Sem 1 2004/2005 : Q2a : Quantum Theory

Sem 1 (2004/2005) : Quantum Theory

Q2 (a) Rydberg equation : $\frac{1}{\lambda} = R_H \left[\frac{1}{n_1^2} - \frac{1}{n_2^2} \right]$

For Lyman Series : $n_1 = 1$
$n_2 = 2, 3, 4 \ldots\ldots$

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

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

$$\frac{1}{\lambda} = R_H - \frac{R_H}{n_2^2} \quad \ldots\ldots\ldots\ldots (i)$$

$\Delta E = \frac{hc}{\lambda}$ and $\Delta E = h\nu$; $c = \lambda \nu$ @ $\frac{1}{\lambda} = \frac{\nu}{c}$

Equation (i) will be :

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

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

or $\nu = -cR_H \left(\frac{1}{n_2^2} \right) + cR_H$

similar to: $y = -mx + c$ (straight line, negative gradient)

ν/Hz	n_2	n_2^2	$1/n_2^2$
2.466×10^{15}	2	4	0.2500
2.923×10^{15}	3	9	0.1111
3.083×10^{15}	4	16	0.06250
3.157×10^{15}	5	25	0.04000
3.197×10^{15}	6	36	0.02780

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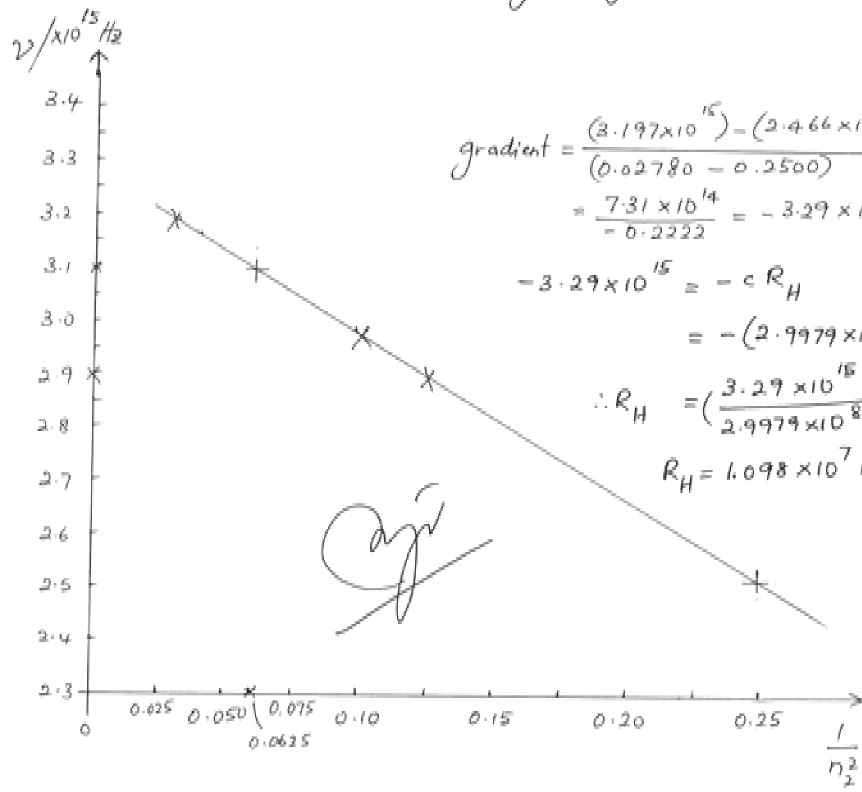
Q_2 (a) Cont.....

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

Plot ν versus $1/n^2$: negative gradient $= -cR_H$



$$\text{gradient} = \frac{(3.197\times10^{15}) - (2.466\times10^{15})}{(0.02780 - 0.2500)}$$

$$= \frac{7.31\times10^{14}}{-0.2222} = -3.29\times10^{15}$$

$$-3.29\times10^{15} = -cR_H$$

$$= -(2.9979\times10^{8})R_H$$

$$\therefore R_H = \left(\frac{3.29\times10^{15}}{2.9979\times10^{8}}\right)\ m^{-1}$$

$$R_H = 1.098\times10^{7}\ m^{-1}\ \text{(Ans)}$$

Q_2 (b) Energy absorbed, $E_1 = \frac{hc}{\lambda_1}$ $\quad(\lambda_1 = 440\ nm)$

Energy emitted, $E_2 = \frac{hc}{\lambda_2}$ $\quad(\lambda_2 = 670\ nm)$

$\therefore$ Energy used for the photosynthesis $= E_1 - E_2 = \Delta E$

$$\therefore \Delta E = \frac{hc}{\lambda_1} - \frac{hc}{\lambda_2} = hc\left(\frac{1}{\lambda_1} - \frac{1}{\lambda_2}\right)\ J\ photon^{-1}$$

OR

$$\Delta E = hc\left(\frac{1}{\lambda_1} - \frac{1}{\lambda_2}\right)(6.022\times10^{23})\ J\ mol^{-1}$$

$$= (6.626\times10^{-34})(2.9979\times10^{8})(6.022\times10^{23})\left(\frac{1}{(440\times10^{-9})} - \frac{1}{(670\times10^{-9})}\right)$$

$$= 9.333\times10^{4}\ J\ mol^{-1}$$

$$\Delta E = 93.33\ kJ\ mol^{-1}\ \text{(Ans)}$$

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Q_2 (c) one electron in one of the 3p orbital.

	n	l	m_l	m_s
	3	1	+1	$+\frac{1}{2}$
or	3	1	+1	$-\frac{1}{2}$
or	3	1	0	$+\frac{1}{2}$
or	3	1	0	$-\frac{1}{2}$
or	3	1	−1	$+\frac{1}{2}$
or	3	1	−1	$-\frac{1}{2}$

Q_2 (d) $\Delta E = h\nu = \frac{hc}{\lambda}$ ...... (i)

Rydberg equation: $\frac{1}{\lambda} = R_H\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$ ..... (ii)

For the ionization energy: $n_1 = 1 \quad n_2 = \infty$

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

$$\frac{1}{\lambda} = R_H \quad \text{-------(iii)}$$

$$\therefore \Delta E = \frac{hc}{\lambda} = R_H hc \ \text{J atom}^{-1}$$

$$\Delta E = (1.09678 \times 10^7)(6.626 \times 10^{-34}\ \text{Js})(2.9979 \times 10^8\ \text{ms}^{-1})$$

$$= 2.179 \times 10^{-18}\ \text{J atom}^{-1}$$

$$= (2.179 \times 10^{-18})(6.022 \times 10^{23})\ \text{J mol}^{-1}$$

$$= 1.312 \times 10^6\ \text{J mol}^{-1}$$

$$\Delta E = 1.312 \times 10^3\ \text{kJ mol}^{-1}\ (\text{Ans})$$

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