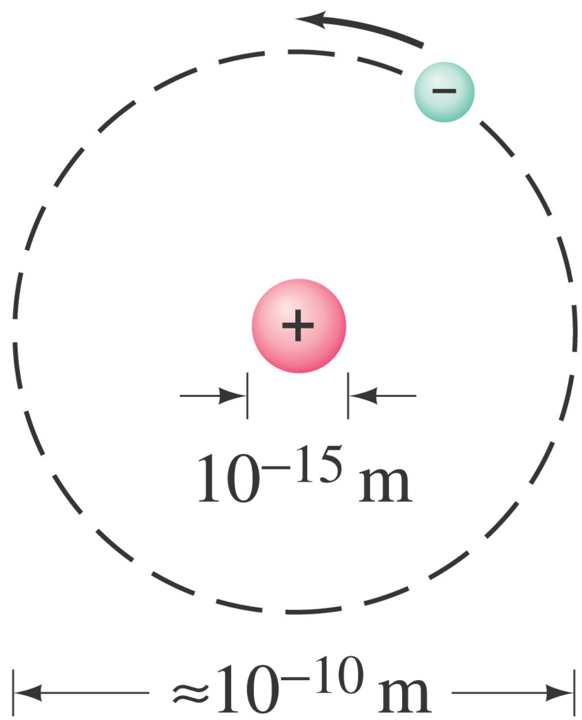
Chapter 27

Early Quantum Theory and Models of the Atom



"Modern" Physics

19th century physics had so many great successes, that most people felt nature was almost fully understood.

Just a few small "clouds on the horizon" remained. Minor details to be worked out, a handful of odd experiments to be explained.

But, as the 20th Century dawned, the entire foundation of science was about to be completely rebuilt.

Relativity and Quantum theory provided a new and deeper understanding of the nature of reality.

The modern world would not exist without it.

Think about how these fields advanced in the last 100 yrs compared to the previous 100,000 years: Medicine, Chemistry, Biology, Astronomy, Technology

Quantum Theory

Key Points:

- → The world is made of atoms.
- → Atomic nature of matter tied to character of EM radiation.
- → wave-particle Duality.
- → Quantum Theory stems entirely from observations!

1880s Wien's Law links blackbody spectrum (λ_{max}) to thermodynamic temperature

1890s Plank's Blackbody radiation Law (Solves UV Catastrophe)

introduced radical view of what atoms "are".

Discovery of electron (Thomson, Millikan) charge is quantized.

1900s photoelectric Effect → Shows light emitted/absorbed

1905 $E = hf$ only in discrete units → Photons.

AND Photon Energy related to frequency

1911 Rutherford discovers the Nucleus.

1913 Quantization of radiation suggests solution to the problem of impossibility of stable orbits for electrons. Bohr's quantum atom model successful at explaining absorption/emission lines in spectrum of Hydrogen. (by assuming orbital angular motion is quantized) However, no justification is apparent, but it works.

1924 DeBroglie argues for symmetry in nature; if light has particle-like properties, matter could have wave-like properties.

$\lambda_D = h/p$ Predicted wavelength for electron — Correct!

With the electron viewed as a wave, only closed 'modes' can exist for an electron-wave inside an atom

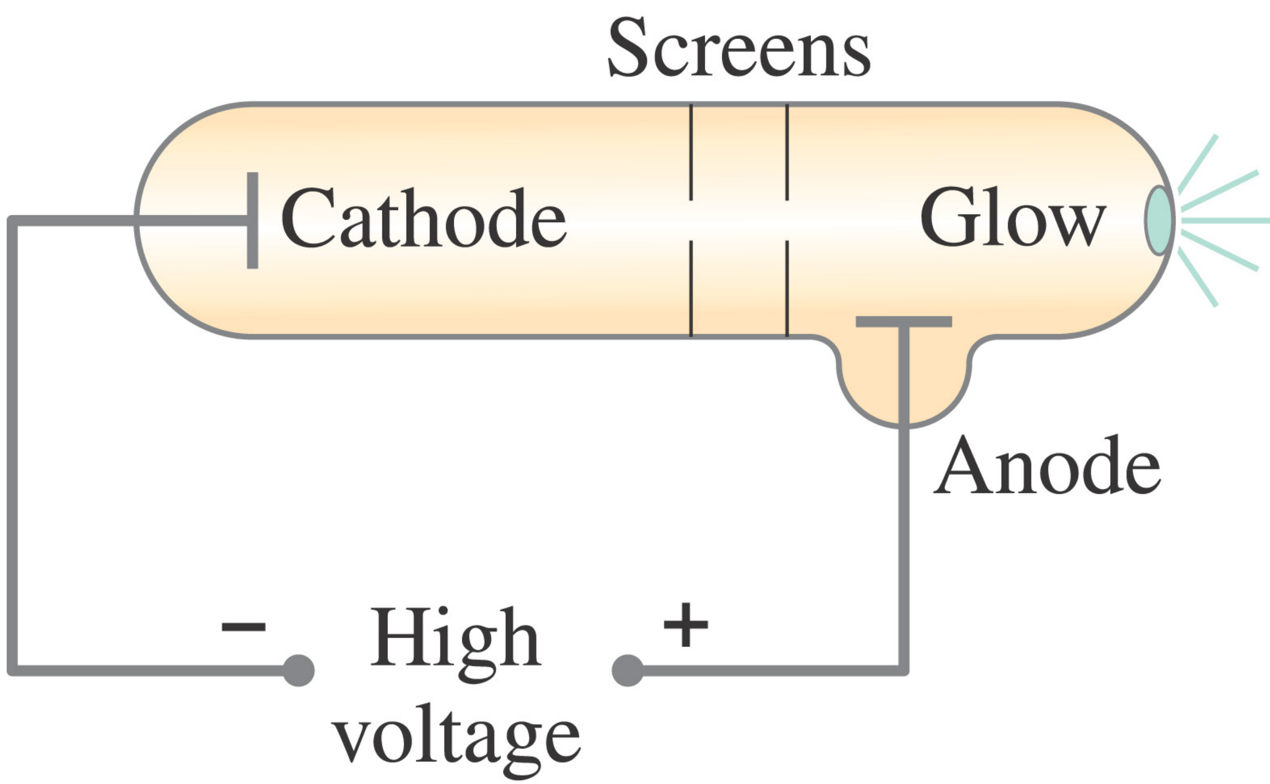
→ Finally provides physical justification for quantum idea of atomic structure / behavior.

Units of Chapter 27

- Discovery and Properties of the Electron
- Planck's Quantum Hypothesis; Blackbody Radiation
- Photon Theory of Light and the Photoelectric Effect
- Energy, Mass, and Momentum of a Photon
- Photon Interactions; Pair Production
- Wave-Particle Duality; the Principle of Complementarity
- Wave Nature of Matter
- Electron Microscopes
- Early Models of the Atom
- Atomic Spectra: Key to the Structure of the Atom
- The Bohr Model
- de Broglie's Hypothesis Applied to Atoms

27.1 Discovery and Properties of the Electron

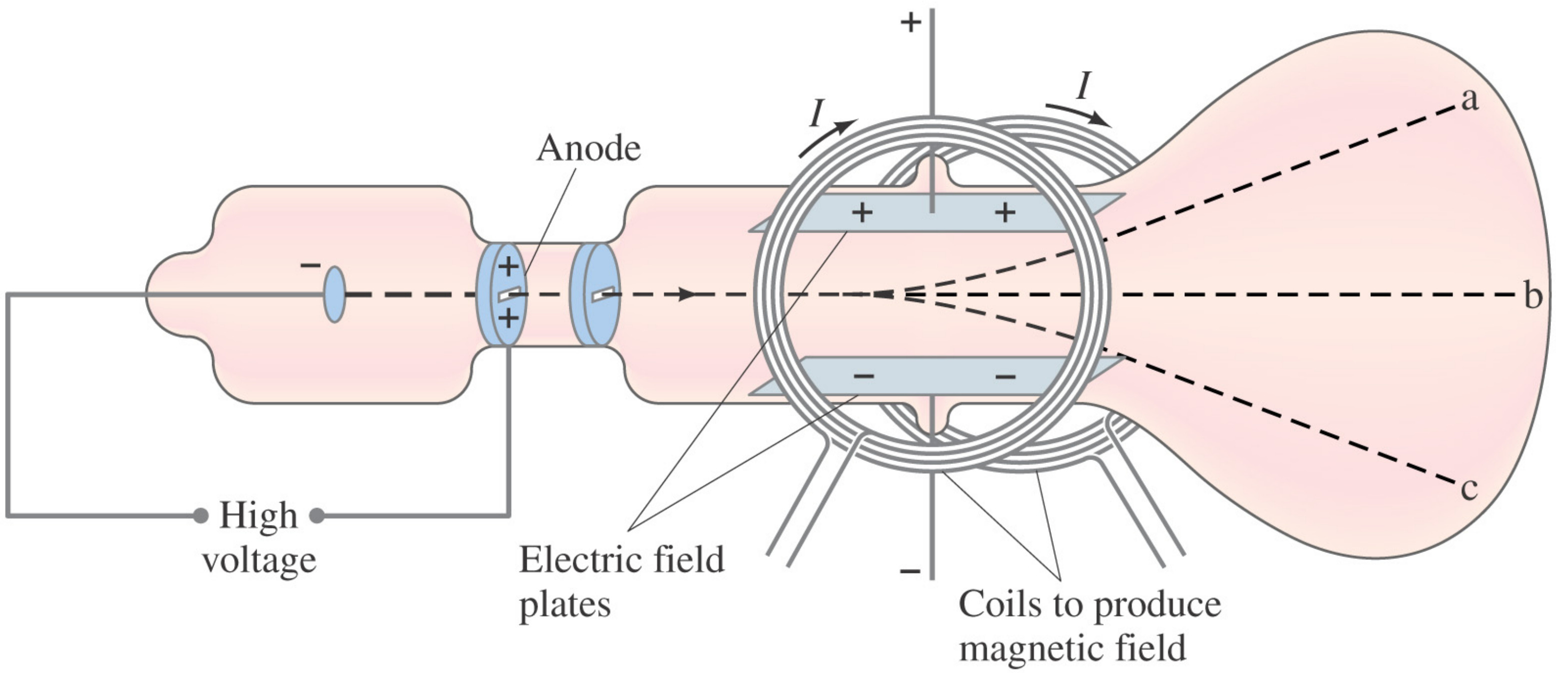
In the late 19th century, discharge tubes were made that emitted "cathode rays."



27.1 J.J Thomson's 1890 apparatus to measure the charge/mass ratio (e/m)

- Cathode rays could be deflected by electric or magnetic fields.
- Measuring the angle of deflection gave e/m

- But finding e and m separately proved elusive



Doesn't this look a lot like a TV tube?

27.1 Discovery and Properties of the Electron

By accelerating cathode rays through a known potential and then measuring the radius of their curved path in a known magnetic field, the charge to mass ratio could be measured:

$$\frac{e}{m} = \frac{E}{B^2 r} \qquad (27\text{-}1)$$

The result is $e/m = 1.76 \times 10^{11}\ \mathrm{C/kg}$

Thomson's result showed that charge is quantized, and is carried by a particle with a small mass. The Electron is born!

27.1 The Charge of the Electron

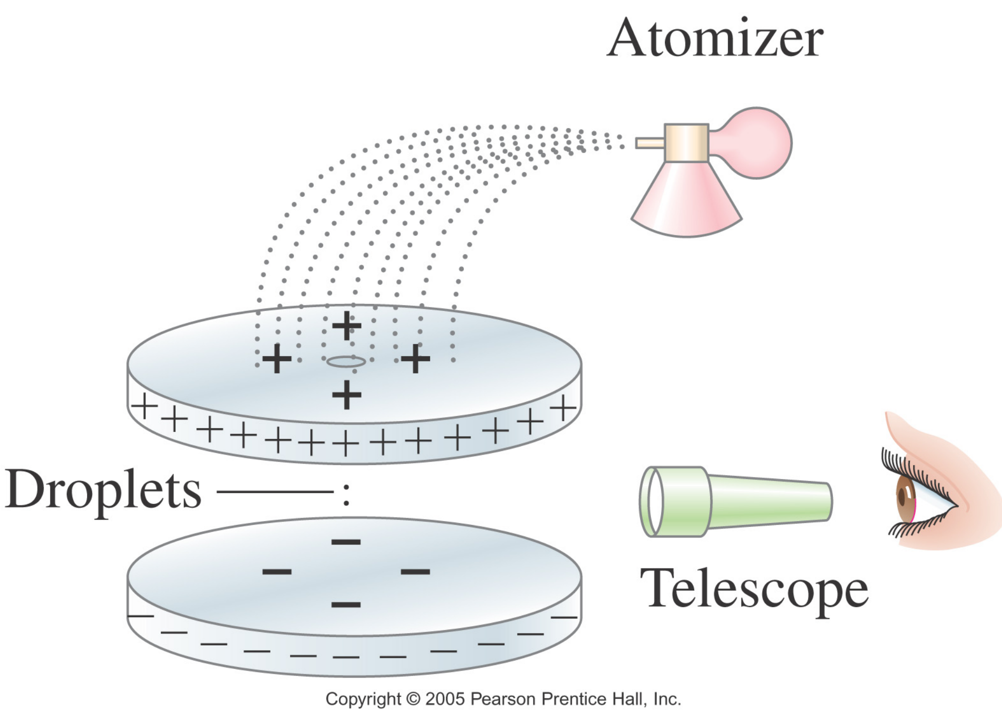
Cathode rays were soon called electrons.

Millikan devised an experiment to measure the charge on the electron by measuring the electric field needed to suspend an oil droplet of known mass between parallel plates.

Millikan's expt. was exactly like the HW problem in Chapter 16 - So you all know how to do it!

27.1 Discovery and Properties of the Electron

The mass and charge of each oil droplet were measured; careful analysis of the data showed that the charge was always an integral multiple of a smallest charge, e.



$$e = 1.602 \times 10^{-19}\ \mathrm{C}$$

Knowing e allows the electron mass to be calculated:

$$m_e = 9.11 \times 10^{-31}\ \mathrm{kg}$$

27.1 Discovery and Properties of the Electron

The currently accepted value of e is:

$$e = 1.602 \times 10^{-19}\,\mathrm{C}$$

Knowing e allows the electron mass to be calculated:

$$m_{\mathrm{e}} = 9.11 \times 10^{-31}\,\mathrm{kg}$$

Radiation and Temperature

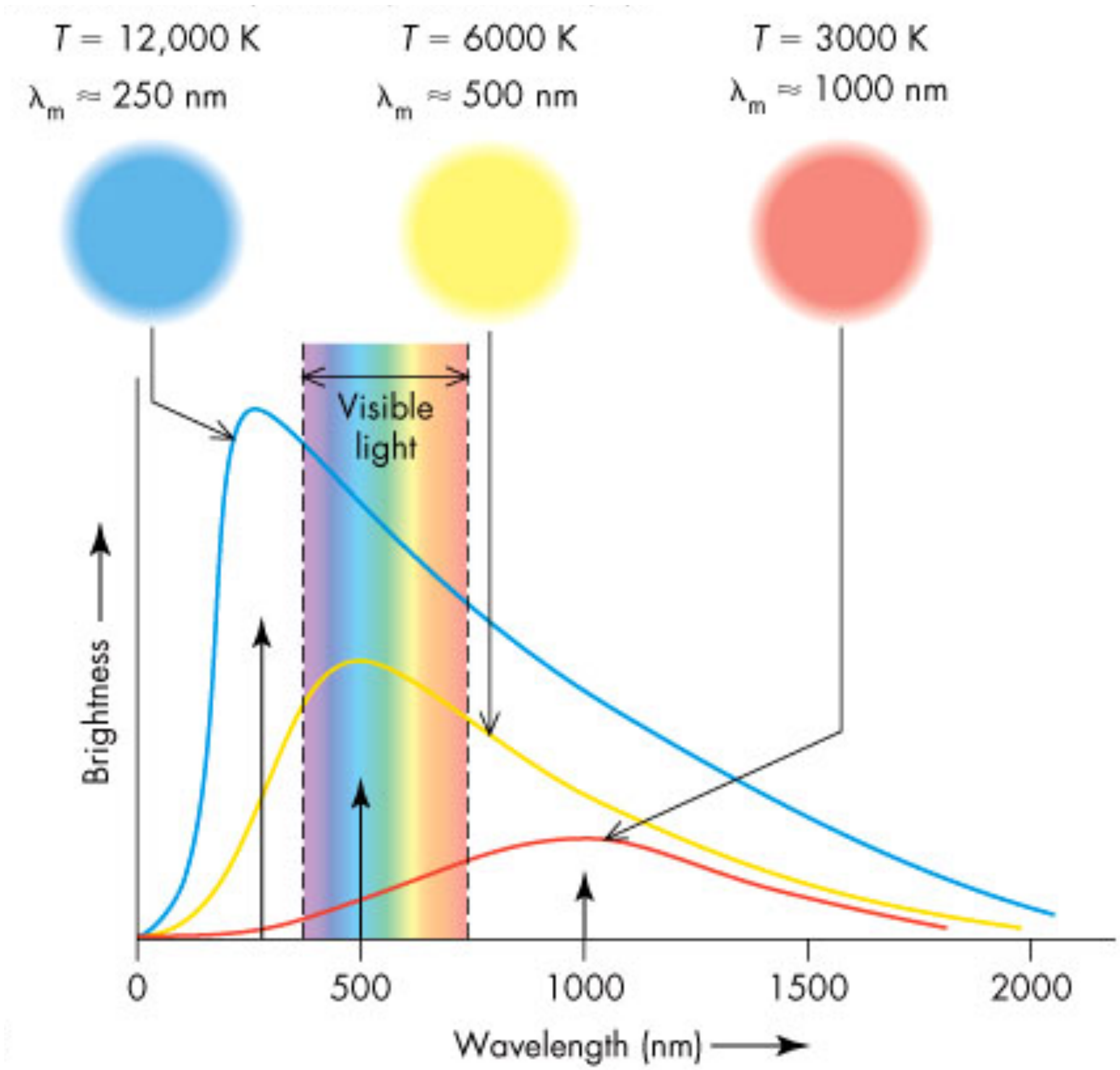
Heated bodies generally radiate across the entire electromagnetic spectrum

There is one particular wavelength, λ_m, at which the radiation is most intense and is given by

Wien's Law:

$\lambda_m = k/T$

Where k is some constant and T is the temperature of the body. For SI units, $k=2.9\times10^{-3}$ m.K



Temperature of the Sun

From the observation that the sun is Yellow (λ_{max}=500nm), what is its surface temperature ?

A. 100 C
B. 5800 K
C. 293 K
D. 30,000 K

27.2 Planck's Quantum Hypothesis; Blackbody Radiation

This spectrum could not be reproduced using 19th-century physics.

All attempts were either purely descriptive (like Wien's law), or led to the so called "ultraviolet catastrophe" - failure to derive the decline in flux at short wavelengths, leading to unrealistic prediction of limitless radiation in the UV.

A solution was proposed by Max Planck in 1900:

The atoms are all radiating, absorbing and redistributing energy between themselves. Each behaves as a harmonic oscillator with discrete modes

The distribution of atomic oscillator energies leads to the black-body spectrum

The oscillations within atoms can only occur at discrete frequencies that are multiples of a certain minimum value. (like modes on a string, or in a pipe)

Introduced a new fundamental constant of nature: Plank's Constant: h

Many ***including Plank*** viewed this model as purely a mathematical device

Today we view it as the first physically correct model in quantum physics!

27.2 Planck's Quantum Hypothesis; Blackbody Radiation

Planck found the value of his constant by fitting blackbody curves:

$$h = 6.626 \times 10^{-34}\,\mathrm{J \cdot s}$$

Planck's proposal was that the energy of an oscillation had to be an integer multiple of hf. This is called the quantization of energy.

Some scientists today believe time and space are also quantized, not just energy and charge.

27.3 Photon Theory of Light and the Photoelectric Effect

Einstein (1905) suggested that, given the success of Planck's theory, light must be emitted in small energy packets:

$$E = hf \qquad (27\text{-}4)$$

These tiny packets, or particles, are called photons.

This theory won Einstein the Nobel Prize

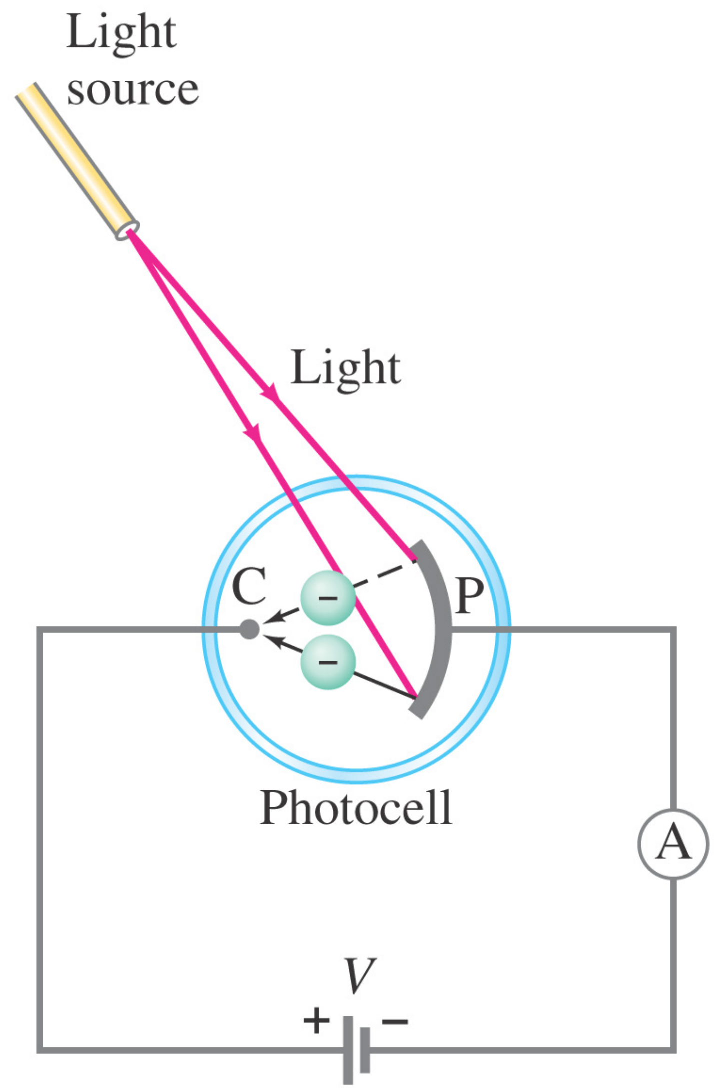
He had resurrected Newton's 400 yr old idea of light as a stream of particles

27.3 Photon Theory of Light and the Photoelectric Effect

The photoelectric effect: If light strikes a metal, electrons are emitted.

But….

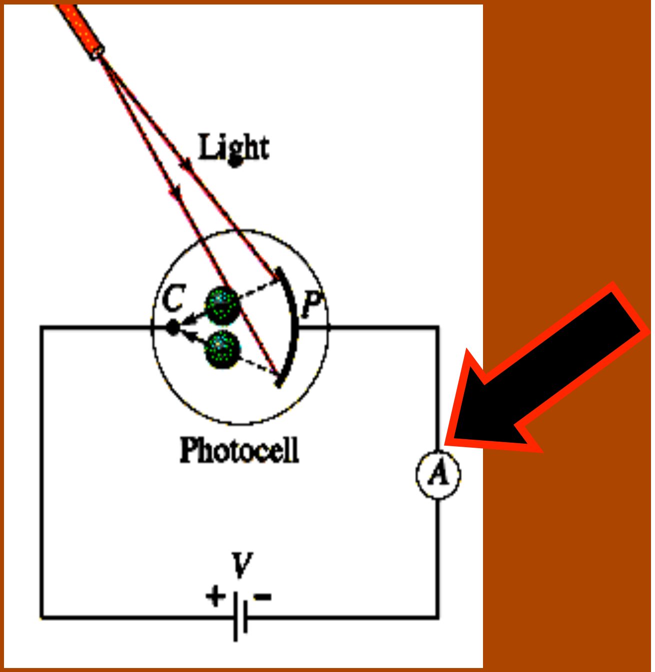
- The effect does not occur if the frequency of the light is too low;
- The kinetic energy of the electrons increases with frequency.



ConcepTest 27.2e Photoelectric Effect V

A photocell is illuminated with light with a frequency above the cutoff frequency. The magnitude of the current produced depends on:

1) wavelength of the light
2) intensity of the light
3) frequency of the light
4) all of the above
5) none of the above



ConcepTest 27.2e Photoelectric Effect V

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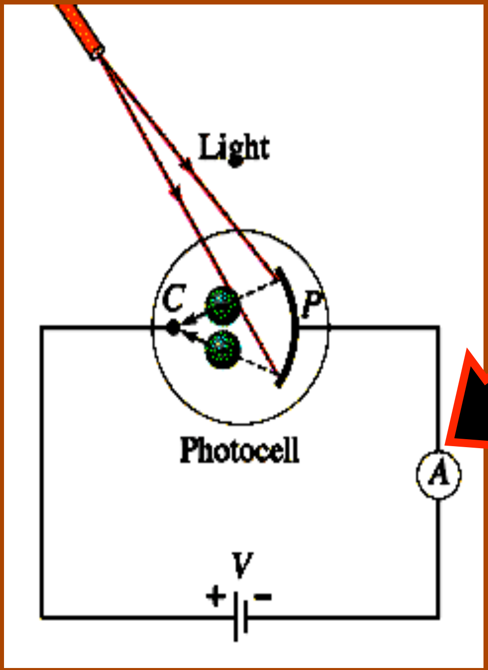
1) wavelength of the light
2) intensity of the light
3) frequency of the light
4) all of the above
5) none of the above

Each photon can only knock out one electron. So to increase the current, we would have to knock out **more electrons**, which means we need **more photons**, which means we need a **greater intensity**!

Changing the **frequency** or **wavelength** will change the **energy** of each electron, but we are interested in the **number of electrons** in this case.



27.3 Photon Theory of Light and the Photoelectric Effect

If light is a wave, theory predicts:

1. **Number of electrons and their energy should increase with intensity**

2. **Frequency would not matter**

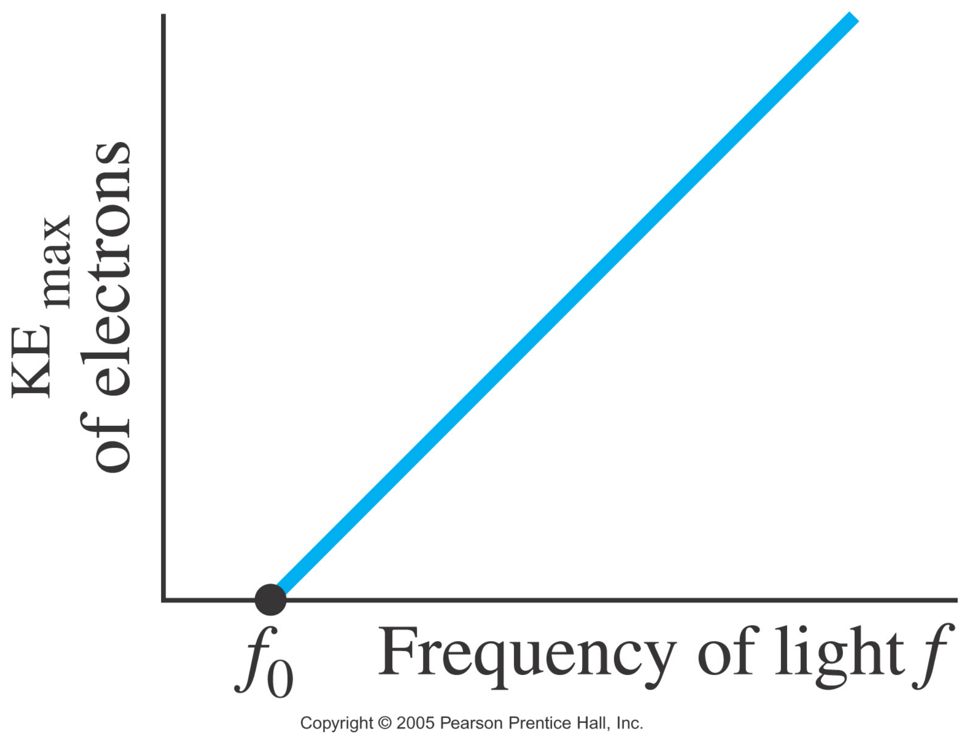
27.3 Photon Theory of Light and the Photoelectric Effect

If light is a stream of particles, one predicts:

- **Increasing intensity increases number of electrons but not energy**
- **Above a minimum energy required to break atomic bond, kinetic energy will increase linearly with frequency**
- **There is a cutoff frequency below which no electrons will be emitted, regardless of intensity**

27.3 Photon Theory of Light and the Photoelectric Effect

The photon theory assumes that each electron absorbs a single photon. Plotting the kinetic energy vs. frequency:



F_0 = "cutoff frequency"

The lowest frequency that will free any electrons at all

W = "work function"

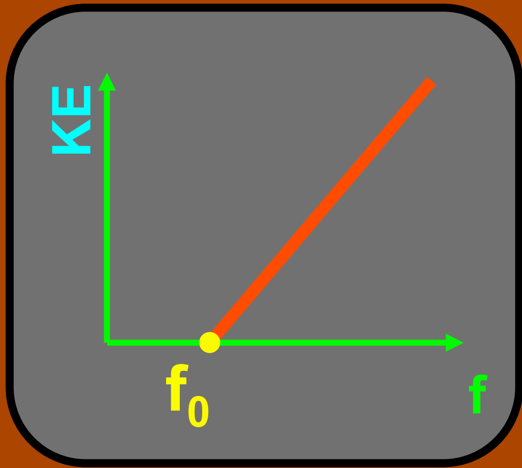
The amount of energy (work) required to free each electron from the metal

This result shows clear agreement with the photon theory, and not with wave theory.

ConcepTest 27.2a Photoelectric Effect I

If the cutoff frequency for light in the photoelectric effect for metal B is greater than that of metal A. Which metal has a greater work function?

1) metal A
2) metal B
3) same for both
4) W_0 must be zero for one of the metals

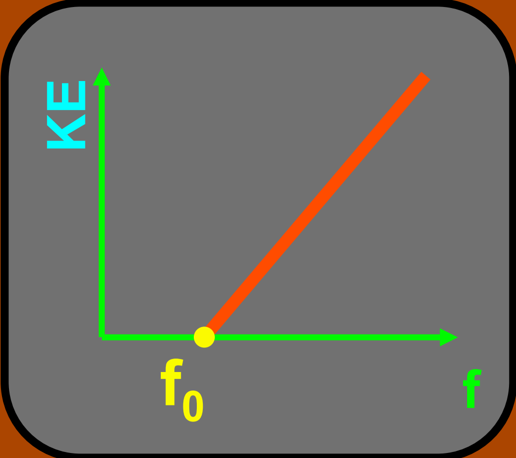


ConcepTest 27.2a Photoelectric Effect I

If the cutoff frequency for light in the photoelectric effect for metal B is greater than that of metal A. Which metal has a greater work function?

1) metal A
2) metal B
3) same for both
4) W_0 must be zero for one of the metals

A greater cutoff frequency means a higher energy is needed to knock out the electron. But this implies that the work function is greater, since the work function is defined as the minimum amount of energy needed to eject an electron.



Follow-up: What would you expect to happen to the work function of a metal if the metal was heated up?

ConcepTest 27.2b Photoelectric Effect II

A metal surface with a work function of $W_0 = hc/550$ nm is struck with blue light and electrons are released. If the blue light is replaced by red light of the same intensity, what is the result?

1) emitted electrons are more energetic
2) emitted electrons are less energetic
3) more electrons are emitted in a given time interval
4) fewer electrons are emitted in a given time interval
5) no electrons are emitted

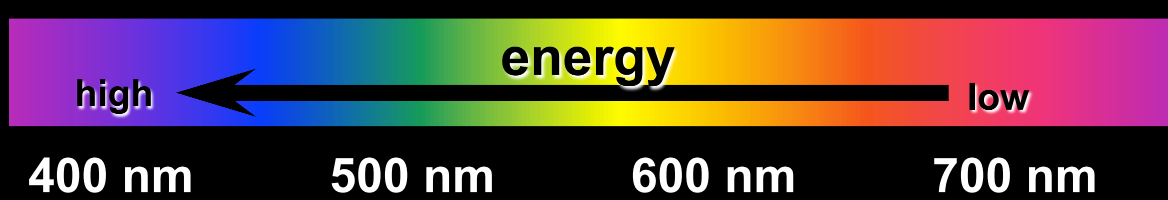
ConcepTest 27.2b Photoelectric Effect II

A metal surface with a work function of $W_0 = hc/550$ nm is struck with blue light and electrons are released. If the blue light is replaced by red light of the same intensity, what is the result?

1) emitted electrons are more energetic
2) emitted electrons are less energetic
3) more electrons are emitted in a given time interval
4) fewer electrons are emitted in a given time interval
5) no electrons are emitted

Red light has a wavelength of about 700 nm. The cutoff wavelength is 550 nm (yellow light), which is the maximum wavelength to knock out electrons. Thus, no electrons are knocked out.

$$E = hc / \lambda$$



ConcepTest 27.2c Photoelectric Effect III

A metal surface is struck with light of $\lambda = 400$ nm, releasing a stream of electrons. If the 400 nm light is replaced by $\lambda = 300$ nm light of the same intensity, what is the result?

1) more electrons are emitted in a given time interval
2) fewer electrons are emitted in a given time interval
3) emitted electrons are more energetic
4) emitted electrons are less energetic
5) none of the above

ConcepTest 27.2c Photoelectric Effect III

A metal surface is struck with light of λ = 400 nm, releasing a stream of electrons. If the 400 nm light is replaced by λ = 300 nm light of the same intensity, what is the result?

1) more electrons are emitted in a given time interval
2) fewer electrons are emitted in a given time interval
3) emitted electrons are more energetic
4) emitted electrons are less energetic
5) none of the above

A **reduced wavelength** means a **higher frequency**, which in turn means a **higher energy**. So the emitted electrons will be ***more energetic***, since they are now being hit with higher energy photons.

Remember that $c = f\lambda$ and that $E = hf$

ConcepTest 27.2d Photoelectric Effect IV

A metal surface is struck with light of λ = 400 nm, releasing a stream of electrons. If the light **intensity** is increased (**without changing λ**), what is the result?

1) more electrons are emitted in a given time interval
2) fewer electrons are emitted in a given time interval
3) emitted electrons are more energetic
4) emitted electrons are less energetic
5) none of the above

ConcepTest 27.2d Photoelectric Effect IV

A metal surface is struck with light of λ = 400 nm, releasing a stream of electrons. If the light intensity is increased (without changing λ), what is the result?

1) more electrons are emitted in a given time interval
2) fewer electrons are emitted in a given time interval
3) emitted electrons are more energetic
4) emitted electrons are less energetic
5) none of the above

A higher intensity means a more photons, which in turn means more electrons. On average, each photon knocks out one electron.

27.3 Photon Theory of Light and the Photoelectric Effect

Electronic light sensors

Digital Cameras

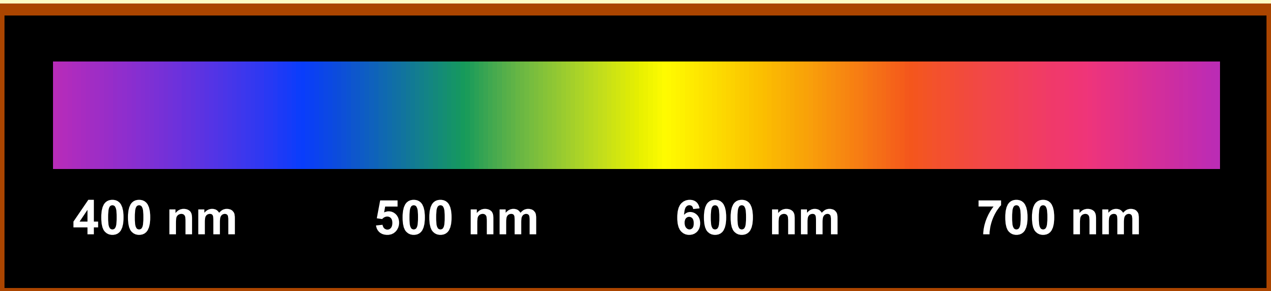
Photosynthesis

Sunburn

ConcepTest 27.1 Photons

Which has more energy, a photon of:

1) red light
2) yellow light
3) green light
4) blue light
5) all have the same energy

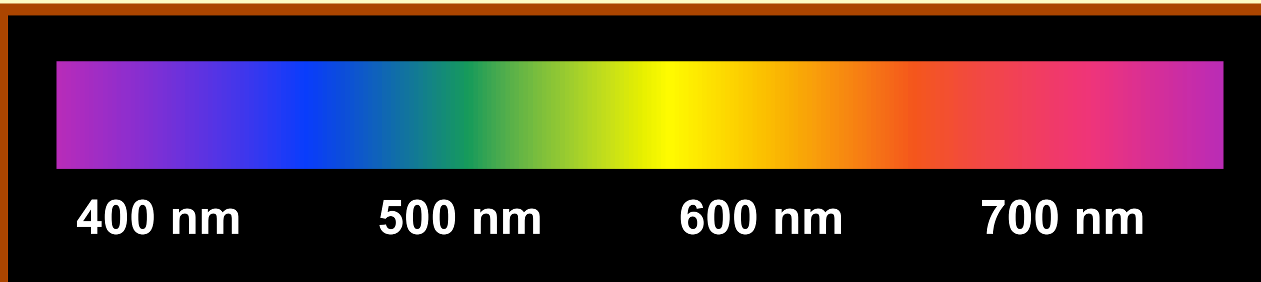


ConcepTest 27.1 Photons

Which has more energy, a photon of:

$$E = h\,f$$

1) red light
2) yellow light
3) green light
4) blue light
5) all have the same energy

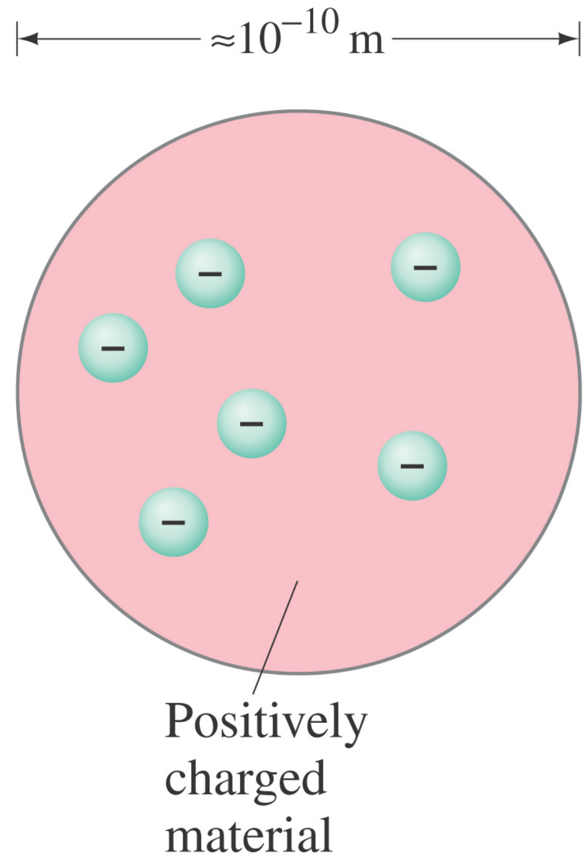


The photon with the **highest frequency** has the **most energy** because $E = hf = hc/\lambda$ (recall that $c = f\lambda$). So a **higher frequency** corresponds to a **lower wavelength**. The highest energy of the above choices is **blue**.

27.10 Early Models of the Atom

It was known that atoms were electrically neutral, but that they could become charged, implying that there were positive and negative charges and that some of them could be removed.

One popular atomic model was the "plum-pudding" model:



“Bullets deflected by tissue paper”

The discovery of the Nucleus

- Rutherford in 1911 did an experiment that showed that the positively charged nucleus must be extremely small compared to the rest of the atom, and contain essentially all of the mass.

- He directed alpha particles (emitted by some radioactive elements) at a thin metal foil and observed the scattering angle.

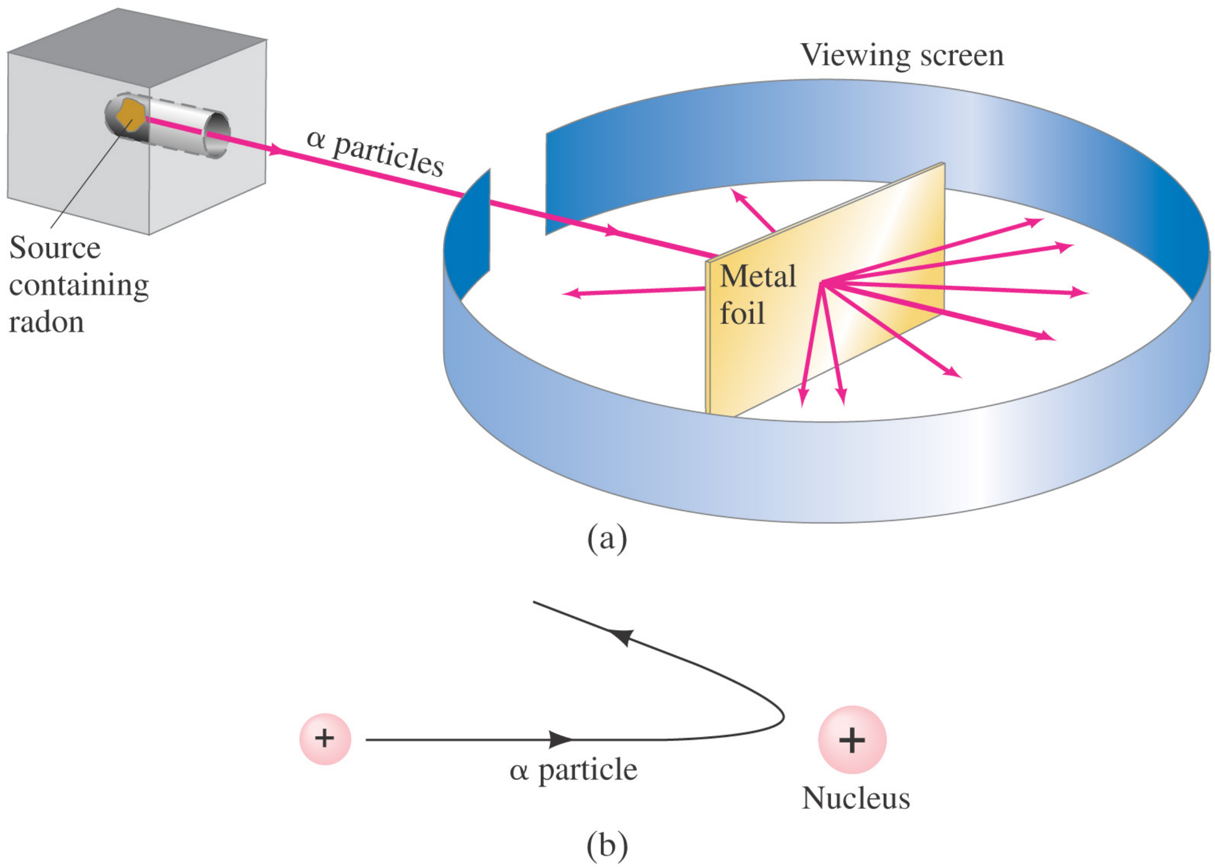
- Rutherford found that a tiny fraction of these “bullets” bounced right back!

- Meaning they had hit something massive, yet tiny.

- the atom was mostly empty space!

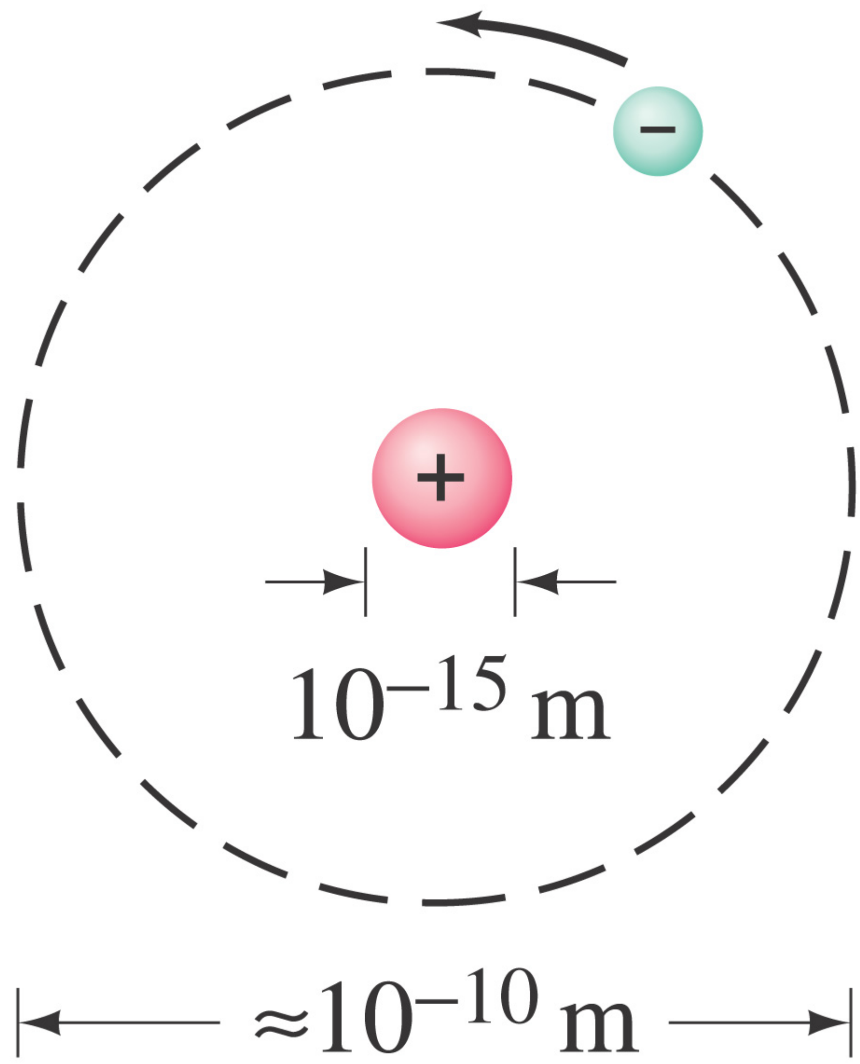
Rutherford’s Experiment

The only way to account for the large angles was to assume that all the positive charge was contained within a tiny volume – now we know that the radius of the nucleus is 1/10000 that of the atom.



Rutherford's Atom

- In Rutherford's model, the electron(s) orbit the nucleus like planets in the solar system
- Held in place by Coulomb force (instead of gravity)
- A BIG problem: Accelerated charges radiate EM radiation
- if the electron loses potential energy it falls closer to the nucleus…..
- So this electron would spiral into the nucleus immediately
- <u>Niels Bohr suggested the Quantum nature of light could make the atom stable</u>



What's wrong with Rutherford's solar system model for the atom?

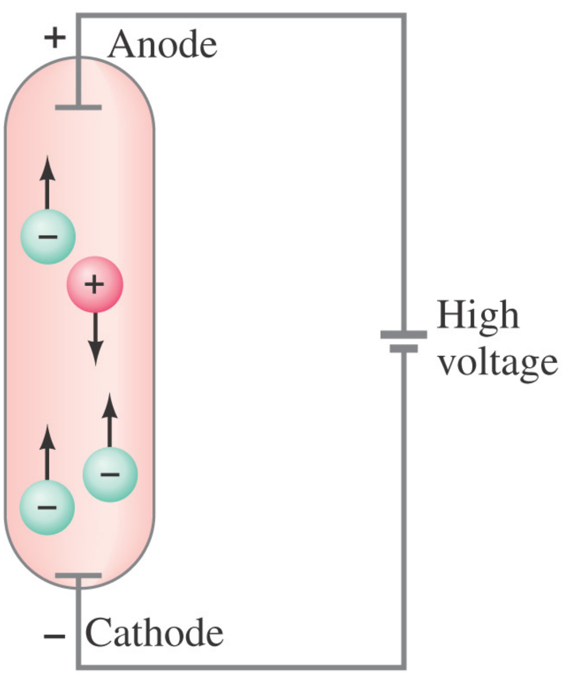
A. The nucleus is too small

B. Electrons appear to be moving

C. It shouldn't be stable

27.11 Atomic Spectra: Key to the Structure of the Atom

A very thin gas heated in a discharge tube emits light only at characteristic frequencies.



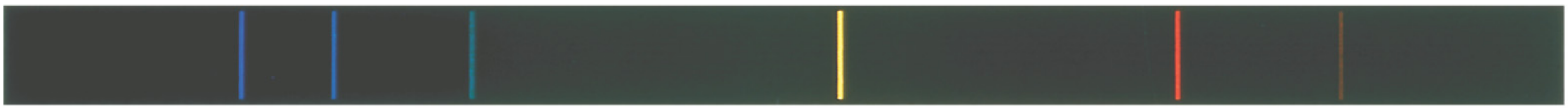
(a)

27.11 Atomic Spectra: Key to the Structure of the Atom

An atomic spectrum is a line spectrum – only certain frequencies appear.

If white light passes through such a gas, it absorbs at those same frequencies.

(a) Hydrogen



(b) Helium

The Sun

Each line is due to photons with a specific energy

What if these energies correspond to special "allowed" level inside the atom?

Rank these “colors” of light by their average photon energy (smallest to largest)

A. Red, Yellow, Blue, Violet

B. Violet, Blue, Yellow, Red

C. Blue, Yellow, Red, Violet

27.11 Atomic Spectra: Key to the Structure of the Atom

The wavelengths of electrons emitted from hydrogen have a regular pattern:

$$\frac{1}{\lambda} = R\left(\frac{1}{2^2} - \frac{1}{n^2}\right), \qquad n = 3, 4, \cdots$$

(27-9)

This is called the Balmer series. R is called the Rydberg constant:

$$R = 1.0974 \times 10^7 \text{ m}^{-1}$$

- A spectral lines occurs for each value of n.
- There are 5 visible lines in the Balmer series: $H\alpha$ (red), $H\beta$, $H\gamma$, $H\delta$, $H\varepsilon$
- This is a 'phenomenological' or 'empirical' theory -it is very accurate but explains nothing. The constant "R" is easily measured with a spectroscope (basically a prism or a diffraction grating)
- But we are still at "WHY?"

27.11 Atomic Spectra: Key to the Structure of the Atom

Other series include the Lyman series (in the Ultraviolet):

$$\frac{1}{\lambda} = R\left(\frac{1}{1^2} - \frac{1}{n^2}\right), \qquad n = 2, 3, \cdots$$

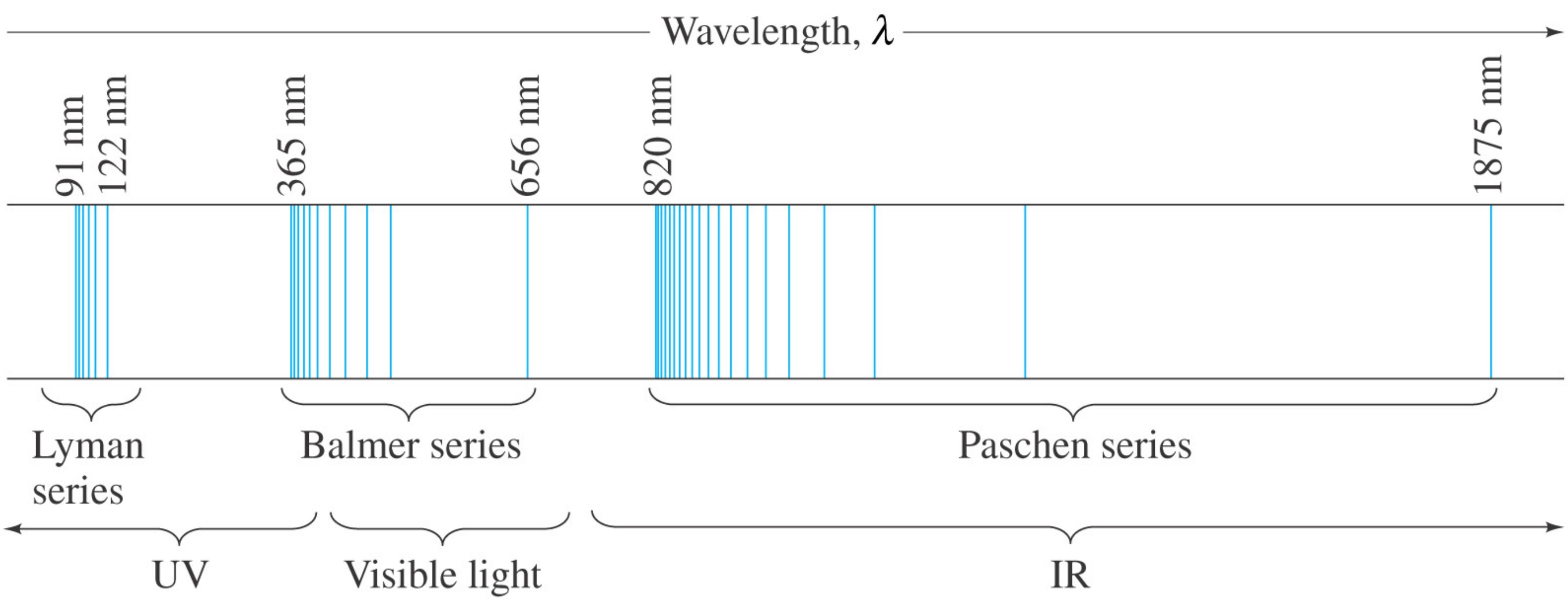
And the Paschen series (in the Infrared):

$$\frac{1}{\lambda} = R\left(\frac{1}{3^2} - \frac{1}{n^2}\right), \qquad n = 4, 5, \cdots$$

Note the series differ only in this number

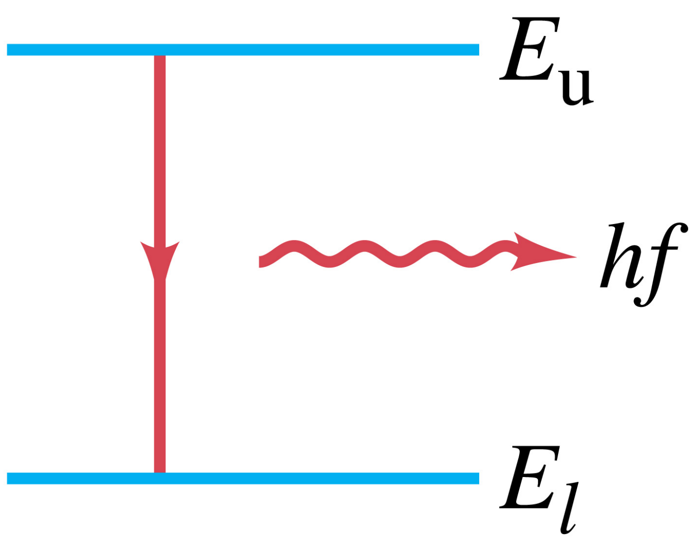
27.11 Atomic Spectra: Key to the Structure of the Atom

A portion of the complete spectrum of hydrogen is shown here. The lines cannot be explained by the Rutherford theory.



27.12 The Bohr Atom

Niels Bohr proposed that the possible energy states for atomic electrons were quantized – only certain values were possible. Then the spectrum could be explained as transitions from one level to another.



The wavelength of light emitted (or absorbed) corresponds to the Rydberg formula, where m and n are the lower and upper energy levels respectively

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

27.12 The Bohr Atom

After trying various ideas….

Bohr found that the angular momentum was quantized:

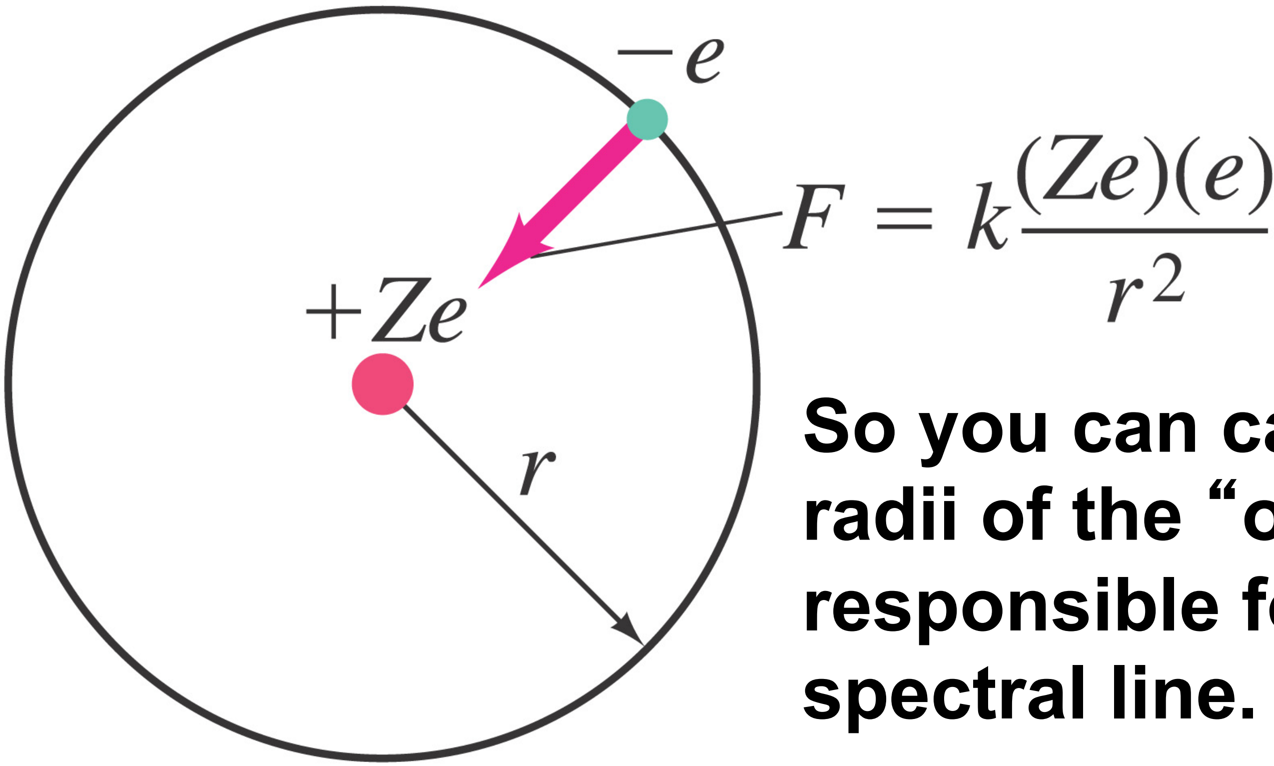
$$L = mvr_n = n\frac{h}{2\pi}, \qquad n = 1, 2, 3, \cdots \qquad \textbf{(27-11)}$$

It worked perfectly!

But there was still no physical motivation…..

27.12 The Bohr Atom

An electron is held in orbit by the Coulomb force:



So you can calculate the radii of the "orbit" responsible for each spectral line.

By equating centripetal force and coulomb force

27.12 The Bohr Atom

Using the Coulomb force, we can calculate the radii of the orbits:

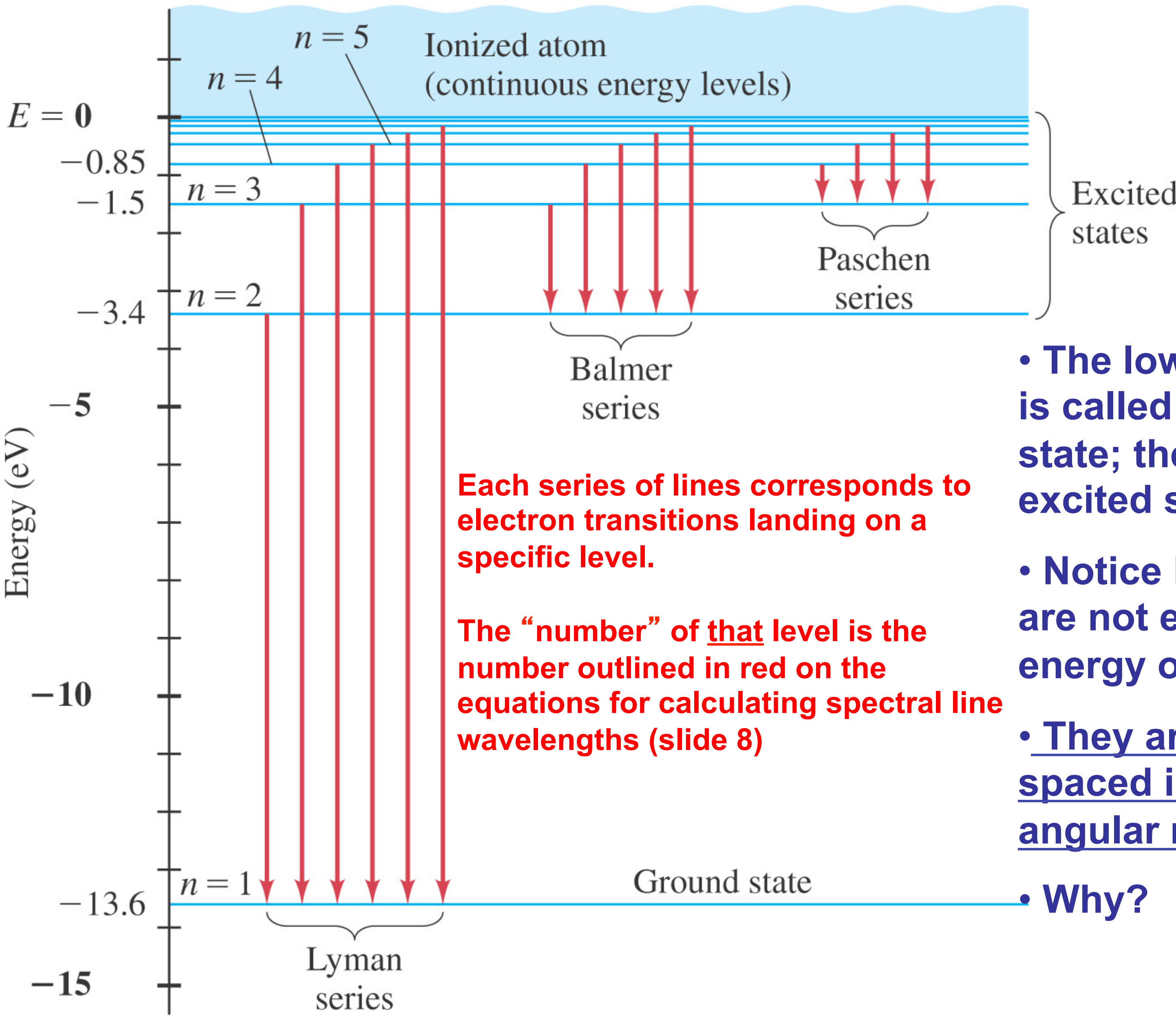
$$r_1 = \frac{h^2}{4\pi^2 mke^2} = 0.529 \times 10^{-10}\,\mathrm{m} \qquad \textbf{(27-13)}$$

$$r_n = \frac{n^2h^2}{4\pi^2 mkZe^2} = \frac{n^2}{Z} r_1 \qquad \textbf{(27-12)}$$

The equation above comes from equating the centripetal force needed to balance the coulomb force.   -a type of calculation you are all familiar with.

The resulting levels are not evenly spaced in radius- just like the spectral lines are not evenly spaced, but follow a distinct pattern

27.12 The Bohr Atom



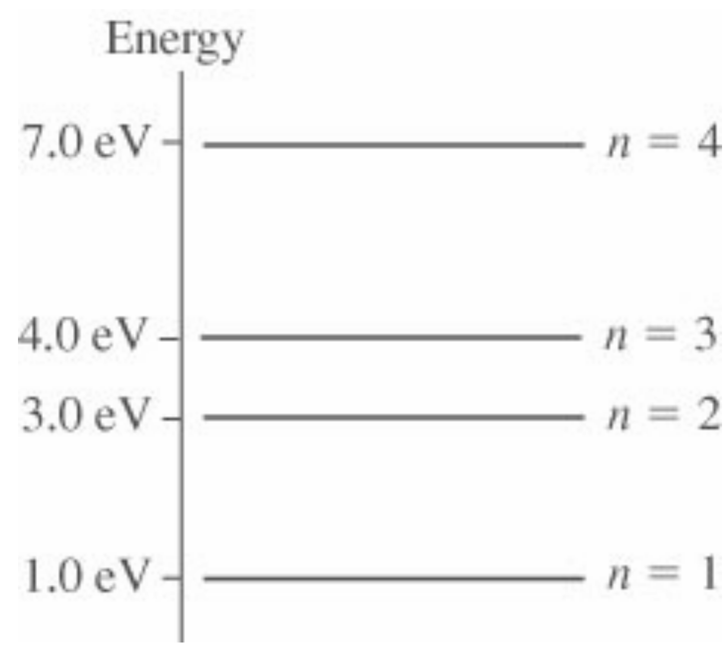
Each series of lines corresponds to electron transitions landing on a specific level.

The "number" of <u>that</u> level is the number outlined in red on the equations for calculating spectral line wavelengths (slide 8)

- **The lowest energy level is called the ground state; the others are excited states.**
- **Notice how the levels are not evenly spaced in energy or in radius.**
- **<u>They are only evenly spaced in terms of angular momentum</u>**
- **Why?**

What is the maximum photon energy in (in electron volts, eV) that could be emitted by the quantum system with the energy level diagram shown here?

A. 7.0
B. 6.0
C. 4.0
D. 3.0
E. 1.0



ConcepTest 27.4 Ionization

How much energy does it take to ionize a hydrogen atom in its ground state?

1) 0 eV
2) 13.6 eV
3) 41.2 eV
4) 54.4 eV
5) 108.8 eV

ConcepTest 27.4 Ionization

How much energy does it take to ionize a hydrogen atom in its ground state?

1) 0 eV
2) 13.6 eV
3) 41.2 eV
4) 54.4 eV
5) 108.8 eV

The energy of the ground state is the energy that binds the electron to the nucleus. Thus, an amount equal to this binding energy must be supplied in order to kick the electron out of the atom.

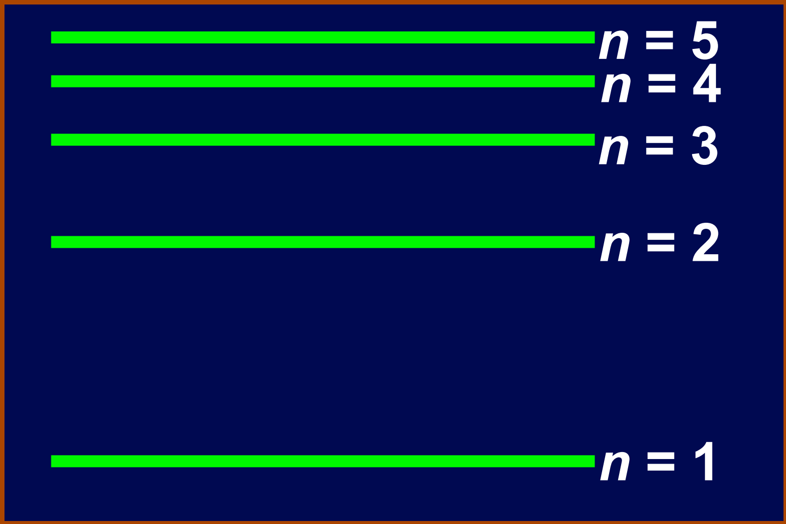
$$E_n = E_1 \frac{Z^2}{n^2}$$

Follow-up: How much energy does it take to change a He^+ ion into a He^{++} ion? Keep in mind that $Z = 2$ for helium.

ConcepTest 27.5a Atomic Transitions I

For the possible transitions shown, for which transition will the electron **gain** the most energy?

1) $2 \rightarrow 5$
2) $5 \rightarrow 3$
3) $8 \rightarrow 5$
4) $4 \rightarrow 7$
5) $15 \rightarrow 7$



ConcepTest 27.5a Atomic Transitions I

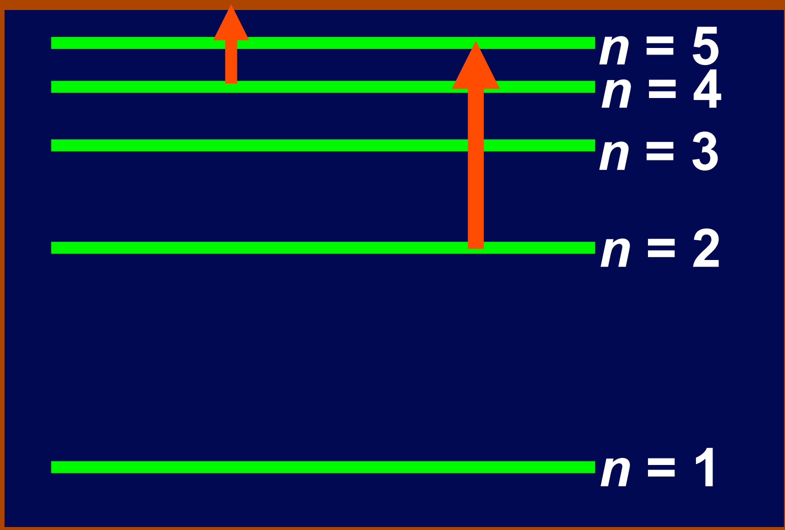
For the possible transitions shown, for which transition will the electron **gain** the most energy?

1) 2 → 5
2) 5 → 3
3) 8 → 5
4) 4 → 7
5) 15 → 7

The electron must go to a **higher** orbit (**higher *n***) in order for the electron to gain energy.

Because of the $1/n^2$ dependence:

$$E_2 - E_5 > E_4 - E_7$$

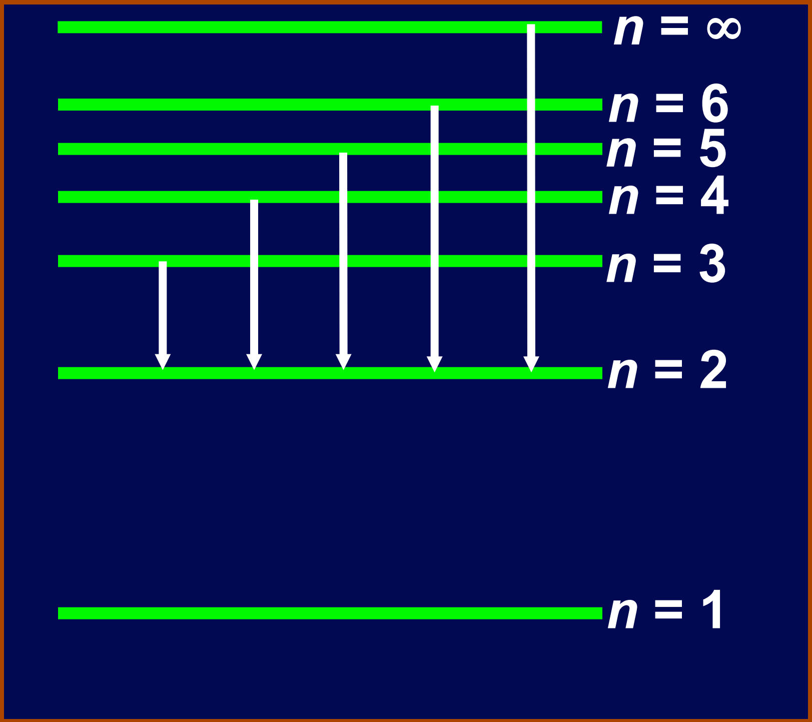


Follow-up: Which transition will emit the shortest wavelength photon?

ConcepTest 27.5b Atomic Transitions II

The Balmer series for hydrogen can be observed in the visible part of the spectrum. Which transition leads to the **reddest** line in the spectrum?

1) 3 → 2
2) 4 → 2
3) 5 → 2
4) 6 → 2
5) ∞ → 2

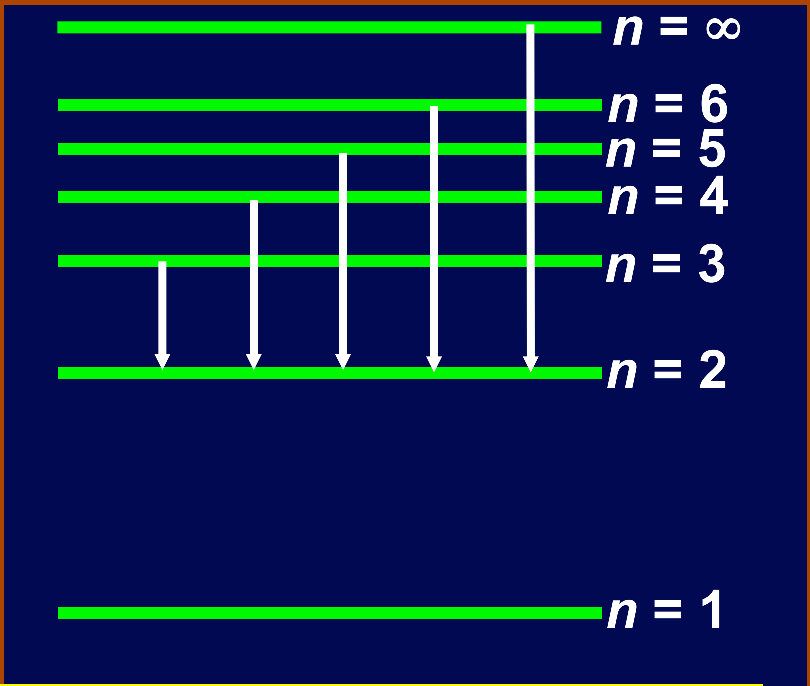


ConcepTest 27.5b Atomic Transitions II

The Balmer series for hydrogen can be observed in the visible part of the spectrum. Which transition leads to the reddest line in the spectrum?

1) $3 \rightarrow 2$
2) $4 \rightarrow 2$
3) $5 \rightarrow 2$
4) $6 \rightarrow 2$
5) $\infty \rightarrow 2$

The transition $3 \rightarrow 2$ has the lowest energy and thus the lowest frequency photon, which corresponds to the longest wavelength (and therefore the "reddest") line in the spectrum.



Follow-up: Which transition leads to the shortest wavelength photon?

Energy, Mass, and Momentum of a Photon

- A photon must travel at the speed of light.
- Quantum theory (photoelectric effect) says

  $E_{photon} = hf$
- Look at the relativistic equation for momentum: $E^2 = p^2c^2 + m^2c^4$
- The photon has no mass, so this reduces to

  $E_{photon} = pc.$
- We combine these equations and find that a photon <u>must</u> carry momentum!

$$p = \frac{E}{c} = \frac{hf}{c} = \frac{h}{\lambda} \qquad \textbf{(27-6)}$$

Consequence of Photons having momentum: - they can exert a force!

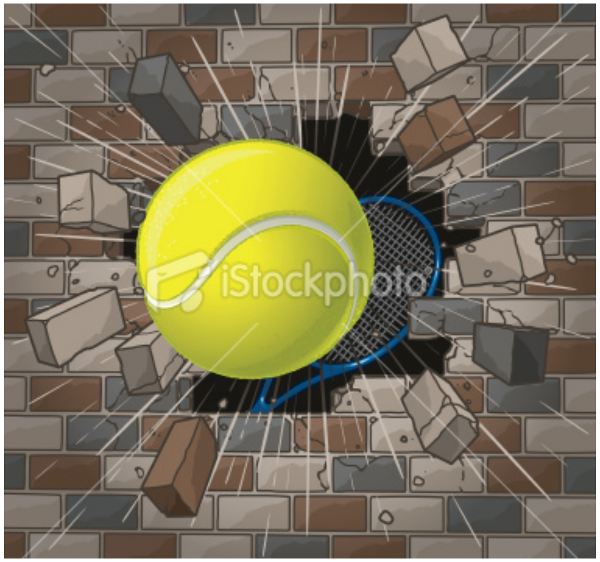
If photons carry momentum, than they should exert
A force when they strike something

Imagine a tennis ball striking a wall - it imparts an impulse, and exerts a force on the wall.

Remember Newton's 2nd law:

F=ma or F = "rate of change of momentum"

This effect is called RADIATION PRESSURE and is another unique prediction of Quantum Theory and Relativity!



Radiation Pressure

How much radiation pressure is exerted by sunlight?

Could the pressure of Sunlight be harnessed to
"sail" across the solar system?

If we know for example the amount of energy emitted by the Sun (1300 W/m^2 at Earth), and the "average wavelength" of sunlight (500 nm)….

Then we can calculate the number of photons per second:

$E_{phot} = hf \quad (c=f\lambda)$

$= hc/\lambda \quad = 6.67\times10^{-34} . 3\times10^{8} / 500\times10^{-9} = 4\times10^{-19}$ J

$N_{phot} = 1300/ 4\times10^{-19}$

$= 3.24\times10^{21}$ photons per second per square meter

$P_{phot} = h/ \lambda \quad = 6.67\times10^{-34} / 500\times10^{-9} = 1.3\times10^{-37}$

So Momentum per second = $N_{phot}\, p_{phot} = 3.24\times10^{21} \times 1.3\times10^{-37}$

which is a Pressure of 4.3×10^{-16} N/m^2

de Broglie and the Wave Nature of Matter

•Louis de Broglie, arguing from the idea of symmetry in nature discovered the following:

• Just as light sometimes behaves as a particle, matter sometimes behaves like a wave.

• The wavelength of a particle of matter is:

$$\lambda = \frac{h}{p} \tag{27-8}$$

De Broglie wavelength depends on the particle's momentum.

The de Broglie wavelength of a very fast moving electron is _______ that of a slow moving electron.

A. Shorter than

B. Longer than

C. The same as

ConcepTest 27.3a Wave-Particle Duality I

The speed of proton A is larger than the speed of proton B. Which one has the longer wavelength?

1) proton A
2) proton B
3) both the same
4) neither has a wavelength

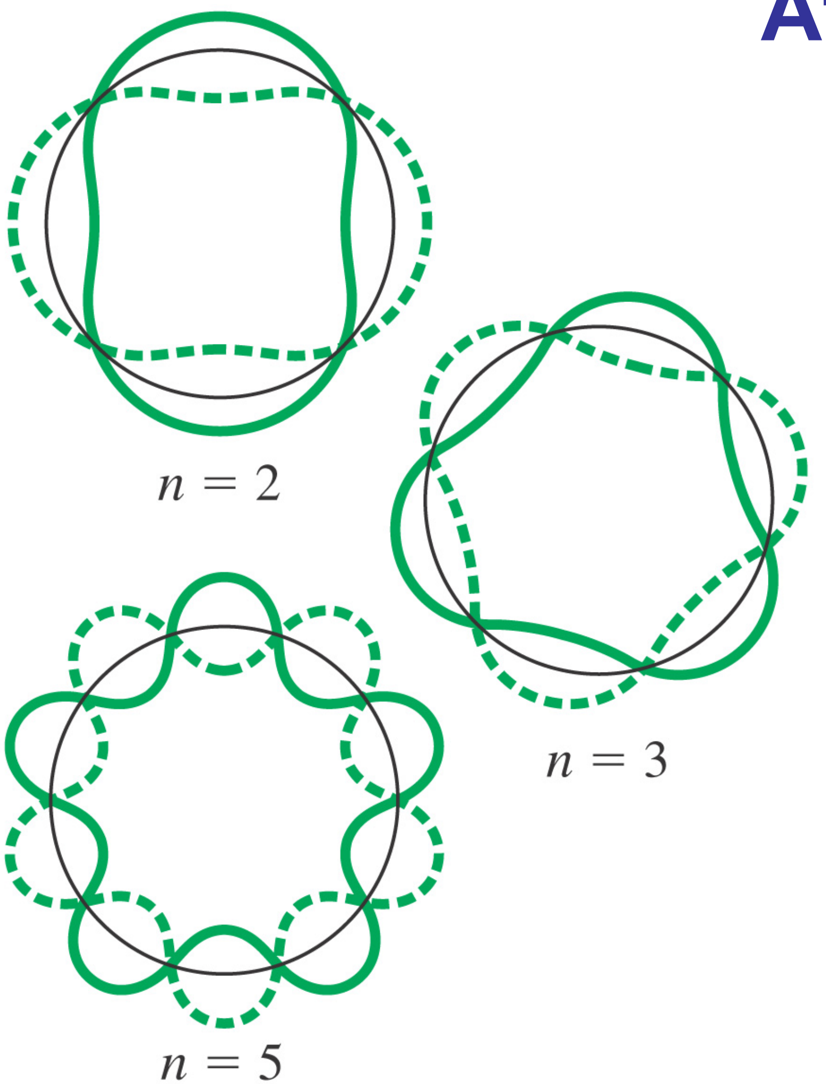
ConcepTest 27.3a Wave-Particle Duality I

The speed of proton A is larger than the speed of proton B. Which one has the longer wavelength?

1) proton A
2) proton B
3) both the same
4) neither has a wavelength

Remember that $\lambda = h/mv$ so the proton with the smaller velocity will have the longer wavelength.

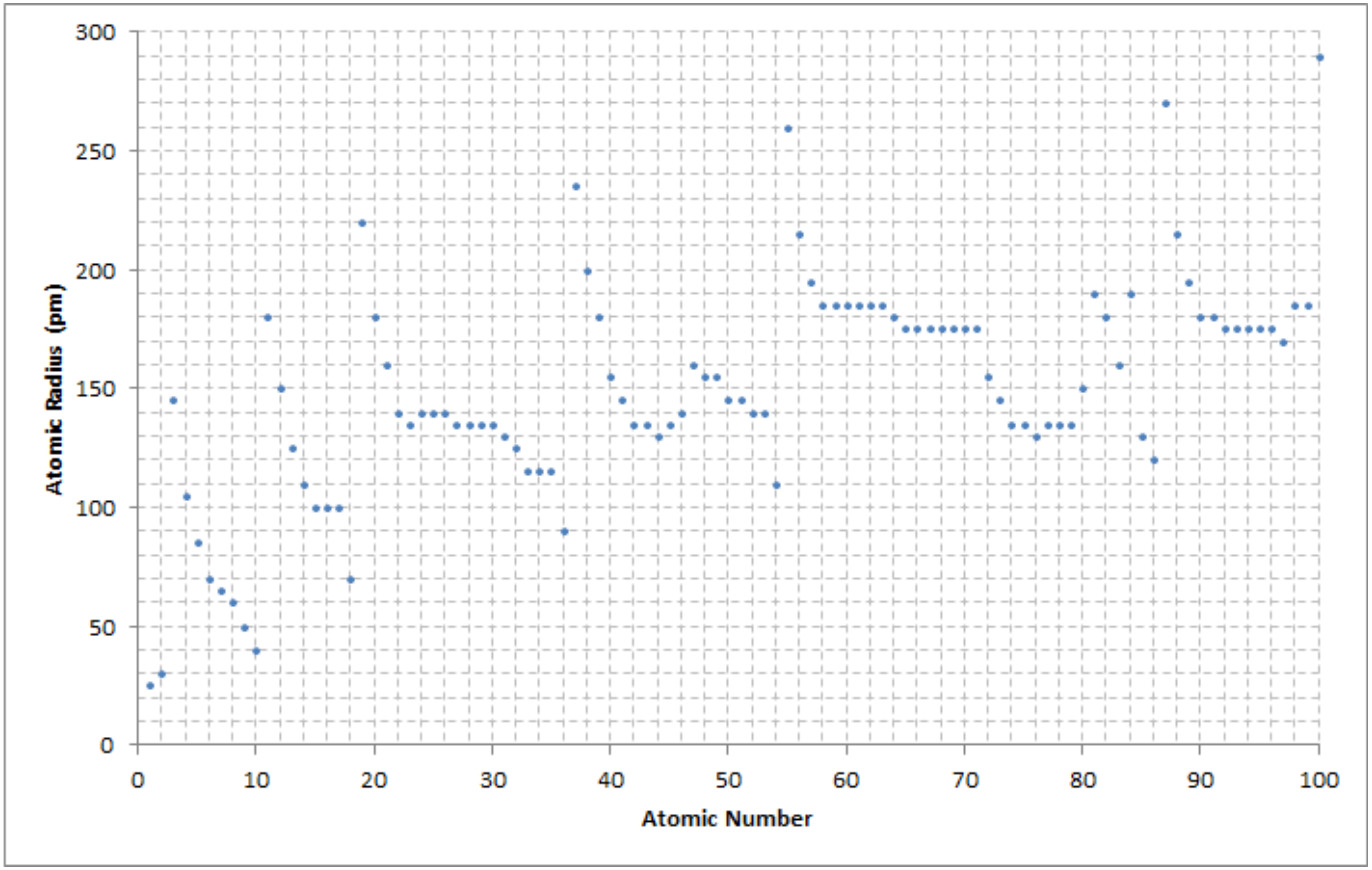
27.13 de Broglie's Hypothesis Applied to Atoms



These are circular standing waves for n = 2, 3, and 5.

- De Broglie's hypothesis associates a wavelength with the momentum of a particle.
- He proposed that only those orbits where the wave would be a circular standing wave will occur.
- This yields graphically the same relation that Bohr had proposed mathematically
- De Broglie's model adds a physical justification to the quantum model of the atom.
- Essentially the electron waves interfere, and the allowed orbits correspond to the condition for constructive interference

The Size of Atoms



- Atomic radii increase with number of protons in the nucleus
- Interesting pattern emerges
- Strange jumps in size, each followed by a steady shrinkage
- Electrons appear to be arranging into “shells”
- Note the even number of electrons per shell

The Size of Atoms

Group (vertical) / Period (horizontal)	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	H 25																	He
2	Li 145	Be 105											B 85	C 70	N 65	O 60	F 50	Ne
3	Na 180	Mg 150											Al 125	Si 110	P 100	S 100	Cl 100	Ar
4	K 220	Ca 180	Sc 160	Ti 140	V 135	Cr 140	Mn 140	Fe 140	Co 135	Ni 135	Cu 135	Zn 135	Ga 130	Ge 125	As 115	Se 115	Br 115	Kr
5	Rb 235	Sr 200	Y 180	Zr 155	Nb 145	Mo 145	Tc 135	Ru 130	Rh 135	Pd 140	Ag 160	Cd 155	In 155	Sn 145	Sb 145	Te 140	I 140	Xe
6	Cs 260	Ba 215	*	Hf 155	Ta 145	W 135	Re 135	Os 130	Ir 135	Pt 135	Au 135	Hg 150	Tl 190	Pb 180	Bi 160	Po 190	At	Rn
7	Fr	Ra 215	**	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cn	Uut	Uuq	Uup	Uuh	Uus	Uuo
Lanthanides	*	La 195	Ce 185	Pr 185	Nd 185	Pm 185	Sm 185	Eu 185	Gd 180	Tb 175	Dy 175	Ho 175	Er 175	Tm 175	Yb 175	Lu 175		
Actinides	**	Ac 195	Th 180	Pa 180	U 175	Np 175	Pu 175	Am 175	Cm	Bk	Cf	Es	Fm	Md	No	Lr		

Atomic Radii range from 25-300 pico meters (10^{-12} m)

Typical size for C, O, Fe etc is 100 pm

27.12 The Correspondence Principle

- The correspondence principle states that the <u>same</u> laws of physics should apply everywhere: from the inside of the atom, to molecules, bacteria, rocks, humans, the solar system, the universe.
- Should be <u>no dividing line</u> between quantum and classical reality
- Yet everyday objects don't appear to exhibit quantum or relativistic properties.
- When the differences between quantum levels are small compared to the energies, quantization should be imperceptible.
- When velocities are much smaller than the speed of light, relativistic effects are imperceptible.
- The laws of quantum physics and relativity simply reduce to the classical theory at "everyday" speeds, energies, sizes, etc.
- In other words, just as relativity is always in effect, so is quantization, but we often don't notice it on the scale of "our world".

Test your understanding: Why does quantization not apply to orbits of planets in the solar system?

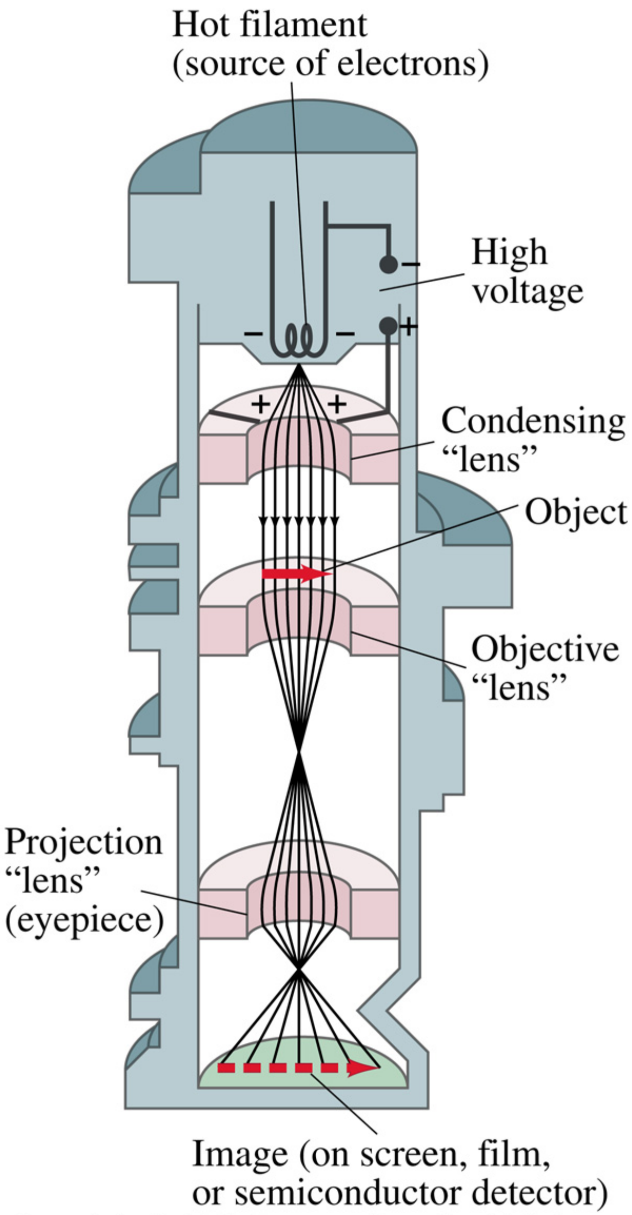
27.9 Electron Microscopes

The wavelength of electrons will vary with energy, but is still quite short. This makes electrons useful for imaging – remember that the smallest object that can be resolved is about one wavelength. Electrons used in electron microscopes have wavelengths of about 0.004 nm.

All the properties of optics apply to particles!

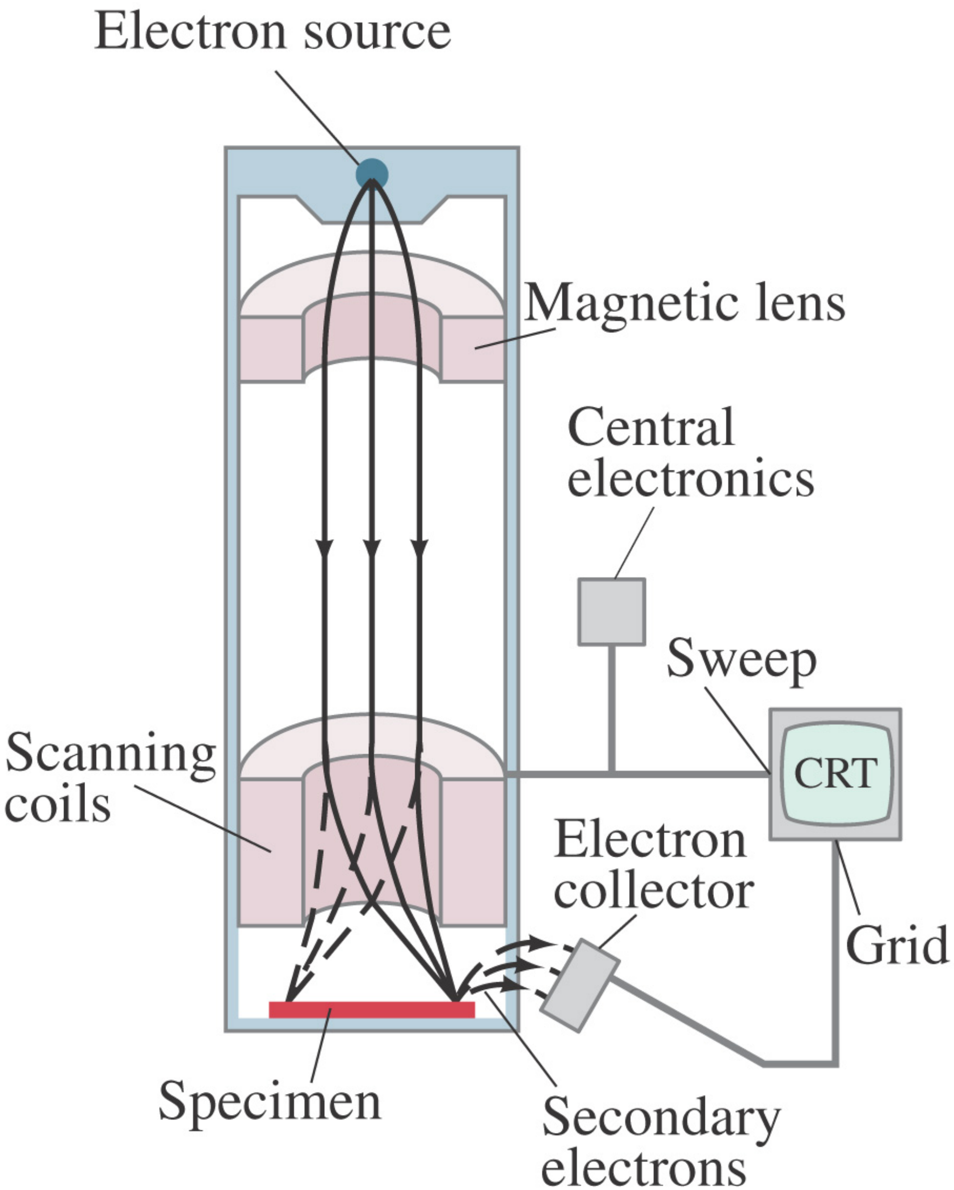
If we use the appropriate de Broglie wavelength!

27.9 Electron Microscopes



- **Transmission electron microscope – the electrons are focused by magnetic coils**
- **The magnetic coils act as lenses to focus the electron beam**
- **The resolution limit (1.22λ/D) still applies**
- **electron wavelength is tiny compared to that of light 0.004nm vs 500nm**
- **So <u>much</u> smaller details can be seen**

27.9 Scanning Electron Microscope



Scanning electron microscope – the electron beam is scanned back and forth across the object to be imaged

ConcepTest 27.3b Wave-Particle Duality II

An electron and a proton have the same speed. Which has the longer wavelength?

1) electron
2) proton
3) both the same
4) neither has a wavelength

ConcepTest 27.3b Wave-Particle Duality II

An electron and a proton have the same speed. Which has the longer wavelength?

1) electron
2) proton
3) both the same
4) neither has a wavelength

Remember that $\lambda = h/mv$ and the particles both have the same velocity, so the particle with the smaller mass will have the longer wavelength.

ConcepTest 27.3c Wave-Particle Duality III

An **electron** and a **proton** are accelerated through the **same voltage**. Which has the **longer** wavelength?

1) electron
2) proton
3) both the same
4) neither has a wavelength

ConcepTest 27.3c Wave-Particle Duality III

An electron and a proton are accelerated through the same voltage. Which has the longer wavelength?

1) electron
2) proton
3) both the same
4) neither has a wavelength

Because $PE_i = KE_f$ both particles will get the same kinetic energy (= 1/2 $mv^2 = p^2/2m$). So the lighter particle (electron) gets the smaller momentum. Because $\lambda = h/mv$ the particle with the smaller momentum will have the longer wavelength.

ConcepTest 27.3d Wave-Particle Duality IV

An **electron** and a **proton** have the **same momentum**. Which has the **longer** wavelength?

1) electron
2) proton
3) both the same
4) neither has a wavelength

ConcepTest 27.3d Wave-Particle Duality IV

An electron and a proton have the same momentum. Which has the longer wavelength?

1) electron
2) proton
3) both the same
4) neither has a wavelength

Remember that $\lambda = h/mv$ and $p = mv$, so if the particles have the same momentum, they will also have the same wavelength.