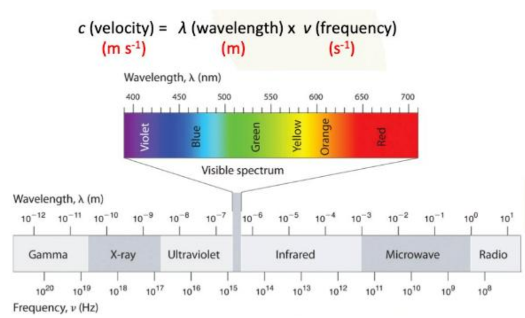
Name: ______________________ Date: __________ Period: ______

IB Chem 1 Unit 3 Lesson 1_ Electronic Configuration (1) - Guided Notes

Slide 3: The Electromagnetic Spectrum

- ____________ 5 of the data booklet
- High energy
- High frequency
- Short wavelength
- Low energy
- Low frequency
- Long wavelength



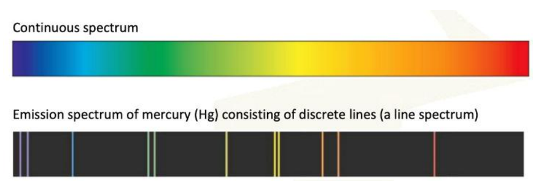
Slide 4: When electrons gain energy, they become excited and move to higher energy levels

Electrons emit energy when they drop back down to lower levels.

Each element produces a unique emission spectrum.

Emission spectra are not continuous (have discrete lines at particular wavelengths)

- Emission Spectra
- Lines represent ____________ between energy levels

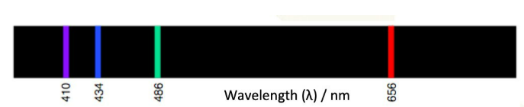


Slide 5: Hydrogen produces other emission spectra that can't be seen with the naked eye (UV or infrared)

All spectra have discrete lines that converge towards the high energy end of the spectrum

Provide evidence that the electron in hydrogen can only exist in certain fixed energy levels.

- ____________ Emission Spectrum

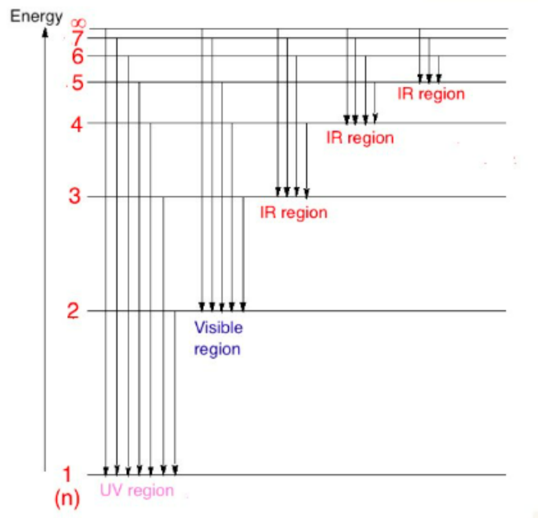


Slide 6: n is the principal quantum number (indicates the energy level/size of orbital)

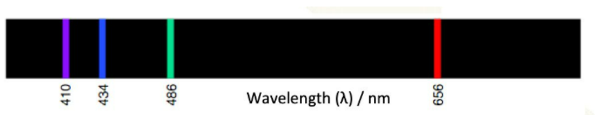
The visible emission spectrum of hydrogen is due to electrons falling to the n=2 level from higher levels.

As the value of n increases, the energy levels get closer together, causing the lines in the spectrum to converge.

- _____________ Emission Spectrum
- *n* = 3 *n* = 2 656 nm
- *n* = 4 *n* = 2 486 nm
- *n* = 5 *n* = 2 434 nm
- *n* = 6 *n* = 2 410 nm



Slide 7: Balmer series



Slide 8: Explain why the series of lines formed when electrons fall to the n = 1 level occur in the ultraviolet region of the spectrum
State which transition will give the line with the shortest wavelength (highest frequency) in the visible region of the spectrum
Explain why the transition from n=4 to n=3 occurs in the infrared region of the spectrum whereas the transition from n=4 to n=2 occurs in the visible region of the spectrum.
Deduce which transition will emit the same amount of energy as the ionization energy of hydrogen (the ionization energy is the energy change when an electron is completely removed from the atom to form a positive ion).

- Let's Practice!

Slide 9: Organization of orbitals

- **Main energy levels (shells)**

- ____________ exist in main energy levels, indicated by the principal quantum number (*n*).
- Only a fixed number of ____________ can exist in each main energy level ($2n^2$)
- **Sub-levels (____________)**
- Main energy levels can be split into sub-levels.
- Each sub-level can only contain a fixed number of ____________ (depending on how many orbitals is has)

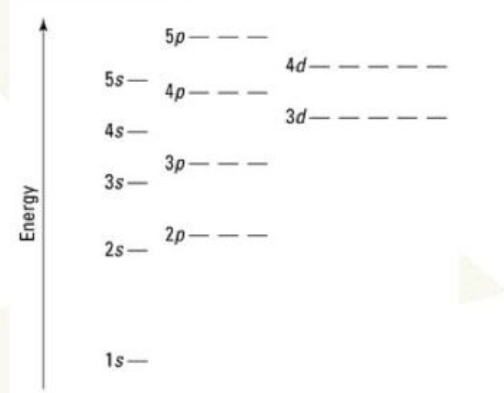
Level	Maximum number of electrons
$n = 1$	2
$n = 2$	8
$n = 3$	18
$n = 4$	32

Slide 10: Electrons exist in orbitals that are located in discrete energy levels
Each orbital can hold a maximum of 2 electrons that are spinning in opposite directions

- ____________ of energy levels

Main level	Number of each type of orbital			
(*n*)	s	p	d	f
1	1	-	-	-
2	1	3	-	-
3	1	3	5	-
4	1	3	5	7

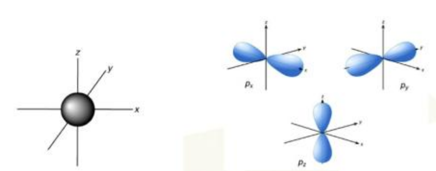
Arrangement of sub-levels within an atom



Note that the 3d sub-level is higher in energy than the 4s sub-level.

Slide 12: Atomic Orbitals

- **Heisenberg's ____________ Principle:** It is impossible to state the exact position of an electron at a precise moment in time (you can know where it is or how fast it is moving, but not both)
- **Describes a volume in space where there is a high ____________ that the electron is located**
- The five d orbitals are oriented at 45° angle to the x, y, z planes and are '____________' shaped.
- ***Note: only 2 ____________ can fit in an atomic orbital**



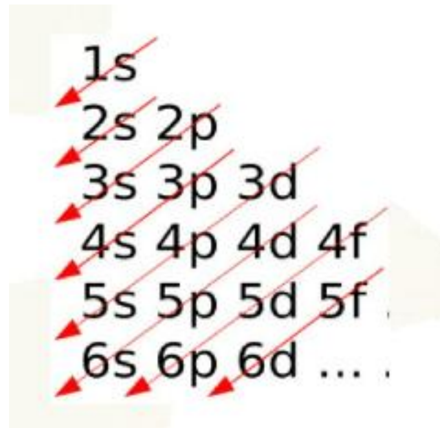
s orbital – spherical in shape

The three p orbitals are orthogonal (at 90° to each other) and are 'dumb-bell' shaped.

Slide 13: f orbitals- A little too much

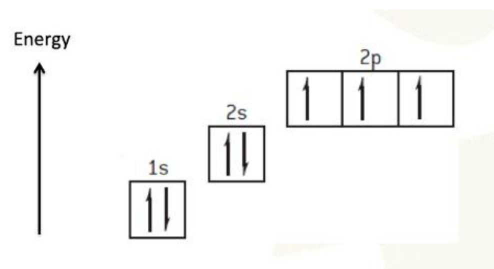
Slide 14: How electrons fill energy levels

- Electron arrangement in atoms (aka **electron ____________**) can be deduced by applying the following three rules:
- **1. Aufbau ____________:** Electrons fill the lowest available energy levels before filling higher energy levels.
- **2. Hund's rule:** If two or more orbitals of equal energy are ____________, electrons will occupy them singly before pairing up.
- **3. Pauli ____________ principle:** No two electrons can have the same four quantum numbers. This means each orbital can only contain a maximum of two electrons (one spin-up and one spin-down).



Slide 15: Orbital Diagrams

- Useful for ____________ Hund's Rule
- **Orbital diagram for ____________:**



Slide 16: Electron Configurations

- Electron ____________ for nitrogen:
- $1s^2 2s^2 2p^3$
- **Energy level**
- **Sublevel**
- **# of ____________ in that energy level**

Slide 17: Electron Configurations for first 3 energy levels

H $1s^1$	Li $1s^2 2s^1$	Na $1s^2 2s^2 2p^6 3s^1$	K $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$
He $1s^2$	Be $1s^2 2s^2$	Mg $1s^2 2s^2 2p^6 3s^2$	Ca $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$
	B $1s^2 2s^2 2p^1$	Al $1s^2 2s^2 2p^6 3s^2 3p^1$	Sc $1s^2 2s^2 2p^6 3s^2 3p^6 3d^1 4s^2$
	C $1s^2 2s^2 2p^2$	Si $1s^2 2s^2 2p^6 3s^2 3p^2$	Ti $1s^2 2s^2 2p^6 3s^2 3p^6 3d^2 4s^2$
	N $1s^2 2s^2 2p^3$	P $1s^2 2s^2 2p^6 3s^2 3p^3$	V $1s^2 2s^2 2p^6 3s^2 3p^6 3d^3 4s^2$
	O $1s^2 2s^2 2p^4$	S $1s^2 2s^2 2p^6 3s^2 3p^4$	Cr $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5 4s^1$
	F $1s^2 2s^2 2p^5$	Cl $1s^2 2s^2 2p^6 3s^2 3p^5$	Mn $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5 4s^2$
	Ne $1s^2 2s^2 2p^6$	Ar $1s^2 2s^2 2p^6 3s^2 3p^6$	Fe $1s^2 2s^2 2p^6 3s^2 3p^6 3d^6 4s^2$
			Co $1s^2 2s^2 2p^6 3s^2 3p^6 3d^7 4s^2$
			Ni $1s^2 2s^2 2p^6 3s^2 3p^6 3d^8 4s^2$
			Cu $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$
			Zn $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2$

Condensed electron configurations build upon the last noble gas e.g. P [Ne] $3s^2 3p^3$ and Fe [Ar] $3d^6 4s^2$

Slide 18: Noble-Gas Configurations

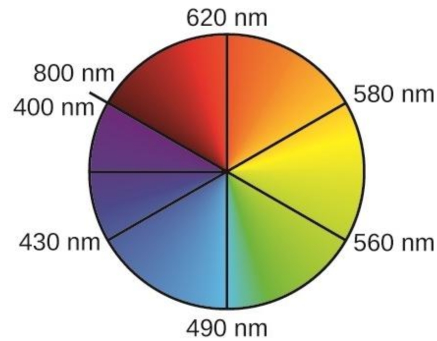
- **Build upon the electron configuration of the last noble gas**
- **Examples:**
- Phosphorus: [Ne] $3s^2 3p^3$
- Iron: [Ar] $4s^2 3d^6$

Slide 19: Transition Metals

- A filled sublevel is the most stable.
- A half-filled sublevel is more stable than a partially filled sublevel.
- Sometimes, an electron will be taken from the s sublevel and moved to the d sublevel in order to maximize stability (aka atom will be in the lowest energy state).
- It requires less energy for the electrons in the s sublevel to be removed than the d sublevel, so those electrons are taken first.
- This is especially true of elements in group 6B and 11B
- **Examples:**
- Chromium: [Ar] $4s^1 3d^5$
- Copper: [Ar] $4s^1 3d^{10}$

Slide 20: Transition Metals

- Copper ____________: [Ar] $4s^13d^{10}$: colored
- Cu^+ ____________: [Ar] $4s^03d^{10}$: not colored
- Cu^{2+} ____________: [Ar] $4s^03d^9$: colored
- Zn ____________: [Ar] $4s^23d^{10}$: not colored
- Zn^{2+} ____________: [Ar] $4s^03d^{10}$: not colored
- Electrons emit light at a wavelength that is ____________ to the wavelength of light they absorb.
- Unpaired electrons in ____________ metals can cause them to have a color in the presence of ligands - this is due to d-orbital splitting and movement of d-orbital electrons
- ____________, Zn is not considered a transition element due to its filled d sublevel.



(b)

Slide 21: You try!

1. Give the full and condensed electron configurations for elements 31 to 36.

2. When transition metals form an ion they lose the 4s electrons before the 3d electrons.
Give the full and condensed electron configurations for the following ions:
Mn^{2+} Fe^{3+} Cu^+ and Cu^{2+}

3. Draw an orbital diagram (arrow-in-box diagram) to show the electron configuration of chromium (only show the 4s and 3d orbitals).