Physics 214 -- PreLab 3

Atomic Spectra

Name ________________________________

Section _____________ **Date** ________

<u>Your TA will collect this paper at the beginning of your lab section.</u>

Readings: **Physics 214 Lectures 4, 11**

Young & Freedman 36.5, 41.1

1. In a lab experiment, a parallel beam of light is normally incident on a diffraction grating. Diffracted light is observed an angle θ (measured with respect to the undiffracted beam). The diffraction grating formula is $d \sin \theta = m \lambda$. If the grating has 6000 lines per centimeter, what is the first-order (i.e. m = 1) diffraction angle for a hydrogen line with photon energy of 1.89 eV?

2. The spectrum of sodium (Na) has two closely-spaced lines at 589.00 nm and 589.59 nm. A typical diffraction grating spectrometer can accommodate a diffraction grating up to 3.0 cm wide. Determine the minimum number of lines per centimeter that is needed for the grating to just resolve these two sodium doublet lines in <u>second order</u>. Assume that the diffraction grating is fully illuminated. (Note: In your setup the grating will *not* be fully illuminated, so you won't have the ability to resolve the lines.)

3. In the Bohr model of the hydrogen atom, the electron can only occupy certain orbits. The allowed energy levels of the hydrogen atom are given by:

$$E_n = \frac{-13.6}{n^2} \ [\mathrm{eV}]$$

with n = 1 corresponding to the ground state, n = 2 the first excited state, and so on. The negative sign ensures that lower quantum numbers have lower energies, and that unbound (i.e. free) electrons have positive energies.

In these orbits the motion of the electron is stable. It does not lose energy by radiating electromagnetic waves. When the electron jumps from one orbit to another, it emits or absorbs a photon whose energy equals the difference in energy levels of the atom. The energy of the photon is given by:

$$E_\gamma = hf = hc/\lambda$$

where $h = 6.626 \times 10^{-34}$ J·s, $c = 2.998 \times 10^8$ m/s and $hc = 1240$ *eV-nm*.

Calculate the first **six** energy levels for the hydrogen atom. **Record your results here *and on page 8 of the lab writeup.***

$E_1 =$ _______________ [eV]

$E_2 =$ _______________ [eV]

$E_3 =$ _______________ [eV]

$E_4 =$ _______________ [eV]

$E_5 =$ _______________ [eV]

$E_6 =$ _______________ [eV]

Of course, observations of emitted photons do not correspond to electron energy levels directly but to *differences* in electron energy levels:

$$E_\gamma = \Delta E = E_{n_i} - E_{n_f} = -13.6\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)\ [eV]$$

For example, for a transition of an electron from the $n = 3$ (2nd excited state) to the $n = 1$ (ground state), the photon energy, E_γ = the electron energy difference, $\Delta E = E_3 - E_1 = 12.09$ eV. Check this value with your calculation(s) above.

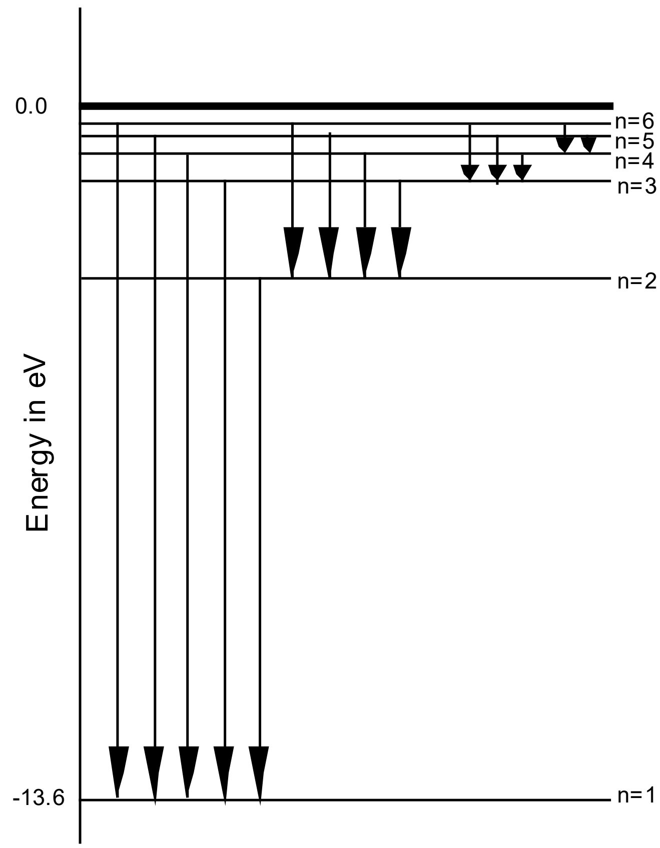
Thus, a photon of energy $E_\gamma = 12.09$ eV has a wavelength, λ of

$$\lambda = \frac{hc}{E_\gamma} = \frac{hc}{\Delta E} = \frac{(6.626\times10^{-34}\,\mathrm{J\cdot s})(2.998\times10^{8}\,\mathrm{m/s})}{(12.09\,\mathrm{eV})(1.602\times10^{-19}\,\mathrm{J/eV})} = \frac{1240eV-nm}{12.09eV} = 102.6\,\mathrm{nm}$$

Alternatively, dividing the above formula for E_γ by hc gives the following relation:

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

where $R = 1.097 \times 10^7$ m^{-1}, known as the Rydberg constant. Note that this formula gives the <u>reciprocal</u> of the wavelength, $1/\lambda$, in m^{-1}.



Hydrogen energy levels

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For hydrogen, calculate some of the energies and wavelengths of some of the transitions that correspond to the lines drawn in the figure above. For each one, indicate if it is in the ultraviolet (UV) or the infrared (IR) region of the electromagnetic spectrum; if it is in the visible region, record the color. (See Figure 7 (lab writeup, page 11) for a chart of the visible spectrum.) **Record your results here *and on page 10 of the lab writeup.***

Transitions to the $n = 1$ (ground state): Region of EM spectrum

$n_i = 2$, $n_f = 1$ $E_\gamma =$ ______ [eV] $\lambda =$ ______ [nm] ________

$n_i = 3$, $n_f = 1$ $E_\gamma =$ ______ [eV] $\lambda =$ ______ [nm] ________

$n_i = 4$, $n_f = 1$ $E_\gamma =$ ______ [eV] $\lambda =$ ______ [nm] ________

Transitions to the $n = 2$ (1st excited state):

$n_i = 3$, $n_f = 2$ $E_\gamma =$ ______ [eV] $\lambda =$ ______ [nm] ________

$n_i = 4$, $n_f = 2$ $E_\gamma =$ ______ [eV] $\lambda =$ ______ [nm] ________

$n_i = 5$, $n_f = 2$ $E_\gamma =$ ______ [eV] $\lambda =$ ______ [nm] ________

$n_i = 6$, $n_f = 2$ $E_\gamma =$ ______ [eV] $\lambda =$ ______ [nm] ________

Transitions to the $n = 3$ (2nd excited state):

$n_i = 4$, $n_f = 3$ $E_\gamma =$ ______ [eV] $\lambda =$ ______ [nm] ________

$n_i = 5$, $n_f = 3$ $E_\gamma =$ ______ [eV] $\lambda =$ ______ [nm] ________

$n_i = 6$, $n_f = 3$ $E_\gamma =$ ______ [eV] $\lambda =$ ______ [nm] ________

The hydrogen atom transitions in the UV light region are collectively known as the Lyman series; the transitions in the visible light region are collectively known as the Balmer series and the transitions in the IR light region are collectively known as the Paschen series, in honor of the physicists who first discovered them.