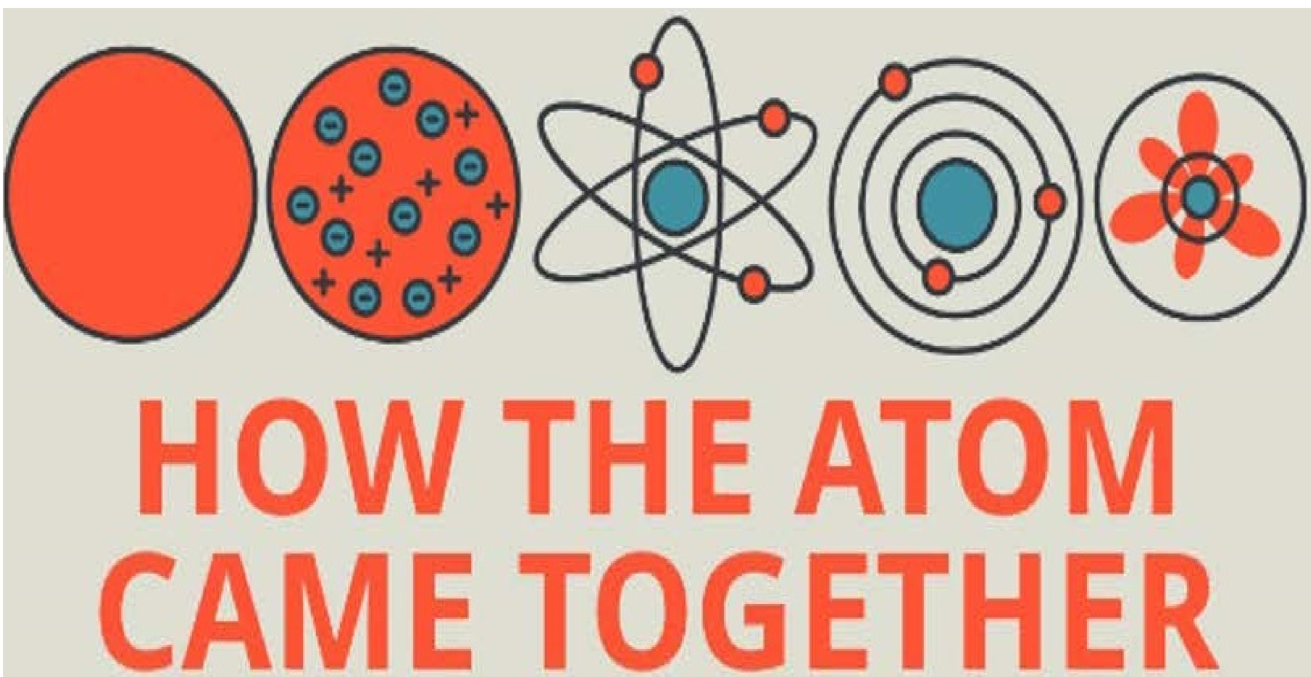
Chemistry 11

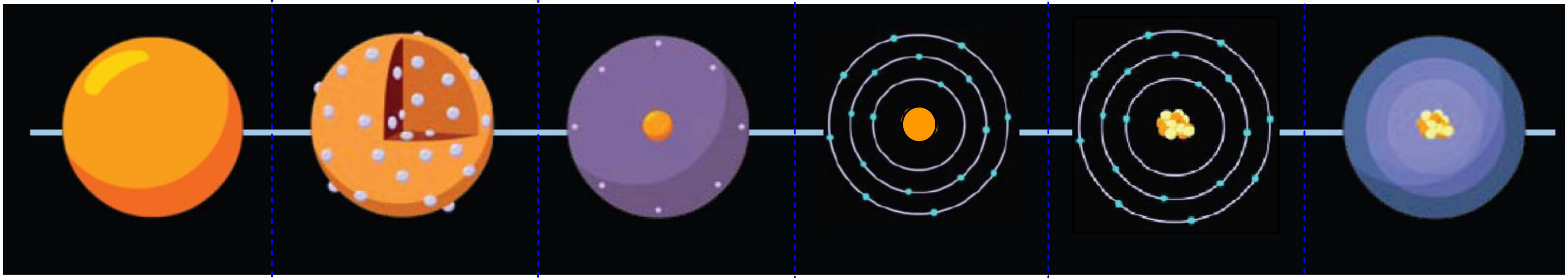
Unit 8 : Atomic Theory

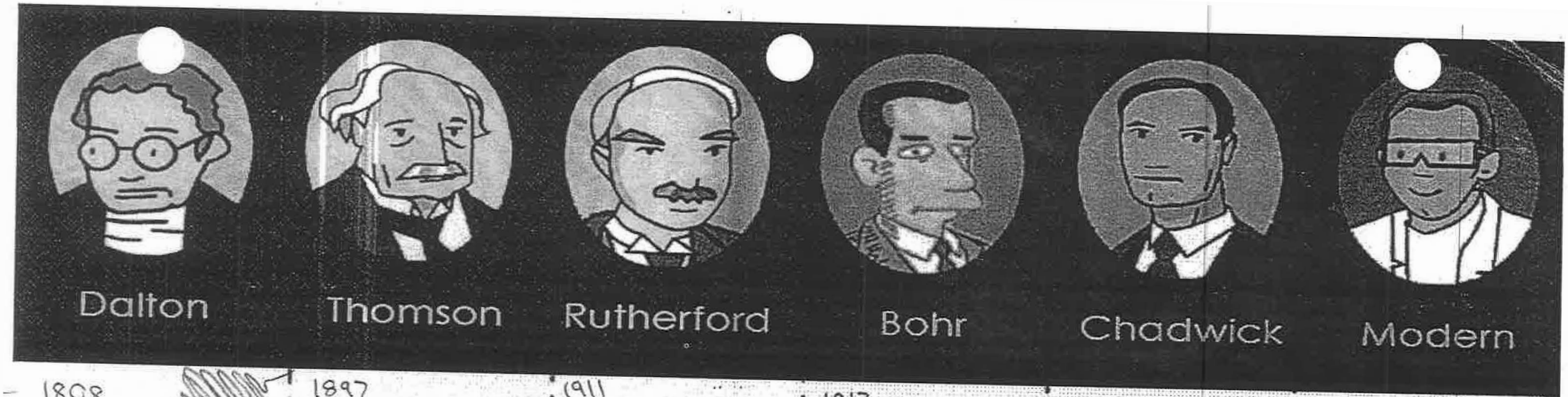


Name: Key

Block:______

Atomic Model Timeline





Dalton

- 1808
- he created the very first atomic theory
- he was an english school teacher who preformed many exp on atoms
- viewed atoms a tiny, small balls.
- atomic theory had 4 Statments
  - tiny invisible particles
  - atoms of 1 element are all the same
  - atoms of different elements are different
  - compounds form by combining atoms

↑ daltons theory

Thomson

1897

- he discovered electrons
- he was the first scientist to show that the atom was made of even smaller things
- also proposed the (+) particle
- his atomic model was known as the rasine bun model.


thomson model

the (−) charged electrons are found inside the (+) dough (mostly made out of (+) charged material)

very dense (+) center → nucleus. Some bounce back (proton) / hits

Rutherford

1911

- discovered protons and the nucleus
- he showed that atoms have (+) particles in the center, and are mostly empty space.
- he called (+) particles protons
- he called the center of atoms the nucleus.

(first person to actual prove anything → experiment)

(Gold Foil)

↓

took a laser and emits beam of alpha particles (+) (laser) at gold foil

most particles went straigh through (empty space)

others deflected (hits ...)

Bohr

1913

- improved rutherfords model
- he proposed that electrons move around the nucleus in specific layers or shells.
- every atom has a specific # of shells

Chadwick

1932

- discovered neutrons
- working with rutherford, he discovered particles with no charge
- called these particles neutrons
- neutrons are also found in the nucleus (No charge)

Modern

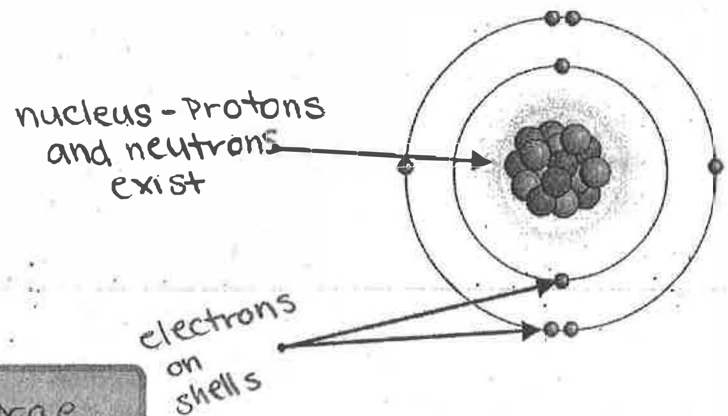
1932 —

- work done since 1920 has changed the model.
- new atomic model has electrons moving around the nucleus in a cloud
- it is impossible to know where an electron is at any given time

Atomic Structure, Isotopes & Atomic Mass

Each element is made up of very tiny particles called atoms, and each element is made up of just one particular type of atom, which is different to the atoms in any other element.

J.J Thomson discovered electrons, and proposed the existence of a (+) particle.
It wasn't until **Rutherford's** famous gold foil experiment that the + proton was discovered, and atoms were thought to me mostly empty space. He named the *centre of atoms* the nucleus
Bohr improved on this model proposing that electrons move around the nucleus in specific layers called shells.
It was James **Chadwick** who discovered particles with **no charge**, which he named neutrons



Particle	Mass	Charge
proton	1	+1
neutron	1	0
electron	almost 0	-1

How many electrons?

Atoms have no overall electrical charge and are neutral

This means atoms must have an equal number of positive protons and negative electrons.

The number of electrons is therefore the same as the atomic number.

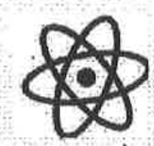
Atoms	Protons	Neutrons	Electrons
helium	2	2	2
copper	29	35	29
iodine	53	74	53

Atomic number is the number of protons rather than the number of electrons, because atoms can lose or gain electrons but do not normally lose or gain protons.

Atomic Number (Z)

The number of protons in an atom is known as the atomic number or proton number.

- always the same for a particular element. doesn't change when ions are formed
- **The number of protons identifies the element!**
- is also equal to the positive charge of the nucleus (*aka* the nuclear charges



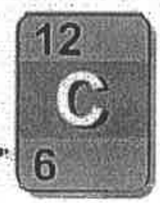
Example:

If an atom has $Z = 12$, then it MUST be an *atom* of Mg #of protons
If an ion has $Z = 41$, then it MUST be an *ion* of Nb
If the nuclear charge of a species is +24, then it MUST *be an atom or ion of* Chromium

The ***overall charge on an atom is zero*** because
the number of protons = number of electrons

The charge on any ion = number of e⁻ lost (⊕ cation)
or gained (⊖ anion).

It is the smaller of the two numbers shown in most periodic tables. *(usually on top…depends where you're looking)*



Mass Number (A)

mass number = number of protons + number of neutrons

Electrons have a mass of *almost zero*, which means that the mass of each atom results almost entirely from the number of protons and neutrons in the nucleus.

- **Does not** technically **appear in the periodic table!** *(not in this exact form)*
- Can be expressed in number of ways: mass number

  Carbon-12 or ^{12}C or $^{12}_{6}C$

  mass number; atomic number (number of protons)
- ***Does not*** uniquely identify the element!

  e.g. ^{3}H : 1 p, 2 n , ^{3}He : 2 p, 1 n

  from periodic table

The larger of the two numbers shown in most periodic tables, which you are probably familiar with is mass….actually shows the ***relative*** atomic mass .

What's the mass number?

mass number = number of protons + number of neutrons

Atomic # - proton

Atoms	Protons	Neutrons	Mass number
helium	2	2	4
copper	29	35	64
cobalt	27	32	59
iodine	53	74	127
germanium	32	41	73

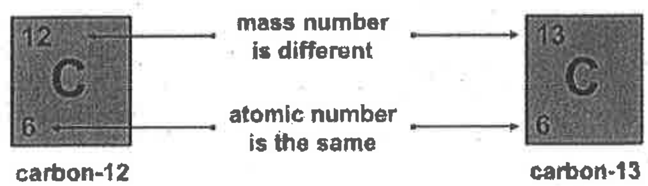
How many neutrons?

neutrons = mass # = atomic #

Atoms	Mass number	Atomic number	Neutrons
helium	4	2	2
fluorine	19	9	10
strontium	88	38	50
zirconium	91	40	51
uranium	238	92	146

What are isotopes?

isotopes are atoms of the same element that contain different numbers of neutrons.



The reactivity of different isotopes of an element is identical because they have the *same number of* electrons.

The **different** masses of the atoms means that physical properties of isotopes *are slightly different*.

atomic #

Z	Name	Symbol	Mass of Atom (u)	% Abundance
15	Phosphorus	^{31}P	30.973762	100
16	Sulphur n=16	^{32}S	31.972071	94.93
	n=17	^{33}S	32.971458	0.76
	n=18	^{34}S	33.967867	4.29
	n=19	^{36}S	35.967081	0.02

# of neutrons is changing

Only a few elements (e.g. phosphorus) are **monoisotopic** *(only have one naturally occurring isotope).*

Most elements (e.g. sulfur) have two or more isotopes.

Question: Why doesn't mass number appear on the periodic table?

49
In
114.82

Answer: because mass numbers are specific to particular isotopes and most elements are actualy a blend of two or more isotopes.

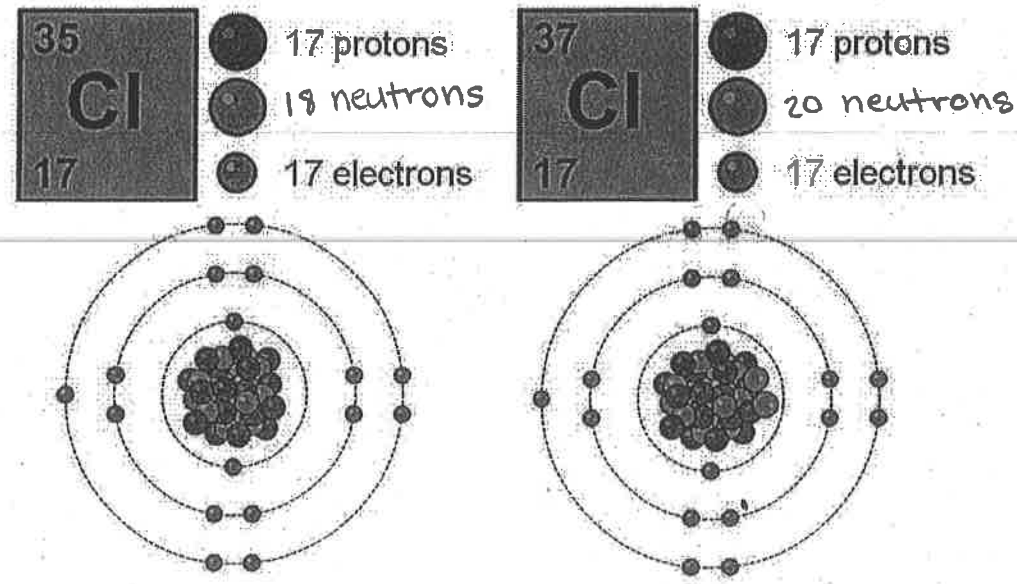
e.g. Indium has two isotopes: **Indium-113 and Indium-115.**

Indium is **4.29%** ^{113}In which has an *isotopic mass of 112.904061 u* and **95.71%** ^{115}In which has an *isotopic mass of 114.903878 u.*

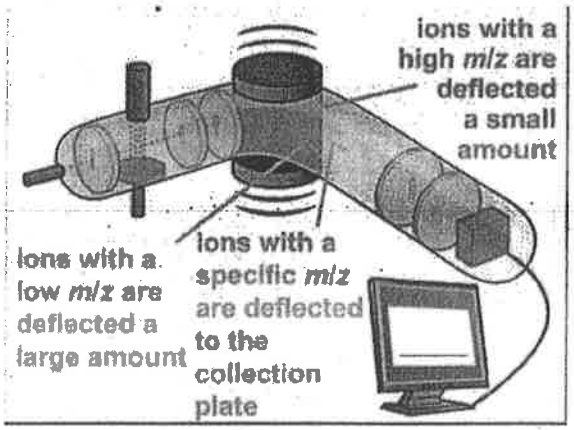
The relative atomic mass in the periodic table are the *weighted averages* of the isotopic masses of each element.

Isotopes of chlorine

About 75% of naturally-occurring chlorine is chlorine-35 (^{35}Cl) and 25% is chlorine-37 (^{37}Cl).



Isotopes & Mass Spectrometry



Mass Spectrometry uses the fact that charged particles moving through a magnetic field are **deflected** from their original path **based on their charge-to-mass ratio** (e/m or e/z)

Carbon - 14 dating
Space exploration.

<u>Describe the Process of Mass Spectrometry:</u>

① Vaporization – sample heated and vaporized

② ionization – vapor passes accross the electron beam, Knocking out the electrons leaving it positively charged.

③ Acceleration – an electric field – the positive ions are attracted to this electromagnetic and accelerate towards it.

④ Deflection – electromagnent creates an electromagnetic field that deflects the ions. the amount by which they are deflected depends on their mass - to charge ratio. Ions with specific m/z for a given strength of the EMF are deflected through a slit and onto a collecting plate.

⑤ dectection –

Calculating A_r: Most elements have more than one isotope.
The relative atomic mass of the element is the average mass of the isotopes taking into account the abundance of each isotope.

Example: what is the A_r of boron?
In a sample of boron, 20% of the atoms are ^{10}B and 80% are ^{11}B
If there are 100 atoms, then 20 atoms would be ^{10}B and 80 atoms would be ^{11}B
The relative atomic mass is calculated as follows:
A_r of B = (0.20)(10) + (0.80)(11)
Ar of B = 10.8 amu

Atomic Mass vs Mass Number

The **atomic mass** (or **atomic weight**) of each element is the weighted* average of the masses of its isotopes where the weighting depends on the abundance of each *isotope in nature*. Roughly speaking, the atomic mass is closest to isotopic mass of the **most abundant isotope.** The units are unified atomic mass units, u.

*** Note: Mass number ≠ atomic mass. The textbook is too simplistic and so the notes above are what you should study.**

Example 1: Magnesium has 3 naturally occurring isotopes which are listed:

12
Mg
24.31

Isotope	Isotopic mass (u)	% Abundance
Mg-24	23.985042	78.99 → sig fig
Mg-25	24.985837	10.00
Mg-26	25.982593	11.01

(23.985042)(0.7899) + (24.985837)(0.100) + (25.982593)(0.1101)

18.9458 + 2.49858 + 2.86068

Ar of Mg = $\boxed{24.31}$ ← found in periodic table
2 decimal

Example 2: Naturally occurring samples of carbon are 98.93% carbon-12 (isotopic mass = 12.0000 u) and 1.07% carbon-13 (isotopic mass = 13.0034 u).
0.9893
0.0107

a) Calculate the expected atomic mass of carbon.

(0.9893)(12.0000 u) + (0.0107)(13.0034 u)

= 11.8716 u + 0.1391 u   * least # d.p = 2

Ar = 12.01 u
2 decimal

answers in textbook

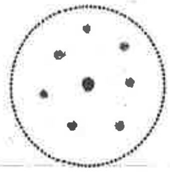
b) Calculate the molar mass of carbon.
↳ g/mol   when 1 mol of substance molar mass = atomic mass
= 12.01 g/mol

Homework

Assignment #1: page 146/147 Exercises # 13-17, 19, 20, 22, 23 a-d, 25
Complete ALL assignments on a separate piece of paper and attach to your booklet when handing in at the end of the unit. Be sure to clearly number each assignment with a heading.

The Electronic Structure of Atoms

Rutherford's Model of the Atom



Recall that, in Rutherford's model, the positively charged protons sat in the nucleus while the negatively charged electrons moved around it.
While it represented a major advance, the model could not be correct.

Here's why:

- Opposite charges attract. So the positively-charged nucleus is very attractive to the orbiting electrons.
- Particles that move in circular paths are accelerating (otherwise, their path would be straight).
- It seemed that the electrons had to be accelerating, and thus losing energy continuously as radiation.

Waves behaving like particles?
Well hit me with a Planck!

Figura 2. Niels Bohr em 1933 com seu mestre, Ernest Rutherford. Os dois se encontraram no Congresso Solvay em Bruxelas. Foto gentilmente cedida pelo Arquivo Bohr, Copenhagen

A serious challenge to Rutherford's atomic model arose almost immediately. By the end of the 1800s the physics available stated that ***accelerating charges should radiate (lose) energy***.

This meant that atoms should collapse in a fraction of a second as their electrons lost energy and spiraled around the nucleus.

Rutherford's model could not explain why the electrons didn't lose energy and spiral into the nucleus *(much like any satellite that loses gravitational energy and crashes back to Earth).*

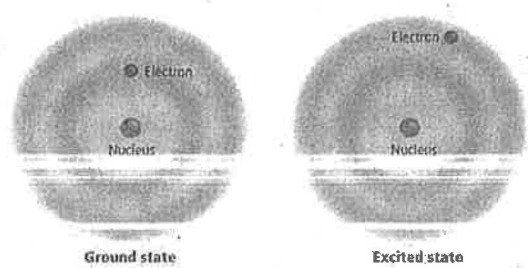
The Bohr model of the atom

In 1900, Max _Planck_____ developed his 'Quantum theory', which states that energy could be shown to behave like particles in fixed amounts he called _quanta_____.

In 1913, Niels _Bohr_____ applied Plank's theory to _electrons_, and improved upon Rutherford's model. He proposed that electrons could only exist in fixed _energy levels_____

The main energy levels are called **principal energy levels** and are given a number called the _principal quantum number_____(_n_) with the lowest in energy being 1.

- Each electron has a _fixed_____**energy = an energy level.**
- Electrons can _jump_ from one energy level to another.
- Electrons **can not** be or exist between energy levels.



A _quantum_____ of energy is the amount of energy needed to _move an electron_____ from one energy level to another energy level.

Understanding Bhor's Experiments

In 1913, Niels Bohr proposed a model that explained why the electrons stay in orbit! To understand the model, however, we first need to get some things straight about electromagnetic energy.

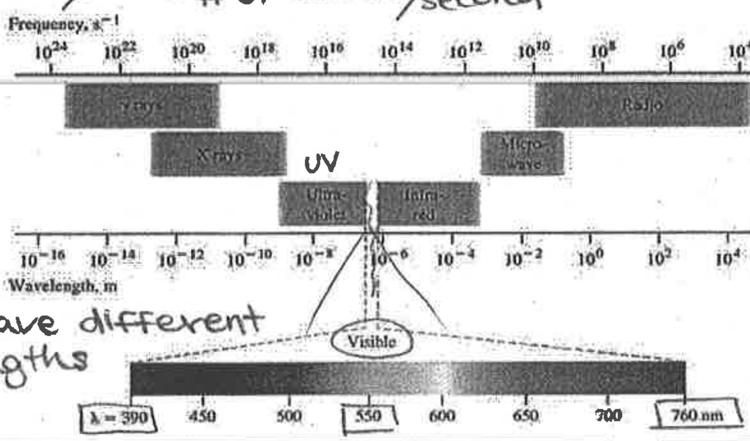
Visible light is a form of ***electromagnetic radiation (EMR)*** and all EMR is made up of photons that travel at the **speed of light, c, where c = 3.00 x 108 m/s,** or the speed of sound, 3.00 x 102 m/s.

$\lambda\uparrow$ then $f\downarrow$

$\lambda\downarrow$ then $f\uparrow$

$v = \lambda\nu$

→ # of waves/second



Colours have different wavelengths

v: velocity (m/s)
λ (lambda): wavelength (m)
ν (nu): frequency (Hz or $\frac{1}{s}$)
f

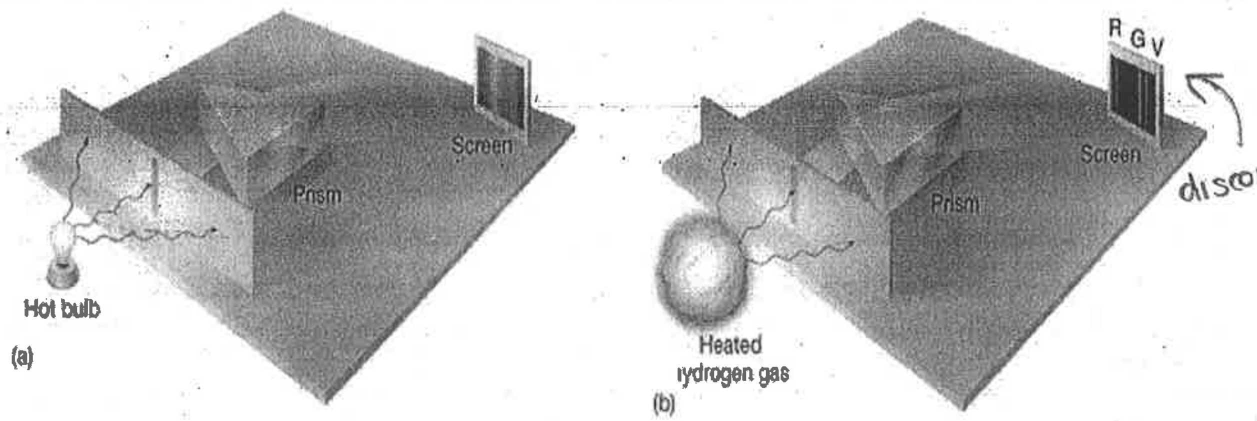
Velocity to Know:

- Speed of Light (c) = 3.00×10^8 m/s
- Speed of Sound = 3.00×10^2 m/s

Solids, liquids and gases when heated under high pressure release continuous spectra. These samples emit light that contain all colours (no gaps).

Bohr worked with hydrogen

Elemental gases when excited at low pressure (low density) release discontinuous spectra. These samples emit light that contain only certain coloured bands (separated by gaps).



discontinous

The discontinuous pattern of light bands came to be known as a bright line emission spectrum or simply an emission spectrum.

Example

Determine the wavelength of sound waves with a frequency of 556 Hz. 3.00×10^2 m/s

$\nu = 556$ Hz $\quad v = 3.00 \times 10^2$ m/s

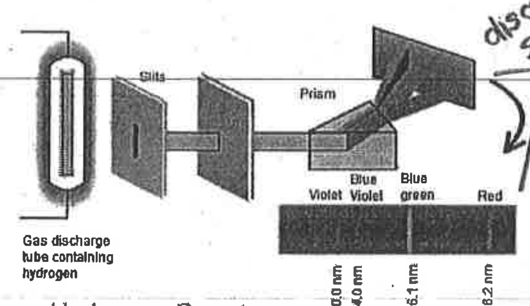
$v = \lambda\nu \quad \therefore \lambda = \frac{v}{\nu}$

$\lambda = \frac{3.00 \times 10^2 \text{ m/s}}{556 \text{ Hz}} = \boxed{0.540 \text{ m}}$

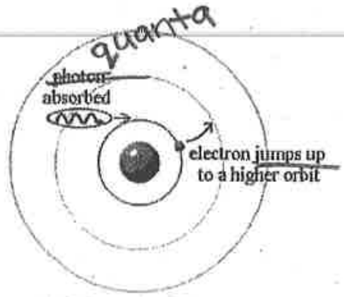
Bhor's experiments…

~~He saw that the spectroscope separated the light into its~~ component wavelengths, and for hydrogen he saw a series of Coloured lines against a black background.

Bhor applied quantum principles to explain the bright-line spectrum he saw for hydrogen.



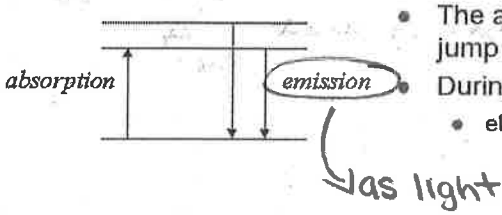
Hydrogen Spectrum



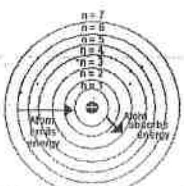
energy levels ✶

Bohr said that electrons travelled in specific paths called **orbits**.
Unlike satellites orbiting the Earth, however, an electron can only change its "altitude" by gaining or losing a specific quantum *(packet)* of energy. $E = h\nu$

- When an electron **absorbs a quantum of energy** ($h\nu$) it "jumps up" to a higher energy level (amout)
- When an electron **emits a quantum of energy** ($h\nu$) it "fall down" to a lower energy level
- The amount of energy must be (at least) some required multiple of quantum #'s or a jump cannot take place.
- During a transition, the movement is instant:
  - electrons are forbidden to be between orbits/energy level



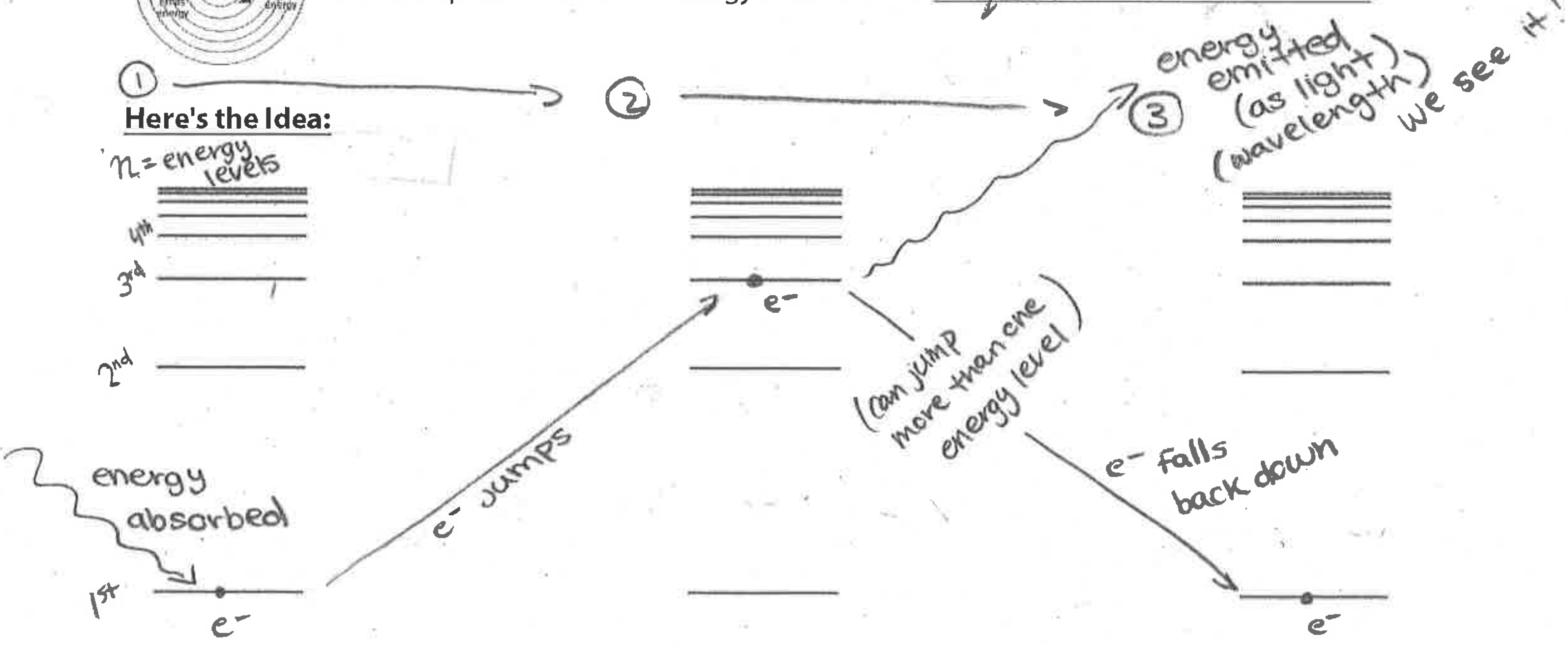
Model of how e⁻ "jump" Energy Levels



To move from one level to another, the electron must gain or lose the right amount of energy.

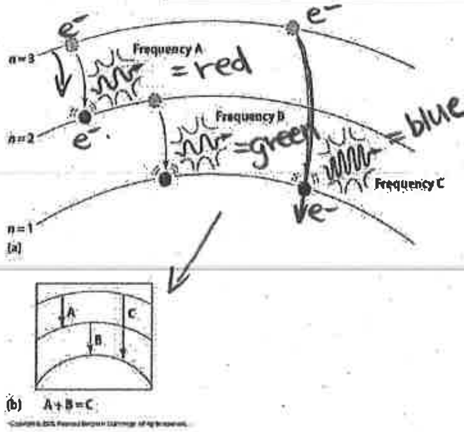
The required amount of energy is known as a quantum.

Here's the Idea:



Quantum Theory rescues the Nuclear Model

To move from one level to another, the electron **must gain or lose the right amount of** energy.



The higher the energy level, the farther it is from the nucleus.

- Gain energy to move to higher energy levels *(away from nucleus)*
- lose energy to move to lower energy levels *(closer to nucleus)*

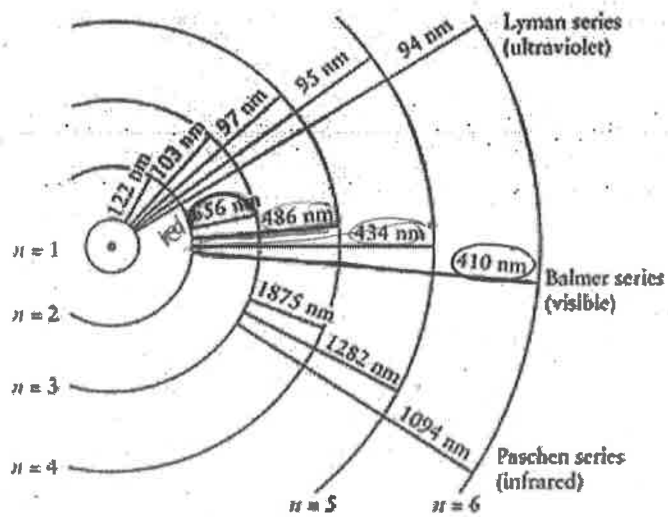
The degree to which they move from level to level determines the frequency of light they emit.

gain energy to move up level → away from nucleus.
lose energy to fall back down → towards nucleus.

Higher energy levels are closer together. This means it *takes* less *energy* to change levels in the higher energy levels.

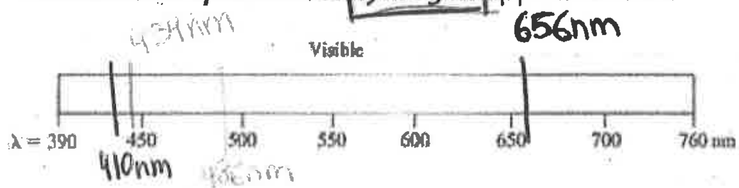
→ bad ladder

Once in a higher energy orbit, any electron could then return to a lower energy orbit by emitting a specific amount of energy corresponding to the energy difference.



- Each transition corresponds to a certain amount of energy, known as a **quantum.**
- The emission spectrum for any given element is the collection of all light-emitting quantum transitions as seen through a prism or diffraction grating.

The ***emission spectrum of hydrogen*** appears below:



Example Calculate the quantum of energy that must be absorbed for an electron in the hydrogen atom to jump from the 2nd energy level to the 3rd energy level (the red band seen in its emission spectrum).

$v = \lambda \cdot f$

$E = h \cdot f$

$v = \lambda \cdot f \therefore f = \frac{v}{\lambda}$

$= \frac{3.0 \times 10^{8}\ m/s}{656 \times 10^{-9} m}$

$f = \boxed{4.57 \times 10^{14}\ Hz}$

$E = (6.63 \times 10^{-34}\ J \cdot s)(4.57 \times 10^{14}\ Hz)$

$E = \boxed{3.03 \times 10^{-19}\ J}$

$\lambda = 636 nm$

$v = 3.0 \times 10^{8} m/s$

Calculate Frequency

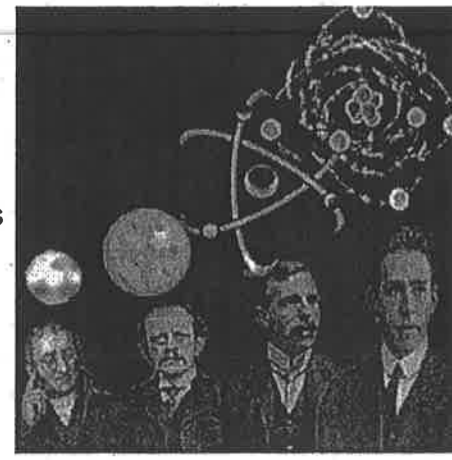
E: Energy (Joules J)
h (Planck's Constant): $6.63 \times 10^{-34}\ J \cdot S$
***v (nu)*:** $f(v)$ = Frequency (Hz or 1/s)

$E = h \cdot f$
$= h \cdot v$

The shortcomings...

Bohr's model of the Hydrogen atom was successful in explaining the mystery of bright line spectra.

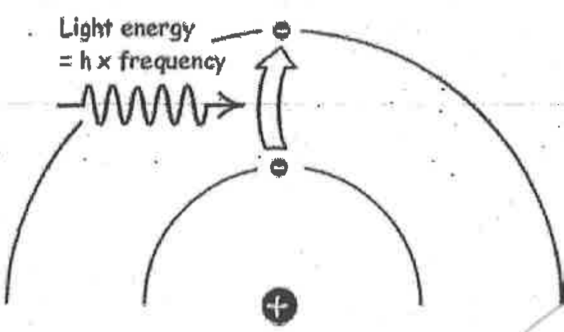
BUT his model failed to explain the energies absorbed and emitted by atoms with more **than one** electron.



Bohr had improved the Rutherford model by incorporating quantum rules.

His model explained why atoms were stable and it successfully predicted the emission spectrum for Hydrogen. Bohr's model was ultimately deficient, however, and was replaced by the quantum mechanical model.

This model successfully predicted the emission spectrum for all the elements and - nearly 100 years later - is still the best model available.



The "birth" of quantum mechanics caused a lot of trouble, however, because accepting it meant letting go of our classical assumptions about the universe.
In the Bohr model, ***electrons behaved as particles*** in orbits around the nucleus.

In the quantum mechanical model, the **electrons only predictable as waves** smeared out in regions of space known as quantum orbitals or simply orbitals.

unpredictable
regions where e⁻ exist (Heisenburg)

Did you know that an element can be identified by its emission spectra?

When atoms absorb energy, electrons move into higher energy levels. These electrons then lose energy by emitting light when they return to lower energy levels.

Mercury

Nitrogen

No two elements have the same pattern of coloured bands in their emission spectra and so ***the emission spectrum for each element acts as its fingerprint.***

Quick Check

1. Describe the appearance of hydrogen's "bright-line" spectrum.

2. Briefly indicate how electrons generate each visible line in hydrogen's emission spectrum.

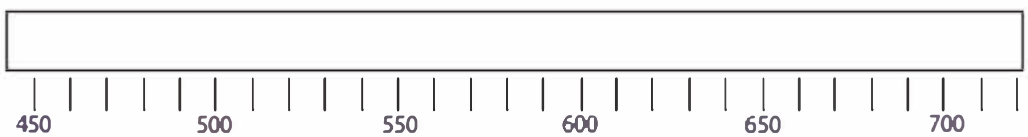
5.2 Activity: The Art of Emission Spectra

When the light emitted by vapourized and then thermally or electrically excited elements is viewed through a spectroscope, a unique line spectrum is observed for each element. In this activity, you will reproduce the emission spectra for lithium, cadmium, sodium, and strontium by colouring in the most visible spectral lines on diagrams representing those spectra.

Materials
- centimetre ruler
- various colours of felt pens or coloured pencils, including: blue, red, yellow and black

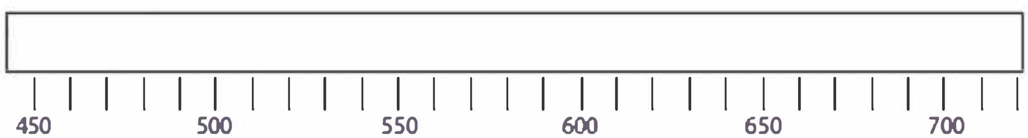
Procedure
1. Using felt pens or coloured pencils, draw vertical lines of the appropriate colour at each of the indicated wavelengths on the spectral diagrams for the elements listed below.
2. Once all the colours are drawn, use a black felt pen or coloured pencil to shade in all of the remaining space on each diagram. Be careful not to blacken out any of the coloured vertical lines.
3. After completing the diagrams, look up the emission spectrum of any other element of your choosing and draw that spectrum on the blank diagram below. You will discover that some atomic spectra include many lines while others contain only a few. Some suggestions are helium, mercury, potassium, or calcium.

Li



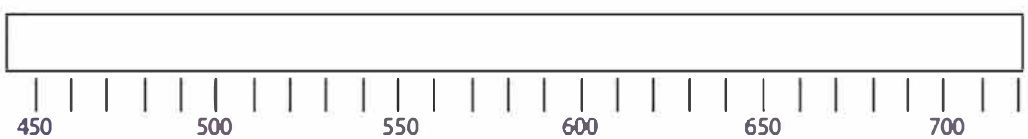
Lithium: 4 coloured lines — one blue at ~460 nm, one blue at ~496 nm, one yellow at ~610 nm, and one red at ~670 nm

Cd



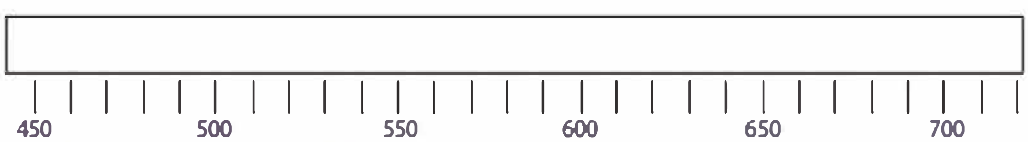
Cadmium: 5 coloured lines — three blue at ~467 nm, ~470 nm, and ~508 nm, one yellow at ~609 nm, and one red at ~642 nm

Na



Sodium: 2 bright yellow lines close together at ~590 nm

Sr



Strontium: 7 coloured lines — 4 blue at ~460 nm, ~482 nm, ~488 nm, and ~495 nm, and three red at ~670 nm, ~686 nm, and ~715 nm.

Homework *Assignment #2: # 1-10*

The Atom: Quantum Mechanics Worksheet

Answers

1. Of γ-rays and μ-waves, which has the longest wavelength? Greatest energy?

μ-waves have longer wavelengths than γ-rays. Smaller wavelength means higher frequency ($c = \lambda\nu$) and thus higher energy ($E = h\nu$), so γ-rays have more NRG.

2. Calculate the energy of one photon of yellow light (λ = 589 nm)

$$E = h\nu = h\frac{c}{\lambda} = \frac{6.63\times10^{-34}\ J\cdot\cancel{s} \times 3.00\times10^{8}\ \cancel{m}\cdot\cancel{s^{-1}}}{5.89\times10^{-7}\ \cancel{m}} = \boxed{3.38\times10^{-19}\ J}$$

3. What is a bright-line spectrum? How is it different from a continuous spectrum?

A bright-line spectrum is a collection of bright coloured lines that are emitted by excited atoms. Unlike continuous spectra, bright-line spectra have big gaps of missing colour.

4. Briefly explain why the bright-line spectrum of hydrogen is composed of discrete lines and is not a continuous spectrum.

The atom can only absorb and give off a limited # of energies. These energies correspond to a very limited # of colours - thus the discrete lines and gaps.

5. What bright-line spectrum did the Bohr model successfully predict? How did it fail?

Hydrogen. The Bohr model could not be used to predict any others.

6. What is a quantum of energy? What relationship exists between all quanta?

A specific "packet" or "bundle" of energy. They are all integer multiples of $h\nu$

7. How does an orbital differ from an orbit?

Orbitals are mathematically determined regions of space where it is likely to find a particular electron. The electron is spread out into an electron cloud and exists everywhere in the orbital all at once. In orbits, electrons behave like satellites going around the Earth in rings

8. How many electrons are in the 3rd energy level when it is exactly half-filled?

The 3rd energy level (see 3rd row of the periodic table) holds a maximum of 8 e^-. Thus it holds 4 e^- when half-filled.

The Quantum Mechanical Model & Electronic Structure of The Atom

Recall, we discussed the Bohr Model, and how at the Chemistry 11 level, this model is slightly different. Instead of electrons being arranged in shells that are a different distance from the nucleus, they are arranged in energy levels, sub-levels and orbitals.

Let's talk about those....

What is ionization energy?

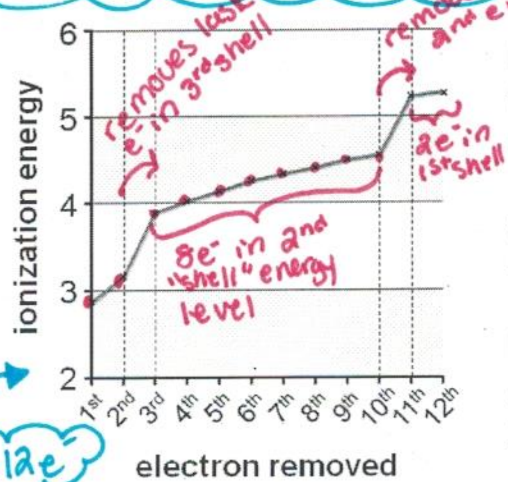
Ionization is a process in which atoms lose or gain electrons and become ions.
The first ionization (I_1) energy of an element is the energy required to REMOVE one electron from a gaseous atom. eg. $M_{(g)} \rightarrow M^- + e^-$

The **second ionization (I_2) energy** involves the removal of a second electron: $M^+ \rightarrow M^{2+} + e^-$

Looking at trends in ionization energies can reveal useful evidence for the arrangement of electrons in atoms and ions.

Evidence for Principal Energy Levels

remove (last) e⁻ in 2nd energy level.



Mg = 12e⁻

Plotting the successive ionization energies of magnesium clearly shows the existence of different energy levels, and the number of electrons at each level.

Successive ionization energies increase as *more electrons are removed.* (the more ⊕ in nucleus, and less e⁻ makes for a stronger attraction)

Large jumps in the ionization energy reveal where ***electrons are being removed from the next principal energy level,*** such as between the 2nd and 3rd, and 10th and 11th ionization energies for magnesium.

down group 2

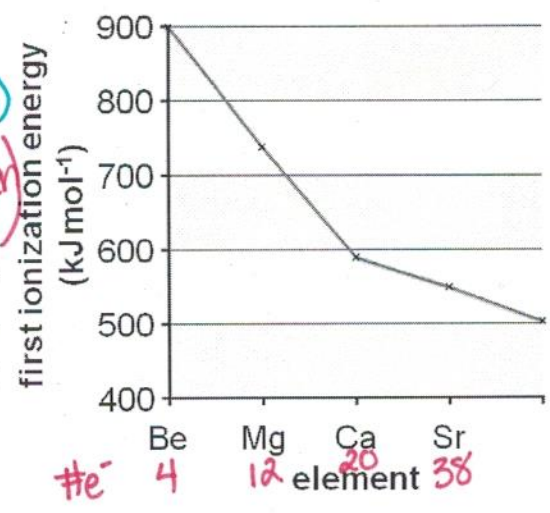
The first ionization energies of group 2 elements also show evidence for the existence of different principal energy levels.

→ # of protons > e⁻ (when removed)

Even though the nuclear charge increases down the group, the **first ionization energy** decreases. (easier to remove when further away)

This means ***electrons are being removed from successively higher energy levels***, which lie further from the nucleus and are less attracted to the nucleus.

→ higher energy levels are fuller (more e⁻) and further from nucleus.



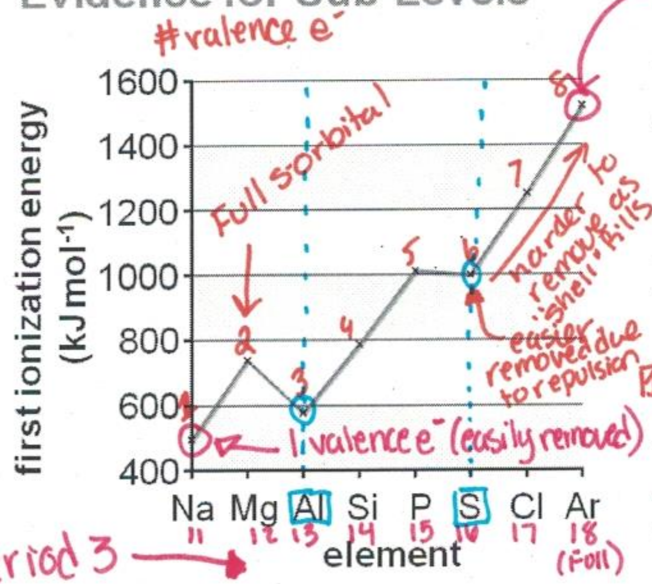
Homework *Assignment #3 Graphing Activity Ionization Energy Trends*

1. On graph paper, construct a grid using the data listed in table 1. get from teacher
2. Plot the **atomic number (x-axis)** of each element against the first **ionization energy (y-axis)** of the same element.
3. Label each point on your graph with the atomic symbol for the element.
4. **Can you observe a periodic trend? If so, describe the trend shown by your graph.**

Table 1

ELEMENT	ATOMIC NUMBER	ATOMIC RADIUS (nm)	FIRST IONIZATION ENERGY (kJ/mol)
hydrogen	1	0.037	1312
helium	2	0.05	2372
lithium	3	0.152	519
beryllium	4	0.111	900
boron	5	0.088	799
carbon	6	0.077	1088
nitrogen	7	0.070	1406
oxygen	8	0.066	1314
fluorine	9	0.064	1682
neon	10	0.070	2080
sodium	11	0.186	498
magnesium	12	0.160	736
aluminum	13	0.143	577
silicon	14	0.117	787
phosphorus	15	0.110	1063
sulfur	16	0.104	1000
chlorine	17	0.099	1255
argon	18	0.094	1519
potassium	19	0.231	418
calcium	20	0.197	590

Evidence for Sub-Levels



The first ionization energies for the elements in period 3 show a general increase

However, aluminium's value is below that of magnesium.

This suggests that the **third principal energy level** is **NOT** *one single energy level.* (levels within levels?!)

All principal energy levels contain one or more SUB-LEVELS, with different but exact energy values.

The sub-levels

There are four sub-levels.

Labelled in order of increasing energy:

s, **p**, **d** and **f**.

Each holds a different number of electrons.

increasing energy

sub-level	max no. electrons
s	2
p	6
d	10
f	14

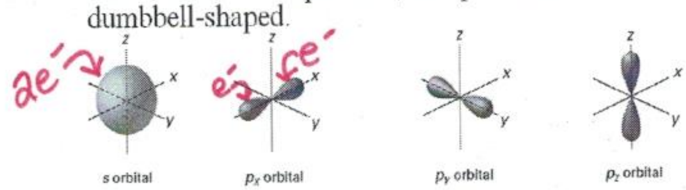
Each **principal energy level** contains a different number of sub-levels.

1	1s	2
2	2s 2p	8
3	3s 3p 3d	18
4	4s 4p 4d 4f	32

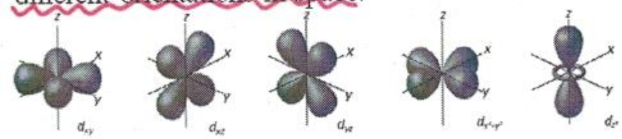
The 3rd "shell" holds more than 8e⁻ (like Bhor thought)

Shapes of Atomic Orbitals

- Different ***atomic orbitals*** are denoted by letters.
- The s orbitals are spherical, and p orbitals are dumbbell-shaped.



- Four of the five d orbitals have the same shape but different orientations in space.

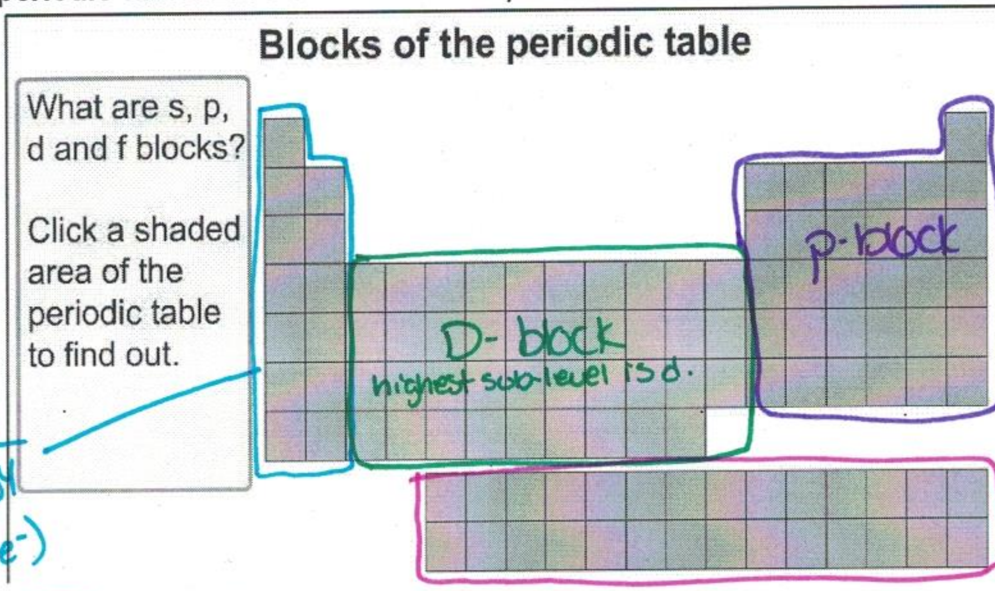


Atomic Orbitals

The numbers and kinds of atomic orbitals depend on the energy sublevel.

Energy Level, n	# of sublevels	Letter of sublevels	# of orbitals per sublevel	# of electrons in each orbital	Total electrons in energy level
1	1	s	1	2	2
2	2	s p	1 3	2 6	8
3	3	s p d	1 3 5	2 6 10	18
4	4	s p d f	1 3 5 7	2 6 10 14	32

Blocks of the periodic table: *Label the block of the periodic table & note the trends:*



S Block
Highest energy level is S
(ie: 1-2 valence e⁻)

highest sub-level is p.
(ie: 3-8 valence e⁻)

F block
·highest energy sub-level is f-orbital

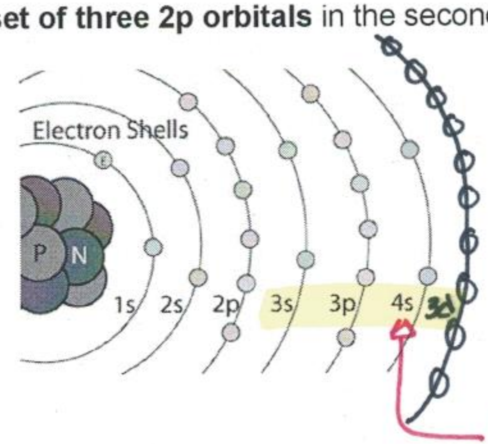
Shells, Orbitals & Sub-shells

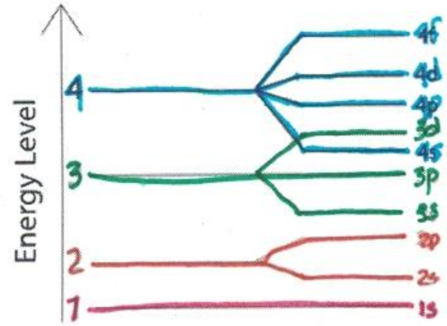
A "shell" is the set of all orbitals found in the same energy level (all have the same n-value)

- The second shell consists of the **2s** and the **2p** orbitals

A sub-shell is a set of ***orbitals of the same type*** within a shell (or principal energy level)

- The **set of three 2p orbitals** in the second energy level **(2px, 2py and 2pz)**





4s is filled BEFORE 3d

Where do the electrons go?

Types of Orbitals in each Energy Level	
	$n = 5$: s, p, d, f
	$n = 4$: s, p, d, f
	$n = 3$: s, p, d
	$n = 2$: s, p
	$n = 1$: s

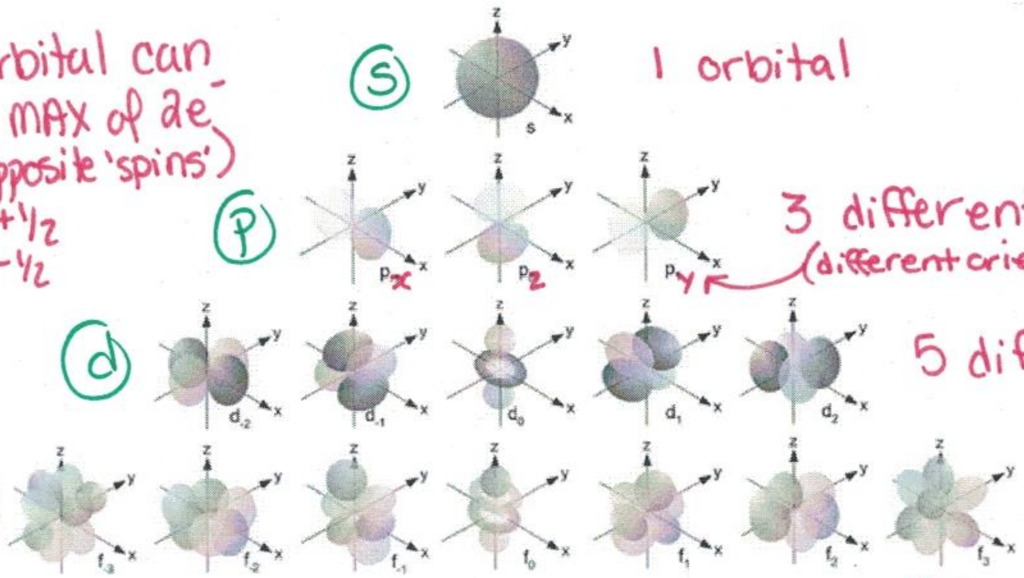
s - type subshells contain MAX $2e^-$

p - type subshells contain MAX $6e^-$

d - type subshells contain MAX $10e^-$

f - type subshells contain MAX $14e^-$

Each orbital can hold a MAX of $2e^-$ (with opposite 'spins')
e^- $+\frac{1}{2}$
e^- $-\frac{1}{2}$



1 orbital

3 different orbitals (different orientations)

5 different orbitals

7 different orbitals

*no $2e^-$ can have the same set of quantum numbers

The Aufbau principle:

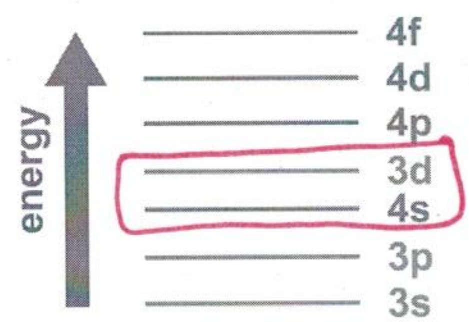
As part of his work on electron configuration, Niels Bohr developed the **Aufbau principle**, which states how electrons occupy sub-levels.

The Aufbau principle states that the lowest energy sub-levels are occupied (filled) first.

This means the 1s sub-level is filled first, followed by 2s, 2p, 3s and 3p.

However, the **4s sub-level** is LOWER than the **3d**, so this will fill FIRST.

*very important exception to remember



Electron configuration & Energy Level Diagrams

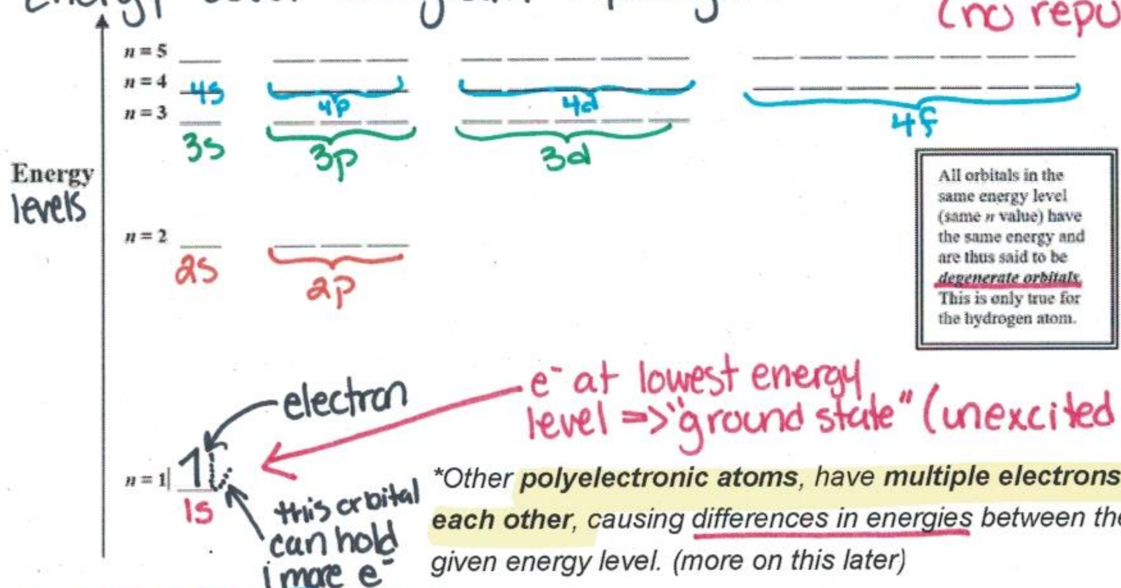
1s
2s 2p
3s 3p 3d
4s 4p 4d 4f
5s 5p 5d 5f
6s 6p 6d 6f

1s 2s 2p 3s 3p 4s 3d 4p 5s 4d 5p 4f 5d

Electrons fill lower energy sub-levels first.

Although the 3d sub level is in a lower principal energy level than the 4s sub level it is actually higher in energy.

Energy-Level Diagram: Hydrogen ← unique b/c of single e^- (no repulsive forces)



e^- at lowest energy level ⇒ "ground state" (unexcited e^-)

Other polyelectronic atoms*, have **multiple electrons which repel each other*, causing differences in energies between the subshells in a given energy level. (more on this later)*

Writing electron configuration:

- The electron configuration of an atom is a shorthand method of writing the location of electrons by energy level + orbital
- The sublevel is written followed by a superscript with the number of electrons ________ in the sublevel.

Example If the p sublevel contains electrons it is written p

E A P E *Write the electron configuration for the following elements:*

Sodium
11
Na
22.990

2e⁻ in 1s sublevel — 6e⁻ in 2p sublevel

$1s^2 2s^2 2p^6 3s^1$ ⟹ $2+2+6+1 = 11e^-$ (Atomic number) for a neutral Na

1e⁻ in 3s sublevel

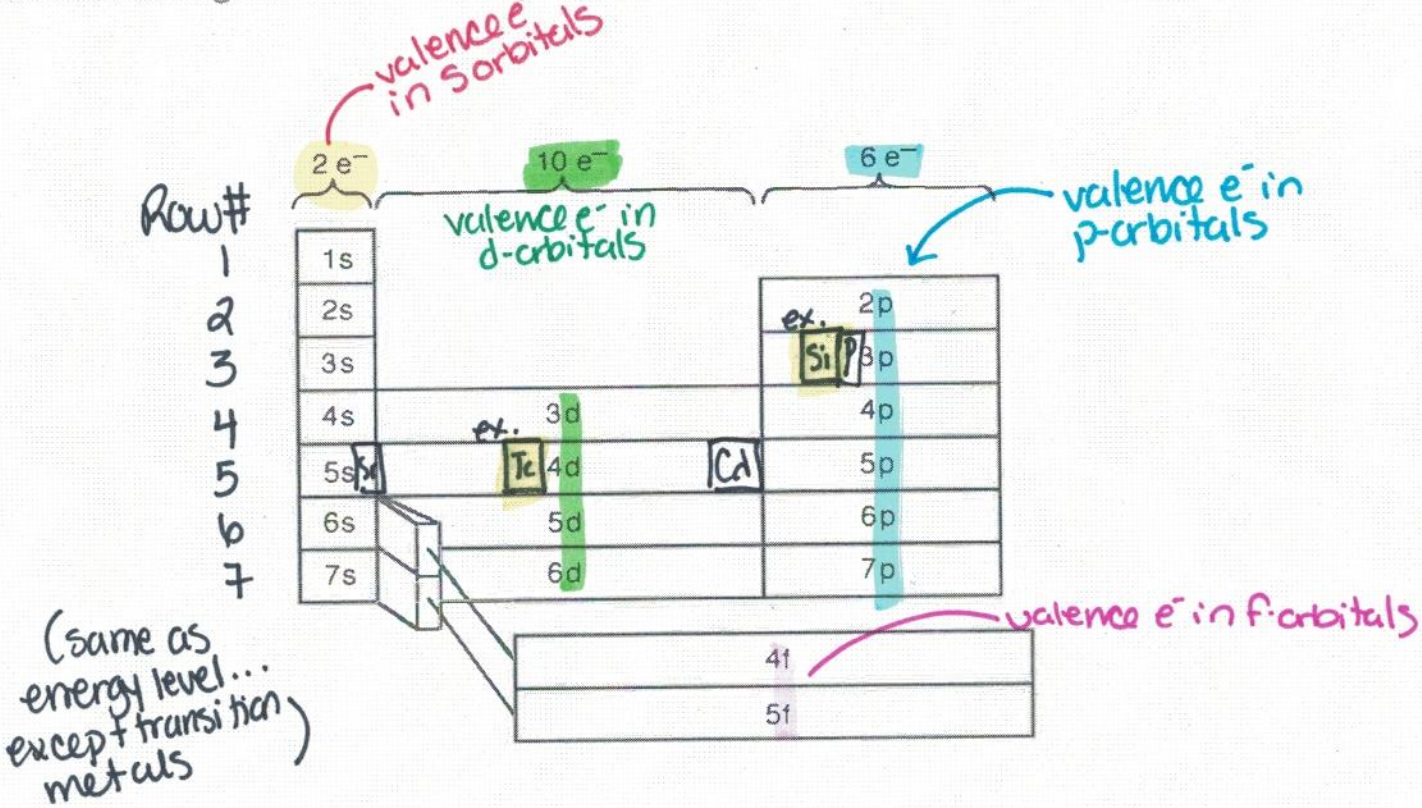
2e⁻ in 2s sublevel

Oxygen
8
O
15.999

8e⁻ total:

$1s^2 2s^2 2p^4$ ← sub-levels don't have to be filled; depends on # of valence e^- for that element

Orbital Arrangements with the Periodic Table



> The ***electron configuration*** is a listing of the electrons in that atom (or ion) in order of increasing energy.

① Find element in Periodic Table # note atomic number
② Start @ 'H' and work towards element
③ count exponents = e⁻ along the way.

Example 1 Silicon (Si) $1s^2\, 2s^2 2p^6\, 3s^2\, 3p^2$

Atomic #(Z) = 14

Example 2 Technetium (Tc)

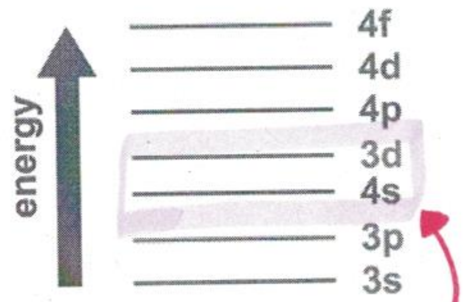
Atomic #(Z) = 43 $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^5$

15 P 38 Sr 48 Cd

Practice: Predict the electron configurations for *phosphorus, strontium, and cadmium.*

Atomic #	Symbol	Electron Configuration
15	P	$1s^2 \mid 2s^2 2p^6 \mid 3s^2 3p^3$
38	Sr	$1s^2 \mid 2s^2 2p^6 \mid 3s^2 3p^6 \mid 4s^2 3d^{10} 4p^6 \mid 5s^2$
48	Cd	$1s^2 \mid 2s^2 2p^6 \mid 3s^2 3p^6 4s^2 \mid 3d^{10} 4p^6 \mid 5s^2 4d^{10}$

Electron configuration of transition metals



Although the 3d sub-level is in a lower principal energy level than the 4s sub-level, it is actually higher in energy.

This means that the 4s sub-level is filled before the 3d sub-level.

Example: What is the electron structure of vanadium? V

1. Count number of electrons in atom 23
2. Fill sub-levels, remembering 4s is filled before 3d $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$

Electron configuration of Cr and Cu

The electron configurations of chromium and copper are **exceptions to the normal rules** of orbital filling:

Cr: chromium "promoted" e^-

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$

Cu: copper e^- Filling the 4 energy orbit makes it more stable.

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

d orbitals are more stable (i.e. have lower energy) when exactly half-filled (5) or filled (10).

For this reason, an electron may be promoted from the s to the d in certain cases. (to 1/2 fill or fill d-orbit)

With larger atoms like this it can be useful to shorten the electron arrangement. *(core notation)*

Copper can be shortened to **$[Ar]4s^1 3d^{10}$**.

last noble gas.

Electron configuration of ions

When writing the electron configuration of ions, it is important to ADD or subtract the appropriate number of electrons.

⊖ anions ADD electrons

⊕ cations REMOVE electrons

For non-transition metals, the sub-levels are then filled as for atoms.

(transition metal ions present exceptions)

Example: what is the electron structure of O^{2-}?

1. Count number of electrons in atom $8e^-$ add $2e^-$
2. Add or remove electrons due to charge $8 + 2 = 10$ electrons in ion
3. Fill sub-levels as for uncharged atom $1s^2 2s^2 2p^6$

removed from highest energy level ($n=4$ quantum #)

Electronic configuration of transition metal ions

When transition metals form ions, it is the 4s electrons that are removed BEFORE the **3d electrons**. *(just as how 4s electrons are filled before 3d e⁻)*

Example: what is the electron structure of Ni^{2+}? cation⊕ remove $2e^-$

1. Count number of electrons in atom 28
2. Fill sub-levels, remembering 4s is filled before 3d $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^8 = 28e^-$
3. Count number of electrons to be removed +2 charge = remove $2e^-$ 28-2 = 26 (ION)
4. Remove electrons starting with 4s $1s^2 2s^2 2p^6 3s^2 3p^6 3d^8$ ($4s^2$ removed) $= 26e^-$

Examples: Electron configuration of ions

1. Negative Ions (Anions): Add appropriate # of electrons to last subshell, starting with the configuration of the neutral atom.

Examples:

O^{2-} ⊖ add $2e^-$

O atom: $1s^2 2s^2 2p^4 + 2e^-$

O^{-2} ion: $1s^2 2s^2 2p^6$

isoelectronic to Ne ($10e^-$)

N^{3-} ⊖ add $3e^-$

N atom: $1s^2 2s^2 2p^3 + 3e^-$

N^{3-} ion: $1s^2 2s^2 2p^6$

isoelectronic to Ne ($10e^-$)

2. Positive Ions (Cations): Uh, oh… there are some rules!

   i) the **electrons in the outermost shell** (largest *n*-value) **are removed first.**
   ii) after that, **removal order is p before s before d** (within the outermost shell)

Examples:

Sn^{2+} ⊕ remove $2e^-$

Sn: [Kr] $5s^2 4d^{10} 5p^2$ from highest energy level

Sn^{2+} ion: [Kr] $5s^2 4d^{10}$

isoelectronic to Cd ($48e^-$)

"Core Notation" (electron config shorthand)

Fe^{3+} ⊕ remove 3 e⁻

Fe: [Ar] $4s^2 3d^6$

① remove from highest energy level FIRST - $2e^-$

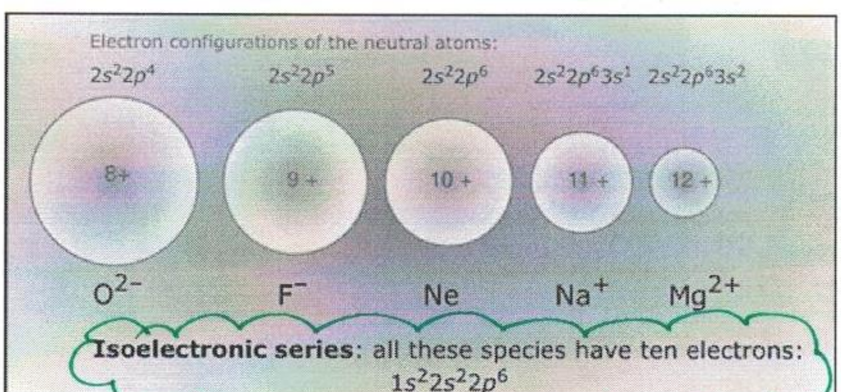
② remove the 3rd e⁻ from next highest energy level

Fe^{3+}: [Ar] $3d^5$

Isoelectronic

This word is an adjective meaning "has the same numbers of electrons".

For example, Na+ is isoelectronic to Ne since they both have 10 electrons.



Electronic configuration & Core Notation

After a noble gas, a New period begins in the periodic table and so too, a **new** Energy Level. We can **condense electron configurations** using core notation, in which the configuration of the previous NOBLE GAS is *represented by that noble gas symbol in* SQUARE BRACKETS then outer electrons are indicated in **bold** type.

Example:

inner core = full shells

valence e⁻ in outer "open shell"

Element	Full Electron Configuration	Core Notation
sodium	$1s^2 2s^2 2p^6 3s^1$	$[Ne] 3s^1$
silicon		$[Ne] 3s^2 3p^2$
argon		$[Ne] 3s^2 3p^6$
calcium		$[Ar] 4s^2$
scandium		$[Ar] 4s^2 3d^1$

Homework ***Assignment #4 Practicing Electron Cconfiguration & Core Notation***

Use the periodic table to complete the following table:

*remove e⁻ from HIGHEST energy level first!

Atom or Ion	Full Electron Configuration	Core Notation
Ge		$[Ar] 4s^2 3d^{10} 4p^2$
Zn^{2+} − 2e⁻		$[Ar] 3d^{10}$ (remove 2e⁻ from $4s^2$)
Sr		$[Kr] 5s^2$
Br^- + 1e⁻		$[Ar] 4s^2 3d^{10} 4p^6$ = [Kr]
Sn		$[Kr] 5s^2 3d^{10} 5p^2$
In^{3+} − 3e⁻		$[Kr] 4d^{10}$ (remove 2e⁻ from $5s^2$ and 1e⁻ from $5p^1$)

Use the periodic table to identify the neutral atoms having the following electron configurations:

Electron Configuration	Element Name
$[Ne] 3s^2$	Mg
$[Ar] 4s^2 3d^5$	Mn
$[Kr] 5s^2 4d^{10} 5p^3$	Sb
$[Xe] 6s^2 4f^7$ (lanthanide series)	Gd

Notice where each of these elements is located on the periodic table. Look at the highest energy sublevel being filled (**bold type**) in each of the atoms in the table, and identify the four different sections of the periodic table associated with each of these four sublevels.

Key on next page

Homework

Assignment #4
Electron Configuration Worksheet

right now...SKIP QUESTIONS #1 & 3...we will come back to these after we study "energy level diagrams"

Complete the following questions on a separate piece of paper.

1) Draw the energy level diagrams for the following atoms.
   a. C
   b. Chlorine
   c. Zn
   d. B
   e. Ca

2) Write the electron configurations for the following atoms.
   a. P
   b. Co
   c. Fe
   d. Osmium
   e. Nd
   f. Sr
   g. Beryllium
   h. Si

3) Draw the energy level diagrams for the following ions.
   a. Li^{+}
   b. C^{4-}
   c. Al^{3+}
   d. H^{+}
   e. Cl^{-}

4) Write the electron configurations for the following ions.
   a. N^{3-}
   b. Na^{+}
   c. Fluoride ion
   d. Mg^{2+}
   e. P^{3-}
   f. Iodide ion
   g. Tin (2^{+})
   h. Sulphide ion

5) Using core notation, write the electron configurations for the following atoms and ions.
   a. K
   b. O^{2-}
   c. Cr
   d. V
   e. Calcium
   f. Tellurium
   g. Xe
   h. Hg
   i. Cl^{-}
   j. Zn^{2+}

6) How many valence electrons are in each of the atoms/ions from #5?

Name: Key Block:______ Date:____________

Chemistry 11

Electron Configuration Worksheet Key

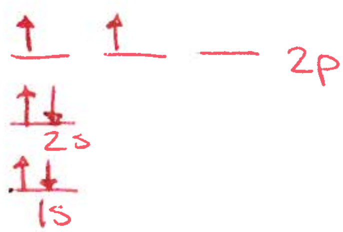
(46 marks)

Assignment

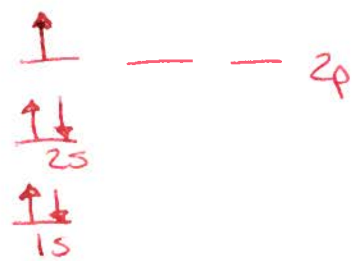
Complete the following questions on a separate piece of paper.

1) Draw the energy level diagrams for the following atoms. (5 marks)

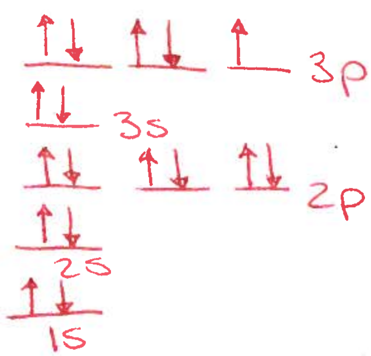
a. C = $6e^-$



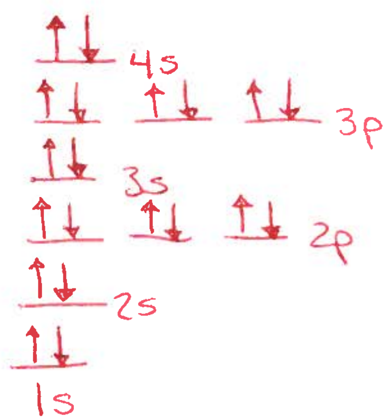
d. B = $5e^-$



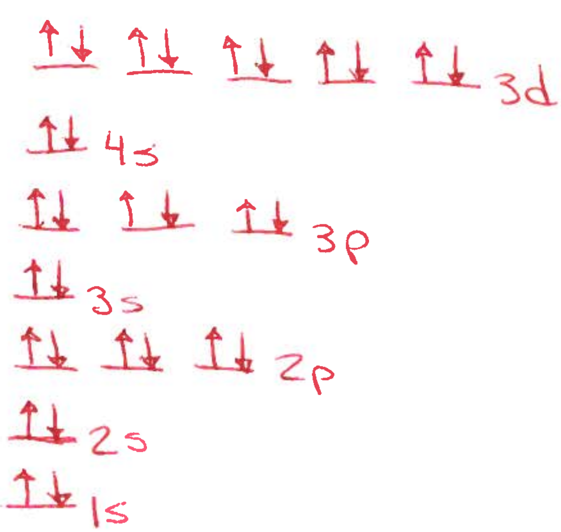
b. Chlorine = $17e^-$



e. Ca = $20e^-$



c. Zn



2) Write the electron configurations for the following atoms (8 marks).

a. P $1s^2 2s^2 2p^6 3s^2 3p^3$

b. Co $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^7$

c. Fe $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$

d. Osmium $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^6$

e. Nd $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^4$

f. Sr $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2$

g. Beryllium $1s^2 2s^2$

h. Si $1s^2 2s^2 2p^6 3s^2 3p^2$

3) Draw the energy level diagrams for the following ions. (5 marks)

a. Li^{+} = 2e⁻

↑↓ 1s

b. C^{4-} = 10e⁻

↑↓ ↑↓ ↑↓ 2p

↑↓ 2s

↑↓ 1s

c. Al^{3+} = 10e⁻

— 3s

↑↓ ↑↓ ↑↓ 2p

↑↓ 2s

↑↓ 1s

d. H^{+} = 0e⁻

e. Cl^- = $18e^-$

↑↓ ↑↓ ↑↓ 3p
↑↓ 3s
↑↓ ↑↓ ↑↓ 2p
↑↓ 2s
↑↓ 1s

4) Write the electron configurations for the following ions. (8 marks)

a. N^{3-} $1s^2 2s^2 2p^6$

b. Na^+ $1s^2 2s^2 2p^6$

c. Fluoride ion $1s^2 2s^2 2p^6$

d. Mg^{2+} $1s^2 2s^2 2p^6$

e. P^{3-} $1s^2 2s^2 2p^6 3s^2 3p^6$

f. Iodide ion $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6$

g. Tin (2^+) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10}$

h. Sulphide ion $1s^2 2s^2 2p^6 3s^2 3p^6$

5) Using core notation, write the electron configurations for the following atoms and ions. (10 marks)

a. K : $[Ar] 4s^1$

b. O^{2-} : $[He] 2s^2 2p^6$

* c. Cr : $[Ar] 4s^1 3d^5$

d. V : $[Ar]4s^2 3d^3$

e. Calcium : $[Ar]4s^2$

f. Tellurium : $[Kr]5s^2 4d^{10} 5p^4$

g. Xe : $[Kr]5s^2 4d^{10} 5p^6$

h. Hg : $[Xe]6s^2 4f^{14} 5d^{10}$

i. Cl^- : $[Ne]3s^2 3p^6$

j. Zn^{2+} : $[Ar]3d^{10}$

6) How many valence electrons are in each of the atoms/ions from #5? (10 marks)

a. K 1

b. O^{2-} 0

c. Cr 6

d. V 5

e. Calcium 2

f. Tellurium 6

g. Xe 0

h. Hg 2

i. Cl^- 0

j. Zn^{2+} 0

Electron orbitals

As Heisenberg suggested, it is impossible to exactly locate the position of an electron within an energy sub-level. (orbital)

But Schrödinger's equation showed that by measuring the electron density around the nucleus, it is possible to define regions where electrons are most likely to be found *at any one time.*

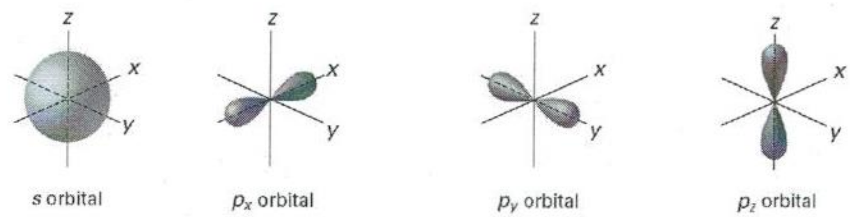
These regions are called orbitals.

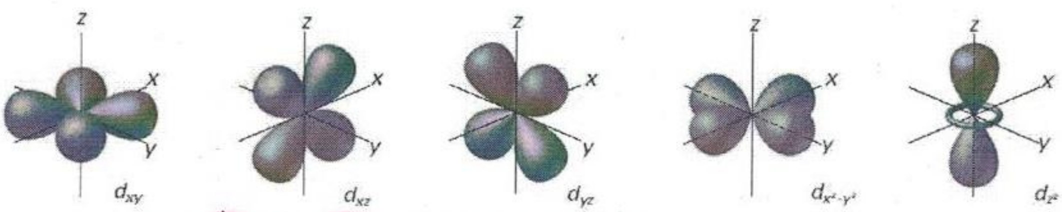
Each energy sub-level has one or more orbitals, each of which can contain a ***maximum*** of ____________________

Shapes of Atomic Orbitals (animation)

Different atomic orbitals are denoted by letters.

The **s orbitals** are circular, and **p orbitals** are dumb-bell-shaped.





Four of the five **d orbitals** have the same shape but different orientations in space.

The Pauli exclusion principle and spin

The Pauli Exclusion Principle states that each orbital may contain NO MORE than 2e⁻

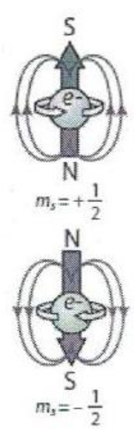
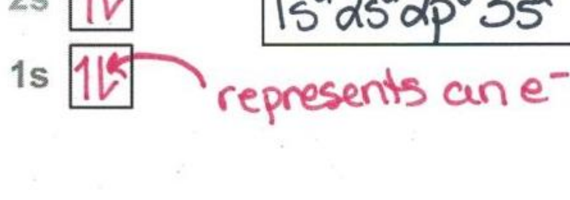


Figure 5.4.5 *The spin quantum number (m_s) identifies which possible spin an electron has.*

It also introduces a property of electrons called spin, which has two states: 'UP' and 'DOWN'.

The spins of electrons in the same orbital must be opposite ↑↓, *i.e. one 'up' and one 'down'.*

A energy level diagram shows how the orbitals are filled. Orbitals are represented by squares, and electrons by arrows pointing up or down.



spin diagram for magnesium,

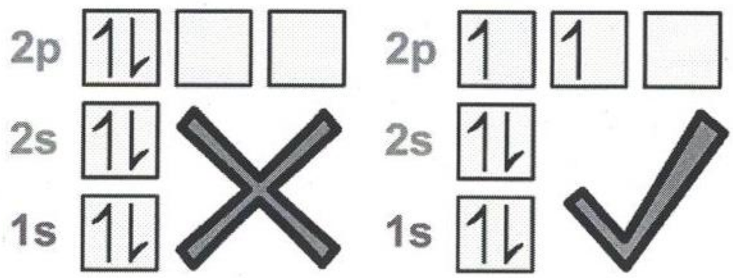
$1s^2 2s^2 2p^6 3s^2$

Rules for filling electrons

When two electrons occupy a p sub-level, they could either completely fill the same p orbital or half fill two different p orbitals.

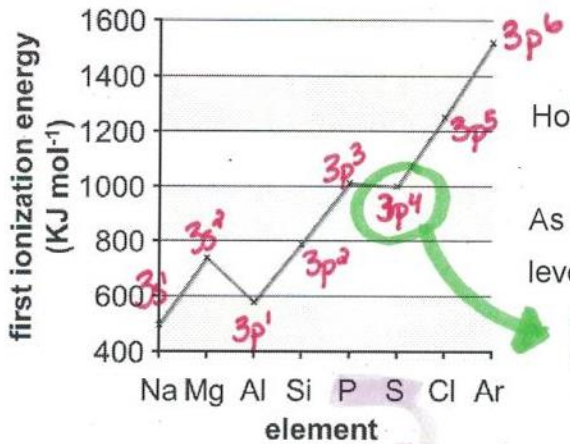
Hunds RULE states that single e^- occupy all empty orbitals within a sub-level before they start to form pairs in orbitals.

If two electrons enter the same orbital there is repulsion between them due to their negative charges. The most stable configuration is with single electrons in different orbitals.



Evidence for Hund's rule

.The first ionization energies for the elements in period 3 show a general increase



However, sulfur's value is below that of phosphorus.

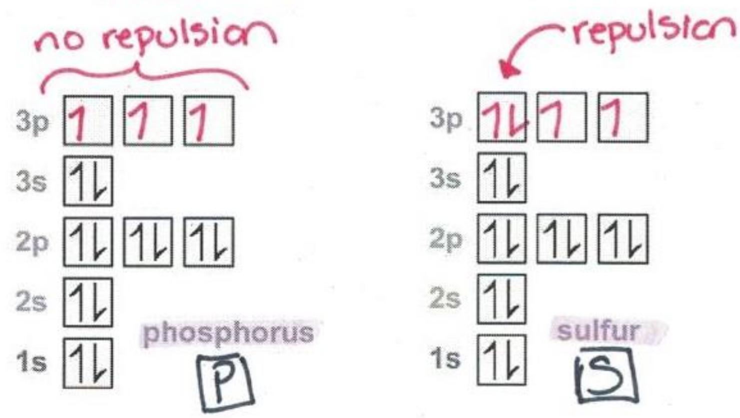
As the highest energy electrons of both are in the 3p sub-level this is evidence for **Hund's rule**.

ionization energy ↓ (easier to remove) b/c now e^- are sharing an orbital so they are repelling.

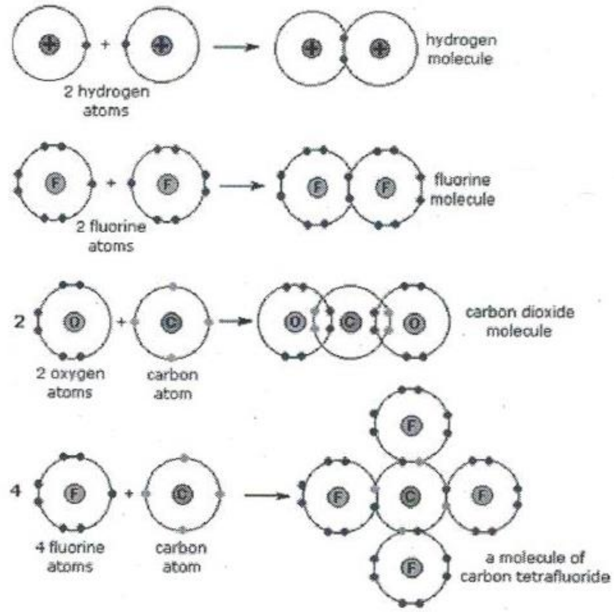
Evidence for Hund's rule: P vs. S

- Phosphorus has three electrons in its 3p sub-level and sulfur has four.

The lower first ionization energy for sulfur is because it has a pair of e^- in one of the 3p orbitals. Mutual repulsion between these two electrons makes it easier to remove 1.



What's the point of all this stuff?



Valence Electrons are the electrons found in the outermost open shell of an atom.

They are the ones that are able to take part in chemical reactions + bonding!.

***Electrons in the core**, or in **full d or f subshells**, are **never** valence electrons.*

Review:

- For any atom, the groundstate refers to the situation where the **electrons are in the lowest energy state possible**. For hydrogen, this is when the electron is in the 1s orbital
- When an electron is excited (absorbs energy) it will – *depending on the amount of energy absorbed* – jump up to any particular energy level above the ground state.
- In the hydrogen atom, all subshells in a given shell are degenerate (of equal energy). This is because there is only a single electron *(no other repulsive forces)*. {e.g. energy of 2s = 2p}
- In **multi-electron atoms**, many orbitals are occupied simultaneously – the electrons "sense" each other through **repulsive forces** and have to be optimally arranged. In such atoms, subshells have different energies within each shell. {e.g. energy of 2s 2p}
- The listing of the electrons in the atom from **lowest energy orbital to highest energy orbital** is known as the ***electron configuration*** of an atom.

Summary of Quantum Numbers

Every electron can be uniquely described using four quantum numbers:

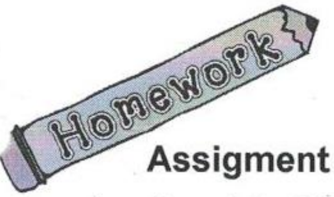
n = principal quantum number – indicates the energy level. Bigger n-value means further from the nucleus.

l = angular momentum quantum number – gives the shape of the orbital.

m = magnetic quantum number – gives orientation of the orbital in space.

s = spin quantum number – Tells you the orientation of the electron.

Quantum number	Name	What it labels	Possible values	Notes
n	principal	electron energy level or shell number	1, 2, 3, ...	Except for d-orbitals, the shell number matches the row of the periodic table.
ℓ	azimuthal	orbital type: s, p, d, f	0, 1, 2, ..., n-1	0 = s orbital 1 = p orbital 2 = d orbital 3 = f orbital
m_ℓ	magnetic	orbital sub-type	integers between and including $-\ell$ and $+\ell$: $-\ell, -\ell+1, \ldots \ell-1, \ell$	$\ell = 0$ (s): 2 e^- in one orbital $\ell = 1$ (p): 2 e^- in each of three sub orbitals (p_x, p_y, p_z) $\ell = 2$ (d): 2 e^- in each of 5 sub orbitals (d_{xy}, d_{xz}, d_{yz}, $d_{x^2-y^2}$, d_{z^2})
m_s	spin	electron spin	$\pm\frac{1}{2}$	Spins in any single sub-orbital must be paired.



answers in back of book.

Assigment #5: Hebden pg155 #26-27 (*every second letter*), #28 (*acegik*) , #29

Complete ALL assignments on a separate piece of paper and attach to your booklet when handing in at the end of the unit. Be sure to clearly number each assignment with a heading.

Electron configurations for the first 18 elements:

Key on next page

Element	*Z*	*1s*	*2s*	*2p*	*3s*	*3p*
H	1					
He	2					
Li	3					
Be	4					
B	5					
C	6					
N	7					
O	8					
F	9					
Ne	10					
Na	11					
Mg	12					
Al	13					
Si	14					
P	15					
S	16					
Cl	17					
Ar	18					

How does this pattern of organizing electrons relate to the periodic table?

Review

- For any atom, the ground state refers to the situation where the electrons are in the lowest energy state possible. For hydrogen, this is when the electron is in the 1s orbital
- When an electron is excited (absorbs energy) it will – *depending on the amount of energy absorbed* – jump up to any particular energy level above the ground state.
- In the hydrogen atom, all subshells in a given shell are degenerate (of equal energy). This is because there is only a single electron (no other repulsive forces). {e.g. energy of 2s = 2p}
- In multi-electron atoms, many orbitals are occupied simultaneously – the electrons "sense" each other through repulsive forces and have to be optimally arranged. In such atoms, subshells have different energies within each shell. {e.g. energy of 2s $\neq$ 2p}
- The listing of the electrons in the atom from lowest energy orbital to highest energy orbital is known as the ***electron configuration*** of an atom.

Electron configurations for the first 18 elements:

Element	***Z***	***1s***	***2s***	***2p***			***3s***	***3p***			
H	1	↑									
He	2	↑↓									1st energy level is full
Li	3	↑↓	↑								
Be	4	↑↓	↑↓								
B	5	↑↓	↑↓	↑							
C	6	↑↓	↑↓	↑	↑						
N	7	↑↓	↑↓	↑	↑	↑					
O	8	↑↓	↑↓	↑↓	↑	↑					
F	9	↑↓	↑↓	↑↓	↑↓	↑					
Ne	10	↑↓	↑↓	↑↓	↑↓	↑↓					2nd energy level is full
Na	11	↑↓	↑↓	↑↓	↑↓	↑↓	↑				
Mg	12	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓				
Al	13	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑			Don't start pairing e^-s until one is in every subshell
Si	14	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑		
P	15	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑	↑	
S	16	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑	
Cl	17	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	
Ar	18	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	3rd energy level is full

How does this pattern of organizing electrons relate to the periodic table?