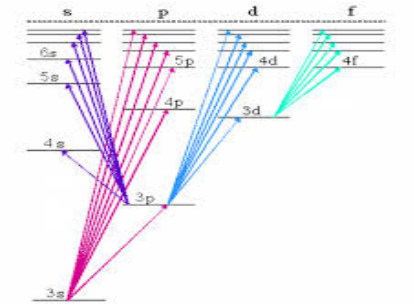
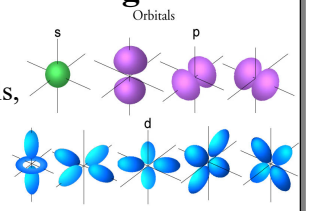
6.3 Atomic Physics



Remember Electron Configuration?

Electrons are organized into energy levels, differently shaped sub-levels, and orbitals, and have a configuration that follows a precise sequence.



Ex: Lead's configuration is [Xe] $4f^{14}5d^{10}6s^{2}6p^{2}$

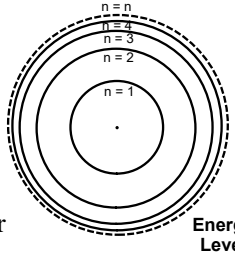
We won't go into that level of detail, but I was hoping to freak you out!

Atomic Energy Levels

Electrons can be thought of as moving on rungs of a ladder (principal quantum numbers: **n**): when they gain energy they move up; when they lose energy they move down.

The lowest level is the ground state (n = 1), and they become excited by adding energy.

Electrons jump between levels willy-nilly, absorbing or emitting photons as they do.



Energy Level Model

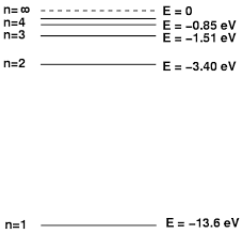
Hydrogen's Energy Levels

In hydrogen, the simplest element, radii and energy at different levels (n = energy level) is calculated:

Radius (m): $r_n = 0.0529E-9m\cdot n^2 \qquad n = 1, 2, 3\ldots$

Energy (eV)*: $E_n = \dfrac{-13.6eV}{n^2} \qquad n = 1, 2, 3\ldots$

Energy Level Diagram
(*Note Negative Energy Values)



*Energy Perspectives

Electrons can be thought of as 'trapped' in an 'energy well'; the negative value is the amount of energy it lost to get there, or how much it needs to get out (called binding energy).

As electrons move between hydrogen's energy levels (up or down), energy change (in eV) is:

$$\Delta E = -13.6eV\left(\frac{1}{n^2}-\frac{1}{n_0^2}\right)$$

E = energy (eV)
n = final energy level
n_0 = inital energy level

To totally escape an atom: set n = infinity: $1/n^2 = 0$.

1. Radius Example

What's the hydrogen atom's radius when its electron is in the 3rd energy level?

$$r_n = 0.0529E-9m\cdot n^2 \qquad n=3$$

$$= 0.0529E-9m\cdot 3^2 = 4.76E-10m$$

Binding Energy Examples

2. How much energy is gained by a hydrogen electron when it jumps from the 4th energy level to the 8th?

$$\Delta E = -13.6eV\left(\frac{1}{n^2}-\frac{1}{n_0^2}\right)$$

$$= -13.6eV\left(\frac{1}{8^2}-\frac{1}{4^2}\right) = 0.638eV$$

3. How much energy would it need to overcome the binding energy from the 8th energy level?

$$\Delta E = -13.6eV\left(\frac{1}{n^2}-\frac{1}{n_0^2}\right) = -13.6eV\left(\frac{1}{\infty^2}-\frac{1}{8^2}\right) = 0.21eV$$

Photonic Wavelength

The wavelength of emitted or absorbed photons can be calculated:

$\lambda = \frac{1.24E-6eV\bullet m}{\Delta E}$	λ = wavelength (m) ΔE = Energy change (eV)

Derivation: Optional?

Planck's Equation (h=4.14 E -15 eV•s) $E = h\bullet f$

Wave Equation $c = \lambda\bullet f$

$\lambda = \frac{c}{f}$

$f = \frac{E}{h}$

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

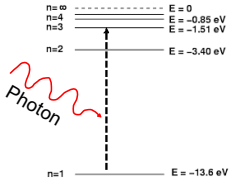
$$= \frac{(4.14E-15eV\bullet s)(3E8\frac{m}{s})}{E}$$

$$= \frac{1.24E-6eV\bullet m}{E}$$

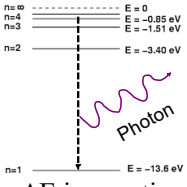
A Note on Energy

Wavelength is always positive, but realize that if the photon is absorbed (n increases), *energy change* is positive.

If the electron gives off a photon (by dropping energy levels), *energy change* is negative.



ΔE is positive

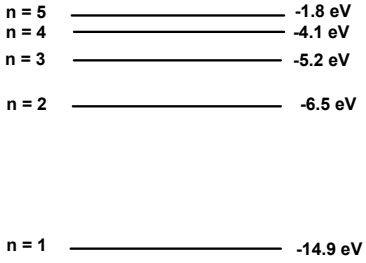


ΔE is negative

4. Energy Level Example

An simple energy level diagram is shown. An electron emits a photon (λ = 516 nm) as it drops from one level to another.

Draw an arrow indicating this transition.



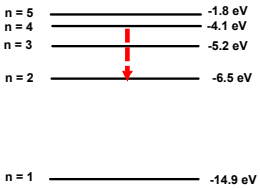
Example

Find energy loss, then compare that to energy differences between levels:

$$\lambda = \frac{1.24E-6eV\bullet m}{-\Delta E}$$

$$-\Delta E = \frac{1.24E-6eV\bullet m}{5.16E-7m} = -2.40eV\,(energy\,lost)$$

-2.40 eV corresponds to a drop from n = 4 to n = 2:
-6.5 eV - (-4.1 eV) = -2.4 eV.



Homework 6.3

Problems 6.3 in your Booklet

Due: Next Class