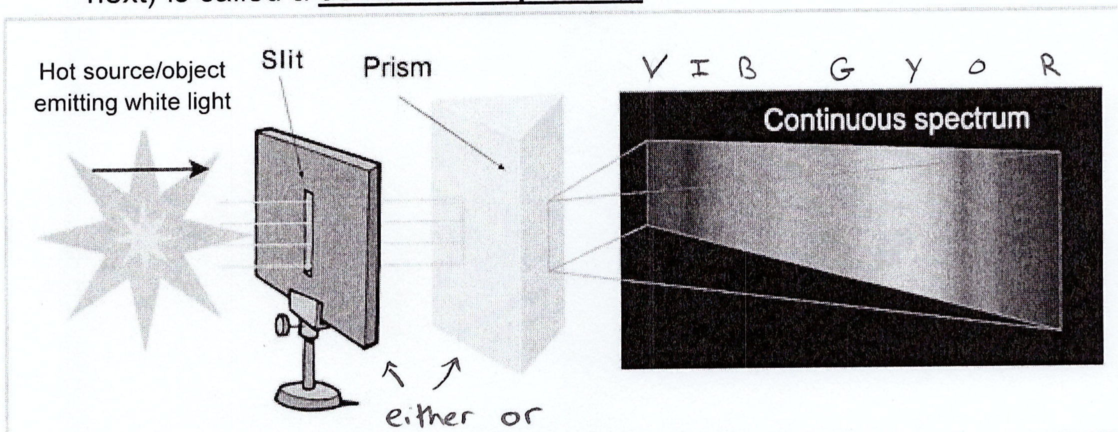
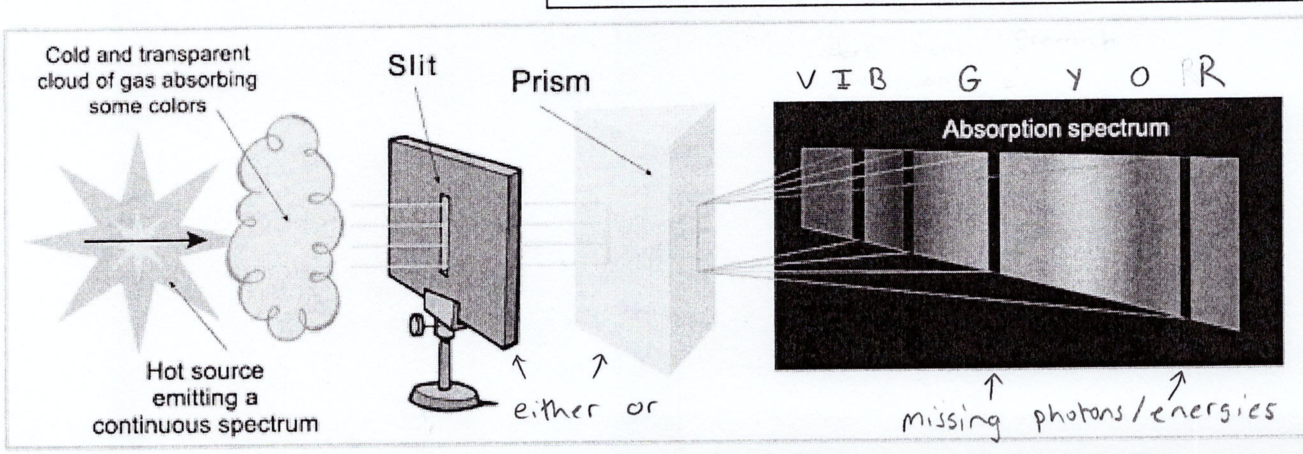
Atomic Spectra and the Quantized Model of the Atom

- Neils Bohr study the *spectra* that was produced from a hydrogen atom, which help him develop his theory of the atom to overcome Rutherford's shortcomings
- From previous scientists, Bohr knew that if you shone white light through a prism or a diffraction grating, light would be *dispersed* (separating white light into its individual wavelengths)
    - The rainbow pattern that is produced (where one color continues into the next) is called a **continuous spectrum**



- Bohr decided to pass white light from a glowing object through a *unexcited gas* (gas that is cooled and has very little energy) first and then through a prism or grating
    - Bohr observed that the same continuous spectrum was produced, but it had dark lines appearing in the spectrum. Those dark bands meant that those particular photons with specific frequencies/wavelengths/energies were missing from the spectrum.

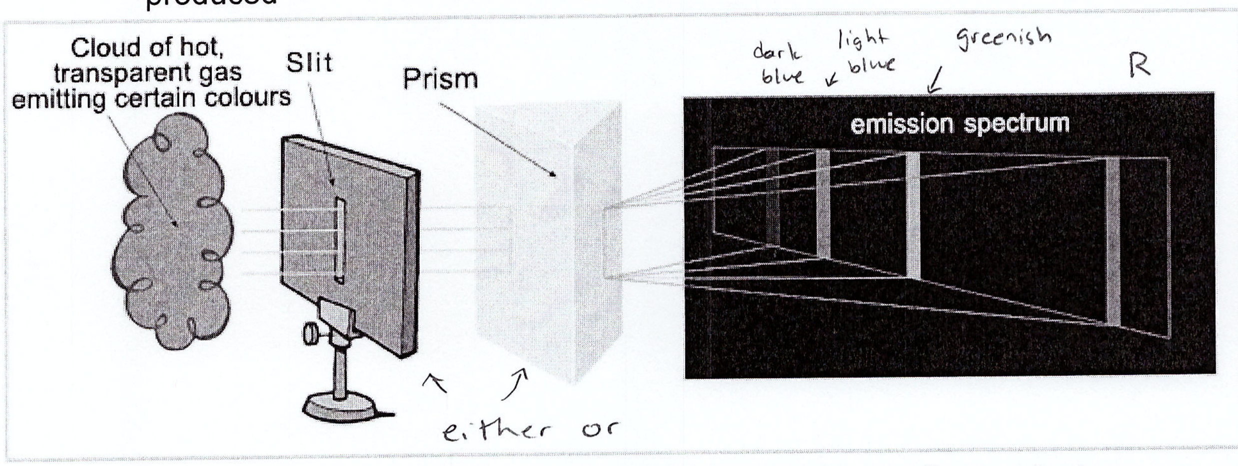
> Recall that the energy, wavelength, and frequency of a photon are all related as explained by the following equation.
>
> $$E = hf = \frac{hc}{\lambda}$$



- This continuous, rainbow pattern with dark bands was termed a **dark line spectrum** or a **line-absorption spectrum**

- ⁎ Bohr concluded that the unexcited gas must only be able to absorb certain photons and not the entire spectrum. This supports Planck's hypothesis that matter is *quantized*.

- Bohr modified the experiment slightly and passed light given from a *glowing/excited gas* (gas that is energized by some form such as heat or electrical energy) through a prism or grating
    - This time, Bohr observed that no continuous spectrum was produced, but only thin bands of certain colors/(frequencies/wavelengths/energies) were produced

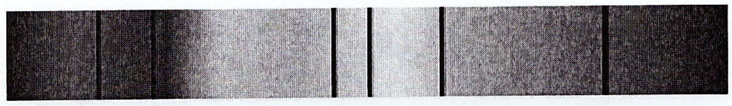


    - This spectrum was called a **bright line spectrum** or a **line emission spectrum**
    - ⁎ Bohr concluded that the excited gas must only be able to emit certain photons, just as a cooled gas can only absorb certain photons.

- Both atomic spectra (line absorption and line emission) help Bohr explain that atoms of a particular element can only absorb and emit photons with certain frequencies/wavelengths/energies.
    - If an element can only absorb certain photons, it makes sense that the same element can only emit those same photons

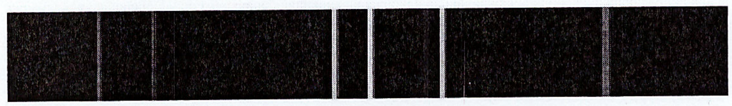
- Bohr also discovered that every element has its own unique atomic spectra

Aluminum Absorption Spectrum



different from hydrogen atom spectrum

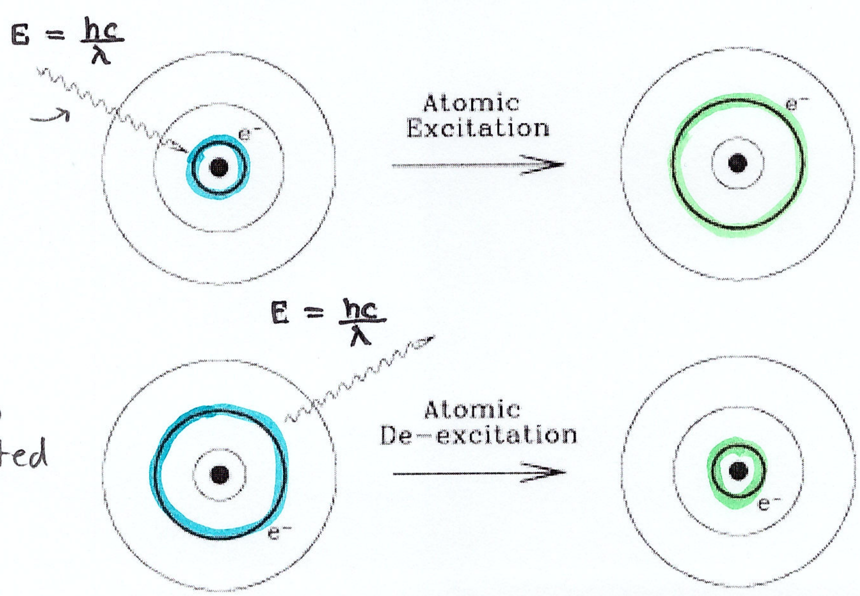
Aluminum Emission Spectrum



- Bohr needed to revise the model of the atom to explain why matter is quantized and can only absorb or emit certain photons
- Bohr proposed that *within the atom*, there are certain *allowed energy orbits* around the nucleus, in which the *electrons can move and orbit* without giving off energy.
    - This meant that the energy of the electron in an atom is *quantized* (just as Planck hypothesized)
    - * For the electron to occupy any one of the allowed energy orbits, it must possess the energy allowed for that orbit; electrons cannot be found between orbits
    - * When an atom *absorbs a photon*, the electron in the atom will move *up* to a higher energy orbit. An atom absorbing a photon corresponds to a dark-line absorption spectrum.
    - * When an atom *emits a photon*, the electron in the atom will move *down* to a lower energy orbit. An atom emitting a photon corresponds to a bright-line emission spectrum.

photon absorbed

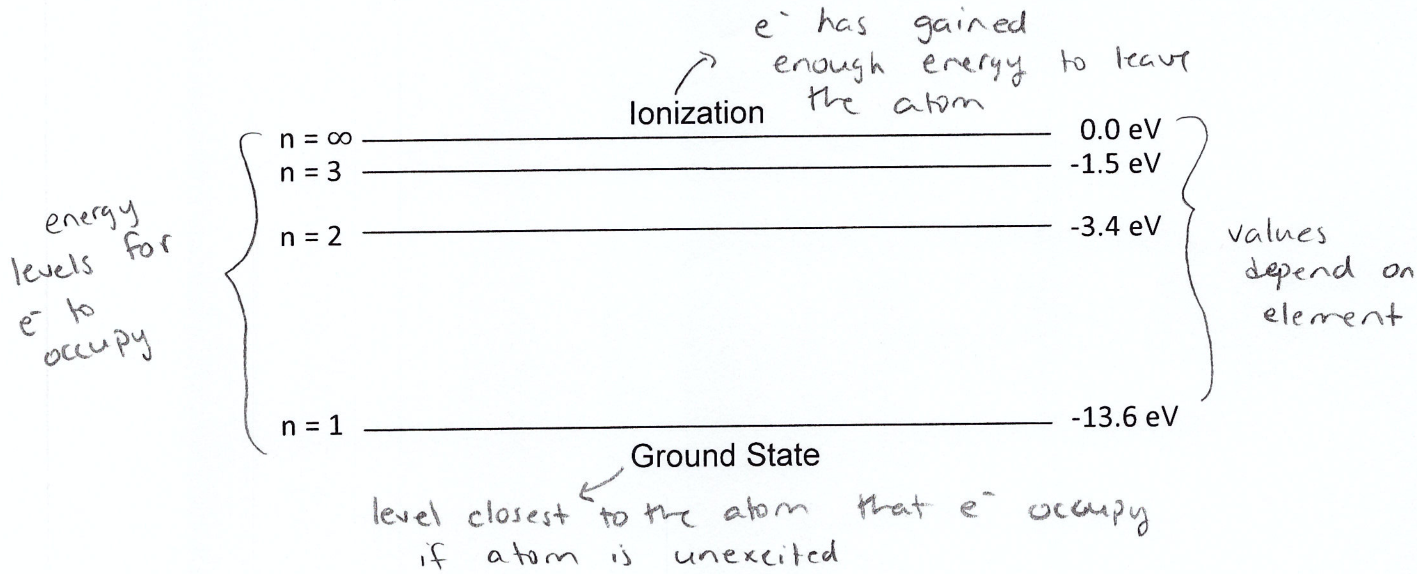
photon emitted



absorption dark-line spectrum

emission bright-line spectrum

- Bohr developed ***energy level diagrams*** to help explain his model of the atom. The energy levels in Bohr's model are not equally spaced and will vary for each element.



- Calculations using Borh's model are based off the conservation of energy (principle # 5)

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

where $\Delta E = E_{photon} = \frac{hc}{\lambda} = hf$

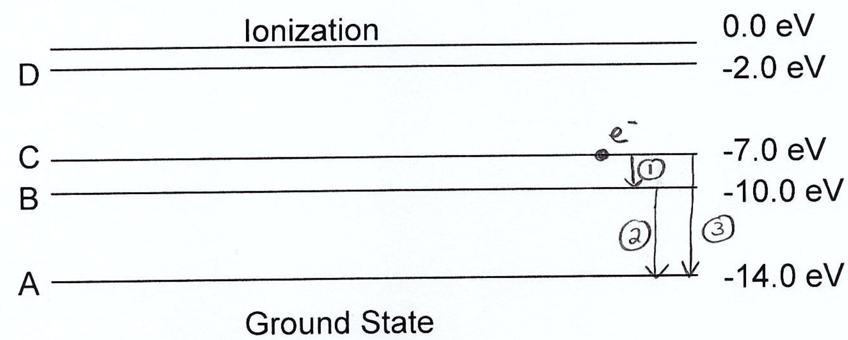
* if ΔE is "−", energy is released ∴ photon is emitted

if ΔE is "+", energy is absorbed ∴ photon is absorbed

EXAMPLES

Use the following information to answer the next 5 questions.

The energies of an electron within an atom are most often given relative to ionization as shown below.



1. Determine the energy of a photon absorbed/released as an electron makes a transition from D to B. Is the photon absorbed or released?

↳ e^- drops down lower

$$\Delta E = E_{photon} = E_B - E_D = (-10.0\,eV) - (-2.0\,eV) = -8.0\,eV$$

↑ photon released

$$\boxed{E_{photon} = 8.0\,eV}$$

2. Determine the number of photons of different wavelengths that could possibly be produced due to an electron at level C moving to ground state.

3 photons with different wavelengths

(see diagram above)

3. Determine the frequency of the photon absorbed/released when the electron in the atom moves from C to A. Is the photon absorbed or released?

↳ e^- drops down lower

$E = hf$ ②

↙

$\Delta E = E_{photon}$ ①

① $\Delta E = E_A - E_C$

$$\Delta E = (-14.0\,eV) - (-7.0\,eV) = -7.0\,eV$$

↑ photon released

② $E = hf \Rightarrow f = \frac{E}{h}$

$$f = \frac{7.0\,eV}{4.14 \times 10^{-15}\,eV \cdot s} = 1.6908... \times 10^{15}\,Hz$$

$$\boxed{f = 1.7 \times 10^{15}\,Hz}$$

4. Determine the wavelength of a photon required to ionize an electron that was initially in ground state. Is the photon absorbed or released?

↳ e^- moves up higher

② $E = \frac{hc}{\lambda}$

① $E_{photon} = \Delta E$

① $\Delta E = E_{ionization} - E_{ground}$

$\Delta E = (0.0eV) - (+14.0eV) = +14.0eV$

↑ photon absorbed

② $E = \frac{hc}{\lambda} \Rightarrow \lambda = \frac{hc}{E} = \frac{(4.14\times10^{-15}eV\cdot s)(3.0\times10^{8}m/s)}{14.0eV}$

$\lambda = 8.87142\ldots\times10^{-8}\ m$

$\boxed{\lambda = 8.9\times10^{-8}m}$

5. An electron at ground state in the atoms is struck by a passing electron having an energy of 9.0eV and the atom's electron absorbs as much of the energy as possible. Determine the kinetic energy of the passing electron after it collides with the electron in the atom.

$E_{in} = E_{out}$ ②

(conservation of energy principle #5)

$E_{k,e^-\,initial} = E_{k,e^-\,final} + E_{absorbed}$

↳ can only move up to an energy level that has a gain in energy of 9.0eV or less. ①

① $\Delta E = E_C - E_A = (-7.0eV) - (-14.0eV) = +7.0eV$

↑ absorbed

② $E_{k,e^-\,initial} = E_{k,e^-\,final} + E_{absorbed}$

$E_{k,e^-\,final} = E_{k,e^-\,initial} - E_{absorbed} = 9.0eV - 7.0eV$

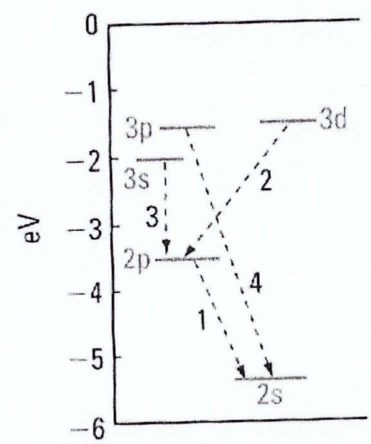
$\boxed{E_{k,e^-\,final} = 2.0eV}$

*****Now try pg. 292 #1, 2-8, 10, & Practice Problems*****

Practice Problems

1. A helium-neon laser produces photons of wavelength 633nm when an electron in a neon atom drops from an excited energy state to a lower state. What is the energy difference between these two states? Express your answer in eV.
[1.96 eV]

2. The diagram shows some of the energy levels for the lithium atom. The designations 2s, 2p, etc. are a common notation for spectroscopy.



   a. Without doing any calculations, sort the four transitions shown from the shortest wavelength to the longest wavelength. **[4, 2, 1, 3]**

   b. Calculate the wavelengths produced in these transitions. Indicate which ones are in the visible part of the spectrum, along with their colours. **[λ_1=654nm, visible (red); λ_2=592nm, visible (yellow); λ_3=829nm; λ_4=319nm]**