Announcements – 10/25/00

- **<u>Reminder:</u>** **Special Quiz on Friday**
- **<u>Reminder:</u>** **Special *Demo* on Friday**
- ***Belated* Reminder:**
  We missed Mole Day!

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Einstein's Explanation

- View EMR as a collection of *particles* (called **<u>photons</u>**), with each photon having the following energy:

$$E = h\nu$$

  - Each photon will cause an electron to be *ejected* **IF** the energy of the photon is above a minimum (threshold) value (called the "work function" of the metal)
  - Any energy of the photon *above* that needed to eject the electron would be transferred to the electron as *kinetic energy*
  - Increased EMR intensity translates to an increase in the number of photons (thereby increasing the number of electrons ejected)

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Photon Energies

- Calculate the *energy* of one photon of light at a wavelength of **700 nm**.

$$E = h\nu \quad \textit{and} \quad c = \lambda\nu$$

So: $\nu = c/\lambda$

Substituting: $\mathbf{E = hc/\lambda}$

$$E = \frac{(6.626 \times 10^{-34}\ \text{J-s})(2.9979 \times 10^{8}\ \text{m/s})}{(700 \times 10^{-9}\ \text{m})}$$

$$= \underline{\mathbf{2.84 \times 10^{-19}\ J}}$$

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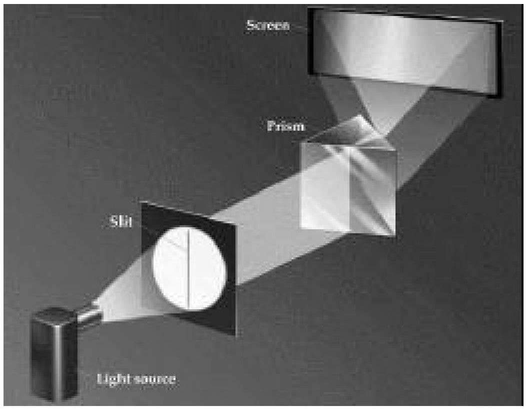
Separating Light into Colors

- To figure out what *wavelengths* of light are emitted from a light source, we need to obtain a **spectrum** of the light.

  This can be done with a spectrograph:



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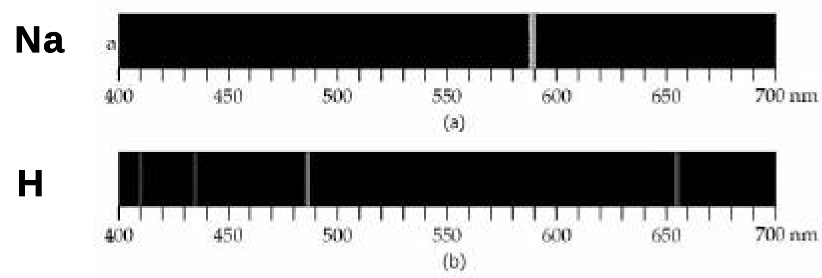
Line Spectra

- The light emitted by elements heated in a flame or electric discharge looks like this:



Line Spectra: *light emitted at* ***discrete*** wavelengths

Each element has a unique spectrum!

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Balmer and the Hydrogen Spectrum

- **1885**: Johann Balmer, a Swiss schoolteacher, *empirically* deduced a formula which predicted the wavelengths of emission for *Hydrogen*:

$$\lambda \text{ (in Å)} = 3645.6 \times \frac{n^2}{n^2 - 4}$$ for n = 3, 4, 5, 6

•Predicts the wavelengths of the 4 *visible* emission lines from Hydrogen (which are called the **Balmer Series**)

•Implies that there is some underlying *order* in the atom that results in this deceptively simple equation.

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The Bohr Atom

- **1913:** Niels Bohr uses quantum theory to **explain** the origin of the line spectrum of hydrogen

1. The electron in a hydrogen atom can exist only in **discrete orbits**
2. The **orbits** are circular paths about the nucleus at varying **radii**
3. Each **orbit** corresponds to a particular **energy**
4. **Orbit** energies *increase* with *increasing* radii
5. The *lowest* energy **orbit** is called the ground state
6. After *absorbing* energy, the e^- jumps to a *higher energy* **orbit** (an excited state)
7. When the e^- drops down to a *lower energy* **orbit**, the energy lost can be given off as a *quantum of light*
8. The ***energy*** of the photon emitted is equal to the difference in energies of the two orbits involved

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Mohr Bohr

- **Mathematically, Bohr equated the two forces acting on the orbiting electron**:

coulombic attraction = *centrifugal accelleration*

$$-(Z/4\pi\varepsilon_o)(e^2/r^2) = m(v^2/r)$$

- **Rearranging and making the *wild* assumption**:

$$mvr = n(h/2\pi)$$

  - e^- angular momentum can only have certain quantified values in whole multiples of $h/2\pi$

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Hydrogen Energy Levels

- Based on this model, Bohr arrived at a simple equation to calculate the electron *energy levels* in hydrogen:

$$\mathbf{E_n = -R_H(1/n^2)}$$ for n = 1, 2, 3, 4, ...

Where:

R_H = 2.179 x 10^{-18} Joules (*the Rydberg constant)*

n is the *Principal Quantum Number*

Radii can be calculated, too:

$$\mathbf{r_n = n^2 a_o}$$ (a_o = 0.529 Å)

Transitions Between Energy Levels

- Now, the energy change associated with a *transition* between electron energy levels can be quantified:

$$\Delta E = E_{final} - E_{initial} = h\nu$$

$$h\nu = \frac{-R_H}{n^2_f} - \frac{-R_H}{n^2_i}$$

Collecting terms:

$$\nu = (R_H/h)\ (1/n_i^2 - 1/n_f^2)$$

Bohr versus Balmer

- With some rearranging, the ***Balmer*** equation looks like this:

$$\nu = 3.29 \times 10^{15}\ s^{-1}\ (1/2^2\ -\ 1/n^2)$$

-This is the equation we just derived, but with n_f fixed at a value of 2

-So, the Bohr model also accurately predicts the frequencies of the *Balmer Series* emission lines

-BUT, it also predicts *other* emission lines (for n_f = 1, 3, 4, etc.)

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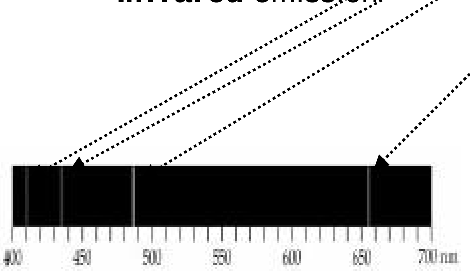
Hydrogen's Energy Level Diagram

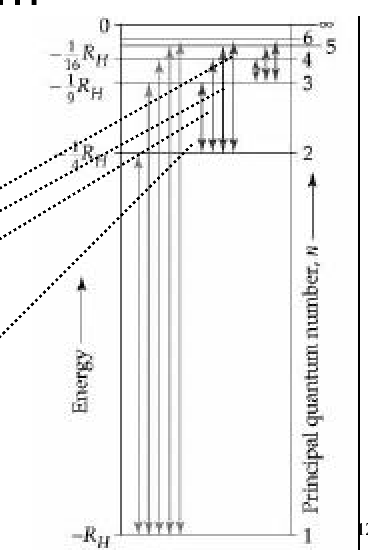
When n_f = 2: *Balmer Series*

-**visible** emission

When n_f = 3: *Paschen Series*

-**infrared** emission





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Sample Calculation

- Calculate the wavelength at which the *least energetic* emission spectral line of the **Lyman Series** (n_f = 1) is observed.

Lowest energy transition will be 2→1:

$$\Delta \mathbf{E} = (R_H)\ (1/2^2 - 1/1^2)$$

$$\Delta \mathbf{E} = (2.179 \times 10^{-18}\ \mathrm{J})(1/4 - 1)$$

$$\Delta E = -1.63425 \times 10^{-18}\ J \quad \text{(energy lost by atom)}$$

Converting to wavelength:

$$\lambda = hc/\Delta E$$

$$= (6.626 \times 10^{-34}\ \mathrm{J\text{-}s})(2.9979 \times 10^{8}\ \mathrm{m/s})/(1.63425 \times 10^{-18}\ \mathrm{J})$$

$$= 1.215486 \times 10^{-7}\ \mathrm{m} = 121.549\ \mathrm{nm} \rightarrow \mathbf{121.5\ nm}\ \text{(vac UV)}$$

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