Emission of Light from Hydrogen and Metal Atoms

Lab Report 2

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Name

1:00 pm
Time

M T W R F

PART A.

Calibration of Wavelength Scale with Helium Emission Lines

Line	Color	Wavelength (nm)	Scale Reading
1	red	707	439nm
2	red	668	461nm
3	yellow	588	485nm
4	light green	501	495nm
5	darker green	492	582nm
6	blue-green	471	667nm
7	blue-violet	447	705nm

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Hydrogen Emission Lines

Line	Scale Reading	Wavelength (from calibration curve)	Assigned Transition
1	658 nm	660	3→2
2	480 nm	487	4→2
3	428 nm	437	5→2
4	400	410 nm	6→2

$y = 0.971x + 21.5$
-21.5

$\frac{-21.5 + y}{0.971} = \frac{0.971x}{0.971}$

PART B.
Flame Tests

Solution	Color and Intensity	Duration
$BaCl_2$	Yellow weak	1.53
$CuCl_2$	Green strong	2.96
KCl	Pink/red weak	3.67
LiCl	Pink strong	13.62
NaCl	orange strong	51.09
$SrCl_2$	dark orange/Red weak	13.87
Unknown	light pink weak	5.57

Unknown number: 2

Identity of unknown: KCl

1. Using the calibration graph and the scale readings for hydrogen, determine the wavelengths for hydrogen emission; enter the results on your data sheet.

2. The four hydrogen lines correspond to the $3 \rightarrow 2$, $4 \rightarrow 2$, $5 \rightarrow 2$, and $6 \rightarrow 2$ transitions. Assign a transition to each wavelength and enter the results on your data sheet. Hint: The lowest wavelength, 410 nm, corresponds to the highest-energy transition. The $3 \rightarrow 2$ transition emits the least energy. $3 \rightarrow 2$ least energy

$6 \rightarrow 2$ most energy

3. Determine the Rydberg constant from the data. Since for $n_\ell = 2$:

$$\frac{1}{\lambda} = R_H\left(\frac{1}{n_\ell^2} - \frac{1}{n_h^2}\right) \quad \text{Equation 1}$$

$n_\ell = 2$

↳ constant

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

$$\frac{1}{\lambda} = -R_H\left(\frac{1}{n_h^2}\right) + \frac{R_H}{4} \quad \text{Equation 2}$$

$$\underbrace{\frac{1}{\lambda}}_{y} = \underbrace{-R_H}_{m}\underbrace{\left(\frac{1}{n_b^2}\right)}_{x} + \underbrace{\frac{R_H}{4}}_{b}$$

which is of the form y = mx + b

That is, Equation 2 is of the form of a straight line; hence the Rydberg constant is the negative of the slope. Plot $y = 1/\lambda$ vs. $x = 1/n_h^2$ on the graph paper provided, then determine the slope $m = -R_H$. Be sure to label the axes of the graph. As a check, compare this value of R_H with the y-intercept, which should equal $R_H/4$. To make it easy, set up a table (next page) for the data. Use meters for wavelength; the units of R_H will be meter^{-1}. Show your calculations in the space below:

660 nm $\cdot \frac{1}{10^9}$

487 nm $\cdot \frac{1}{10^9}$

437 nm $\cdot \frac{1}{10^9}$

410 nm → Given in doc

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λ (meter)	n_h	n_h^2	$\frac{1}{\lambda} = y(m^{-1})$	$\frac{1}{n_h^2} = x$
$6.60 \cdot 10^{-7}$	3	$3^2 = 9$	$1.51 \cdot 10^6$	$\frac{1}{9} = 0.111$
4.87×10^{-7}	4	$4^2 = 16$	$2.05 \cdot 10^6$	$\frac{1}{16} = 0.0625$
4.37×10^{-7}	5	$5^2 = 25$	$2.29 \cdot 10^6$	$\frac{1}{25} = 0.04$
$410 \cdot 10^{-9}$	6	$6^2 = 36$	$\frac{1}{410 \cdot 10^{-9}}$ $2.43902 \cdot 10^6$	$\frac{1}{36} = 0.0278$

4. Calculate the change in energy for each transition, $n_h = 3, 4, 5, 6$ and $n_l = 2$. Show your work and give the results in units of joules.

$h = 6.626 \cdot 10^{-34}$ J·s $\quad R_H = 1.097 \times 10^7$

$c = 3.00 \cdot 10^8$ m/s

$\Delta E = \frac{hc}{m} = hcR_H\left(\frac{1}{n_l^2} - \frac{1}{n_h^2}\right)$ Equation 3

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

$3 \rightarrow 2$: $\Delta E = 3.028 \cdot 10^{-19}$ J

$4 \rightarrow 2$: $\Delta E = 4.088 \cdot 10^{-19}$ J

$5 \rightarrow 2$: $\Delta E = 4.579 \cdot 10^{-19}$ J

$6 \rightarrow 2$: $\Delta E = 4.848 \cdot 10^{-19}$ J

$6 \rightarrow 2$ h c

$\frac{6.626 \cdot 10^{-34} \text{ J·s} \times 3.00 \cdot 10^8 \text{ m/s}}{410 \cdot 10^{-9}} = 4.848 \cdot 10^{-19}$

$5 \rightarrow 2$

$6.626 \cdot 10^{-34}$ J·s $\times 3.00 \cdot 10^8$ m/s $\cdot 1.097 \cdot 10^7 \left(\frac{1}{4} - \frac{1}{25}\right)$

$4 \rightarrow 2$

$6.626 \cdot 10^{-34}$ J·s $\times 3.00 \cdot 10^8 \cdot 1.097 \cdot 10^7 \left(\frac{1}{4} - \frac{1}{16}\right)$

$3 \rightarrow 2$

$6.626 \cdot 10^{-34}$ J·s $\times 3.00 \cdot 10^8 \cdot 1.097 \cdot 10^7 \left(\frac{1}{4} - \frac{1}{9}\right)$

equation 1

5. Derive Equation 3 from Equation 1:

$\frac{1}{\lambda} = R_H\left(\frac{1}{n_l^2} - \frac{1}{n_h^2}\right) \rightarrow \Delta E = \frac{hc}{m} = hcR_H\left(\frac{1}{n_l^2} - \frac{1}{n_h^2}\right)$

$\frac{hc}{\lambda} = hcR_H\left(\frac{1}{n_l^2} - \frac{1}{n_h^2}\right) \rightarrow \Delta E = \frac{hc}{\lambda} \rightarrow \Delta E = \frac{hc}{m} = hcR_H\left(\frac{1}{n_l} - \frac{1}{n_h^2}\right)$ equation 3

6. Determine the ionization energy for the hydrogen atom. The ionization energy is the energy

6. Determine the ionization energy for the hydrogen atom. The ionization energy is the energy necessary to take an electron from its most stable state (n_i = 1 for hydrogen), to a state infinitely far from the nucleus. Use Equation 3 to make the calculation. What is n_b? Show your work. $\Delta E = \frac{hc}{m} = hcR_H \left(\frac{1}{n^2} - \frac{1}{n_b^2}\right)$

from 1 to ∞

$n_b = \infty$

$R_H Z^2 \left(\frac{1}{n^2} - \frac{1}{n_b^2}\right)$

$13.6 \cdot 1^2 \left(\frac{1}{1^2} - \frac{1}{\infty^2}\right)$

$13.6 \cdot 1.6 \cdot 10^{-19}$

$2.176 \cdot 10^{-18}$

n_b E → energy

n_b → large #

horizontal energy

square

$\frac{Z^2}{n^2}$ eV

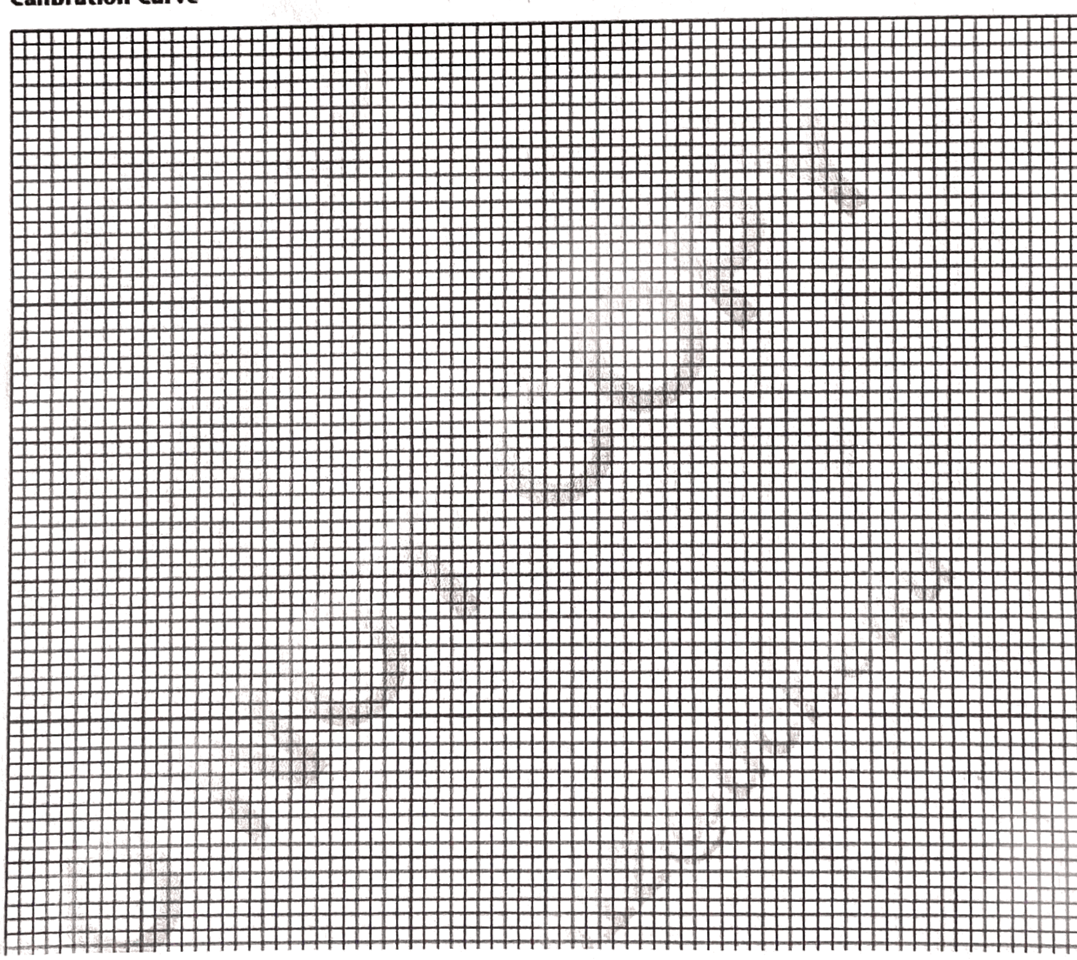
13.60s

$E_n = -13.60 \text{Sel}$

Spread sheet

Calibration Curve

wavelength



Scale Reading

Spread Sheet

Graph to Obtain Rydberg Constant