Chapter 6 Part 2

Dr. Turner

Orbitals

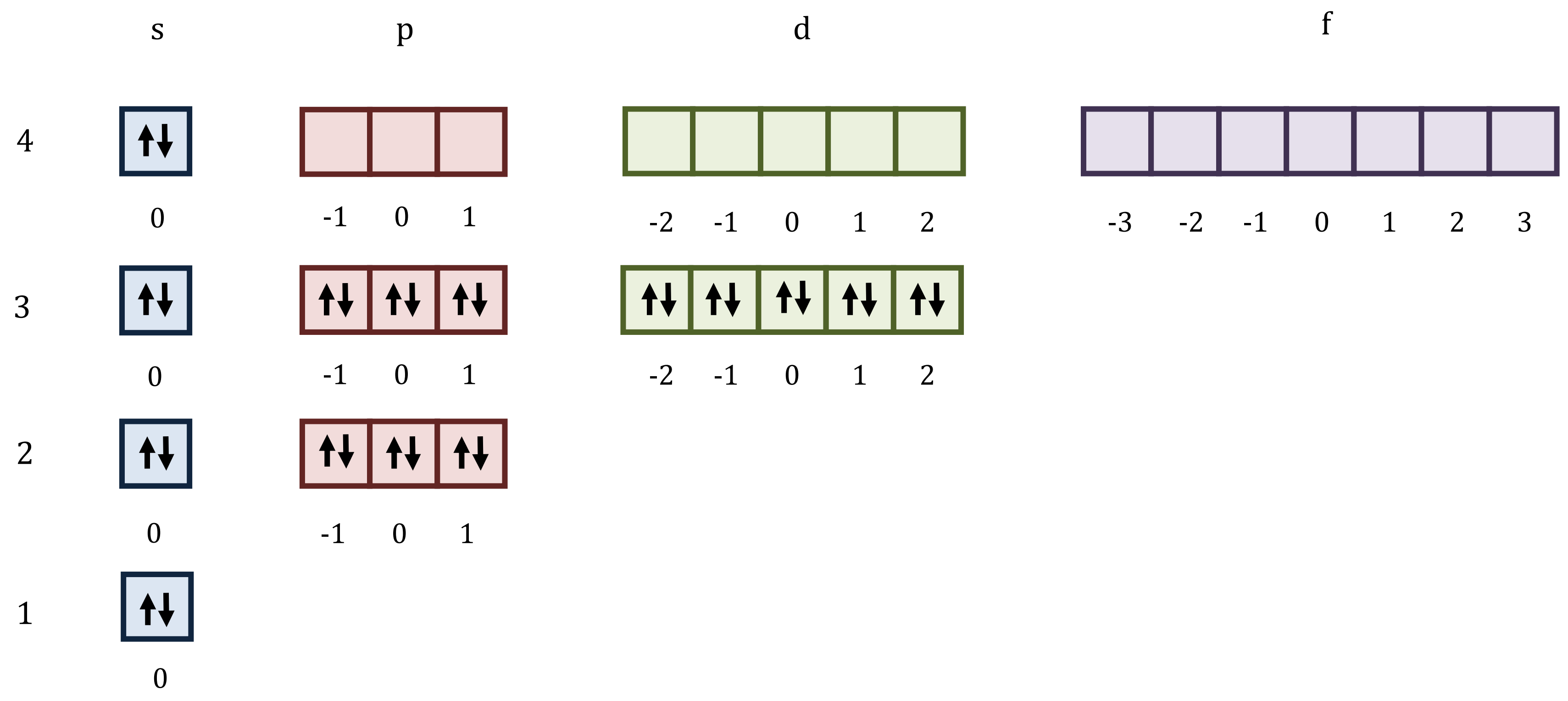
- Orbitals describe the most probable area of finding an electron in an atom
- An orbital can hold two electrons

Orbitals

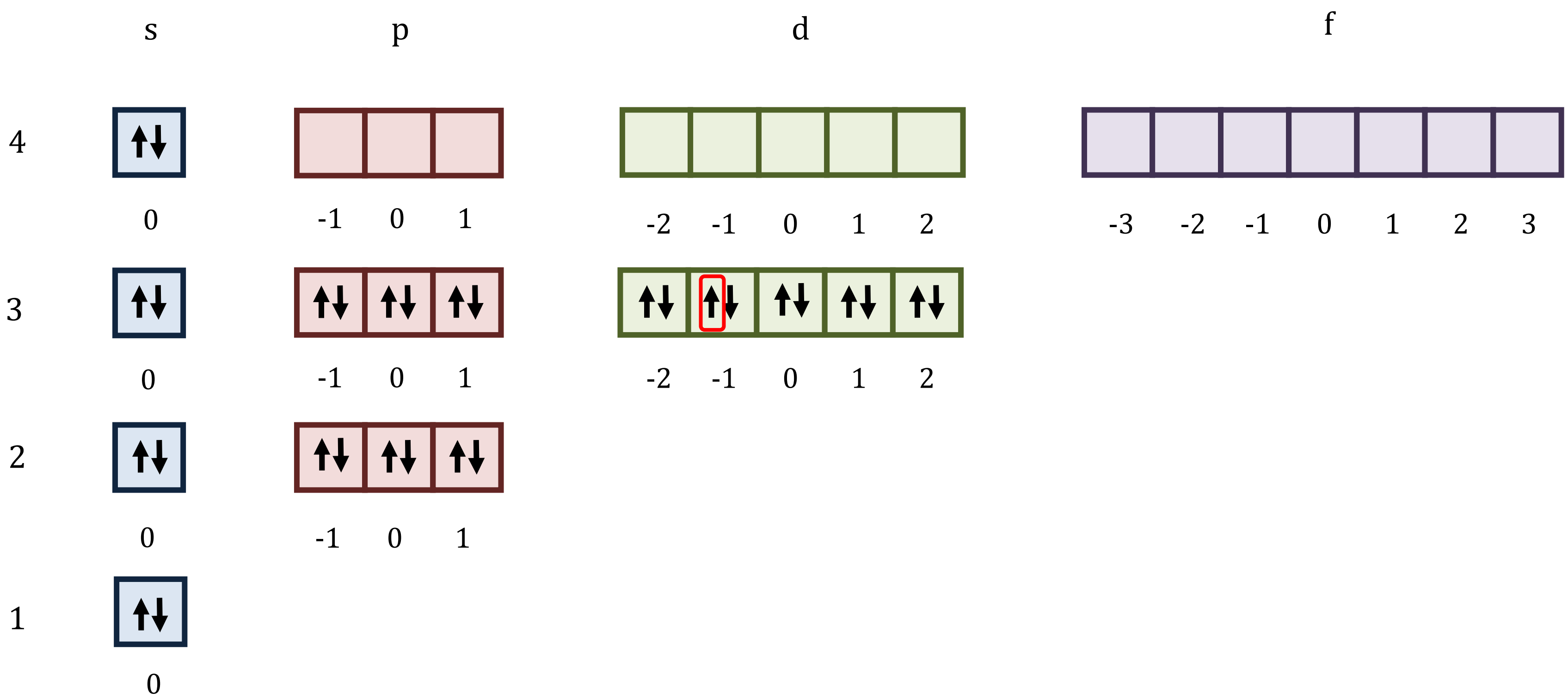
What is the maximum number of electrons any orbital can hold?

A. 1
B. 2
C. 3
D. 6
E. 10

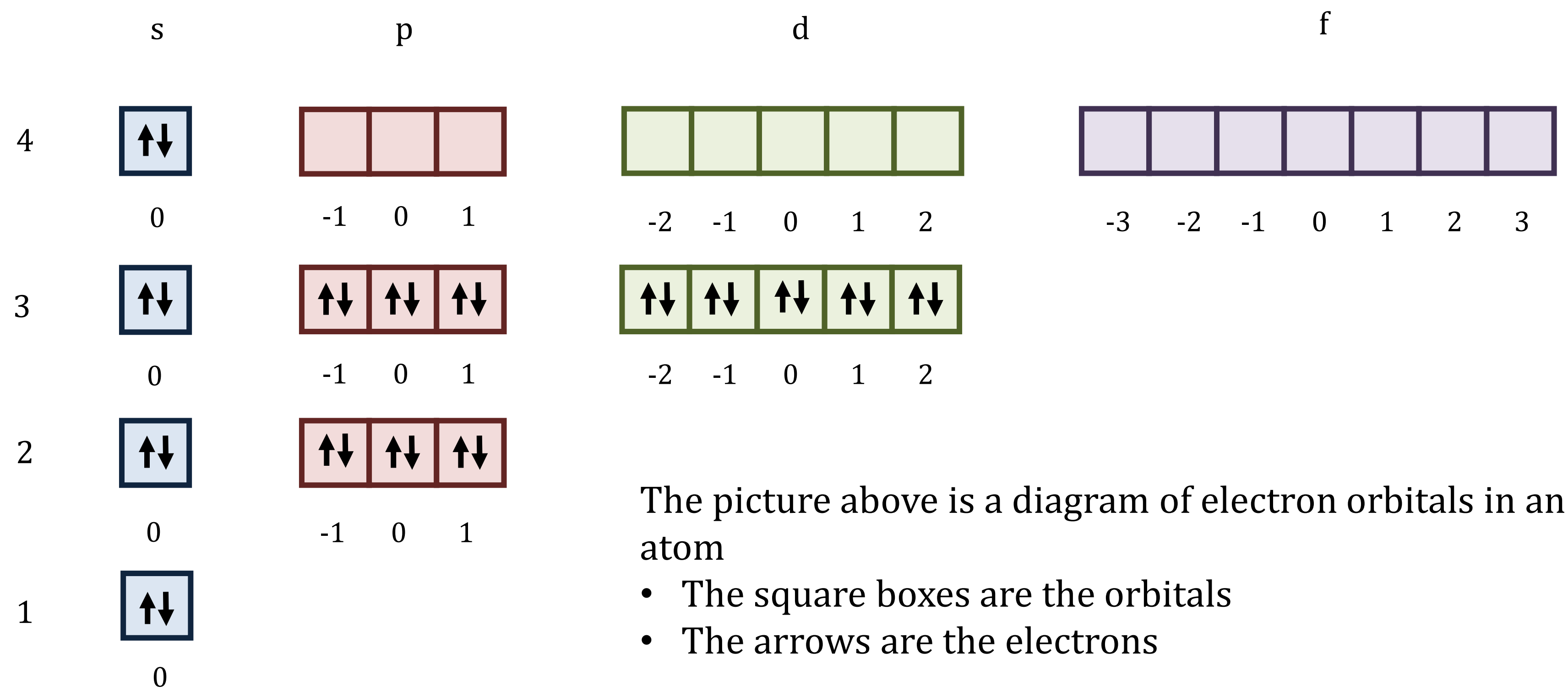
A picture



Describe which arrow is circled



Describe which arrow is circled



The picture above is a diagram of electron orbitals in an atom

- The square boxes are the orbitals
- The arrows are the electrons

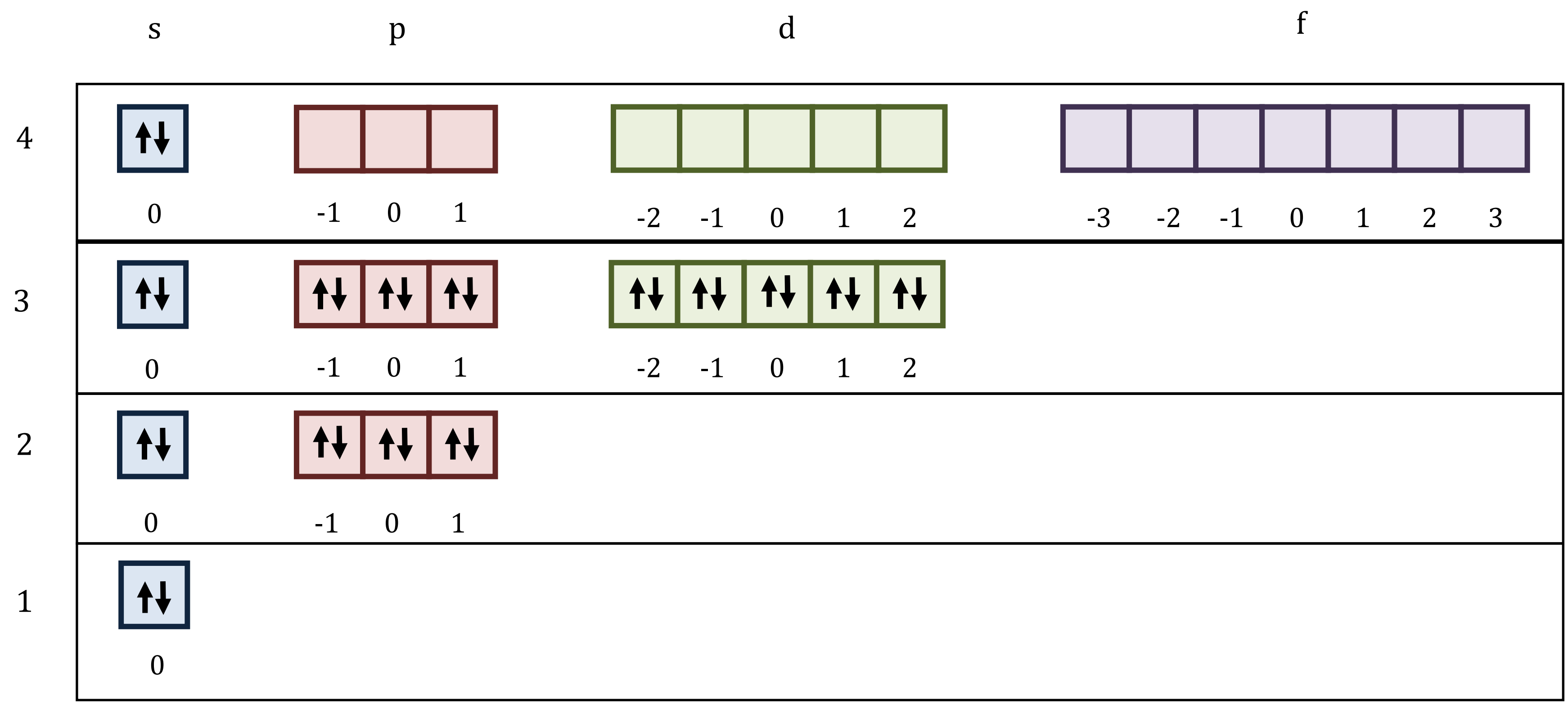
Quantum numbers

- Quantum numbers may be thought of as the way to specify an individual electron in an atom
- Like we observed before, 4 things must be specified to identify the electron
  - Shell is identified with n
  - Subshell is identified with l
  - Orbital is identified with m_l
  - Up/Down arrow is identified with m_s

n, the principle quantum number

- The shell number is equal to the principal quantum number, n

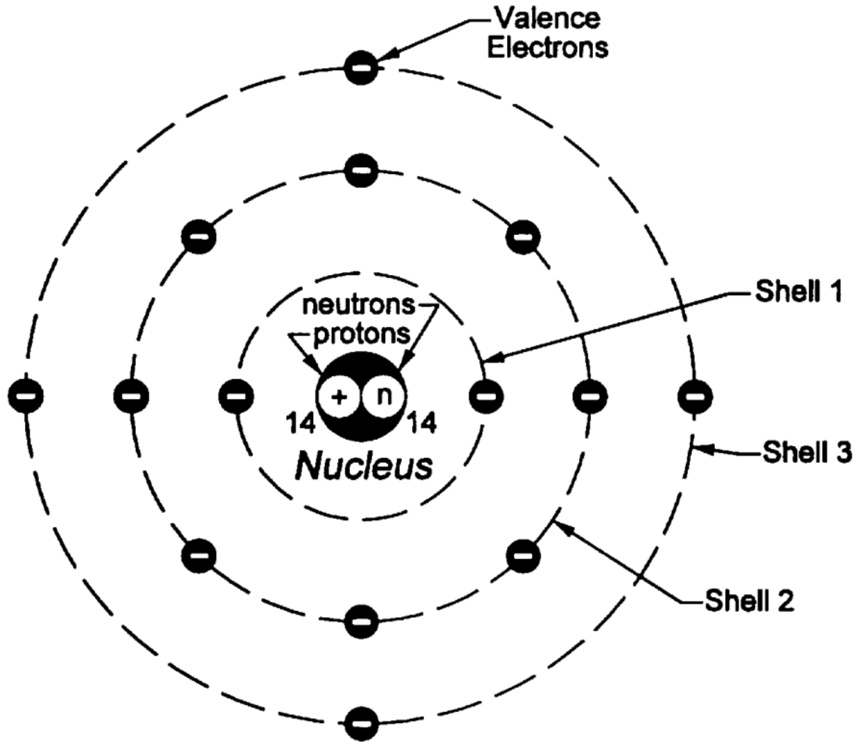
Shells, rows



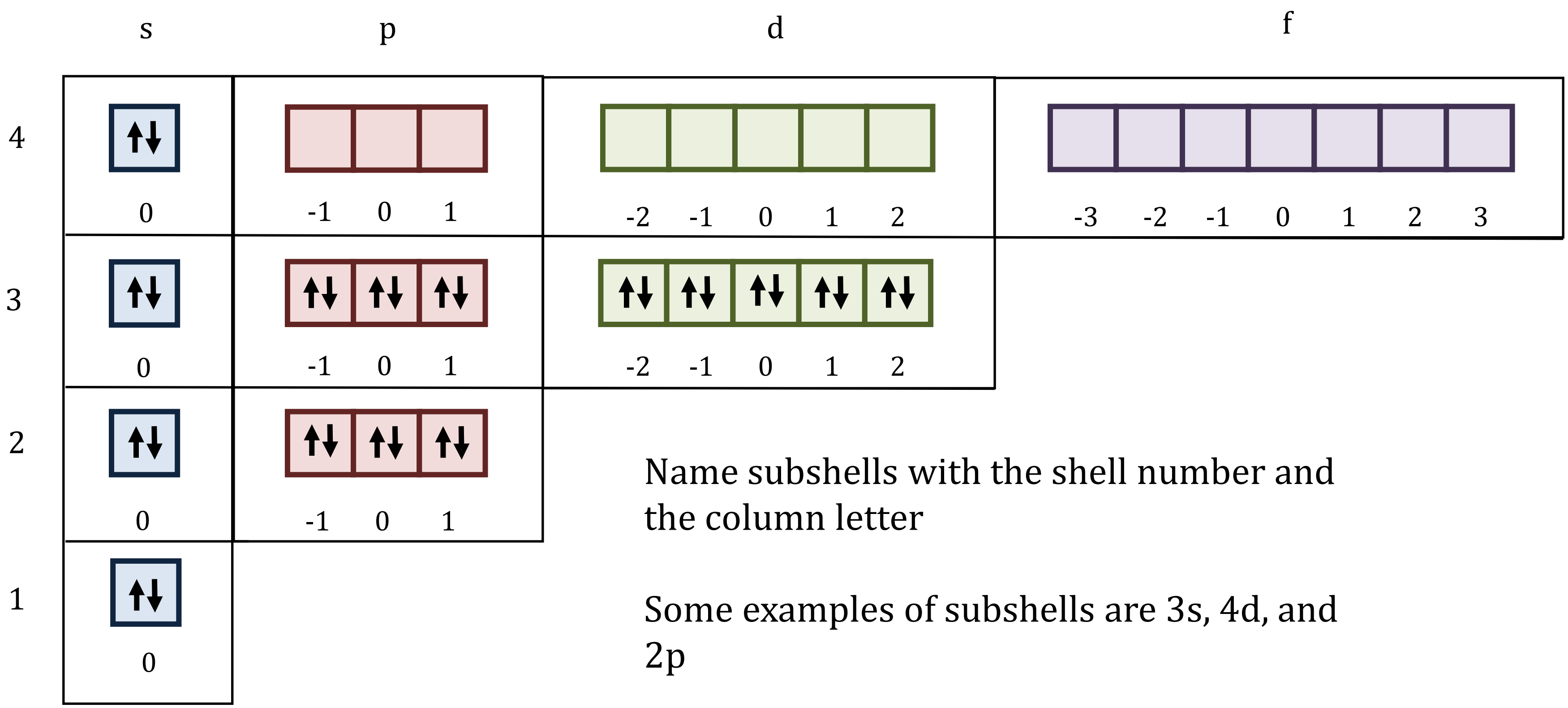
Shell numbers are indicated on the left side of the row

Shells and the Bohr model

- These are the same shells and principal quantum number (n) denoted in the Bohr model



Subshells, each cluster of boxes



Name subshells with the shell number and the column letter

Some examples of subshells are 3s, 4d, and 2p

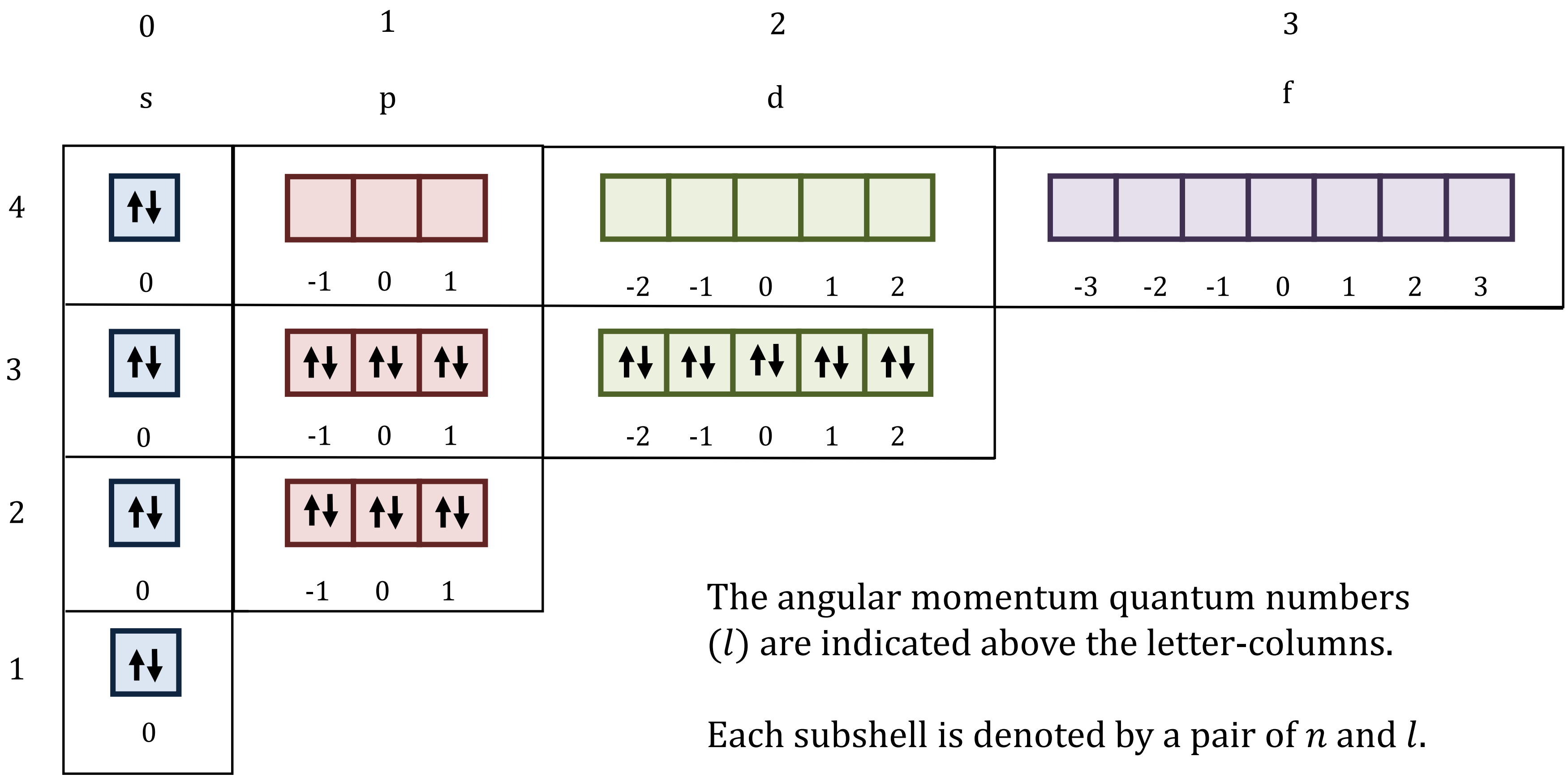
l, Angular momentum quantum number

- Designates which subshell
- *l* values are assigned as such

Subshell type	*l* value
s	0
p	1
d	2
f	3

- For any given shell n, l can have integer values from 0 to $n - 1$
  - For $n = 1$, l can be 0 only. Meaning the first shell only has an *s* orbital.
  - For $n = 2$, l can be 0 or 1. Meaning the second shell has *s* and *p* orbitals.
  - For $n = 3$, l can be 0, 1, or 2. Meaning the third shell has *s, p,* and *d* orbitals.

Subshells, each cluster of boxes

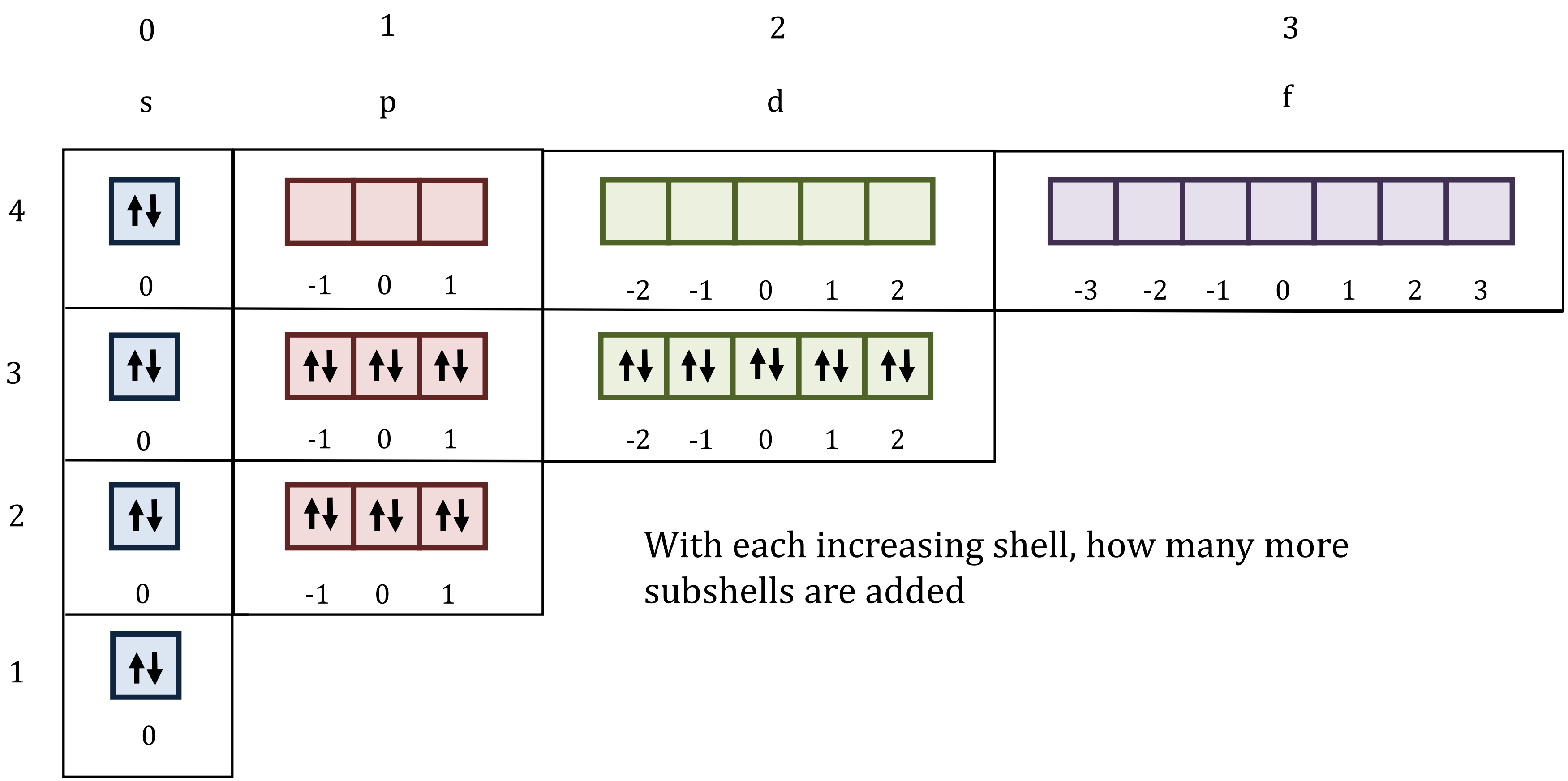


The angular momentum quantum numbers (l) are indicated above the letter-columns.

Each subshell is denoted by a pair of n and l.

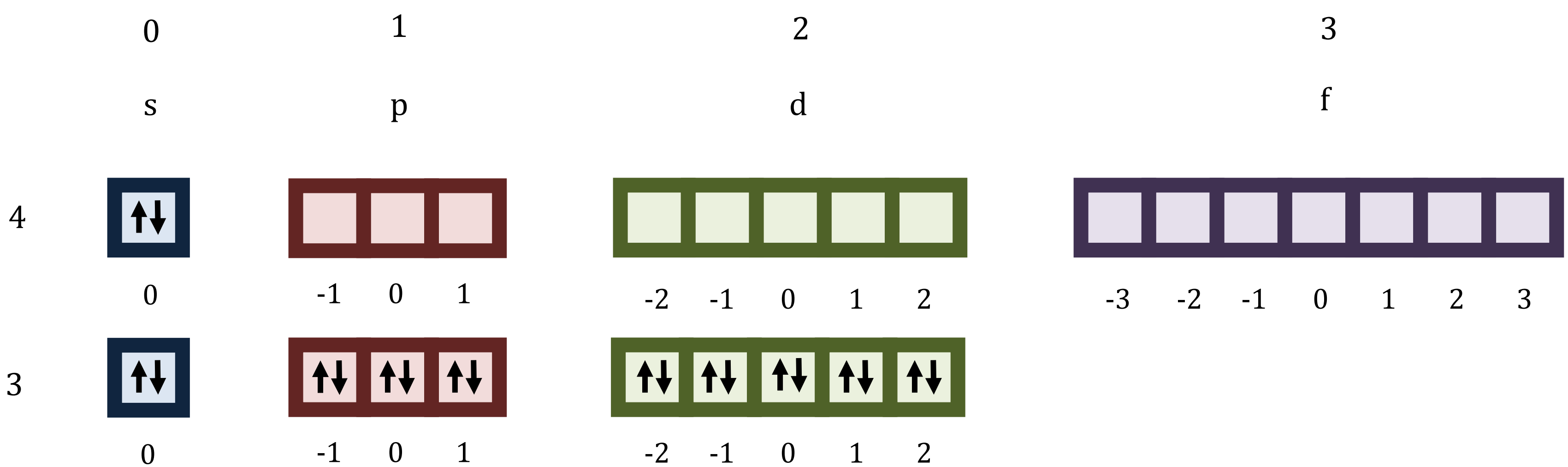
For example, 1s, 3p, 4d

Subshells, each cluster of boxes



With each increasing shell, how many more subshells are added

Orbitals, each box within a subshell



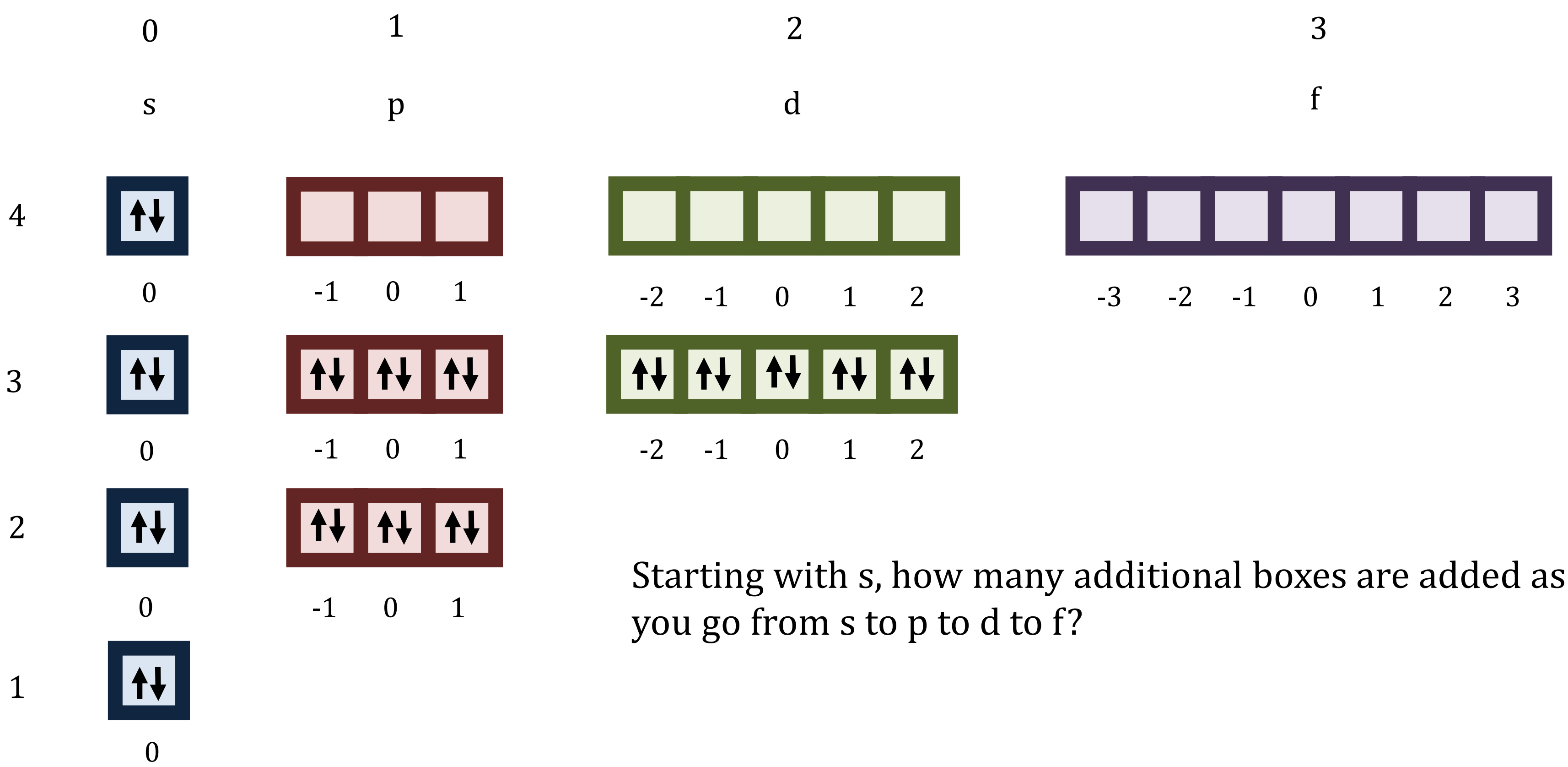
The magnetic quantum numbers (m_l) are located below the individual boxes and range from l to $-l$

Magnetic quantum number, m_l

- Designates which particular orbital within a subshell you are talking about
- For any given value of l, m_l can have values of the integers between $(-l)$ to $(+l)$

Orbital type	l value	m_l values
s	0	0
p	1	$1, 0, -1$
d	2	$2, 1, 0, -1, -2$
f	3	$3, 2, 1, 0, -1, -2, -3$

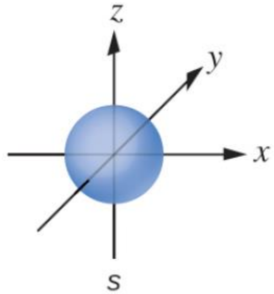
Orbitals, each box within a subshell



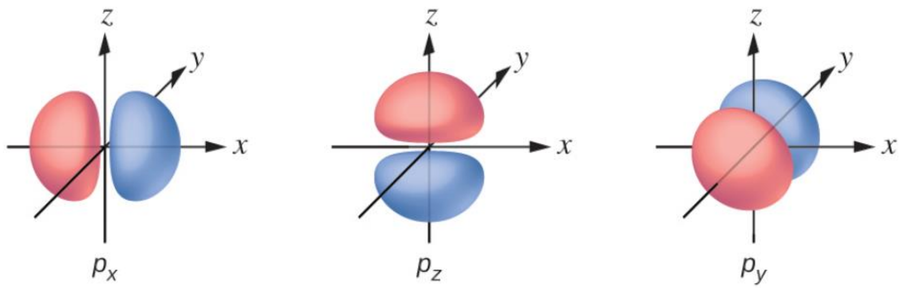
Starting with s, how many additional boxes are added as you go from s to p to d to f?

How the orbitals look

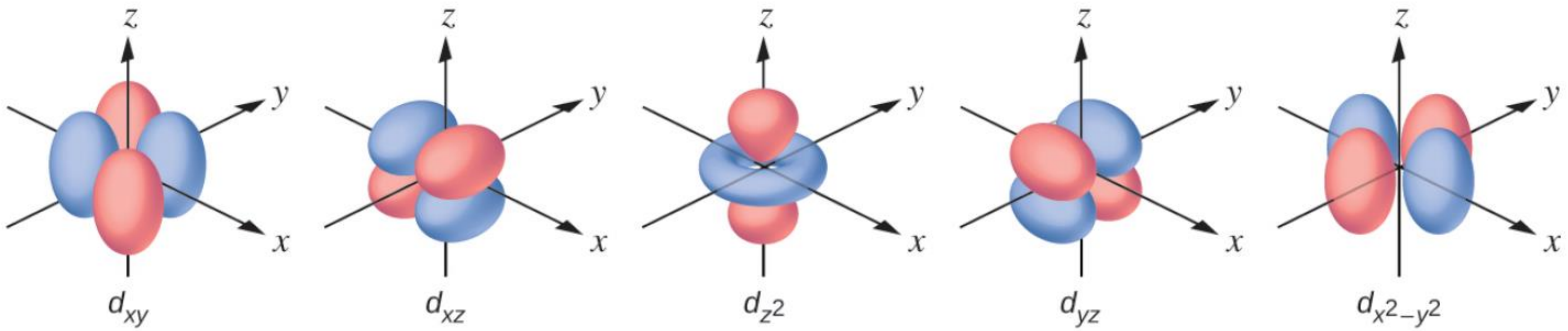
- 1 type of *s* orbital



- 3 types of *p* orbitals



- 5 types of *d* orbitals

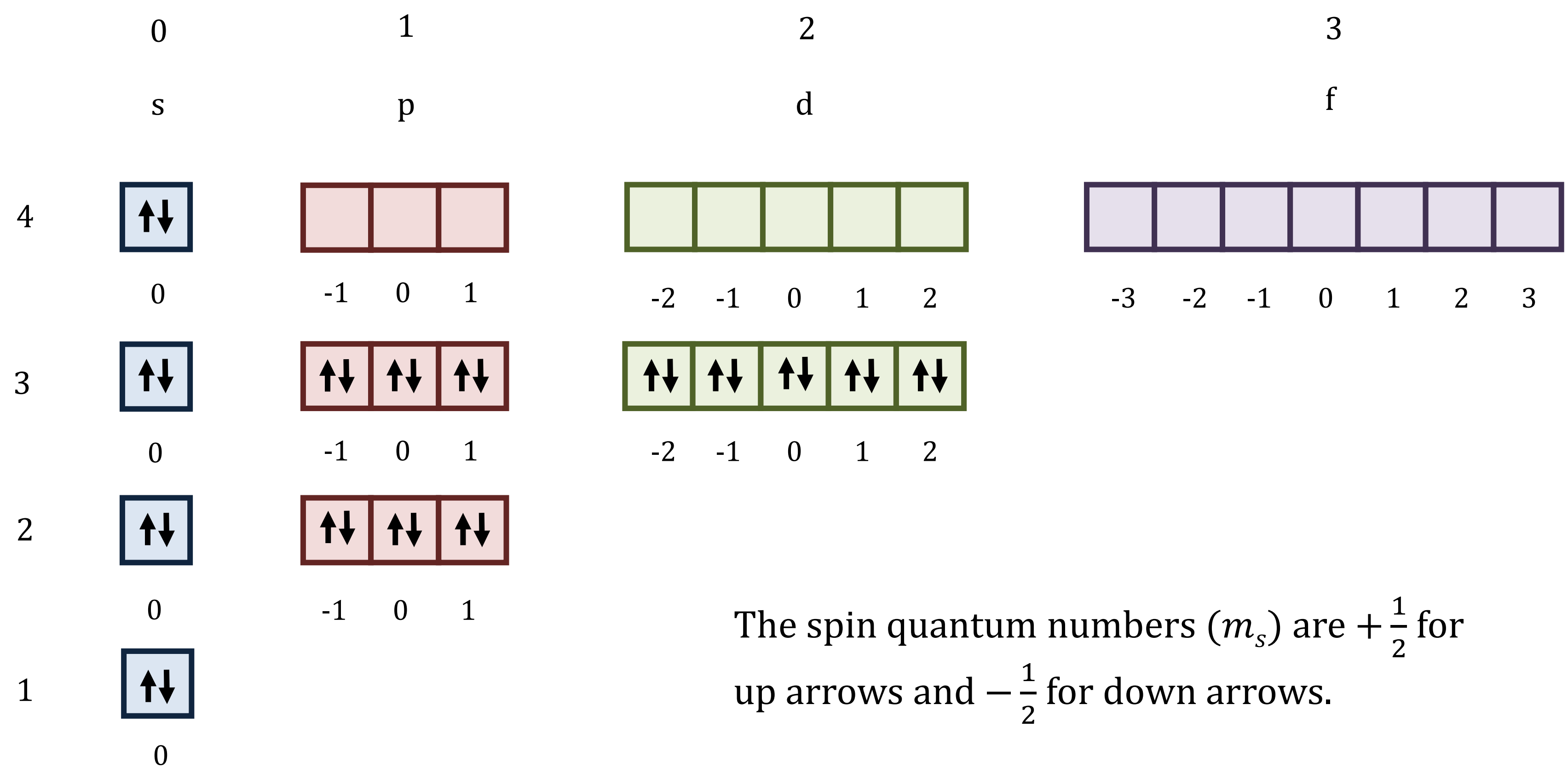


Orbitals

Which type of orbital is shaped like a sphere?

A. *s*
B. *p*
C. *d*
D. *f*

Electrons, each of the arrows



The spin quantum numbers (m_s) are $+\frac{1}{2}$ for up arrows and $-\frac{1}{2}$ for down arrows.

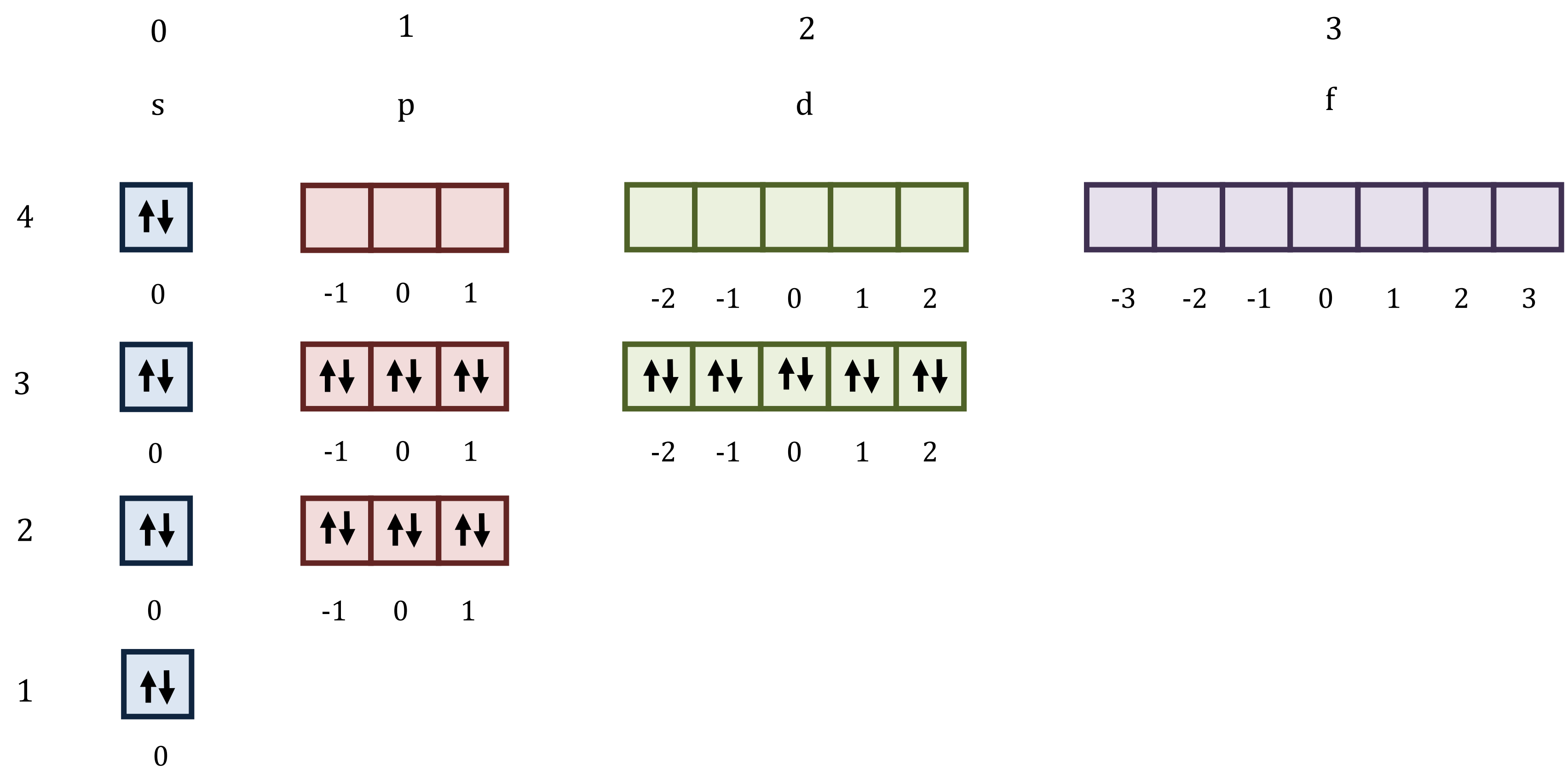
Spin quantum number, m_s

- Designates which of the electrons in an orbital you are talking about
- Describes the spin of the electron
- Electrons can be spin up or spin down
- m_s can have values of $+\frac{1}{2}$ (denoting spin up) or $-\frac{1}{2}$ (denoting spin down)
- Spin up electrons are denoted "↑" and spin down electrons are denoted "↓"

Pauli exclusion principle

- The Pauli exclusion principle states that no two electrons may have the same four quantum numbers (n, l, m_l, and m_s)
- Thus, there cannot be two spin up electrons in one orbital or two spin down electrons in one orbital.

Draw the $n = 5$ and $n = 6$ shells filled with electrons. Label the n, l, and m_l quantum numbers.



Quantum numbers

Provide all (2) possible sets of quantum numbers to describe an electron in a 3s orbital

Quantum numbers

Provide all (6) possible sets of quantum numbers to describe an electron in a $2p$ orbital

Quantum numbers

State which of the following sets of quantum numbers would be possible for an electron in an atom.

1. $n = 2, l = 0, m_l = 0, m_s = +\frac{1}{2}$
2. $n = 1, l = 1, m_l = 0, m_s = -\frac{1}{2}$

A. 1
B. 2
C. 1 & 2
D. Neither 1 nor 2

Quantum numbers

Answer the following questions concerning quantum numbers.

I. What is the maximum number of electrons that can occupy an $l = 5$ subshell?
II. How many subshells are in the $n = 4$ shell?
III. What is the maximum number of electrons that can occupy the $n = 3$ shell?
IV. How many orbitals are in an $l = 4$ subshell?