Topic 7: Atomic, nuclear and particle physics

7.1 – Discrete energy and radioactivity

Understandings:

- Discrete energy and discrete energy levels
- Transitions between energy levels

Applications and skills:

- Describing the emission and absorption spectrum of common gases
- Solving problems involving atomic spectra, including calculating the wavelength of photons emitted during atomic transitions

Universal Wave Equation

$$v = f \cdot \lambda$$

Speed of light

$c = 3 \times 10^8$ m/s

In nuclear physics, the joule is not a very convenient unit of measure to use for energy. Instead, a much smaller unit, called the mega-electron volt (MeV), is more useful. An electron-volt (eV) is defined as the amount of energy given to an electron when it is accelerated through a potential difference of 1 V. One electron-volt is equal to 1.602×10^{-19} J. The **mega-electron volt** is one million times this value:

$1\ \text{MeV} = 1.602 \times 10^{-13}\ \text{J}$

$$V = \frac{E}{q} \qquad E = (1)(1.6x10^{-19})$$

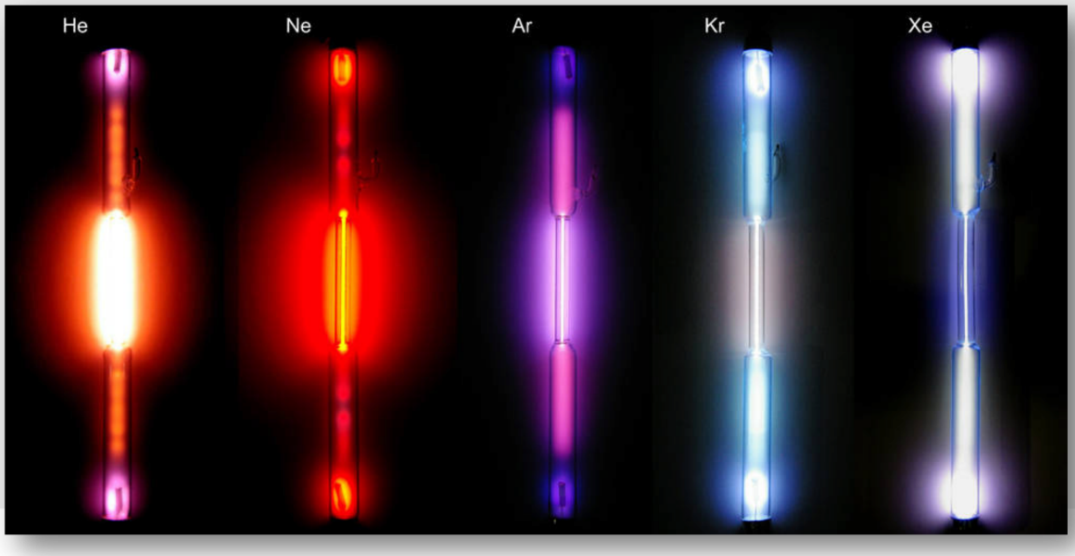
$$E = Vq \qquad E = 1.6x10^{-19} J$$

$$1\ \text{MeV} = 1.602 \times 10^{-13}\ \text{J}$$

A "volt" is a unit used to describe electrical "pressure" (like water).

Describing the emission and absorption spectrum of common gases

- When a gas in a tube is subjected to a voltage, the gas ionizes, and emits light.

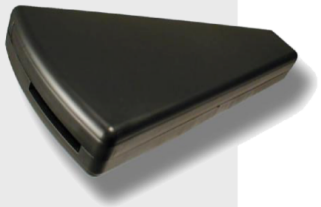


Link - Phet demo

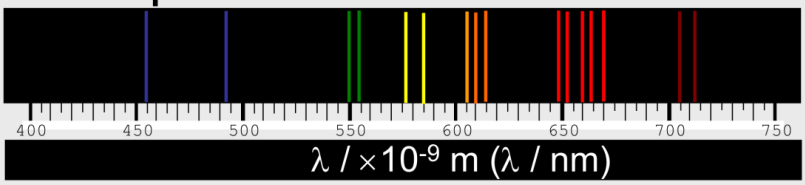
Part 1 - demo - electron movement and voltage

Part 2 - absorption and emission (later)

Describing the emission and absorption spectrum of common gases



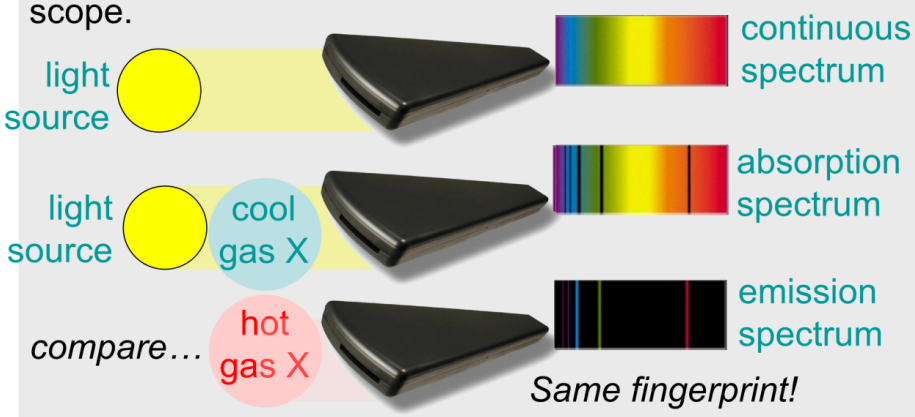
- We can analyze that light by looking at it through a **spectroscope**.
- A spectroscope acts similar to a prism, in that it separates the incident light into its constituent wavelengths.
- For example, heated barium gas will produce an **emission spectrum** that looks like this:



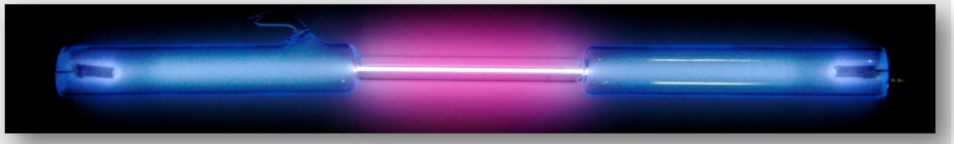
- An emission spectrum is an elemental fingerprint.

Describing the emission and absorption spectrum of common gases

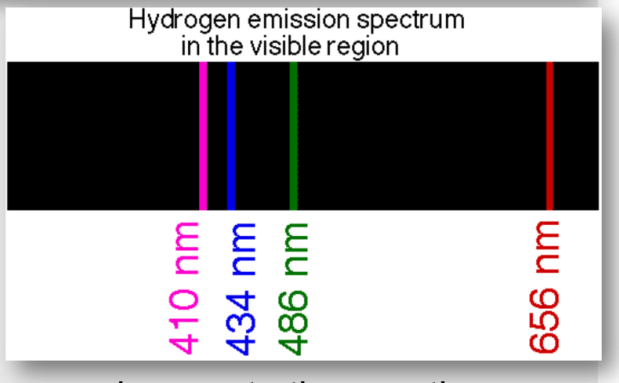
- Each element also has an **absorption spectrum**, caused by cool gases between a source of light and the scope.



Transitions between energy levels



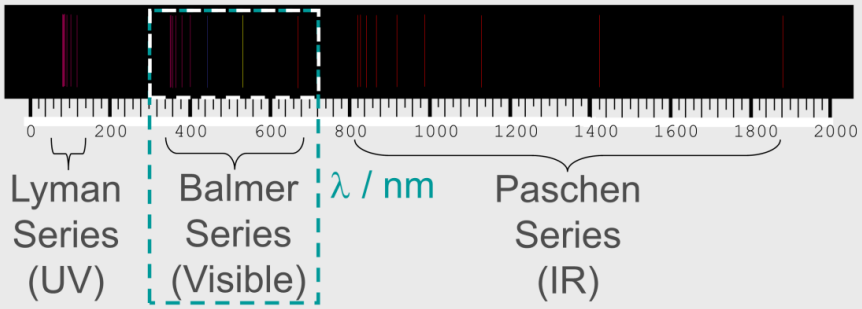
- In the late 1800s a Swedish physicist by the name of J.J. Balmer observed the spectrum of hydrogen – the simplest of all the elements:



- His observations gave us clues as to the way the negative charges were distributed about the nucleus.

Transitions between energy levels

- In reality, there are many additional natural groupings for the hydrogen spectrum, three of which are shown here:

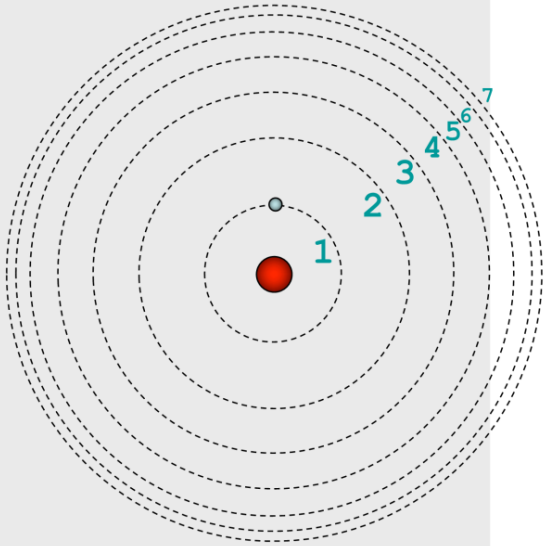


- These groupings led scientists to imagine that the hydrogen's single electron could occupy many different **energy levels**, as shown in the next slide:

Big Idea

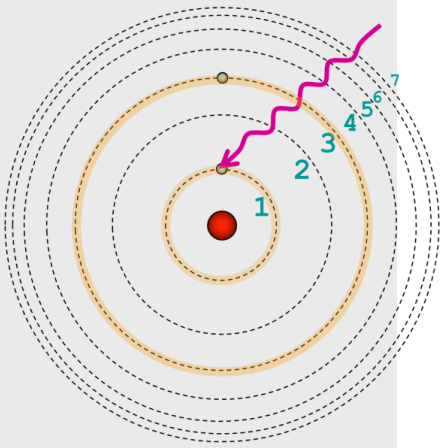
Transitions between energy levels

- The first 7 energy levels for hydrogen are shown here:
- The energy levels are labeled from the lowest to the highest as $n = 1$ to $n = 7$ in the picture.
- n is called the **principal quantum number** and goes all the way up to infinity (∞)!
- In its **ground state** or **unexcited state**, hydrogen's single electron is in the 1st energy level ($n = 1$):



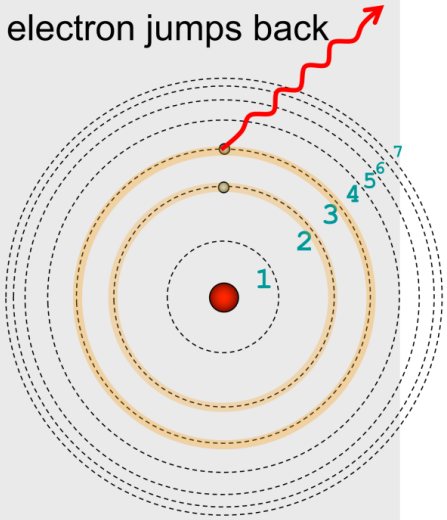
Transitions between energy levels

- As we will see later, light energy is carried by a particle called a **photon**.
- If a photon of <u>just the right energy</u> strikes a hydrogen atom, it is **absorbed** by the atom and stored by virtue of the electron jumping to a new energy level:
- The electron jumped from the $n = 1$ state to the $n = 3$ state.
- We say the atom is **excited**.



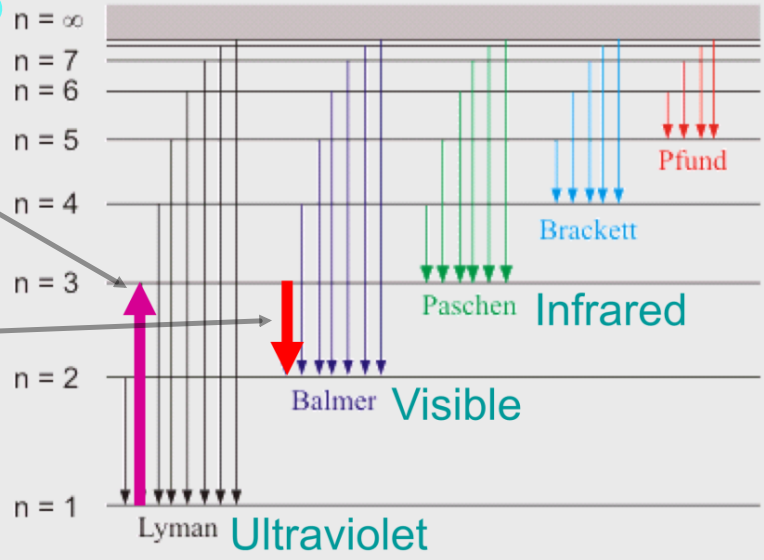
Transitions between energy levels

- When the atom **de-excites** the electron jumps back down to a lower energy level.
- When it does, it emits a photon of <u>just the right energy</u> to account for the atom's energy loss during the electron's orbital drop.
- The electron jumped from the $n = 3$ state to the $n = 2$ state.
- We say the atom is **de-excited**, but not quite in its ground state.



Transitions between energy levels

- The graphic shown here accounts for many of the observed hydrogen emission spectra.
- The excitation illustrated looked like this:
- The de-excitation looked like this:



Note:

Lyman 7 to 1

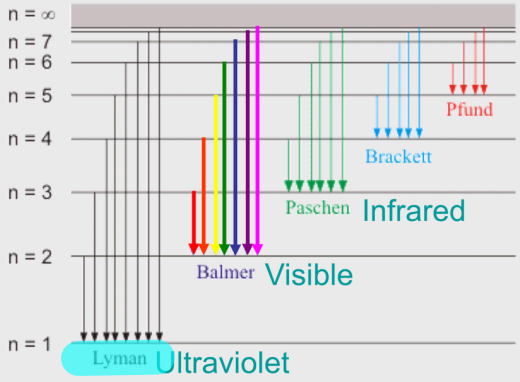
(only series that ends at 1)

Balmer 7 to 2

(only series that ends at 2)

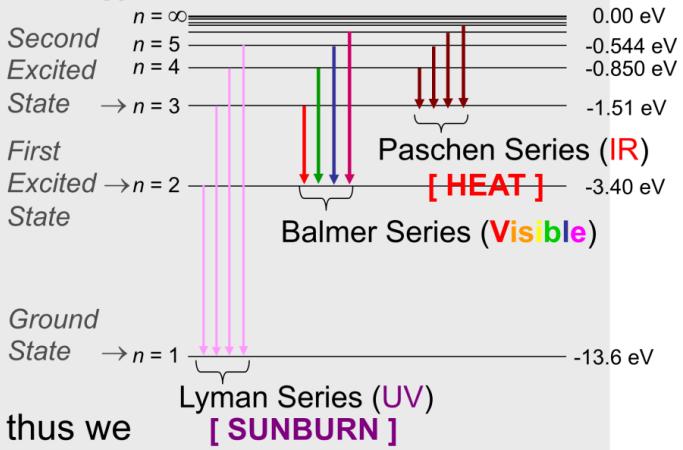
Transitions between energy levels

- The human eye is only sensitive to the Balmer series of photon energies (or wavelengths):



Transitions between energy levels

- The previous energy level diagram was NOT to scale. This one is. Note that none of the energy drops of the other series overlap those of the Balmer series, and thus we cannot see any of them.
- But we can still sense them!

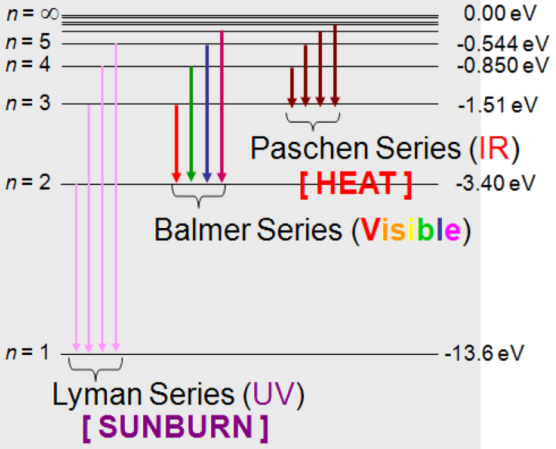


FYI • Transition energy is measured in eV because of the tiny amounts involved.

Note: negative energy, attractive force (like gravity), as n increases, electric energy increases, as n decreases, energy is released.

Transitions between energy levels

- Because of **wave-particle duality**, we have discovered that light not only acts like a wave, having a wavelength λ and a frequency *f*, but it acts like a particle (called a **photon**) having an energy *E* given by



$E = hf$ Where $h = 6.63 \times 10^{-34}$ Js and is called Planck's constant.

energy *E* of a photon having frequency *f*

Transitions between energy levels

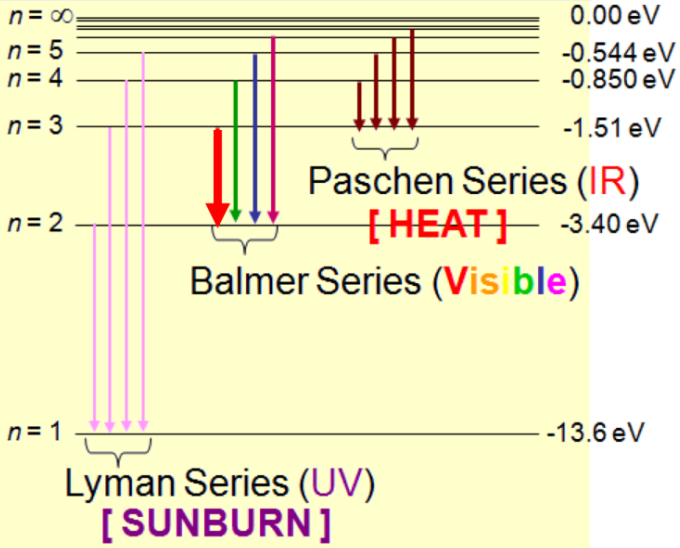
EXAMPLE:

An electron jumps from energy level $n = 3$ to energy level $n = 2$ in the hydrogen atom.

(a) What series is this de-excitation in?

SOLUTION:

- Find it on the diagram:
- This jump is contained in the Balmer Series, and produces a visible photon.



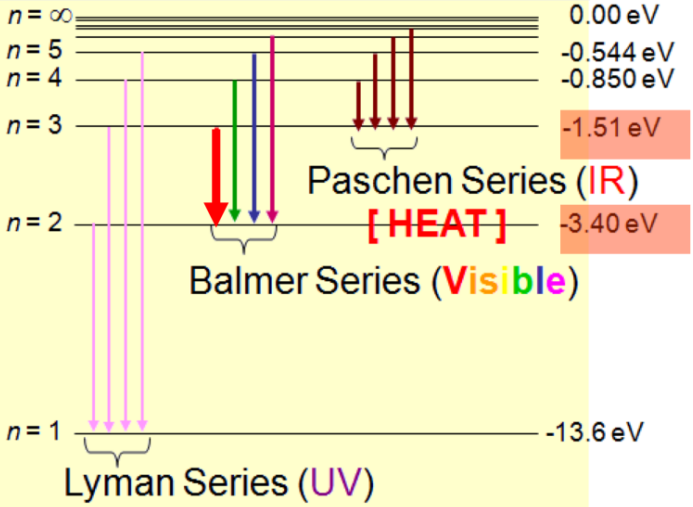
Transitions between energy levels

EXAMPLE:

An electron jumps from energy level $n = 3$ to energy level $n = 2$ in the hydrogen atom.

(b) Find the atom's change in energy in eV and in J.

SOLUTION:



- $\Delta E = E_f - E_0$
  $= -3.40 - (-1.51) = -1.89$ eV.
- $\Delta E = (-1.89 \text{ eV})(1.60 \times 10^{-19} \text{ J/eV})$
  $= -3.02 \times 10^{-19}$ J.

Energy is released (electric energy decreases, hence negative value)

An increase in electric energy would result in a delta E being positive

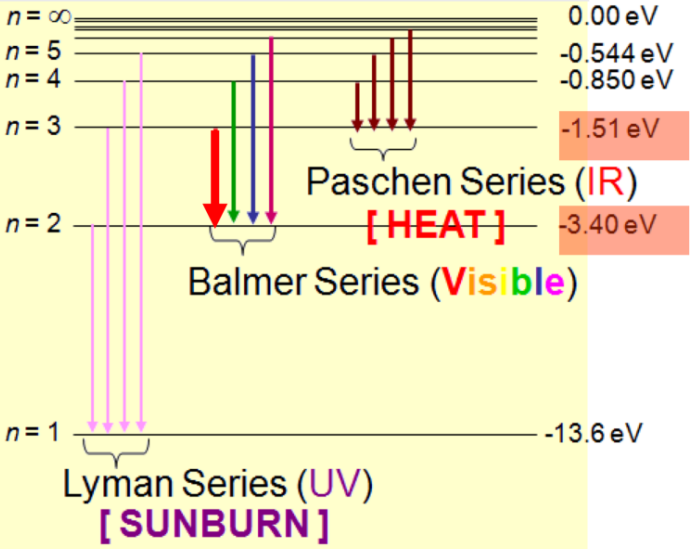
Transitions between energy levels

EXAMPLE:

An electron jumps from energy level $n = 3$ to energy level $n = 2$ in the hydrogen atom.

(c) Find the energy (in J) of the emitted photon.

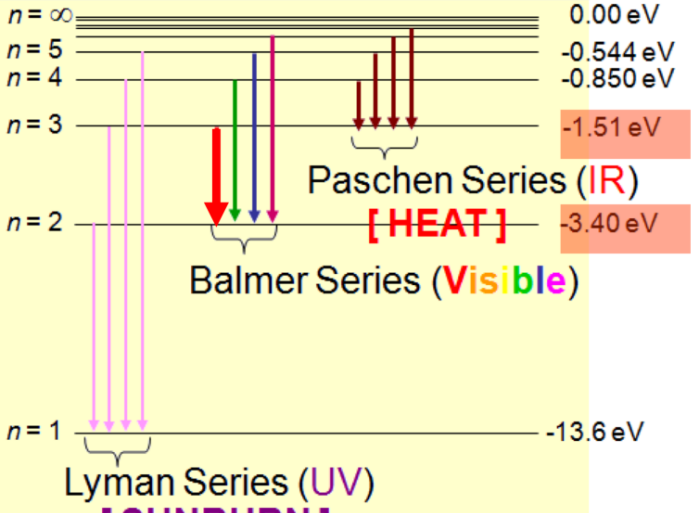
SOLUTION:



- The hydrogen atom lost 3.02×10^{-19} J of energy.
- From conservation of energy a photon was created having $E = 3.02 \times 10^{-19}$ J.

Transitions between energy levels

EXAMPLE:
An electron jumps from energy level *n* = 3 to energy level *n* = 2 in the hydrogen atom.
(d) Find the frequency of the emitted photon.



SOLUTION:

- From $E = hf$ we have

$3.02\times10^{-19} = (6.63\times10^{-34})f$, or

$$f = \frac{3.02\times10^{-19}}{6.63\times10^{-34}}$$

$$f = 4.56\times1014 \text{ Hz.}$$

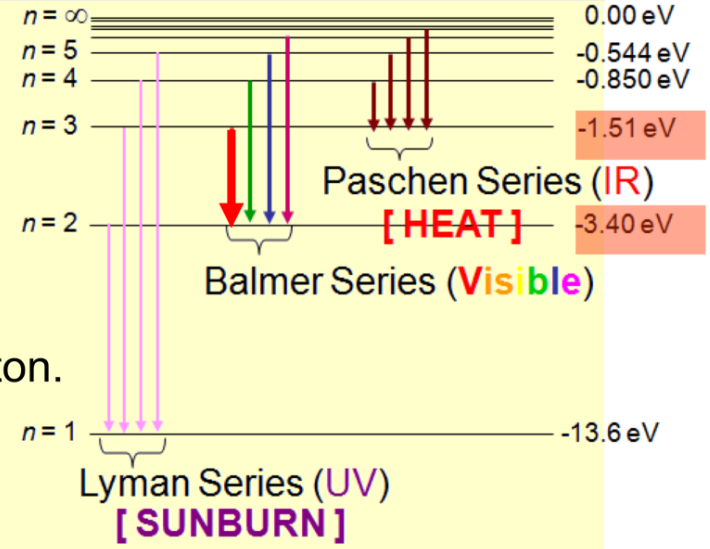
FYI **When finding *f*, be sure *E* is in Joules, not eV.**

Transitions between energy levels

EXAMPLE:

An electron jumps from energy level $n = 3$ to energy level $n = 2$ in the hydrogen atom.

(e) Find the wavelength (in nm) of the emitted photon.



SOLUTION:

- From $v = \lambda f$ where $v = c$ we have $3.00\times10^{8} = \lambda(4.56\times10^{14})$, or $\lambda = 6.58\times10^{-7}$ m.
- Then $\lambda = 6.58\times10^{-7}$ m

$= 658\times10^{-9}$ m $= 658$ nm.

Practice IB MC problems.

Discrete energy and discrete energy levels

- **Discrete** means discontinuous, or separated.

PRACTICE: Which one of the following provides direct evidence for the existence of discrete energy levels in an atom?

A. The continuous spectrum of the light emitted by a white hot metal.

B. The line emission spectrum of a gas at low pressure.

C. The emission of gamma radiation from radioactive atoms.

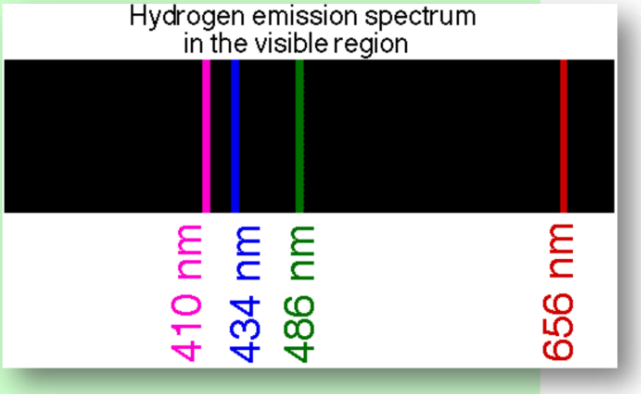
D. The ionization of gas atoms when bombarded by alpha particles.

SOLUTION: B

- Just pay attention!

Solving problems involving atomic spectra

PRACTICE: A spectroscopic examination of glowing hydrogen shows the presence of a 434 nm blue emission line.
(a) What is its frequency?



SOLUTION:
Use $c = \lambda f$
where $c = 3.00\times10^{8}$ m s^{-1} and $\lambda = 434\times10^{-9}$ m:

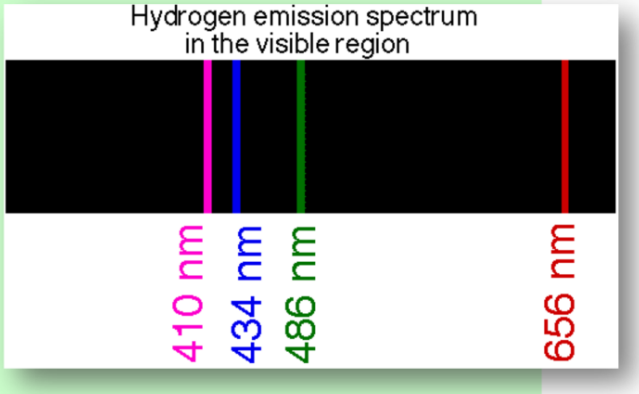
$$3.00\times10^{8} = (434\times10^{-9})f$$

$$f = \frac{3.00\times10^{8}}{434\times10^{-9}}$$

$$= 6.91\times10^{14} \text{ Hz.}$$

Solving problems involving atomic spectra

PRACTICE: A spectroscopic examination of glowing hydrogen shows the presence of a 434 nm blue emission line.
(b) What is the energy (in J and eV) of each of its blue-light photons?



SOLUTION: Use $E = hf$:

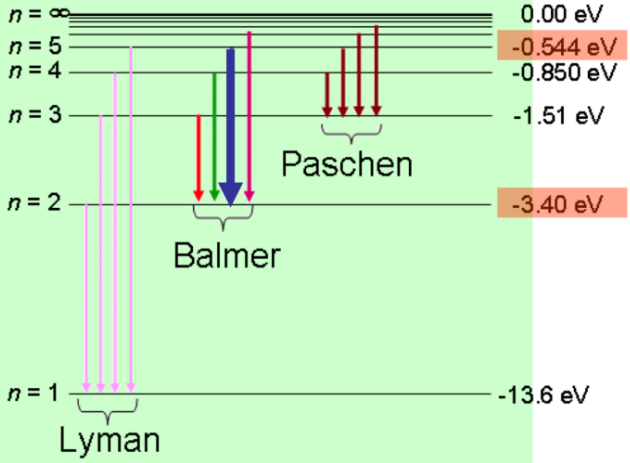
$$E = (6.63\times10^{-34})(6.91\times10^{14})$$

$$E = 4.58\times10^{-19}\text{J}.$$

$$E = (4.58\times10^{-19}J)\left(\frac{1eV}{1.60\times10^{-19}J}\right) = 2.86\text{ eV}.$$

Solving problems involving atomic spectra

PRACTICE: A spectroscopic examination of glowing hydrogen shows the presence of a 434 nm blue emission line.
(c) What are the energy levels associated with this photon?



SOLUTION:

- Because it is visible use the Balmer Series with ΔE = -2.86 eV.
- Note that $E_2 - E_5$ = -3.40 – -0.544 = -2.86 eV.
- Thus the electron jumped from n = 5 to n = 2.

Solving problems involving atomic spectra

PRACTICE: The element helium was first identified by the *absorption spectrum* of the sun.
(a) Explain what is meant by the term *absorption spectrum*.

continuous spectrum

SOLUTION:

absorption spectrum

emission spectrum

- An absorption spectrum is produced when a cool gas is between a source having a continuous spectrum and an observer with a spectroscope.
- The cool gases absorb their signature wavelengths from the continuous spectrum.
- Where the wavelengths have been absorbed by the gas there will be black lines.

Solving problems involving atomic spectra

PRACTICE:

One of the wavelengths of the absorption spectrum for helium occurs at 588 nm.

(b) Show that the energy of a photon having a wavelength of 588 nm is 3.38×10^{-19} J.

SOLUTION: This formula can be used directly:

$$E = \frac{hc}{\lambda}$$

Where $h = 6.63\times10^{-34}$ Js and is called Planck's constant.

energy E of a photon having wavelength λ

- From $E = \frac{hc}{\lambda}$ we see that

$$E = \frac{(6.63\times10^{-34})(3.00\times10^{8})}{588\times10^{-9}}$$

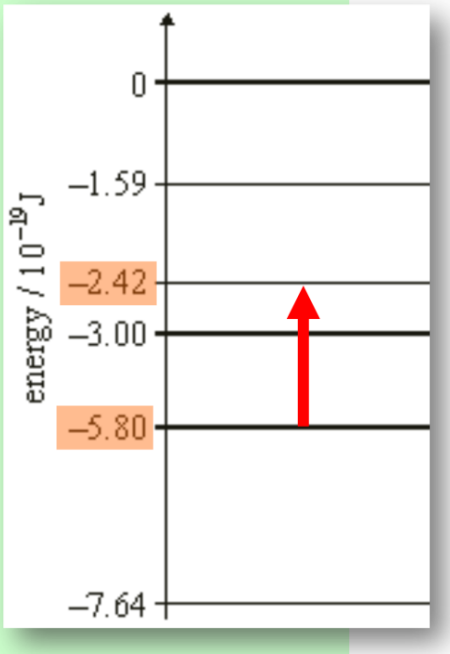
$E = 3.38\times10^{-19}$ J.

where c/lambda is frequency

Solving problems involving atomic spectra

PRACTICE: The diagram represents some energy levels of the helium atom.
(c) Use the information in the diagram to explain how absorption at 588 nm arises.
SOLUTION: We need the difference in energies between two levels to be 3.38×10^{-19} J.

- Note that 5.80 – 2.42 = 3.38.
- Since it is an absorption the atom stored the energy by jumping an electron from the -5.80×10^{-19} J level to the -2.42×10^{-19} J level, as illustrated.



Task - handout #2