Atoms

and the

Bohr Model

Reading: Gray: (1-1) to (1-7)
OGN: (15.1) and (15.4)

Outline of First Lecture

I. General information about the atom

II. How the theory of the atomic structure evolved

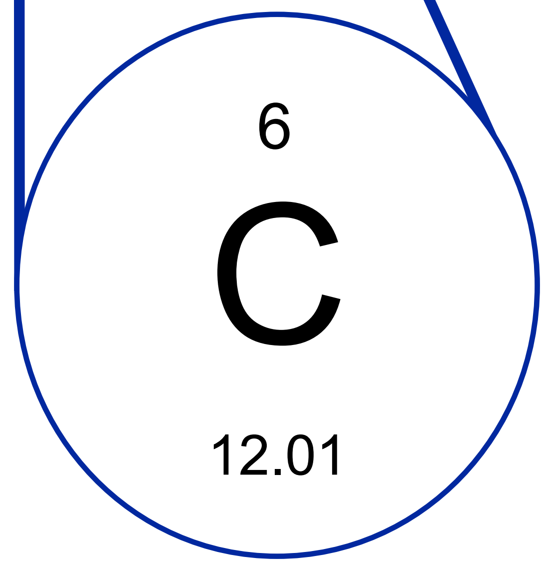
A. Charge and Mass of the atomic particles
   1. Faraday
   2. Thomson
   3. Millikan

B. Rutherford's Model of the atom

Reading: Gray: (1-1) to (1-7)
OGN: (15.1) and (15.4)

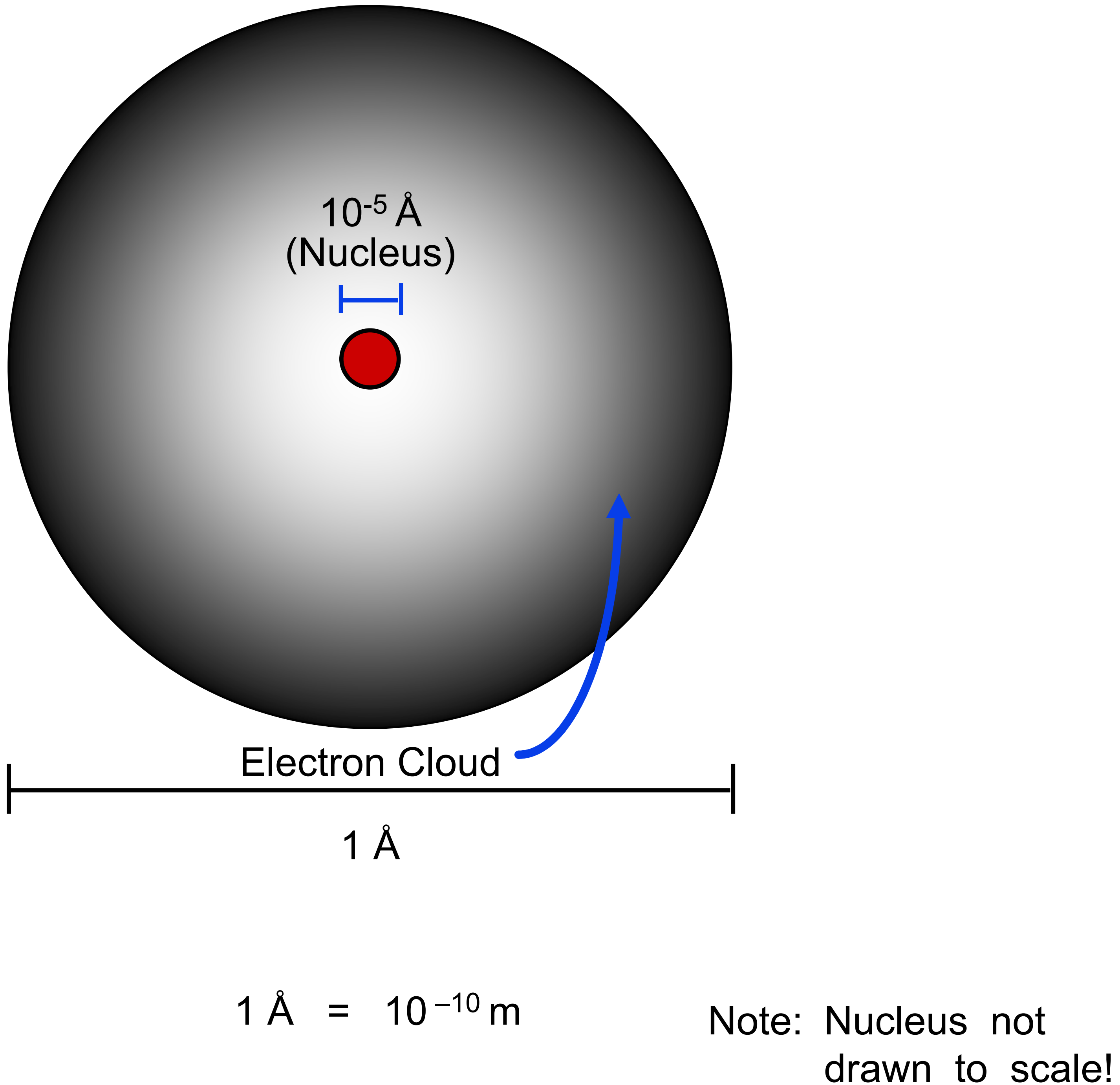
H																	He
Li	Be											B	C	N	O	F	Ne
Na	Mg											Al	Si	P	S	Cl	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
Fr	Ra	Ac															

Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu
Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr



Atoms consist of:

	Mass (a.m.u.)	Charge (eu)
Protons	~1	+1
Neutrons	~1	0
Electrons	1/1836	-1


10^{-5} Å
(Nucleus)
Electron Cloud
1 Å
1 Å = 10^{-10} m
Note: Nucleus not drawn to scale!

HOW DO WE KNOW:

- **Atomic Size?**
- **Charge and Mass of an Electron?**
- **Charge and Mass of a Proton?**
- **Mass Distribution in an Atom?**

Reading: Gray: (1-1) to (1-7)
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A Timeline of the Atom

← **400 BC …… 0 ……. 1800 1850 1900 1950**

-450 Empedocles- earth, air, fire, water
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1661 Robert Boyle: disputes 4 element theory; postulates an element is a substance that can not be reduced into simpler substances
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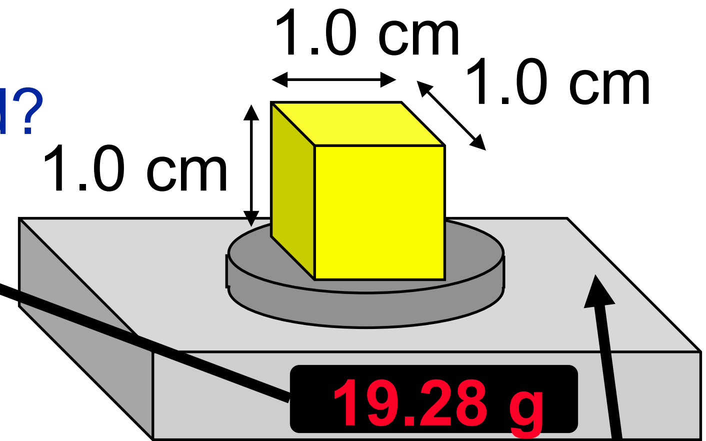
Calculating the Number of Atoms in One Cubic Centimeter of Gold



How many moles of gold?

Atomic weight of gold = 197.0 g / mole

$$\frac{19.28\ \text{g}}{197.0\ \text{g / mole}} = 0.09786\ \text{moles}$$

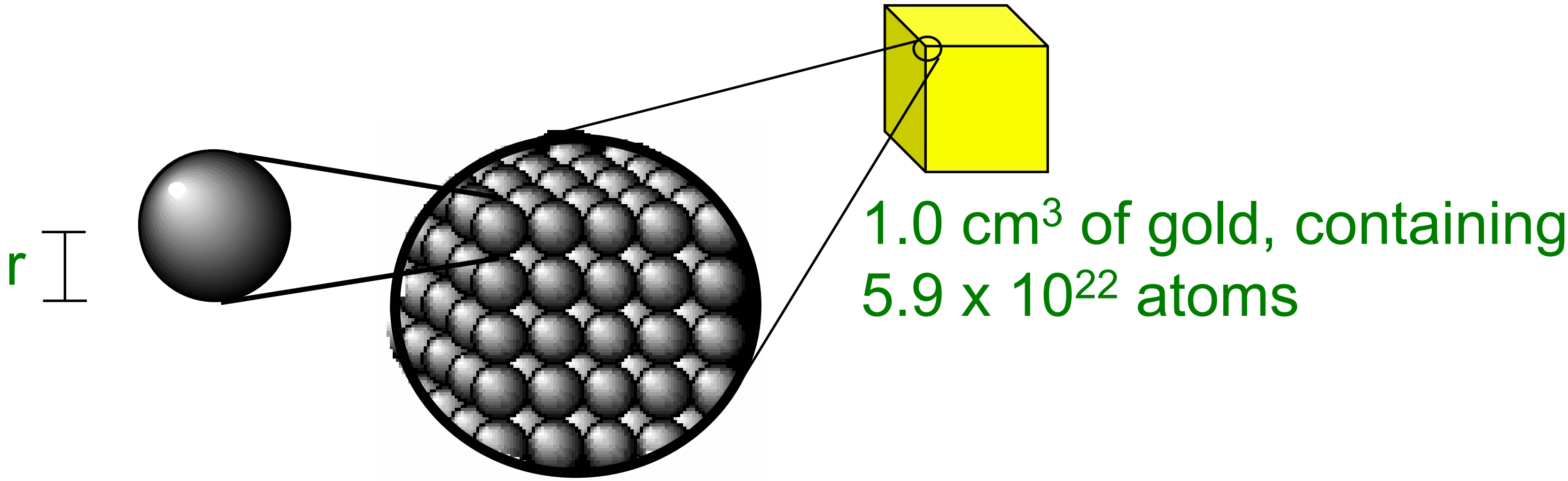
How many gold atoms?

$N_A = 6.022 \times 10^{23}$ atoms / mole

(0.09786 moles)x(6.022×10^{23} atoms / mole)=5.893×10^{22} atoms

There are approx. 5.9×10^{22} atoms in 1 cm^3 of gold

Calculating Atomic Size



Assuming each atom takes up a volume of $\frac{4}{3}\pi r^3$, we can calculate the radius of a single atom:

$$\frac{1.0\ \text{cm}^3}{(5.9 \times 10^{22})} = \frac{4}{3}\pi r^3$$

total volume

number of atoms

volume of one atom

radius of one gold atom
= 1.6×10^{-8} cm, or
1.6 angstoms

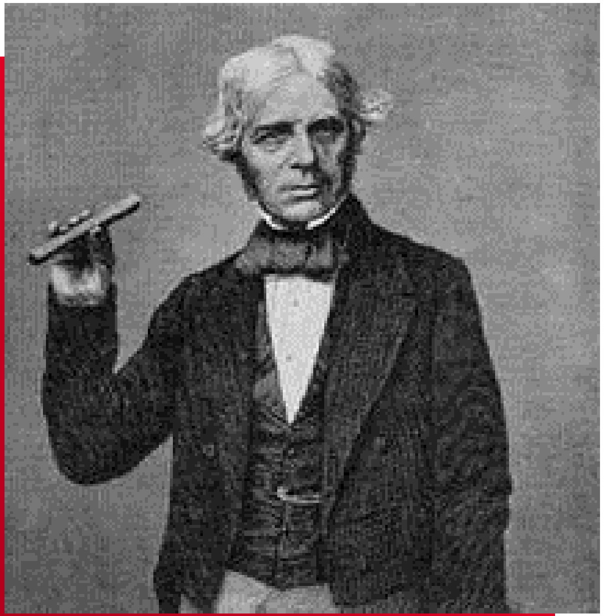
The Person Behind The Science

Michael Faraday

1791-1867

Highlights

- Bookbinder turned self-taught scientist
- Discovered magnetic optical rotation
- Invented the *Dynamo*, a device capable of converting electricity into motion (1821)



Moments in a Life

- Began experimenting on electricity in 1813 under Sir Humphrey Davy
- Discovered electromagnetic induction (1831)
- Published a three volume treatise titled *Experimental Researches in Electricity* (1839-1855)

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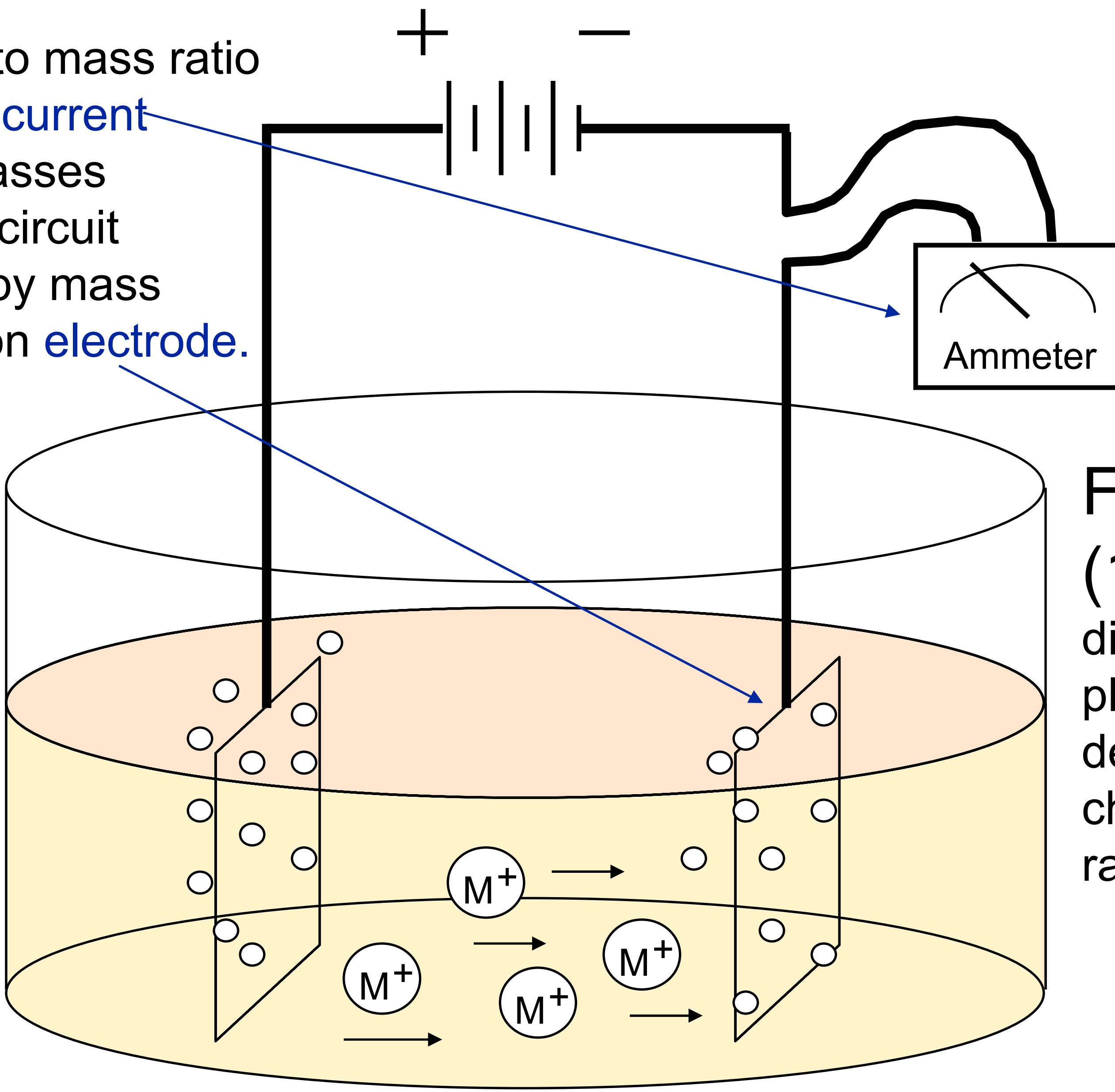
Charge per Unit Mass of Ions

Charge to mass ratio of ions: current which passes through circuit divided by mass gained on electrode.



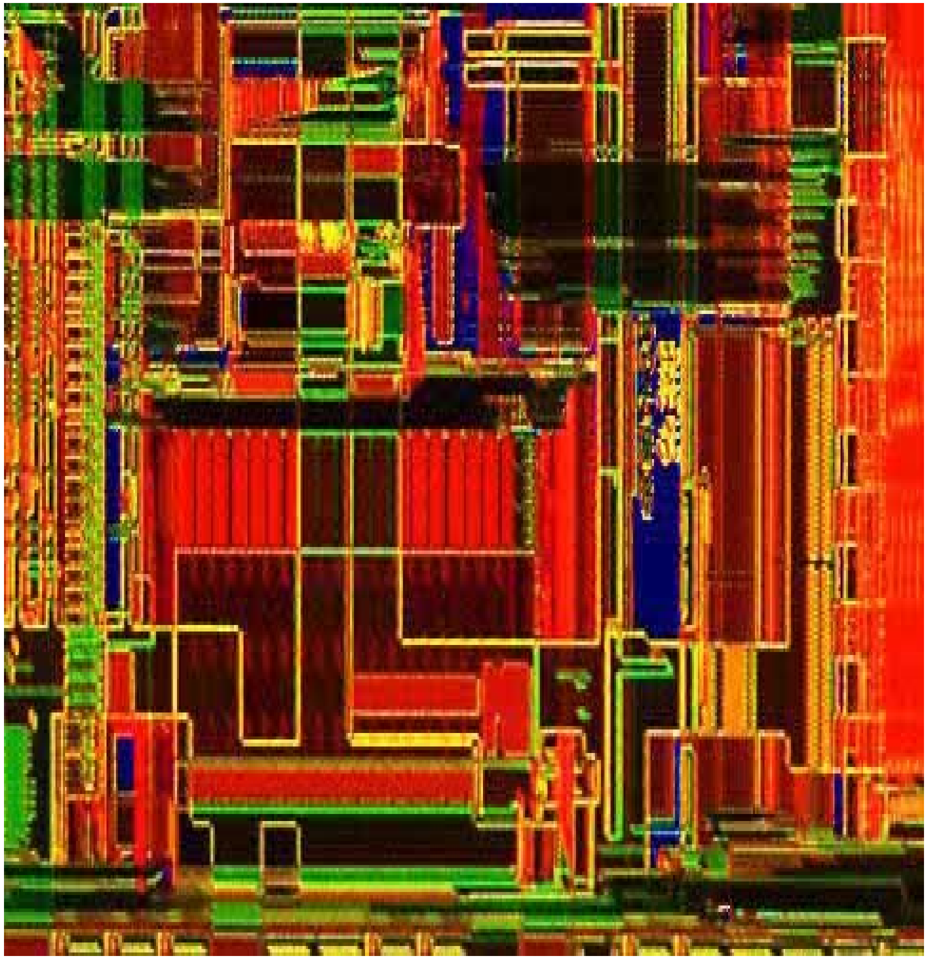
Faraday (1830):
discovered this phenomenon and determined the charge to mass ratio of ions

What Faraday Found

- Charge to mass ratio for various ions.
- e.g. e/m for $H^+ = 10^8$ C/kg

Reading: Gray: (1-1) to (1-7)
OGN: (15.1) and (15.4)

Electroplating



The Person Behind The Science

Sir J.J. Thomson

1856-1940

Highlights

- Cavendish Professor of Experimental Philosophy at Cambridge University (most important position in physics at the time)
- Received his degree from Trinity College in mathematics (1880)

Moments in a Life

- Appointed Cavendish Professor (1884)
- Won the Nobel Prize in Physics (1906) for his work on the properties of the electron

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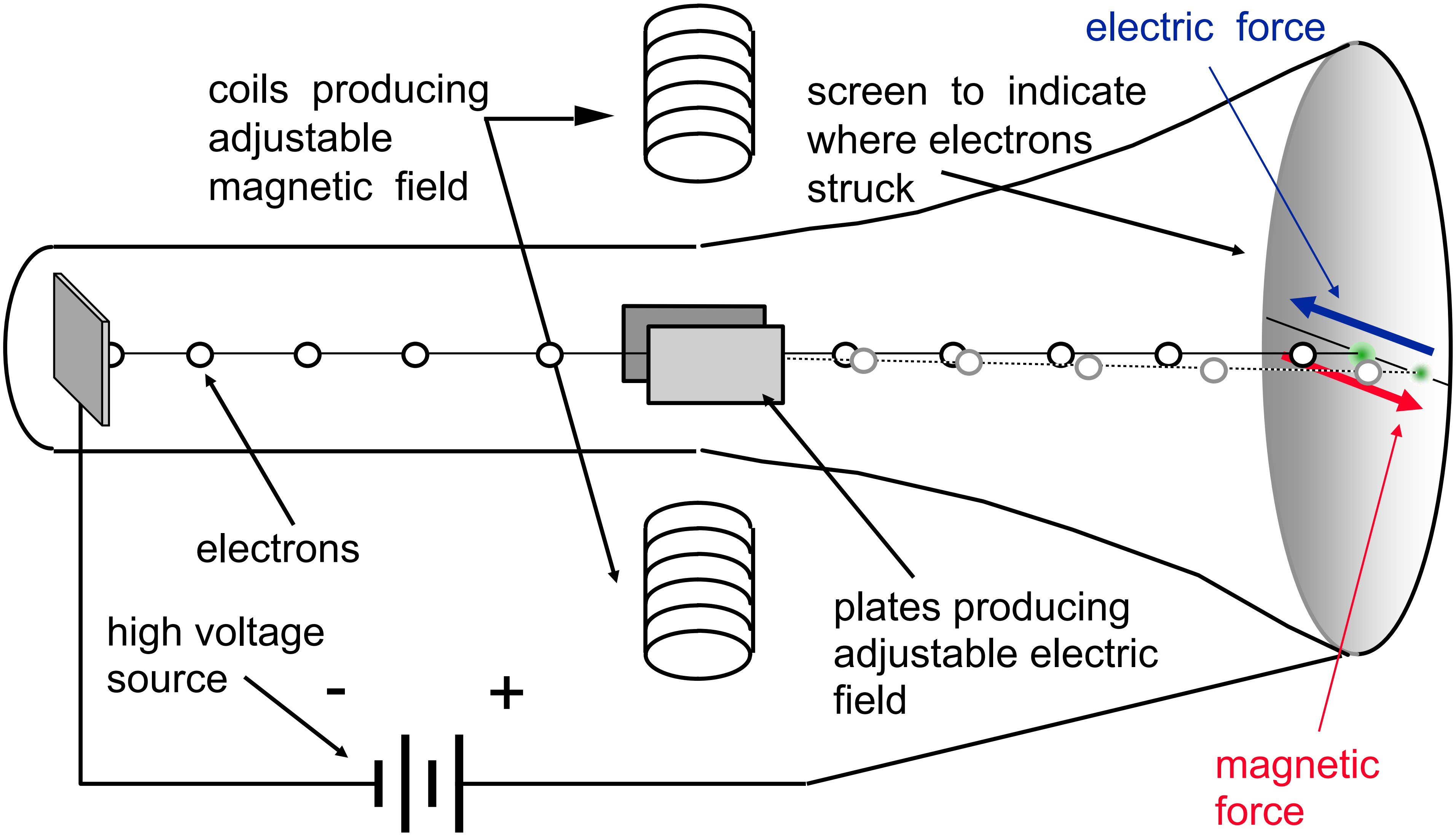
Charge to Mass Ratio of Electrons



Thomson: determined the charge-to-mass ratio of electrons by balancing the force laws: $\vec{F}_{mag} = q\vec{v}\times\vec{B}$, $\vec{F}_{elec} = q\vec{E}$

What Thomson Found (1897)

- Found charge/mass ratio for electron: $\sim 1.2 \times 10^{11}$ C/kg

- e/m for electron = 1.2×10^{11} C/kg, e/m for proton = 10^{8} C/kg.

- Since 1.2×10^{11} C/kg >> 10^{8} C/kg, either the electron has a far greater charge than the proton, or it has far less mass.

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The Person Behind The Science

Robert Millikan

1868-1953

Highlights

- Appointed Director of the Norman Bridge Laboratory of Physics, Caltech (1921)
- Received Nobel Prize in Physics (1923)
- Worked on experimental aspects of photoelectric effect

Moments in a Life

- Became a professor at the University of Chicago (1910)
- Performs his famous oil drop experiments (1909)

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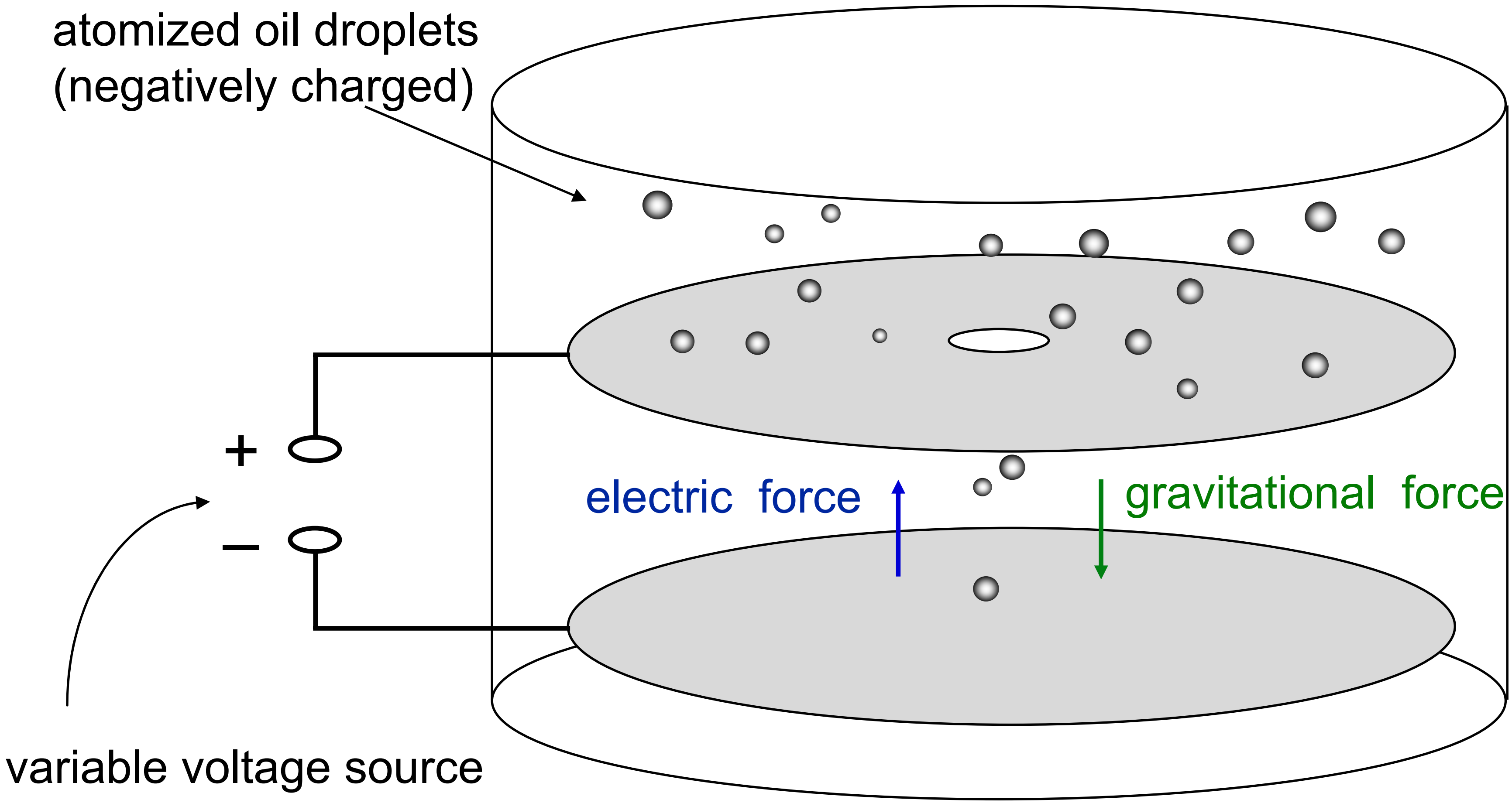
Absolute Charge on an Electron



Millikan: determined that charges occur in multiples of 1.60×10^{-19} C (the charge of an electron) by balancing gravitational and electrical forces and using the previously-obtained charge to mass ratio of electrons

Conclusions

We know...	• Atomic radius ≈ 10^{-8} cm • (e/m) for proton ≈ 10^{8} C/kg • (e/m) for electron ≈ 1.2×10^{11} C/kg • Charge of an electron ≈ 1.6×10^{-19} C
We assume...	• Charge of a proton is equal and opposite that of an electron

WE DEDUCE...

• Mass of the electron:

$$\frac{e^- \text{ charge}}{e/m \text{ for } e^-} = \frac{1.6 \times 10^{-19}\ \text{C}}{1.2 \times 10^{11}\ \text{C/kg}} \approx \boxed{1.3 \times 10^{-30}\ \text{kg}}$$

(The currently accepted value is: 9.1×10^{-31} kg)

• Mass of the proton:

$$\frac{p^+ \text{ charge}}{e/m \text{ for } p^+} = \frac{1.6 \times 10^{-19}\ \text{C}}{1 \times 10^{8}\ \text{C/kg}} \approx \boxed{1.2 \times 10^{-27}\ \text{kg}}$$

(The currently accepted value is: 1.7×10^{-27} kg)

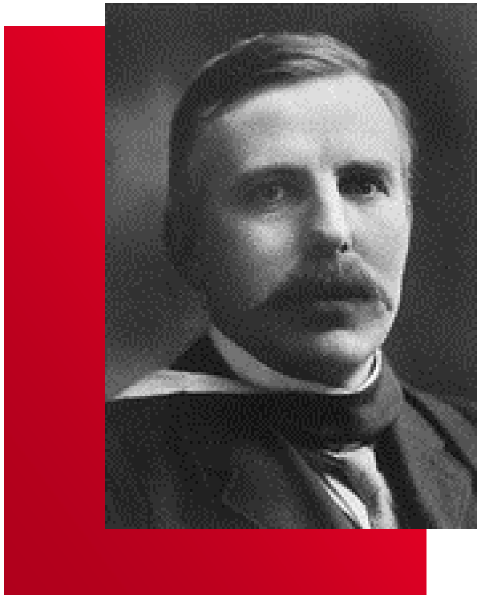
The Person Behind The Science

Ernest Rutherford

1871-1937

Highlights

- Worked for J.J. Thomson (1895)
- The *scientific father* of 10 Nobel Prize Laureates
- Worked on alpha, beta, and gamma particles



Moments in a Life

- Performs particle deflection experiments (1907)
- Awarded Nobel Prize in Chemistry (1908)
- Assumes position at Cambridge formerly held by J.J. Thomson (1919)

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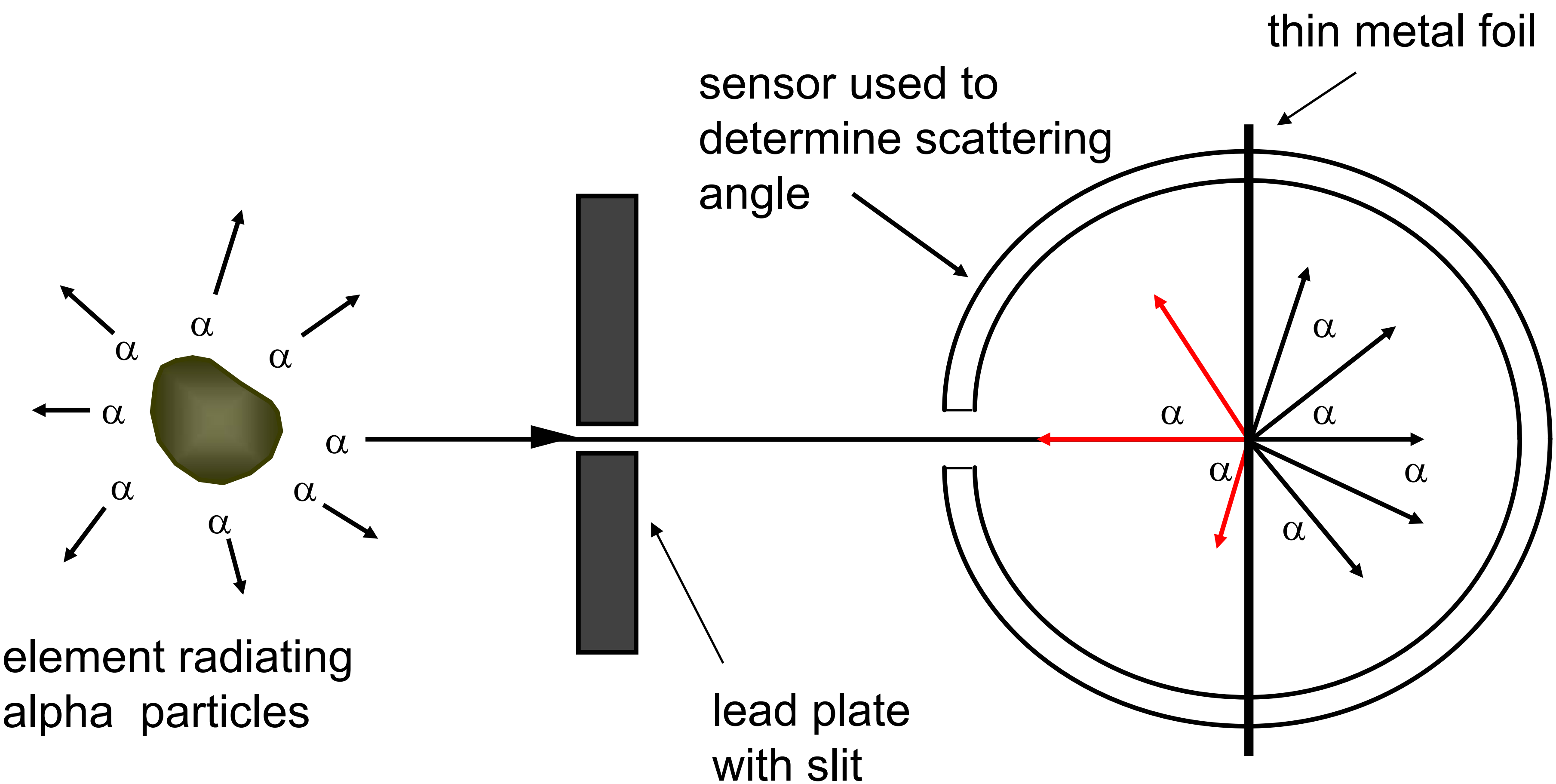
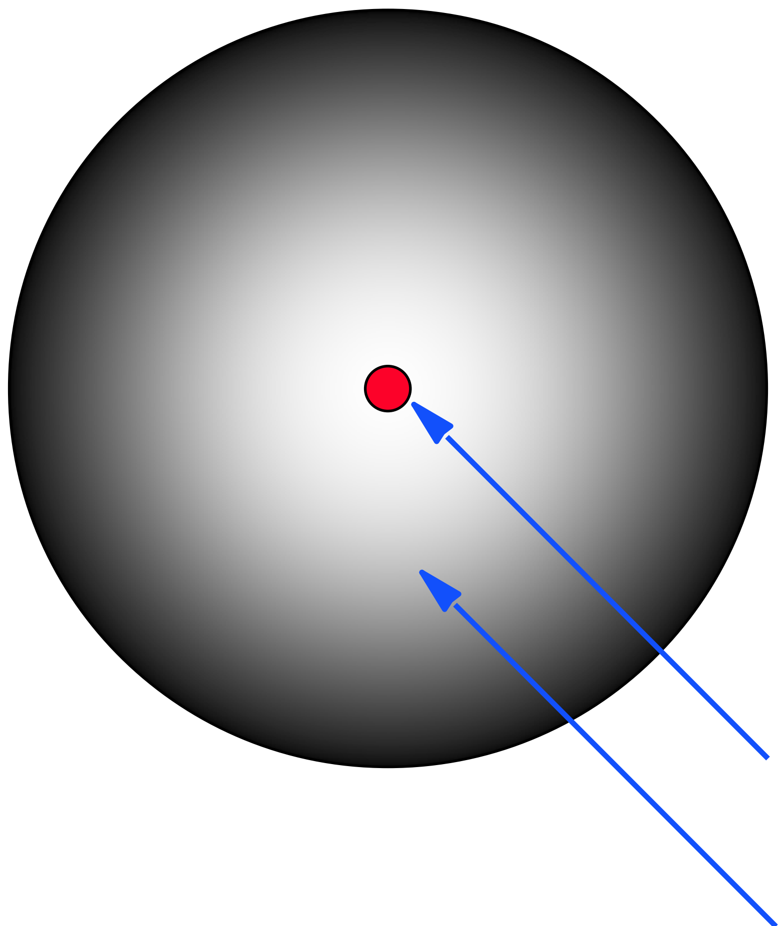
Mass Distribution in an Atom



Rutherford: showed that the mass of an atom is not distributed evenly (otherwise every angle of scattering would be close to zero)

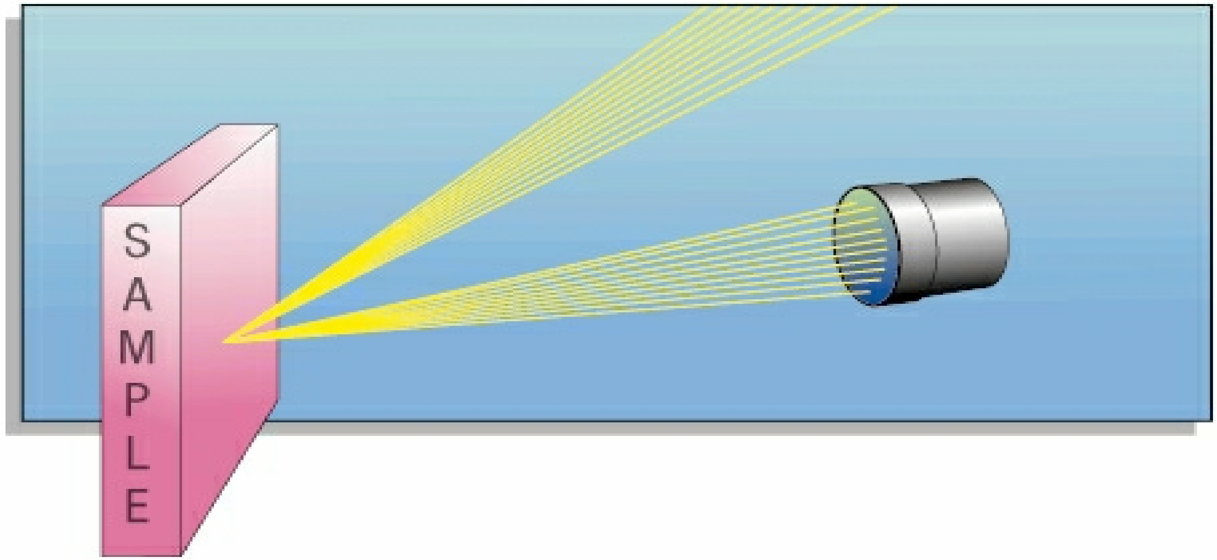
Rutherford's Model of the Atom (1911)

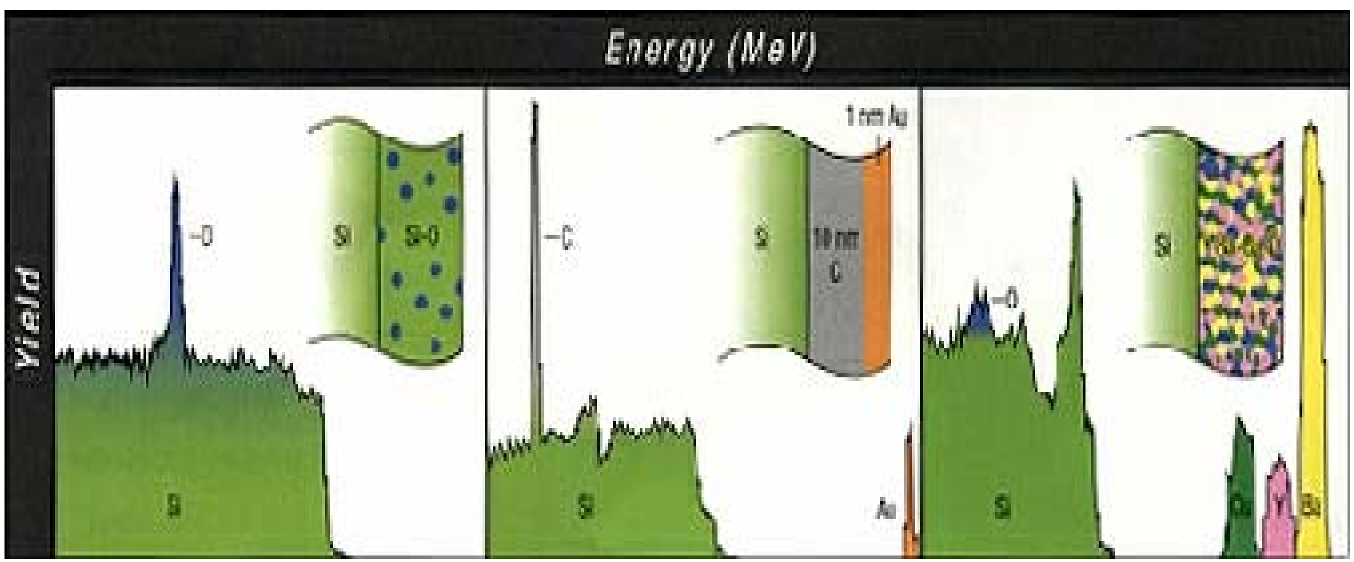


Rutherford thus showed that the nucleus is very dense

and that the remainder is virtually all vacuum

Rutherford Backscattering Spectroscopy



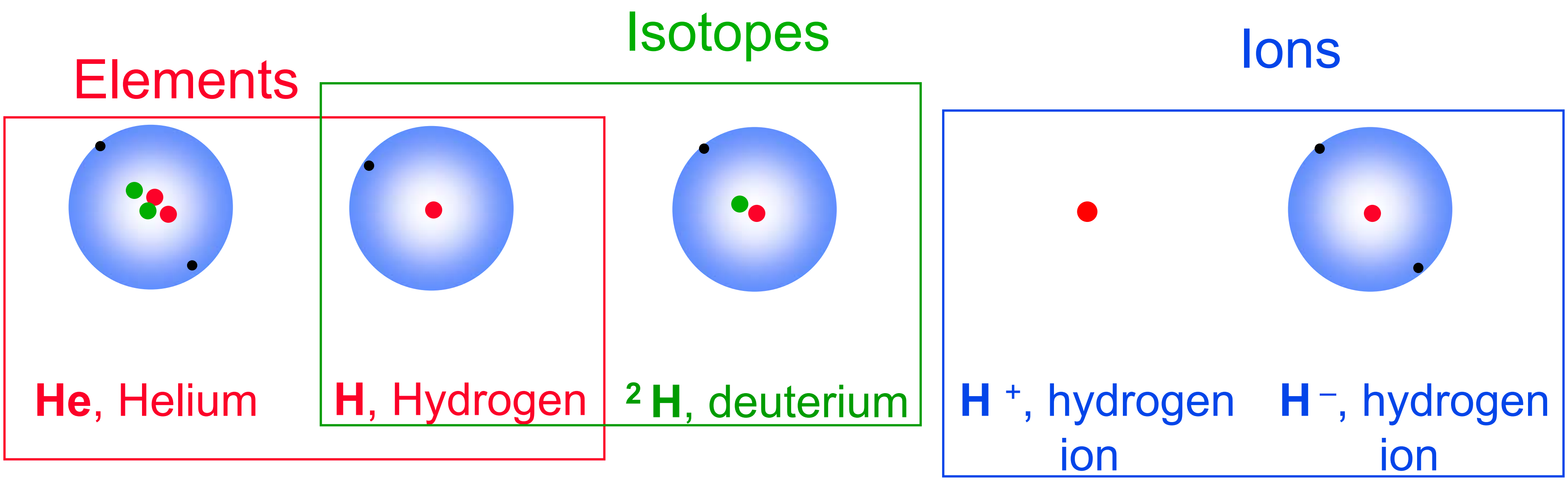


Useful Definitions:

Element: A substance containing atoms of only one type, i.e. atoms with the same number of **protons**.

Isotopes: Atoms of the same element having different masses—having different numbers of **neutrons**.

Ions: Positively or negatively charged atoms—have different numbers of **electrons**.



REVIEW

- **Atoms Consist of Protons, Neutrons, Electrons**
- **Protons: 1 a.m.u. +1 charge**
- **Neutrons: 1 a.m.u. 0 charge**
- **Electrons: 1/1836 a.m.u. -1 charge**
- **Atomic Radius: ≈1 Å (= 10^{-10} m)**
- **Most of the Mass is in the Nucleus; 10^{-5} Å**

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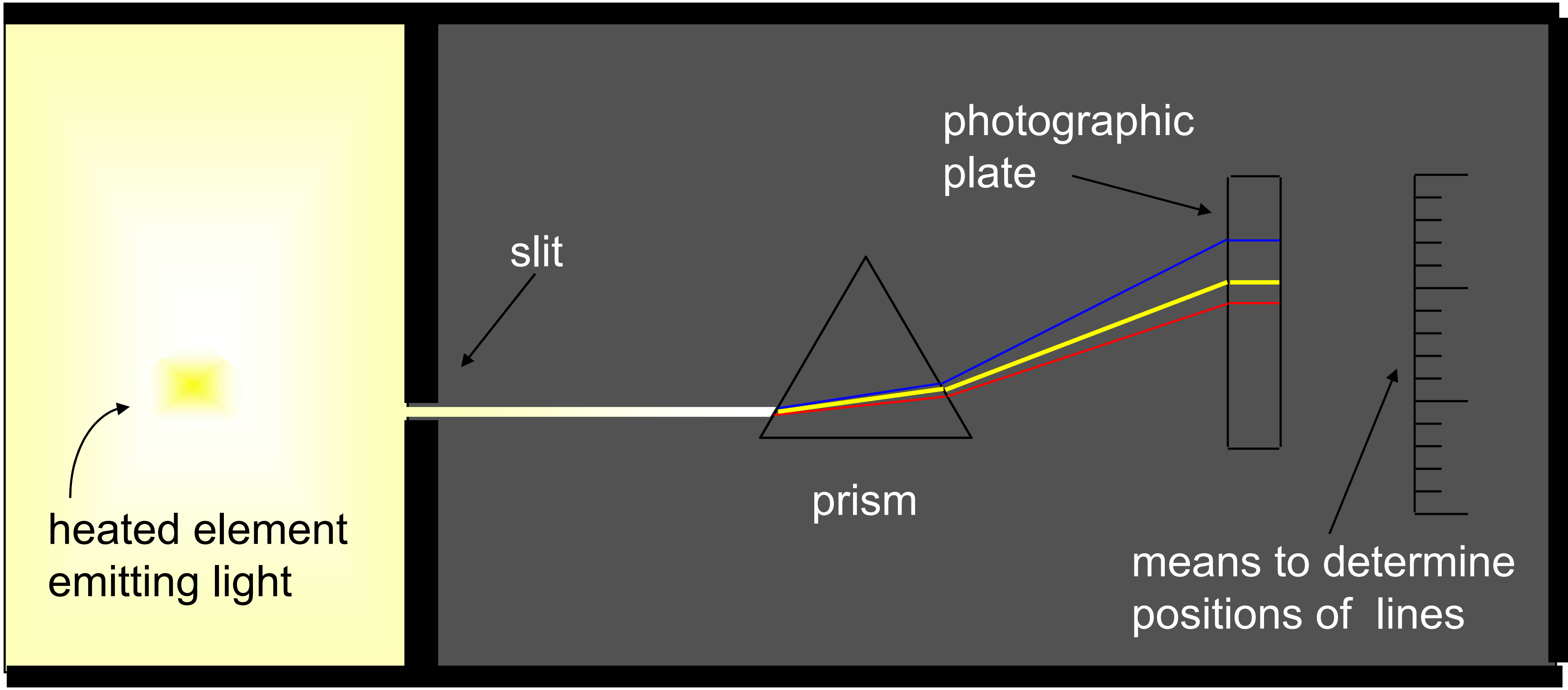
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
Fr	Ra	Ac															
				Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu
				Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr

How do we know which element is which?

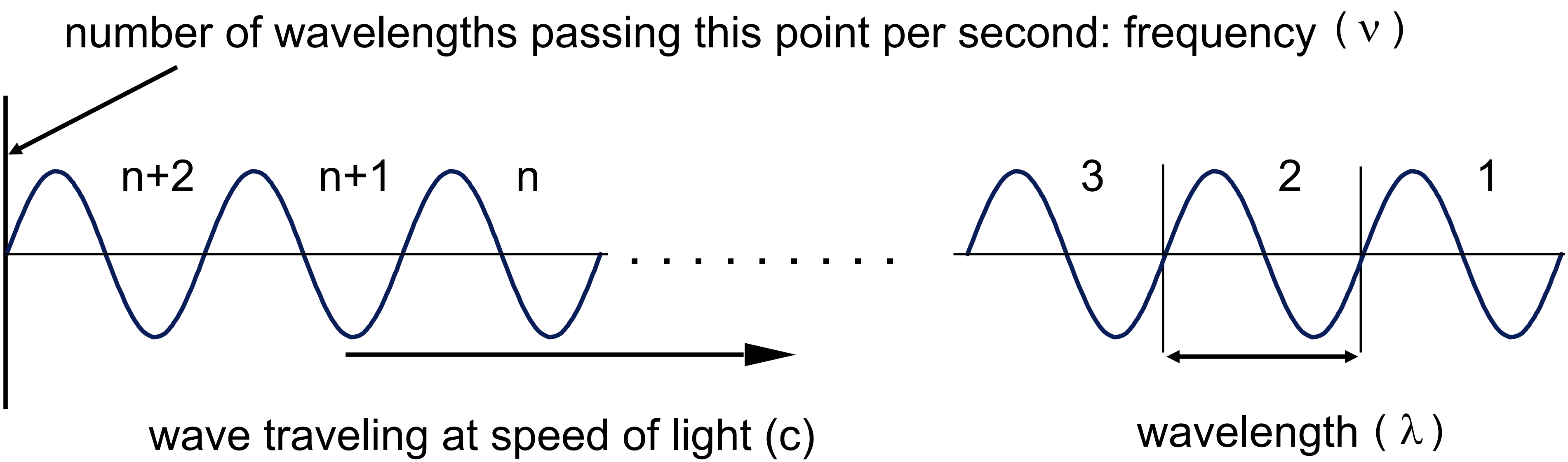
Reading: Gray: (1-1) to (1-7)
OGN: (15.1) and (15.4)

Emission Spectra



*Each element has characteristic emission lines

Properties of Light



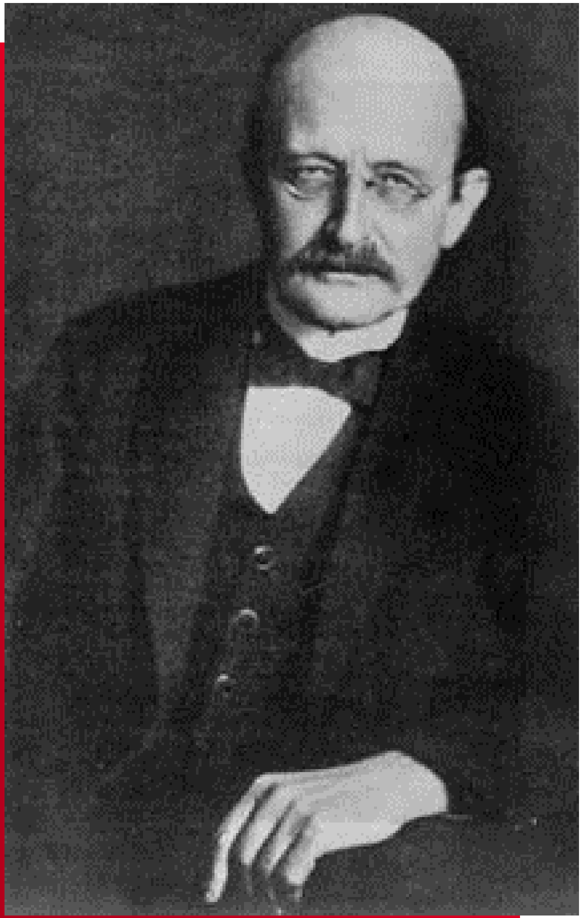
General relationship: $\mathbf{c} = \nu\lambda$

Max Planck

1858-1947

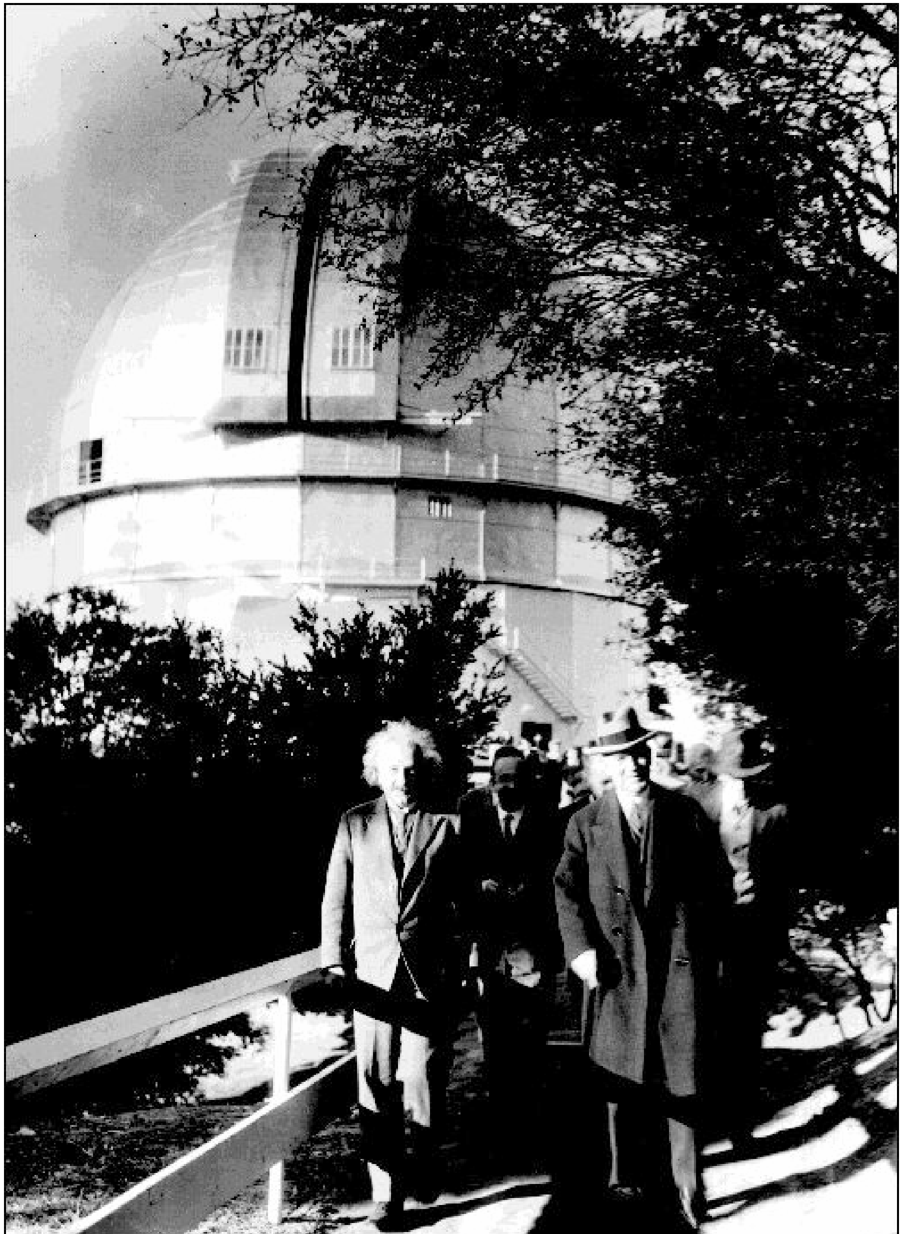
Highlights

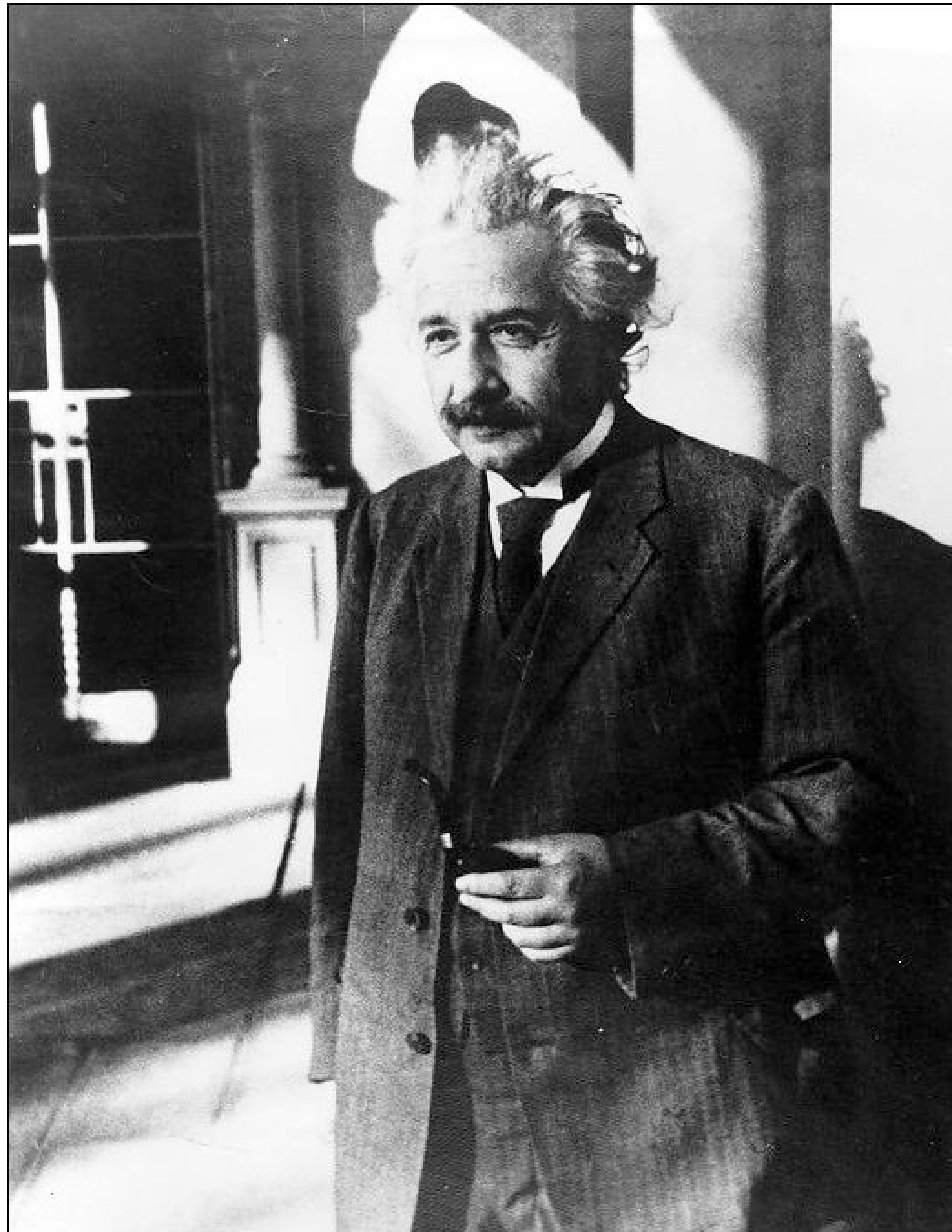
- Early worked focused on thermodynamics
- Received Nobel Prize in Physics (1918)
- Father of the quantum revolution



Moments in a Life

- Received his doctorate from Universities of Munich and Berlin under the guidance of Kirchhoff and Helmholtz (1879)
- Published his work on quanta (1900)





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Other Useful Relations:

$$E = h\nu$$

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

E: energy

h: Planck's constant

ν: frequency

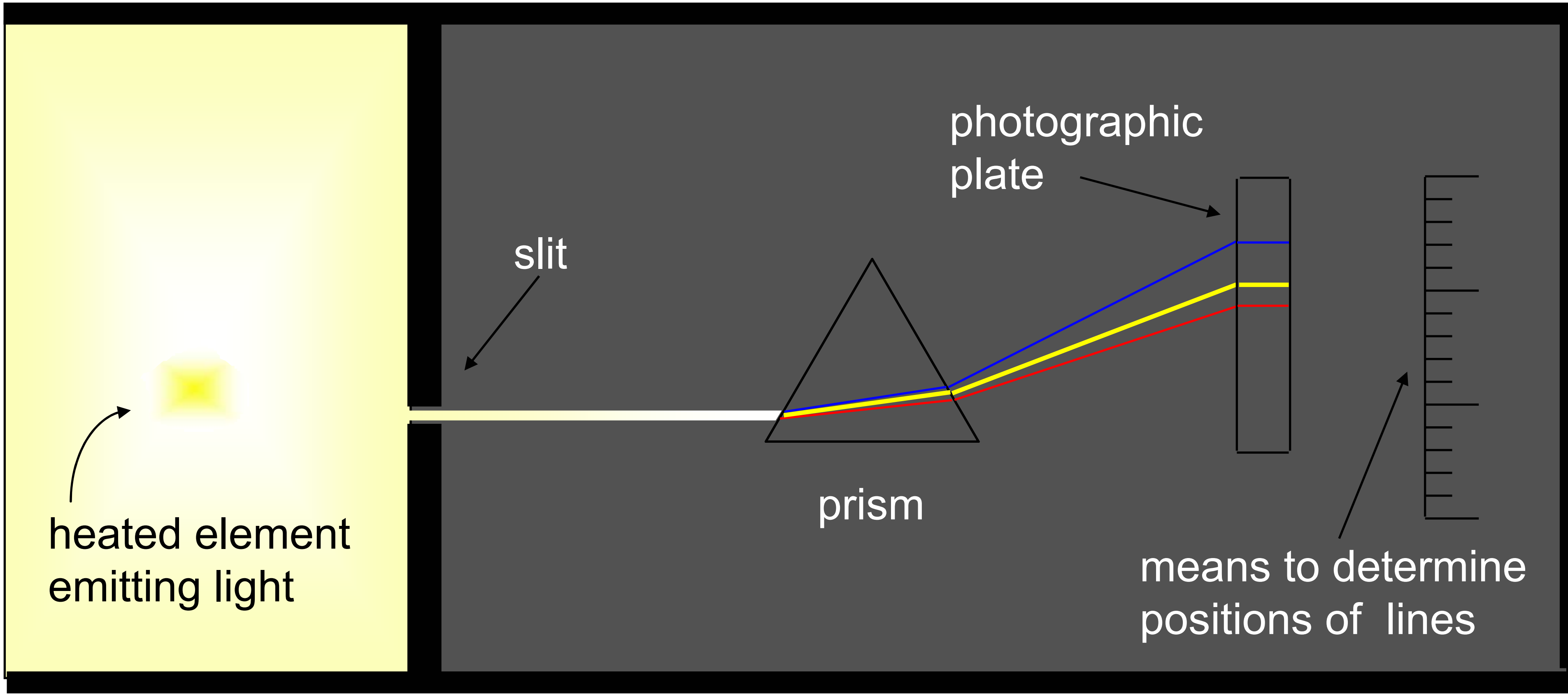
$\left(\nu = \frac{c}{\lambda} \right)$

$$\bar{\nu} \equiv \frac{1}{\lambda}$$

$\bar{\nu}$: wavenumber

$$E = hc\bar{\nu}$$

Emission Spectra



*Each element has characteristic emission lines

E = hν gives E for each line

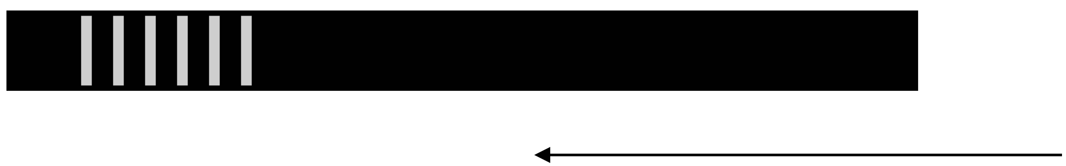
Lyman Lines:

Examples:

$$E = 13.6 \text{ eV} \left(1 - \frac{1}{n^2}\right) \quad n = 2$$

$$E = 13.6 \text{ eV} \left(1 - \frac{1}{n^2}\right) \quad n = 3$$

General Case:



shorter wavelength,
higher energy

$$E = 13.6 \text{ eV} \left(1 - \frac{1}{n^2}\right) \quad n \in \mathbf{Z} > 1$$

Johann Balmer (1885)

Johann Balmer's doctorate from Basel was for a dissertation on the cycloid.
He taught in Basel all his life both as a school teacher and as a
university lecturer at the University of Basel. His main field of interest was geometry.

However Balmer is best remembered for his work on spectral series and his formula,
given in 1885, for the wavelengths of the spectral lines
of the hydrogen atom. The reason why the formula holds was not understood until the
work of Niels Bohr in 1913.

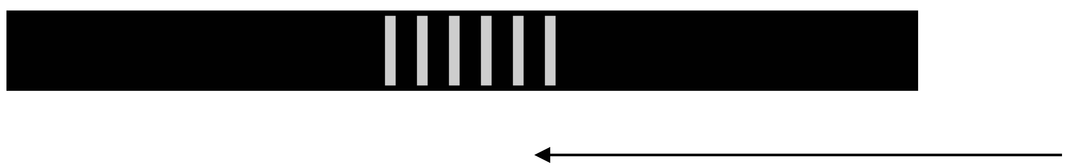
Balmer Lines:

Examples:

$$E = 13.6 \text{ eV} \left(\frac{1}{2^2} - \frac{1}{n^2}\right) \quad n = 3$$

$$E = 13.6 \text{ eV} \left(\frac{1}{2^2} - \frac{1}{n^2}\right) \quad n = 4$$

General Case:



$$E = 13.6 \text{ eV} \left(\frac{1}{2^2} - \frac{1}{n^2}\right) \quad n \in \mathbf{Z} > 2$$

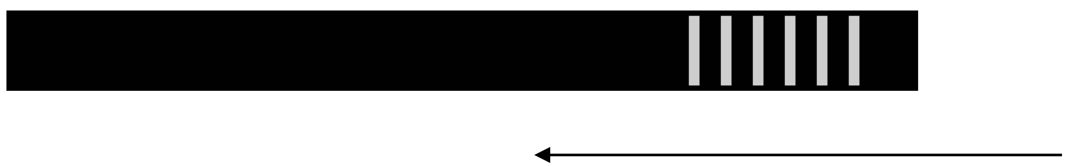
Paschen Lines:

Examples:

$$E = 13.6\ \text{eV}\left(\frac{1}{3^2} - \frac{1}{n^2}\right) \quad n = 4$$

$$E = 13.6\ \text{eV}\left(\frac{1}{3^2} - \frac{1}{n^2}\right) \quad n = 5$$

General Case:



shorter wavelength,
higher energy

$$E = 13.6\ \text{eV}\left(\frac{1}{3^2} - \frac{1}{n^2}\right) \quad n \in \mathbf{Z} > 3$$

Combined Formula for E:

$$E = 13.6\text{eV}\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right) \quad , \quad n_1, n_2 \in Z \middle| n_2 > n_1$$

With available values, the Rydberg constant was determined

$$E = hcR_H\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

$$R_H = \frac{e^2}{2hc\ 4\pi\varepsilon_o\ a_o} = 109\ 677.581 \pm 0.007\text{cm}^{-1}$$

Forms of the Rydberg constant:

$$E = R_H\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

$R_H = 13.6\ \text{eV}$ (energy)

$E = \frac{hc}{\lambda}$ ➡ $R_H = \frac{hc}{13.6\text{eV}} = 9.118 \times 10^{-6}\ \text{cm}$ (wavelength)

$E = h\nu$ ➡ $R_H = \frac{13.6\text{eV}}{h} = 3.288 \times 10^{15}\ \text{Hz}$ (frequency)

$E = hc\bar{\nu}$ ➡ $R_H = \frac{13.6\text{eV}}{hc} = 1.097 \times 10^{5}\ \text{cm}^{-1}$ (wavenumber)

The Classical Paradox of Atoms

$$\text{Potential Energy} \equiv PE = \frac{(-e)(+e)}{4\pi\varepsilon_0 r} = \frac{-e^2}{4\pi\varepsilon_0 r}$$

$$\text{Kinetic Energy} \equiv KE = \frac{1}{2} m_p v_p^2 + \frac{1}{2} m_e v_e^2$$

Total Energy = PE + KE

But, lowest energy, i.e. ground state, is when?

$PE \rightarrow 0$ as $r \rightarrow \infty$, but $PE \rightarrow -\infty$ as $r \rightarrow 0$

$KE = 0$ at $v_p=0$ and $v_e=0$

So, lowest E is - ∞ when electron is at nucleus!

Alternatively, why doesn't atom radiate EM waves?

The Person Behind The Science

Niels Bohr

1885-1962

Highlights

- Worked with J.J. Thomson (1911)
- Awarded Nobel Prize in Physics (1922)

Moments in a Life

- Received his doctorate degree from Copenhagen University (1911)
- Professor of Theoretical Physics at Copenhagen University (1916)

A Timeline of the Atom

← **400 BC …… 0 ……. 1800 1850 1900 1950**

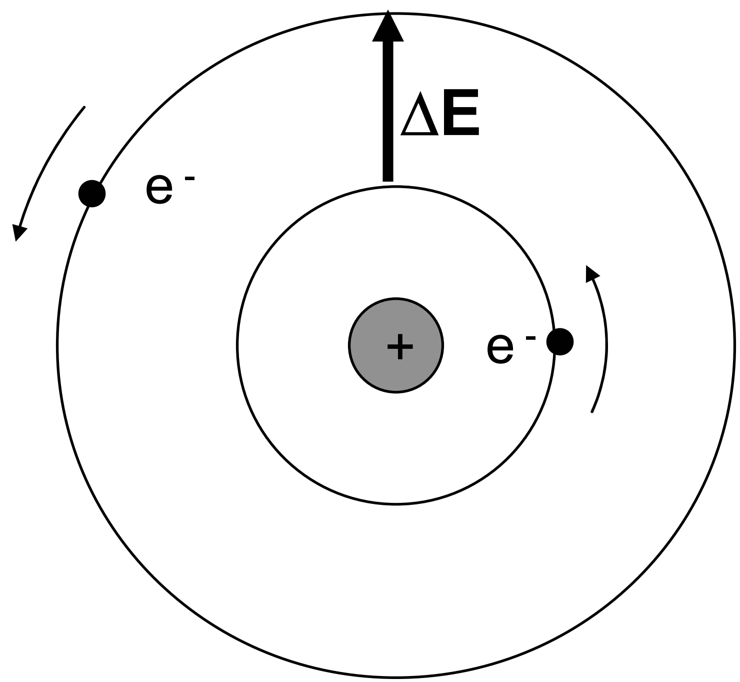
-450 Empedocles- earth, air, fire, water
-400 Democritus: idea of an atom; is water like sand: smooth from afar, coarse up close? atoms had no mass, just filled space
1661 Robert Boyle: Disputes 4 element theory; postulates an element is a substance that can not be reduced into simpler substances
1808 John Dalton postulates that atoms have weight and combine to form substances
1811 Amadeo Avogadro; Postulates that compounds are formed from molecules.
1820 Faraday: charge/mass ratio of protons
1885 E. Goldstein: discovers a positively charged sub-atomic particle
1898 J. J. Thompson finds a negatively charged particle called an electron.
1909 Robert Millikan experiments to find the charge and mass of the electron.
1911 Ernest Rutherford discovers the nucleus of an atom.
1913 Niels Bohr introduces his atomic theory.
1919 The positively charged particle identified by Goldstein is found to be a proton.
1920s Heisenburg, de Broglie, and Schrodinger.
1932 James Chadwick finds the neutron.
1964 The Up, Down, and Strange quark are discovered.
1974 The Charm quark is discovered.
1977 The Bottom quark is discovered.
1995 The Top (and final) quark is discovered.

Bohr's Theory

1. Atoms have well-defined electron orbits.
2. They don't radiate.
3. Circular orbits: only specific orbits with specific angular momenta, l = n( ), are allowed {quantization postulate}
4. Transitions in energy: electrons go from one orbit to the next

$$l = n\frac{h}{2\pi}$$

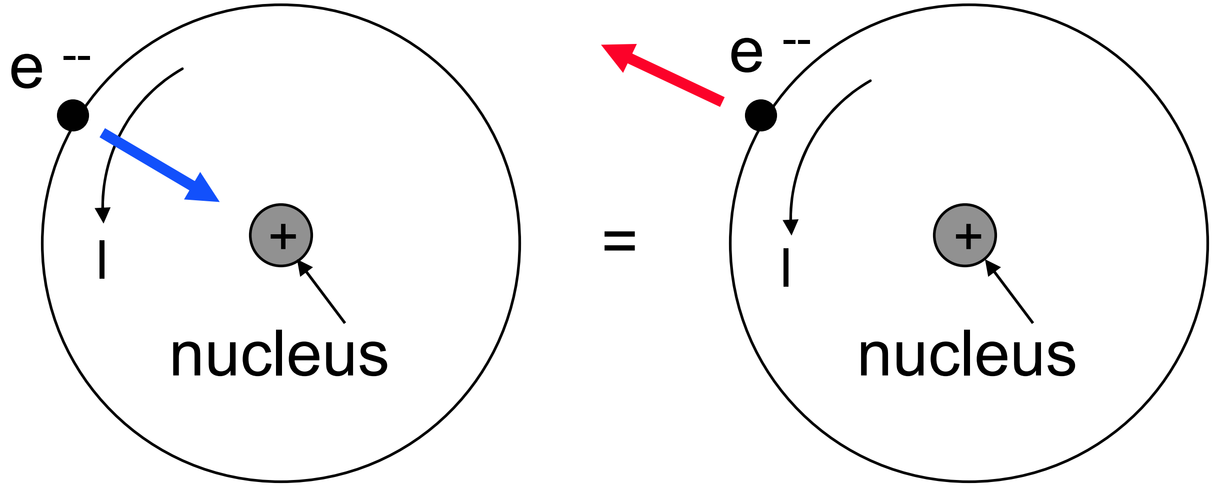
$$n \in \mathbf{Z}$$



Balance of forces within an atom

$$F_{coulombic} = F_{centrifugal}$$

$$\frac{Ze^2}{4\pi\varepsilon_0 r_n^2} = \frac{mv^2}{r_n}$$



(Z = nuclear charge)

So, if we know v, we know r_n

Angular momentum: solve for **v**

$$l = m\,v\,r_n = \frac{n\,h}{2\,\pi} \longrightarrow v = \frac{n\,h}{2\,\pi\,m\,r_n}$$

since we also have: $$\frac{Ze^2}{4\pi\varepsilon_0\,r_n^{\,2}} = \frac{m\,v^2}{r_n}$$

we can substitute for **v** and solve for $\mathbf{r_n}$:

$$\frac{Ze^2}{4\pi\varepsilon_0\,r_n^{\,2}} = \frac{mn^2h^2}{4\pi^2m^2r_n^{\,3}} \Rightarrow r_n = \frac{n^2h^2\varepsilon_0}{\pi mZe^2} = \frac{n^2a_0}{Z}$$

where $\mathbf{a_o}$ is the Bohr radius:

$$a_o = \frac{h^2\,\varepsilon_o}{\pi\,m\,e^2} = 0.529\ \text{Å}$$

Radii of Hydrogen Atom Orbitals

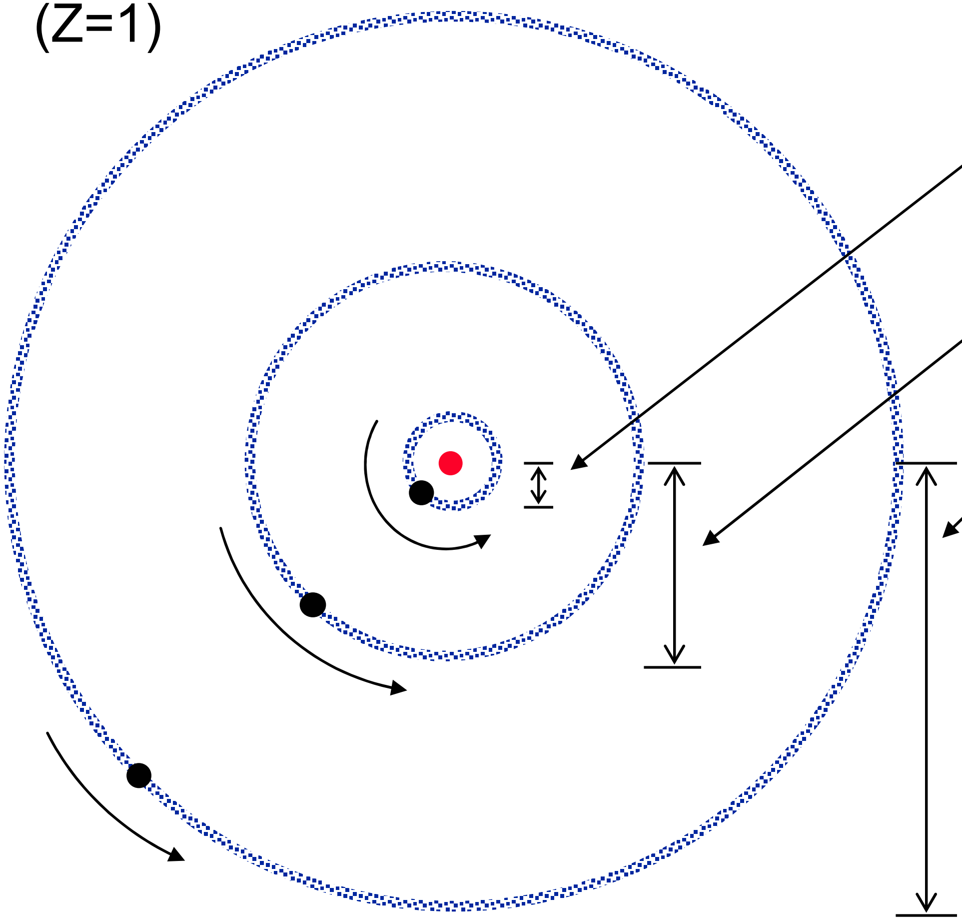
$r_n = n^2 a_0$

(Z=1)

$n = 1, \quad r = 1^2\,(0.529) = 0.529$ Å

$n = 2, \quad r = 2^2\,(0.529) = 2.116$ Å

$n = 3, \quad r = 3^2\,(0.529) = 4.761$ Å



Orbital Energies:

E = kinetic energy + potential energy

$$= \frac{1}{2}mv^2 - \frac{Ze^2}{4\pi\varepsilon_o r_n} \qquad \text{(Z = nuclear charge)}$$

$$= \left(\frac{e^2}{2r_n} - \frac{e^2}{r_n}\right)\left(\frac{Z}{4\pi\varepsilon_o}\right) \qquad \left(\text{since} \quad \frac{Ze^2}{4\pi\varepsilon_o r_n^2} = \frac{mv^2}{r_n}\right)$$

$$= \frac{-e^2}{2r_n}\left(\frac{Z}{4\pi\varepsilon_o}\right)$$

$$= \frac{-Z^2me^4}{8n^2h^2\varepsilon_o^2} \qquad \left(\text{since} \quad r_n = \frac{n^2h^2\varepsilon_o}{\pi mZe^2}\right)$$

$$= \frac{Z^2 \times \text{constant}}{n^2}$$

constant = -13.6 eV

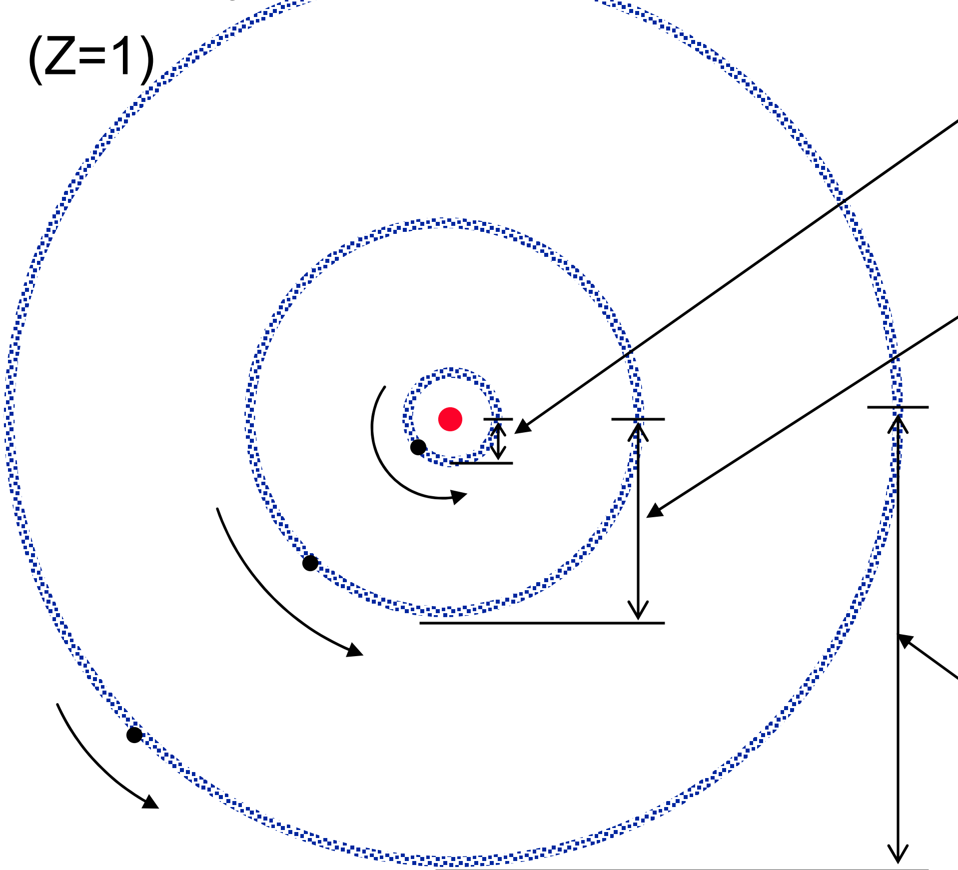
Z = 1 for hydrogen

Energies of Hydrogen Atom Orbitals

$$E_n = \frac{-13.6}{n^2}\ [eV]$$

$r_n = n^2 a_0$

(Z=1)



$n = 1,\ \ r = 1^2(0.529) = 0.529\ Å$

$$E = \frac{-13.6}{n^2} = \mathbf{-13.6\ eV}$$

$n = 2,\ \ r = 2^2(0.529) = 2.116\ Å$

$$E = \frac{-13.6}{n^2} = \mathbf{-3.4\ eV}$$

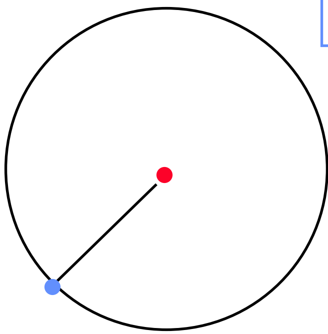
$n = 3,\ \ r = 3^2(0.529) = 4.761\ Å$

$$E = \frac{-13.6}{n^2} = \mathbf{-1.5\ eV}$$

Radii and Energies with Differing Nuclear Charge

$$r = \frac{n^2 a_0}{Z} \qquad E = \frac{-13.6\ Z^2}{n^2}$$

H

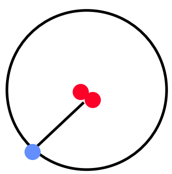


n = 1 , Z = 1

$$r = \frac{1^2 a_0}{1} = \mathbf{0.529\ \text{Å}}$$

$$E = \frac{-13.6\ (1^2)}{1^2} = \mathbf{-13.6\ eV}$$

He$^+$

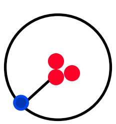


n = 1 , Z = 2

$$r = \frac{1^2 a_0}{2} = \mathbf{0.265\ \text{Å}}$$

$$E = \frac{-13.6\ (2^2)}{1^2} = \mathbf{-54.4\ eV}$$

Li^{2+}



n = 1 , Z = 3

$$r = \frac{1^2 a_0}{3} = \mathbf{0.176\ \text{Å}}$$

$$E = \frac{-13.6\ (3^2)}{1^2} = \mathbf{-122\ eV}$$

The Rydberg Constant:

As measured: $R_H = 109\,677.6\ cm^{-1}$ (correct value)

As predicted: $R_H = 109\,737.8\ cm^{-1}$

In order to obtain a more accurate predicted value for $\mathbf{R_H}$ the reduced mass must be used:

$$m_{reduced} = \frac{m_e m_p}{m_e + m_p}$$

(to account for the fact that a proton has mass also)

Success of the Theory

- Also Gets All Other 1e$^-$ Atoms Right!
- So Bohr Gets the H atom Right!
- Go to Stockholm, Niels!

Failures of the Theory

- **So Bohr Gets the H atom Right!**
- **Also Gets All Other 1e$^-$ Atoms Right!**
- **Go to Stockholm, Niels!**

BUT

There are two major flaws:

(a) Can't explain any multi-electron atoms!

(b) Is completely wrong, according to quantum mechanics!

End

Atoms

and the

Bohr Model

Reading: Gray: (1-1) to (1-7)
OGN: (15.1) and (15.4)