The Physics of Emission & Absorption

Atoms & The Chemical Elements

The Classical Elements

Democritus and Atomism

Earth, Air, Fire, Water, (and Quintessence)

Discovering the Elements: Chemical Properties

Recognizing and naming the elements

Hydrogen, Oxygen, Gold, Lead, ...

Physical Properties of the Chemical Elements

Arranging the elements

The Atomic Number: Z

The Atomic Weight: A

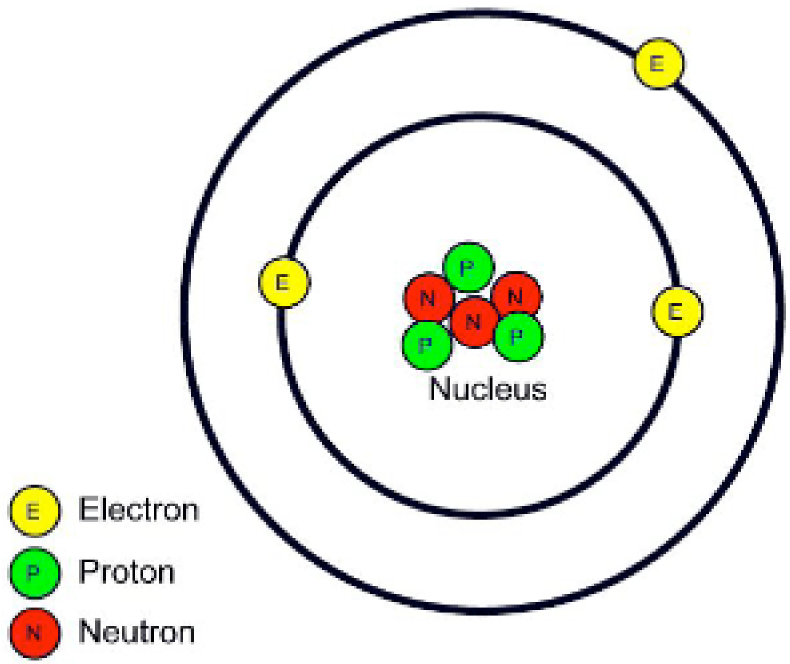
also "Electrical Properties" as "Valence" (Faraday)

...and, yes, Democritus was right about atoms too.

<u>Atoms and Atomic Structure</u>

Electrons and the Nucleus

Example: Lithium (Z = 3, A = 6, N = 3)



Protons have one unit of positive (+) charge
Electrons have one unit of negative (-) charge
Neutrons have no (0) charge.

Note: Electrons are "bound" to the nucleus by the <u>Coulomb Force</u> ($F = -Ze^2/r^2$)

Atoms, Ions, and Isotopes: Nomenclature

Atomic Number, Z
Atomic Weight, A
Neutron Number, N = A - Z
Proton Number, Z
Electron Number, Z (in neutral atom)

Ions

An **ion** is an atom with one or more of its Z electrons removed.
(It may have chemical properties similar to a lighter atom.)

Examples

Hydrogen (Z = 1): H and H^+

Helium (Z = 2): He, He^+, and He^{++}

...

Oxygen (Z = 8): O, O^+, O^{++},..., O^{+8}

Note that ions are usually (but not always) positively charged.

<u>Isotopes</u>

The chemical properties of an atom are largely determined by its electron number (Z). An element's <u>name</u> corresponds to a given Z.

Hydrogen Z = 1, Helium Z = 2, Lithium Z = 3, ..., Uranium Z = 92

Isotopes of an element (Z) have a common atomic number (Z) but differing atomic weights (A) and neutron numbers (A - Z). But isotopes of a given element have <u>similar</u> chemical properties.

H^1, H^2, H^3 or U^{234}, U^{235}, U^{238}

(Z = 1) **(Z = 92)**

Other physical properties (*e.g.*, atomic masses, radioactivity) can be <u>very</u> different.

*"Naturally occurring" isotopes. Short-lived isotopes also include U^{217} through U^{242}.

The Emission and Absorption of Light* by Atoms

Basic Idea

1. Electrons in an atom are only "allowed" in certain orbits:

Label them: n = 1, 2, 3, ...

2. Each orbit corresponds to a different electron energy:

Say: E_1, E_2, E_3, E_4, E_5, ...

3. Moving an electron between orbits requires the addition or subtraction of the energy difference:

$$\Delta E_{12} = E_1 - E_2$$

4. That energy can be supplied (absorption) or carried off (emission) by a photon of just the right frequency, ν:

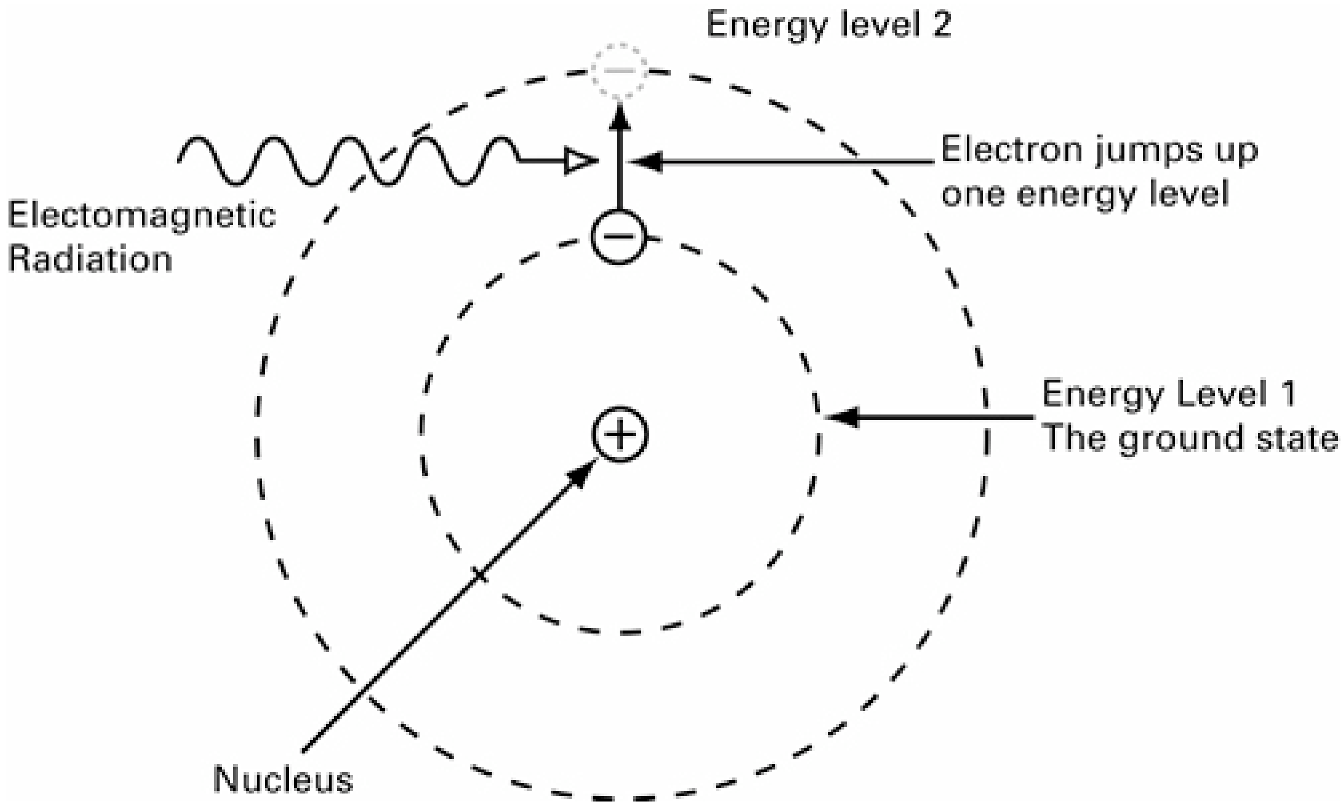
$$\Delta E = h\nu$$ (where h is "Planck's Constant")

5. The atom can emit or absorb <u>only</u> at frequencies that correspond to the energy differences between level-pairs.

... giving rise to the spectral "fingerprint" characteristic of the atom.

*** Note that "Light" includes all forms of electromagnetic radiation!**

Example: The absorption of light by Hydrogen

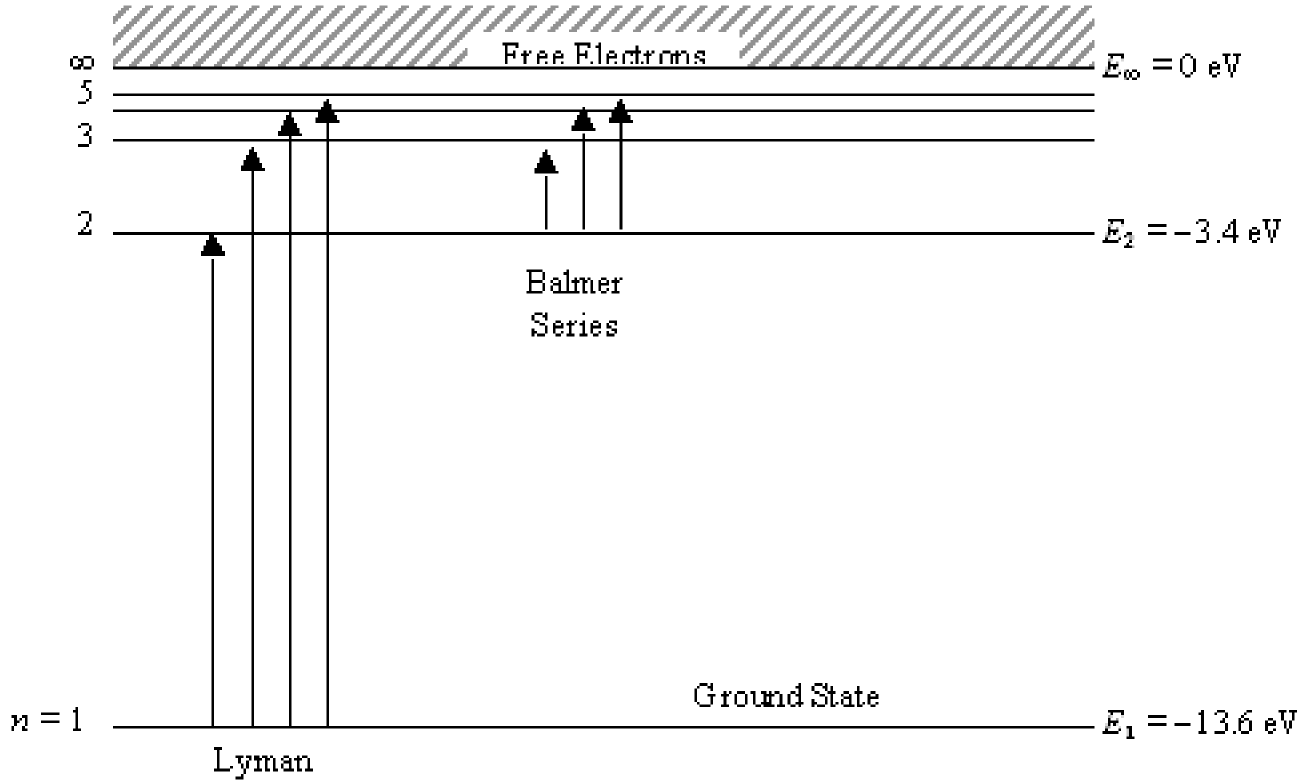


Note: The energy difference between hydrogen levels #1 and #2 corresponds to light of wavelength $\lambda = 121.5$ nm

or frequency $\nu = 2.447 \times 10^{15}$ Hz.

This is in the far ultraviolet part of the spectrum

Example: The absorption of light by Hydrogen
(another view)
Orbits and Energy Levels



Hydrogen Energy Level Diagram

Molecules and Their Spectra

The Chemical Bond

Atoms can be bound to other atoms to form molecules.

Examples: H_2, H_2O, CH_4, C_2H_5OH, H_2S, ...

The binding force is generally the electric (Coulomb) force.

ionic bonding, sharing electrons, van der Waals attraction, *etc.*

Na & Cl *vs.* Na^+ & Cl^-

Molecules can appear as solid, liquid, gaseous, or plasma.

Molecular Spectra

Energy levels can be Electronic, Vibrational, or Rotational.
One can have Electronic, Vibrational, or Rotational transitions
- or combinations thereof!

Molecular spectra are generally very complex.

but, like atoms, each molecule has a unique spectral fingerprint!