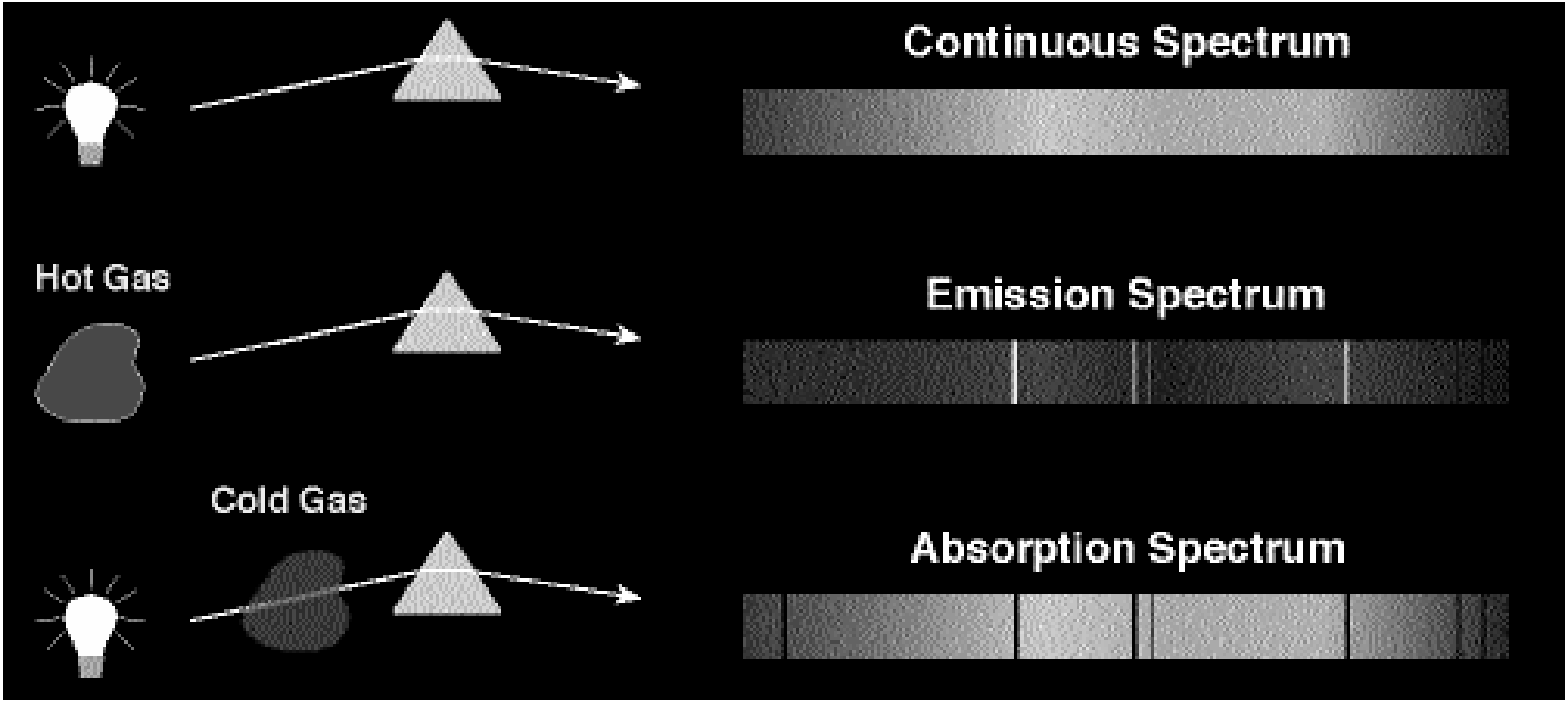
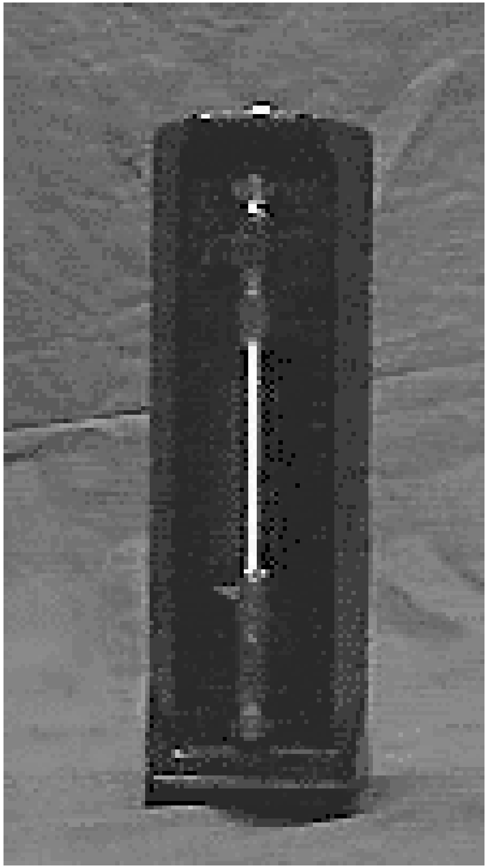
Forces between atoms and molecules

Optical Spectra – way of interrogating atoms and matte

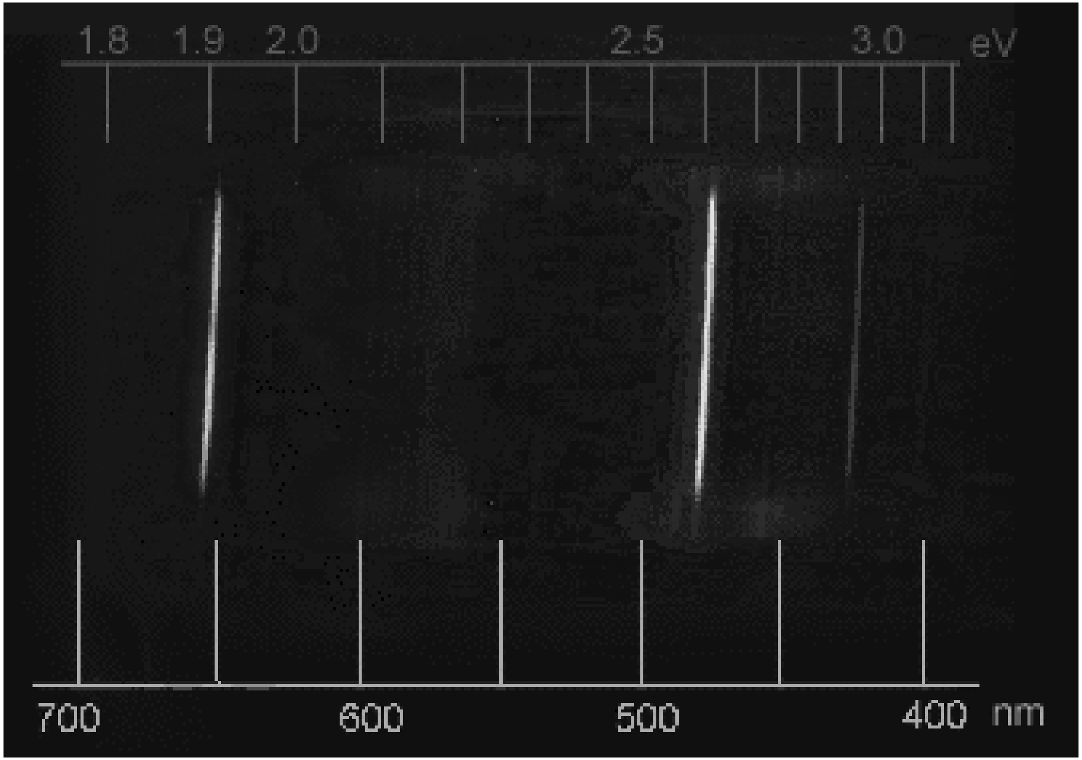


http://csep10.phys.utk.edu/astr162/lect/light/absorption.html

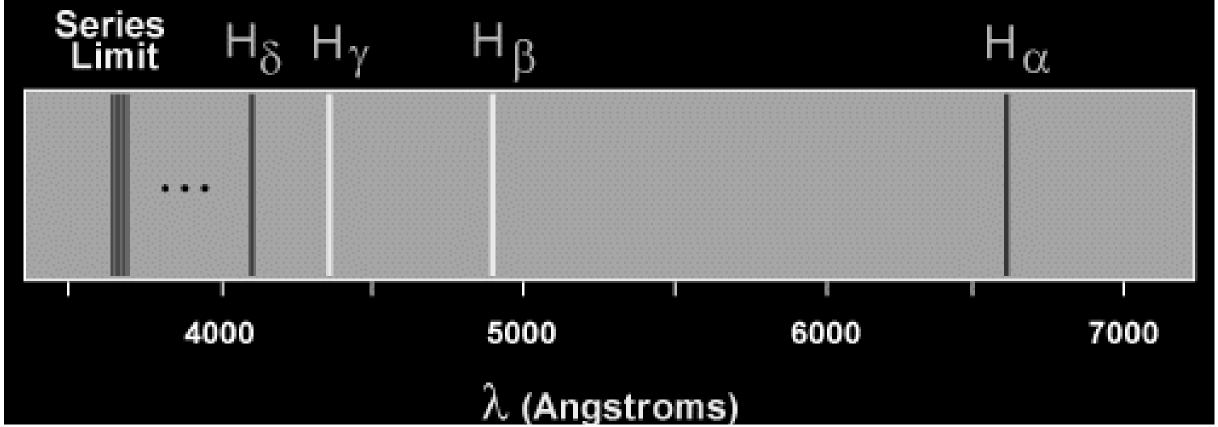
Hydrogen spectra



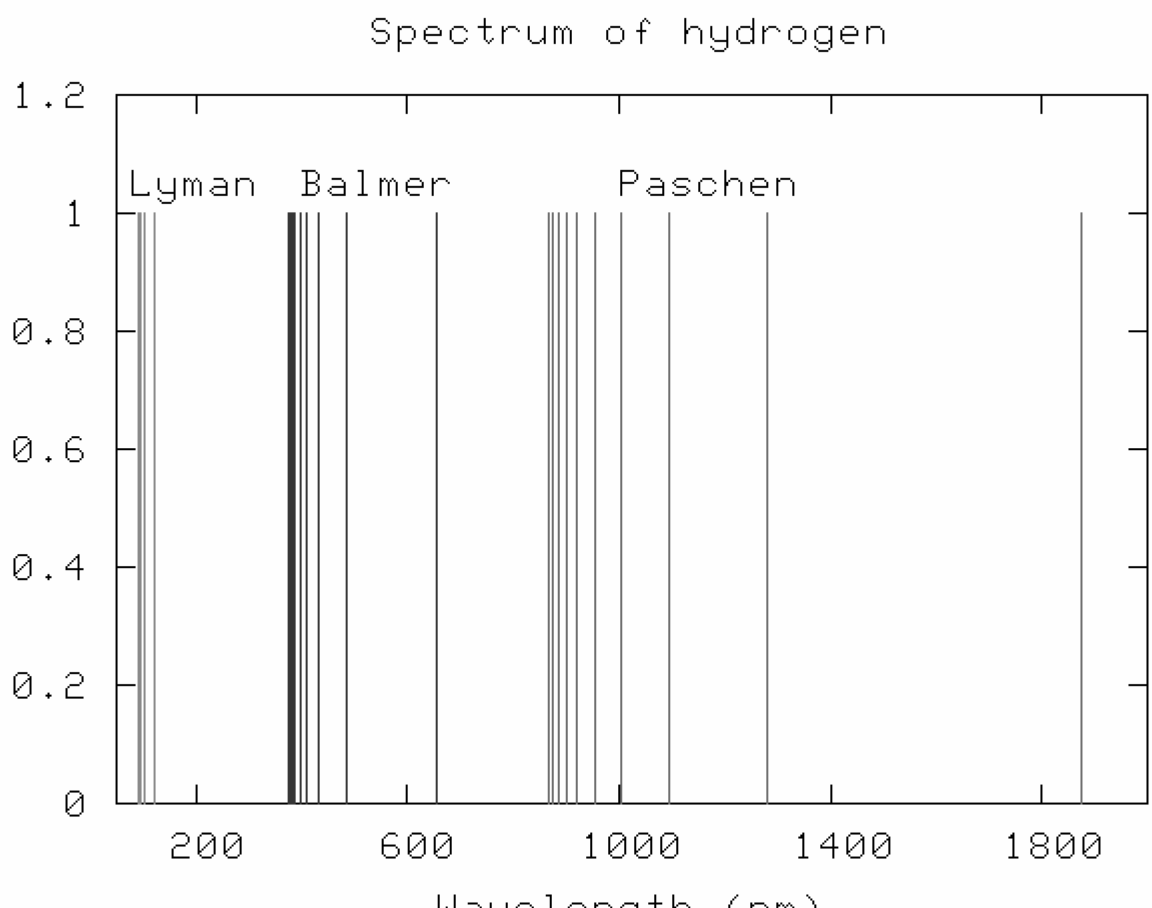
Hydrogen lamp



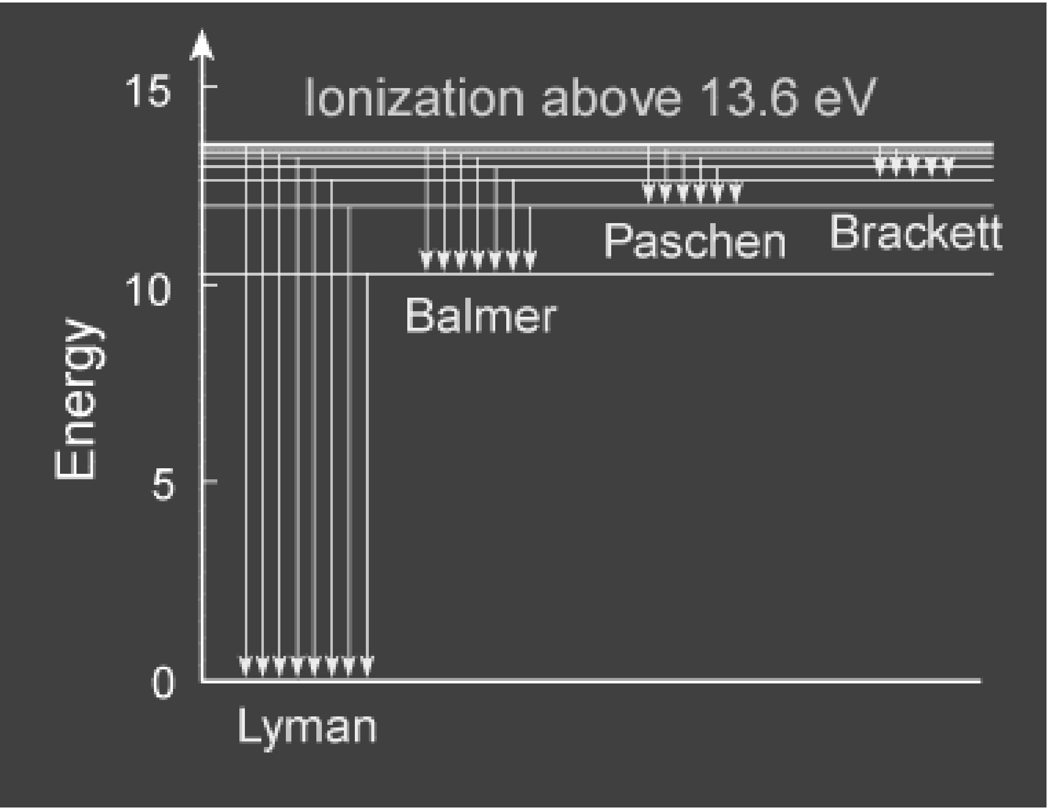
Hydrogen spectrum



Series Limit
Hδ
Hγ
Hβ
Hα
4000
5000
6000
7000
λ (Angstroms)



Spectrum of hydrogen
Lyman
Balmer
Paschen
1.2
1
0.8
0.6
0.4
0.2
0
200
600
1000
1400
1800



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

The Balmer series arises from transition down to n = 2.

Other series are:

Lyman series (transitions to n = 1)
Paschen series (transitions to n = 3)
Brackett series (transitions to n = 4)
Pfund (transitions to n = 5)

THE BOHR MODEL OF ATOM

Niels Bohr in 1913 developed his theory of the **hydrogenic** (one-electron) atoms from **four postulates**:

1. An electron in an atom moves in a **circular orbit** about the nucleus under the influence of the Coulomb attraction between the electron and the nucleus, obeying the laws of classical mechanics.

2. Instead of the infinity of orbits, which would be possible in classical mechanics, it is only possible for an electron to move in an orbit for which its **orbital angular momentum *L* is and integral multiple of h**.

3. Despite the fact that it is constantly accelerating, an electron moving in such an allowed orbit **does not radiate electromagnetic energy**. Thus, its total energy **E** remains constant.

4. Electromagnetic radiation is emitted if an **electron**, initially moving in an orbit of total energy, $\mathbf{E}_i$, **discontinuously changes its motion** so that it moves in an orbit of total energy $\mathbf{E}_f$. The frequency of the emitted radiation $\boldsymbol{v}$ is equal to the quantity ($\mathbf{E}_i - \mathbf{E}_f$) divided by **h**.

THE BOHR MODEL OF ATOM

The **second postulate** can be written mathematically

$$L = nh, \qquad n = 1,2.3,\ldots$$

The **energy levels E_n and orbit radii** for hydrogenic atom are also **quantized**:

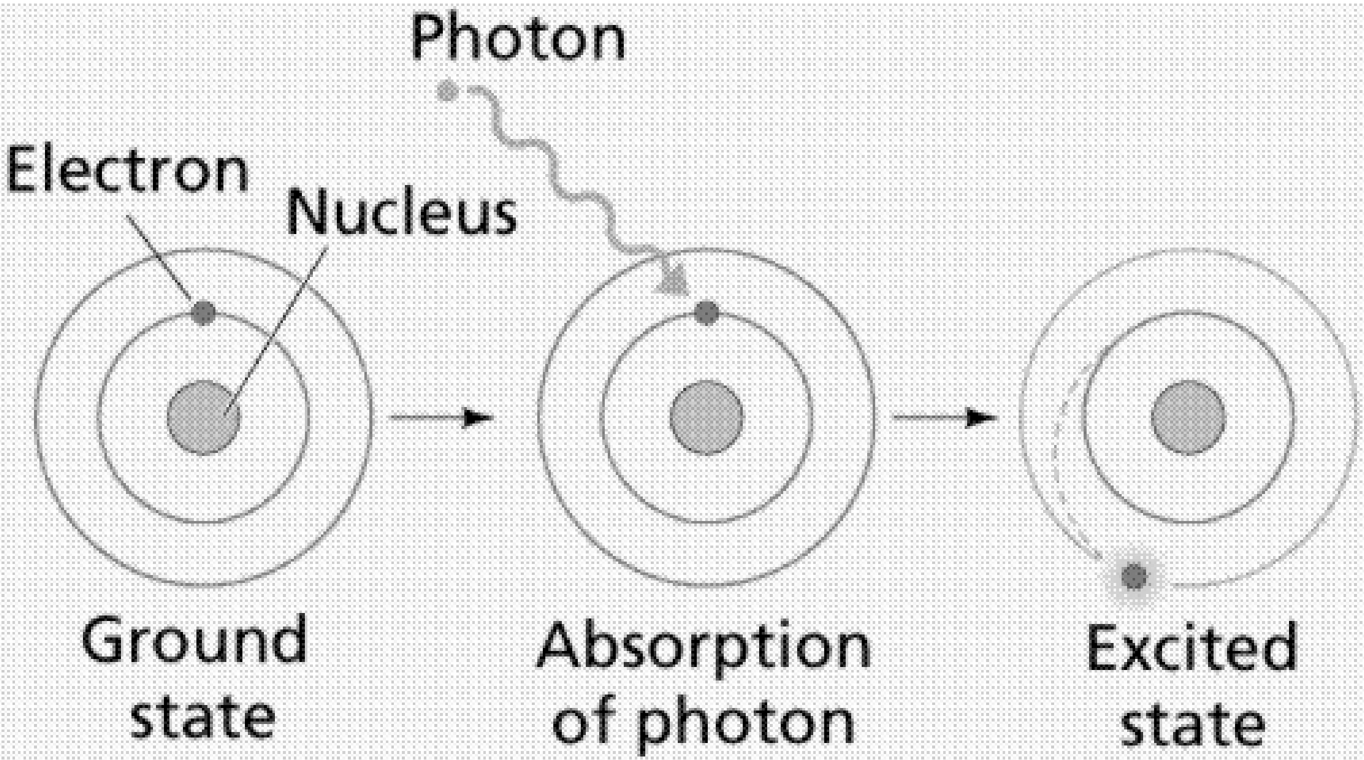
$$E = -13.6\ \text{eV} \times (Z^2/n^2).$$

$$r = (n^2/Z)a_0$$

a_0 = 0.529 Å = 0.529 ×10^{-10} m - **Bohr radius**

Z = charge of nucleus

THE BOHR MODEL OF ATOM



http://www.estrellamountain.edu/faculty/farabee/biobk/BioBookCHEM1.html

Pauli exclusion principle

1. No two electrons in an atom can be at the same time in the same state or configuration.

2. Accounts for the observed patterns of light emission from atoms.

3. Those particles to which the Pauli exclusion principle applies are called ***fermions***; those that do not obey this principle are called ***bosons***.

4. **Two electrons**, and more generally two fermions, **cannot have the same quantum state** (position, momentum, mass, spin).

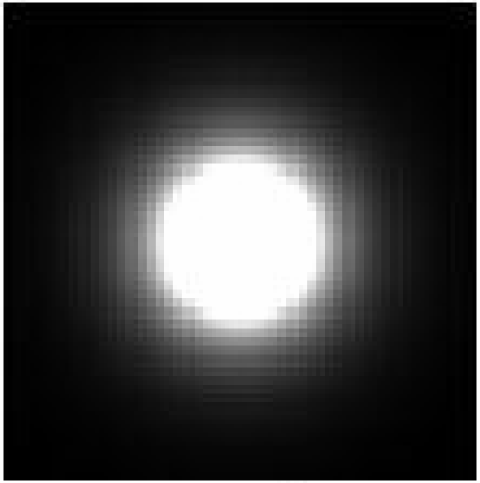
Pauli exclusion principle

5. ***Orbitals*** are regions around nuclei, each of which comprises only two distinct states. The Pauli exclusion principle indicates that, if one of these states is occupied by an electron of spin one-half, the other may be occupied only by an electron of opposite spin, or spin negative one-half.

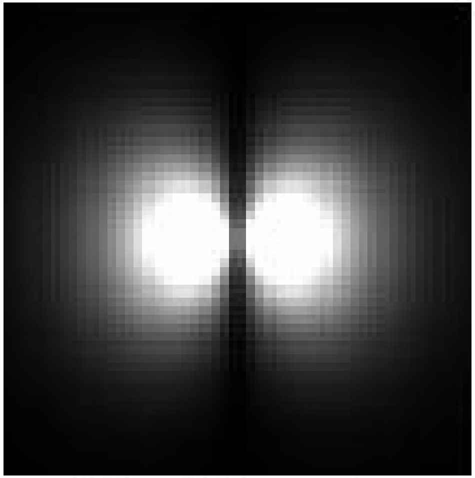
6. An orbital occupied by a pair of electrons of opposite spin is filled: no more electrons may enter it until one of the pair vacates the orbital.

7. An alternative version of the exclusion principle as applied to atomic electrons states that **no two electrons can have the same values of all four quantum numbers**.

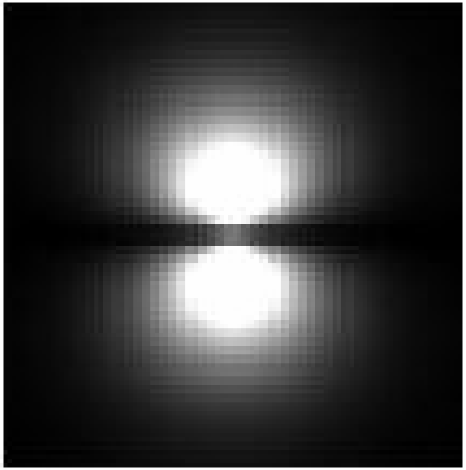
Quantum numbers



n=1



n=2 , *l* =1 , m=0



n=2 , *l* =1 , m=1

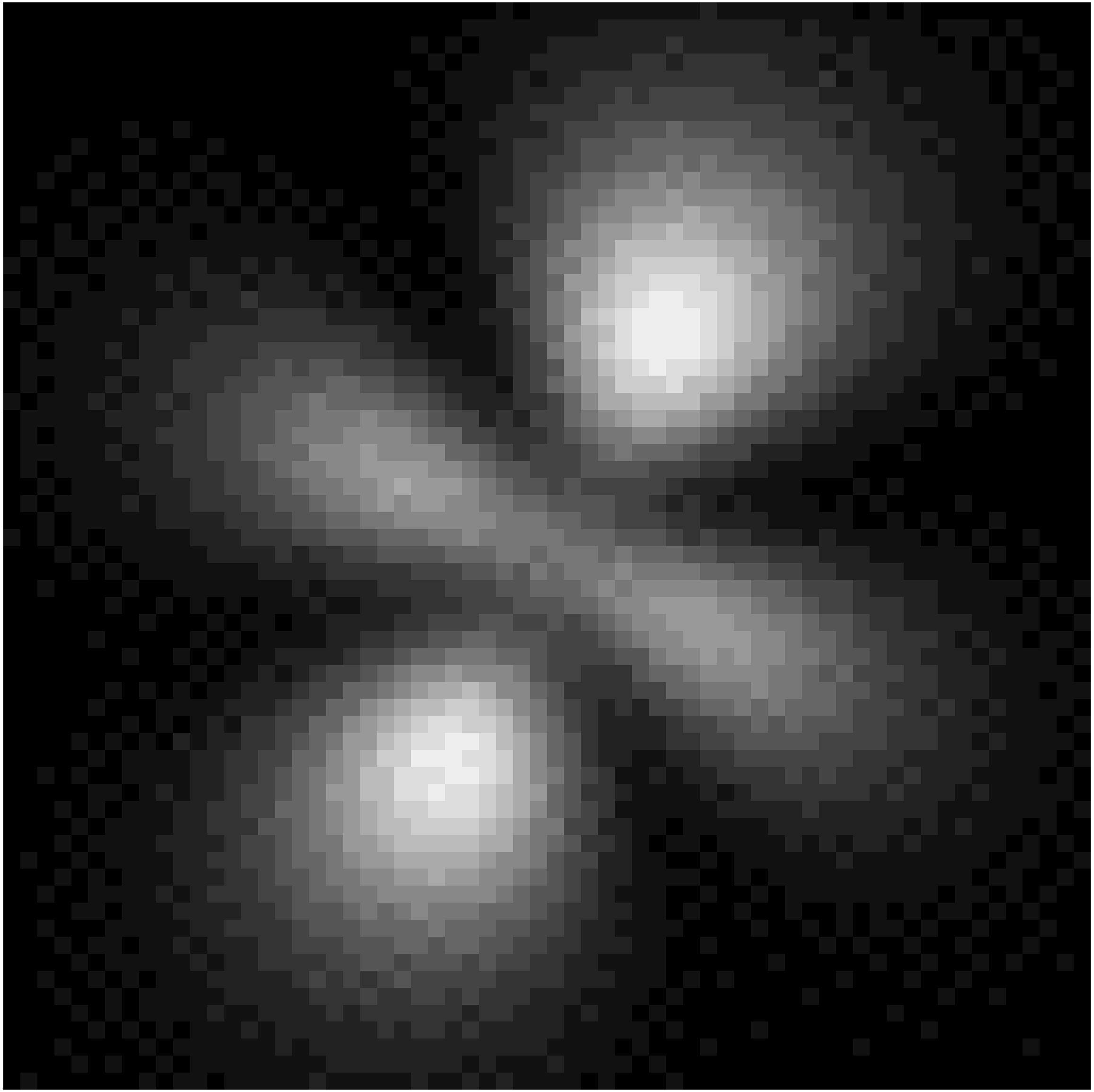
n = 1, 2, 3, ..., denoting energy;

l **= 0, 1, ..., n-1, denoting angular momentum;**

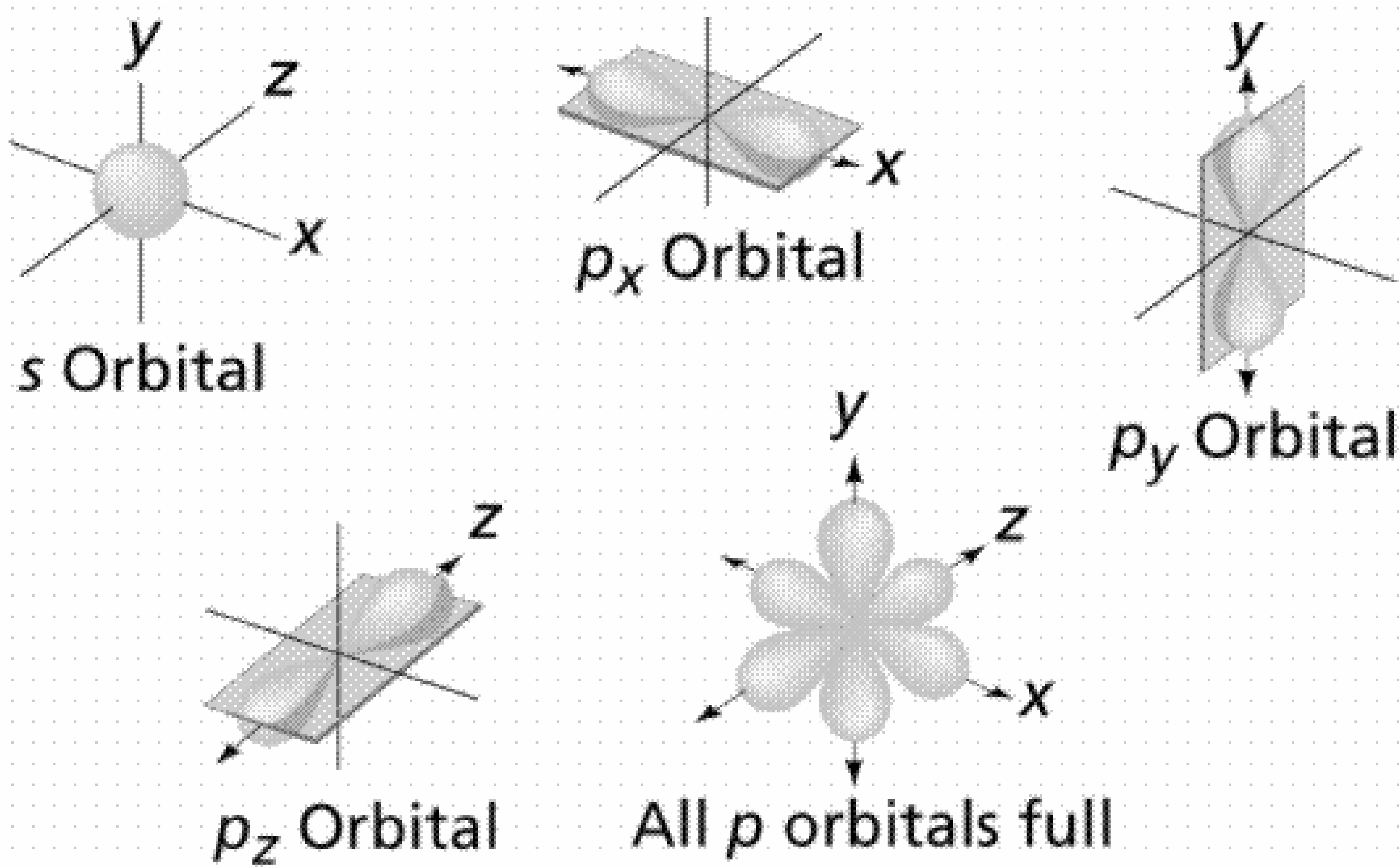
m = -*l*, ..., *l*, denoting orientation (the "magnetic quantum number"

s = -1/2, 1/2, denoting spin.

Orbital – probability distribution of finding an electron in space



Orbital – probability distribution of finding an electron in space



Orbital – probability distribution of finding an electron in space