Unit 2

UNDERLYING STRUCTURE OF MATTER

USMLT1: Use standard atomic notation to represent and describe atoms and isotopes. Calculate atomic mass.

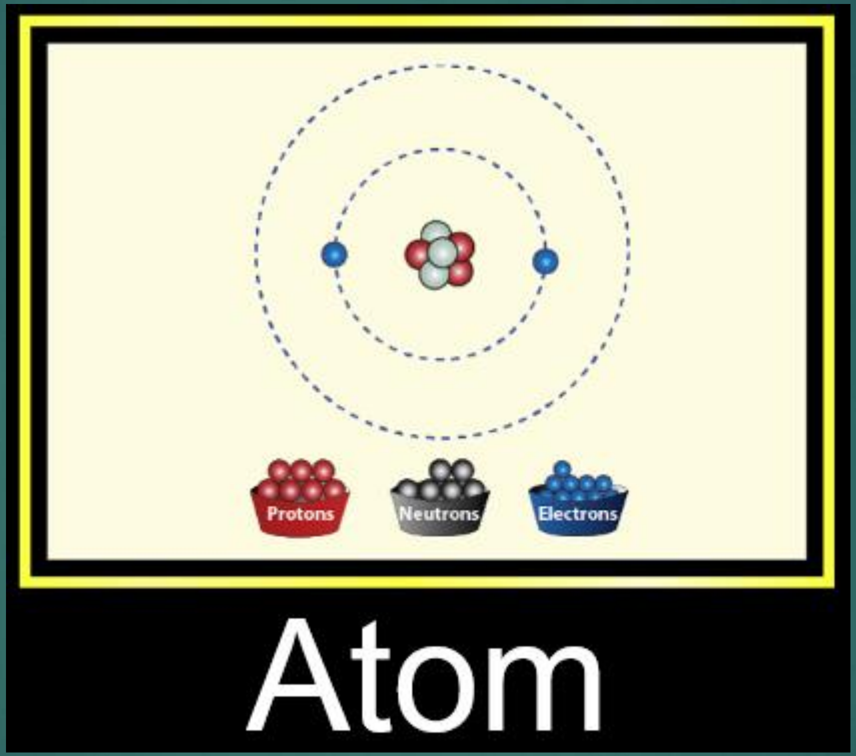
Be able to define, explain, identify or provide examples of each of the following:

- Atoms
- Protons
- Electrons
- Neutrons
- Isotope
- Atomic mass
- Atomic number
- Mass number
-

Textbook Practice

- Page 112 # 17
- Page 117 #s 23, 24
- Page 119 #s 25, 27, 30 – 32
- Page 122 – 123 #s 47, 49 – 52, 55, 65

The atomic nucleus



Subatomic Particles

Particle	Charge	Mass (g)	Location
Electron (e^-)	-1	9.11×10^{-28}	Electron cloud
Proton (p^+)	+1	1.67×10^{-24}	Nucleus
Neutron (n^o)	0	1.67×10^{-24}	Nucleus

Complete Symbols

- Contain the symbol of the element, the mass number and the atomic number.

Superscript → Mass number

Total # of protons and neutrons

X

Subscript → Atomic number

# of protons

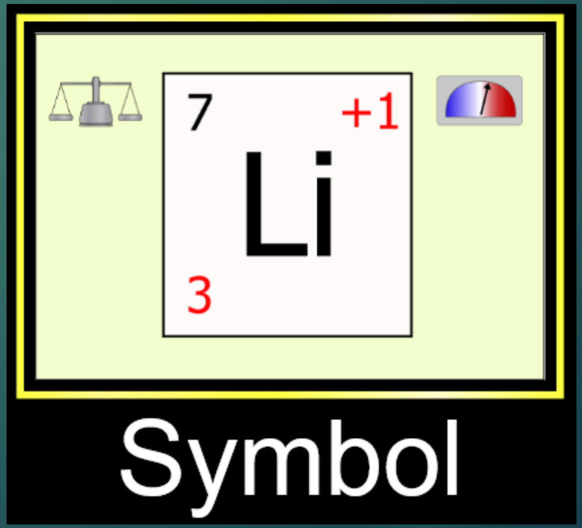
Information from Symbols

- Find each of these:
  a) number of protons
  b) number of neutrons
  c) number of electrons
  d) Atomic number
  e) Mass Number

$^{80}_{35}Br$

Isotopes

- Multiple atoms that have the same number of protons, but a different number of neutrons. i.e. the same atomic number but different mass numbers.
- We can also put the mass number *after* the name of the element:
  - carbon-12
  - carbon-14
  - uranium-235



Atomic Mass

- How heavy is an atom of oxygen?
  - It depends, because there are different *kinds* of oxygen atoms.
- We are more concerned with the *average atomic mass.*
- This is based on the abundance (percentage) of each variety of that element in nature.
  - We don’t use grams for this mass because the numbers would be too small.

Measuring Atomic Mass

- Instead of grams, the unit we use is the Atomic Mass Unit (amu)
- It is defined as one-twelfth the mass of a carbon-12 atom.
- Each isotope has its own atomic mass, thus we determine the average from percent abundance.

Atomic Masses

Atomic mass is the average of all the naturally occurring isotopes of that element.

Isotope	Symbol	Composition of the nucleus	% in nature
Carbon-12	^{12}C	6 protons 6 neutrons	98.89%
Carbon-13	^{13}C	6 protons 7 neutrons	1.11%
Carbon-14	^{14}C	6 protons 8 neutrons	<0.01%

Carbon = 12.011

Calculating atomic mass

- The two most abundant isotopes of carbon are carbon-12 (mass 12.00 amu) and carbon-13 (mass 13.00 amu). Their relative abundances are 98.9% and 1.10%, respectively. Calculate the atomic mass of carbon.

Another example

- Using the information below, calculate the approximate average atomic mass of silicon.

$^{28}_{14}Si$	$^{29}_{14}Si$	$^{30}_{14}Si$	$^{31}_{14}Si$
92.2%	4.7%	3.1%	trace

Approximating Relative Abundance

- Copper, Cu, forms naturally with 34 and 36 neutrons. If its average atomic mass is 63.546, calculate the relative abundance found naturally.

Review Questions

- See Learning Target Guide

USMLT2: Describe model of the atom over the past 100 years and compare them to the current quantum mechanical model. Explore and summarize Rutherford's experiment.

Be able to define, explain, identify or provide examples of each of the following:

- Dalton's Model
- Thompson's Model
- Rutherford's Model
- Bohr Model
- Quantum Mechanical Model
- Quantum
- Energy Level
- Orbital
- Orbital Shape

Textbook Practice

- Page 108 #s 9, 12 – 14
- Page 122 – 124 #s 42, 43, 45, 74, 76
- Page 132 #s 1 – 7
- Page 149 #s 22 – 29

The Atom

- The smallest part of an element.
  - If you could zoom in on elements, like iron, oxygen, helium, plutonium, etc., you would see the atoms that make up that element.
- Theorized by Democritus around 2500 years ago.
  - Not based on a scientific investigation.
  - Could not explain chemical properties of matter.
  - Would remain undeveloped until the early 1800s.

John Dalton's Atomic Theory: 1803

1. All elements are composed of tiny indivisible particles called atoms.

2. Atoms of the same element are identical. The atoms of any one element are different from those of any other element.

John Dalton's Atomic Theory: 1803

3. Atoms of different elements can physically mix together or chemically combine in simple whole-number ratios to form compounds (like H_2O, CO_2).

4. Chemical reactions occur when atoms are separated, joined, or rearranged. Atoms of one of the element, however, are never changed to atoms of another element as a result of a chemical reaction. (Nuclear reactions change atoms from one type to another – happens in the naturally in the Sun and on Earth.

Incorporation of Electrons

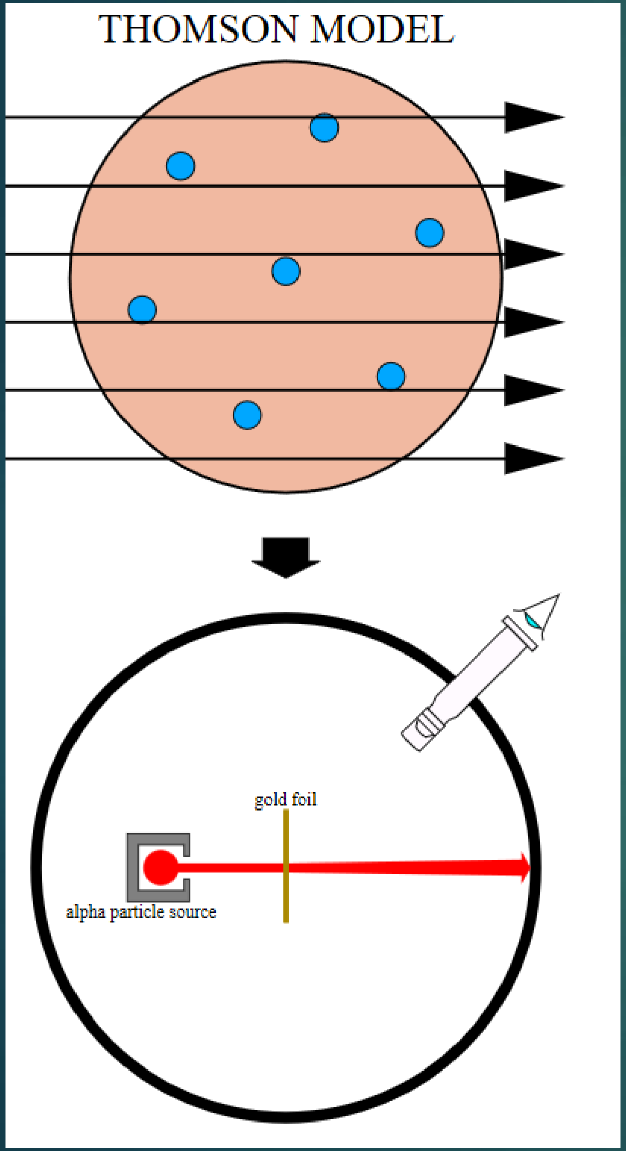
- Electrons, as particles, were first theorized in 1897 by English physicist J.J. Thompson. He invented the cathode ray tube to test for charges. That work eventually became the CRT television.

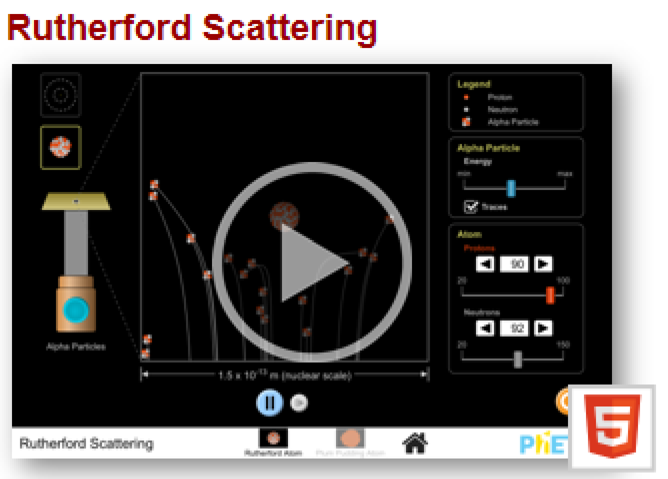
- Thompson adjusted the model of the atom to incorporate electrons; he proposed the atom is a lump of positive charge with electrons evenly spaced within it – dubbed the "plum pudding" model of the atom.

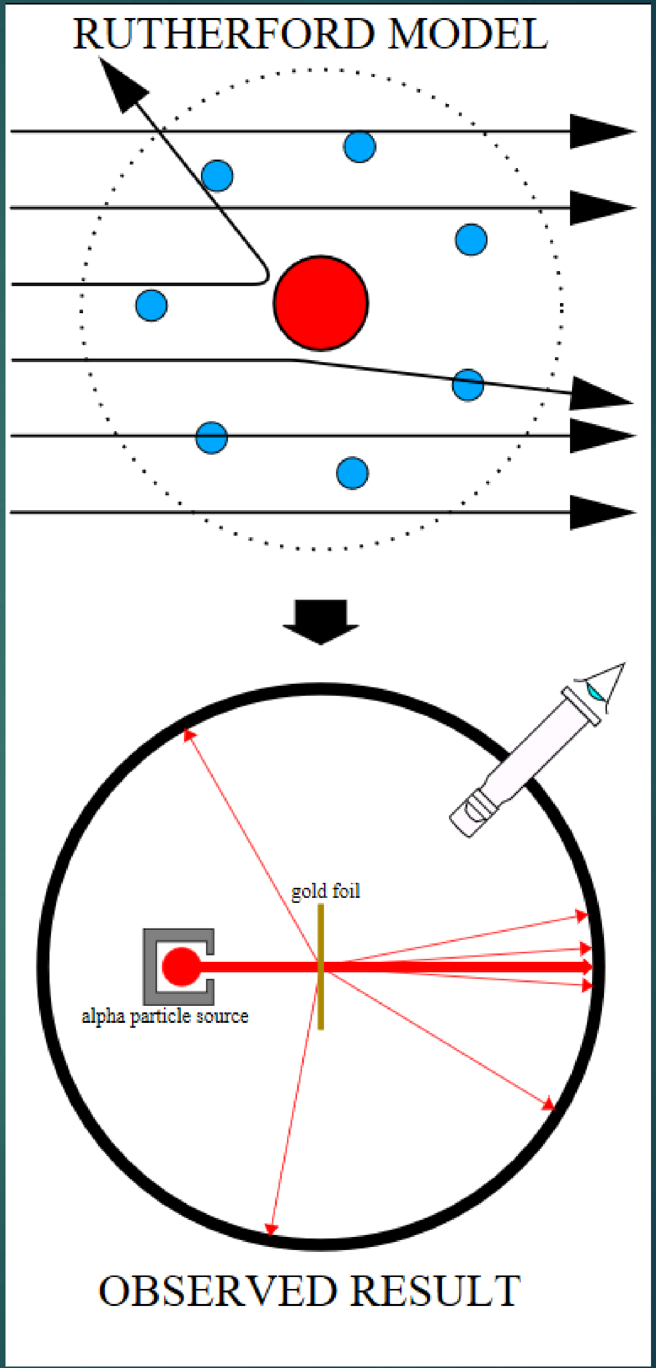
Discovery of the Nucleus

- Earnest Rutherford and coworkers at University of Manchester, England, were the first to theorize, based on experimental evidence, the existence of the atomic nucleus.
- In 1911 he performed the "Gold-Foil" experiment.
- His discovery changed the model of the atom significantly – the first evidence of the positively charged atomic nucleus and that atoms are mostly empty space.

Rutherford Experiment

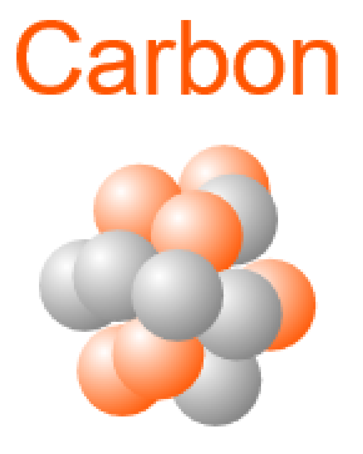


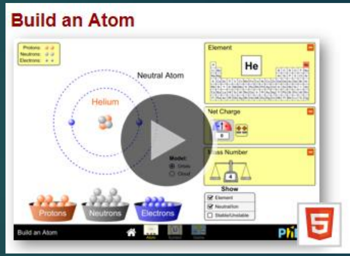




The Atomic Nucleus

- Rutherford's experiment confirmed the presence of a small, dense, area of positive charge. The term ***proton*** was used to name the unseen positive particles. Also, that atoms are mostly empty space.
- It would be 21 years later, in 1932, when physicist James Chadwick discovered the ***neutron***, which also exists within the nucleus to keep protons apart. The neutron is neutral in charge (a charge of zero). The # of neutrons does not have to equal the # of protons in an atom.
- Protons (p^+) and neutrons (n^0) are very close to the same size and mass. Both have a much, much larger mass than the electron (e^-).





The Nucleus: Summary Video

The Nucleus: Crash Course Chemistry

CrashCourse

- But what about ***electrons***?
- Chemical and physical properties are the result of electrons in the atom.

Atomic Structure – Electron Orbitals

- 1897: Thompson theorized electrons were static in a clump of positive charge.
- 1911: Rutherford’s experimental results support the nucleus and he agreed that electrons orbit the nucleus like planets around the Sun. However, it could not explain properties of elements, like why heated metal glowed red/orange.
- 1913: New Zealand physicist Niels Bohr adjusts the model such that electron’s have fixed distances from the nucleus, but that electrons can change where they are in the atom by gaining or losing energy.

The Bohr Model of the Atom

Niels Bohr

Electrons orbit the nucleus much like planets orbiting the sun.

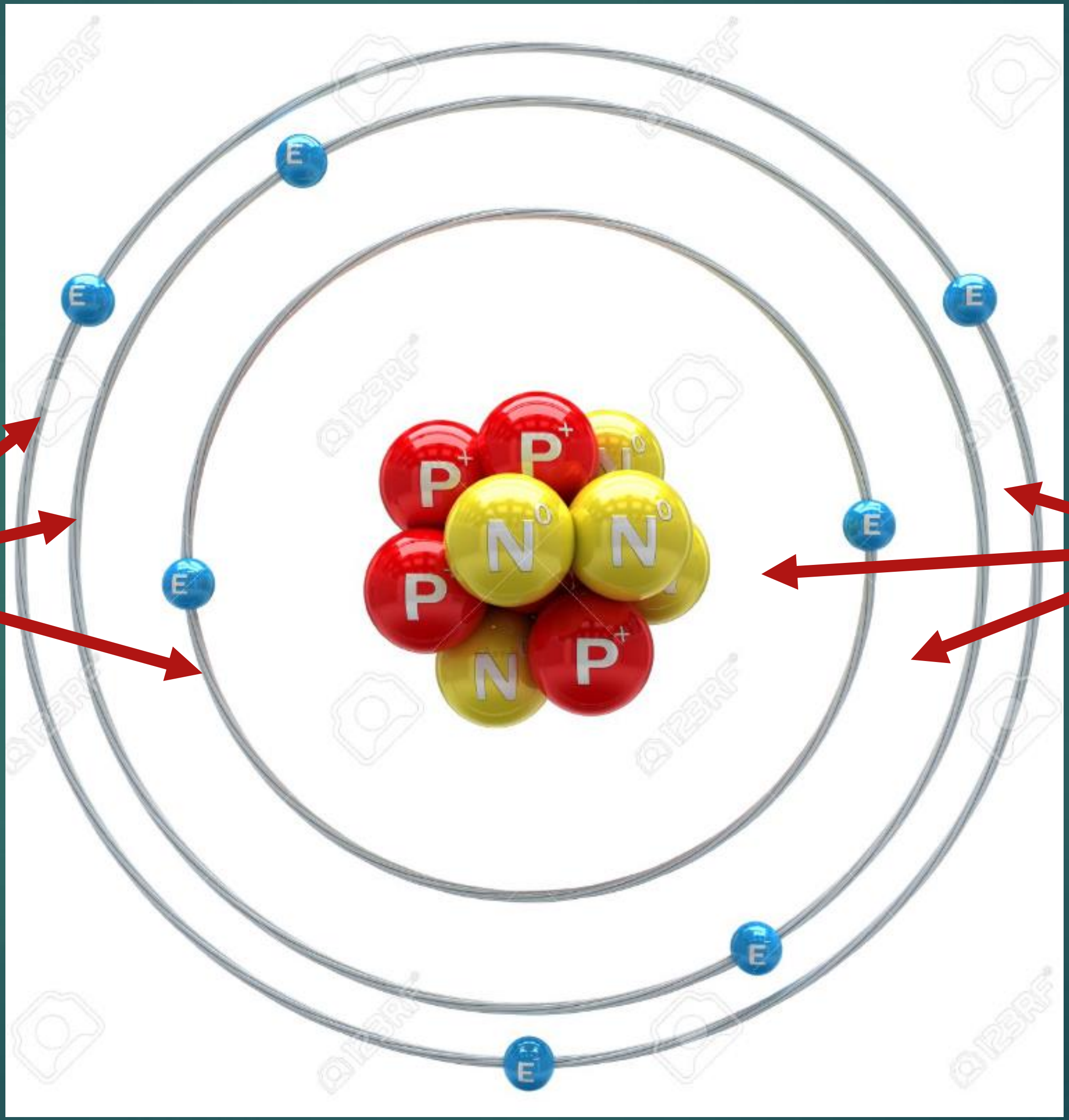
However, electrons are found in specific circular paths around the nucleus and can jump from *one level to another*.

Bohr Model of the Atom

- Explains observations of light coming from the simplest element, hydrogen, but failed for larger atoms, like metallic elements change color when heated.
- The energy electrons have is ***quantized***, they can only have a specific amount of energy at each orbit and they cannot be found at any other orbit.
- Electrons gain or lose a ***quantum*** of energy to change orbital locations around the nucleus.

Bohr Model of the Atom

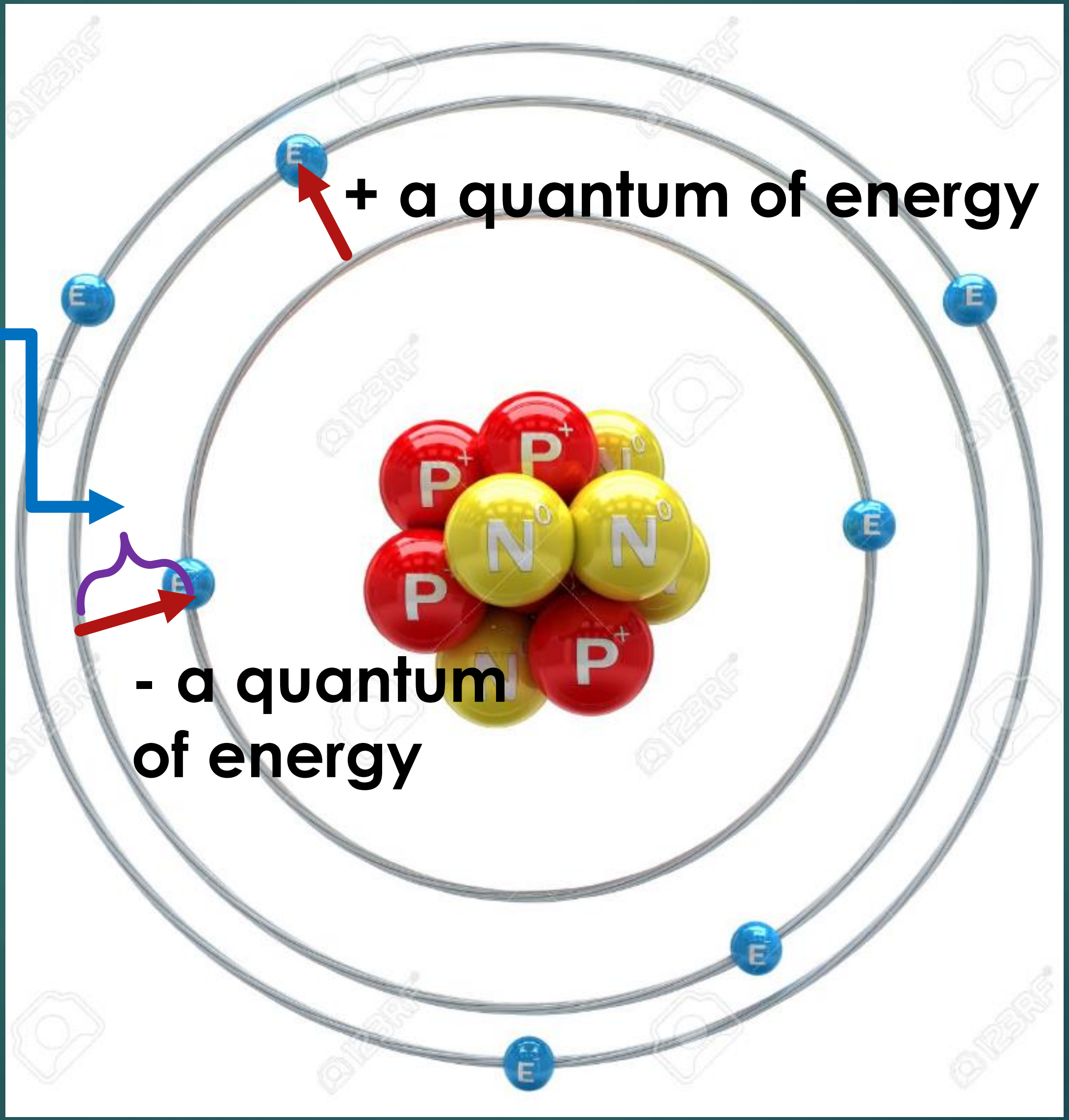
Quantized energy orbitals



Electrons can’t be here!

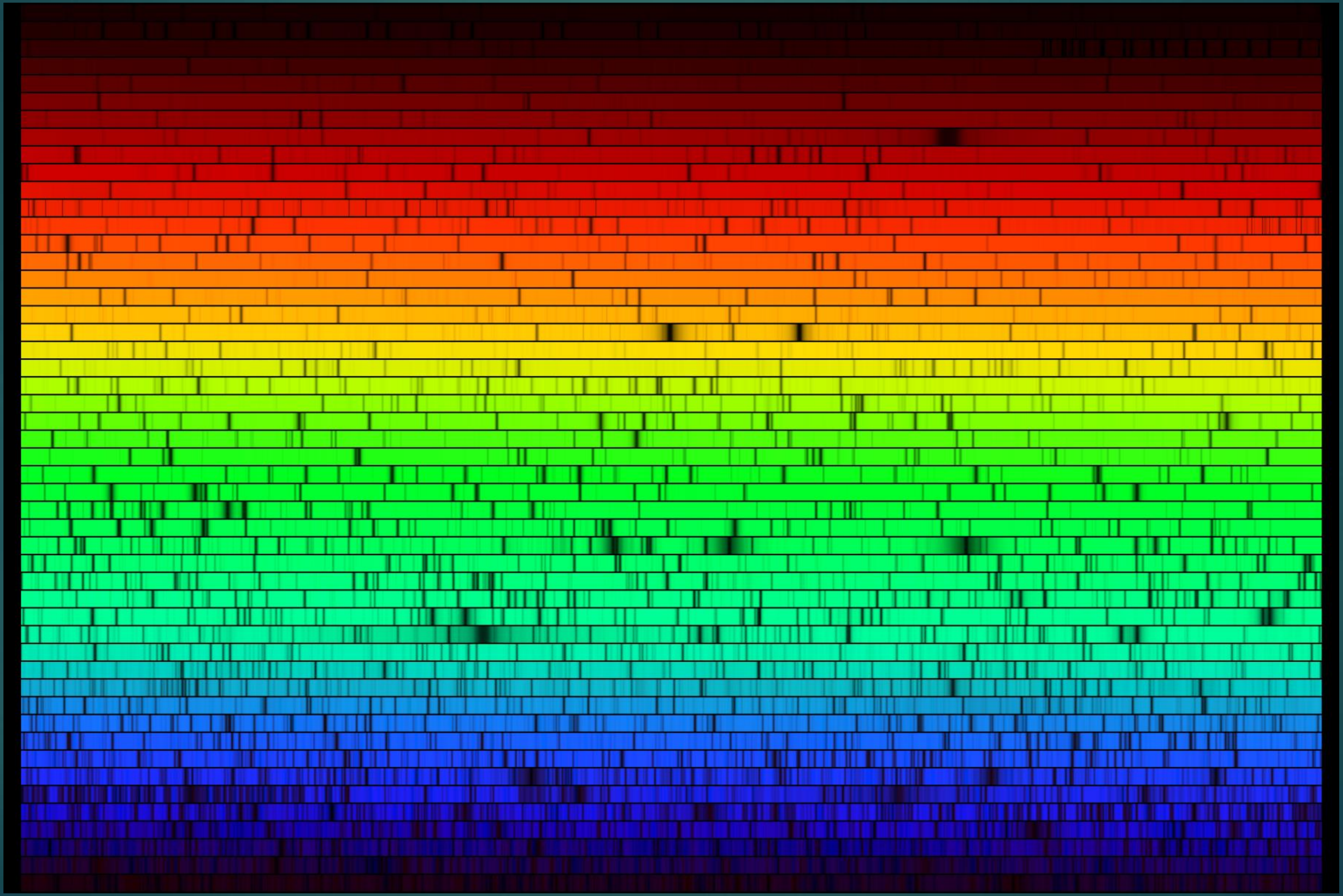
Bohr Model of the Atom

Change in energy is released as radiation, in the case of metals, it could be orange light!



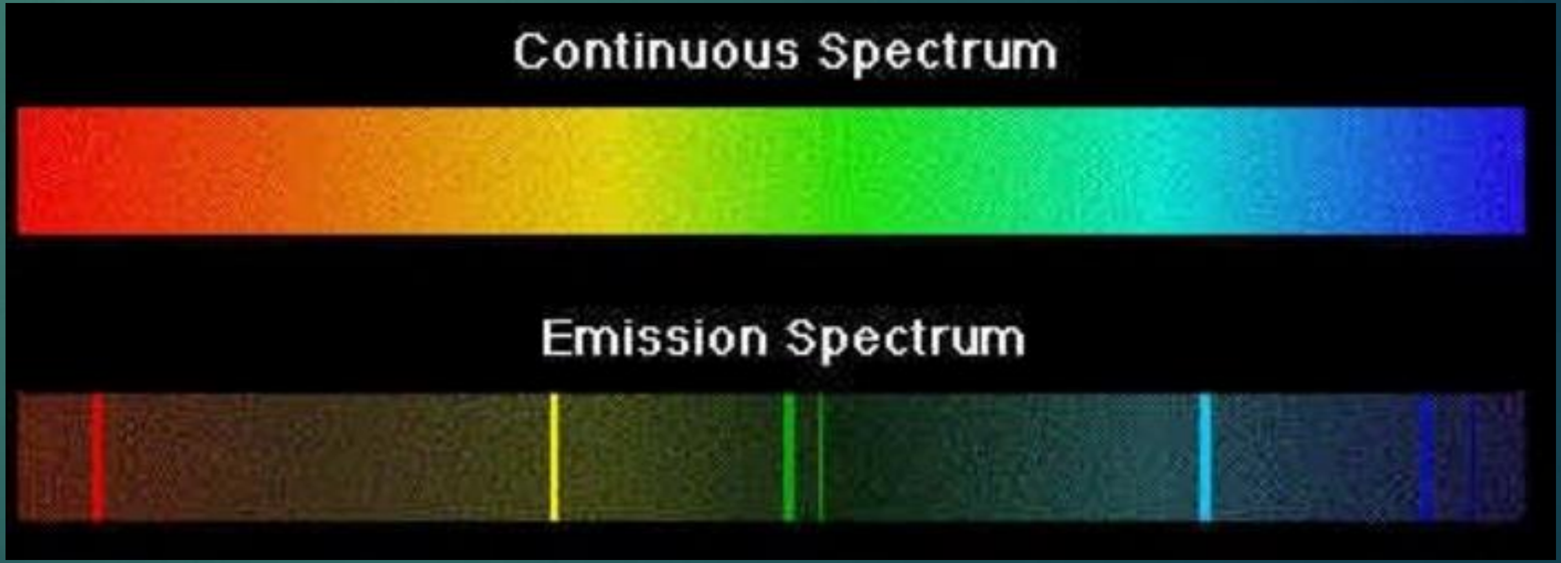
- Again, this model explained hydrogen, but failed for the larger atoms.
- This theory was refined by Erwin Schrodinger in 1926.

Absorbing Energy: Absorption Spectra

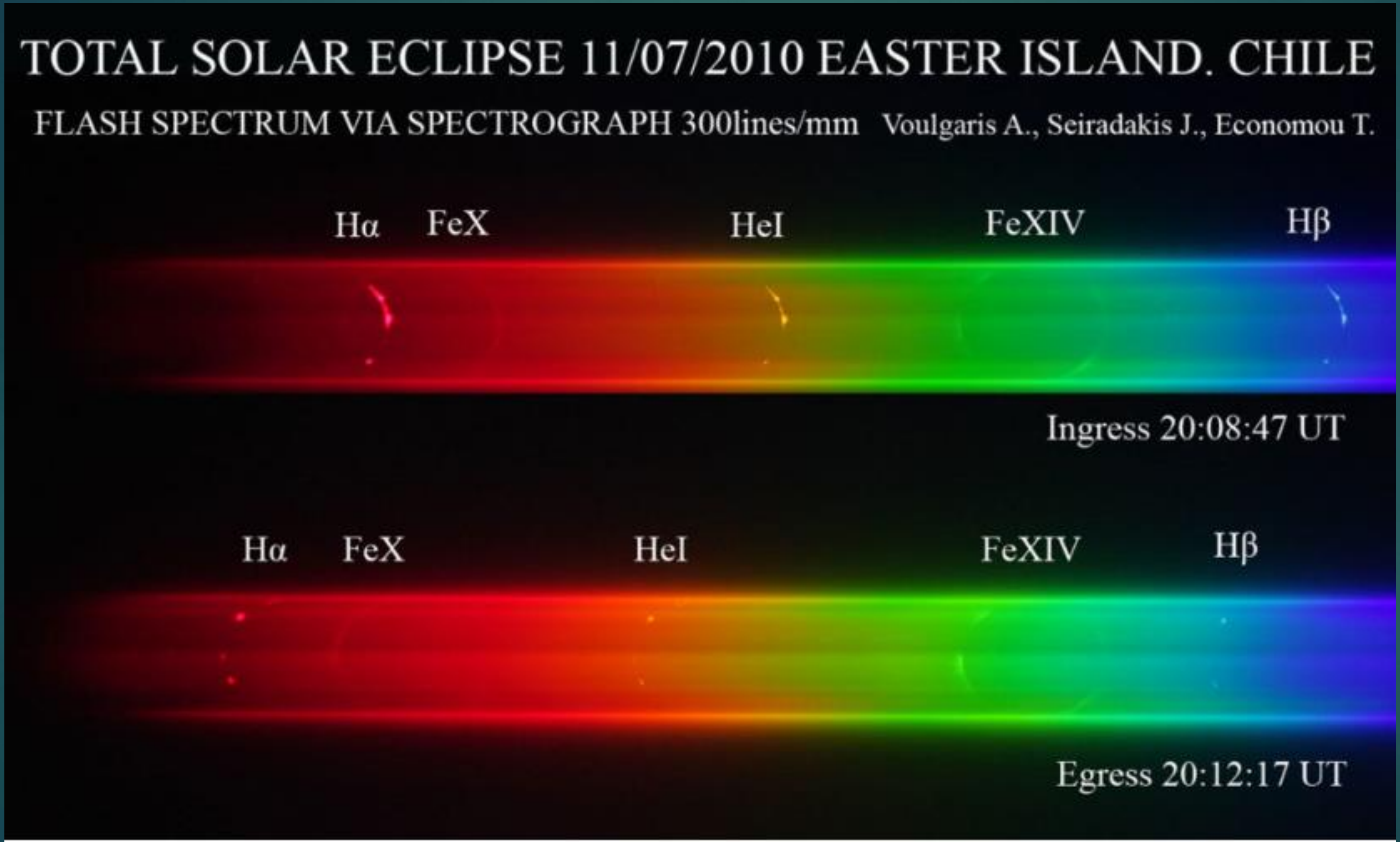


Emitting Energy: Emission Spectrum

- Helium was discovered on the Sun before it was found on Earth.
- During a solar eclipse, extra energies were found being emitted from the hot gas, called the solar corona, surrounding the Sun.



Emitting Energy: Emission Spectrum



The Flash Spectrum after the 2nd and before 3rd contact from TSE 2010, Easter Island

Quantum Mechanical Model

- In 1926 Austrian physicist Erwin Schrodinger spearheaded the development of the basis for our current model of the atom, the ***Electron Cloud Model*** or ***Quantum Mechanical Model***.
- His model of the atom was mathematical. A mathematical representation of the atom based on all know numerical and scientific information at the time.
- This resulted in a model where electrons have a high ***probability*** of being found in a specific region around the nucleus, called an ***atomic orbital***.

Schrodinger's Wave Equation

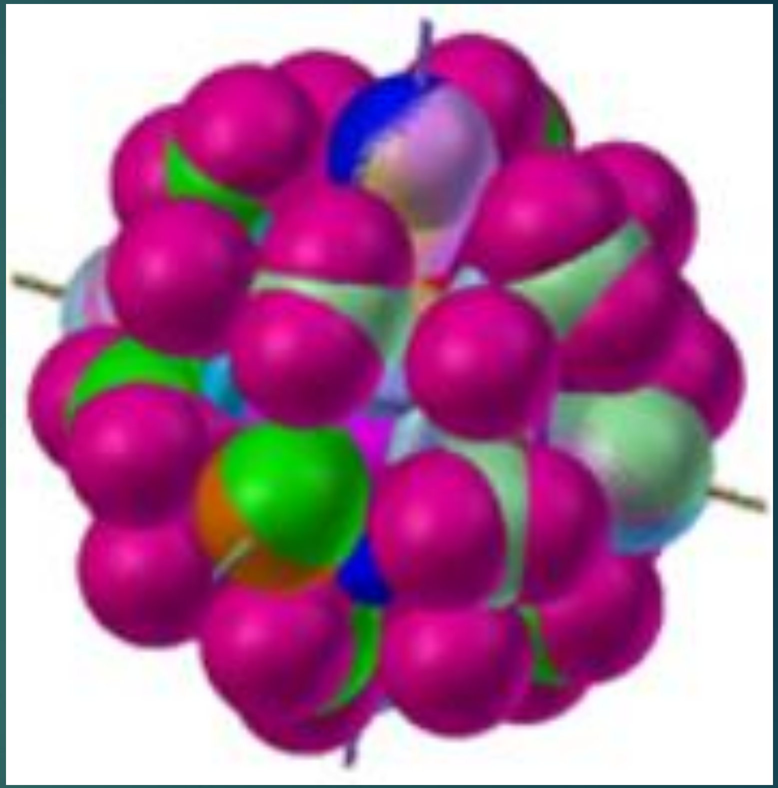
$$-\frac{h^{2}}{8\pi^{2}m}\frac{d^{2}\psi}{dx^{2}}+V\psi=E\psi$$

Erwin Schrodinger
Erwin Schrodinger

Equation for the <u>probability</u> of a single electron being found along a single axis (x-axis)

Quantum Mechanical Model

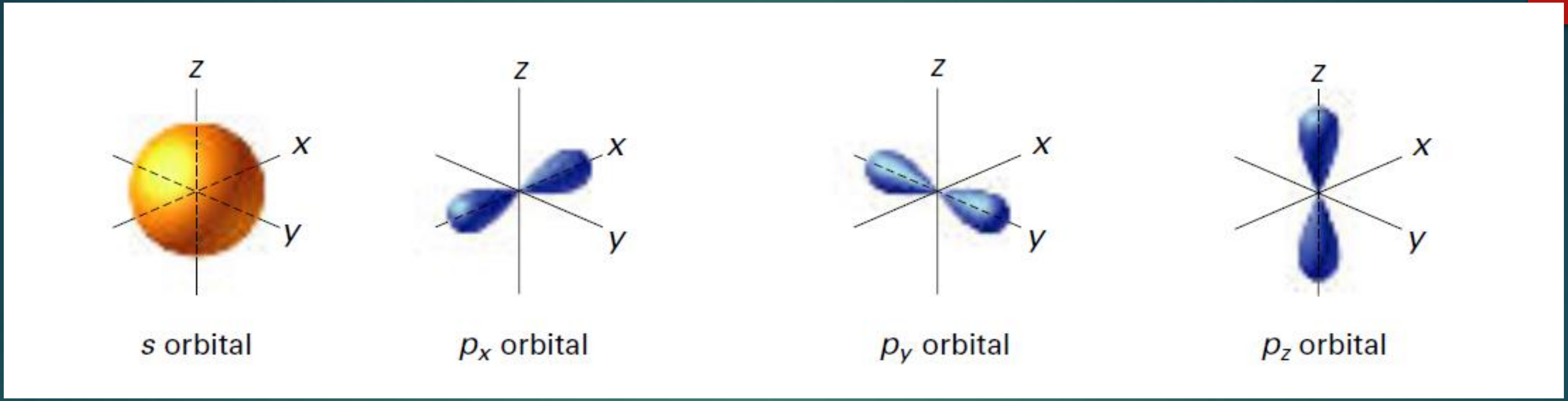
- The regions were not circular orbitals, but blob-shaped 3D spaces.
- Electrons were found to have wave and particle properties.
- Expanded upon Bohr's work and explained the properties of all elements (to a degree, they did not have any electronically powered computers for calculations).

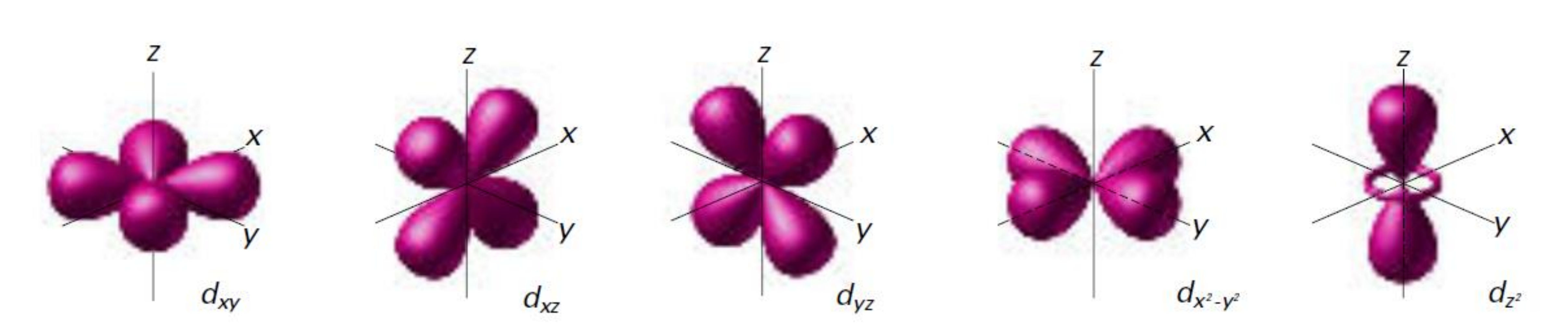


Atomic Orbitals

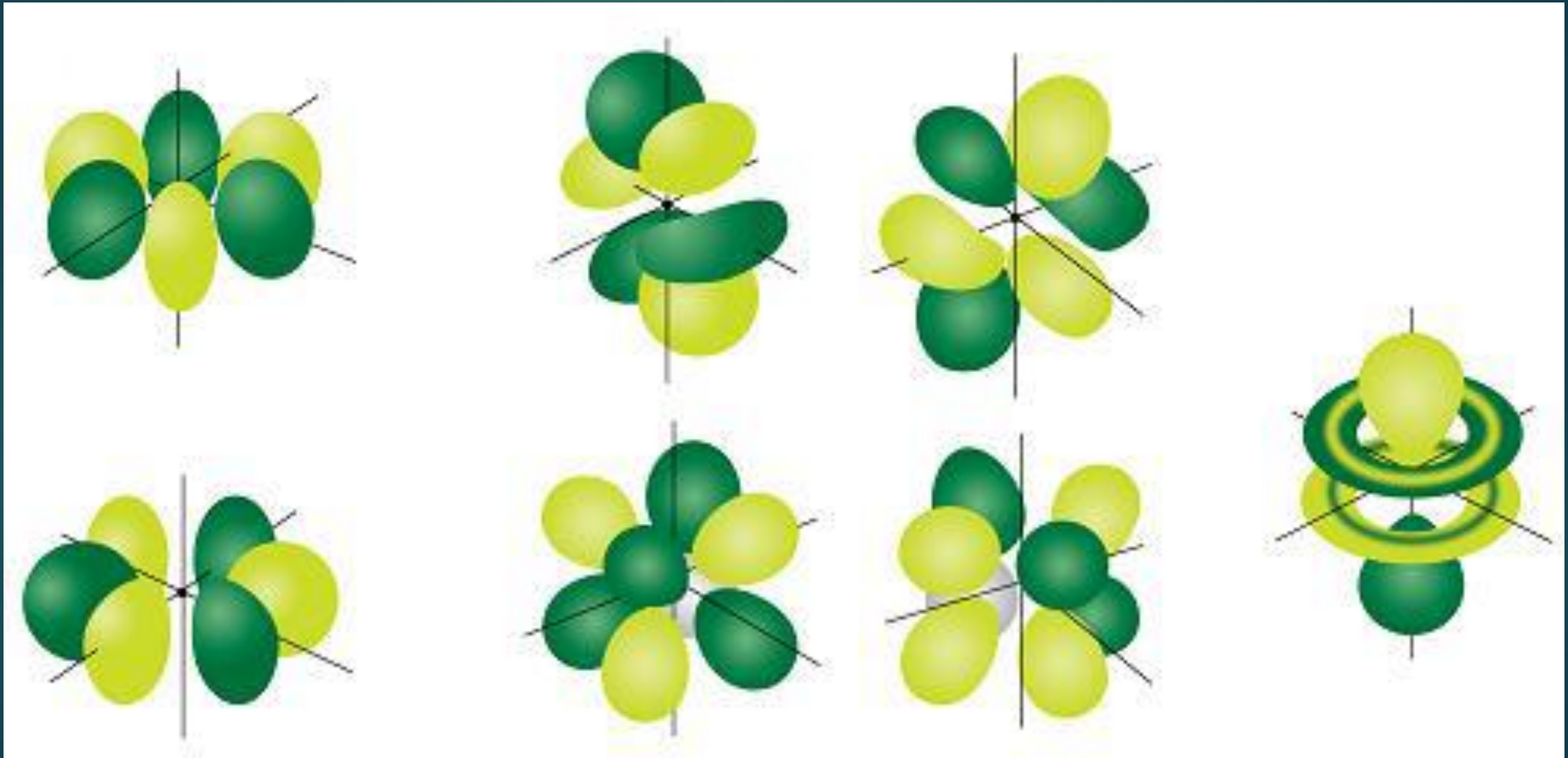
- *Principal Quantum Number* (n) = the energy level of the electron: 1, 2, 3, etc.
- Within each energy level, the complex math of Schrodinger's equation describes several shapes.
- These are called atomic orbitals - regions where there is a high probability of finding an electron.
- Sublevels- like theater seats arranged in sections: letters s, p, d, and f

Quantum Mechanical Model

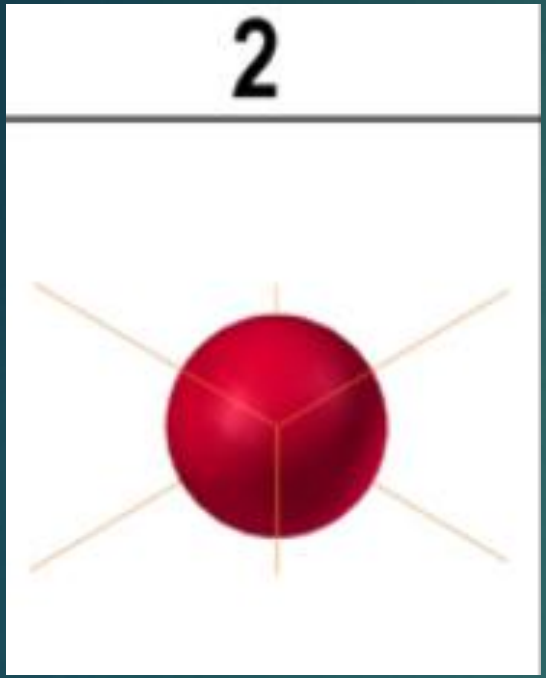


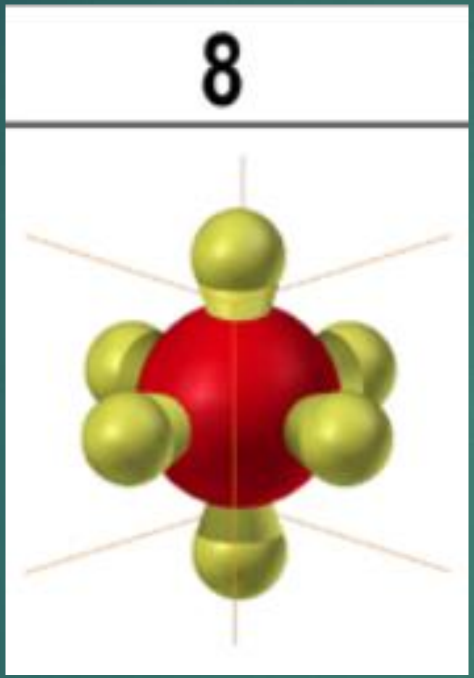


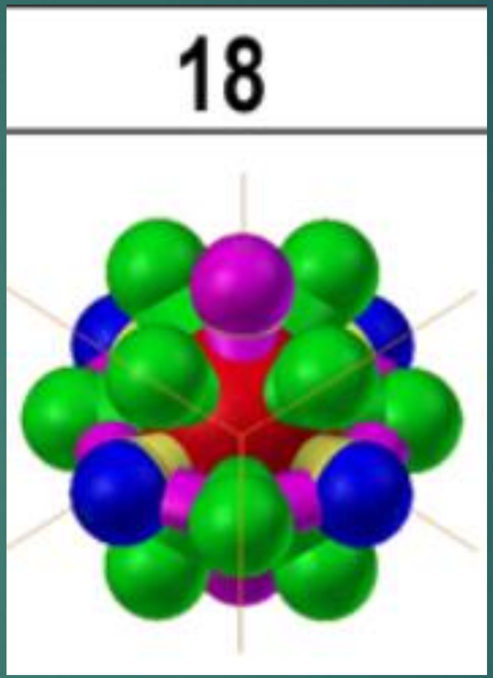
Quantum Mechanical Model

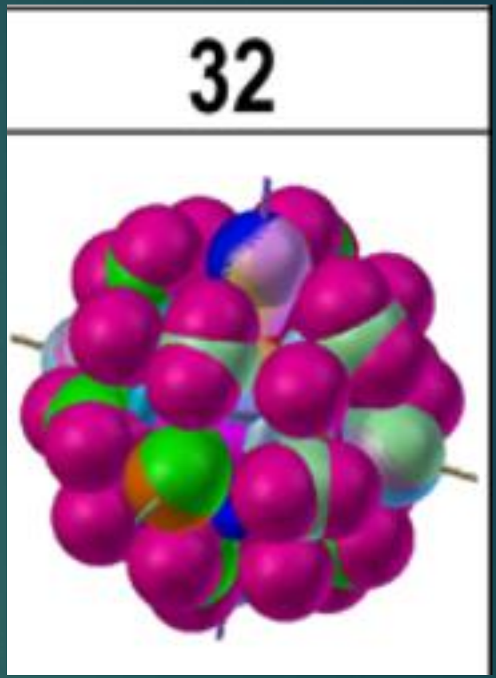


Overlapping Orbitals = Electron Cloud









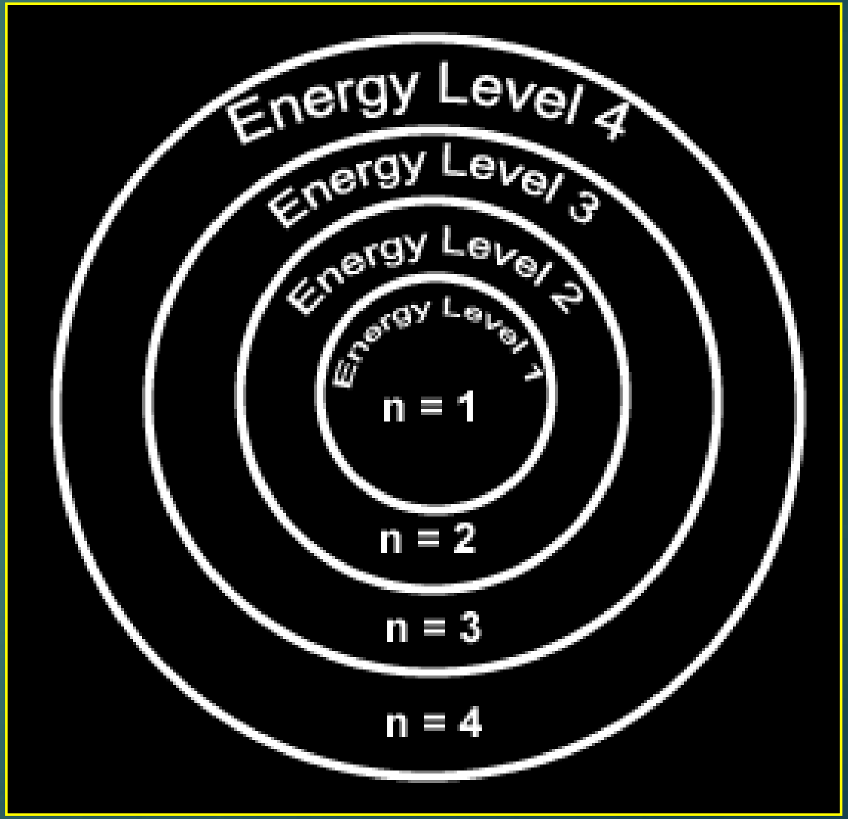
The Electron: Summary Video

The Electron: Crash Course Chemistry

CrashCourse

Principal Quantum Number

Generally symbolized by "n", it denotes the shell (energy level) in which the electron is located.



Maximum number of electrons that can fit in an energy level is:

$2n^2$

How many e^- in level 2? 3?

Summary

	# of shapes (orbitals)	Maximum electrons	Starts at energy level
s	1	**2**	1
p	3	**6**	2
d	5	**10**	3
f	7	**14**	4

By Energy Level

- First Energy Level
- Has only s orbital
- only **2** electrons
- $1s^2$

- Second Energy Level
- Has s and p orbitals available
- 2 in s, 6 in p
- $2s^22p^6$
- 8 total electrons

By Energy Level

- Third energy level
- Has s, p, and d orbitals
- 2 in s, 6 in p, and 10 in d
- $3s^2 3p^6 3d^{10}$
- 18 total electrons

- Fourth energy level
- Has s, p, d, and f orbitals
- 2 in s, 6 in p, 10 in d, and 14 in f
- $4s^2 4p^6 4d^{10} 4f^{14}$
- 32 total electrons

USMLT2: Describe model of the atom over the past 100 years and compare them to the current quantum mechanical model. Explore and summarize Rutherford's experiment.

Be able to define, explain, identify or provide examples of each of the following:

- Dalton's Model
- Thompson's Model
- Rutherford's Model
- Bohr Model
- Quantum Mechanical Model
- Quantum
- Energy Level
- Orbital
- Orbital Shape

Textbook Practice

- Page 108 #s 9, 12 – 14
- Page 122 – 124 #s 42, 43, 45, 74, 76
- Page 132 #s 1 – 7
- Page 149 #s 22 – 29

Electron Arrangement in Atoms

OBJECTIVES:

- Describe how to write the ***electron configuration*** for an atom.
- Explain why the actual electron configurations for some elements *differ* from those predicted by the Aufbau principle.

By Energy Level

- The orbitals do not fill up in a neat order.
- The energy levels overlap
- Lowest energy fill first.

Quantum Numbers

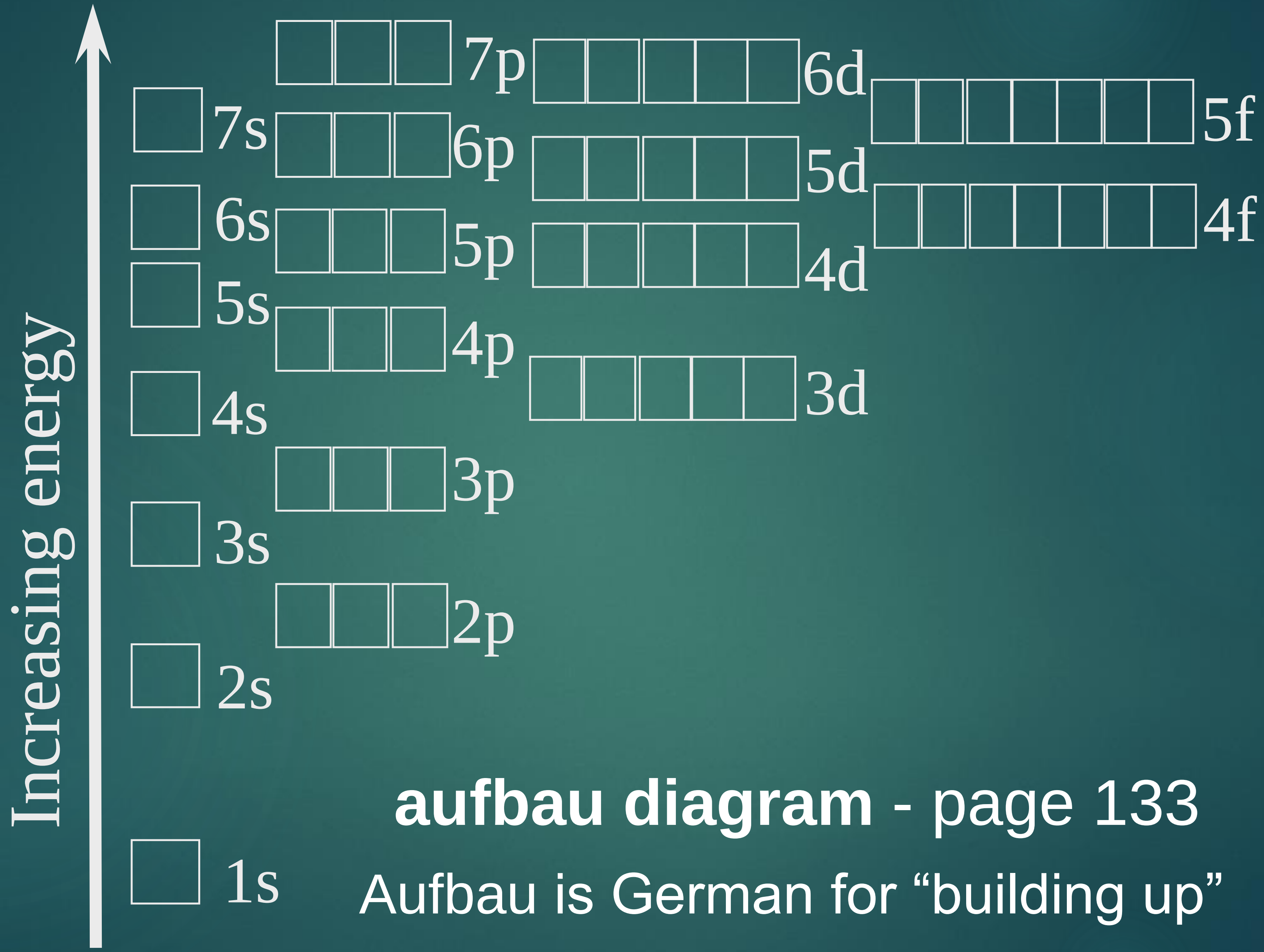
Each electron in an atom has a unique set of <u>4 quantum numbers</u> which describe it.

1) **Principal quantum number**
2) **Angular momentum quantum number**
3) **Magnetic quantum number**
4) **Spin quantum number**

Electron Configurations…

- …are the way electrons are arranged in various orbitals around the nuclei of atoms. *Three rules tell us how:*

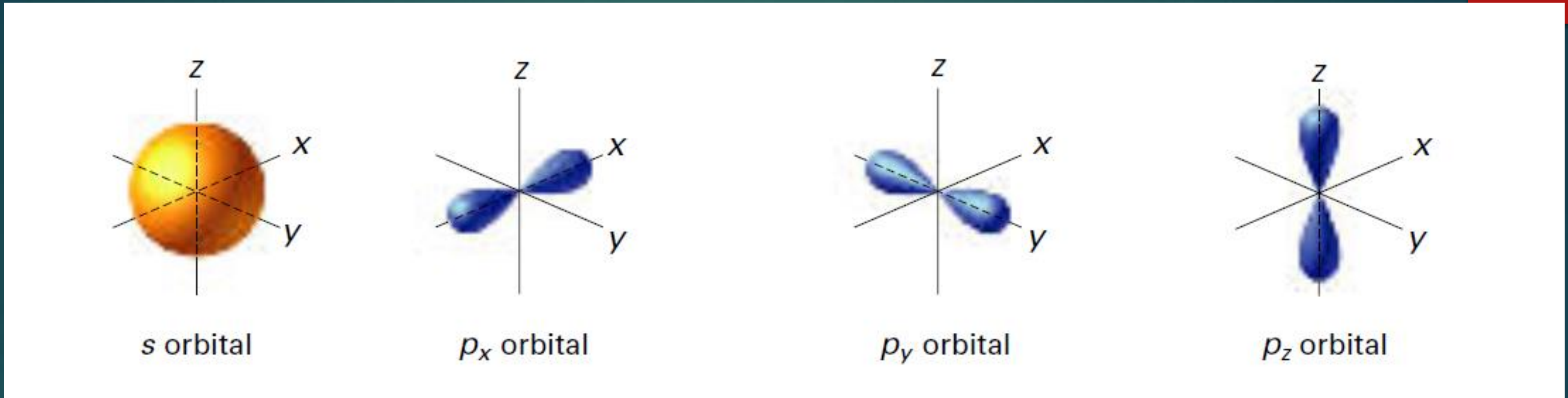
1) Aufbau principle - electrons enter the lowest energy first.
   - This causes difficulties because of the overlap of orbitals of different energies – follow the diagram!

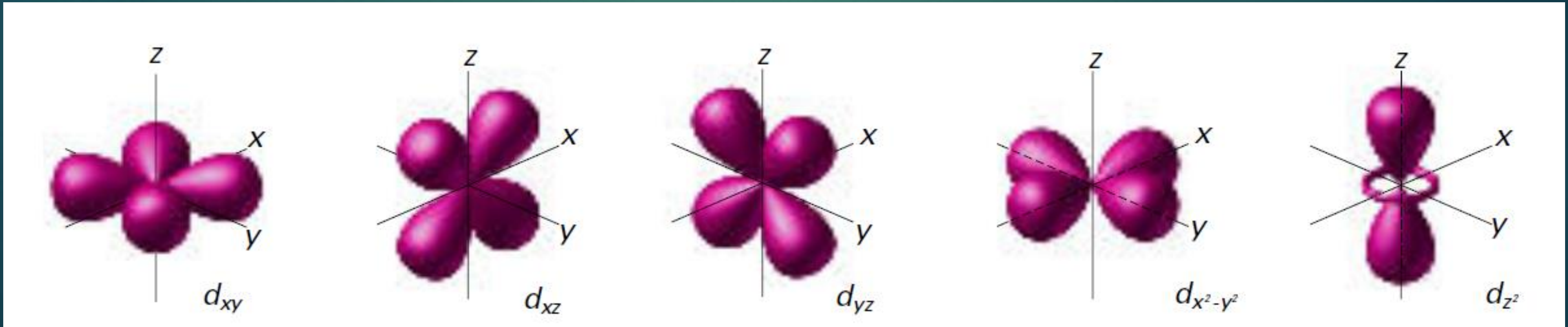


aufbau diagram - page 133

Aufbau is German for “building up”

Orbital shapes





↙ = filling order

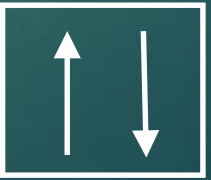
n=1	$1s^2$				$2e^-$
n=2	$2s^2$	$2p^6$			$8e^-$
n=3	$3s^2$	$3p^6$	$3d^{10}$		$18e^-$
n=4	$4s^2$	$4p^6$	$4d^{10}$	$4f^{14}$	$32e^-$
n=5	$5s^2$	$5p^6$	$5d^{10}$	$5f^{14}$	...
n=6	$6s^2$	$6p^6$	$6d^{10}$	...	...
n=7	$7s^2$	$7p^6$	...	...	...

Rule 2: Pauli Exclusion Principle

Wolfgang Pauli

No two electrons in an atom can have the same four quantum numbers.

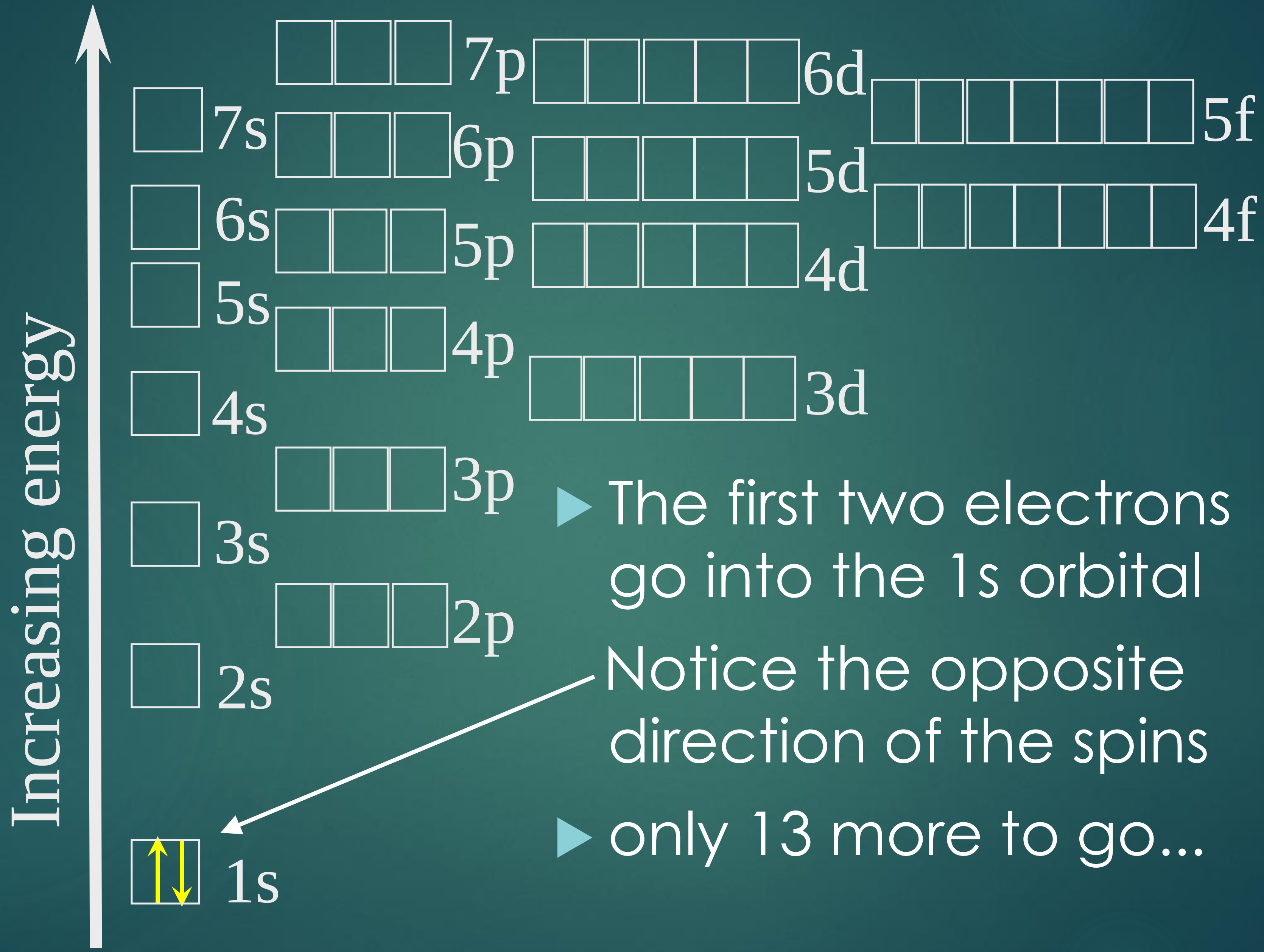
To show the different direction of spin, a pair in the same orbital is written as:



Electron Configurations

3) <u>Hund's Rule-</u> When electrons occupy orbitals of equal energy, they don't pair up until they have to.

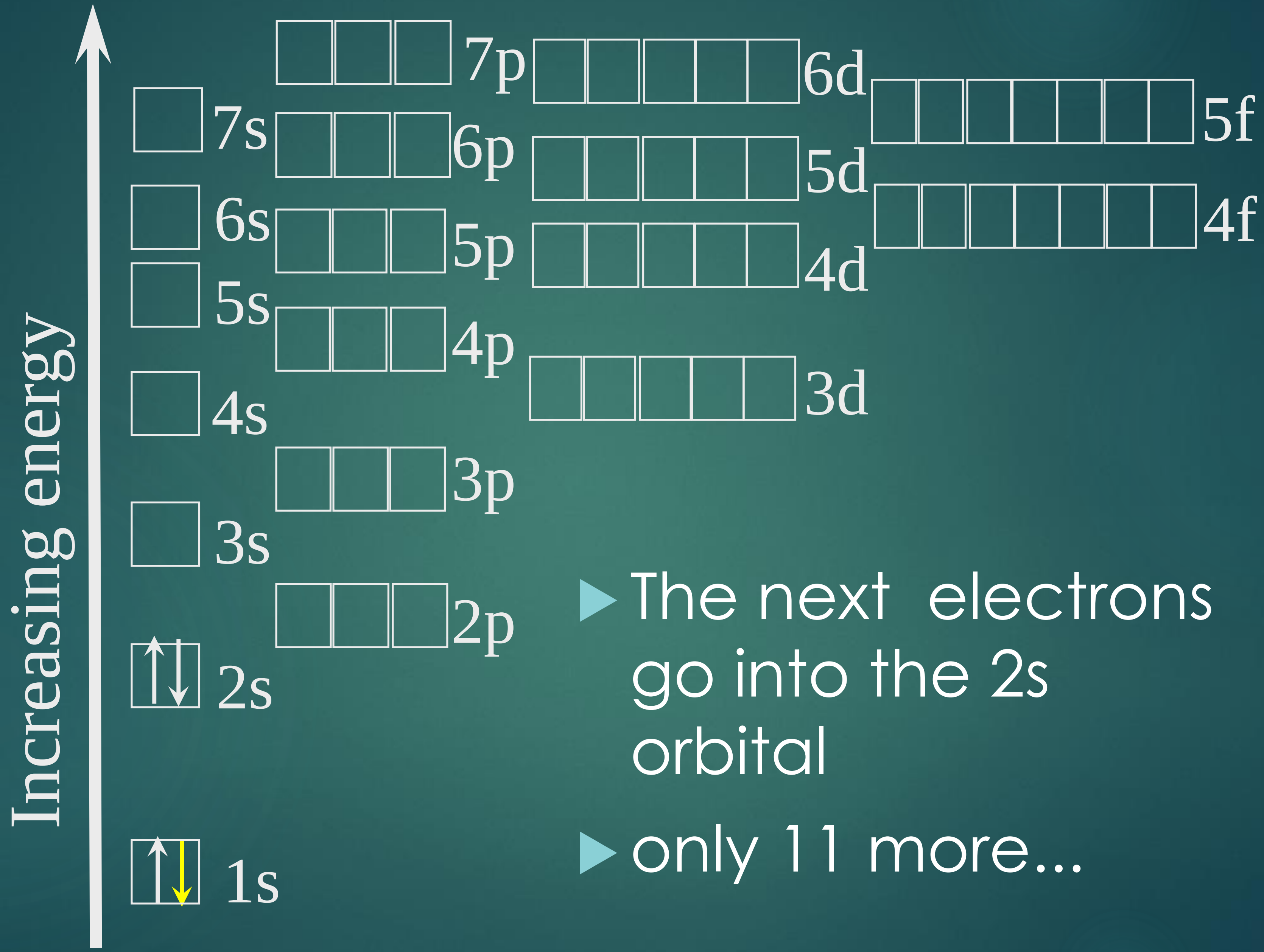
- Let's write the electron configuration for Phosphorus
  - We need to account for all 15 electrons in phosphorus



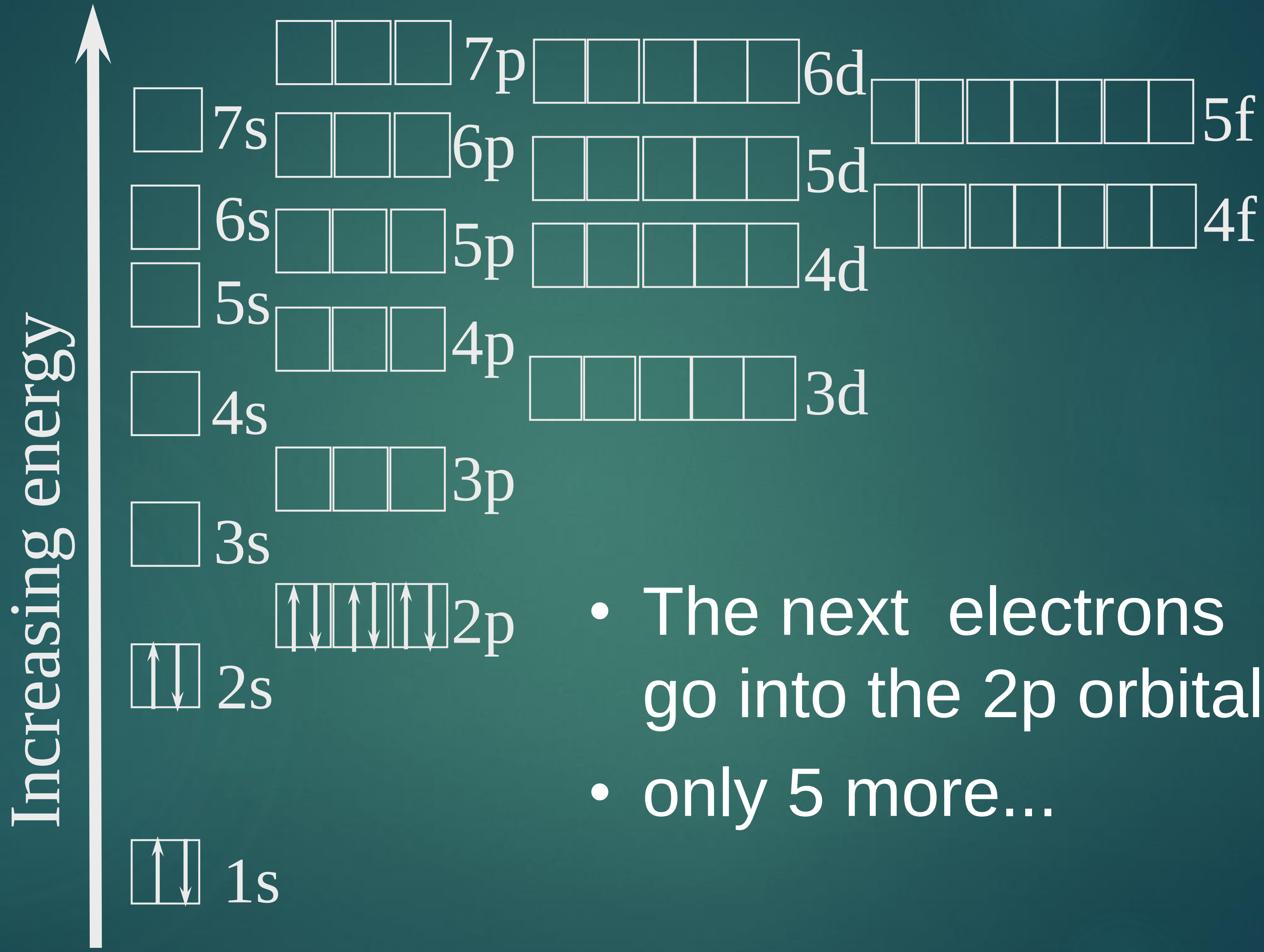
- The first two electrons go into the 1s orbital

Notice the opposite direction of the spins

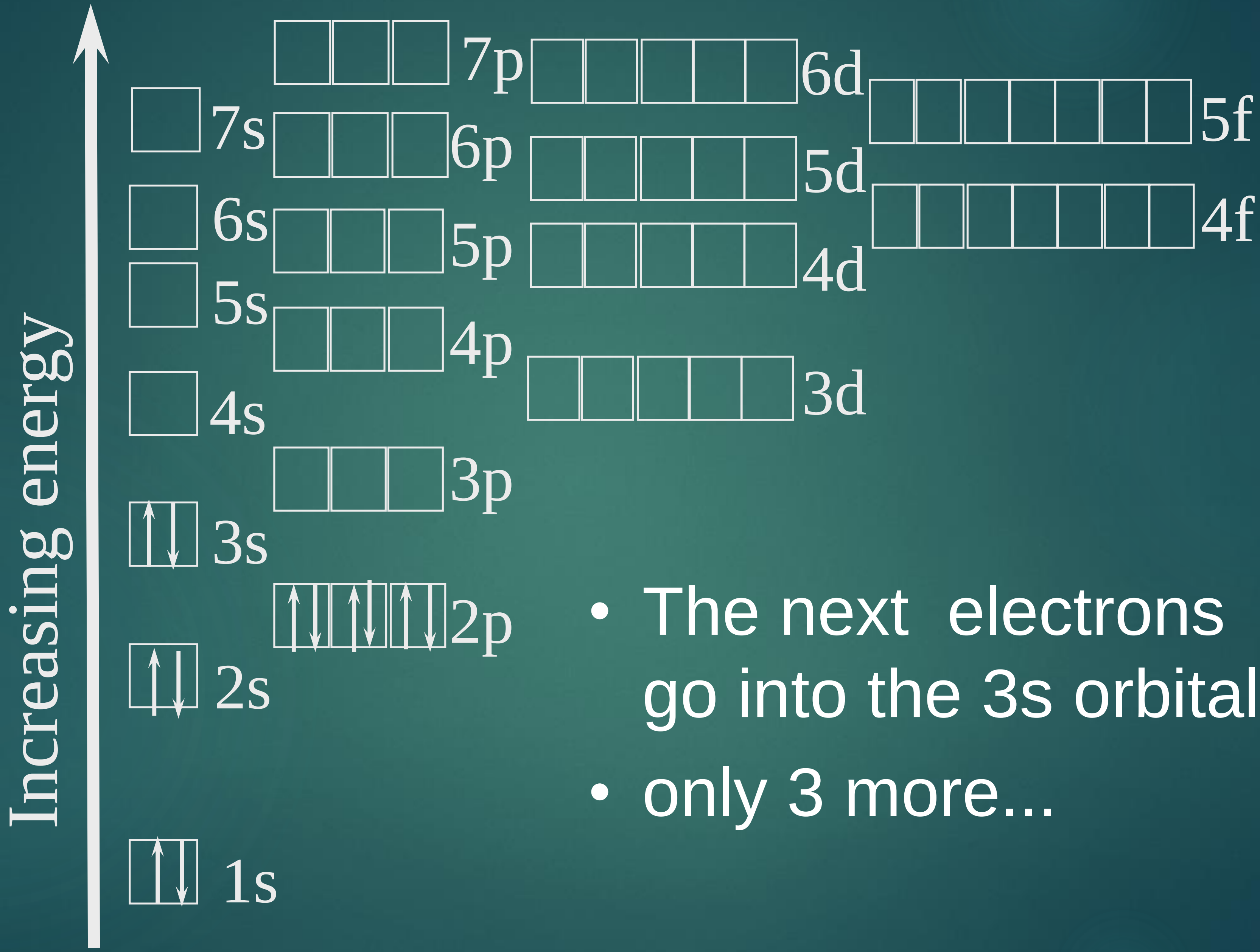
- only 13 more to go...



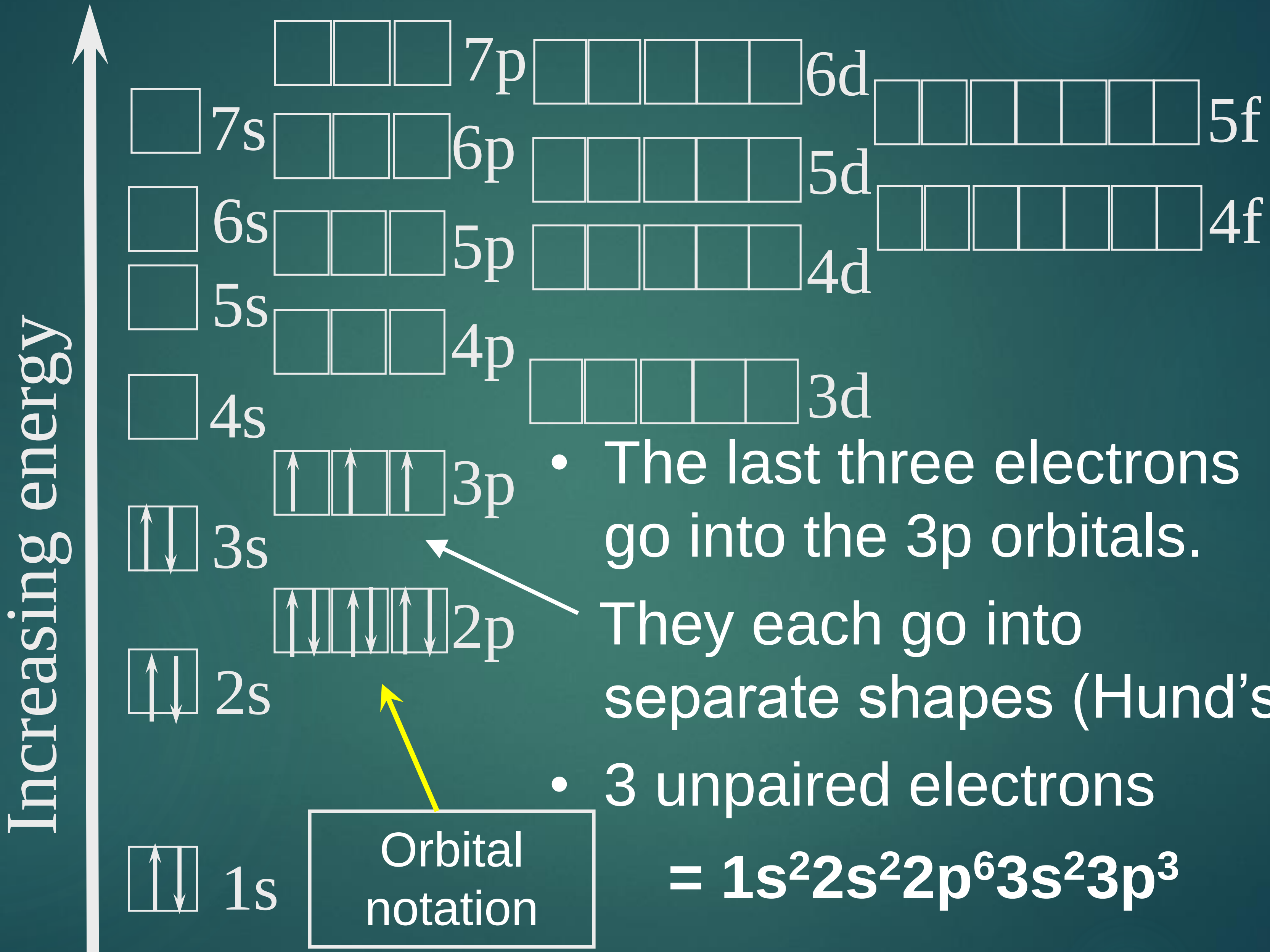
- The next electrons go into the 2s orbital
- only 11 more...



- The next electrons go into the 2p orbital
- only 5 more...



- The next electrons go into the 3s orbital
- only 3 more...



- The last three electrons go into the 3p orbitals.

  They each go into separate shapes (Hund's)

- 3 unpaired electrons

  $= 1s^2 2s^2 2p^6 3s^2 3p^3$

Practice questions

- Refer to Learning Target Guide

Orbitals fill in an order

- Lowest energy to higher energy.
- Adding electrons can change the energy of the orbital. Full orbitals are the absolute best situation.
- However, **half filled** orbitals have a lower energy, and are next best
  - Makes them more stable.
  - Changes the filling order

Write the electron configurations for these elements:

- Titanium - 22 electrons
  - $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$
- Vanadium - 23 electrons
  - $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$
- Chromium - 24 electrons
  - $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^4$ **(expected)**
    - But this is not what happens!!

Chromium is actually:

- $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$
- Why?
- This gives us two half filled orbitals (the others are all still full)
- Half full is slightly lower in energy.
- The same principal applies to copper.

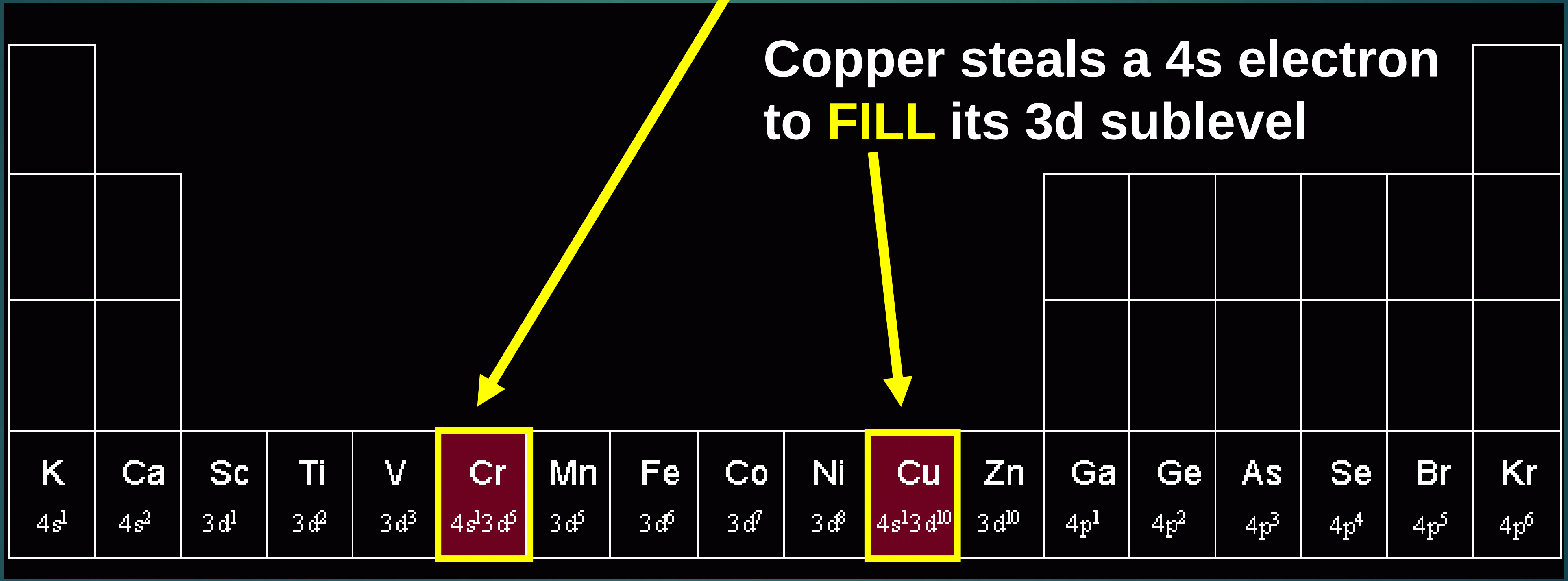
Copper's electron configuration

- Copper has 29 electrons so we expect: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^9$
- But the *actual configuration* is:
- $1s^2 2s^2 2p^6 3s^2 3p^6 \underline{4s^1 3d^{10}}$
- This change gives one more filled orbital and one that is half filled.
- Remember these exceptions: **d^4, d^9**

Irregular configurations of Chromium and Copper

Chromium steals a 4s electron to make its 3d sublevel HALF FULL

Copper steals a 4s electron to FILL its 3d sublevel



Electron configuration in groups

- Noble gases
  - Elements in group 8A.
  - The highest energy levels are completely filled with electrons.
  - That leads to them being relatively inert.
- Representative Elements
  - Groups 1A – 7A.
  - Group number is the number of electrons in the highest energy level.

1 H Hydrogen 1.00794 1							2 He Helium 4.002602 2
3 Li Lithium 6.941 2 1	4 Be Beryllium 9.012182 2 2	5 B Boron 10.811 2 3	6 C Carbon 12.0107 2 4	7 N Nitrogen 14.0067 2 5	8 O Oxygen 15.9994 2 6	9 F Fluorine 18.9984032 2 7	10 Ne Neon 20.1797 2 8
11 Na Sodium 22.98976928 2 8 1	12 Mg Magnesium 24.3050 2 8 2	13 Al Aluminium 26.9815386 2 8 3	14 Si Silicon 28.0855 2 8 4	15 P Phosphorus 30.973762 2 8 5	16 S Sulfur 32.065 2 8 6	17 Cl Chlorine 35.453 2 8 7	18 Ar Argon 39.948 2 8 8
19 K Potassium 39.0983 2 8 8 1	20 Ca Calcium 40.078 2 8 8 2	31 Ga Gallium 69.723 2 8 18 3	32 Ge Germanium 72.64 2 8 18 4	33 As Arsenic 74.92160 2 8 18 5	34 Se Selenium 78.96 2 8 18 6	35 Br Bromine 79.904 2 8 18 7	36 Kr Krypton 83.798 2 8 18 8
37 Rb Rubidium 85.4678 2 8 18 8 1	38 Sr Strontium 87.62 2 8 18 8 2	49 In Indium 114.818 2 8 18 18 3	50 Sn Tin 118.710 2 8 18 18 4	51 Sb Antimony 121.760 2 8 18 18 5	52 Te Tellurium 127.60 2 8 18 18 6	53 I Iodine 126.90447 2 8 18 18 7	54 Xe Xenon 131.293 2 8 18 18 8
55 Cs Caesium 132.9054519 2 8 18 18 8 1	56 Ba Barium 137.327 2 8 18 18 8 2	81 Tl Thallium 204.3833 2 8 18 32 18 3	82 Pb Lead 207.2 2 8 18 32 18 4	83 Bi Bismuth 208.98040 2 8 18 32 18 5	84 Po Polonium (208.9824) 2 8 18 32 18 6	85 At Astatine (209.9871) 2 8 18 32 18 7	86 Rn Radon (222.0176) 2 8 18 32 18 8
87 Fr Francium (223) 2 8 18 32 18 8 1	88 Ra Radium (226) 2 8 18 32 18 8 2	113 Uut Ununtrium (284) 2 8 18 32 32 18 3	114 Uuq Ununquadium (289) 2 8 18 32 32 18 4	115 Uup Ununpentium (288) 2 8 18 32 32 18 5	116 Uuh Ununhexium (292) 2 8 18 32 32 18 6	117 Uus Ununseptium	118 Uuo Ununoctium (294) 2 8 18 32 32 18 8

USMLT3: Explain and write electron configuration diagrams using Hund's rule, Pauli exclusion principle and the Aufbau principle.

Be able to define, explain, identify or provide examples of each of the following:

- Quantum Numbers
- Electron Configuration
- Noble Gases
- Representative Elements
- Hund's Rule
- Pauli Exclusion Principle
- Aufbau Principle
- Exceptions to Aufbau Principle

Textbook Practice

- Page 135 #s 8, 9
- Page 136 #s 10 – 13
- Page 149 #s 30 – 34, 36, 37, 39