Electricity & Magnetism
Lecture 30

Today's Concepts:

Bohr Atomic Model

Presentation Created by Dr. Adam Lark

Things you asked about:

"This actually kind of made sense considering I have done this all in chemistry."

"Is this really necessary to learn this in this course?? "

"Could you please explain the relationship between the wavelengths and atomic orbitals??"

"I liked this one."

Question

"Please explain all 4 quantum numbers that they listed in the prelecture."

Quantum Mechanics

Numbers describing the state of the electron

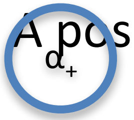
Quantum Number Name	Symbol	Function	Values
Principal	n	Specifies the energy level (i.e., electron shell number)	$1, 2, 3, \ldots$
Angular Momentum	ℓ	Specifies the magnitude of the orbital angular momentum	$0, 1, 2, \ldots, n-1$
Magnetic	m_ℓ	Describes the orientation of the subshell	$-\ell, -\ell+1, \ldots, 0, 1, \ldots, \ell$
Spin	m_s	Describes the spin angular momentum	$+\frac{1}{2}$ ("spin up"), $-\frac{1}{2}$ ("spin down")

Pauli Exclusion Principle

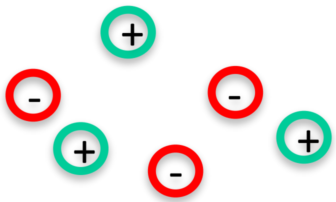
No two electrons in the same atom can have the same quantum numbers.

Rutherford Scattering

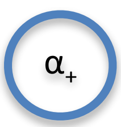
If positive and negative charges are randomly distributed in atoms



A positively charged alpha particle will not be deflected because the net charge is close to zero.

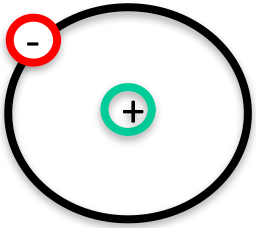


If the positive and negative charge are separated in an atom



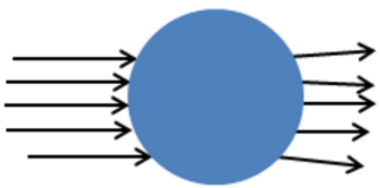
A positively charged alpha particle will be deflected some of the time.

Electrons must orbit the Protons!

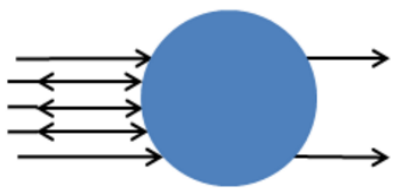


Which of the following scattering patterns correspond to the model of the atom proposed by Rutherford?

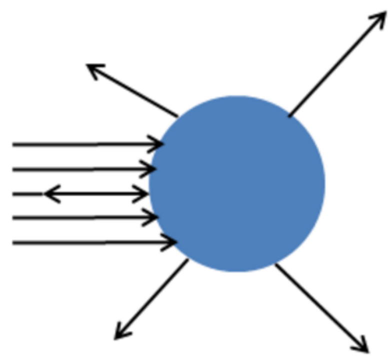
Case A



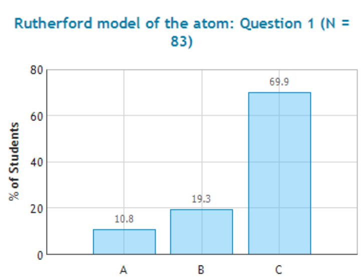
Case B



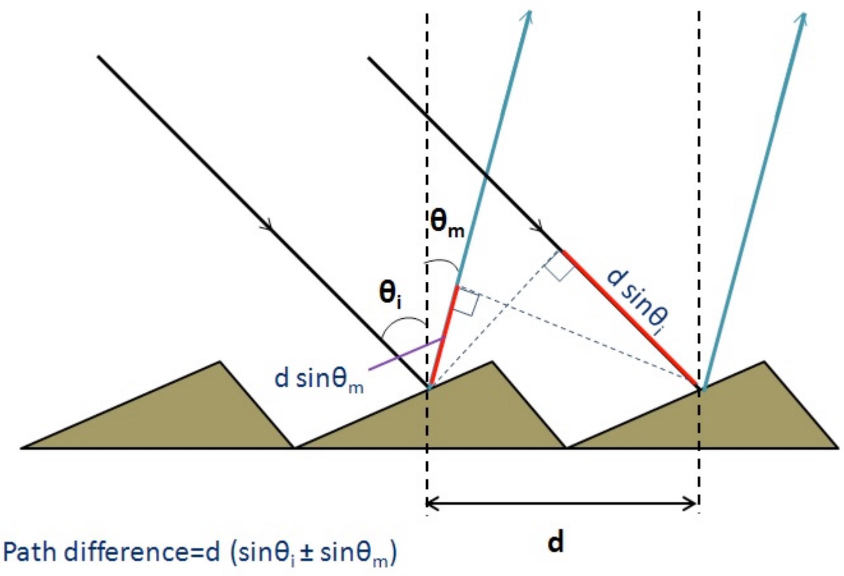
Case C



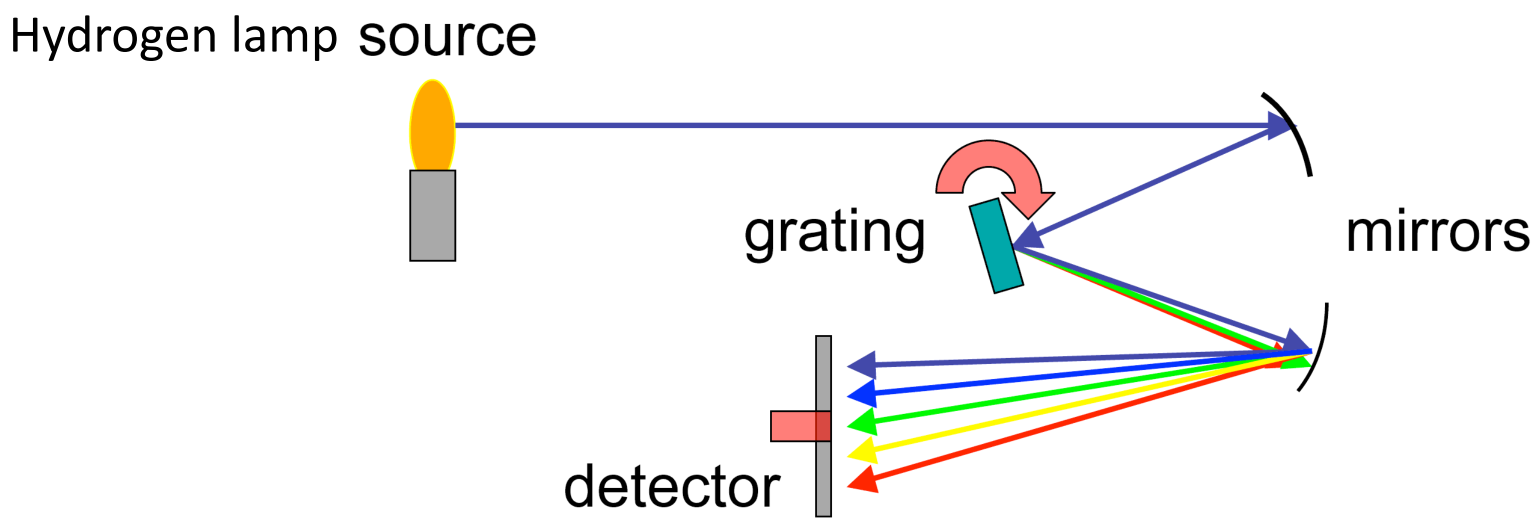
The alpha particles must scatter in all directions.



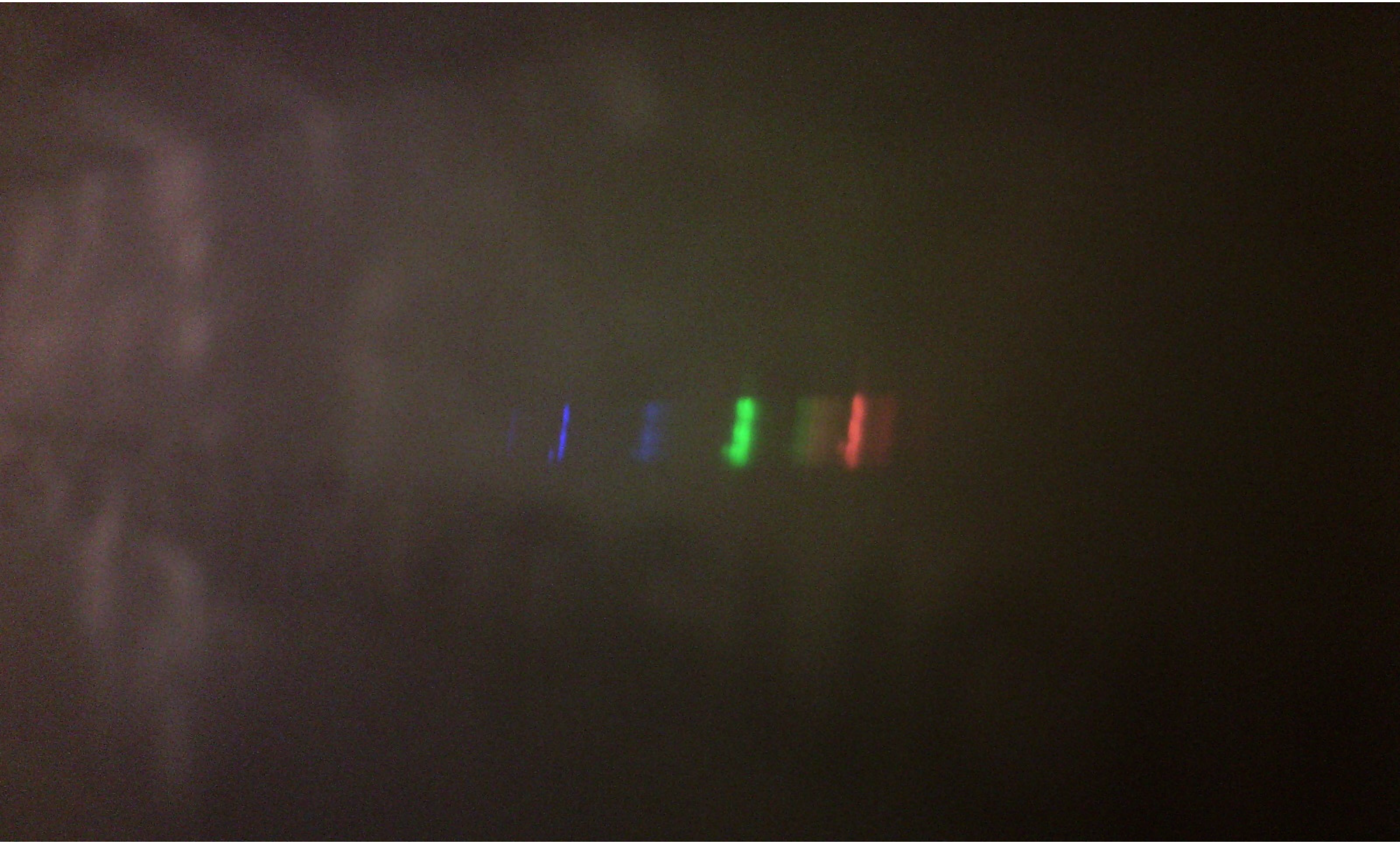
Diffraction Grating Spectrometer



Path difference=d ($\sin\theta_i \pm \sin\theta_m$)



Diffraction Grating spectrometer



Visible Spectral Lines of Hydrogen

A regular pattern of lines was observed.

Balmer Series, n = 2

m	λ, air (nm)
3	656.3
4	486.1
5	434.0
6	410.2
7	397.0
∞	364.6

$$\frac{1}{\lambda} = R_H\left(\frac{1}{n^2} - \frac{1}{m^2}\right)$$

The wavelength of that light can be found using the Rydberg formula:

$$\frac{1}{\lambda} = R_H\left(\frac{1}{n^2} - \frac{1}{m^2}\right)$$

$R_H = 1.097 \times 10^7 m^{-1}$

m = initial orbit

n = final orbit

Rydberg Formula

In the Bohr model for the hydrogen atom, what is the wavelength of the light produced by an electron falling from the state, n = 3, to the ground level, n = 1?.

(A) 102 nm

(B) 1367 nm

(C) 1823 nm

(D) 102 nm

(E) None of the above

$$\frac{1}{\lambda} = R_H\left(\frac{1}{n^2} - \frac{1}{m^2}\right)$$

R_H=1.097 x 10^7m^{-1}

λ apx= 1x 10^{-7} m

Electrons in an atom are confined to specific orbits

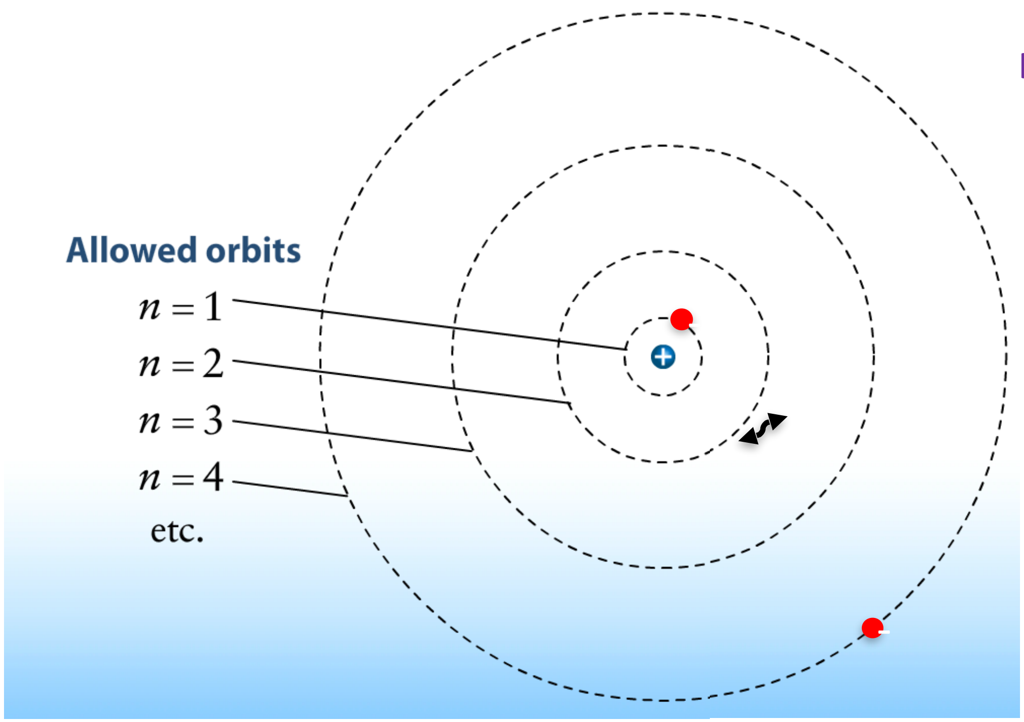
Each orbit has an **energy** associated with it:

$$E_{\text{atom},\,n} = -\frac{hcR_H}{n^2}$$

<u>Rydberg Energy</u>

for n=1: $E = E_0$

$E_0 = 2.18 \times 10^{-18}$ J



Each orbit has a **radius** associated with it:

$$r_n = \frac{n^2\hbar^2}{m_e k e^2}$$

<u>Bohr Radius</u>

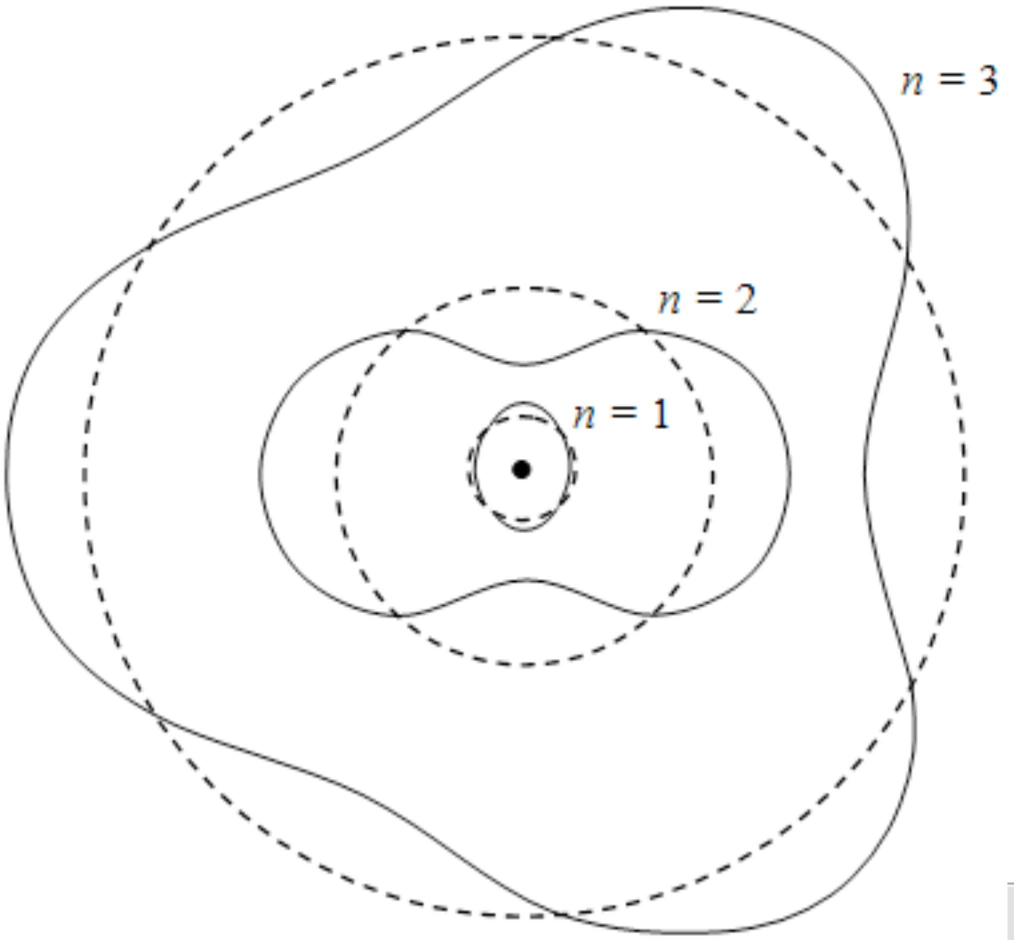
for n=1: $r_n = a_0$

$a_0 = 0.529$ nm

When electrons jump from one orbit to another, energy is gained or lost in the form of light!

de Broglie orbits

Hypothesis: Wavelengths of electrons are related to energy and only orbits with integer wavelengths are allowed.



$E = hf = hc/\lambda$

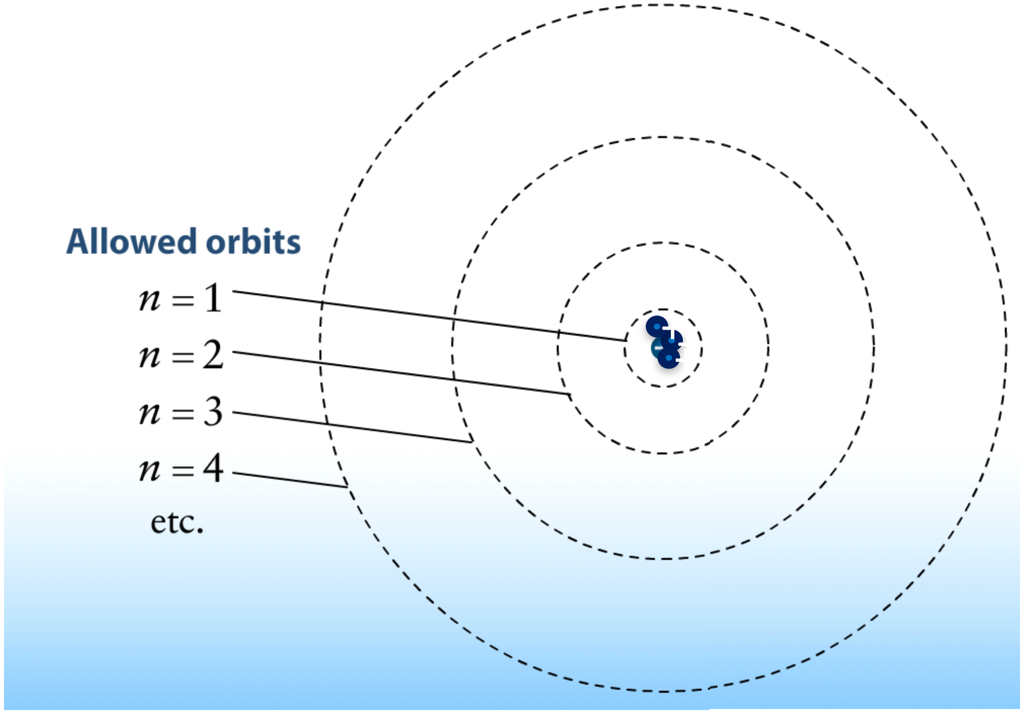
$\lambda = hc/E$

Radius and Energy Scale with Z, the number of protons in the nucleus

Energy:

$$E_n = \frac{Z^2}{n^2} E_0$$

$E_0 = 2.18 \times 10^{-18}$ J

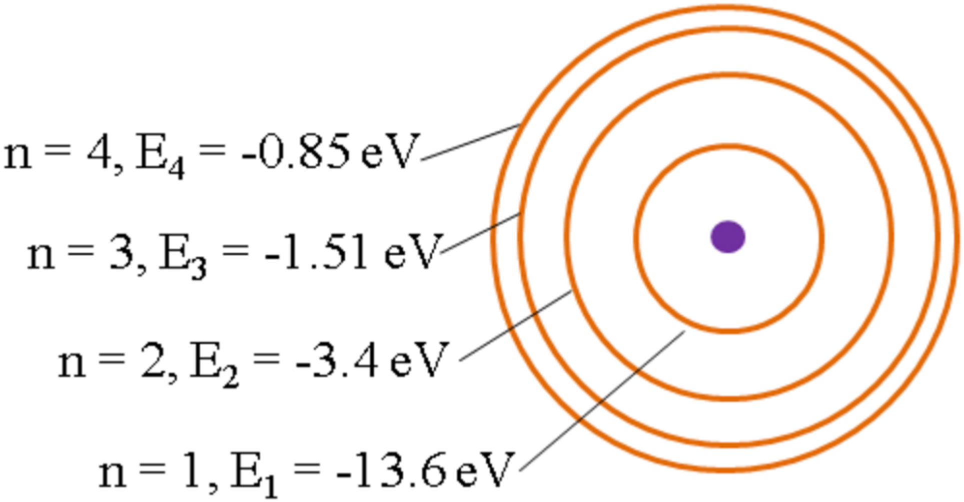


Radius:

$$r_n = \frac{n^2}{Z^2} a_0$$

$a_0 = 0.529$ nm

A drawing of the Bohr model of electron orbits in the Hydrogen atom is shown.



Which of the following will cause an electron transition from n = 1 to n = 2?

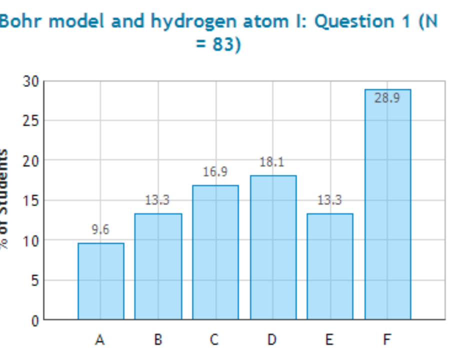
A. A photon with energy E_{photon} = 13.6 eV

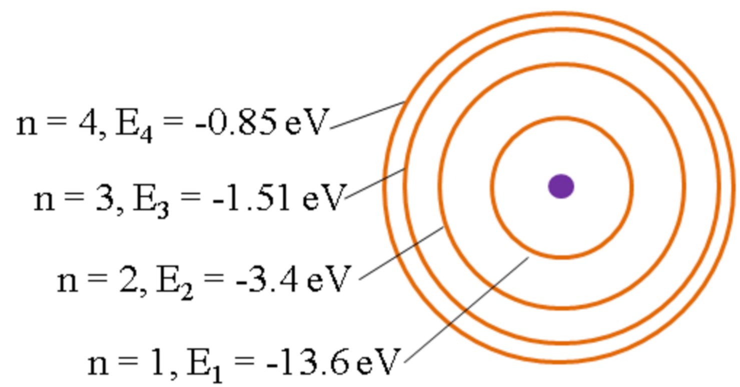
B. A photon with energy E_{photon} = 3.4 eV

C. Any photon with energy E_{photon} < 10.2 eV

D. A photon with energy E_{photon} = 11 eV

E. None of the above



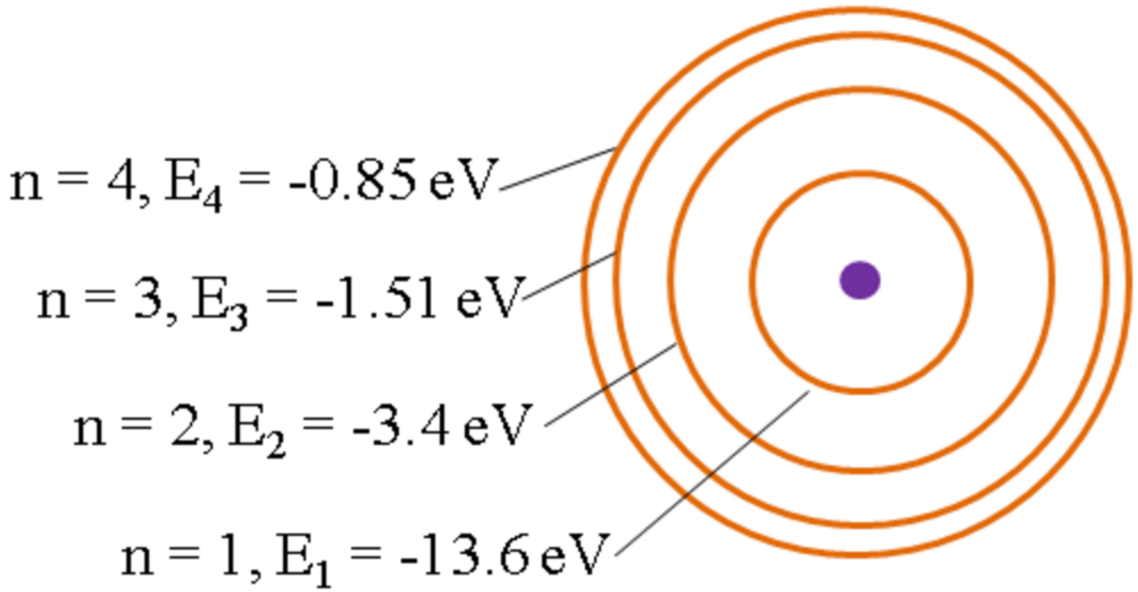


$E_F = -3.4\text{ eV}$

$- \quad E_I = -13.6\text{ eV}$

$E_{photon} = 10.2\text{ eV}$

~~Answer C?~~

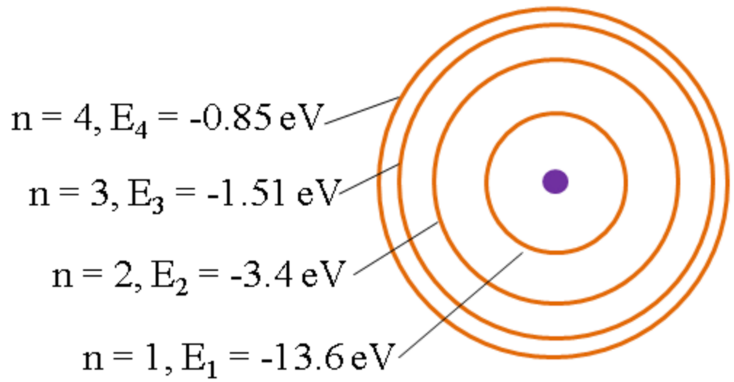


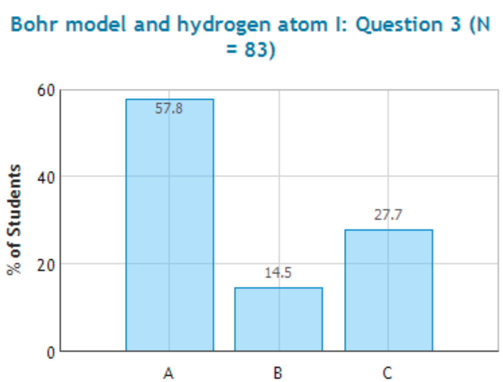
The photon must have energy exactly equal to 10.2 eV

Answer E!

Which of the following transitions will emit a photon with the longest wavelength?

- From n = 4 to n = 3
- From n = 3 to n = 2
- From n = 2 to n = 1





$$E_{photon} = hf$$

Lower energy = lower frequency

lower frequency = longer wavelength

$h = 4.13 \times 10^{-15}\ eV \cdot s = 6.63 \times 10^{-34}\ J \cdot s$

Bridge Question 3

How many possible wavelengths could be observed when a hydrogen atom initially excited to the n = 4 state returns to the ground state?

Orbits:	4	3	2	1
	n=4	to n=3		
	n=4		to n=2	
	n=4			to n=1
		n=3	to n=2	
		n=3		to n=1
			n=2	to n=1

A. 2
B. 3
C. 4
D. 5
E. 6

Bridge Question 4

According to the Bohr model, which of the ions listed in the table below would have the smallest distance between an electron in the lowest energy level and the nucleus?

A. He

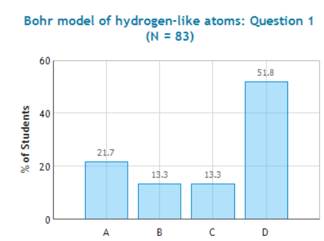
B. Li

C. Be

D. B

$$r_n = \frac{n^2}{Z^2} a_0$$

Largest Z, smallest r!



Ion	He^{+}	Li^{2+}	Be^{3+}	B^{4+}
Atomic number	2	3	4	5

[Hydrogen atom simulation]