UNIT 1

UNDERLYING STRUCTURE OF MATTER

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

Subscript → Atomic number

X

Symbols

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

$$^{80}_{35}Br$$

NAMING ISOTOPES

- 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.
  - Carbon-12 chosen because of its isotope purity.
- 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

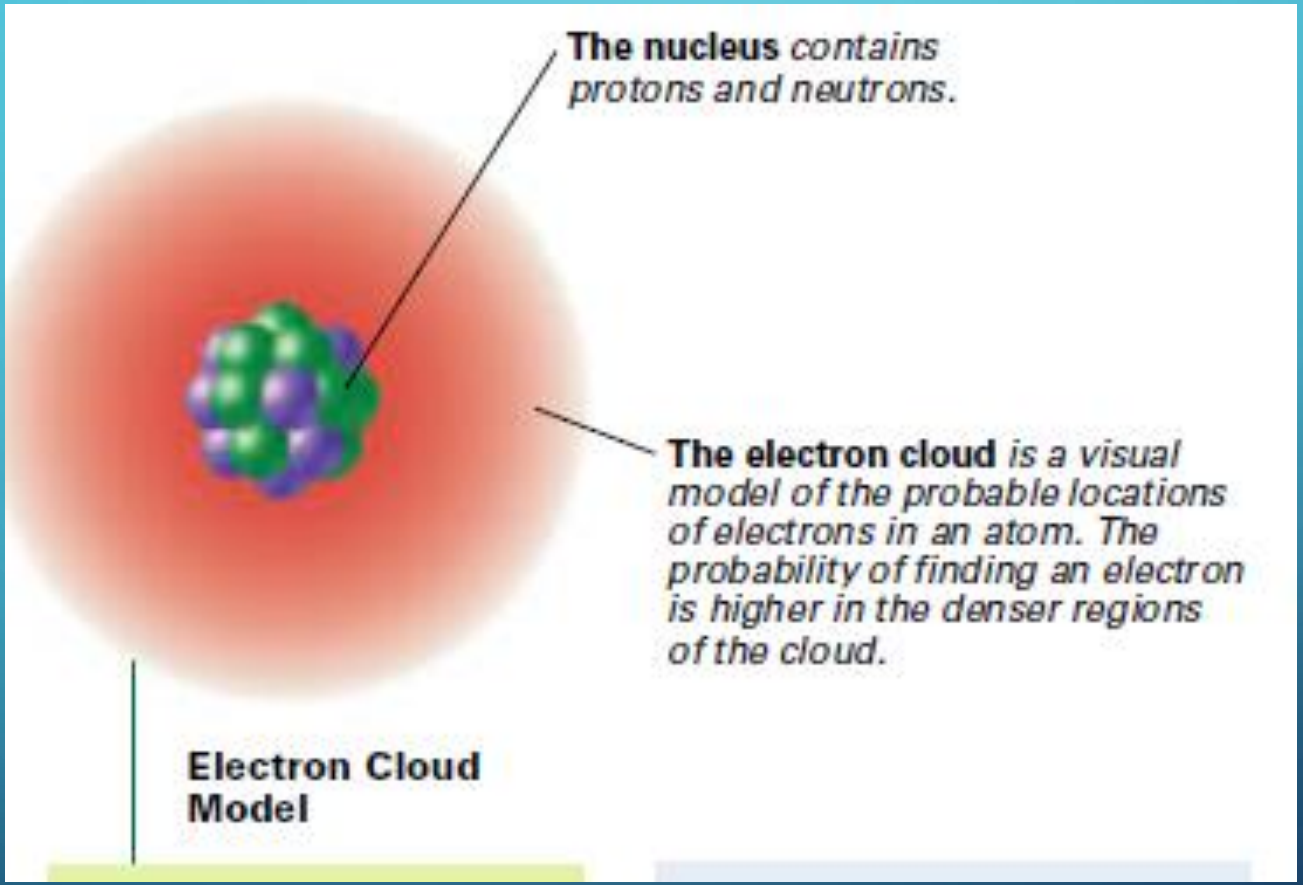
- Using the information below, calculate the approximate atomic mass of silicon.

$^{28}_{14}Si$	$^{29}_{14}Si$	$^{30}_{14}Si$	$^{31}_{14}Si$
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.

ELECTRONS IN ATOMS – CHAPTER 5

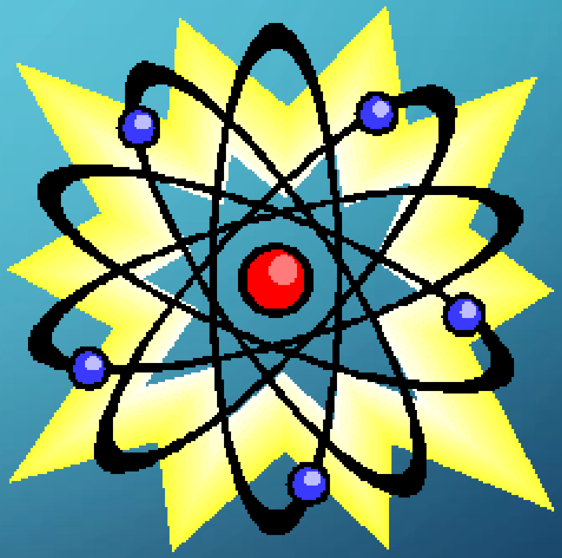


SECTION 5.1
MODELS OF THE ATOM

- OBJECTIVES:
  - Identify the inadequacies in the Rutherford atomic model.
  - Identify the new proposal in the Bohr model of the atom.
  - Describe the energies and positions of electrons according to the quantum mechanical model.
  - Describe how the shapes of orbitals related to different sublevels differ.

ERNEST RUTHERFORD'S MODEL

- Discovered dense positive piece at the center of the atom- "nucleus"
- Electrons would surround and move around it, like planets around the sun
- Atom is mostly empty space
- It did not explain the *chemical properties* of the elements – a better description of the electron behavior was needed



NIELS BOHR'S MODEL

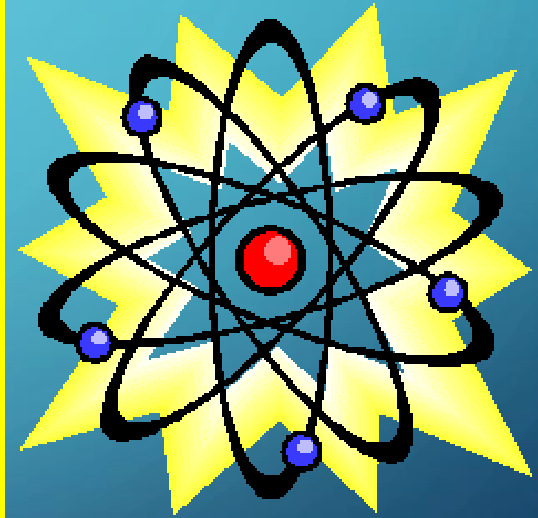
- Why don't the electrons fall into the nucleus?
- Move like planets around the sun.
  - In specific circular paths, or orbits, at different levels.
  - An amount of fixed energy separates one level from another.

THE BOHR MODEL OF THE ATOM

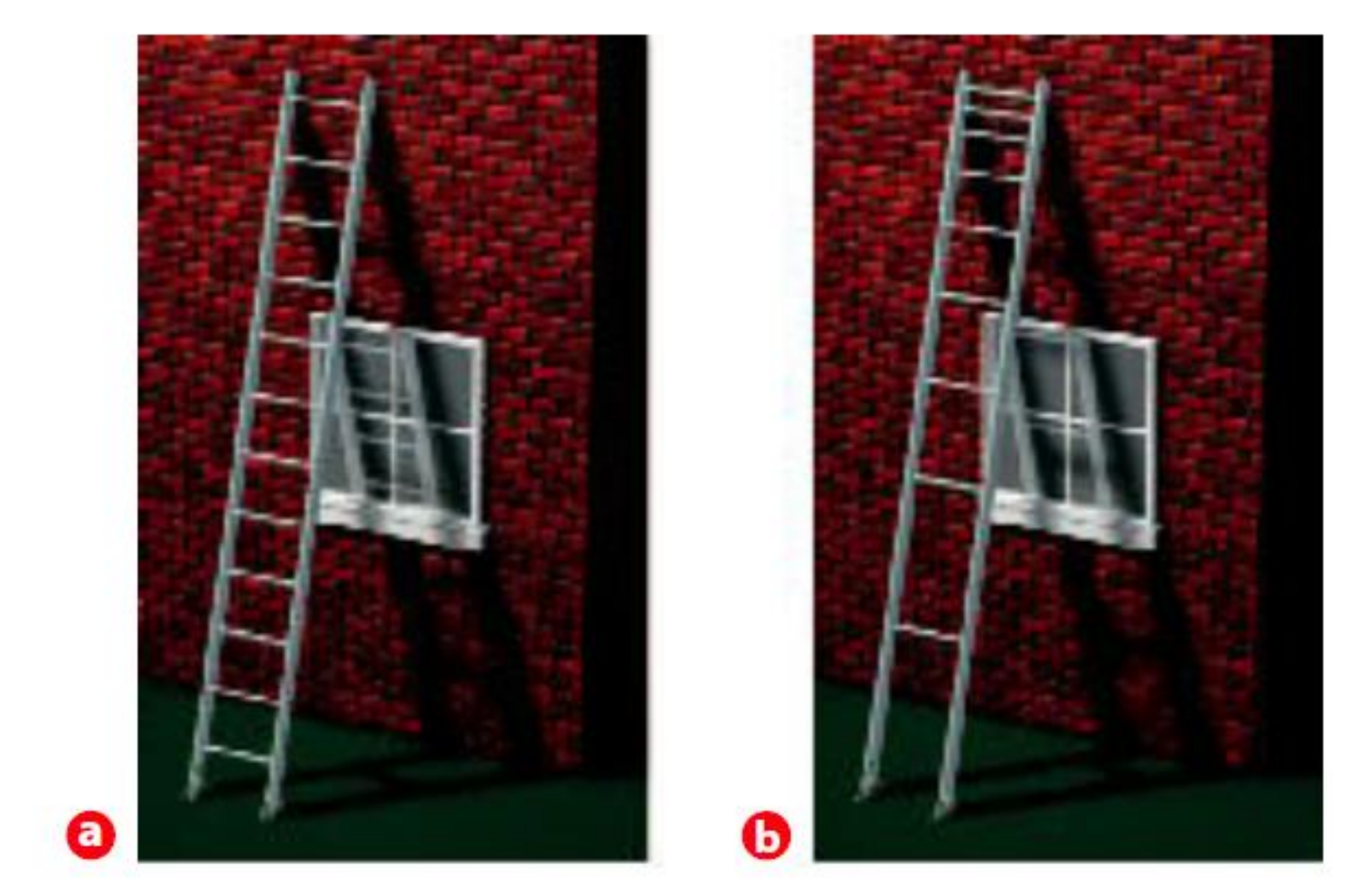
Niels Bohr

I pictured the electrons orbiting 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*.



QUANTUM ENERGY LEVELS



BOHR'S MODEL

- **Energy level** of an electron
  - analogous to the rungs of a ladder
- The electron cannot exist between energy levels, just like you can't stand between rungs on a ladder
- A **quantum** of energy is the amount of energy required to move an electron from one energy level to another

THE QUANTUM MECHANICAL MODEL

- Energy is "quantized" - It comes in chunks.
- A **quantum** is the amount of energy needed to move from one energy level to another.
- Since the energy of an atom is never "in between" there must be a quantum leap in energy.
- In 1926, **Erwin Schrodinger** derived an equation that described the energy and position of the electrons in an atom

SCHRODINGER'S WAVE EQUATION

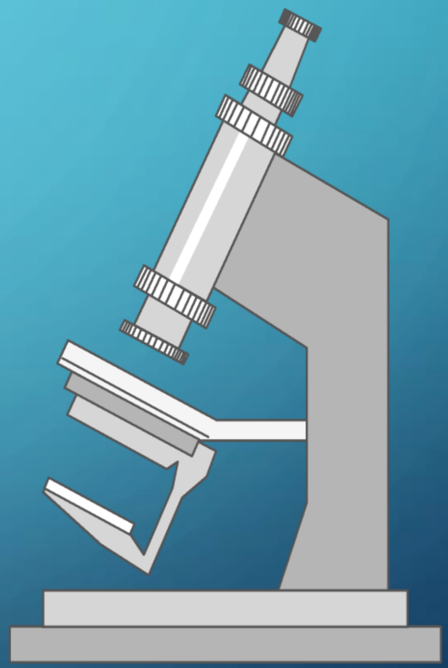
Erwin Schrodinger
Erwin Schrodinger

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

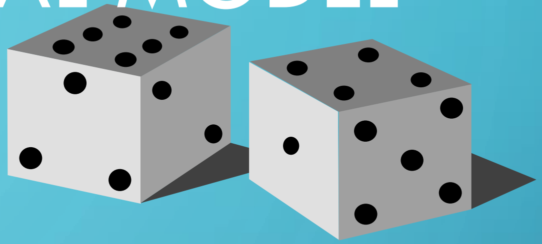
Equation for the <u>probability</u> of a single electron being found along a single axis (x-axis)

THE QUANTUM MECHANICAL MODEL

- Things that are very small *behave differently* from things big enough to see.
- The *quantum mechanical model* is a mathematical solution
- It is not like anything you can see (like plum pudding!)



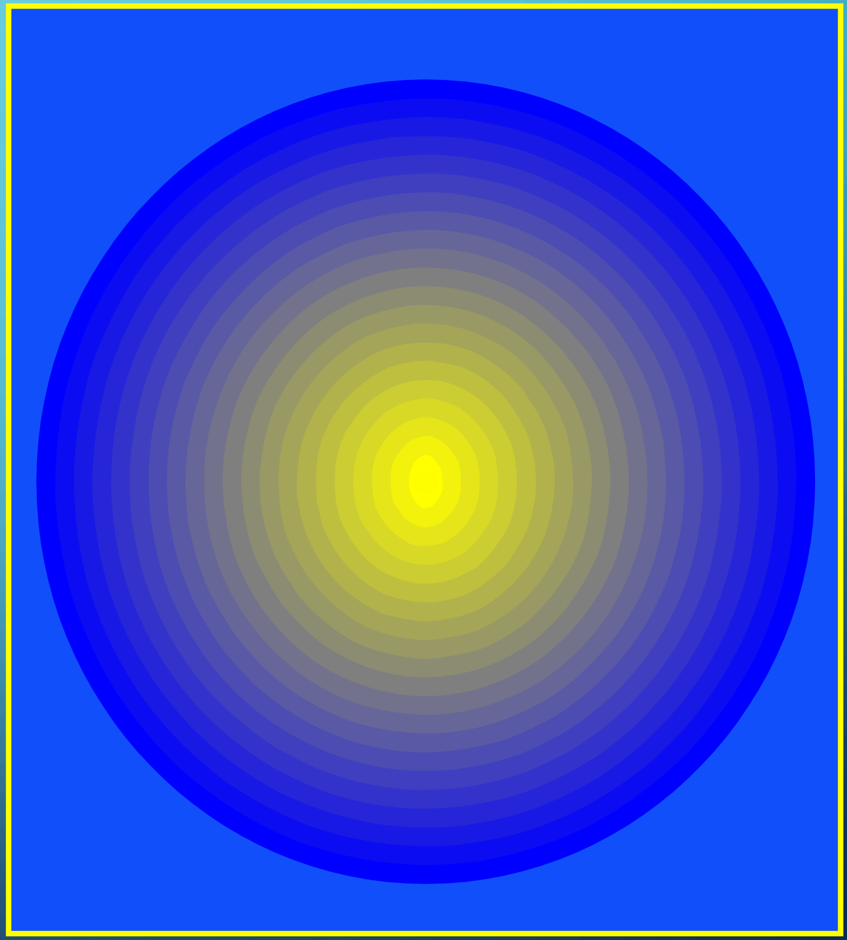
THE QUANTUM MECHANICAL MODEL



- Has energy levels for electrons.
- Orbits are not circular.
- It can only tell us the **probability** of finding an electron a certain distance from the nucleus.

THE QUANTUM MECHANICAL MODEL

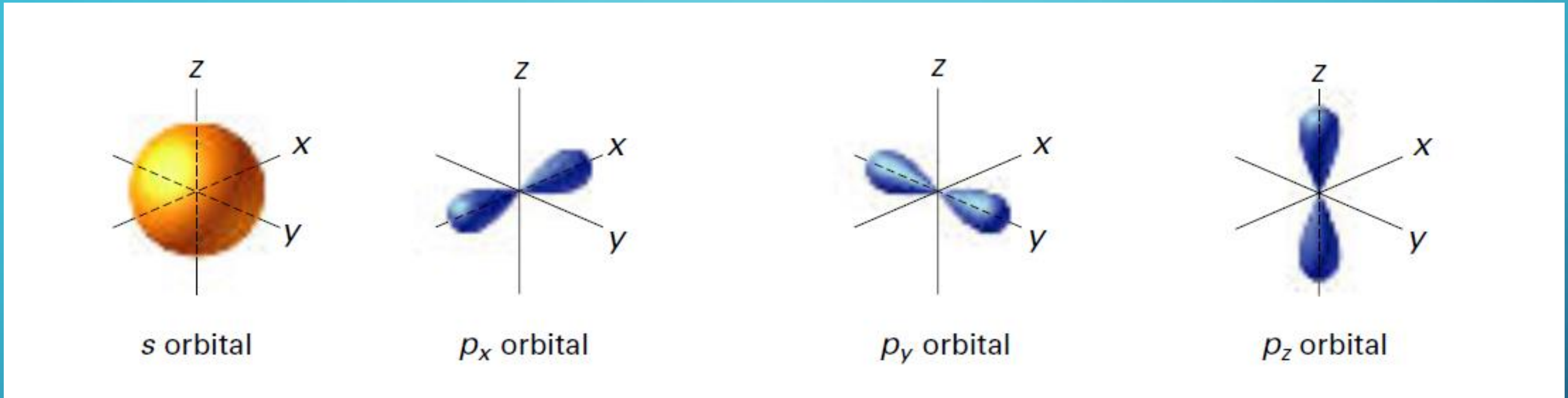
- The atom is found inside a blurry “electron cloud”
- An area where there is a *chance* of finding an electron.
- Think of fan blades

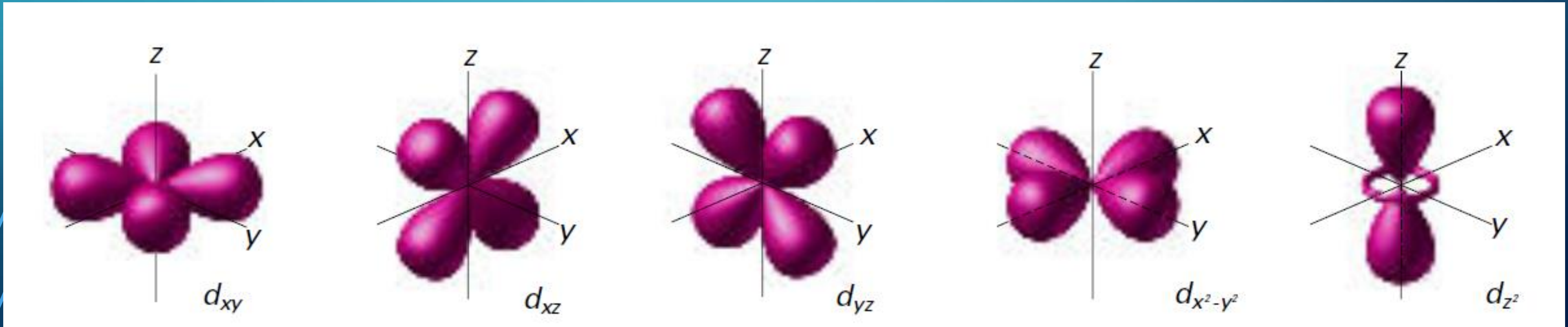


ATOMIC ORBITALS

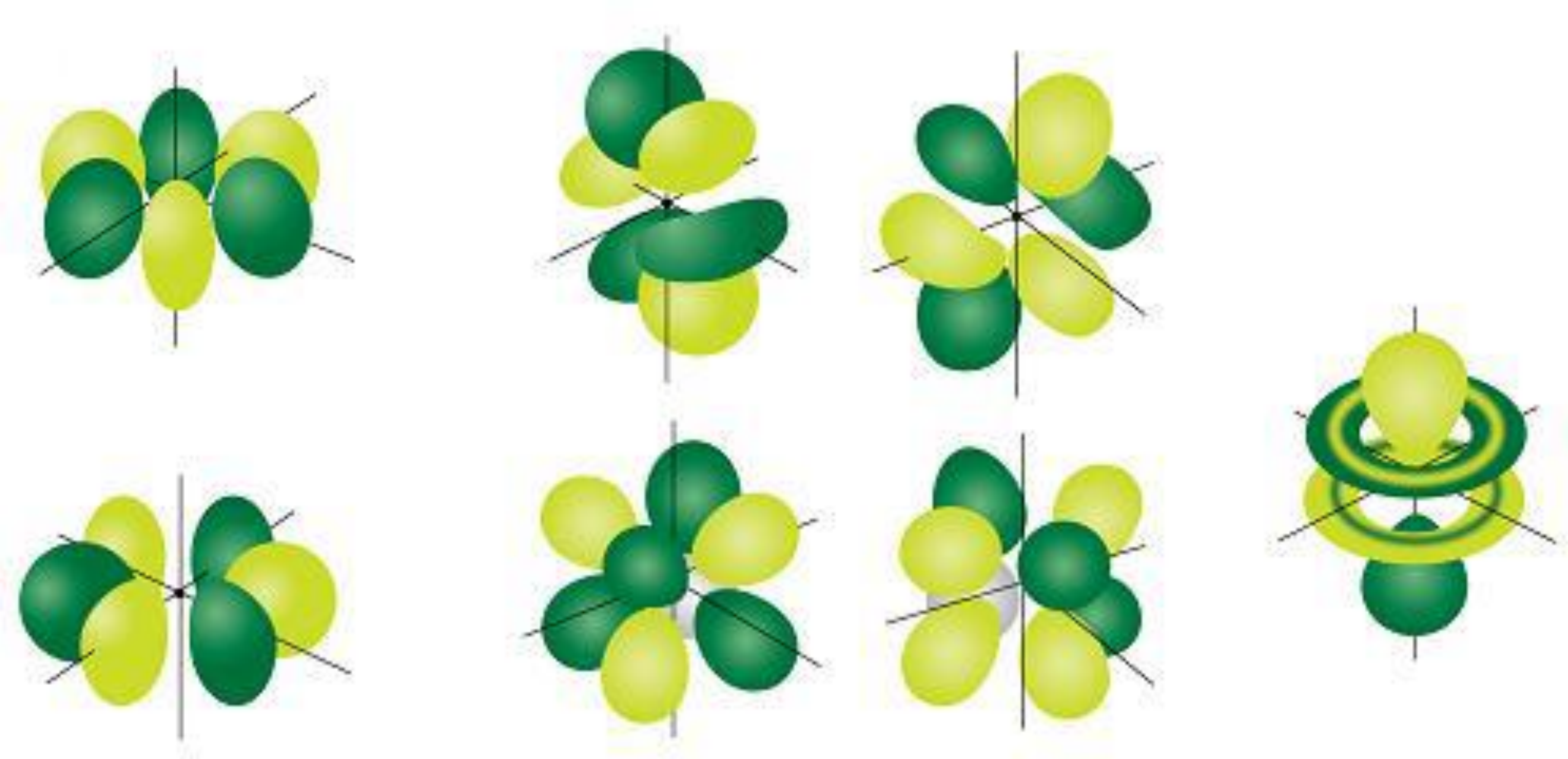
- Principal Quantum Number (n) = the energy level of the electron: 1, 2, 3, etc.
- Within each energy level, the complex math of Schrodinger's equation describes several shapes.
- These are called atomic orbitals - regions where there is a high probability of finding an electron.
- Sublevels- like theater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

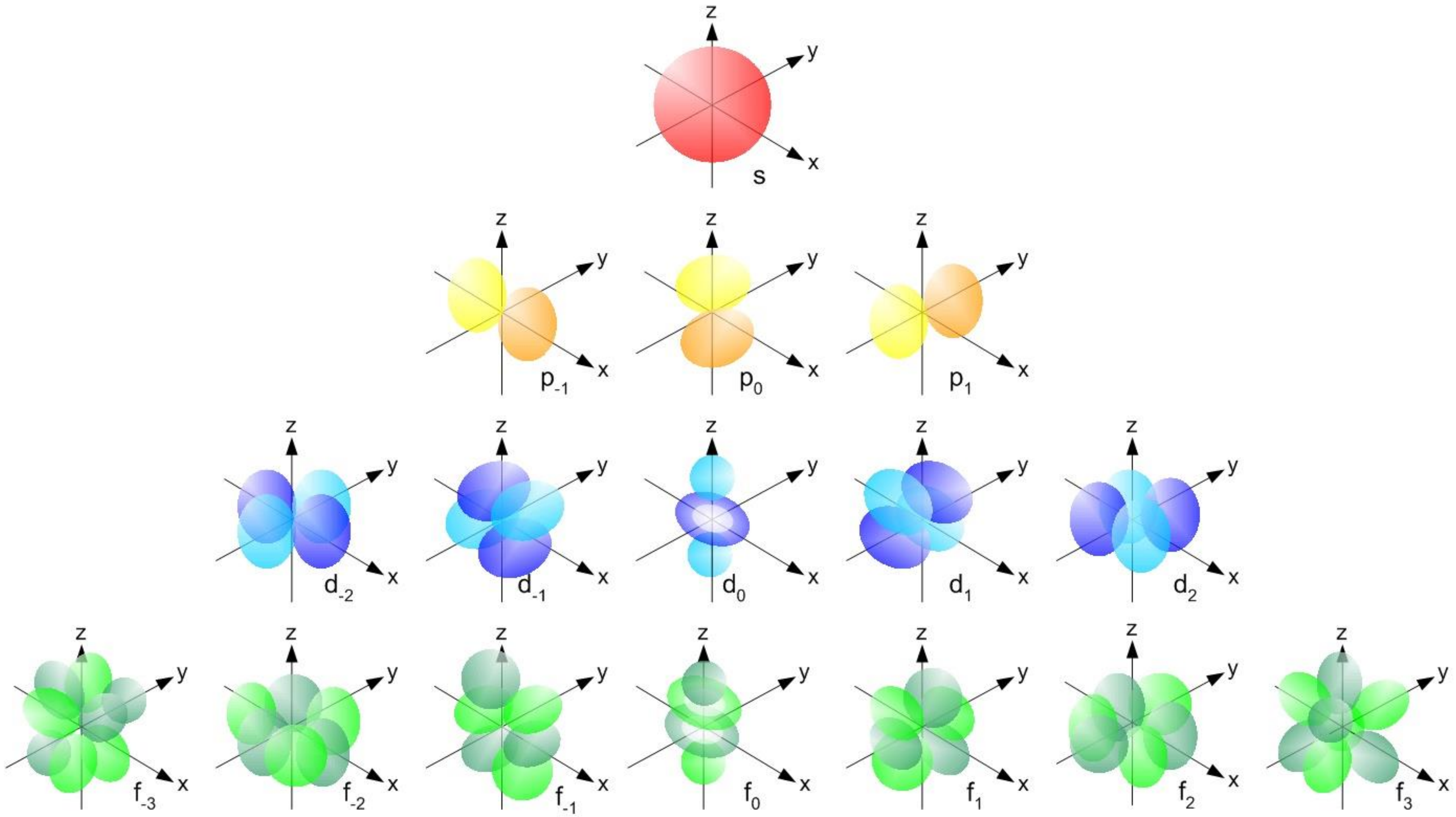
ORBITAL SHAPES





F- ORBITALS

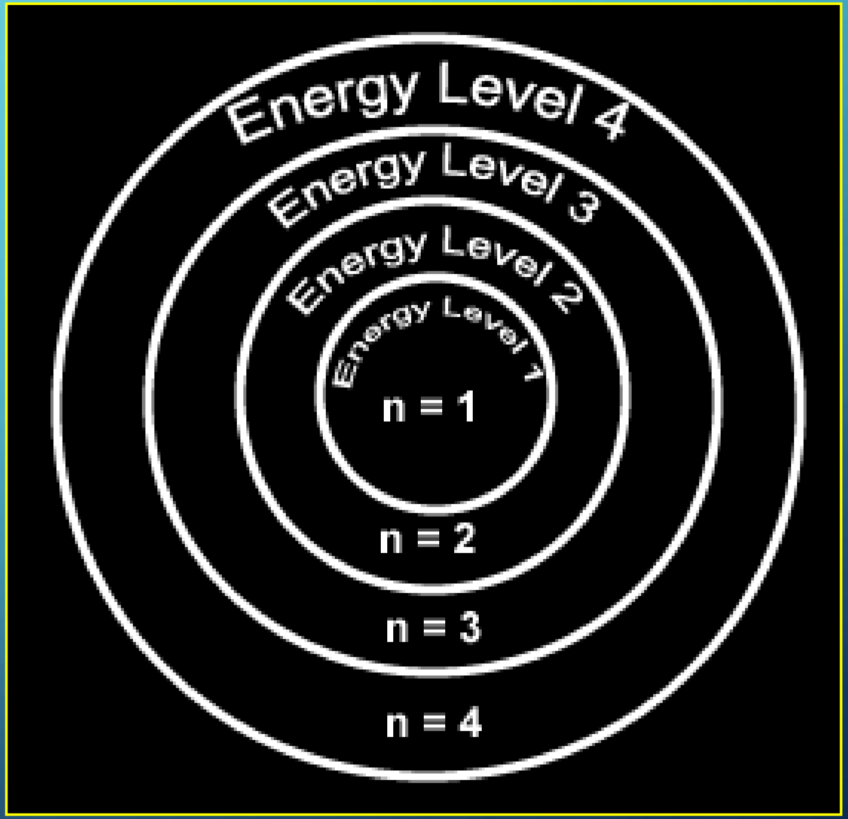




z
y
x
s
p-1
p0
p1
d-2
d-1
d0
d1
d2
f-3
f-2
f-1
f0
f1
f2
f3

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
- $2s^2 2p^6$
- 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

SECTION 5.2
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

- Any more than the fourth and not all the orbitals will fill up.
- You simply run out of electrons

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

QUANTUM NUMBERS

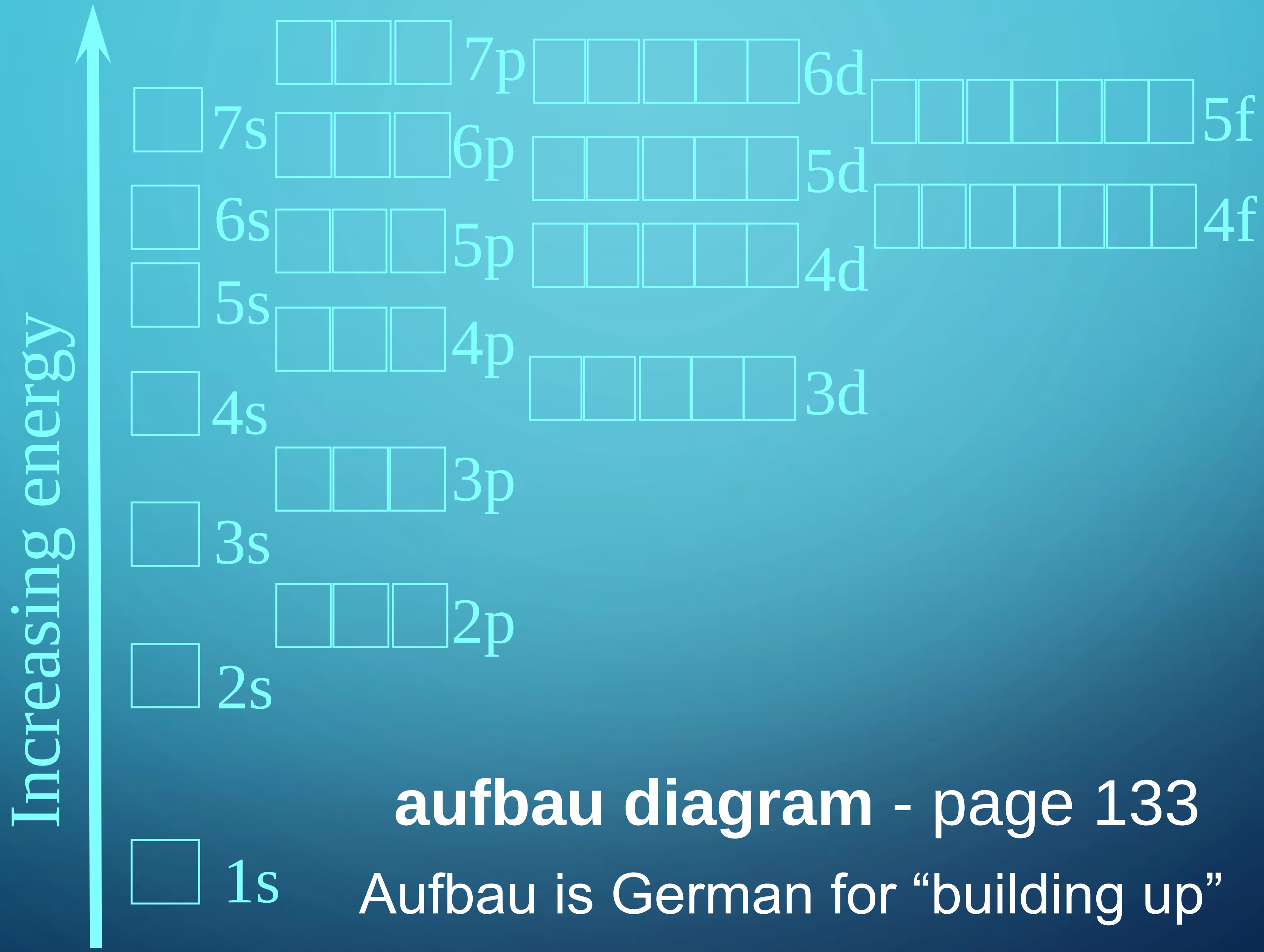
Each electron in an atom has a unique set of 4 quantum numbers which describe it.

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”

= filling order

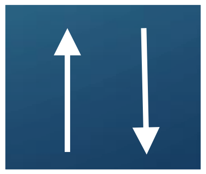
n=1	$1s^2$				$2e^-$
n=2	$2s^2$	$2p^6$			$8e^-$
n=3	$3s^2$	$3p^6$	$3d^{10}$		$18e^-$
n=4	$4s^2$	$4p^6$	$4d^{10}$	$4f^{14}$	$32e^-$
n=5	$5s^2$	$5p^6$	$5d^{10}$	$5f^{14}$	...
n=6	$6s^2$	$6p^6$	$6d^{10}$	...	...
n=7	$7s^2$	$7p^6$	...	...	...

RULE 2: PAULI EXCLUSION PRINCIPLE

Wolfgang Pauli

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

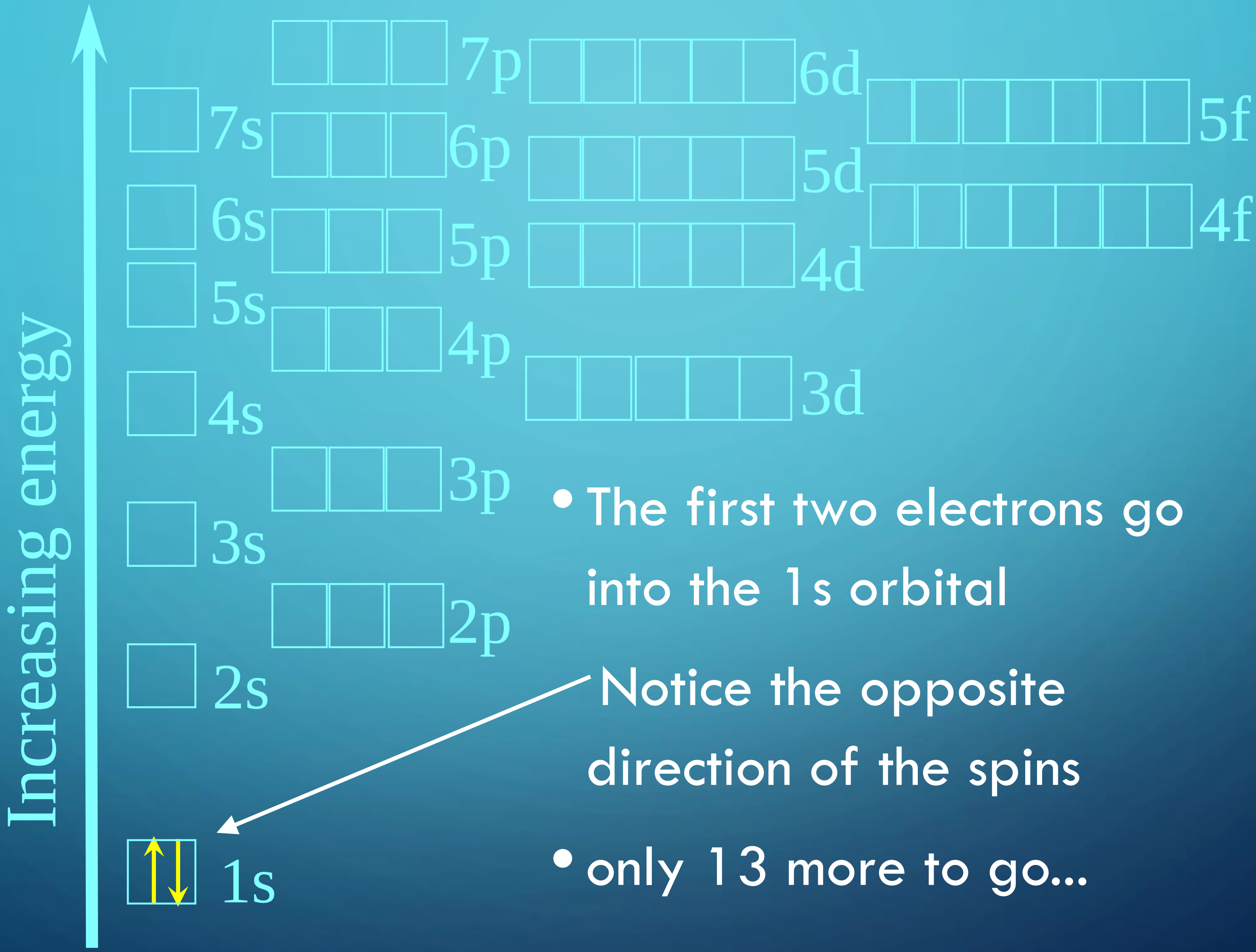
To show the different direction of spin, a pair in the same orbital is written as:



ELECTRON CONFIGURATIONS

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

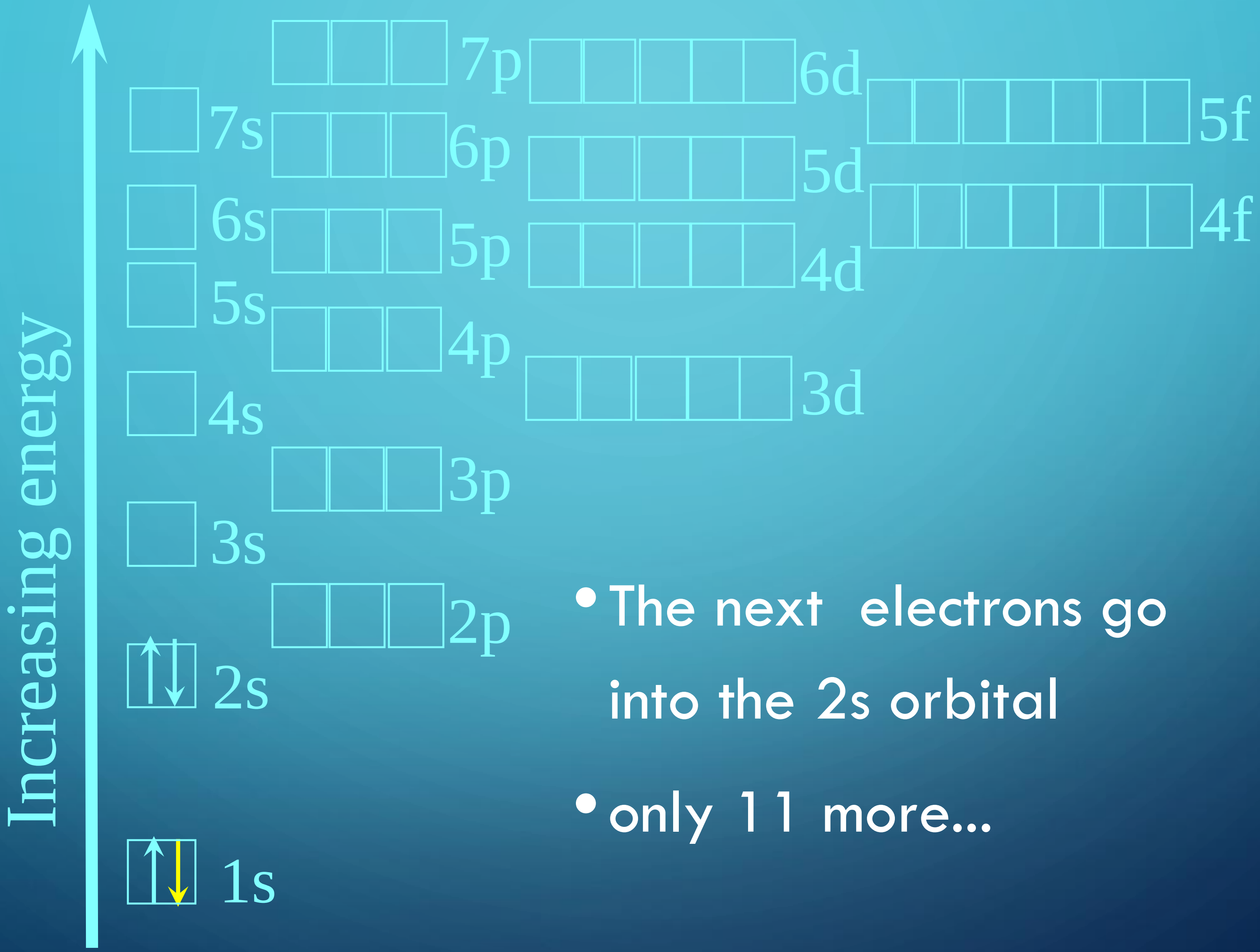
- Let's write the electron configuration for Phosphorus
  - We need to account for all 15 electrons in phosphorus



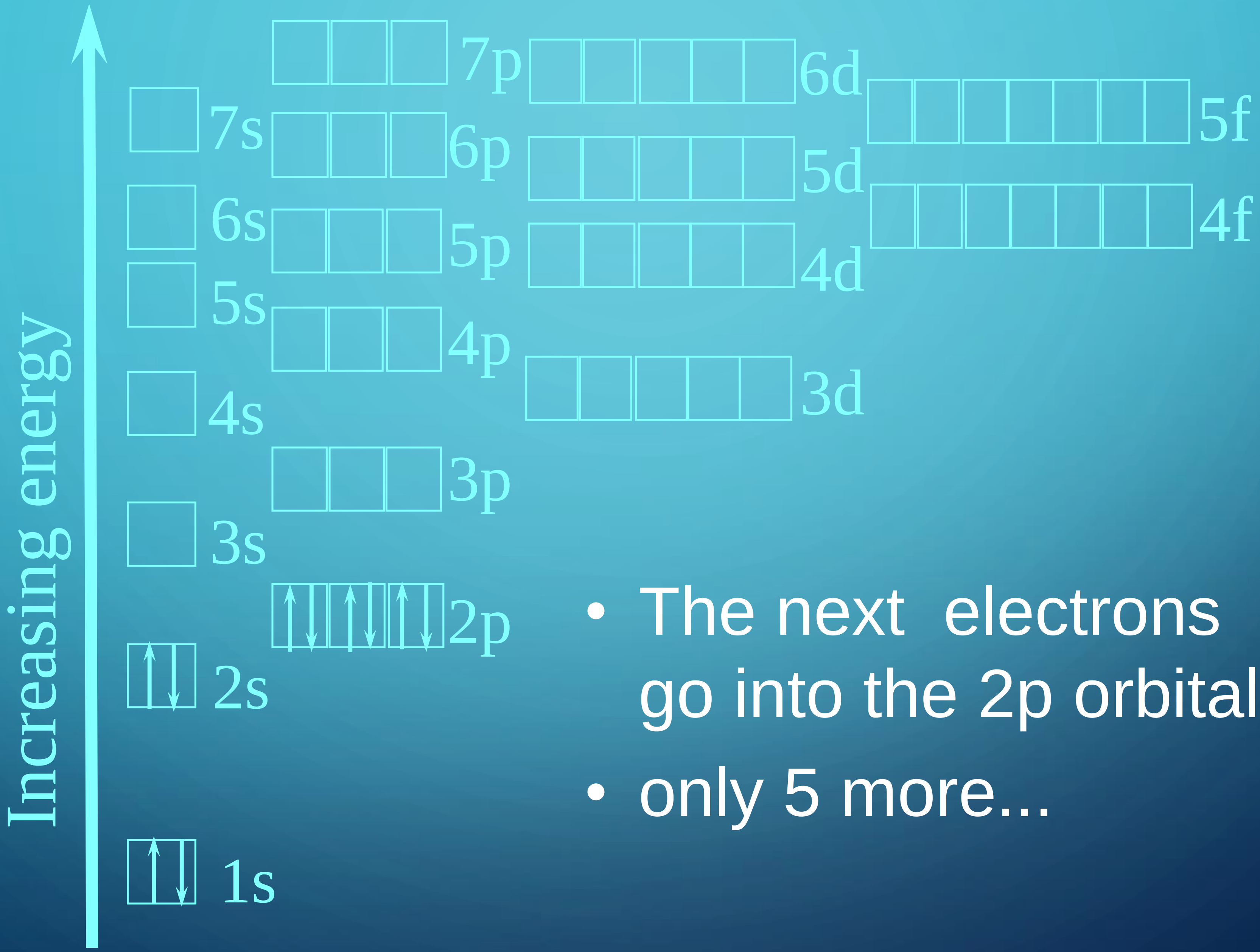
- The first two electrons go into the 1s orbital

Notice the opposite direction of the spins

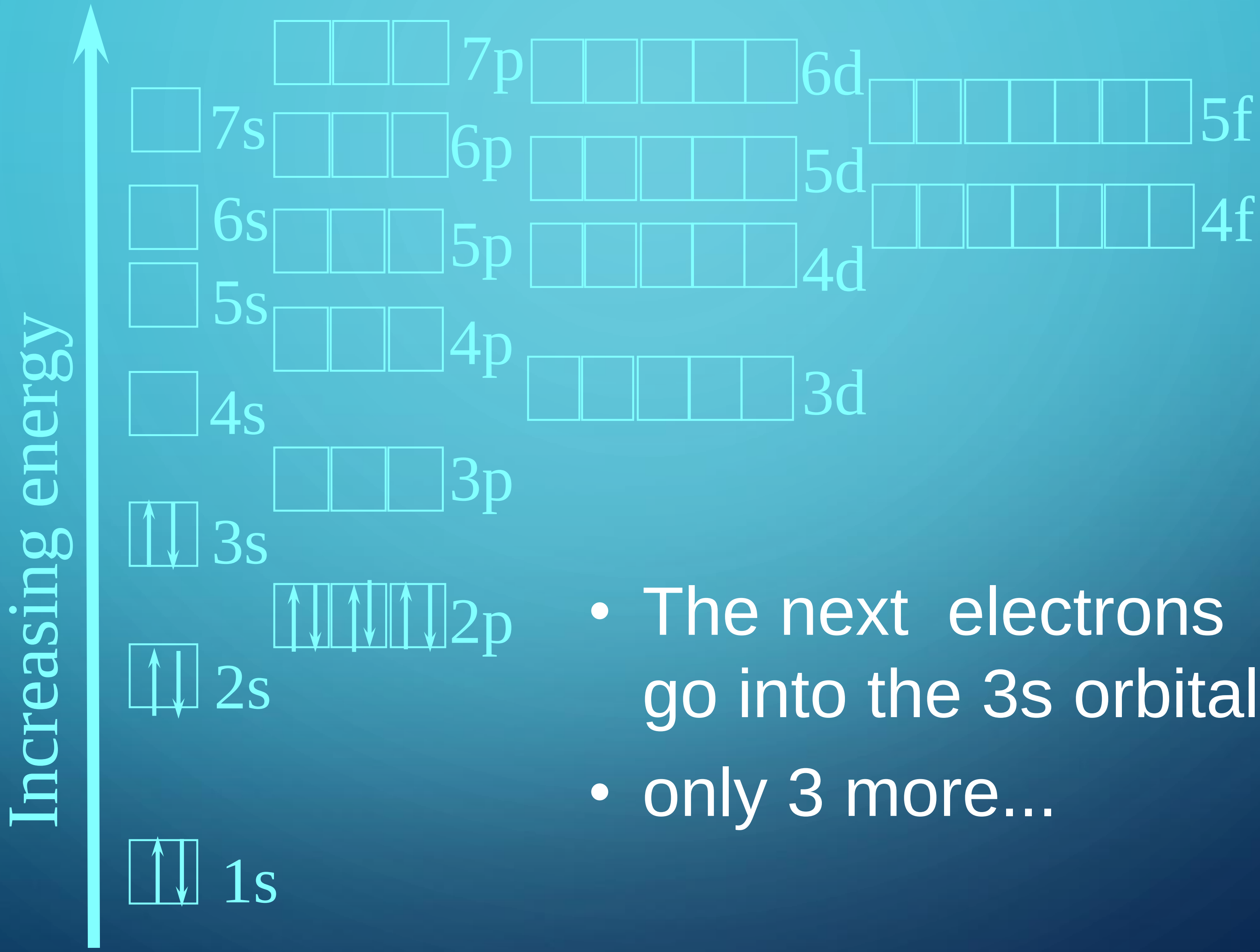
- only 13 more to go…



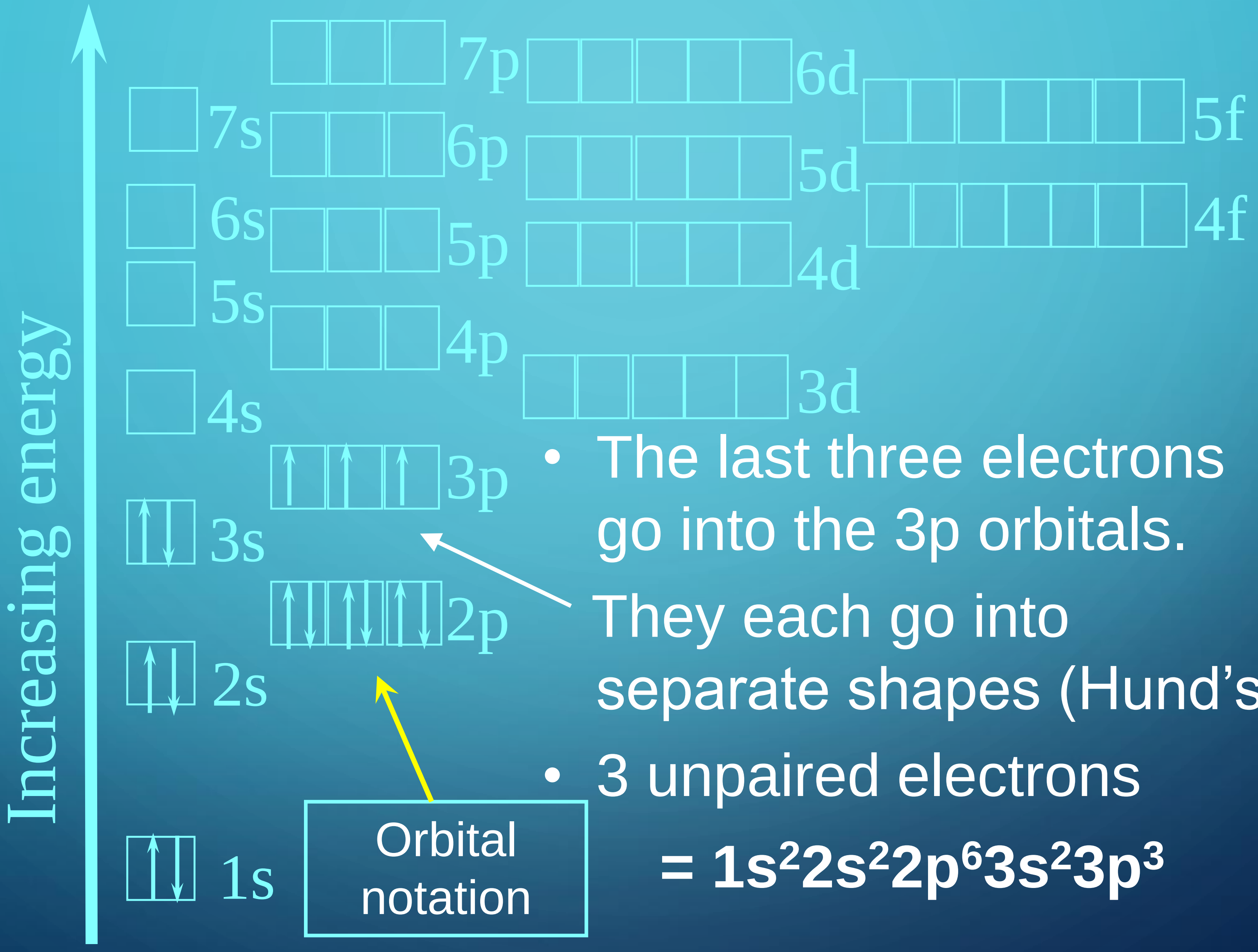
- The next electrons go into the 2s orbital
- only 11 more...



- 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

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

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:

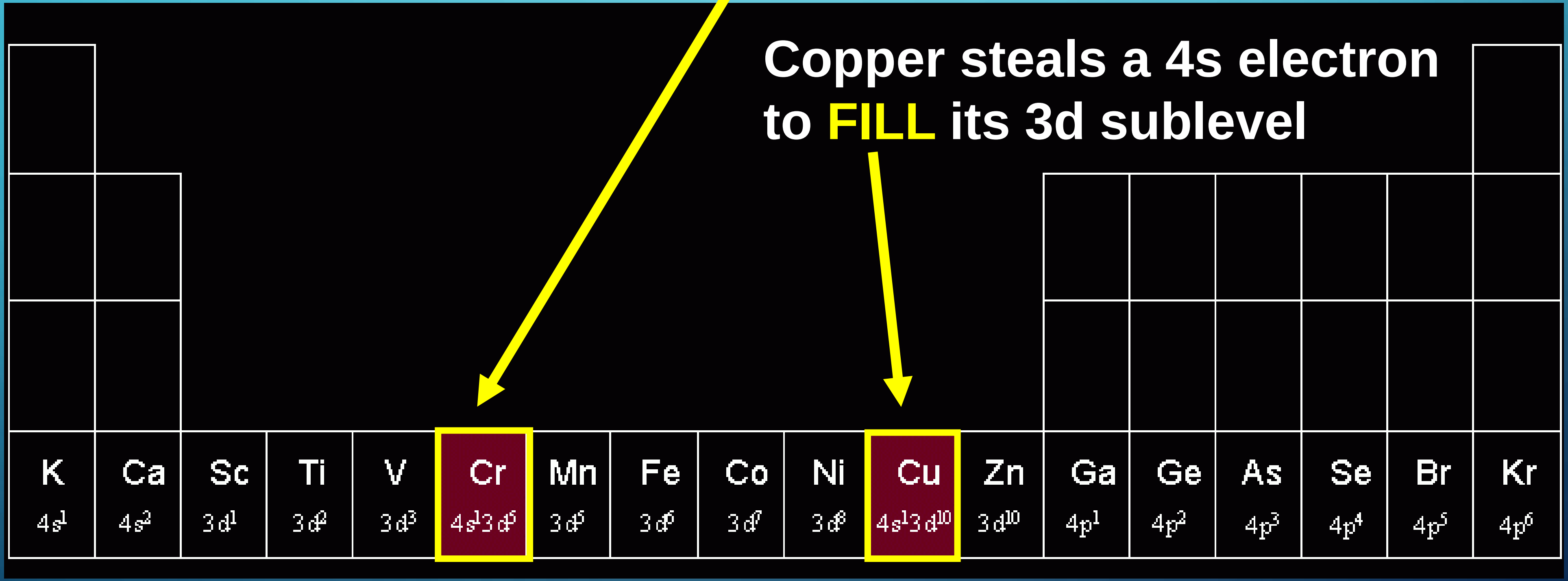
- $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$
- Why?
- This gives us two half filled orbitals (the others are all still full)
- Half full is slightly lower in energy.
- The same principal applies to copper.

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:
- $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$
- 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



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

REVIEW QUESTIONS

- Page 149
  - #s 22, 23, 26 - 34, 36, 37, 39

IONS (CHAPTER 7.1)

- OBJECTIVES:
  - Determine the number of ***valence electrons*** in an atom of a representative element.
  - Explain how the octet rule applies to atoms of metallic and nonmetallic elements.
  - Describe how **cations** form.
  - Explain how **anions** form.

KEEPING TRACK OF ELECTRONS

- Atoms in the same column…
  1) Have the same outer electron configuration.
  2) Have the same valence electrons.
- The number of valence electrons (electrons in the outer most energy level) are easily determined. It is the group number for a representative element
- Group 2A: Be, Mg, Ca, etc.
  - have 2 valence electrons

ELECTRON DOT DIAGRAMS ARE…

- A way of showing & keeping track of valence electrons.
- How to write them?
- Write the symbol - it represents the nucleus and inner (core) electrons
- Put one dot for each valence electron (**8 maximum**)
- They don’t pair up until they have to (Hund’s rule)

The Electron Dot diagram for Nitrogen

- Nitrogen has 5 valence electrons to show.
- First we write the symbol.
- Then add 1 electron at a time to each side.
- Now they are forced to pair up.
  - We have now written the electron dot diagram for Nitrogen.

The Octet Rule

- In Chapter 6, we learned that noble gases are unreactive in chemical reactions
- In 1916, Gilbert Lewis used this fact to explain why atoms form certain kinds of ions and molecules
- The Octet Rule: in forming compounds, atoms tend to achieve a noble gas configuration; 8 in the outer level is stable
  - Each noble gas (except He, which has 2) has 8 electrons in the outer level

FORMATION OF CATIONS

- Metals *lose* electrons to attain a noble gas configuration.
- They make positive ions (**cations**)
- If we look at the electron configuration, it makes sense to lose electrons:
- Na **$1s^2 2s^2 2p^6 3s^1$** 1 valence electron
- Na^{1+} **$1s^2 2s^2 2p^6$** This is a noble gas configuration with 8 electrons in the outer level.

ELECTRON DOTS FOR CATIONS

- Metals will have few valence electrons (usually 3 or less); calcium has only 2 valence electrons

ELECTRON DOTS FOR CATIONS

- Metals will have few valence electrons
- Metals will *lose the valence electrons*

ELECTRON DOTS FOR CATIONS

- Metals will have few valence electrons
- Metals will *lose the valence electrons*
- Forming positive ions

Ca^{2+}

This is named the "calcium ion".

NO DOTS are now shown for the cation.

ELECTRON DOTS FOR CATIONS

- Let's do Scandium, #21
- The electron configuration is: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^1$
- Thus, it can lose $2e^-$ (making it 2+), or lose $3e^-$ (making 3+)

- •Sc• = Sc^{2+}
  Scandium (II) ion
- •Sc• (with a third dot below) = Sc^{3+}
  Scandium (III) ion

ELECTRON DOTS FOR CATIONS

- Let's do Silver, element #47
- *Predicted* configuration is: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^9$
- *Actual* configuration is: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^1 4d^{10}$

Ag • = Ag^{1+} (can't lose any more, charges of 3+ or greater are uncommon)

ELECTRON DOTS FOR CATIONS

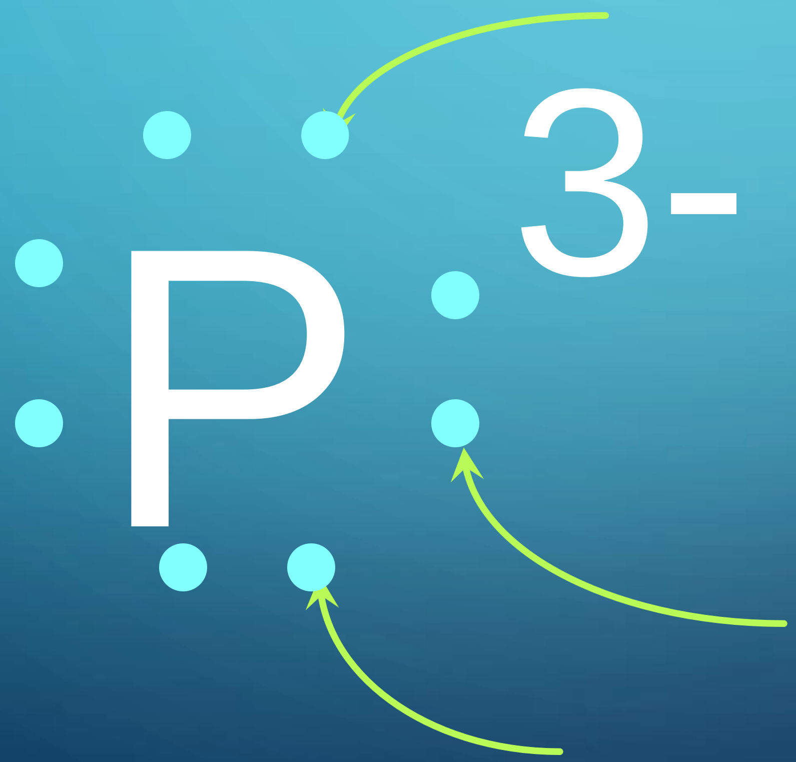
- Silver did the best job it could, but it did not achieve a ***true*** Noble Gas configuration
- Instead, it is called a **“pseudo-noble gas configuration”**

ELECTRON CONFIGURATIONS: ANIONS

- Nonmetals **gain** electrons to attain noble gas configuration.
- They make negative ions (**anions**)
- **S = $1s^2 2s^2 2p^6 3s^2 3p^4$** = 6 valence electrons
- **S^{2-} = $1s^2 2s^2 2p^6 3s^2 3p^6$** = noble gas configuration.
- ***Halide ions*** are ions from chlorine or other halogens that gain electrons

ELECTRON DOTS FOR ANIONS

- Nonmetals will have many valence electrons (usually 5 or more)
- They will *gain* electrons to fill outer shell.



STABLE ELECTRON CONFIGURATIONS

- All atoms react to try and achieve a noble gas configuration.
- Noble gases have 2 s and 6 p electrons.
- 8 valence electrons = already stable!
- This is the *octet rule* (8 in the outer level is particularly stable).

REVIEW QUESTIONS

- Page 193
  - #s 3 – 10.

TEST PREPARATION QUESTIONS

- Page 122
  - #s 38 – 45, 47, 49 – 52, 55, 65.
- Page 150 - 152
  - #s 50 – 53, 57, 60, 62, 64, 65, 68, 70, 72
- Page 207
  - #s 1 – 40.