What we know so far:

Rutherford: Atoms have a tiny, but heavy core surrounded by a cloud of electrons.

Discharge lamps: Atoms struck by fast electrons emit light of distinct colors.

Energy levels: Electrons in atoms are found only in discrete energy levels. When they jump down to a lower level, a photon is emitted carrying away the energy.

Different Atoms: Different types of atoms have different level structures (seen by the distinct set of colors they emit):
- Neon: Strong red line; Sodium: strong yellow line...

Remember the ‘Electron Volt’ (eV)?

Define new energy unit: The electron-volt (eV)

1 eV = kinetic energy gained/lost by an electron when accelerated/decelerated by traversing an area with 1 Volt potential difference (such as in a plate capacitor)

1 eV = 1.6 · 10^{-19} J

Energy level diagrams represent energy levels the electron can go to. Different height = different energy

No light emitted with colors in this region because no energy levels spaced with this energy.

Some handy relations:

Current flowing through a wire:

1 Ampere = 1 Coulomb / Second
1 Watt = 1 Ampere · 1 Volt = 1 C/s · 1 V
1 Joule = 1 Watt · 1 Second = 1 C · 1 V

i.e: An electric potential of 1 V puts in one Joule of energy into a charge of 1 C → kinetic energy increases by 1 Joule.

An e⁻ has a charge of 1.6 · 10^{-19} C, therefore, it gains a kinetic energy of

1 eV = 1.6 · 10^{-19} J

Applications of atomic spectroscopy

1. Detecting what kind of atoms are in a material. (excite by putting in discharge lamp or heating in flame to see spectral lines)
2. Detecting what the sun and the stars are made of: Look at the light from a star through a diffraction grating. See what lines there are; Match up to atoms on earth.
3. Making much more efficient lights! Incandescent light bulbs waste >90% of the electrical energy that goes into them! (<10% efficient) Streetlight discharge lamps (Na or Hg) ~80% efficient. Fluorescent lights ~ 40-60% efficient.
Atomic spectra in astronomy

Application: Designing a better light

What is important?
(Well, we have to be able to see the light!)

What color(s) do you want?
(Choice of atom)

How do you excite them to desired level?
(Electron collisions)

How to get electrons with desired energy when hit atoms? What determines energy of electrons?

Florescent Lights.

How to do this?
Converting UV light into visible photons with “phosphor”.
Phosphor converts 180 nm UV to red+ green+blue.

Florescent Lights. Discharge lamp, White= Red + green +blue
40-60% efficient (electrical power⇒ Visible light)

Converting 180nm UV light into visible photons with “phosphor”.

Hubble and the big bang

Spectral lines from Hydrogen

Edwin Hubble, PNAS March 15, 1929 vol. 15 no. 3 168-173

Incanescent light (hot filament)
Temperature = 2500-3000K
Hot electrons jump between many very closely spaced levels (solid metal). Produce all colors. Mostly infrared at temp of normal filament. >90% is worthless
Infrared radiation (IR = longer than ~700nm)

~10% of energy is useful visible light

Discharge lamp

Energy levels in isolated atom:

Streetlight discharge lamps (Na or Hg) 80% efficient.

Florescent Lights.
Discharge lamp, but want to have it look white.
White = red + green + blue
40-60% efficient (electrical power⇒ visible light)

Converting 180nm UV light into visible photons with “phosphor”.

180nm → 6.9 eV energy per photon
633nm (red) → 2 eV / photon
532nm (green) → 2.3 eV / photon
475nm (blue) → 2.6 eV / photon

\{ 6.9 \text{ eV} \}

phosphor wastes 20-30% energy ⇒ heat

phosphor coating

energy of electron in phosphor molecule
### Summary of important Ideas

1) Electrons in atoms are found at specific energy levels.
2) Different set of energy levels for different atoms.
3) One photon emitted per electron jump down between energy levels. Photon color determined by energy difference.
4) If electron not bound to an atom: Can have any energy. (For instance free electrons in the PE effect.)

### Now we know about the energy levels in atoms. But how can we calculate/predict them?

→ Need a model

Step 1: Make precise, quantitative observations!
Step 2: Be creative & come up with a model.
Step 3: Put your model to the test.

### Balmer series: A closer look at the spectrum of hydrogen

Balmer (1885) noticed wavelengths followed a progression

\[ \lambda = \frac{91.19\text{nm}}{2^2 - \frac{1}{n^2}} \]

where \( n = 3, 4, 5, 6, \ldots \)

As \( n \) gets larger, what happens to wavelengths of emitted light?

→ \( \lambda \) gets smaller and smaller, but it approaches a limit.

### Hydrogen atom – Rydberg formula

**Rydberg’s general formula**

\[ \lambda = \frac{91.19\text{nm}}{\frac{1}{m^2} - \frac{1}{n^2}} \]

Predicts \( \lambda \) of \( n \rightarrow m \) transition:

\[ n \rightarrow (n > m) \]

\[ m = 1, 2, 3, \ldots \]

### Hydrogen atom – Lyman series

**Rydberg’s formula**

\[ \lambda = \frac{91.19\text{nm}}{\frac{1}{m^2} - \frac{1}{n^2}} \]

Predicts \( \lambda \) of \( n \rightarrow m \) transition:

\[ n \rightarrow (n > m) \]

\[ m = 1, 2, 3, \ldots \]

Can Rydberg’s formula tell us what ground state energy is?
The Balmer/Rydberg formula is a mathematical representation of an empirical observation.

It doesn't explain anything, really.

How can we calculate the energy levels in the hydrogen atom?

→ A semi-classical explanation of the atomic spectra (Bohr model)