Unit 7 - IB

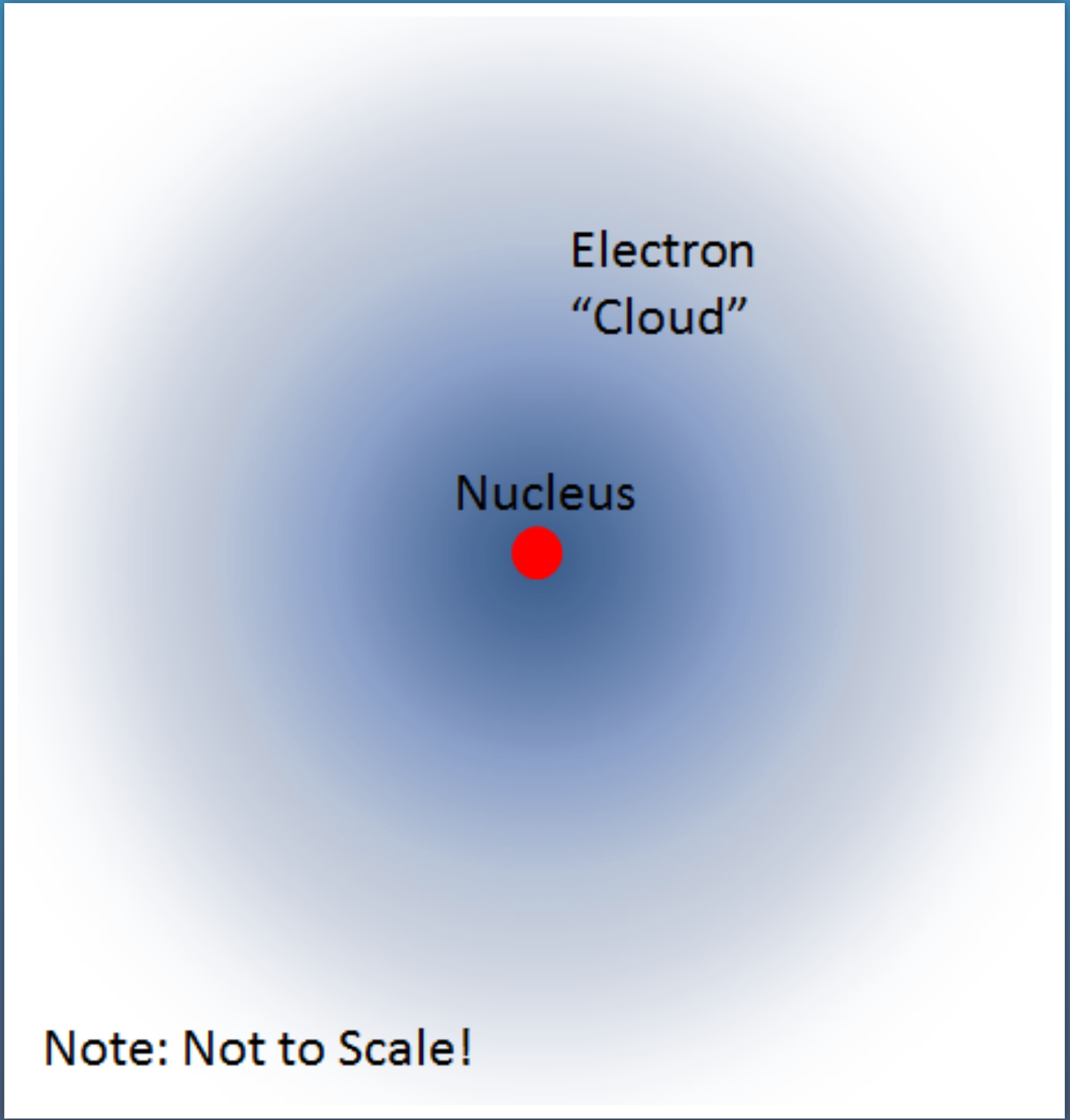
Atomic Structure & Periodic Table

SOLVAY CONFERENCE 1927

Quantum Mechanical Model

- Each orbital can only hold 2 electrons, each with opposite spins

Pauli Exclusion Principle



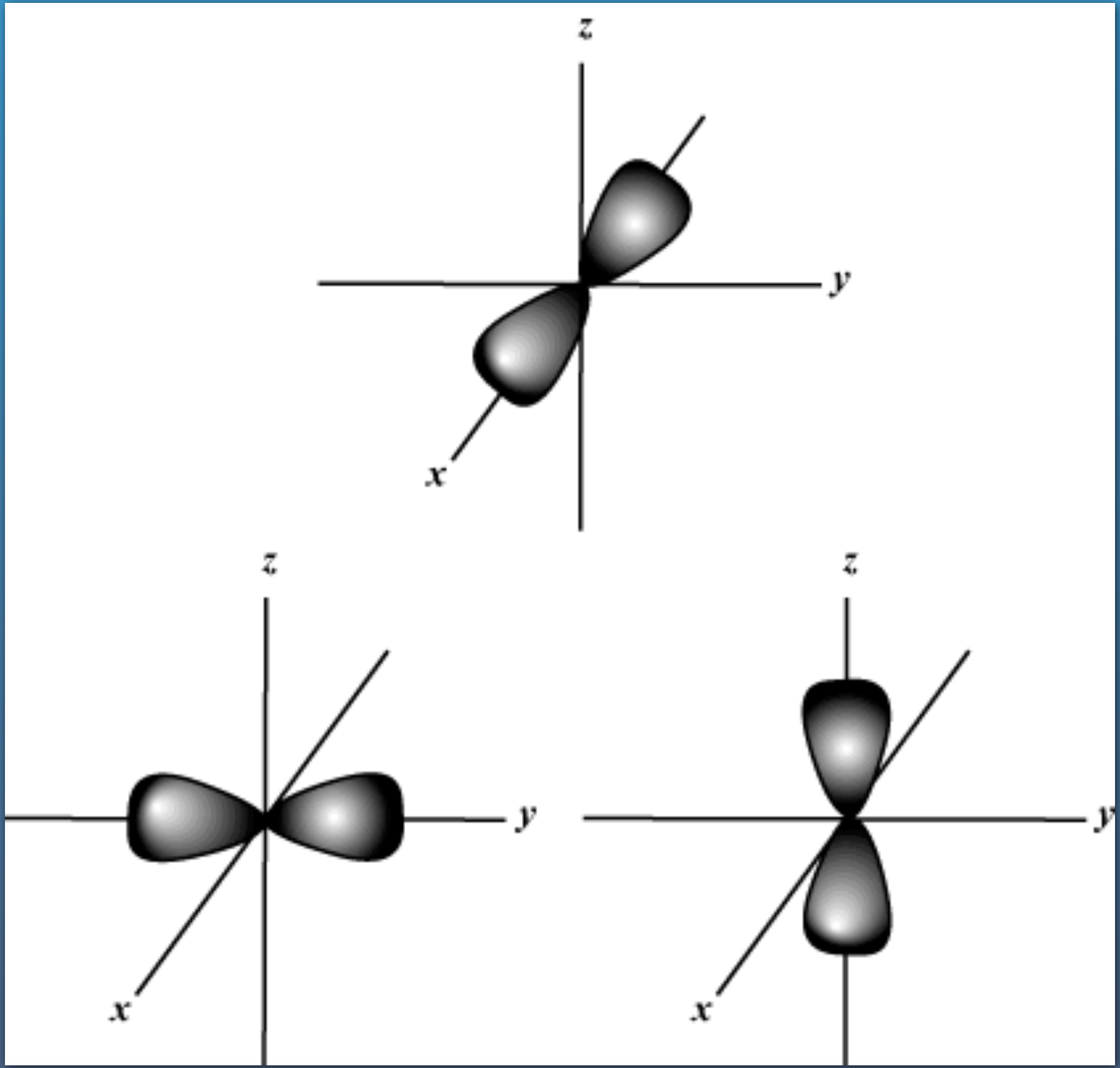
Electron Sub-Energy Levels (s-orbital)

Electrons fill orbitals
<u>lowest energy first</u>

spherical in shape
and holds only 2 e^-
total.

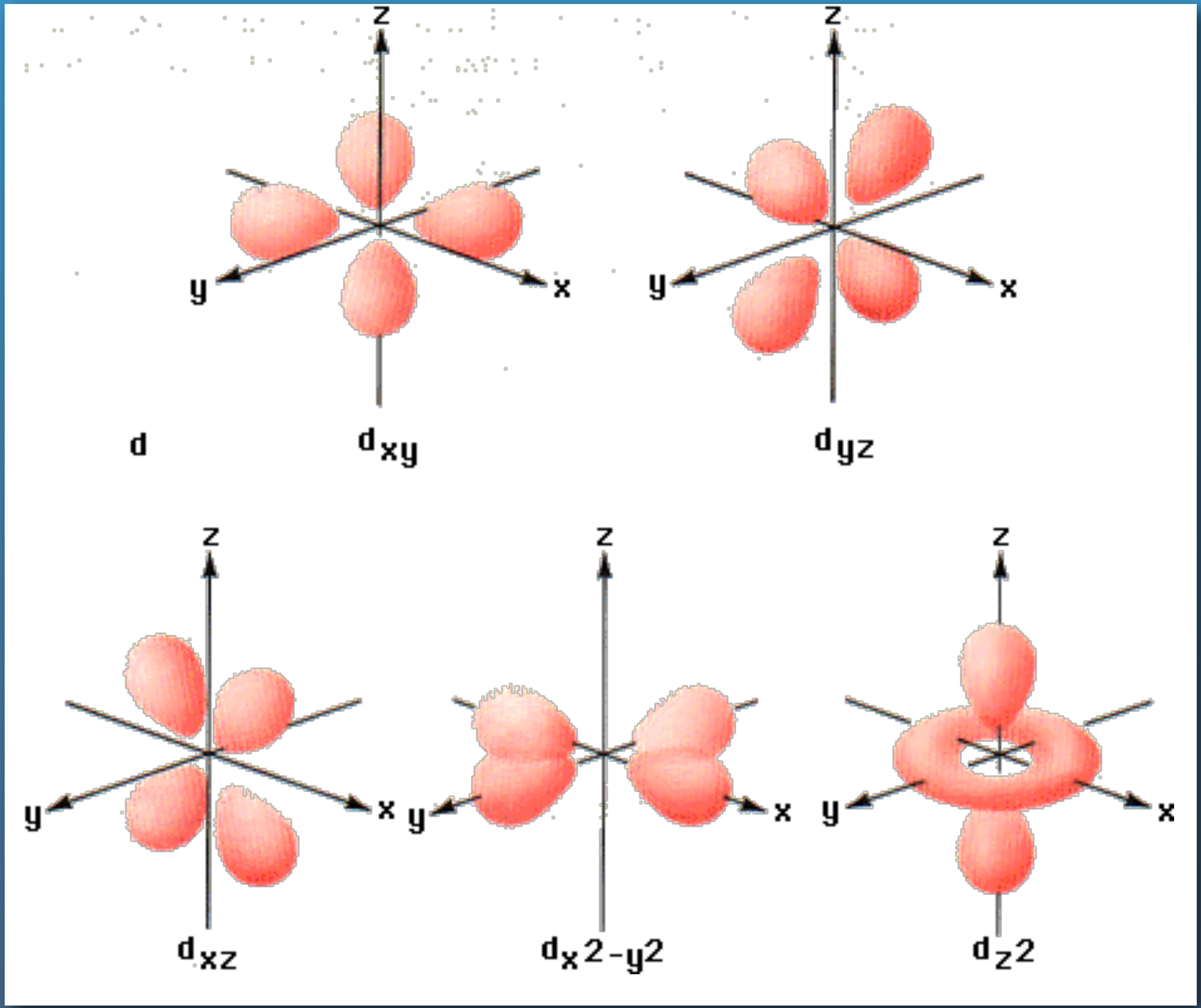
p-Orbital

- (x, y, z) are dumbbell shaped
- Each holds 2 e^-
- hold up to 6 e^- total



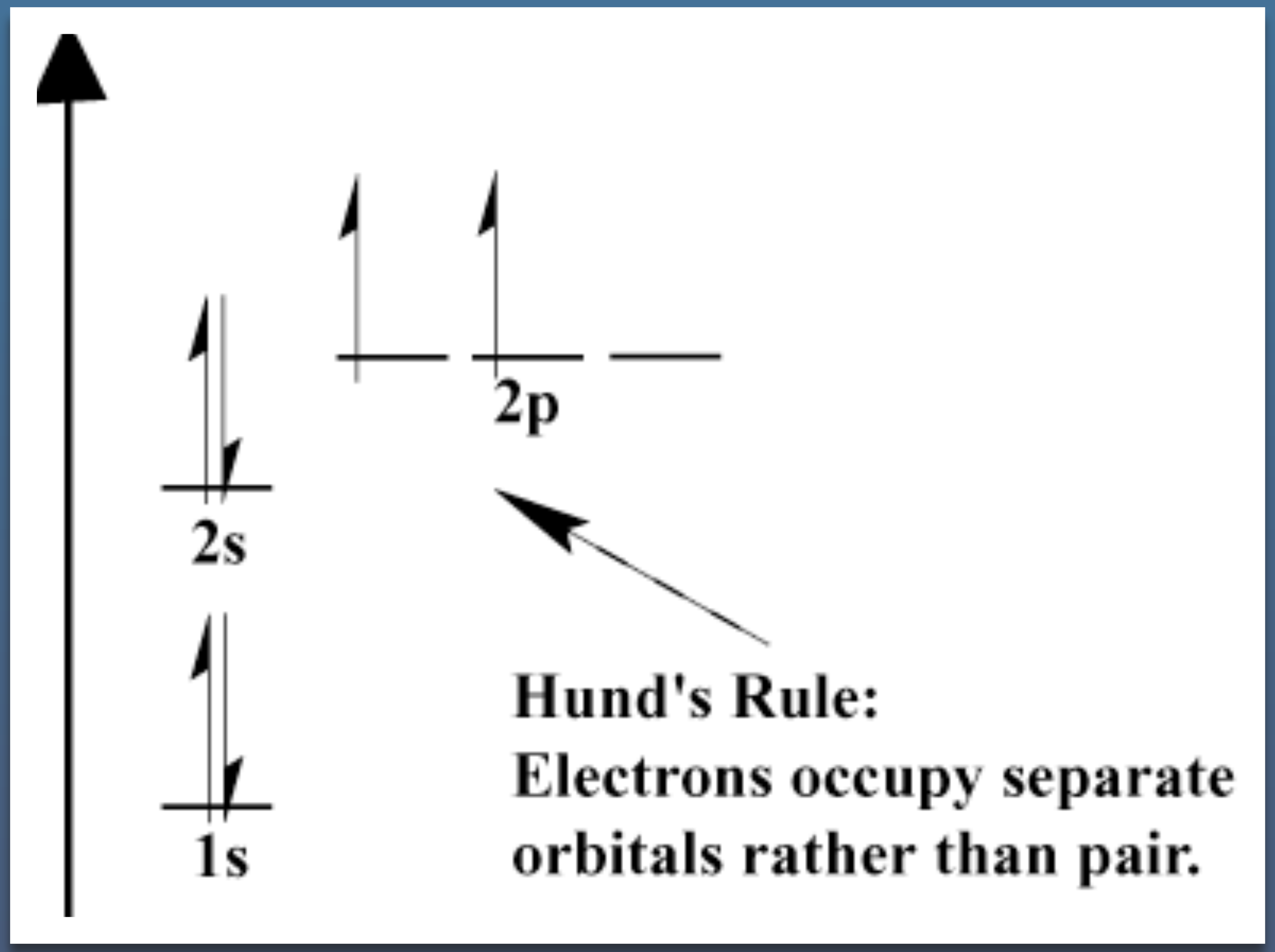
d-Orbitals

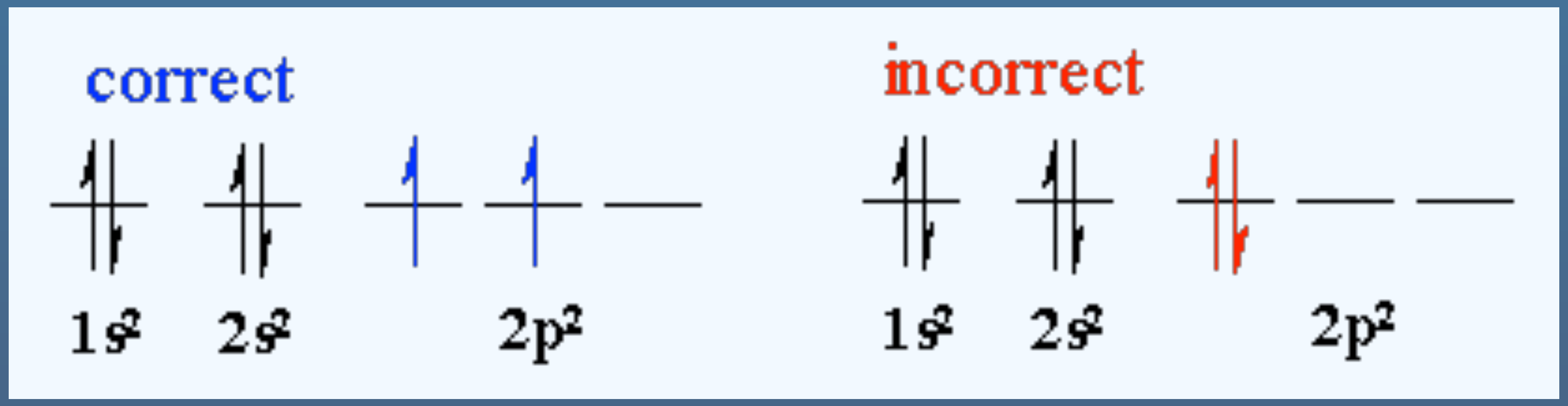
- 4-leaf clover shaped
- 5 types of d-orbitals; each holding 2 electrons
- total of 10 e-'s



Hund's Rule!

- Sub-shells (s, p, d) are most stable when they are half full or completely filled with electrons.
- Electrons fill orbitals one electron at a time (because they repel)
- All seats get filled with one person each first, then they double up.





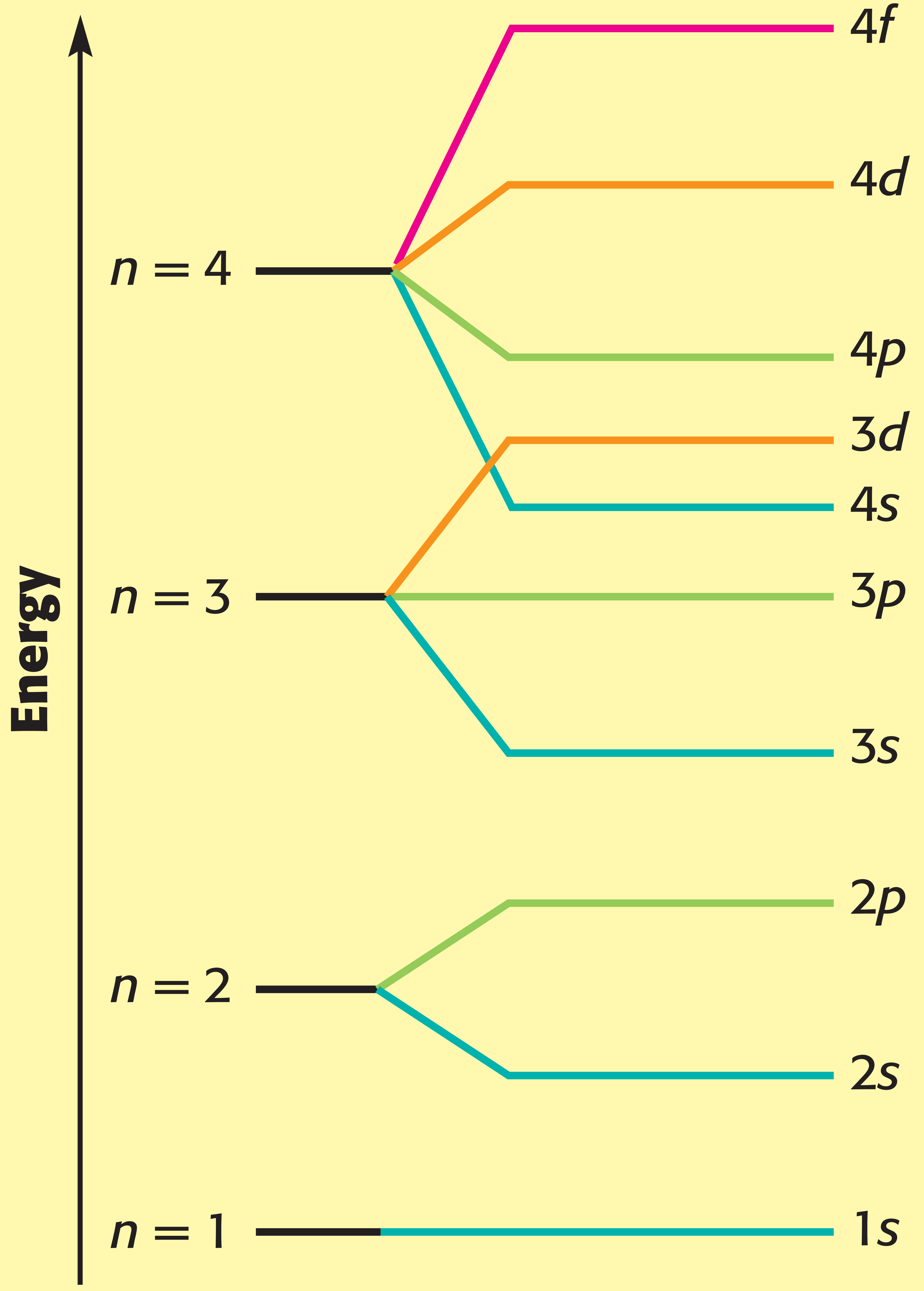
Electron Configuration Continued

Example: $1s^2$

- 1 = energy level, s = orbital type, 2 = # of e^- are in it.
- Writing electron configuration: go from left to right across periods of the periodic table, write all symbols from each **'block'** (***s, p, d, or f***)
- Li = $1s^2 2s^1$
- Na = $1s^2 2s^2 2p^6 3s^1$
- Ti = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$

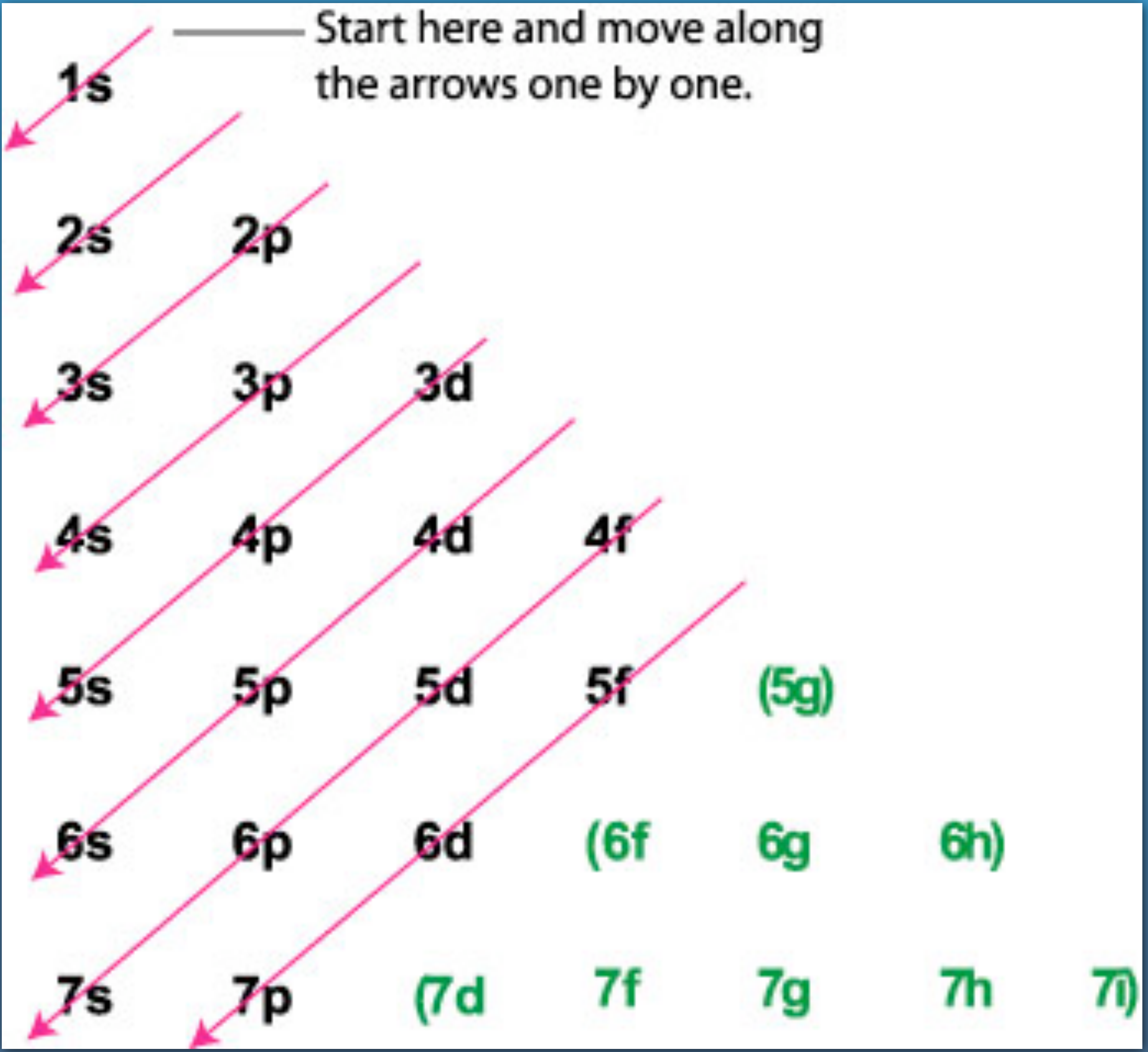
Aufbau Principle

- Aufbau is the German word for 'building up'.
- Electrons fill orbitals that have the lowest energy first.



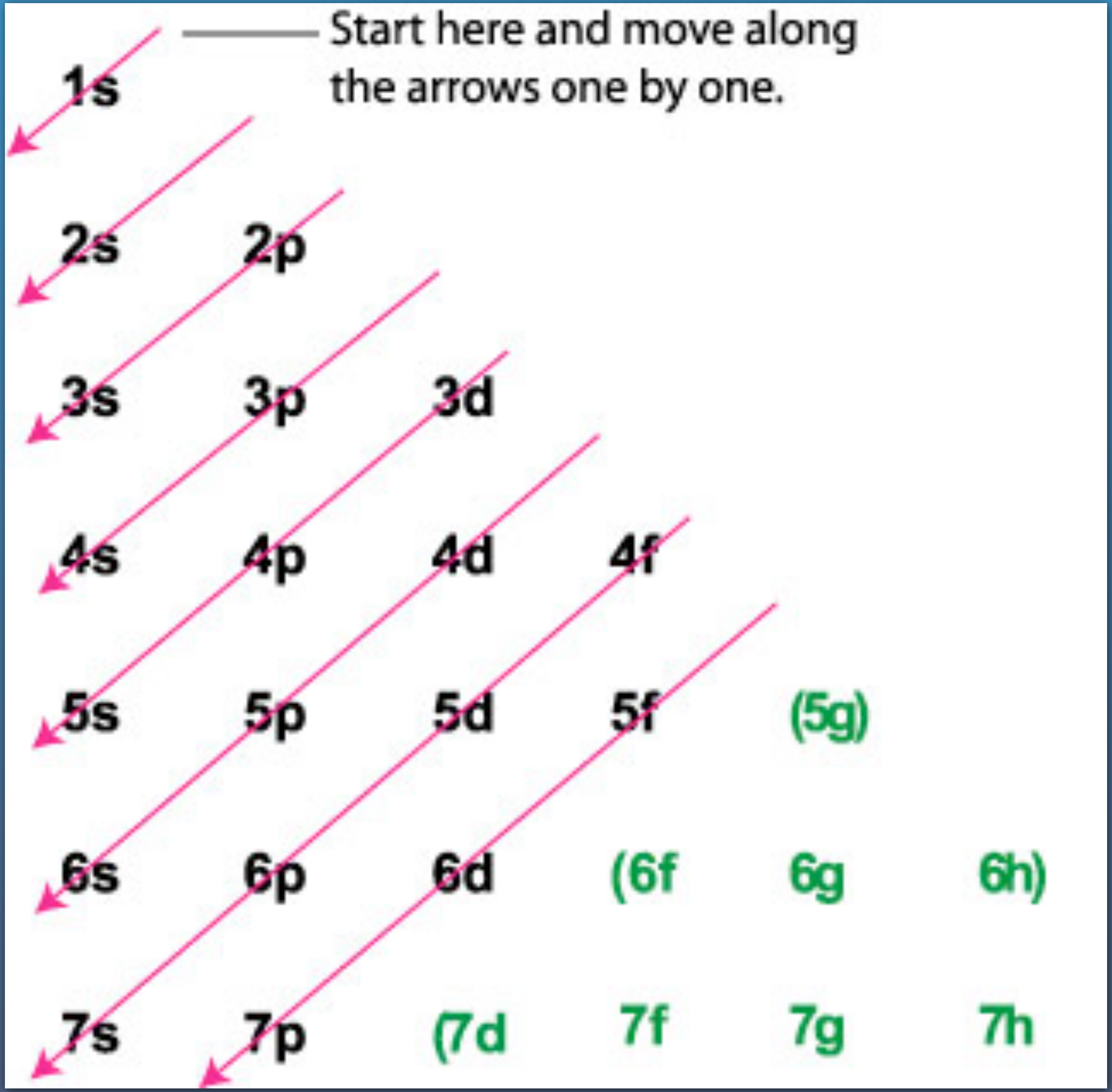
Another Approach

- Can also 'draw' electron configuration as levels with arrows to represent electrons.
- Ex: sodium (11 electrons)
- 1s ↑↓ 2s ↑↓ 2p ↑↓ ↑↓ ↑↓ 3s ↑_



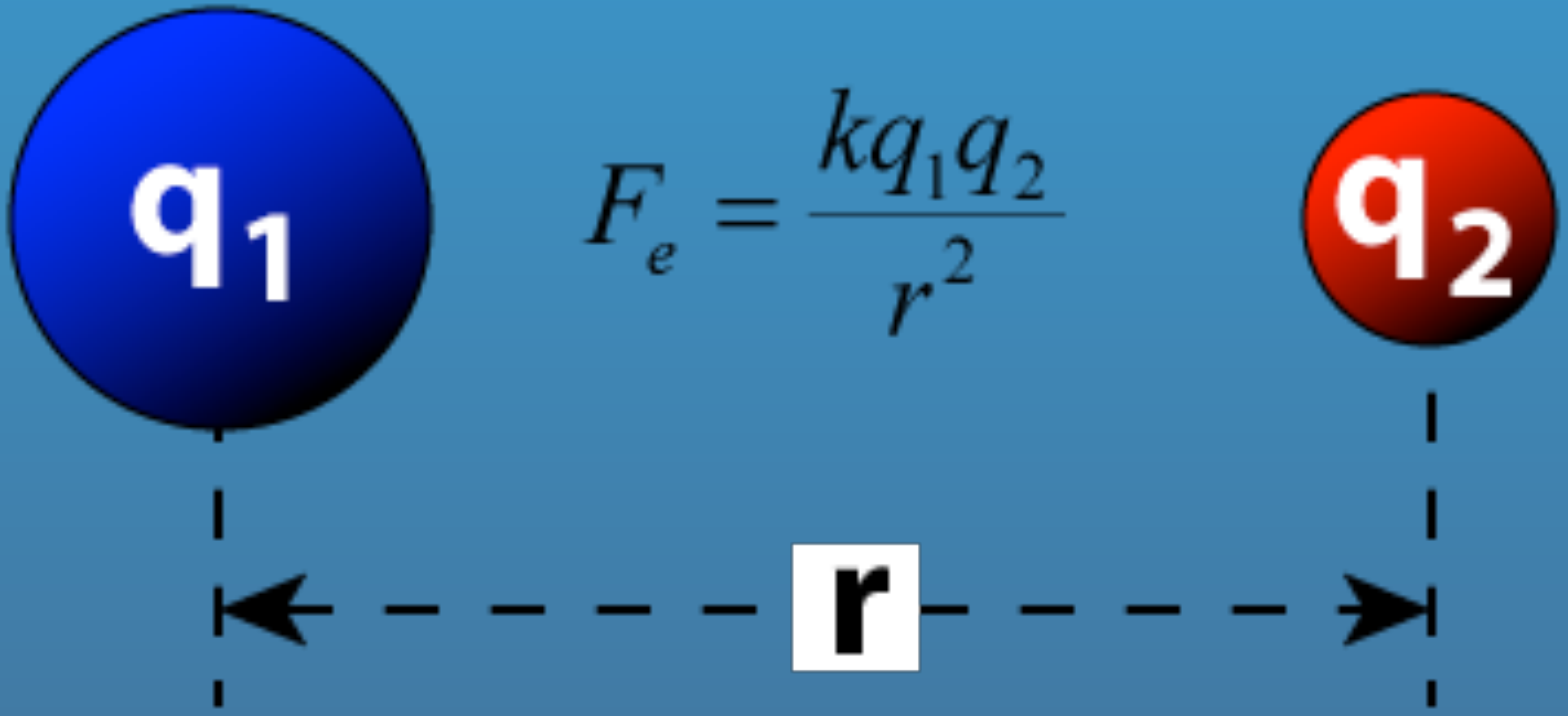
Let’s try some other examples

- Nitrogen
- Oxygen
- Magnesium
- Bromine
- Iron



Coulomb's Law

$$F = \frac{k\,Q_1 Q_2}{r^2}$$



The force of attraction or repulsion (F) between 2 particles is dependent on the product of the charges of the particles (Q_1 and Q_2) divided by the distance between the two particles (r) squared.
*(k is a constant dependent on the nature of the particles.)

Extras!

- Isoelectronic: Same # of electrons
- Noble Gas orbital notation: $[Ne]3s^1$
- Effective Nuclear Charge: Coulomb's Law, practical applications

Effective Nuclear Charge

- Nuclear Charge: Given by the atomic number and increases by one as you go across a period.
- Outer electrons do not feel all of this attractive force because they are ‘shielded’ by the inner electrons.
- Therefore, the ‘effective’ charge the outer electrons feel is less than the nuclear charge (# of protons).
- Let’s look at an example…

Effective Nuclear Charge

- Consider, for example, a sodium atom. The nuclear charge is given by the atomic number of element ($Z = 11$). The outer electron in the 3s orbital is, however, shielded from these 11 protons by the 10 electrons in the first and second principal energy levels ($1s^2 2s^2 2p^6$).

Element	Na	Mg	Al	Si
Nuclear Charge	11	12	13	14
Electron Configuration	$[Ne]3s^1$	$[Ne]3s^2$	$[Ne]3s^2 3p^1$	$[Ne]3s^2 3p^2$
Effective Nuclear Charge	+1	+2	+3	+4
Atomic Radius	160	140	124	114