According to Classical Physics, the nuclear atom cannot exist.

Quantum mechanics (wave mechanics) was developed to deal with the properties of microscopic particles. Such particles have wave characteristics as well as particulate properties.

\[ \hat{A} \psi = E \psi \]

The exact path of a subatomic particle can not be known.

All knowledge of the location an electron in an atom is contained in the wave function, \( \psi \) (psi). The probability of finding the electron in a given region of space is related to the value of \( |\Psi^* \Psi| \) (psi-squared).
\( \Psi \) is evaluated in a spherical coordinate system, defined by \( r, \theta, \) and \( \phi. \)

\[
\Psi = r^{n+l-1} \cos^l(m_\theta) \exp(-im_\phi)
\]

A typical wave function might have the structure:

The quantum numbers \( n, l, m_\ell \) can take on only certain allowed values (those which give meaningful values to psi-squared).
\( n \) principal quantum number

Tells you the average distance between the electron and the nucleus.

Is the main factor in determining the energy of the electron.

(Note: The electron has a negative energy when it is "bound" in the atom by attraction to the nucleus. The electron has \( E = 0 \) when the electron is free of the nucleus' attractive force.)

\( n = 1, 2, 3, ... \)
l  azimuthal quantum number

gives directional preferences of the electron

\[ l = 0 \]

\[ l = 1 \]

fine tunes the energy in multi-electron atom

\[ l = 0, 1, 2, \ldots n-1 \]

The \( n \)th energy level is divided into \((nl)\) sublevels.

\[ n = 1 \]
\[ l = 0 \]

\[ n = 2 \]
\[ l = 0 \]

\[ n = 3 \]
\[ l = 0 \]

\[ n = 2 \]
\[ l = 1 \]

\[ n = 3 \]
\[ l = 1 \]

\[ n = 3 \]
\[ l = 2 \]
\( n = 3 \quad l = 2 \quad m_l = -2, -1, 0, 1, 2 \)

\( n = 3 \quad l = 1 \quad m_l = -1, 0, 1 \)

\( n = 3 \quad l = 0 \quad m_l = 0 \)

\( n = 2 \quad l = 1 \quad m_l = -1, 0, 1 \)

\( n = 2 \quad l = 0 \quad m_l = 0 \)

\( n = 1 \quad l = 0 \quad m_l = 0 \)
$m_l$ magnetic quantum number
tells how being placed in a magnetic field will affect the
energy of the electron (magnetic fields are created when
atoms approach one another to bond).

Each $(nl)$ sublevel contains orbitals which are degenerate
(have the same energy) in the absence of an external
magnetic field.

$m_l = -l, -l+1, \ldots, 0, \ldots, l-1, l$

When an electron is moving in accord with the $\Psi | nlm_l >$,
We say it is in the $nl$ orbital.

The value of $l$ is specified by a letter.

<table>
<thead>
<tr>
<th>letter</th>
<th>s</th>
<th>p</th>
<th>d</th>
<th>f</th>
</tr>
</thead>
<tbody>
<tr>
<td>Value of $l$</td>
<td>0</td>
<td>1</td>
<td>2</td>
<td>3</td>
</tr>
</tbody>
</table>
In a multi-electron atom, we assume the orbitals will be defined in the same way as they are defined in the hydrogen atom. The question is: “How many electrons can occupy a single orbital”?

Answer: We need to consider another quantum number, $m_s$.

This is the spin quantum number and it has the values of $+1/2$ and $-1/2$. Spin is a characteristic of the electron and not of the atom.

Pauli Principle: No two electrons in a single atom can have an identical set of the four quantum numbers:

$$n, l, m_l, m_s.$$  

We can use this rule to explain the structure of the Periodic Table (aufbau principle).

\[ n = 1, \quad l = 0, \quad m_l = 0, \quad m_s = +\frac{1}{2} \]

\[ n = 1, \quad l = 0, \quad m_l = 0, \quad m_s = -\frac{1}{2} \]

in the is orbital

only 2 electrons can occupy a single orbital (in an atom)
The 1s energy is lower than H 1s energy
\( Z^2 = 2 \text{ for } \text{He} \) + \( Z = 1 \text{ for } H \)

\( e^-e^- \) repulsion works against extra \( Z \) \( e^- \) attraction