Atomic Emission Spectra Online Lab Submission

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Class: Chem 1212

Section: Mondays, 10:30 AM - 1:25 PM

Report Sheet

Observed Colors

- NaCl -- `Yellow Orange`
- $CaCl_2$ -- `Red Orange`
- KCl -- `Light Purple`
- $SrCl_2$ -- `Bright Red`
- $BaCl_2$ -- `Yellow Green`
- $CuCl_2$ -- `Teal`

Unknown Code Letter -- `A`

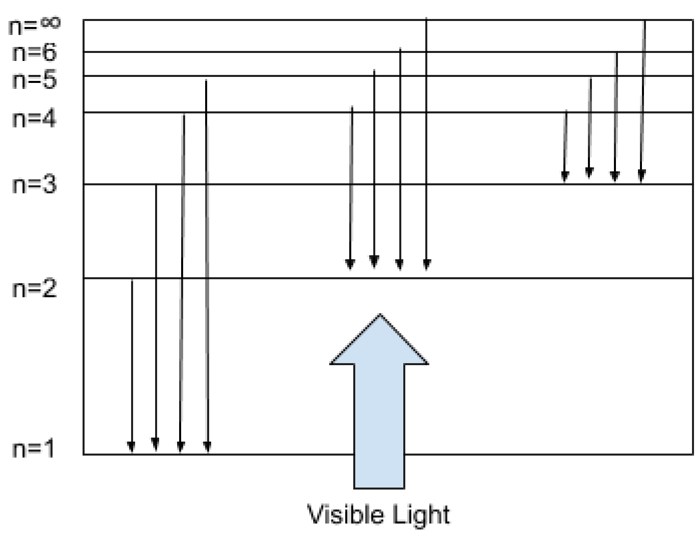
Identification of Unknown -- `Barium Chloride`

Pre-Lab Questions

1.

$$\lambda = \frac{hc}{\Delta E} = \frac{(6.626 \times 10^{-34} J \cdot s) \times (3.00 \times 10^{8} m/s)}{-2.178 \times 10^{-18} \times ((\frac{1}{4})^2 - (\frac{1}{2})^2) J} = 4.87 \times 10^{-7} m \times \frac{10^9 nm}{1m} = 487nm$$

2.



Post-Lab Questions

1. A. The spectrum would be a black strip, with a few colored lines in the visible light spectrum that correspond to the wavelengths that the hydrogen emits. These wavelengths would likely match up with 410 nm (violet), 434 nm (blue), 486 nm (green), and 656 nm (red). B. The spectrum would be an inverse of the previous spectrum, with black lines corresponding to the color being absorbed, and colored regions for non-absorbed wavelengths.
2. Listed below
   - 4 to 1 -- not visible
   - 4 to 2 -- visible
   - 4 to 3 -- not visible
   - 3 to 1 -- not visible
   - 3 to 2 -- visible

$$\lambda = \frac{c}{f} = \frac{3.00 \times 10^{8} m/s}{104.1 \times 10^{6} Hz} = 2.88m$$

$$E_{photon} = \frac{hc}{\lambda} = \frac{(6.626 \times 10^{-34} J \cdot s) \times (3.00 \times 10^{8} m/s)}{2.88m} = 6.90 \times 10^{-27} J$$

$$56 \times 10^{6} km \times \frac{10^{3} m}{1km} \times \frac{1}{3.00 \times 10^{8} m/s} = 190s$$

$$E_{photon} = \frac{hc}{\lambda} = \frac{(6.626 \times 10^{-34} J \cdot s) \times (3.00 \times 10^{8} m/s)}{450 \times 10^{-9} m} = 4.4 \times 10^{-19} J$$

$$E_{total} = E_{photon} \times \#\ of\ photons = 4.4 \times 10^{-19} J \times 7.25 \times 10^{17} = 0.32J$$

$$E_{emittedphoton} = \frac{hc}{\lambda} = \frac{(6.626 \times 10^{-34} J \cdot s) \times (3.00 \times 10^{8} m/s)}{520 \times 10^{-9} m} = 3.8 \times 10^{-19} J$$

$$E_{absorbedphoton} = \frac{hc}{\lambda} = \frac{(6.626 \times 10^{-34} J \cdot s) \times (3.00 \times 10^{8} m/s)}{520 \times 10^{-9} m} = 6.2 \times 10^{-19} J$$

$$E_{heat} = (E_{absorbedphothon} - E_{emittedphoton}) \times \frac{6.0221 \times 10^{23} photons}{1mole} \times \frac{1kJ}{1000J} = 145kJ/mole$$