table below shows some data about fundamental particles.

Particle	**Proton**	**Neutron**	**Electrons**
Mass /g	1.6725×10^{-24}	1.6748×10^{-24}	0.0009×10^{-24}
Relative charge			

(a) Complete the table by giving a value for the relative charge of each particle.

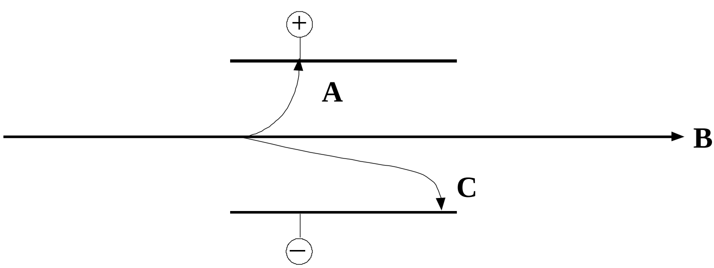
(b) Calculate the mass of an atom of hydrogen which is made from a proton and an electron.

(c) Calculate the mass of one mole of such hydrogen atoms giving your answer to four decimal places (the Avogadro's constant = $L = 6.0225 \times 10^{23} \text{mol}^{-1}$).

(d) An accurate value for the mass of one mole of hydrogen atoms is 1.0080g. Give one reason why this value is different from your answer to part (c).

Question 80

The diagram in the figure below shows the behaviour of the three fundamental particles when passes through an electric field.



(i) Identify the particles represented by **A**, **B** and **C**.

(ii) Explain the shapes and directions of the path traced by the fundamental particles as they pass through the electric field.

Question 81

The following questions refer to the operation of a mass spectrometer.

(a) Name the device used to ionize atoms in a mass spectrometer.
(b) Why is it necessary to ionize atoms before acceleration?
(c) What deflects the ions?
(d) What is adjusted in order to direct ions of different mass to charge ratio onto the detector?

Question 82

A sample of copper contain the two isotope $^{63}\mathrm{Cu}$ and $^{65}\mathrm{Cu}$ only. It has a relative atomic mass, $\mathrm{A_r}$, less than 64. The mass spectrum of this sample shows major peaks with m/z value of 63 and 65, respectively.

(i) Explain why the $\mathrm{A_r}$ of this sample is less than 64.
(ii) Explain how Cu atoms are converted into $\mathrm{Cu^+}$ ions in a mass spectrometer.
(iii) In addition to the major peaks at $\mathrm{m/z} = 63$ and 65, much smaller peaks at $\mathrm{m/z} = 31.5$ and 32.5 are also present in the mass spectrum. Identify the ion responsible for the peak at $\mathrm{m/z} = 31.5$ in the mass spectrum.
(iv) Explain why your chosen ion has this $\mathrm{m/z}$ value and suggest one reason why this peak is very small.

Question 83

(a) Ozone in the upper atmosphere absorbs light with wavelength of 220 to 290nm. What are the frequency (in Hz) and energy (in J) of the most energetic of these photons?
(b) Carbon-carbon bonds form the backbone of nearly every organic and biological molecule. The average bond energy of $\mathrm{C-C}$ bond is $347\mathrm{kJmol^{-1}}$. Calculate the wavelength (in nm) of the least energetic photon that can break this bond.

Compare this value to that absorbed by ozone and comment on the ability of the ozone layer to prevent $\mathrm{C-C}$ bond disruption.

Question 84

(a) Boron has an atomic mass of 10.81amu according to the periodic table. However, no single atom of boron has a mass of 10.81amu.Suggest a reason of this.

(b) Naturally occurring iodine has an atomic mass of 126.9045. A 12.3849g sample of iodine is accidentally contaminated with 1.00007g of I-129, a synthetic radioisotope of iodine used in the treatment of certain diseases of thyroid gland. The mass of I-129 is 128.9050amu. Find the apparent “atomic mass” of the contaminated iodine.

Question 85

Two important concepts that relate to the behaviour of electrons in atomic system are the **Heisenberg uncertainty principle** and the **wave-particle duality** of matter.

(a) State the Heisenberg uncertainty principle as it relates to determining the position and momentum of an object.

(b) What is the meaning of “Wave-particle duality” and derive the equations which justify the concept.

(c) What aspect of the Bohr Theory of the atom is considered unsatisfactory as result of Heisenberg uncertainty?

(d) Explain why the wave nature of particles is not significant when describing the behaviour of macroscopic objects, but is very significant when describing the behaviour of electrons.

Question 86

The emission spectrum of hydrogen consists of several series of sharp emission lines in the ultraviolet, visible and infrared regions of the spectrum.

(a) Mention series which are found in the infrared regions

(b) What feature of the electronic energies of the hydrogen atom explains why the emission spectrum consists of discrete wavelengths rather than a continuum of wavelengths?

(c) Account for existence of several series of lines in the spectrum. What quantity distinguishes one series of lines from another?

(d) Draw an electronic energy level diagram for the hydrogen atom and indicate on it the transition corresponding to the line of lowest frequency in the Balmer series

(e) What is the difference between an **emission spectrum** and an **absorption spectrum**? Explain why the absorption spectrum of atomic hydrogen at room temperature has only the lines of the Lyman series.

Question 87

The Rydberg equation enables you to calculate the frequency of a line in the hydrogen spectrum. The version of the Rydberg equation in terms of frequency is:

$$f = cR_H\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

(a) Calculate the frequency of the line produced when and electron falls back from the infinity level to the $n = 1$

(b) Write the equation which relates the energy gap between two levels and the frequency of the light emitted

(c) Use the equation you have written in (b) to calculate the energy in eV needed to move an electron form the 1- level to the infinity level. State clearly any assumption you have made in your calculations.

(d) Use the result obtained in (c) to calculate the ionisation energy of hydrogen.

Question 88

(a) What is the maximum number of electrons that an orbital with a magnetic quantum number of 2 could hold?

(b)

(i) How many electrons on an atom of argon have m=+1?

(ii) Explain why are up to ten electrons on a set of d-orbitals?

(iii) What is the maximum number of electrons that are allowed to have the following set of quantum numbers in one atom? $n = 4$ and $m_l = +2$

(iv) Which element has the last electron in its atom with the following quantum numbers?

$$n = 3, \quad l = 1, \quad m_l = -1, \quad m_s = -\frac{1}{2}$$

Question 89

(a) Mention the rule violated in each of the following electronic configuration

1s	2s	2p		
↑	↑↓			

1s	2s	2s		
↑↓	↑↓	↑↓		

1s	2s	2p		
↑↓	↑↓			

(b)

Give set of quantum numbers of

(i) Last electron in copper

(ii) 26^{th} electron in cobalt

Question 90

(c) Explain in your own words what is meant by:
(i) The Pauli Exclusion principle
(ii) Hund's rule
(iii) A line in an atomic spectrum
(iv) The principle quantum number

(d) An element **X** has 2 electrons in K shell, 8 electrons in L shell, 13 electrons in M shell and 2 electrons in N shell. Deduce the following:
(i) Its full electronic configuration
(ii) Atomic number of **X**
(iii) Total number of principle quantum numbers with electrons in an atom of **X** at ground state
(iv) Total number of sublevels with electrons in an atom of **X** at ground state
(v) Total number of unpaired electrons in an atom of **X** at ground state
(vi) In which block of periodic table does element **X** belong?

Question 91

(a) *"The electron in the atom is located at a definite distance from the nucleus in an orbit which has discrete energy"*
(i) Which atomic model is this statement based?

(ii) Challenge this statement according to wave mechanical model of the atom?

(b) Give two examples to show that electronic configuration of half-filled and completely-filled orbitals are more stable.

(c) Calculate the energy emitted when electrons of 1g atoms of hydrogen undergo transition giving line of lowest energy in the visible region of its atomic spectrum. Show your work clearly including manipulation of units.

Question 92

(a) Define the following terms:
(i) Nuclides
(ii) Isotones
(iii) Isotopes

(b) Paul Exclusion principle may be stated as; *"No two electrons in an atom may have all four quantum numbers the same."* Explain two pieces of information which may be extracted from the principle.

(c) The first ionisation energy of sodium atom is 145.684kJ/mol. Calculate the wavelength of electromagnetic radiation in angstrom which is sufficient to ionize the atom.

Question 93

(a) Define the following:
(i) Atom
(ii) Element

(b) What is the relationship between the two terms in (a) above?

(c)
(i) How does the atomic spectrum differ from continuous spectrum?
(ii) Calculate the wave number for the longest wavelength in the Balmer series of atomic hydrogen.

Question 94

(a) Define the following terms:
(i) Quantum number
(ii) Atomic spectrum
(iii) Mass spectrometer

(b) Naturally occurring bromine consist of two isotopes, ^{79}Br and ^{81}Br. In determination of relative atomic mass of bromine, the following mass spectrum data was obtained, which showed peaks at masses 158, 160 and 162 a.m.u. The heights of the peaks were in the ratio of 6:1:5 respectively.
(i) Sketch the mass spectrum data and label it properly
(ii) Calculate the average atomic mass of bromine
(iii) Why RAM of bromine is not a whole number
(iv) Calculate the relative abundancies of ^{79}Br and ^{81}Br

Question 95

(a) State the following:
(i) Planck's quantum theory
(ii) Heisenberg uncertainty principle
(iii) Four shortcomings of Bohr's atomic theory

(b)
(i) What is ionisation energy?
(ii) By using Rydberg equation, calculate ionisation energy of hydrogen

(c) If the wavelength of the first line in the Balmer series in a hydrogen spectrum is $6863A^{O}$, calculate the wavelength of the first line in the Lyman series in the same spectrum.

Question 96

(a) At different times scientists have proposed various descriptions or models of the atom to match experimental evidence available. The model that Thomson proposed was called the plum-pudding model.
(i) Why the model is known as plum-pudding model?
(ii) Describe the model.
(iii) List at least two advantages and two disadvantages of the model.

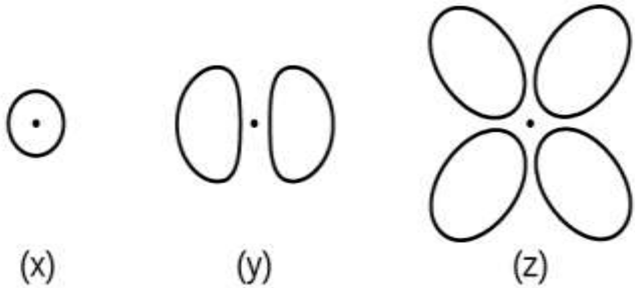
(b)
(i) When you charged a strip of plastic by rubbing it through your fingers or on cloth, you actually transferred electrons onto the plastic strip. Using the terms **"electrons"** and **"protons"**, and **"negative"** and **"positive"**, explain why the strip was attracted to your fingers or the cloth you rubbed it on.

(ii) When you rub a balloon on your hair, electrons are transferred onto the balloon. Using the terms **"electrons"** and **"protons"**, and **"negative"** and **"positive"**, explain why a rubbed balloon is attracted to and sticks to a wall even though you didn't rub the balloon on the wall.

Question 97

(a) Define the following:
- (i) Electronic configuration
- (ii) Azimuthal quantum number
- (iii) Stark effect
- (iv) Quantum number

(b) Consider orbitals shown in the figure below:



- (i) What is the maximum number of electrons contained in each of orbitals of type **Y**?
- (ii) How many orbitals of type **Z** are found in a shell with $n = 2$?
- (iii) An electron residing in a shell with principal quantum number of 3, in an orbital of type **X** is spinning in clockwise direction. Write set of all four quantum numbers for the electron.
- (iv) What is the smallest possible value of n for an orbital of type **Y**?
- (v) What are the possible l and m_l values for an orbital of type **Z**?

(c) An organic fertiliser was analysed using a mass spectrometer. The spectrum showed that the nitrogen in the fertiliser was made up of 95.12% ^{14}N and 4.88% ^{15}N.

(i) Calculate the relative atomic mass of the nitrogen found in this organic fertiliser.

(ii) In a mass spectrometer, under the same conditions, monopositive ions of ^{14}N and ^{15}N follow different paths.

A: State the property of these ions that causes them to follow different paths.

B: State **one** change in the operation of the mass spectrometer that will change the path of an ion.

(iii) Organic fertilisers contain a higher proportion of ^{15}N atoms than are found in synthetic fertilisers. State and explain whether or not you would expect the chemical reactions of the nitrogen compounds in the synthetic fertiliser to be different from those in the organic fertiliser. (Assume that the nitrogen compounds in each fertiliser are the same).

Question 98

(a) State the following:

(i) Aufbau principle

(ii) Paul's exclusion principle

(iii) Hund's rule of maximum multiplicity

(iv) Half-filled and full-filled orbital rule

(b) A diode laser emits a wavelength of 987nm. All of radiations it emits is measured with a total energy of 0.52J over a period of 32 seconds. Calculate the number of photons per second emitted by the laser.

(c) For each of the following statement, state whether you agree or disagree and give reason(s) to defend your answer:

(i) If you know the atomic number of an element in the periodic table, you will also know the number of neutrons in any atom of that element.

(ii) Different atoms of the same element can have different atomic masses.

Question 99

(a) (i) State four postulates of Dalton's atomic theory.

(ii) Why different atoms have different chemical properties? Briefly explain.

(b) Write the chemical symbol $\left({}^{Z}_{A}X\right)$ and orbital electronic configuration for the atoms described in the following table:

SN	Number of Neutrons	Number of Electrons
(i)	13	11
(ii)	7	8
(iii)	17	18
(iv)	16	16

(c) Calculate the minimum energy required to remove an electron from the hydrogen atom in its ground state.

Question 100

(a) Define the following:
- (i) Mass number
- (ii) Relative atomic mass

(b) Clearly state four postulates of Planck's quantum theory as derived from black body radiation.

(c) Indium is in Group IIIA in the Periodic Table and exists as a mixture of the isotopes: $\text{In} - 113$ and $\text{In} - 115$. A sample of indium must be ionised before it can be analysed in a mass spectrometer.
- (i) State what is used to ionise a sample of indium in a mass spectrometer.
- (ii) Write an equation, including state symbols, for the ionisation of indium that requires the minimum energy.
- (iii) State why more than the minimum energy is **not** used to ionise the sample of indium.
- (iv) Give two reasons why the sample of indium must be ionised.
- (v) By reference to the relevant part of the mass spectrometer, explain how the abundance of an isotope in a sample of indium is determined.
- (vi) A mass spectrum of a sample of indium showed two peaks at *m/z* = 113 and *m/z* = 115. The relative atomic mass of this sample of indium is 114.5. Calculate the percentage abundances of the two isotopes.

Part two
SOLUTIONS

Question 1

(a)

(i) By adding number of protons and number of neutrons of the atom.

(ii) In the atom, protons and neutrons are heavier sub-atomic particles contributing to almost the entire mass of the atom. Consequently, their total number is almost equal to the mass of the atom in atomic mass unit (a.m.u).

(b) Using; $A_r = \frac{\sum(\text{ Isotopic mass} \times \text{intensity})}{\text{Total intensity}}$

$$= \frac{23.985 \times 39.35 + 24.985 \times 5.065 + 25.982 \times 5.585}{39.35 + 5.065 + 5.585} = 24.3 \text{ a.m.u}$$

The average atomic mass is 24.3 a.m.u

Question 2

(a) The number of electrons is equal to the number of protons.

(b) The negative charge of an electron is opposite to the positive charge of proton; but they are of equal magnitude. So to have the neutral atom there must be the same number of protons and electrons.

$$A_r = \frac{\sum(\text{Isotopic mass} \times \text{percentage adunbance})}{\text{Total percentage adunbances}}$$

$$= \frac{19.992 \times 90.22 + 20.994 \times 0.257 + 21.991 \times 8.82}{90.22 + 0.257 + 8.82} = 20.17$$

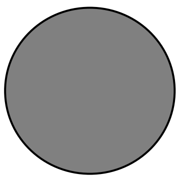
So the relative atomic mass of **X** is 20.17

Question 3

(a)

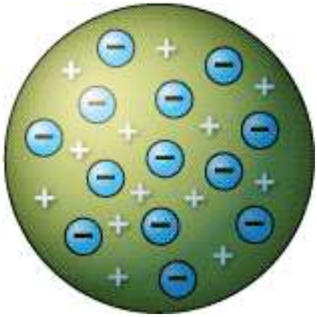
(i) **Name:** Billiard ball atomic model

Diagram:

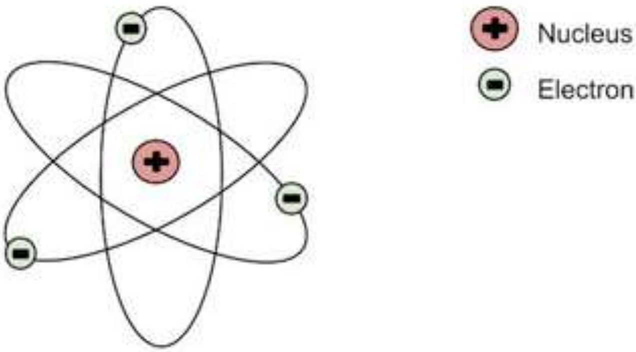


(ii) **Name:** Plum-pudding atomic model

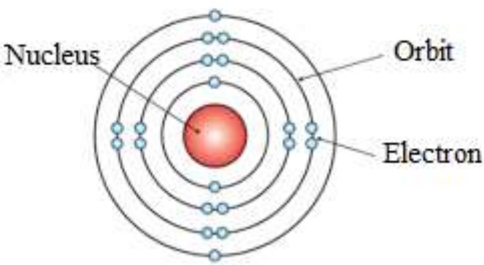
Diagram:



(iii) **Name:** Nuclear atomic model (or planetary atomic model)
Diagram:

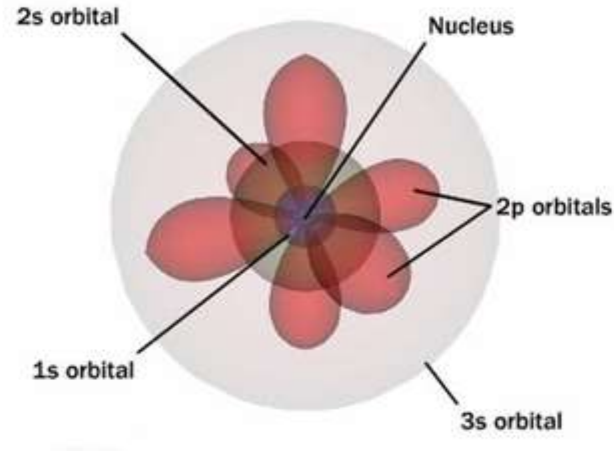


(iv) **Name:** Solar system atomic model
Diagram:



(v) **Name:** Quantum mechanical atomic model (or electron cloud atomic model)

Diagram:



(b) $A_r = \frac{\sum(\text{isotopic mass} \times \text{proportion})}{\text{Total proportions}} = \frac{69\times60+71\times40}{60+40} = 69.8$ (Under assumption that A=RIM)

The average atomic mass is 69.8

Question 4

(i) **Isotope** *is an atom of an element with the same atomic number as that of another atom of the same element but differ in mass number*.

(ii) $^{79}_{35}\text{X}$ and $^{81}_{35}\text{X}$

(iii) $A_r = \frac{\sum(\text{isotopic mass} \times \text{percentage abundance})}{100}$

$$= \frac{79\times50.5+81\times49.5}{100} = 79.99$$

So the relative atomic mass of **X** is 80.0 (To three significant figures)

Question 5

(a) Atomic number of an element is fixed (All isotopes of Cl have the same atomic number of 17) unlike mass number whose value is different for different isotopes and hence it is essential to indicate mass number of Cl so as to differentiate to another Cl atom with mass number of 37.

(b) $A_r = \frac{\sum(\text{isotopic mass} \times \text{percentage abundance})}{100}$

$= \frac{51.88 \times 106.91 + 48.12 \times 108.91}{100} = 107.87\text{g/mol}$

The average atomic mass of Ag 107.87g/mol

Question 6

(a) Rutherford performed alpha rays scattering experiment to demonstrate the structure of atom. Heavy atoms have a heavy nucleus carrying a large amount of positive charge. Hence some alpha particles are easily deflected back on hitting the nucleus. Also a number of alpha particles are deflected through small angles because of large positive charge on the nucleus. If light atoms are used, their nuclei will be light and moreover, they will have small positive charge on the nucleus. Hence, the number of particles deflected back and those deflected through some angle will be negligible.

(b) Percentage abundance of another isotope
$= (100 - 69.15)\%$ or 30.85%
(since summatioin of percentage abundance must be 100%)

Then using $A_r = \frac{m_1P_1 + m_2P_2}{100}$

It follows that $63.546 = \frac{69.15 \times 62.9296 + 30.85m_2}{100}$

From which $m_2 = 64.9277$ amu

Hence the atomic mass of the second isotope is 64.9277 amu

Question 7

(a)

1. Gold is the most malleable metal and therefore thinnest metal layer can be obtained from it.

2. Gold nucleus is heavy and therefore can produce large deflection of α-particles.

(b) $A_r = \frac{\sum(\text{Isotopic mass} \times \text{intensity})}{\text{Total intensity}} = \frac{24\times63+25\times8.1+26\times9.1}{63+8.1+9.1} = 24.3$

Therefore, the relative atomic mass of magnesium is 24.3

Question 8

(a) Isodiapheres are atoms (of different elements) with different atomic number (proton number) and neutron number but the same difference between the number of neutrons and protons. For example $^{39}_{19}K$ and $^{31}_{15}P$ are isodiapheres where the difference between the number of neutrons (20 for K and 16 for P) and protons (19 in K and 15 in P) is 1 ($20 - 19 = 16 - 15 = 1$).

(b) Using $A_r = \frac{\sum(\text{Relative abundance} \times \text{isotopic mass})}{\sum \text{Abundance}}$

But abundance of isotopic is directly proportional to the number of atoms of the isotopes

It follows that: $A_r = \frac{\sum(\text{Number pf atoms of isotope x isotopic mass})}{\text{Total number of atoms in the sample}}$

$$= \left(\frac{370\times7.016+30\times6.015}{400}\right)\text{g/mol} = 6.941\text{g/mol}$$

Hence the average atomic mass of lithium is 6.941g/mol

Question 9

(a) Nuclear charge is defined as the net positive charge experienced by an electron in the orbital of a multi-electron atom. The closer the orbital, the greater is the nuclear charge experienced by the electron (s) in it.

(i) 2s is closer to the nucleus than 3s.Hence 2s will experience larger effective nuclear charge.

(ii) 4d will experience greater nuclear charge than 4f since 4d is closer to the nucleus than 4f.

(iii) 3p will experience greater nuclear charge since it is closer to the nucleus than 3f because 3p is closer to nucleus than 3f.

(b) Using $A_r = \frac{\sum(\text{Number of atoms of isotope} \times \text{isotopic mass})}{\text{Total number of atoms of isotopes}}$

Given that: $\frac{\text{Number of atoms of Rb}-85}{\text{Number of atoms of Rb}-87} = \frac{2.591}{1}$

That is for every 2.591 atoms of $\mathrm{Rb}-85$ there is 1 atom of $\mathrm{Rb}-87$.

So substituting $85.4678 = \frac{2.591\times84.9117+1\times m}{2.591+1}$

Where m stands for isotopic mass of $\mathrm{Rb}-87$

- When solving for 'm' in the above equation, gives;
  $m = 86.9087$ amu

Hence the mass of $\mathrm{Rb}-87$ is 86.9087 amu

Question 10

(a) Quantized energy means that the electrons can possess only certain discrete energy values; values between those quantized values are not permitted.

(b) $^{35}Cl - {}^{35}Cl$ show peak at mass of 70

$^{35}Cl - {}^{37}Cl$ shows peak at mass of 72

$^{37}Cl - {}^{37}Cl$ shows peak at mass of 74

Thus % o f $^{35}Cl = \left(\frac{9+\left(\frac{1}{2}\times6\right)}{9+6+1}\right)\times 100\% = \frac{12}{16}\times 100\% = 75\%$

% of $^{37}Cl = \left(\frac{1+\left(\frac{1}{2}\times6\right)}{9+6+1}\right)\times 100\% = \frac{4}{16}\times 100\% = 25\%$

Thus the relative abundance of ^{35}Cl is 75%

The relative abundance of ^{37}Cl is 25%

Using;

$$A_r = \frac{\sum(\text{isotopic mass} \times \text{percentage abundance})}{100}$$

$$= \frac{75\times35+25\times37}{100} = \frac{2625+925}{100} = 35.5\ \text{a.m.u}$$

Therefore average mass of chlorine atom is 35.5 a.m.u

Question 11

(a) E< G < B = D < C = F < A

(b) Using $\frac{1}{\lambda} = R_H\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$

For Balmer series $n_1 = 2$

So the first line in Balmer series is n= 3, thus n_2= 3 substituting n_1, n_2 and given constant (R_H) to the above equation gives;

$\frac{1}{\lambda} = 1.097 \times 10^7 m^{-1}\left(\frac{1}{2^2} - \frac{1}{3^2}\right)$

From which $\lambda = 6.56 \times 10^{-7}$m

Hence the wavelength of the first line in Balmer series is 6.56×10^{-7}m

Question 12

(a)

(i) $^{36}_{18}Ar^{+}$ having lighter mass, has smaller mass-charge ratio and hence will be more deflected.

(ii) $^{36}_{18}Ar^{2+}$ having greater positive charge has smaller mass-charge ratio and hence will be more deflected.

(b) Using Planck's equation: E = nhf

Where E is the radiation energy emitted

n is the number of photons

h is the Planck constant

f is the frequency of radiations

But $f = \frac{c}{\lambda}$ where λ is the wavelength in metres

Thus $E = \frac{nhc}{\lambda}$ where n = 1 (a photon) and $\lambda = 589 \times 10^{-5}$cm $= 5.89 \times 10^{-7}$m

$$E = \frac{1 \times 6.62 \times 10^{-34} \times 3 \times 10^{8}}{5.89 \times 10^{-7}} \text{J/photon}$$
$$= 3.372 \times 10^{-19} \text{J/photon}$$

Hence energy of a photon of radiation is 3.372×10^{-19}J

Question 13

(a) Most of the instruments used for investigating the structure of the atom are based on the use of measurement of electric charge. As the **neutron is uncharged** particle, it was not detected by these instruments.

(b)

(i) For n = 1: $E = hf = \frac{hc}{\lambda}$;

Where $\lambda = 242.4\text{nm} = 242.4 \times 10^{-9}\text{m}$

$c = 2.998 \times 10^{8}\text{m/s}$ and $h = 6.626 \times 10^{-34}\text{Js}$

Then $E = \frac{6.626 \times 10^{-34} \times 2.998 \times 10^{8}}{242.4 \times 10^{-9}}\text{J/ photon} = 8.195 \times 10^{-19}\text{J/ photon}$

Energy is $8.195 \times 10^{-19}\text{J/photon}$

(ii) One mole of photons has 6.02×10^{23}photons

Thus in the formula $E = nhf = \frac{nhc}{\lambda}$; $n = 6.02 \times 10^{23}$ photons/mol

But $\frac{hc}{\lambda} = 8.195 \times 10^{-19}$ (from (a) above)

So $E = 8.195 \times 10^{-19}\text{J/photon} \times 6.02 \times 10^{23}$ photon/mol

$= 4.9 \times 10^{5}$ J/mol or 4.9×10^{2} kJ/mol

Hence energy of a mole of photon is 4.9×10^{2} kJ/mol.

Question 14

(a) Although potassium atoms have one more proton than argon atoms, the most abundant isotope of argon has 22 neutrons which are 2 more neutrons than potassium which has only 20 neutrons. Thus the increase in number of neutrons of Ar outweighs the increase in number of protons (atomic number) of K, making average atomic weight of Ar close to 40 while that of K is close to 39.

(b) Using $\frac{1}{\lambda} = R_H\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$

For Balmer series $n_1 = 2$

So the first line in Balmer series is n = 3, thus $n_2 = 3$

Given that $\lambda = 6563$ A°

But 1 A° = 10^{-10}m

Therefore, 6563 A° $= 6563 \times 10^{-10}$m $= 6.563 \times 10^{-7}$m

Then substituting value of n_1, n_2 and λ in above equation so as to get the value of R_H

$$\frac{1}{6.563\times10^{-7}\text{m}} = R_H\left(\frac{1}{2^2} - \frac{2}{3^2}\right) = R_H\left(\frac{9-4}{36}\right)$$

From which $R_H = 1.097 \times 10^7 m^{-1}$

But for Lyman series $n_1 = 1$ and $n_2 = 2$ for first member

Then $\frac{1}{\lambda} = 1.097 \times 10^7 m^{-1}\left(\frac{1}{1^2} - \frac{1}{2^2}\right) = 1.097 \times 10^7 m^{-1}\left(\frac{3}{4}\right)$

From which $\lambda = 1.215 \times 10^{-6}$m or 1215A°

Hence the wavelength of the member in Lyman series is 1215A°

Question 15

(a) $\text{Ca} - 40$ is probably the most abundant isotope (because the average atomic mass is usually close to the mass number of the most common isotope).

(b) Energy of one photon is given by the following Planck's equation $E = hf = \frac{hc}{\lambda}$

Substituting $E = \frac{6.63\times10^{-34}\text{J}\times3\times10^8\text{m/s}}{987\times10^{-9}\text{m}} = 2.0152 \times 10^{-19}\text{J/photon}$

But total energy emitted in $32sec = 0.52$J

And the total energy=Energy per photon×Number of photons

Thus Number of photons emitted (in 32s)

$$= \frac{\text{Total energy}}{\text{Energy per photon}} = \frac{0.52\text{J}}{2.0152\times10^{-19}\text{J/photon}} =$$

2.58×10^{18} photons

And number of photons emitted per second

$$= \frac{\text{Total number of photons emitted}}{\text{Time taken}} = \frac{2.58 \times 10^{18}\text{ photons}}{32sec}$$

$$= 8.0625 \times 10^{16}\text{photon/sec}$$

Hence 8.0625×10^{16} photons were emitted in one second.

Question 16

(a) Isotopes are ${}^{238}_{92}\mathrm{U}$, ${}^{234}_{92}\mathrm{U}$

Isobars are: i. ${}^{238}_{92}\mathrm{U}$ and ${}^{238}_{93}\mathrm{Np}$,

ii. ${}^{234}_{90}\mathrm{Th}$, ${}^{234}_{92}\mathrm{U}$, ${}^{234}_{91}\mathrm{Pa}$

(b) $\Delta E = hf = \frac{hc}{\lambda}$; But $\Delta E = E_4 - E_2$

Thus $E_4 - E_2 = \frac{hc}{\lambda}$

$$(-1.362 - -5.498) \times 10^{-19} = \frac{6.63\times10^{-34}\times3\times10^{8}}{\lambda}$$

From which; $\lambda = 4.8 \times 10^{-7}$m or 4800A°

Hence the wavelength is 4.8×10^{-7}m or 4800A°

Question 17

From $\Delta E = hf$ where $f = \frac{c}{\lambda}$ (for n = 1)

$$\Delta E = \frac{hc}{\lambda}$$

But $\frac{1}{\lambda} = R_H\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$

Then $\Delta E = hcR_H\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right) {}^{J}/_{Photon}$

$$\Delta E = 6.63 \times 10^{-34} \times 3 \times 10^{8} \times 1.097 \times 10^{7}\left(\frac{1}{2^2} - \frac{1}{5^2}\right) {}^{J}/_{Photon}$$

$$\Delta E = 2.18 \times 21 \times 10^{-20}\, {}^{J}/_{Photon}$$

But 1 mole of photo contains 6.02×10^{23}photons

Thus energy radiated in J/mol is;

$$2.18 \times 21 \times 10^{-20}\, {}^{J}/_{photon} \times 6.02 \times 10^{23}\, {}^{photon}/_{mol}$$

In kJ/mol = $\frac{2.18\times21\times10^{-20}\times6.02\times10^{23}}{1000}$ kJ/mol = 275.6 kJ/mol

Hence energy radiation in kJ/mol is 275.6kJ/mol

(b) From $\Delta E = hf$

$$f = \frac{\Delta E}{h} = \frac{2.18 \times 21 \times 10^{-20} J}{6.63 \times 10^{-34} Js} = 6.9 \times 10^{14} Hz$$

Hence the frequency of is 6.9×10^{14}Hz

(c) Using $\lambda = \frac{c}{f} = \frac{3\times10^8}{6.9\times10^{14}} m = 4.35 \times 10^{-7} m$

Hence the wavelength of the radiation is 4.35×10^{-7}m

Alternative solution:

Using $\Delta E = \frac{hc}{\lambda}$

From which $\lambda = \frac{hc}{\Delta E} = \frac{6.63\times10^{-34}\times3\times10^8}{2.18\times21\times10^{-20}} m$

$$= 4.34 \times 10^{-7} m$$

Hence the wavelength o f the radiation is 4.34×10^{-7}m.

Question 18

(a)

- Carbon-14 refers to the isotope of carbon with mass number of 14.
- Silver-108 refers to the isotope of silver with mass number of 108.

(b) Using $\Delta E = hf = \frac{hc}{\lambda}$

But $\frac{1}{\lambda} = R_H \left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$

So $\Delta E = hc\, R_H \left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$

In ionisation; the most loosely held electrons is removed i.e. an electron which is found in outmost energy level. For hydrogen this electron is found at energy level of n = 1

So in ionizing hydrogen an electron is removed from energy level of n= 1 to infinity. Thus $n_1 = 1$ and $n_2 = \infty$

Substituting given value of h, c and R_H including values of n_1 and n_2 in above equation of ΔE gives:

$$\Delta E = 6.63 \times 10^{-34} \times 3 \times 10^{8} \times 1.097 \times 10^{7}(1-0)\ {}^{J}/_{electron}$$

$$= 2.18 \times 10^{-18}\ {}^{J}/_{electron}$$

Thus ΔE is 2.18×10^{-18}J for one electron

But from definition of ionisation energy: **ionisation energy must be for one mole of electrons**, that is it must be given in: Units of energy/mol

And one mole of electrons contain Avogadro's number (6.02×10^{23}) of electrons

Therefore ionisation energy of hydrogen:

$= 2.18\times 10^{-18}\ {}^{J}/_{electron} \times 6.02 \times 10^{23}\ {}^{electron}/_{mol}$

= 1312360J/mol or 1312.36kJ/mol

Hence ionisation energy of hydrogen is 1312.36kJ/mol

Alternative solution:

Energy associated with an electron of any energy level of hydrogen is given by:

$E = \frac{-13.6ev}{n^2}$ where n is the energy level

In ionizing hydrogen an electron jump from n = 1 to infinity,

So the ionisation energy for one electron = $E_\infty - E_1 = \Delta E$

But $E_\infty = \frac{-13.6ev}{\infty^2} = 0ev\ (n = \infty)$

And $E_1 = \frac{-13.6ev}{1^2} = -13.6ev\ (n = 1)$

So in ionizing one electron the energy which is required is:

$0 - -13.6ev = 13.6ev$ per one electron

For one mole of electrons the energy become;

$13.6 \times 6.02 \times 10^{23}$ev per one mole

But 1ev = 1.6×10^{-19}J

Hence the ionisation energy is;

$13.6 \times 6.02 \times 1.6 \times 10^{4}$J/mol = 1309952J/mol or 1309.952kJ/mol

Question 19

(a) The average atomic mass is always closer to the nominal mass. From the given data, the mass number of most abundant isotope is 20 and hence the average atomic mass will be about 20.

(b) Generally: $\Delta E = nhf$

But for n = 1 (An electron), the formula becomes; $\Delta E = hf$

But f = $\frac{c}{\lambda}$

Thus $\Delta E = \frac{hc}{\lambda} = \frac{6.626\times10^{-34}\times3\times10^{8}}{4.34\times10^{-7}} J = 4.58 \times 10^{-19} J = \frac{4.58\times10^{-19}}{41.6\times10^{-19}} ev = 2.8625ev$

Hence the energy difference is 4.58×10^{-19}J or 2.8625ev

Question 20

(a) The statement is **correct.**

Explanation:
Due to presence of isotopes, RAM listed in the periodic table is just an average of relative isotopic mass of isotopes of magnesium.

(b) Using $\Delta x \Delta p = \frac{h}{4\pi}$ $\Delta p = 2.5 \times 10^{-14}$gcmsec^{-1}

Converting given Δp into kgmsec^{-1}

1g = 10^{-3}kg 1cm = 10^{-2}m

Thus 2.5×10^{-14}gcmsec^{-1} = $2.5\times 10^{-14} \times 10^{-3} \times 10^{-2}$kgmsec^{-1}

= 2.5×10^{-19}kgmsec^{-1}

Then from $\Delta x \Delta p = \frac{h}{4\pi}$

$\Delta x = \frac{h}{4\pi\Delta p} = \frac{6.626\times10^{-34}}{4\times3.14\times2.5\times10^{=19}}$ m = 2.11×10^{-16}m

Hence the approximate position is 2.11×10^{-16}m from the nucleus.

Question 21

(a) When electrons are excited by the addition of energy, they absorb this energy and move out further from the nucleus. When the electrons drop back closer to the nucleus, they give the off added energy as light. Depending on how far electrons were from the nucleus, they will give off light of different wavelengths in different colours when they drop back.

(b) According to De Broglie's equation

$$\lambda = \frac{h}{mv}$$

Where λ is De Broglie's wavelength

V is the speed o f neutron

m is the mass of neutron

But v = $\frac{5}{100}\times 3\times 10^{8}$m/s = 1.5×10^{7}m/s

$$\lambda = \frac{6.626\times10^{-34}}{1.61\times10^{-27}\times1.5\times10^{7}}$$

= 2.74×10^{-14}m

Hence de – Broglie wavelength is 2.74×10^{-14}m

Question 22

(a) In order for an electron to be ejected from a metal surface the electron must be struck by a single photon with at least the minimum energy to knock the electron loose and hence the proton act as the light particle.

(b) From the uncertainty principle, the minimum uncertainty in the position may be found from the following equation; $\Delta x\Delta p = \frac{h}{4\pi}$

From which: $\Delta x = \frac{h}{4\pi(\Delta p)} = \frac{h}{4\pi(m\Delta c)}$

Where $m = 100g = 0.1kg$ and $\Delta c = 1m/s$

Then $\Delta x = \frac{6.63\times10^{-34}}{4\times3.14\times0.1\times1} = 5.28\times10^{-34}m$

Hence the minimum uncertainty in the position is $5.28\times10^{-34}m$

Question 23

(a) A photon can be emitted when an electron moves either from:

- An excited state to its ground state or
- A higher energy excited state to a lower energy excited state.

(b) $\Delta v = \frac{0.0025mile}{hour} = \frac{0.0025\times1.6\times10^{3}m}{3600s} = 1.11\times10^{-3}m/s$

Then $\Delta p = m\Delta v = 3\times10^{3}\times1.11\times10^{-3}m/s = 3.33kgm/s$

And $\Delta x = 0.01mile = 0.01\times1.6\times10^{3}m = 16m$

It follows that; $\Delta x\Delta p = 16\times3.33 = 53.28$

But; $\frac{h}{4\pi} = \frac{6.626\times10^{-34}}{4\times3.14} = 5.275\times10^{-35}$

Since $\Delta x\Delta p$ is very large compared to $\frac{h}{4\pi}$, the uncertainty principle is not valid here.

Question 24

(a) Because the energies of the orbits in which electron revolve are fixed making the orbits stable.

(b)

n	3								
l	0	1			2				
m_l	0	-1	0	+1	-2	-1	0	+1	2
m_s	$\pm{}^1/_2$	$\pm{}^1/_2$	$\pm{}^1/_2$	$\pm{}^1/_2$	$\pm{}^1/_2$	$\pm{}^1/_2$	$\pm{}^1/_2$	$\pm{}^1/_2$	$\pm{}^1/_2$

Thus maximum number of electrons is 18

Question 25

(a) **They are similar in the following manner:**

- They both include electron as the sub-atomic particle.
- They both include proton as another sub-atomic particle
- Neither of their models included neutrons as the sub-atomic particles
- They both include the fact that an atom is electrically neutral

They are different in the following manner:

- Rutherford's model recognised the presence of nucleus **while** Thomson's model did not
- In Rutherford's model, the atom is mostly empty space and this is where the electrons are found **while** Thomson's model suggested that the atom was a solid sphere, with the electron embedded in among the positive protons.
- Rutherford's model suggested that protons are contained in very small core of an atom **while** Thomson's model suggested that protons occupy the whole space of the atom as positive fluid.

(b)

(i) $3s - \text{orbital}$

(ii) $2p_y - \text{orbital}$

(iii) $2p - \text{orbitals(which may be } 2p_x, 2p_y, \text{or } 2p_z \text{ orbital)}$

(iv) $3d_{z^2} - \text{orbital}$

Question 26

(a) From $\lambda = \frac{h}{mv}$

The De Broglie wavelength varies inversely proportional to the mass and hence electrons having smaller mass (than protons) have longer the De Broglie wavelength.

(b)

(i) $n = 3, l = 0 \; m_l = 0$

(ii) $n = 3, l = 1 \; m_l = +1$

(iii) $n = 2, l = 1 \; m_l = -1$

(iv) $n = 3, l = 2\ m_l = +1$

Question 27

(a) Light carries **quanta** (fixed amount) of energy and travels in packets called photons. Atoms can **absorb** or emit **photons** of specific energies.

(b)

(i) $[Ne]3s^2 3p^1$

(ii) $[He]\ 2s^2 2p^5$

(iii) $[Ar]4s^2 3d^7$

(iv) $[Ar]4s^2 3d^1$

Question 28

(a) In the daily life, macroscopic objects are more common. Those macroscopic objects have too small de Broglie wavelength to be detected and hence they do not exhibit wave characteristic.

(b)

(i) Chromium has atomic number of 24. Its electronic configuration is:

1s	2s	2p			3s	3p			4s	3d				
↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑	↑	↑	↑	↑

(ii) Copper has atomic number of 29. Its electronic configuration is;

1s	2s	2p			3s	3p			4s	3d				
↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑↓	↑↓	↑↓	↑↓	↑↓

Question 29

(a)

(i) Allowed

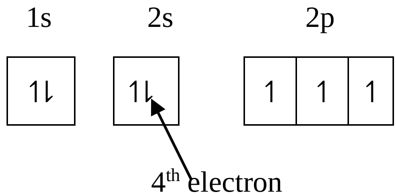
(ii) Not allowed because l must be less than n(l cannot be equal to n)

(iii) Allowed

(iv) Not allowed because magnitude of m_l cannot be greater than that of l(m_l must take values from $-l$ to $+l$)

(b)

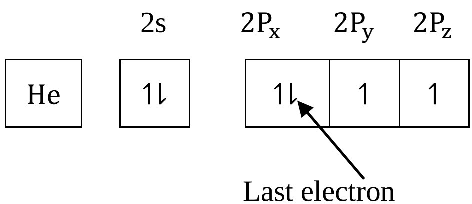
(i) Electronic configuration of nitrogen is



Thus quantum numbers for the fourth electron in the nitrogen are as follows:

$$n = 2, l = 0, m_l = 0, m_s = {}^{-1}/_{2}$$

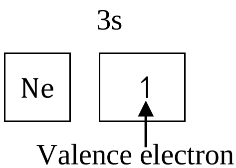
(ii) Electronic configuration of oxygen is



Thus quantum numbers for the last electron in the oxygen are as follows:

$$n = 2, l = 1, m_l = -1, m_s = {}^{-1}/_{2}$$

(iii) Electronic configuration of sodium is



Thus quantum numbers for the valence electron in the sodium are as follows:

$$n = 3, l = 0, m_l = 0\ , m_s = {}^{+1}/_{2}$$

Question 30

(a)

	ORBIT		ORBITAL
01	It is a defined circular (more specific, elliptical path followed by electron around the nucleus	01	Is the region around the nucleus of an atom where there is a maximum probability of finding an electron
02	It represents two dimension motion of electron around the nucleus	02	It represents three dimensional motion of electron around the nucleus
03	Orbit is elliptical in shape	03	Orbital have different shapes
04	The maximum number of electrons in an orbit is $2n^2$ where n is the orbit number	04	The maximum number of electrons in an orbital is 2

(b)

(i) Aluminium has atomic number of 13

$$\begin{array}{lcccc} & 3s & & 3p & \\ [Al] = [Ne] & \boxed{1⇂} & \boxed{1} & \boxed{} & \boxed{} \end{array}$$ ..At ground state

At excited state, one s-electron is promoted to higher energy p-orbital as follows:

$$\begin{array}{lcccc} & 3s & & 3p & \\ [Al] = [Ne] & \boxed{1} & \boxed{1} & \boxed{1} & \boxed{} \end{array}$$ …At excited state

(ii) Silicon has atomic number of 14

(iii)

	3s	3p			
[Si] = [Ne]	1⇂	1	1		….At ground state

At excited state, one s-electron is promoted to higher energy p-orbital as follows:

	3s	3p			
[Si] = [Ne]	1	1	1	1	...At excited state

(iv) Phosphorous has atomic number of 15

At ground state:

	3s	3p			3d				
[P] = [Ne]	1⇂	1	1	1					

Therefore, at excited state the configuration will be as follows:

	3s	3p			3d				
[P] = [Ne]	1	1	1	1	1				

Question 31

(a)

1. 1p because 1p means $n = 1, l = 1$ and l can never be equal to l.
2. 3f because 3f means $n = 3, l = 3$ and l can never be equal to l.
3. 2f because 2f means $n = 2, l = 3$ and I can never be greater than n.

(b) Sulphur has atomic number of 16

	3s	3p			3d				
[S] = [Ne]	1⇂	1⇂	1	1					

(i) To increase number of unpaired electrons to four so as to have covalency of four in S, one p-electron must be promoted to higher energy d-orbital at excited state as follows:

	3s	3p			3d				
[S] = [Ne]	1⇂	1	1	1	1				

(ii) To have six unpaired electrons (covalency of 6), one s-electron and one p-electron (from the paired orbital) must be promoted to higher energy d-orbital as follows:

	3s	3p			3d				
[S] = [Ne]	1	1	1	1	1	1			

Question 32

(a) $n = 4, l = 2, m_l = -2, -1, 0, +1$ or $+2$

(b) Atomic number of chlorine is 17

	3s	3p			3d				
[Cl] = [Ne]	1⇂	1⇂	1⇂	1					

(i) To increase number of unpaired electrons to three so as to have covalency of three in Cl, one p-electron must be promoted to higher energy d-orbital at excited state as follows:

	3s	3p			3d				
[Cl] = [Ne]	1⇂	1⇂	1	1	1				

(ii) To have five unpaired electrons (covalency of 5), two p-electrons (from paired p-orbitals) must be promoted to higher energy d-orbital as follows:

	3s	3p			3d				
[Cl] = [Ne]	1⇂	1	1	1	1	1			

(iii) To have seven unpaired electrons (covalency of 7), one s-electron and two p-electrons (from paired p-orbitals) must be promoted to higher energy d-orbital as follows:

[Cl] = [Ne]

3s	3p			3d				
↑	↑	↑	↑	↑	↑	↑		

Question 33

(a)

(i) This is accordance to Aufbau principle which states that: Electrons tend to occupy the orbitals of lower energy before those of higher energy.

(ii) This is accordance to Hund's rule which states that: Pairing of electrons in orbitals of the same type with the same energy is not allowed until each orbital is singly occupied with electron

(iii) This is accordance to Pauli's Exclusion principle which may be stated as: No two electrons in the same atom may have identical values for all four of their quantum numbers.

(b)

(i) Iron has atomic number of 26

So, [Fe] = [Ar]

3d					4s
↑↓	↑	↑	↑	↑	↑↓

Here outermost shell is n = 4, so two electrons which are to be removed so as to form Fe^{2+} must be from 4s-orbital.

Hence $[Fe^{2+}]$ = [Ar]

3d					4s
↑↓	↑	↑	↑	↑	

(ii) Zinc has atomic number of 30

So, [Zn] = [Ar]

3d					4s
↑↓	↑↓	↑↓	↑↓	↑↓	↑↓

Again the outermost shell is n = 4, so two electrons which are to be removed so as to form Zn^{2+} must be from 4s-orbital.

Hence $[Zn^{2+}] = [Ar]$

3d					4s
↿⇂	↿⇂	↿⇂	↿⇂	↿⇂	

(iii) From [Fe] in (i) above, $[Fe^{3+}]$ can be deduced as follows; (After removing 3 electrons from Fe).

Hence $[Fe^{3+}] = [Ar]$

3d					4s
↿	↿	↿	↿	↿	

Question 34

(a) With the configuration, $[Ar]3d^5 4s^1$, the configuration of chromium is said to violate Aufbau principle because electrons are filled in the higher energy 3d orbitals before totally occupying the lower energy $4s -$ orbital.

(b)

(i) Not valid: Violate Aufbau principle.
(ii) Not valid: Violate Paul Exclusion principle and Aufbau principle.
(iii) Not valid: Ra is not a noble gas and therefore it cannot be used in shortening configuration of other atoms.
(iv) Valid
(v) Not valid: Configuration of noble gas cannot used to shorten the configuration of the noble gas itself.

Question 35

(a)

(i) $1s < 2s = 2p < 3s = 3p = 3d < 4s$
(ii) $1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p$

(b) Using $A_r = \frac{m_1P_1 + m_2P_2 + m_3P_3}{100}$

Substituting $28.0855 = \frac{27.9769P_1 + 28.9765 \times 4.67 + 29.9738P_3}{100}$

From which: $27.9769P_1 + 29.9738P_3 = 2673.2297 \ldots\ldots\ldots\ldots\ldots$ (i)

But: $p_1 + p_2 + p_3 = 100$ Or

$P_1 + 4.67 + P_3 = 100$ or $P_1 + P_3 = 95.33 \ldots\ldots\ldots\ldots\ldots\ldots..\ldots\ldots$ (ii)

Solving (i) and (ii) simultaneously gives:

$P_1 = 92.23\%$ and $P_3 = 3.1\%$

Hence natural abundance of $Si-28$ and $Si-30$ are 92.23% and 3.1% respectively

Question 36

(a) Electronic configuration of Xe is

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6$

(i) Total number of electrons in $n = 4(4s^2 4p^6 4d^{10})$ is

$2 + 6 + 10 = 18$

Hence 18 electrons have $n = 4$ in the xenon

(ii) $n = 4, l = 2$ is for 4d orbitals which have 10 electrons $(4d^{10})$

Hence 10 electrons have$n = 4, l = 2.$

(iii) $ml = 0$ stand for s-orbital which are $1s^2\ 2s^2 3s^2 4s^2\ 5s^2$ for Xe with total number of $2+2+2+2+2$ or 10 electrons

Hence 10 electrons have $ml = 0$ in the xenon

(iv) Four quantum numbers specify only one electron in the Xe

Hence only 1 electron has $n = 2$ $l = 2,$ $ml = -1,$ $s = {}^1\!/_2$ in the xenon

(v) $n = 4, l = 3$ stands for 4f orbital which has no electron in xe

Hence 0 electrons has $n = 4,$ $l = 3$ in the xenon

(b) Using $A_r = \frac{M_1P_1 + M_3P_3}{100}$

Substituting $207.19g = \frac{205.98P_1 + 206.98P_2 + 207.98P_3}{100}$

From which: $205.98P_1 + 206.98P_2 + 207.98P_3 = 20719 \ldots\ldots\ldots$ (i)

Also: $P_1 + P_2 + P_3 = 100$

But:

$P_2 = P_3 = P(\text{Pb} - 207 \text{ and Pb} - 208 \text{ present in equal amount})$

Then equation (i) becomes:

$205.98P_1 + 206.98P + 207.98P = 20719$

Or $205.98P_1 + 414.96P = 20719 \ldots \ldots \ldots \ldots \ldots (\text{iii})$

And equation (ii) becomes:

$P_1 + P + P = 100$ or $P_1 + 2P = 100 \ldots \ldots \ldots \ldots . (\text{iv})$

Solving (i) and (ii) simultaneously gives:

$P_1 = 19.3\%, P = 40.3\%$

Hence percentage abundance of each isotope is as follows:

$\text{Pb} - 206{:}\ 19.3\%, \text{Pb} - 207{:}\ 40.3\%, \text{Pb} - 208{:}\ 40.3\%$

Question 37

(a) Electronic configuration of calcium is

$$1s^2\ 2s^2 2p_x^2 2p_y^2\ 2p_z^2\ 3s^2 3p_x^2 3p_y^2 3p_z^2 4s^2$$

(i) For calcium atom, $m_l = 0$ represents s and p_y orbitals in the following:

$$1s^2 2s^2 2p_y^2\ 3s^2 3p_y^2 4s^2$$

Thus the total number of electrons is $2 + 2 + 2 + 2 + 2 + 2 = 12$

Hence the number of electrons in calcium with quantum number $m_l = 0$ is 12.

(ii) $l = 1$ stands for the p orbitals which for calcium are:

$$2p_x^2 2p_y^2\ 2p_z^2\ 3p_x^2\ 3p_y^2\ 3p_z^2$$

Thus the total number of electrons is $2 + 2 + 2 + 2 + 2 + 2 = 12$

Hence the total number of electrons in calcium with quantum number $l = 1$ is 12.

(b) Using $n = \frac{m}{mr}$;

Number of moles of $\text{Ne}(\text{both Ne} - 20 \text{ and Ne} - 22)$

$= \dfrac{12.55}{20.18}\text{mol}$

$= 0.6219\text{mol}(\text{Notice the use of average atomic mass } 20.18)$

Using $N = nNA$

Total number of Ne atoms (Again both $Ne - 20$ and $Ne - 22$ atoms)

$= 0.6219 \times 6.02 \times 10^{23}$ atoms $= 3.743838 \times 10^{23}$ atoms

Now, let number of $Ne - 22$ be X

Then number of atom of $Ne - 20$ will be $3.743838 \times 10^{23} - X$

And by using,

$$Ar = \frac{\sum(\text{Number of atoms of isotope} \times \text{isotopic mass})}{\text{Total number of atoms}}$$

$$20.18 = \frac{21.99X + (3.743838 \times 10^{23} - X)19.99}{3.743838 \times 10^{23}}$$

Solving the above equation, gives $X = 3.5567 \times 10^{22}$

Hence there were 3.5567×10^{22} atoms of $Ne - 22$ in the given sample.

Question 38

(a)

(i)

- Horizontal lines represent quantized energy levels.
- Vertical lines represent transition of electrons between different quantized energy levels.

(ii)

- Through excitation of electron, quanta of radiation energy are absorbed and therefore giving absorption line spectra.
- When the **excited electron**s return **to ground state or to lower energy excited state** quanta of radiation energy is emitted and therefore giving emission spectrum.

(b)

Wavelength of a line in the hydrogen emission spectrum of is given by the following Rydberg equation:

$$\frac{1}{\lambda} = R_H\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

Substituting given values to the above equation gives

$$\frac{1}{\lambda} = 1.1 \times 10^7 \left(\frac{1}{2^2} - \frac{1}{5^2}\right),$$

From which $\lambda = 4.329 \times 10^{-7}$m or 432.9nm

From the wavelength corresponding to the given transition is approximately 433nm which is almost the same with the given value of 434nm. So as the transition of electron obey Rydberg equation for the hydrogen, the given wavelength is correct for hydrogen atom.

(c)

Assume an electron in hydrogen atom jump from its ground state of n= 1to higher energy of n_1 with respective energy of E and E_1.

From Rydberg equation:

$$\frac{1}{\lambda} = R_H \left(\frac{1}{1^2} - \frac{1}{n_1^2}\right)$$

Where λ is the wavelength of the energy absorbed,

R_H is Rydberg constant = $1.097 \times 10^7 m^{-1}$.

From Bohr's theory;

$$\Delta E = hf = \frac{hc}{\lambda} \qquad (f = \frac{c}{\lambda})$$

But $\Delta E = E_1 - E$

Thus $E_1 - E = \frac{hc}{\lambda}$

But from Rydberg equation

$$\frac{1}{\lambda} = R_H \left(\frac{1}{1^2} - \frac{1}{n_1^2}\right)$$

Then $E_1 - E = hc R_H \left(\frac{1}{n^2} - \frac{1}{n_1^2}\right)$

Where h is the plank's constant = 6.626×10^{-34}Js

c is the velocity of radiation = 2.998×10^8m/s

So $E_1 - E = 6.626 \times 10^{-34} \times 2.998 \times 10^8 \times 1.097 \times 10^7 \left(\frac{1}{1^2} - \frac{1}{n_1^2}\right)$

$$E_1 - E = 2.179163 \times 10^{-18} \left(\frac{1}{1^2} - \frac{1}{n_1^2}\right)$$

When n is too large, i.e. $n_1 = \infty$, $E_1 = 0$ and $\frac{1}{n_1^2} \approx 0$

Then $\Delta E = E_1 - E = 0 - E = -E$

Thus $-E = 2.179163 \times 10^{-18} \left(\frac{1}{1^2} - 0\right)$

From which; $-E = \frac{2.179163 \times 10^{-18}}{1}$ J

But 1ev = 1.6×10^{-19}J

Then $\frac{-2.179163 \times 10^{-18}}{1}\mathrm{J} = \frac{-2.179163 \times 10^{-18}}{1 \times 1.6 \times 10^{-19}}\mathrm{ev} = -13.6\mathrm{ev}$

Hence the energy is -13.6ev

Question 39

(a) $n = 2$(where Balmer series if formed).

(b) Dark lines are formed when electrons in $n = 2$ (lower energy excited state) hydrogen atoms are further excited to higher energy level by absorbing radiation energy of wavelength in the visible spectrum.

This is because when excited electron from higher energy level to $n = 2$ (Balmer series), it forms emission visible spectrum which inturn implies that only an electron in $n = 2$ will be excited by absorption of the same amount of energy in the visible spectrum, turning the red line into dark line.

(c) Ground state in hydrogen has $n = 1 = n_1$

Therefore second excited state has $n = 3 = n_2$

Using $f = \frac{c}{\lambda}$

But from Rydberg equation:

$$\frac{1}{\lambda} = R_H\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

It follows that; $f = cR_H\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$

Substituting $f = 3\mathrm{x}10^8 \times 1.1 \times 10^7\left(\frac{1}{1^2} - \frac{1}{3^2}\right)$ Hz

$= 2.93 \times 10^{15}$Hz

Hence the frequency is 2.93×10^{15}Hz

(d) Using $f = cR_H\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$

Where $n_1 = 1$(an electron dropped to the ground state after the first jump)

$n_2 =$ Is the higher energy level attained after the first jump which is unknown.

Substituting $3.2 \times 10^{15} = 3 \times 10^{8} \times 1.1 \times 10^{7} = \left(\frac{1}{1^2} - \frac{1}{n_2^2}\right)$

From which $n_2 = 5.7 \approx 6$

So for first jump: $n_1 = 4, n_2 = 6$(as the same energy absorbed in jumping from n=4 to n=6 will be emitted).

Then using again

$f = cR_H\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$

$= 3 \times 10^{8} \times 1.1 \times 10^{7}\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)\mathrm{Hz}$

$= 3 \times 10^{8} \times 1.1 \times 10^{7}\left(\frac{1}{4^2} - \frac{1}{6^2}\right)\mathrm{Hz}$

$= 1.146 \times 10^{14}\mathrm{Hz}$

Hence the frequency of photon required for first jump is $1.146 \times 10^{14}\mathrm{Hz}$

Question 40

(a) The lowest frequency in the Balmer series is obtained when an electron fall from $L = 3$. (Since the frequency is direct proportional to the energy, the lowest frequency is obtained when the energy difference between the two energy level is minimum which is $E_3 - E_2$ for Balmer series).

Then using:

$\frac{1}{\lambda} = R\left(\frac{1}{S^2} - \frac{1}{L^2}\right)$

$\frac{1}{\lambda} = 1.1 \times 10^{7}\left(\frac{1}{2^2} - \frac{1}{3^2}\right)$

From Which $\lambda = 6.55 \times 10^{-7}\mathrm{m}$

Hence the wavelength is $6.55 \times 10^{-7}\mathrm{m}$

(b) The light is formed when the excited electron in the hydrogen atom fall from L=3 to $S = 2$ and therefore emitting energy in terms of radiation (light).

(c) For first jump:

L=7: S is unknown

Substituting $\frac{1}{2.165\times10^{-6}} = 1.1 \times 10^{7}\left(\frac{1}{S^2} - \frac{1}{7^2}\right)$

From which: $S = 4$

For second jump;

S = 1(Ground state)

L = 4 (S of the first jump is the L or excited state of the second jump)

Then substituting that value in:

$$\frac{1}{\lambda} = R\left(\frac{1}{S^2} - \frac{1}{L^2}\right)$$

$$\frac{1}{\lambda} = 1.1 \times 10^7 \left(\frac{1}{1^2} - \frac{1}{4^2}\right)$$

From which $\lambda = 9.697 \times 10^{-8}$m

Hence the wavelength of the second photon is 9.697×10^{-8}m

Question 41

(a) Ultraviolet region

(b) Using Rydberg equation $\frac{1}{\lambda} = 1.1\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$

Substituting $\frac{1}{\lambda} = 1.1 \times 10^7 \left(\frac{1}{1^2} - \frac{1}{2^2}\right)$

From which $\lambda = 1.212 \times 10^{-7}$m

The wavelength of photon emitted is 1.212×10^{-7}m

(c) Electron transition from higher energy excited state to $n = 2$

(d) The amount of radiation energy absorbed in the excitation of the electron is equal to the amount of radiation energy emitted when the electron return to the ground state and is given by the following equation.

$$\Delta E = hf = \frac{hc}{\lambda}$$

But from Rydberg equation:

$$\frac{1}{\lambda} = R_H\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

Then $\Delta E = hcR_H\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$

$$= 6.63 \times 10^{-34} \times 3 \times 10^8 \times 1.1 \times 10^7 \left(\frac{1}{3^2} - \frac{1}{5^2}\right) \mathrm{J}$$

$= 1.556 \times 10^{-19}\text{J}$

Hence the energy of the photon is $= 1.556 \times 10^{-19}\text{J}$

Question 42

(a) Thomson used it to discover the electron and that the electron is negatively charged.

(b) Rutherford shot alpha particles through gold foil and discovered the atomic nucleus.
He learned that all atoms have a dense, positively charged tiny core and that electrons are found in space around the nucleus.

(c) Bohr discovered that electrons can only be certain, discrete distance from the nucleus.
The electrons orbit the nucleus in fixed predictable orbit whose energy is quantized rather than circling it randomly in an electron cloud.

Question 43

(a)

(i) $1s^2 2s^2 2p^6$

(ii) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

(iii) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

(b) Energy of one photon is given by the following Planck's equation:
$E = hf = \frac{hc}{\lambda}$

Substituting $E = \frac{6.63\times10^{-34}\text{Js}\times3\times10^{8}\text{ m/s}}{510\times10^{-9}\text{m}} = 3.9 \times 10^{-19}\text{J/photon}$

Using number of photons $= \frac{\text{Total energy}}{\text{Energy per photon}}$

$$= \frac{3.15\times10^{-17}\text{J}}{3.9\times10^{-19}\text{J/photons}} = 81 \text{ photons}$$

Hence the given energy corresponds to 81 photons

(Don't forget to approximate the answer to the nearest whole number because photons is discrete data, it can never be a fraction).

Question 44

(a) Dipositive ion (ion with a charge of +2) is formed after removing two electrons from an atom.

So the given electronic configuration with 10 electrons was obtained after removing 2 electrons from the atom.
Thus the actual total number of electrons in the neutral atom is $10 + 2 = 12$ electrons.
And for neutral atom;

$$\text{Number of electrons} = \text{Atomic number}$$

Hence atomic number of the atom is 12

(b)

(i) Two (2) electrons; the three set of quantum numbers specify specific type of an orbital and maximum number of electrons in the orbital is 2.

(ii) Fourteen (14) electrons (This is 5f orbitals and maximum number of electrons in f orbitals is 14).

(iii) Four (4) electrons (Maximum number of electrons in $n = 2$ 1s $2n^2 = 2 \times 2^2 = 8$, a half of them which is $^1/_2 \times 8 = 4$ spinning in anti-clockwise direction with $m_s = -^1/_2$).

Question 45

(a) The anion x^{3-} with electronic configuration of 10 electrons is formed after x gaining three electrons.

So number of electron in the neutral atom of x is $10 - 3$ or 7 electrons.
But for neutral atom; $\text{Atomic number} = \text{Number of electrons}$

Hence the atomic number of x is 7.

(b) Since radiation energy varies directly proportional to its frequency maximum energy corresponds to the maximum frequency which is$1\text{GHz} = 10^9\text{Hz}$.

Using $E = hf$: (Planck′s equation)

$$E = 6.63 \times 10^{-34}\text{Js} \times 10^9\text{Hz} = 6.63 \times 10^{-25}\text{J/photon}$$

Hence maximum energy, a photon can have $6.63 \times 10^{-25}\text{J}$

Converting the energy to $\mathrm{J/mol}$

Using total energy=Energy per photon × Number of photons

But for one mole of photons:

Number of photons=Avogadro's number=$6.02 \times 10^{23}\mathrm{photons/mol}$

Thus energy in $\frac{\mathrm{J}}{\mathrm{mol}} = 6.63 \times \frac{10^{-25}\mathrm{J}}{\mathrm{photons}} \times 6.02 \times 10^{23}\mathrm{photon/mol}$

$= 0.399\mathrm{J/mol}$

Hence the maximum energy per mole basis is approximately$0.4\mathrm{J/mol}$.Comparison

$0.4\mathrm{J/mol}$(Maximum radiation energy from all phones) is very small compared to the energy required to break the covalent bond and therefore cell phone radiation is highly unlikely to cause chemical changes in the body.

Question 46

(a) This is because there are three possibilities of forming diatomic molecule from the two isotopes which are:

- Two $\mathrm{Cl} - 35$ atoms combine to form $\mathrm{Cl_2}\left({}^{35}\mathrm{Cl} - {}^{35}\mathrm{Cl}\right)$
- One $\mathrm{Cl} - 35$ atom combine with another atom of $\mathrm{Cl} - 37$ to form $\mathrm{Cl_2}\ \left({}^{35}\mathrm{Cl} - {}^{37}\mathrm{Cl}\right)$
- Two $\mathrm{Cl} - 37$atoms combine to form$\mathrm{Cl_2}\left({}^{37}\mathrm{Cl} - {}^{37}\mathrm{Cl}\right)$.

(b)

- ${}^{35}\mathrm{Cl} - {}^{35}\mathrm{Cl}$ has $(35 + 35)\mathrm{amu} = 70\mathrm{amu}$
- ${}^{35}\mathrm{Cl} - {}^{37}\mathrm{Cl}$ has $(35 + 37)\mathrm{amu} = 72\mathrm{amu}$
- ${}^{37}\mathrm{Cl} - {}^{37}\mathrm{Cl}$ has $(37 + 37)\mathrm{amu} = 74\mathrm{amu}$

(c)

- Intensity of corresponding ${}^{35}\mathrm{Cl} - {}^{35}\mathrm{Cl}$ is proportional to probability of getting first ${}^{35}\mathrm{Cl}$ and second ${}^{35}\mathrm{Cl} = \frac{75}{100} \times \frac{75}{100}$
- Intensity corresponding to ${}^{35}\mathrm{Cl} - {}^{37}\mathrm{Cl}$ is proportional to the probability of getting first ${}^{35}\mathrm{Cl}$ and second ${}^{37}\mathrm{Cl}$ or first ${}^{37}\mathrm{Cl}$ and second ${}^{37}\mathrm{Cl} = \left(\frac{75\times25}{100}\right)$

- Intensity corresponding to $^{37}Cl - ^{37}Cl$ is proportional to the probability of getting first ^{37}Cl and second $^{37}Cl = \frac{25}{100} \times \frac{25}{100}$

Thus the relative amount of the three isotopes is:

$\frac{75\times25}{100} : 2\left(\frac{75\times25}{100}\right) : \frac{25}{100} \times \frac{25}{100}$

Or $75 \times 75 : 2 \times 75 \times 25 : 25 \times 25$ or $\frac{75\times75}{25\times25} : \frac{2\times75\times25}{25\times25} : \frac{25\times25}{25\times25}$

$= 9 : 6 : 1$

Hence the relative amount of three isotopes is $9:6:1$ for $70,72$ and 74amu respectively

Question 47

(i) If $n = 2$ and $m_l = 0$, l may be either 0 or 1 in the following two set of quantum numbers.

1: $n = 2, l = 0, m_l = 0 \Rightarrow 2s -$ orbital

2 : $n = 2, l = 1,\ m_l = 0 \Rightarrow 2P_y -$ orbital

Each set of the first three quantum numbers specify a particular orbital which may contain maximum of two electrons making a total of $4(2 \times 2)$electrons in the two sets.

Hence maximum number of electrons with given quantum number is 4.

(ii) If $n = 3$ and $L = 1$, ml may be $-1{,}0$ or $+ 1$ in the following three sets of quantum numbers:

1: $n = 3, l = 1,\ m_l = -1 \Rightarrow 3P_x -$ orbital

2: $n = 3, l = 1, m_l = 0 \Rightarrow 3P_y -$ orbital

3: $n = 3, l = 1, m_l = +1 \Rightarrow 3P_z -$ orbital

Each set of the first three quantum number specify a particular orbital with maximum of two electron in the orbital thus making a total of 2×3 or 6 electrons in the three sets.

Hence the maximum number of electrons with the given quantum numbers is 6.

(iii) If $n = 2$, and $l = 1$, ml may be either $-1{,}0$ or $+ 1$ in the following three sets of quantum numbers:

1: $n = 2, l = 1, m_l = -1 \Rightarrow 2P_x -$ orbital

2: $n = 2, l = 1, m_l = 0 \Rightarrow 2P_y -$ orbital

3: $n = 2, l = 1, m_l = +1 \Rightarrow 2P_z -$ orbital

Each set of the first three quantum numbers specify a particular orbital with maximum of two electrons in the orbital. This makes a total of 2×3 or electrons in the three sets.

Hence maximum number of electrons with the given quantum numbers is 6.

(iv) The given first three quantum numbers specify a particular orbital which is $2P_x -$ orbital which may contain a maximum of two electrons.

Hence maximum number of electrons with given quantum numbers is 2.

(v) The given four quantum numbers specify a particular electron.

Hence maximum number of electrons with given quantum number is 1.

Question 48

(a)

(i) m_s cannot be whole number, can be $+{}^1/_2$ or $-{}^1/_2$

(ii) m_l must lie between $-l$ and $+l$, its magnitude cannot be greater than that of l.

(b) From De-Broglie equation: $\lambda = \frac{h}{mc}$

But from: $K{:}E = {}^1/_2\, mc^2$ or $c = \sqrt{\frac{2\ K{:}E}{m}}$

Then substituting the above expression for c in the De-Broglie equation gives:

$$\lambda = \frac{h}{m\sqrt{\frac{2\ K{:}E}{m}}} = \frac{h}{\sqrt{2mK{:}E}}$$

For electron: $\lambda_e = \frac{h}{\sqrt{2m_eK{:}E}}$

For proton: $\lambda_p = \frac{h}{\sqrt{2m_pK{:}E}}$

Then $\frac{\lambda e}{\lambda_p} = \frac{h}{\sqrt{2meK{:}E}} \div \frac{h}{\sqrt{2m_pK{:}E}} = \sqrt{\frac{2m_pK{:}E}{2m_eK{:}E}}$

$$\frac{\lambda_e}{\lambda_p} = \sqrt{\frac{m_p}{m_e}} = \sqrt{\frac{1.67\times10^{-27}kg}{9.11\times10^{-31}kg}} = 42.8$$

Hence the ratio of De-Broglie wavelength of electron to that of proton is 42.8.

Question 49

(a) From De-Broglie equation; $\lambda = \frac{h}{mc} = \frac{h}{P}$

For electron: $\lambda_e = \frac{h}{P_e}$

For proton: $\lambda p = \frac{h}{P_p}$

But $\lambda_e = \lambda_p$ (The two particles have the same De-Broglie wavelength)

It follows that:

$$\frac{h}{P_e} = \frac{h}{p_p}$$

From which $P_e = P_p$

Hence both particles will have the **same momentum**

(b) From $P = mc$ or $c = \frac{P}{m}$.

Thus for a given momentum, speed varies inversely to the mass.

Since mass of electron is smaller than that of proton and the two have the same momentum (momentum is constant), speed of electron must be greater than that of proton.

(c) Using $K.E = \frac{1}{2}mc^2$

Then by multiplying m to both denominator and numerator of the above equation gives:

$$K.E = \frac{(mc^2)\times m}{2\times m} = \frac{m^2c^2}{2m} = \frac{(mc)^2}{2m} = \frac{P^2}{2m}$$

Then equation suggest that, for given momentum, kinetic energy varies inversely proportional to the mass.

Since mass of the electron is smaller than that of the proton and the two particles have the same momentum, kinetic energy of the electron must be greater than that of the proton

Hence the particles with greater kinetic energy is electron.

(d) From Planck's theory:

$E \propto f$ (Energy varies directly proportional to frequency) thus the particle with greater energy will have greater frequency too.

Hence the particle with greater frequency is the electrons.

Question 50

(a) (i) $2s-$ orbital (ii) $1s-$ orbital (iii) $3d-$ orbitals (iv) $4P_x-$ orbital

(b) From Heisenberg equation: $\Delta x \times \Delta p = \frac{h}{4\pi}$

Where $\Delta p = m\Delta c$:

And $\Delta c = \frac{0.2}{100} \times 1 \times \frac{10^6 m}{s} = 2000 m/s$

If follows that: $\Delta x = \frac{h}{(4\pi)p} = \frac{h}{(4\pi)m\Delta c}$

Substituting $\Delta x = \frac{6.63\times10^{-34}}{4\times3.14\times9.11\times10^{-31}\times2000} = 2.9 \times 10^{-8} m$

Hence maximum precision in the position measurement is $2.9 \times 10^{-8} m$

Question 51

(a) In Rutherford's $\alpha-$scattering experiment, very few $\alpha-$particles were either deflected or reflected and most of them passed through the gold foil without being deflected or reflected**.(0.5)** This led to the following conclusion:

(i) Since some $\alpha-$ particles were deflected**,** the nucleus is **positively** charged.

(ii) Since some $\alpha-$ particles were reflected, the nucleus is **heavy**.

(iii) Since most of $\alpha-$ particles passed through the gold foil with neither deflection nor reflection, nucleus occupies **small volume**.

(b) From Heisenberg equation: $\Delta x \Delta p = \frac{h}{4\pi}$

But $\Delta p = m\Delta c$

Then $\Delta x(m\Delta c) = \frac{h}{4\pi}$

From which: $\Delta c = \frac{h}{4\pi \times m \times \Delta x}$

Substituting $\Delta c = \frac{6.63\times10^{-34}}{4\times3.14\times1.67\times10^{-27}\times1.5\times10^{-11}} = 2107 m/s$

Thus the uncertainty in the position after 2 seconds is $2107 m/s \times 2s = 4214 m$

Question 52

(a) An excited electron returns to ground state or to lower energy excited state, emitting quanta of radiation energy as visible light of a specific wavelength.

(b) The energy difference between successive energy levels decreases with increasing energy which means the higher energy levels get closer and closer together.

(c)

(i) From Bohr theory: energy of quanta ('n' photons) is given by;

$$\Delta E = nhf$$

But for 1 mol, $n = 6.02 \times 10^{23}$ photons

Thus energy in J/mol

$$= 6.02 \times 10^{23} \times 6.63 \times 10^{-34} \times 1.26 \times 10^{15} J/mol$$

$$= 502899 J/mol \text{ Or } 502.899 kJ/mol$$

Hence the energy is $502.899 kJ/mol$

(ii) The energy represents first **ionisation energy** of the element.

Question 53

(a) $n = 3$ to $n = 2$

Explanation

Line A represents the transition of smallest energy gap because has longer wavelength than all the others shown. And as for Balmer series the lower energy level is $n = 2$, the smallest energy gap will be obtained when the higher energy level is $n = 3$.

(b) From Bohr theory: energy of quanta is given by:

$$\Delta E = nhf = \frac{nhc}{\lambda}$$

Where $n = 6.02 \times 10^{23}$ protons for one mole

Then substituting $\Delta E = \frac{6.02\times10^{23}\times6.63\times10^{-34}\times3\times10^{8}}{656\times10^{-9}} J/mol$

$$= 182527 J/mol \text{ Or } 182.527 kJ/mol$$

Hence the energy difference is $182.527 kJ/mol$.

Question 54

(a) An electron where the upward arrow represents an electron spinning in clockwise direction while the downward arrow represents an electron spinning in anti-clockwise direction.

(b) Each letter represents an orbital orientated along the *x*-, *y*- or *z*-axis.

(c) (i) s orbitals are spherical and symmetrical around the nucleus.

p orbitals are dumb-bell shaped and are symmetrical around each axis.

(ii) The p orbitals are arranged mutually at right angles.

(d) According to Hund's rule; electrons are placed singly in degenerate orbitals before pairing occurs in one orbital.

(e) Mistake 1: The '4p' orbitals should be labelled '3p'.

Mistake 2: Violate Hund's rule which states that electrons will occupy degenerate orbitals singly before any one is doubly filled.

Question 55

(a) (ii) is wrong because the 2s electrons have all four quantum numbers the same as they have parallel spin in the same orbital. This violates Pauli's exclusion principle.

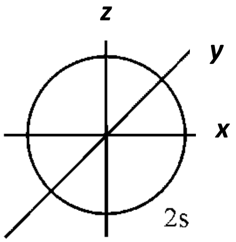
(b) (iv) is wrong because:

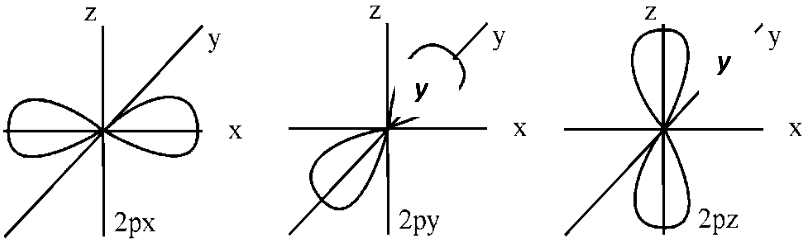
- Pairing in the $2p_x$ has been done before the two 2p electrons occupying two degenerate 2p orbitals (with parallel spin). This violate Hund's rule.
- Also the two 2p electrons have parallel spin in the same orbital. This violates Pauli's exclusion principle.

(c) (vi) is wrong because the 2p sublevel which has higher energy than 2s sublevel is filled with electrons before1s is completely filled with electrons. This violates the Aufbau principle.

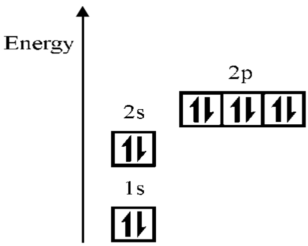
Question 56

(a)





(b) A region where one or (at most) two electrons are likely to be found.
(c) It signifies the second energy level.
(d) Of equal energy.
(e)



Question 57

(a)

ION	m/z
CH_4O^+	32
CH_3O^+	31
CH_2O^+	30
CHO^+	29
CO^+	28

The parent ion is CH_4O^+.
(b) The relative abundance of each ion formed.

(c)

ION	m/z where z=2
CH_4O^{2+}	16
CH_3O^{2+}	15.5
CH_2O^{2+}	15
CHO^{2+}	14.5
CO^{2+}	14

Question 58

(a)

(i) Only a small part of the diagram in the question represents the visible spectrum. Other parts represents invisible spectrum.

(ii) Balmer series.

(b) An electron is excited from its ground state to a higher energy level. On returning to the ground state, it emits energy in forms of radiation with amount equal to the energy difference of the two energy levels involved.

(c) From Bohr theory:

$$\Delta E = nhf = \frac{nhc}{\lambda}$$

Where $\frac{1}{\lambda}$ = Wave number

Substituting;

$$\Delta E = 6.02 \times 10^{23} \times 6.63 \times 10^{-34} \times 3 \times 10^{8} \times 11 \times 10^{6} \text{J/mol}$$

$$= 1317116\text{J/mol Or } 1317.116\text{kJ/mol}$$

Hence the energy is 1317.116kJ/mol

(d) The ionisation energy of hydrogen.

Question 59

(a) $1s^3 2s^3 2p^9$

(b) The electronic configuration of the element with a complete n=3 shell $1s^3 2s^3 2p^9 3s^3 3p^9 4s^3 3d^{15}$ (Notice that n=3 has 3s, 3p and

3d sublevels which must be completely filled with electrons for n=3 to be completely filled with electrons).

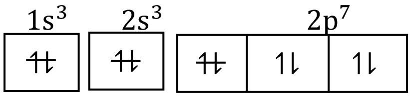
Thus total number of electrons in the atom $= 3 + 3 + 9 + 3 + 9 + 3 + 15 = 45$

But for neutral atom:

$$\textbf{Total number of electrons} = \textbf{Atomic number}$$

Hence the atomic number of the element is 45

(c) The electronic configuration of aluminium would be:



Hence there are **0** unpaired electrons.

Question 60

(a)

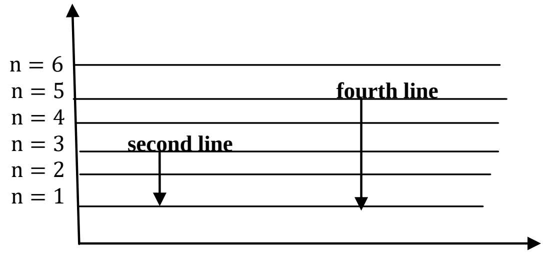
(i) In the ground state, electrons tend to occupy the orbital with minimum energy.

(ii) Electrons are not allowed to pair up unless the empty degenerate orbitals are singly occupied with parallel spins electrons.

(iii) It is impossible at any moment to predict simultaneously the exact position and velocity of an electron in an atom.

(b)

(i) Diagram to show the origin of second and fourth line in Balmer series.



(ii) It is given that: $\frac{1}{\lambda} = R_H\left(\frac{1}{2^2} - \frac{1}{n^2}\right)$

Frequency, $f = \frac{c}{\lambda}$

Then: $f = cR_H\left(\frac{1}{2^2} - \frac{1}{n^2}\right)$

For the first line in the balmier series, $n = 3$

$$f = 3 \times 10^8 \times 1.09678 \times 10^7\left(\frac{1}{2^2} - \frac{1}{3^2}\right) = 4.57 \times 10^{14}\ \text{Hz}$$

The frequency for the first line in the Balmer series is 4.57×10^{14} Hz

Question 61

(a) Due to stability for completely a half filled and full filled electronic structure in d-orbitals for Cr and Cu respectively which in turn is caused by the intersection of d and s sub-energy level. This make the electronic configuration of Cr to be $[Ar]3d^5 4s^1$ and not $[Ar]3d^4 4s^2$ while that of Cu is $[Ar]3d^{10} 4s^1$ and not $[Ar]3d^9 4s^2$.

(b) Number of orbitals presents in given principal quantum number, n, is n^2 And maximum number of electrons in an orbital is 2

Hence total number of electrons that can be held in all orbitals having the same principal quantum number is n^2.

Question 62

(a) Maximum frequency and minimum wavelength is obtained when maximum energy is emitted and (in this case) this occur when an electrons falls from $n = 3$ to $n = 1$

Energy remitted $= E_1 - E_3 = \Delta E$

$$= -20 \times 10^{-19} - (-15 \times 10^{-19})$$

$= -5 \times 10^{-19}$(Negative sign means energy was emitted)

But $\Delta E = hf$

From which $f = \frac{\Delta E}{h} = \frac{5\times10^{-19}}{6.63\times10^{-34}}\text{Hz} = 7.54 \times 10^{14}\text{Hz}$

Highest frequency of radiation is 7.54×10^{14}Hz

Using $\lambda = \frac{c}{f} = \frac{3\times10^{8}}{7.54\times10^{14}} = 3.98 \times 10^{-7}\text{m}$

Minimum wavelength of the radiation is 3.98×10^{-7}m

(b) An electrons is said to be complete ionised if it jumps from n=1 to infinity;(for and atom with only one electron, valency shell is n=1)

But $E_\infty = 0\text{J}$

Thus the energy absorbed in ionizing one electron

$$= 0 - E_1 = 20 \times 10^{-19}\text{J/electron}$$

Energy absorbed per one mole of electrons

$= 20 \times 10^{-19}\text{J/electron} \times 6.02 \times 10^{23}\text{J/mol}$

$$= 1204000\text{J/mol} = 1204\text{kJ/mol}$$

Hence the ionisation energy of the element is 1204kJ/mol

(c) Shortest the wavelength is obtained when the absorbed energy is maximum. If the electron is in $n = 4$ the maximum energy absorbed without causing ionisation is obtained when the electron absorb the energy so as to jump up to just before reaching the convergent limit where the energy is almost 0J.

Thus maximum energy absorbed

$$= 0 - 11 \times 10^{-19} = 11 \times 10^{-19}\text{J} = \Delta E$$

But $\Delta E = hf = \frac{hc}{\lambda}$ or$\lambda = \frac{hc}{\Delta E} = \frac{6.63\times10^{-34}\times3\times10^{8}}{11\times10^{-19}} = 1.8 \times 10^{-7}\text{m}$

Hence the shortest wavelength is 1.8×10^{-7}m.

Question 63

(i) The energy is negative to indicate that the electron (has negative charge) in the hydrogen atom is **attracted** to the nucleus (has positive) in electron-nucleus electrostatic force of attraction.

$\Delta E = E_2 - E_1$

$= \frac{-2.178x10^{-18}}{2^2} - \frac{-2.178x10^{-18}J}{1^2} = 1.6335 \times 10^{-18}J$

Hence $1.6335 \times 10^{-18}J$ must be absorbed

(ii) Using;

$$\Delta E = \frac{hc}{\lambda};\ \lambda = \frac{hc}{\Delta E} = \frac{6.63 \times 10^{-34} \times 3 \times 10^{8}m}{1.6335 \times 10^{-18}} = 1.22 \times 10^{-7}m$$

Hence the wavelength is $1.22 \times 10^{-7}m$

Question 64

(a) Electronic configuration of calcium is as follows: $1s^2 2s^2 2P_x^2 2P_y^2 2P_z^2 3s^2 3P_x^2 3P_y^2 3P_z^2 4s^2$

(i) Total number of electrons in n = 3 of $3s^2 3P_x^2 3P_y^2 3P_z^2$ is $2 + 2 + 2 + 2$ or 8 electrons.

Hence 8 electrons have quantum number of $n = 3$

(ii) $m_l = 0$ represents s and P_y orbitals in $1s^2 2s^2 2P_y^2 3s^2 3P_y^2 4s^2$ with total of $2 + 2 + 2 + 2 + 2 + 2$ or 12 electrons.

Hence 12 electrons have quantum number of $m_l = 0$

(iii) $l = 1$ represents P orbitals in $2P_x^2 2P_y^2 2P_z^2 3P_x^2 3P_y^2 3P_z^2$ with total of $2 + 2 + 2 + 2 + 2 + 2$ or 12 electrons.

Hence 12 electrons have quantum number of $l = 1$

(b) Thickness of soap bubble = wavelength = $6nm = 6 \times 10^{-9}m$

Using $f = \frac{c}{\lambda} = \frac{3\times10^8}{6\times10^{-9}} = 5 \times 10^{16}Hz$

Question 65

(a)

(i) $[Ar]3d^{10}4s^1$

(ii) $1s^2 2s^2 2p^6$

(iii) $[Ar]3d^5 4s^1$

(iv) $1s^2$

(b) **Postulates of Bohr's atomic model**

1. An atom consists of very tiny positively charged nucleus. Nearly the entire mass of the atom is concentrated in the nucleus.
2. The electrons in an atom revolve around the nucleus in a certain permitted circular orbits. The electrostatic force of attraction between the electron and the nucleus provides the necessary centripetal force.
3. In an atom there are energy levels or energy states called stationary states in which electrons do not emit (radiate) energy, the energy of the orbits is quantized, that is they have fixed value of energy to enable electrons keeps on moving in the same orbit.
4. The angular momentum of the electrons in stationary states (ground states) quantized and it is an integral multiple of $\frac{h}{2\pi}$
5. The electron can jump from one orbit to the other when it does so it emits or absorb energy. Radiation energy is emitted when electrons move from higher to lower energy levels while it is absorbed when electrons jump from lower energy level and in both cases the radiation emitted or absorbed is equal to the energy difference between the two energy levels and is quantized.

Shortcomings of Bohr's atomic model

1. It does not explain the spectrum of multi-electron atoms.
2. It does not explain the relative intensity of spectral lines (that is why some spectral lines are brighter while others are dimmer).
3. The model does not explain how covalent bonding makes a molecule stable because the model does not recognise the existence of sub-energy levels.
4. The model viewed an electron as being placed at certain distance from the nucleus but it was proved by Werner Heisenberg in his Heisenberg's uncertainty principle, which states that,
5. No clear justification was given for the quantization of angular momentum of electron and presence of stationary states.

6. Bohr's model suggests that the electron move in circular orbit in two dimensions while it has been proved that (in quantum mechanical atomic model) the motion is in three dimensions on the elliptical path.
7. It does not explain presence of hyperfine spectral lines.
8. It does not explain the Zeeman and stark effect.

Question 66

(a)

(i) Is the phenomenon which occurs when electron revolve around the nucleus whereby angular momentum can only take discrete values.

(ii) Is the energy sublevel in which electron is placed and specifies the shape of an orbital of an electron.

(iii) Is the property of material objects to exhibit both wave and particle characteristics.

(b) When the electron is in $n = \infty$ where $E = 0$ implies that it experiences zero nuclear attraction and therefore the electron is completely ionised from the atom. So $n = \infty$ and $E = 0$ is very useful in the determining the ionisation energy of hydrogen.

Question 67

(a) $\text{Number of protons} = \text{the number in the subscript} = 7$

$\text{Number of electrons} = \text{Number of protons} = 7$

$\text{Number of neutrons}$

$= \text{Mass number (in the superscript)} - \text{Number of protons}$

$$= 15 - 7 = 8$$

(b) $E = hf = \frac{hc}{\lambda} = \frac{6.63 \times 10^{-34} \times 3 \times 10^{8}}{2420 \times 10^{-10}} = 8.219 \times 10^{-19}\text{J/electron}$

For one mole of electron

$E = 8.219 \times 10^{-19}\text{J/electron} \times 6.02 \times 10^{23}\text{electron/mole}$

$= 494783.8\text{J} = 494.7838\text{kJ/mol}$

Question 68

(a)

1. **Probable distance of an electron from the nucleus**
   Greater value of principal quantum number means greater distance from the nucleus.
2. **Energy of an electron**
   Greater value of principal quantum number means greater energy of an electron with that principal quantum number.
3. **Stability of an electron**
   Greater value of principal quantum number means less stability of an electron with that principal quantum number.

4. **Size of an orbital**
   Greater value of principal quantum number means greater size of an orbital with that principal quantum number.

(b) For Balmier series, the ground state is n=2; the longest wavelength is obtained when there is minimum energy emitted and this occur when an electron falls from n=3 to n=2(for Balmer series)

Using $\frac{1}{\lambda} = R_H(\frac{1}{n_1^2} - \frac{1}{n_2^2})$

But $\frac{1}{\lambda} = \text{wave number}$

Wave number= $R_H\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$

Where $R_H = 1.1 \times 10^7 m^{-1}$; $n_1 = 2$ and $n_2 = 3$

Then wave number= $1.1 \times 10^7\left(\frac{1}{2^2} - \frac{1}{3^2}\right) = 1527778m^{-1}$

Thus the wave number of longest wavelength is $1527778m^{-1}$

Question 69

(i) $\Delta E = hf$; but $f = \frac{c}{\lambda}$

It follows that: $\Delta E = \frac{hc}{\lambda}$

From which $\lambda = \frac{hc}{\Delta E} = \frac{6.63\times10^{-34}\times3\times10^{8}}{4.071\times10^{-19}} = 4.886 \times 10^{-7}\text{m}$

Wavelength of green light is $4.886 \times 10^{-7}\text{m}$

(ii) $\Delta E = E_1 - E_2$

Where;

$\Delta E = -4.071 \times 10^{-19}$J(negative sign is included because the energy is released)

$E_1 = \frac{-13.6\text{eV}}{2^2}$ (Using $E = \frac{-13.6\text{eV}}{n^2}$ (And the light was in Balmer series for which n=2)

$1\text{eV} = 1.6 \times 10^{-19}\text{J}$

$$E_1 = \frac{-13.6 \times 1.6 \times 10^{-19}}{2^2}\text{J} = -5.44 \times 10^{-19} = \Delta E + E_2$$
$$= -4.071 \times 10^{-19}\text{J} + E_2$$

From which $E_2 = -1.369 \times 10^{-19}\text{J}$

Thus the higher energy level is $-1.369 \times 10^{-19}\text{J}$

Question 70

(a) λ represents the wavelength in metres

R_H represents Rydberg constant

n_1 represents lower energy level which is n_1=2 for visible region

n_2 represent the higher energy level, that is $n_2 > 2$

(b) For visible spectrum , n_1=2

Thus the third line in visible spectrum is $n_2 = 5$

Thus the given equation becomes:

$$\frac{1}{\lambda} = R_H\left(\frac{1}{2^2} - \frac{1}{5^2}\right)$$

But $f = \frac{c}{\lambda}$

It follows that:

$$f = cR_H\left(\frac{1}{2^2} - \frac{1}{5^2}\right) = 3 \times 10^8 \times 1.1 \times 10^7 \times \left(\frac{1}{4} - \frac{1}{25}\right)$$

$= 6.93 \times 10^{14}$Hz.

The frequency of the third line of the visible spectrum is 6.93×10^{14}Hz

Question 71

(a) Due to presence of isotopes, relative atomic mass is found from the weighted average of mass of different isotopes of silicon. Abundances (relative amounts) of these isotopes is different depending on the sources (region where silicon are is mined) and hence the mass will not be exactly the same.

(b) $\Delta E = hf = \frac{hc}{\lambda} = \frac{6.63\times10^{-34}\times3\times10^{8}}{1875\times10^{-9}}J = 1.0608 \times 10^{-19}J$

(Negative sign will be included in calculations because the ΔE is the energy emitted)

Using $E = \frac{-13.6eV}{n^2}$

For n =3

$E = \frac{-13.6\times1.6\times10^{-19}J}{3^2} = -2.42 \times 10^{-19}J$

Let the higher energy level be E_n

Then $\Delta E = E_3 - E_n = -2.42 \times 10^{-19}J - E_n = -1.0608 \times 10^{-19}J$

From which; $E_n = -1.3592 \times 10^{-19}J$

But $E_n = \frac{-13.6eV}{n^2}$

Then $= -1.3592 \times 10^{-19} = \frac{-13.6\times1.6\times10^{-19}J}{n^2}$

From which n=4 (To nearest integer)

Thus for the higher energy level, n=4.

Question 72

(a)

(i) 3
(ii) 0
(iii) 5
(iv) 1

(b) From Bohr Theory;

$\Delta E = nhf = \frac{nhc}{\lambda}$ where $n = 6.02 \times 10^{23}$ photons (or electrons) if the energy is in J/mol.

Substituting $145.684 \times 10^{3}\text{J/mol} = \frac{6.02\times10^{23}\times6.63\times10^{-34}\times3\times10^{8}}{\lambda}$

From which; $\lambda = 8.219 \times 10^{-7}\text{m}$

But $1\text{A}° = 10^{-10}\text{m}$

So $8.219 \times 10^{-7}\text{m} = \frac{8.219\times10^{-7}\text{m}}{10^{-10}}\ \text{A}° = 8219\text{A}°$

The wavelength of radiation is $8219\text{A}°$

Question 73

(a) From the Heisenberg's equation;

$$\Delta x \times m\Delta v \approx \frac{h}{4\pi}$$

Rearranging the above equation gives;

$$\Delta x \approx \frac{h}{4\pi \times m\Delta v}$$

From which it can be deduced that the uncertainty in determining position (Δx) of an object of mass, m, varies inversely proportional to the uncertainty in the determination of velocity (Δv) of the object.

That means if the velocity of the object is determined with more accuracy then its position will determined with less accuracy as suggested in Heisenberg uncertainty principle.

(b) De-Broglie wavelength is given by;

$\lambda = \frac{h}{mv}$

But $\text{K.E} = \frac{1}{2}mv^{2}$

From which $v = \sqrt{\frac{2\text{K.E}}{m}}$

Thus $\lambda = \frac{h}{m\sqrt{\frac{2K.E}{m}}} = \frac{h}{\sqrt{2mK.E}}$

Where $m = 6.6 \times 10^{-27}kg$

$K.E = 4.8MeV = 4.8 \times 10^{6} \times 1.6 \times 10^{-19}J$

$h = 6.63 \times 10^{-34}Js$

$\lambda = \frac{6.63\times10^{-34}Js}{\sqrt{2\times6.6\times10^{-27}\times4.8\times10^{6}\times1.6\times10^{-19}}} = 6.58 \times 10^{-15}m$

The De-Broglie wavelength is $6.58 \times 10^{-15}m$

Question 74

(a) With given isotopes, there are following possibilities of forming triatomic molecule:

(i) Three atoms of ^{m}X

(ii) Three atoms of ^{n}X

(iii) One atom of ^{m}X and two atoms of ^{n}X

(iv) One atom of ^{n}X and two atoms of ^{m}X

Hence four peaks will be shown in the mass spectrum.

(b) Visible hydrogen spectrum is formed when an electron fall from higher energy level to $n = 2$.

Thus from Rydberg equation;

$\frac{1}{\lambda} = R_H\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$;

n_1=2 for visible spectrum.

The lowest energy will be emitted when the electron falls from $n = 3$, that is $n_2 = 3$

From Bohr Theory, energy of one quantum is given by;

$$\Delta E = hf = \frac{hc}{\lambda}$$

It follows that:

$$\Delta E = hf = hcR_H\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right);$$

Substituting $\Delta E = 6.63 \times 10^{-34} \times 3 \times 10^{8} \times 1.1 \times 10^{7}\left(\frac{1}{2^2} - \frac{1}{3^2}\right)$ J/electron $= 3.0388 \times 10^{-19}$J/electron

Using $n = \frac{m}{M_r}$;

$$\text{Number of moles of H atoms} = \frac{1g}{1g/mol} = 1mol$$
$$= 6.02 \times 10^{23}\text{atoms}$$

But each atom of hydrogen contains one electron;

Thus total number of electrons in 1g of H is 6.02×10^{23}electrons

Hence the total energy emitted is $3.0388 \times 10^{-19} \times 6.02 \times 10^{23}$ J
$= 182935.76$J or 182.93576kJ

Question 75

(a) Using; $Z = A - \text{Number of neutrons}$
Atomic number of x $= 35 - 18 = 17 =$ Atomic number

For neutral atom; atomic number = number of electrons

Thus **number of electrons in Y is 17**

Using Number of neutrons = A – Z

Number of neutrons in Y = 37 – 17 = 20

Thus **number of neutrons in Y is 20**

(b)

(i) Not permissible because l cannot be equal to n.

(ii) Not permissible because magnitude of m_l cannot be greater than that of l; m_l must range between –l and +l.

(iii) Permissible

(iv) Not permissible because m_s can only be either $+{}^{1}/_{2}$ or $-{}^{1}/_{2}$. It cannot take whole number.

Question 76

The electrically heated metal coil produces enough amount of heat to initiate the ionisation as per equation:

$$\mathrm{M(g) + Heat \rightarrow M^{+}(g) + e}$$

The knocked off electrons collide with un-ionised atoms to cause more ionisation as per equation:

$$\mathrm{M(g) + e \rightarrow M^{+}(g) + 2e}$$

Alternatively it can be done by electron bombardment through employment of an electron gun.

$$\mathrm{M(g) + \underset{\uparrow}{e} \longrightarrow M^{+}(g) + 2e}$$

An electron from the electron gun
to knock off electron from $\mathrm{M(g)}$

Reasons for the ionisation:

Ionisation of the sample is necessary because it enable particles (gaseous ions) to be:

(i) **Accelerated** by electric field in the acceleration chamber
(ii) **Deflected** by magnetic field in the deflection chamber
(iii) **Detected** in the detector

Question 77

(i) Carbon has an atomic number of 6

Thus in $^{12}\mathrm{C}^{+}$: $\mathrm{A = 12, z = 6 = Number\ of\ protons}$

And number of neutrons $= \mathrm{A - z = 12 - 6 = 6}$

Using mass of an atom = mass of protons + mass of neutrons

And if we let:

$$\text{Mass of one proton } = \mathrm{p}$$
$$\text{Mass of one neutrons} = \mathrm{n}$$

It follows that:

$$6\mathrm{p} + 6\mathrm{n} = 1.993 \times 10^{-23} \ldots\ldots\ldots\ldots\ldots \text{(i)}$$

And for $^{13}\mathrm{C}^{+}$: $\mathrm{A = 13, z = 6}$

Then number of neutrons $= A - z = 13 - 6 = 7$
And therefore $6p + 7n = 2.158 \times 10^{-23}$ … … … … … (iii)
Taking (iii) − (i) gives:
$$n = 1.65 \times 10^{-24}\text{g}$$
Hence the mass of one neutron is 1.65×10^{-24}g

Alternative solution for (i)

^{12}C and ^{13}C are isotopes with the same number of protons but different number of neutrons.
Thus ^{13}C has one more neutron compared to ^{12}C.
So the difference in mass between $^{13}C^{+}$ and $^{12}C^{+}$ is mass of one extra neutron present in the $^{12}C^{+}$.
Hence mass of one neutron $= (2.158 \times 10^{-23} - 1.993 \times 10^{-23})\text{g} = 1.65 \times 10^{-24}\text{g}$

(ii) Numerically, the relative atomic mass of ^{13}C is equal to the mass of one mole of ^{13}C atoms.

Thus relative atomic mass of ^{13}C = Mass of one atom of $^{13}C \times$ Avogadro's constant $= 2.158 \times 10^{-23} \times 6.02 \times 10^{23} = 12.9912$
Hence the relative atomic mass of ^{13}C is 12.9912

(iii) Using $A_r = \frac{m_1P_1 + m_2P_2}{P_1 + P_2}$

Where $m_2 = 12.9912, P_2 = 1.1\%, P_1 = 98.9\%$
And $m_1 = 1.993 \times 10^{-23} \times 6.02 \times 10^{-23} = 11.9979$
Then $A_r = \frac{(11.9979 \times 98.9) + (12.9912 \times 1.1)}{(98.9 + 1.1)} = 12.0088\text{amu}$
Hence the relative atomic mass of carbon in sample is 12.0088amu

Question 78

(i) Electric field from charged plates.
(ii) Electromagnet (or simply magnet).
(iii) Presence of different isotopes of K_r with different $\frac{m}{z}$ ratio.
(iv) Reducing strength of the magnetic field.
(v) 1: **Mass to charge ratio** ($^{m}/_{z}$ ratio) of isotope (or simply isotopic mass).
2: **Abundance** of each isotope

(vi) **Speed** of the ionised particle.

Question 79

(a)

Particle	**Proton**	**Neutron**	**Electrons**
Mass /g	1.6725×10^{-24}	1.6748×10^{-24}	0.0009×10^{-24}
Relative charge	$+1$	0	-1

(b) Mass of the H atom $=$ mass of a proton $+$ mass of an electron

$$= (1.6725 \times 10^{-24} + 0.0009 \times 10^{-24})\text{g}$$
$$= 1.6734 \times 10^{-24}\text{g}$$

Hence the mass of an atom of hydrogen is 1.6734×10^{-24}g

(c) Mass of one mole of H atoms $=$ mass of one atom $\times$ Avogadro's constant

$$= 1.6734 \times 10^{-24} \times 6.0225 \times 10^{23} = 1.0078\text{g}$$

Hence mass of one mole of H atoms is 1.0078g

(d) Due to presence of other heavier **isotopes** of hydrogen.

Question 80

(i) **A** is electron
B is neutron
C is proton

(ii) **Explanation for the directions:**

- An electron (**A**) being **negatively charged** is attracted (deflected) towards positive plate.
- A proton (**C**) being **positively charged** is attracted (deflected) towards negative plate.
- A neutron (**B**) having **no charge** is undeflected.

Explanation for the shape:

The shapes of path traced by A and C suggest that A is more deflected. This is because **A** (electron) has less mass than **C** (proton).

Question 81

(a) Metal coil (or electron gun).
(b) Only charged particles (ions) can be attracted or accelerated by electric (or magnetic) field.
(c) Magnetic field (or simply magnet).
(d) Magnetic field

Question 82

(i) ^{63}Cu is **more abundant** than ^{65}Cu and therefore A_r become close to 63.

(ii) There are two ways of converting Cu atoms to Cu^+ ions:

First way: Heat energy from strongly heated metal coil as per equation;

$$Cu(g) + heat \longrightarrow Cu^+(g) + e$$

Second way: Strong energetic collision between high speed electron (which has high kinetic energy) and un-ionised Cu atoms leading to the loss of electron in the Cu atoms as per equation:

$$Cu(g) + e \longrightarrow Cu^+(g) + 2e$$

(iii) **Dipositive ion** of ^{63}Cu which is $^{63}Cu^{2+}$

(iv) $\frac{m}{z}$ for $^{63}Cu^{2+}$: $m = 63, z = 2$; Therefore $\frac{m}{z} = \frac{63}{2} = 31.5$

Reason:

More energy is needed to remove second electron and therefore making the process of removing two electrons in Cu atoms more difficulty than removing one electron only.

Question 83

(a) Since wavelength varies inversely proportional to energy, most energetic photon has shortest wavelength which is given as 220nm (or 220×10^{-9}m)

Then using $f = \frac{c}{\lambda} = \frac{3\times10^{8}m/s}{220\times10^{-9}m} = 1.3636 \times 10^{15}$Hz

Hence the frequency of the most energetic photon is 1.3636×10^{15}Hz

Using energy,

$E = hf = 6.63 \times 10^{-34} \times 1.3636 \times 10^{15} J = 9.0407 \times 10^{-19} J$

Hence the energy of most energetic photon is 9.0407×10^{-19}J

(b) Using $E = nhf = \frac{nhc}{\lambda}$

Where;

E= $347 kJ/mol = 347 \times 10^{3} J/mol$, $c = 3 \times 10^{8} m/s$

$n = 6.02 \times 10^{23}$ photon/mol, $h = 6.63 \times 10^{-34} Js$

Substituting $347 \times 10^{3} = \frac{6.02\times10^{23}\times6.63\times10^{-34}\times3\times10^{8}}{\lambda}$

From which $\lambda = 3.45 \times 10^{-7} m = 345 nm$

Hence the wavelength of the least energetic photon that can break the $C - C$ bond is 345nm

Comparison:

The photon with 345nm is not blocked by ozone, (it is not energetic enough to be absorbed by ozone which absorb a light with wavelength in ranged from 220 to 290nm) so $C - C$ bond disruption is still possible even with presence of the ozone layer.

Question 84

(a) Due to existence for of isotopes, the value recorded in the periodic table is just an average of isotopic masses.

(b) Using $n = \frac{m}{mr}$:

Number of moles of iodine with atomic mass of 126.9045

$$= \frac{12.3849 g}{126.9045\ g/mol} = 0.09759\ mol$$

Number of moles of iodine with atomic mass of 128.9050 (I − 129)

$$= \frac{1.00007 g}{128.9050\ g/mol} = 0.007758\ mol$$

Then the average atomic mass can be found by using the following formula:

$$A_r = \frac{\sum(\text{Number of moles of isotope } \times \text{isotopic mass})}{\text{Total number of moles}}$$

Substituting $A_r = \frac{(0.09759 \times 126.9045)+(0.007758 \times 128.9050)}{(0.09759+0.007758)}$

$$= 127.0518 \text{ amu}$$

Hence the apparent atomic mass of the contaminated iodine is 127.0518 amu

Question 85

(a) It is impossible at any moment to predict simultaneously the exact position and momentum of an electron in an atom.

(b) **Meaning:** Is the property of material objects to exhibit both wave and particle characteristics.

Derivation:

Consider a photon of light; in wave model its energy is given by the following Planck's equation:

$$E = hf$$

Where h is Planck's constant

f is the frequency

In the particle model the energy is given by the following Einstein's equation:

$$E = mc^2$$

Where m is the mass of the photon

c is the velocity.

Equating Plank's and Einstein's equation (Law of conservation of energy)

$$hf = mc^2$$

But $f = \frac{c}{\lambda}$

Then $\frac{hc}{\lambda} = mc^2$

$\frac{h}{\lambda} = mc$ or $\lambda = \frac{h}{mc}$

Hence $\lambda = \frac{h}{mc}$

(c) Bohr considered an electron being orbiting the nucleus in the fixed path with fixed distance from the nucleus while according to the Heisenberg's uncertainty principle, "It is impossible at any moment to predict simultaneously the exact position and velocity of an electron in an atom."

(d) From the De-Broglie equation;

$$\lambda = \frac{h}{mc}$$

Where De-Broglie wavelength varies inversely proportional to the mass of particle; so for macroscopic object whose mass is large, the wavelength becomes too small to be measured unlike in electrons whose mass is very small.

Question 86

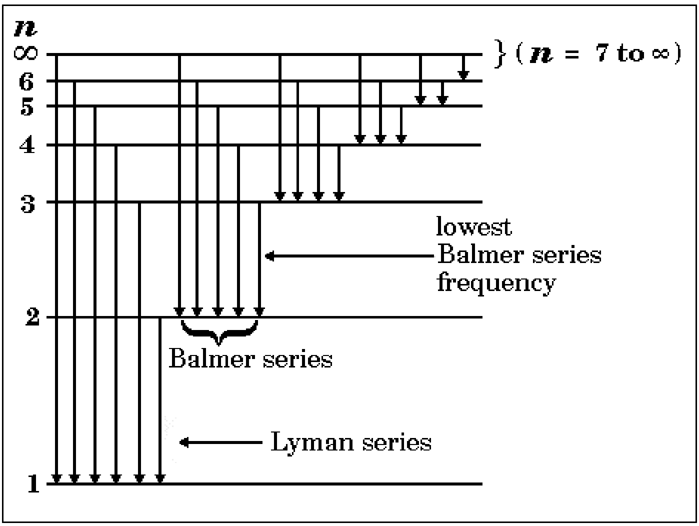
(a) Paschen series, Bracket series and P-fund series.

(b) Electronic energies are quantized.

(c)

- Excited electron may fall to the ground state either directly or in steps.
- Different energy of energy levels.

(d)



(e)

- Refer to the text for the difference between emission spectrum and absorption spectrum.
- The room temperature is not sufficient to excite an electron in the hydrogen atom and therefore the electron is at ground state where $n = 1$.

Question 87

(a) $n_1 = 1, n_2 = \infty$

Substituting $f = 3 \times 10^8 \times 1.1 \times 10^7 \left(\frac{1}{1^2} - \frac{1}{\infty^2}\right)$

$$= 3.3 \times 10^{15}(1 - 0) = 3.3 \times 10^{15}\text{Hz}$$

Hence the frequency of the line produced is $3.3 \times 10^{15}\text{Hz}$

(b) From Bohr theory

$$\Delta E = hf$$

But $f = cR_H\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$ (Given)

Hence $\Delta E = hcR_H\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$

(c) Substituting

$$\Delta E = 6.63 \times 10^{-34} \times 3 \times 10^8 \times 1.1 \times 10^7\left(\frac{1}{1^2} - \frac{1}{\infty^2}\right)$$

$$= 2.1879 \times 10^{-18}\text{J}$$

But $1\text{ev} = 1.6 \times 10^{-19}\text{J}$

Then $2.1879 \times 10^{-18}\text{J} = \frac{2.1879\times10^{-18}}{1.6\times10^{-19}}\ \text{ev} = 13.67\text{ev}$

Hence the energy needed is 13.67ev

Assumption made:

The energy absorbed in moving the electron from 1-level to infinity is the same as the energy emitted in moving the electron from the infinity to the 1-level.

(d) From (c), the energy required to remove the electron from 1-level to infinity is $2.1879 \times 10^{-18}\text{J}$.This corresponds to the ionisation if it is converted to J/mol.

But one mole $=6.02 \times 10^{23}$ photons

So the energy is $\text{J/mol} = 2.1879 \times 10^{-18} \times 6.02 \times 10^{23}\text{J/mol}$

$$= 1317116\text{J/mol} \text{ or } 1317.116\text{kJmol}$$

Hence the ionisation energy of hydrogen is $1317.116 \mathrm{kJ/mol}$

Question 88

(a) Two electrons (magnetic quantum numbers specify a particular orbital and maximum number of electrons in an orbital is always two).

(b)

(i) Electronic configuration of argon is:

$$1s^2\ 2s^2\ 2p_x^2\ 2p_y^2\ 2p_z^2\ 3s^2\ 3p_y^2\ 3p_z^2$$

In this case, $m = +1$ stands for P_z-orbital in $2P_z^2$ and $3P_z^2$ with total of 2+2=4 electrons.

Hence there are **4 electrons** with m=+1 in the argon

(ii) For d orbitals l=2

Thus total number of d-orbitals is $2l + 1 = (2 \times 2) + 1 = 5$

But each orbital may contains maximum of two electrons:

Thus total number of electron in d-orbitals $= (2 \times 5)$ or 10 electrons

(iii)

$n = 4$ and $m_l = +2$ may occur when $l = 3$ or 2 in following two sets of quantum numbers:

1. $n = 4, l = 3, m_l = +2 \Rightarrow$ for first orbital
2. $n = 4, l = 2, m_l = +2 \Rightarrow$ for second orbital

With maximum of two electrons in each orbital, maximum number of electrons with the given quantum numbers in the single atom is 4.

(iv) The given quantum numbers represent the last electron in $3p_x^2$.

The element which end with $3p_x^2$ in its electronic configuration have the following full electronic configuration:

$1s^2 2s^2 2p_x^2\ 2p_y^2\ 2p_z^2\ 3s^2 3p_x^2\ 3p_y^1\ 3p_z^1$

The configuration gives a total of $2 + 2 + 2 + 2 + +2 + 2 + 2 + 1 + 1$ or 16 electrons.

With a total of 16 electrons, the element has atomic number of 16

The element with atomic number of 16 is sulphur.

Hence the element is **sulphur**.

Question 89

(a)

(i) The rule violated is **Aufbau principle**

(ii) The rule violated is **Hund's rule**

(iii) It is correct. No rule is violated

(b)

(i) Electronic configuration of copper is

	3d						4s	
[Ar]	↑↓	↑↓	↑↓	↑↓	↑↓		↑	← Last electron in copper

Hence the quantum numbers for the last electron in copper are:

$$n = 4,\ l = 0,\ m_l = 0,\ m_s = +{}^{1}/_{2}$$

(ii) Electronic configuration of cobalt is

	3d						4s
[Ar]	↑↓	↑↓	↑	↑	↑		↑↓

26th electron in cobalt

Hence the quantum numbers for the 26th electron in cobalt are:

$$n = 4, l = 0, m_l = 0, m_s = +{}^{1}/_{2}$$

Question 90

(a) (i) Is the rule which govern the arrangement of electrons in an orbital which instructs that maximum number of electrons in the orbital is two and the two electrons must have opposite spin.

(ii) Is the rule which govern the arrangement of electrons in degenerated orbitals which restricts pairing of electrons in degenerated orbitals until each orbital is singly occupied with electron spinning in parallel direction.

(iii) Represents wavelength of energy emitted when excited electron return to lower energy excited state or to ground state.

(iv) Is the level used to specify the main energy level in which electron is placed. It is usually represented by using whole numbers whereby greater value of the number means higher energy level.

(b) (i) K,L,M and N stands for $\mathrm{n} = 1, \mathrm{n} = 2, \mathrm{n} = 3$ and $\mathrm{n} = 4$ principal numbers respectively.

Thus the full configuration is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5\ 4s^2$

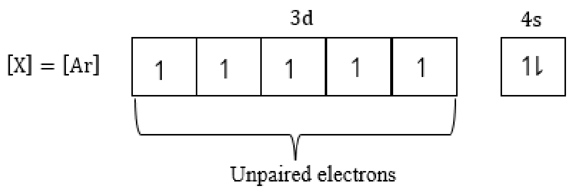
(ii) Atomic number of X is the total number of electrons in the atom. That is $Z = 2 + 8 + 13 + 2 = 25$

Hence the atomic number is 25

(iii) **Four** principle quantum numbers (represented by letters: K, L, M and N).

(iv) Sublevels with electrons are $1s, 2s, 2p, 3s, 3p, 3d$ and $4s$. Hence there are **seven** sublevels with electrons.

(v) The sublevel with unpaired electrons is 3d-sublevel as illustrated below:



Hence the total number of unpaired electrons is **five**

(vi)The block letter, is the letter representing sublevel with highest energy which is 3d-sublevel. Hence the element belongs to **d-block.**

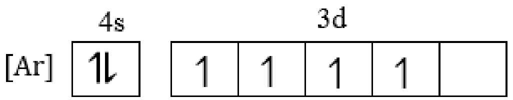
Question 91

(a)(i) Bohr's atomic model.

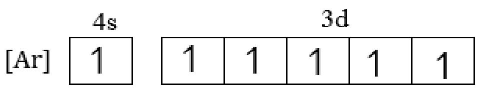
(ii) The model viewed an electron as being placed at certain distance from nucleus; this is against Heisenberg uncertainty principle (one of wave mechanical atomic model) which states that, "It is impossible at any moment to predict simultaneously the exact position and velocity of an electron in an atom".

(b) **First example:** Electronic configuration of Cr (Atomic number = 24)

In accordance to Aufbau principle, the electronic configuration of Cr was expected to be;

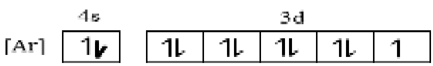


However due to stability of half-filled orbitals electronic configuration, the practical electronic configuration of Cr is;

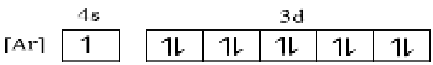


Second example: Electronic configuration of Cu (Atomic number = 29)

In accordance to Aufbau principle, the electronic configuration of Cu was expected to be:



However due to stability of completely filled orbitals electronic configuration, the practical electronic configuration is;



(d) Using $\Delta E = nhf$

But $f = \frac{c}{\lambda}$

Then $\Delta E = \frac{nhc}{\lambda}$

But from Rydberg equation: $\frac{1}{\lambda} = R_H\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$

Thus; $\Delta E = nhcR_H\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$

Since 1g of hydrogen atoms is equivalent to one mole of hydrogen atoms which by definition of mole, it represents 6.02×10^{23} atoms, $n = 6.02 \times 10^{23}$

Also by giving line of lowest energy in visible spectrum means that $n_1 = 2$ and $n_2 = 3$.

Thence substituting;

$$\Delta E = 6.02 \times 10^{23} \times 6.626 \times 10^{-34} J\cancel{s} \times 3 \times 10^8 m\cancel{s^{-1}} \times 1.097 \times 10^7 \cancel{m^{-1}} \times \left(\frac{1}{2^2} - \frac{1}{3^2}\right)$$

$$= 182324J = 182.324kJ$$

Hence the energy emitted was $182.324kJ$

Question 92

(a)

(i) Nuclide is a **nucleus** of a particular atom characterized by a defined **atomic number** and **mass number.**

(ii) Isotones are **atoms** of different **elements** with different atomic numbers but the **same neutron number.**

(iii) Isotopes are **atoms** of the **same element** with the same atomic number but **differ in their mass number.**

(b)

First information:

If two electrons occupy the same orbital, the electrons must have the opposite spin. This because, if two electrons are in the same orbital, have the same first three quantum numbers (n, l and m_l); so they must spin in opposite direction in order to have different fourth quantum numbers (m_s)

Second information:

Maximum number of electrons in an orbital is two. This because, there are only two directions of spinning and therefore only two values for m_s. Having more than two electrons in an orbital would make more than one electron to have the same value of m_s and hence all four quantum numbers would be the same.

(c)

From Bohr theory;

$\Delta E = nhf = \frac{nhc}{\lambda}$

Where $n = 6.02 \times 10^{23}mol^{-1}$ if the energy is in J/mol

Substituting

$145.684 \times 10^3 J\cancel{mol^{-1}} = \frac{6.02\times10^{23}\cancel{mol^{-1}}\times6.63\times10^{-34}\cancel{Js}\times3\times10^8 m\cancel{s^{-1}}}{\lambda}$

From which $\lambda = 8.219 \times 10^{-7}m$

But $1A^0 = 10^{-10}m$

So $8.219 \times 10^{-7}m = \frac{8.219\times10^{-7}}{10^{-10}}A^0 = 8219A^0$

Hence the wavelength of radiation is $8219A^0$

Question 93

(b)

(i) Atom is the smallest individual particle of an element that can be take part in the chemical reaction.

(ii) Element is the substance which cannot, by any known chemical process, be split up into two or more simpler substances.

(c) An element is made of one type of atoms. Thus atom is the smallest amount of an element.

(d)

(i) Atomic spectrum consists of separate frequency lines with very little or no grouping at all while continuous spectrum consists of all possible lines of all frequencies over a wide range of frequencies with no clear cut between the lies.

(ii) Minimum energy to cause excitation of electron in the hydrogen atom is obtained when the electron transist from $n = 1$ to $n = 2$.

This energy has maximum wavelength possible for causing the excitation.

Then from Rydberg equation:

$$\frac{1}{\lambda} = R_H\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

Substituting $\frac{1}{\lambda} = 1.1 \times 10^7\left(\frac{1}{1^2} - \frac{1}{2^2}\right)$

From which $\lambda = 1.21 \times 10^{-7}$m

Thus the maximum possible wavelength to cause the excitation of the electron in the hydrogen atom is 1.21×10^{-7}m

But the given wavelength of 1.22×10^{-6}m is greater than maximum possible wavelength for the excitation to occur and hence **no excitation** will occur because the given wavelength has too small energy to cause the excitation.

Question 94

(c)

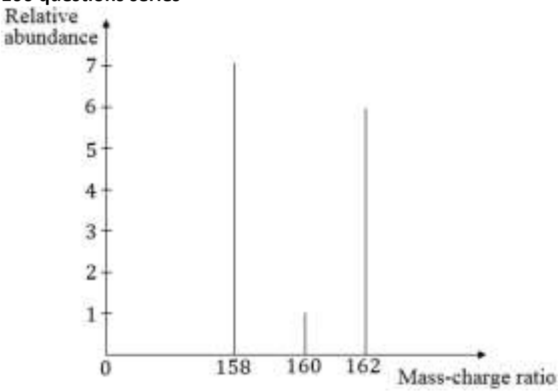
(i) Are levels used to describe electrons at their stationary states.

(ii) Is the frequency spectrum of separate lines with very little or no grouping at all which are formed during transition of electrons between quantized energy levels within an atom.

(iii) Is an instrument which measures masses and relative concentrations of atoms or molecules.

(b)

(i)



(ii) With two isotopes, bromine can form three different diatomic molecules (Br_2) as follows:

$^{79}Br - {}^{79}Br$ with peak at 158 ($M_1 = 158, I_1 = 6$)

$^{79}Br - {}^{81}Br$ with peak at 160 ($M_2 = 160, I_2 = 1$)

$^{81}Br - {}^{81}Br$ with peak at 162 ($M_3 = 162, I_3 = 5$)

Then average molecular mass, $M_r = \frac{M_1I_1+M_2I_2+M_3I_3}{I_1+I_2+I_3}$

$= \frac{(158\times6)+(160\times1)+(162\times5)}{6+5+1}$ a.m.u= 159.83 a.m.u

But since the bromine molecule is diatomic;

Average atomic mass $= \frac{\text{Average molecular mass}}{2}$

$$= \frac{159.83}{2}\text{a.m.u} = 79.915\text{ a.m.u}$$

Hence the average atomic mass is 79.915 a.m.u

(iii) Because it is weighted **average** of mass number of ^{79}Br and ^{81}Br isotopes.

(iv) Let the relative abundance of ^{79}Br be y%

Then that of ^{81}Br will be (100-y)%

Then using $A_r = \frac{A_1P_1+A_2P_2}{100}$

Substituting $79.915 = \frac{79y+81(100-y)}{100}$

From which $y = 54.25$ and $100 - y = 45.75$

Hence the relative abundance is
54.25% for ^{79}Br and 45.75% for ^{81}Br

Question 95

(a)

(i) Radiant energy is emitted or absorbed in small packets as quanta.

(ii) It is not passible at any moment to predict accurately both position and momentum of an electron in an atom simultaneously.

(iii)

1. It does not explain the spectrum of multi-electron atoms.
2. It does not explain the relative intensity of spectral lines.
3. It does not explain why covalent bonding makes a molecule stable.
4. It viewed an electron as placed at certain fixed distance from nucleus which is against Heisenberg's uncertainty principle.
5. No clear justification was given for the quantization of angular momentum of electron and presence of stationary states.
6. It suggested that electron move in circular orbit in two dimensions which is against quantum mechanical atomic model which suggests that the motion is in three dimensions on the elliptical path.
7. It does not explain presence of hyperfine spectral lines.
8. It does not explain Zeeman and Stark effect.

(b)

(i) Is the energy required to remove completely the most loosely bound electron each from one mole of gaseous atom or ion at given conditions of temperature and pressure. It is the measure of strength of nuclear attractive force in an atom.

(ii) When electron jump from one energy level to another, the energy change (ΔE) is given by:

$\Delta E = hf = \frac{hc}{\lambda}$ $\left(f = \frac{c}{\lambda}\right)$

But from Rydberg equation;

$\frac{1}{\lambda} = R_H\left(\frac{1}{{n_1}^2} - \frac{1}{{n_2}^2}\right)$

Then $\Delta E = hcR_H\left(\frac{1}{{n_1}^2} - \frac{1}{{n_2}^2}\right)$

When an electron is ionised in hydrogen atom, the electron jumps from $n = 1$ to infinity (convergence limit where $n = \infty$).

Thus $n_1 = 1$ and $n_2 = \infty$

Substituting:

$\Delta E = 6.626 \times 10^{-34} \times 3 \times 10^{8} \times 1.097 \times 10^{7}\left(\frac{1}{1^2} - \frac{1}{\infty^2}\right)$

$= 2.1806 \times 10^{-18}\text{J/electron}$

But from the definition of ionisation energy, it must be per mole of electrons

And 1 mol of $e^- \equiv 6.02 \times 10^{23}\text{electron/mol}$

Thus the ionisation energy

$= 2.1806 \times 10^{-18}\text{J/electron} \times 6.02 \times 10^{23}\text{electron/mol}$

$1312721.2\text{J/mol} = 1312.7212\text{kJ/mol}$

The ionisation energy of hydrogen is 1312.7212kJ/mol.

(c) Using Rydberg equation:

$\frac{1}{\lambda} = R_H\left(\frac{1}{{n_1}^2} - \frac{1}{{n_2}^2}\right)$

For first line in the Balmer series;

$n_1 = 2, n_2 = 3$ and $\lambda = \lambda_B = 6863A^O$

$\frac{1}{\lambda_B} = R_H\left(\frac{1}{2^2} - \frac{1}{3^2}\right) = \frac{5}{36}R_H \ldots\ldots(i)$

For first line in Lyman series:

$n_1 = 1, n_2 = 2$ and $\lambda = \lambda_L = 6863A^O$

$\frac{1}{\lambda_L} = R_H\left(\frac{1}{1^2} - \frac{1}{2^2}\right) = \frac{3}{4}R_H \ldots\ldots(ii)$

Taking $\frac{(\text{i})}{(\text{ii})}$ gives; $\frac{\lambda_L}{\lambda_B} = \frac{5}{36} \times \frac{4}{3} = \frac{5}{27}$

Or $\lambda_L = \frac{5}{27}\lambda_B = \frac{5}{27} \times 6863A^O = 1271A^O$

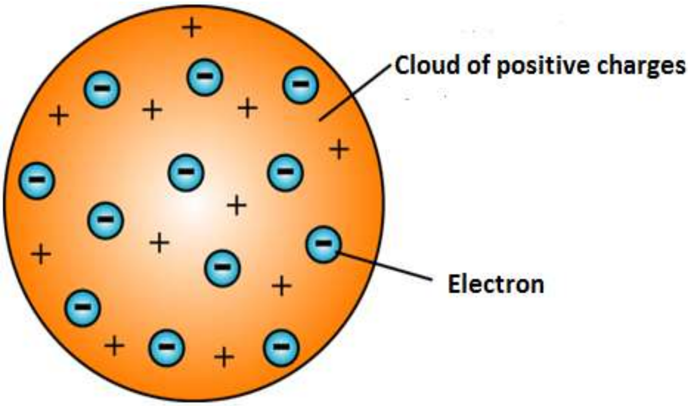
The wavelength is $1271A^O$

Question 96

(c)

(i) This is because the model suggested that the atom to be full of some positive fluid like a **pudding** in which electrons were embedded like **plums** on that pudding.

(ii) According to Thomson; "An atom possesses a spherical shape in which the positive charge is uniformly distributed with electrons embedded over it"



(iii)

Advantages:

1. It explains satisfactorily how heating a substance starts radiating light.
2. It explains neutral character of an atom.

Disadvantages:

1. It does not explain the stability of an atom.
2. It is not supported by Rutherford's $\propto -$ scattering experiment.

(d)

(i) If electrons were transferred to the plastic, the plastic has extra electrons and becomes negative, since electrons are negative. If my fingers lost electrons, they would have more protons than electrons and become positive. Positive and negative attract.

(ii) Since the balloon has extra electrons from being rubbed on your hair, the balloon is negatively charged. When you bring it over to the wall, the negative balloon repels the negative electrons in the wall and leaves an area of positive charge. The negative balloon is then attracted to the positive wall.

Question 97

(a)

(i) Is a particular arrangement of electrons in different orbitals of an atom.

(ii) Is the energy sublevel in which electron is placed and specifies the shape of an orbital of an electron.

(iii) Is the splitting of a spectral line of atoms or molecules subjected to the electric field.

(iv) Are levels used to describe electrons at their stationary states.

(b)

(i) 2

(ii) 0

(iii) $n = 3, l = 0, m = 0, s = +\frac{1}{2}$

(iv) 2

(v) $l = 2, m_l = -2, -1, 0, +1, +2$

(c)

(i) Using; $A_r = \frac{A_1P_1 + A_2P_2}{P_1 + P_2}$

Substituting; $A_r = \frac{(14\times95.12)+(15\times4.88)}{100} = 14.05$

The relative atomic mas of nitrogen is 14.05

(ii)

A. Different mass-charge $\left(^{m}/_{z}\right)$ratio

B. Change in the amount of magnetic field

(iii) No difference

Explanation

Since the two isotopes have the same number of electrons and hence the same electronic configuration, no difference will be observed in their chemical reactions.

Question 98

(c)

(i) In the ground state, electrons tend to occupy the orbital with minimum energy.

(ii) No two electrons in the same atom can have the same all four quantum numbers.

(iii) Electrons are not allowed to pair up unless the empty degenerate orbitals are singly occupied with parallel spins electrons.

(iv) For d and f orbitals the full or half full configurations in the sub –shells are very stable configuration with minimum energy content.

(b) Energy of one photon is given by the following Planck's equation; $E = hf = \frac{hc}{\lambda}$

Substituting $E = \frac{6.63\times10^{-34}\mathrm{J}\times3\times10^{8}\mathrm{m/s}}{987\times10^{-9}\mathrm{m}} = 2.0152\times10^{-19}\mathrm{J/photon}$

But total energy emitted in $32\mathrm{sec} = 0.52\mathrm{J}$

And the total energy=Energy per photon×Number of photons

Thus Number of photons emitted (in 32s)

$= \frac{\text{Total energy}}{\text{Energy per photon}} = \frac{0.52\mathrm{J}}{2.0152\times10^{-19}\mathrm{J/photon}} = 2.58\times10^{18}$ photons

And number of photons emitted per second

$= \frac{\text{Total number of photons emitted}}{\text{Time taken}} = \frac{2.58\times10^{18}\text{ photons}}{32\mathrm{sec}}$

$$= 8.0625\times10^{16}\mathrm{photon/sec}$$

Hence 8.0625×10^{16} photons were emitted in one second.

(c)

(i) Disagree

Reason:

Due to isotopy, it is possible for atoms of the same element to have different number of neutrons.

(ii) Agree

Reason:

Atoms of the same element may have different number of neutrons and hence different atomic masses.

Question 99

(a)

(i)

1. Matter is made up of very small indivisible particles called atoms.
2. Atoms can neither be created nor destroyed.
3. Atoms of the same element are all alike but differ from the atoms of all other elements.
4. Chemical combination takes place between small whole numbers of atoms.

(ii) Different atoms have different atomic number and hence different number of electrons. Different number of electrons means different electronic configuration and therefore different chemical properties.

(b)

s/n	Number of neutrons	Number of electrons	Atomic number	Mass number	Symbol	Electronic configuration
(i)	13	11	11	24	${}^{24}_{11}X$	$1s^2 2s^2 2p^6 3s^1$
(ii)	7	8	8	15	${}^{15}_{8}X$	$1s^2 2s^2 2p^4$
(iii)	17	18	18	35	${}^{35}_{18}X$	$1s^2 2s^2 2p^6 3s^2 3p^6$
(iv)	16	16	16	32	${}^{32}_{16}X$	$1s^2 2s^2 2p^6 3s^2 3p^4$

(c) From Plank's equation

$$E = hf = \frac{hc}{\lambda}$$

But from Rydberg equation:

$\frac{1}{\lambda} = R_H\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$

Thus $E = hcR_H\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$

In removing an electron from hydrogen in ground state, the electron move from $n = 1$ to $n = \infty$.

Thus, $n_1 = 1$ and $n_2 = \infty$

Substituting:

$E = 6.63 \times 10^{-34}Js \times 3 \times 10^{8}m/s \times 1.09678 \times 10^{7} \times m^{-1}\left(\frac{1}{1^2} - \frac{1}{\infty^2}\right)$

$= 2.18 \times 10^{-18}J$

The minimum energy required is $2.18 \times 10^{-18}J$

Question 100

(a)

(v) In the ground state, electrons tend to occupy the orbital with minimum energy.

(vi) No two electrons in the same atom can have the same all four quantum numbers.

(vii) Electrons are not allowed to pair up unless the empty degenerate orbitals are singly occupied with parallel spins electrons.

(viii) For d and f orbitals the full or half full configurations in the sub –shells are very stable configuration with minimum energy content.

(b) From the Heisenberg's equation;

$$\Delta x \times m\Delta v \approx \frac{h}{4\pi}$$

Rearranging the above equation gives;

$$\Delta x \approx \frac{h}{4\pi \times m\Delta v}$$

From which it can be deduced that the uncertainty in determining position (Δx) of an object of mass, m, varies inversely proportional to the uncertainty in the determination of velocity (Δv) of the object.

That means if the velocity of the object is determined with more accuracy then its position will determined with less accuracy as suggested in Heisenberg uncertainty principle.

(c) Using; $\mathrm{Z = A - Number\ of\ neutrons}$
Atomic number of x $= 35 - 18 = 17 =$ Atomic number

For neutral atom; atomic number = number of electrons

Thus **number of electrons in Y is 17**

Using Number of neutrons = A – Z

Number of neutrons in Y = 37 – 17 = 20

Thus **number of neutrons in Y is 20**

(d) Energy of one photon is given by the following Planck's equation; $\mathrm{E = hf = \frac{hc}{\lambda}}$

Substituting $\mathrm{E} = \frac{6.63\times10^{-34}\mathrm{J}\times3\times10^{8}\mathrm{m/s}}{987\times10^{-9}\mathrm{m}} = 2.0152\times10^{-19}\mathrm{J/photon}$

But total energy emitted in $32\mathrm{sec} = 0.52\mathrm{J}$

And the total energy=Energy per photon×Number of photons

Thus Number of photons emitted (in 32s)

$$= \frac{\text{Total energy}}{\text{Energy per photon}} = \frac{0.52\mathrm{J}}{2.0152\times10^{-19}\mathrm{J/photon}} =$$

2.58×10^{18} photons

And number of photons emitted per second

$$= \frac{\text{Total number of photons emitted}}{\text{Time taken}} = \frac{2.58\times10^{18}\ \text{photons}}{32\text{sec}}$$

$$= 8.0625\times10^{16}\text{photon/sec}$$

Hence 8.0625×10^{16} photons were emitted in one second.

(e)

(iii) Disagree **(0.5)**
Reason:

Due to isotopy, it is possible for atoms of the same element to have different number of neutrons.

(iv) Agree
Reason:

Atoms of the same element may have different number of neutrons and hence different atomic masses.