Modern Physics - Answers

Question 1: Ionisation Energy

A) n is any whole number greater than zero. ie. $n = 1, 2, 3, \ldots$

B) Since $E_n = \frac{chR}{n^2}$, if $n = \infty$ then $E_n = 0$

Therefore ionisation energy of hydrogen $= E_1$

$$E_1 = \frac{chR}{1^2} = \frac{3.0 \times 10^{8} \times 6.63 \times 10^{-34} \times 1.097 \times 10^{7}}{1}$$

$$E_1 = \underline{\underline{21.8 \times 10^{-19}\ J}}$$

Question 2: Atomic Spectra

A) The circles represent the specific energy levels in which the electron in a hydrogen atom can exist. The lines show the transitions between the energy levels that the electron can make as it absorbs or emits energy.

B) As an electron falls from a higher to lower energy level, the energy is lost by the emission of a photon. The energy of the photon emitted is equal to the difference in the atoms energy as the electron is in either state. The wavelength shown on the diagram relates to the frequency of light emitted (since $v = f\lambda$) and therefore to the energy of the photon (since $E = hf$). The wavelength shown are the wavelengths observed in hydrogen spectrum.

C) $n = 5 \rightarrow n = 2$

$$\frac{1}{\lambda} = R\left(\frac{1}{2^2} - \frac{1}{5^2}\right)$$

$$= 1.097 \times 10^{7}\left(\frac{1}{4} - \frac{1}{25}\right)$$

$$= 2303700$$

$$\therefore \lambda = \frac{1}{2303700} = 4.34 \times 10^{-7} = 434 \times 10^{-9} = \underline{434\,nm}$$

d. At ground state, $n=1$

$$\therefore E_n = \frac{-chR}{1^2} = -3.00\times10^8 \times 6.63\times10^{-34} \times 1.097\times10^7$$

$$E_n = -2.18 \times 10^{-18}\ J$$

Question 3: Light emitted from a star.

A) Ultraviolet range of EM spectrum (any transition to $n=1$ is in UV range)

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

$$= 1.097\times10^7\left(\frac{1}{1^2} - \frac{1}{2^2}\right)$$

$$= 1.097\times10^7 \times 0.75$$

$$= 8227500$$

$$\therefore \lambda = \frac{1}{8227500} = 1.2154\times10^{-7} = 121\ nm.$$

c) Lines in the visible spectrum are all associated with the Balmer series where electrons fall to $n=2$ from higher energy states.* The red line has the lowest frequency of hydrogen emission lines in the visible spectrum, therefore it must represent the smallest change in energy (since $E=hf$). The smallest transition in the Balmer series is from $n=3$ to $n=2$.

*Even the smallest transitions to $n=1$ emit photons in the UV range; the largest transitions to $n=3$ emit photons in the infra-red range.

d) In ground state $n=1$

$$E_n = \frac{-hcR}{1^2} = -6.63\times10^{-34}\times3\times10^8\times1.097\times10^7 = 2.18\times10^{-18}\ J.$$

e) $n=5 \rightarrow n=3$

$$E_n = hcR\left(\frac{1}{m^2} - \frac{1}{n^2}\right) = 6.63\times10^{-34}\times3\times10^8\times1.097\times10^7\left(\frac{1}{3^2} - \frac{1}{5^2}\right) = 1.55\times10^{-19}\ J$$

Question 4:

A) When an atom is in an excited state, the electron(s) are in an energy state above the ground state (i.e. $n>1$). The electron is then able to emit a photon in order to lose energy as it falls to a lower energy state.

B) The photon emitted during transition "A" will have the longest wavelength. This is because the difference between the energy states in this transition is less than transitions B and C. Since $E=hf$, the small change in energy will result in a low frequency photon and since $\lambda = \frac{v}{f}$ the lowest frequency photon will have the longest wavelength.

C) Transition C: $f = 3.27 \times 10^{15}$

NB: This frequency is incorrect

$$E = hf$$
$$= 6.63 \times 10^{-34} \times 3.27 \times 10^{15}$$
$$= 2.16 \times 10^{-18}$$

D) Transition B: $\Delta E = (21.8 - 5.4) \times 10^{-19}$

$$= 16.4 \times 10^{-19}\ J$$

$$E = hf \quad \therefore f = \frac{E}{h} = \frac{16.4 \times 10^{-19}}{6.63 \times 10^{-34}} = 2.47 \times 10^{15}$$

$$v = f\lambda \quad \therefore \lambda = \frac{v}{f} = \frac{3.0 \times 10^{8}}{2.47 \times 10^{15}} = \underline{\underline{121\ nm}}.$$

E) Paschen series $(m=3)$

Second line of series: $n=5$.

$$\frac{1}{\lambda} = R\left(\frac{1}{m^2} - \frac{1}{n^2}\right)$$
$$= 1.097 \times 10^{7}\left(\frac{1}{3^2} - \frac{1}{5^2}\right)$$
$$= 780\ 089$$

$$\therefore \lambda = \frac{1}{780089} = \underline{\underline{1.282 \times 10^{-6}\ m}}$$

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F) Bohrs atomic theory says that electrons can only exist in specific orbits around the nucleus. These orbits are characterised by specific quantised energy states. As an electron transitions from a high energy to lower energy state, a photon is emitted. The energy of the photon is equal to the difference between the energy states. For the second line of the Paschen series, an electron in a hydrogen atom is transitioning from the n=5 state to the n=3 state.

G) $n=5 \rightarrow n=3$

$$E_{5\to3} = chR\left(\frac{1}{3^2} - \frac{1}{5^2}\right)$$

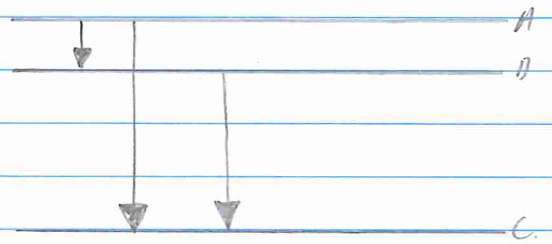
$$= 3{\cdot}00 \times 10^{8} \times 6{\cdot}63 \times 10^{-34} \times 1{\cdot}097 \times 10^{7}\left(\frac{1}{9} - \frac{1}{25}\right)$$

$$= \underline{1{\cdot}55 \times 10^{-19}}$$

Question 5: Atomic Spectra.

A). Excitation energy is the energy required or provided to an atom that causes an electron to transition from a low to higher energy state.

B)



C) Smallest wavelength ≡ highest frequency ≡ greatest ΔE
∴ Smallest wavelength associated with $A \rightarrow C$ transition.

D) ΔE from $A \rightarrow C$ $= (14{\cdot}7 - 0{\cdot}496) \times 10^{-19}$
$= 14{\cdot}2 \times 10^{-19}$ J

$E = hf$ ∴ $f = \dfrac{E}{h} = \dfrac{14{\cdot}2 \times 10^{-19}}{6{\cdot}63 \times 10^{-34}} = 2{\cdot}14 \times 10^{15}$

$v = f\lambda$ ∴ $\lambda = \dfrac{c}{f} = \dfrac{3{\cdot}0 \times 10^{8}}{2{\cdot}14 \times 10^{15}} = \underline{140 \text{ nm}}$

Question 6: The Hydrogen Atom

A) i) $E_2 = \frac{-chR}{2^2}$

$= \frac{-3.00\times10^{8} \times 6.63\times10^{-34} \times 1.097\times10^{7}}{4}$

$= \underline{5.45 \times 10^{-19}}$

ii) $\frac{1}{\lambda} = R\left(\frac{1}{2^2} - \frac{1}{5^2}\right)$

$= 1.097\times10^{7}\left(\frac{1}{4} - \frac{1}{25}\right)$

$= 2303700$

$\therefore \lambda = \frac{1}{2303700}$

$= 434$ nm.

B). $n=2 \rightarrow n=\infty$ (ionised)

Associated $\Delta E = E_2$ (calculated above)

$E = hf$

$\therefore$ minimum $f = \frac{E}{h}$

$= \frac{5.45\times10^{-19}}{6.63\times10^{-34}}$

$= \underline{8.23 \times 10^{14}}$

c) The lowest frequency light is associated with the smallest difference between energy states (since $E=hf$). The smallest difference between energy states is for adjacent energy states. Therefore the lowest frequency photon will be associated with the transition from $n=3$ to $n=2$.