CHEM 205 section 03

LECTURE #17 Thurs., March 6, 2008

LECTURE TOPICS:

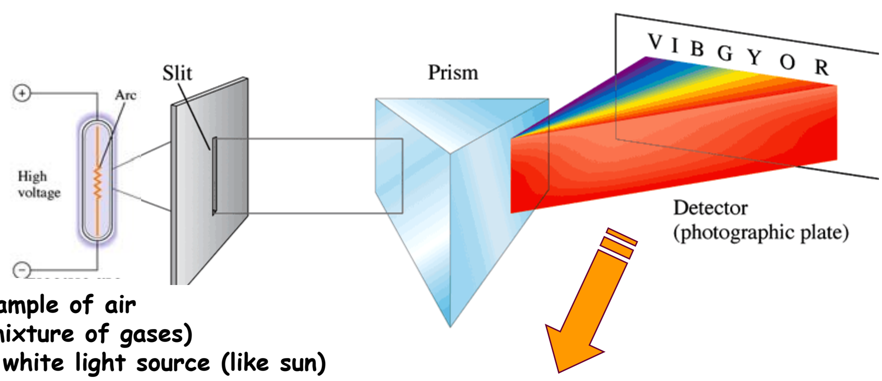
TODAY'S CLASS: continue Ch.7

NEXT CLASS: **finish Ch.7**

(1)

7.3 Atomic line spectra & Niels Bohr…

WHITE LIGHT: combination of ALL visible wavelengths
if separate wavelengths using a prism: "CONTINUOUS SPECTRUM"



Sample of air (mixture of gases) = white light source (like sun)

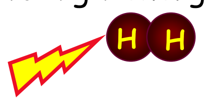
WHY? **Light enters different medium ⇨ rays are refracted**
→ **shorter λ light bends more** (for more detail, take physics)

Zumdahl's Figure 7.4
See K&T Figure 7.7 & so on…

(2)

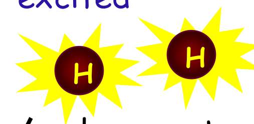
Atomic spectrum of <u>PURE</u> H_2: a LINE spectrum…

Expose hydrogen (H_2)
to high-energy spark



H_2 molecules absorb energy $\Rightarrow$ breaks H-H bonds

Resulting H atoms are "excited"



release extra energy as PHOTONS

"atomic emission spectrum"
= light of **specific** energies (λ's)

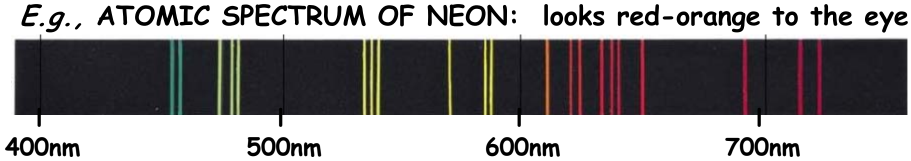
+ other, non-visible lines:
some in UV region,
some in IR region…

Application = glowing electric "neon" signs

- for a molecular substance like elemental hydrogen, relaxed atoms at end reform H_2 or react with other stuff (= loss)
- for noble gases: unreactive free atoms ⇨ no losses ⇨ useful…

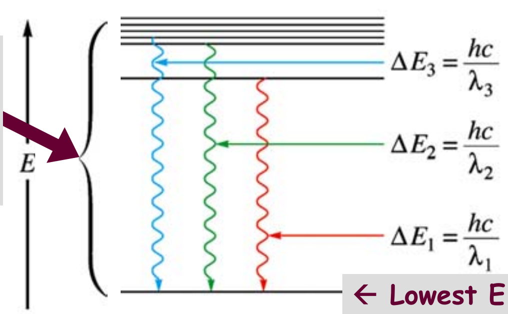
Why are atomic spectra LINE SPECTRA ?

***E.g.*, ATOMIC SPECTRUM OF NEON: looks red-orange to the eye**



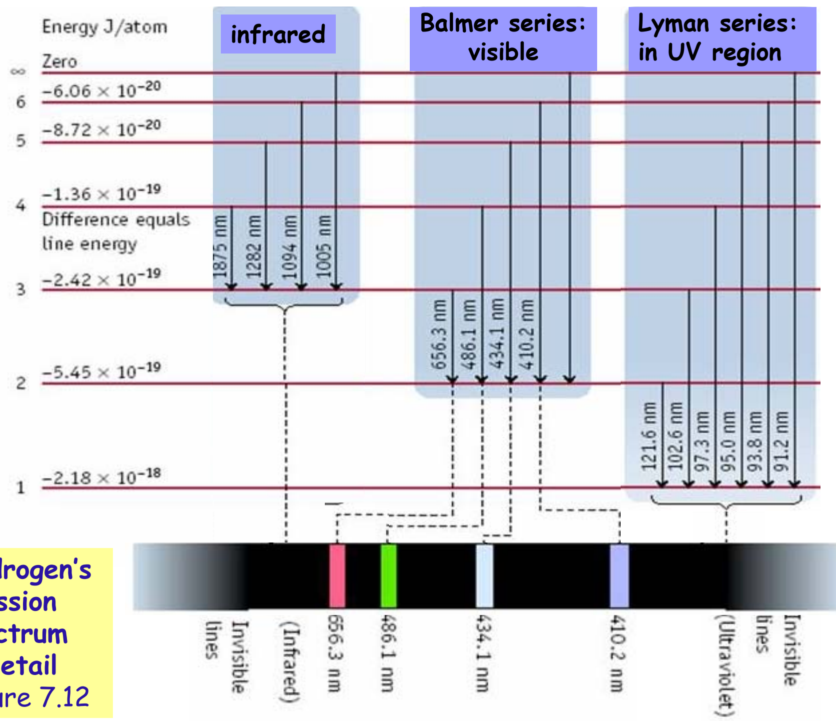
When an atom decreases its energy level by a discrete increment ΔE ⇨ a PHOTON is emitted with $E = h\nu = \Delta E$

Discrete allowed E levels for hydrogen atom (closer & closer together as E ↑)



← Lowest E state = "ground state"

Zumdahl's Figure 7.7



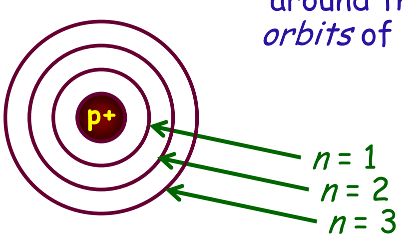
Hydrogen's emission spectrum in detail
Figure 7.12

The Bohr Model of the H atom (incorrect)

…1st recognition that quantization applies to atomic structure!!

Niels Bohr
(1885- 1962)
Danish physicist

BOHR'S MODEL: The e^- in an H atom can move around the nucleus in *circular orbits* of ALLOWED energies



"n" is now known as the *principal quantum number* for the e^- ⇨ describes e^-'s energy

$$E = -\left(\frac{Rhc}{n^2}\right)$$

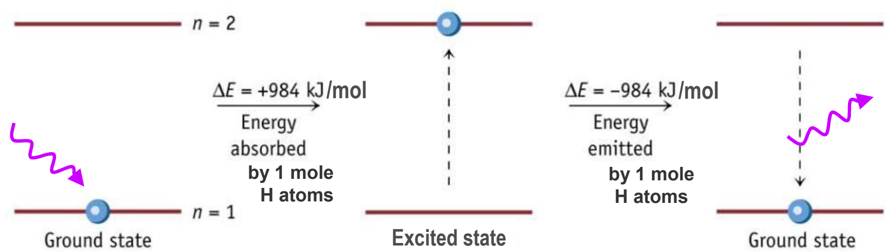
- E = energy of specific energy level, n
- n = an integer (*name* of energy level…)
- R = Rydberg constant = $1.0974 \times 10^7\ m^{-1}$
- h = Planck's constant
- c = speed of light in vacuum

Energy must be ABSORBED (to move e^- to higher E state) or EMITTED (to move e^- to lower E state)

Hydrogen's e^- transition energies can be calculated

$$\Delta E = -Rhc\left(\frac{1}{n^2_{final}} - \frac{1}{n^2_{initial}}\right)$$

Rhc: per mole: 1312 kJ/mol
per electron: 2.181×10^{-18} J



E.g., Exciting an electron: in one H atom from the n=2 to n=4 state:

$\Delta E = -(2.181\times10^{-18}\text{ J}) \times [(1/4^2) - (1/2^2)]$
$= +4.089\times10^{-19}$ J ⇨ **absorb a 486nm photon (bluish-green…)**

c.f., Ionizing an atom: eject e^- from atom (from n=1 to "n=∞" state):

$\Delta E = -(2.181\times10^{-18}\text{ J}) \times [(1/\infty^2) - (1/1^2)]$
$= +2.181\times10^{-18}$ J **= ionization energy (91nm photon, deep UV)**

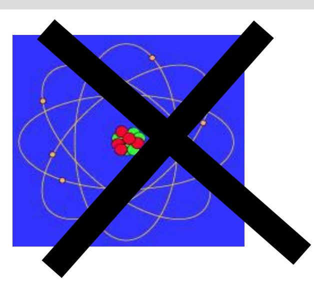
Bohr's model: good and bad…

- Could explain experimental observations for H & He^+:
  - Line spectrum
  - Ionization energy
- & predict wavelengths of unseen lines in spectra (correctly)
- INTRODUCED CONCEPT OF **ENERGY QUANTIZATION** IN DESCRIBING ATOMIC STRUCTURE !

DOWNFALL:

- Model does not work for species with >1 electron

⇨ **Idea of planet-like orbits is incorrect...**



Graphic adapted from http://csep10.phys.utk.eduastr162lectlightbohr.html

ASSIGNED READINGS

- **BEFORE NEXT CLASS:**

  Read Ch.7 up to section 7.3