Atomic Structure Class Notes

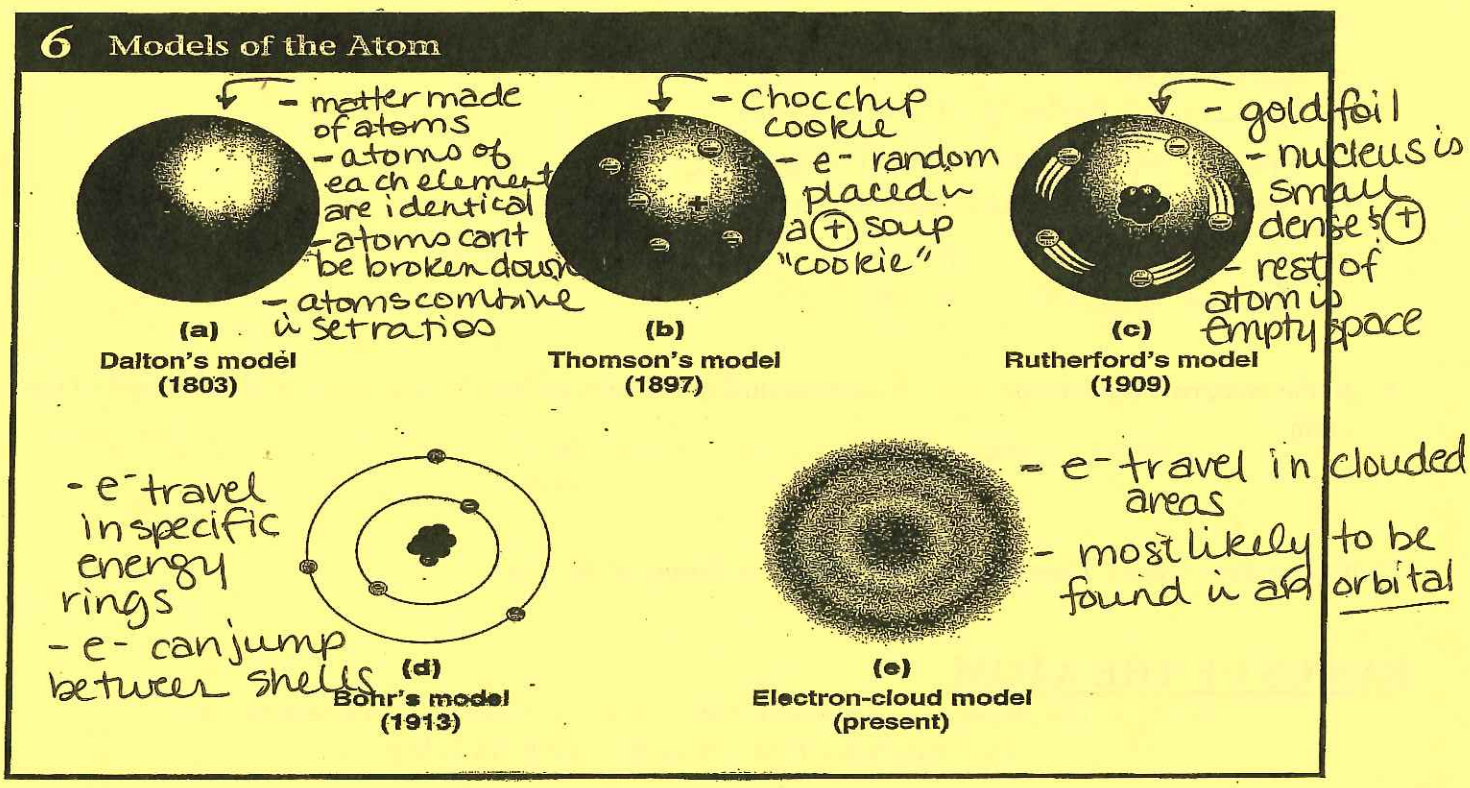
HISTORICAL BACKGROUND OF THE ATOM

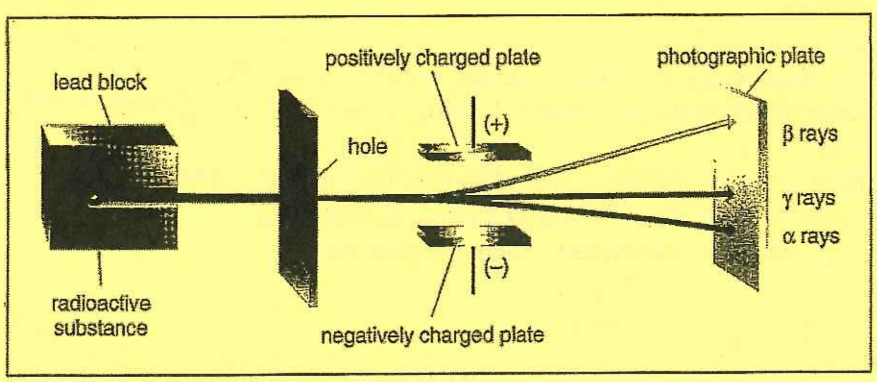
- History of Atomic Structure and Electrical Nature of the Atom and its particles including key scientists (Dalton, Thomson, Rutherford (Gold Foil Experiment), Bohr, etc..)
    - Key ideas:
        - Know contributions to history of the discovery of the structure of the atom
        - Detail flaws in predecessors theories of atomic structure
        - Understand out current structure of the atom

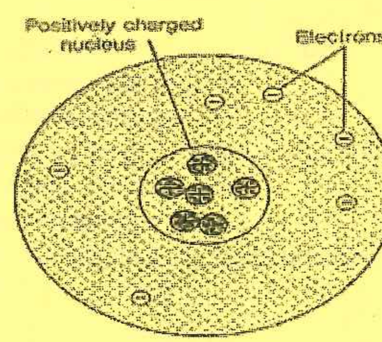
Atomic History Activity (cut + paste music & prezi)

Go to T & The Jhawks > chemistry resources find the <u>Atomic History Prezi</u> Link

View the prezi online take notes on the slide below







Model

Class Activity: *Atomic timeline ~ obtain materials at front of the classroom on teachers desk*

Questions:

1. List the components of ***Dalton's*** *Atomic Theory*
   - indivisible particles
   - smallest matter
   - ball of "stuff"
2. What hypothesis did ***Thomson*** develop based on his experiments?
   - plum pudding / choc chip cookie
   - discovered e- / randomly embedded in (+) model
3. Describe ***Rutherford's*** Gold Foil Experiment. How did this experiment change the view of the atom developed by Thomson?
   - gold foil → used Alpha particles
   - small dense (+) nucleus
   - most of atom empty space
4. Describe the changes that Niels ***Bohr*** did on the model of the atom.
   - e- travel in specific energy paths
   - Solar system model
5. In the ***modern*** atomic model what is the relationship between probability and electron cloud model of the atom? - most likely place to find an e- is w an ORBITAL

What do we know now? Take the Edmodo Quizlet ~ History of the Atom

BASICS OF THE ATOM

Go to T & The Jhawks > chemistry resources find the basics of atom structure Link
View the prezi online and supplement notes below

- Parts of the Atom

- Nucleus:
  Nucleons- all particles in the nucleus
  - ➢ Protons- positive charge, found in nucleus, mass 1 AMU = Atomic #
  - ➢ Neutrons- neutrally charge, found in nucleus, mass= 1 AMU = Atomic mass - At #
- Shells/ Principal Energy (quantum) Levels:
  - ➢ Electrons- negative charge, orbit around the nucleus, has very little mass (1/1836 of a proton or amu) = Atomic # in neutral atoms
- Overall charge of an atom = 0 (because protons (+) = electrons (-) therefore neutral)

Ex: C $12p^+$ Ex:

~~Ex~~ $12e^-$

- Nuclear charge- positively charged based on the # of protons in the nucleus

Ex: C $12p^{+}$

- Atomic number- # of protons and in a neutral atom also the # of electrons = overall charge of All atoms = 0

- Atomic mass/ Mass Number ~ # of protons + # of neutrons

$C \Rightarrow 6p^+ + 6n^0 = $ At mass $= 12$ AMU

- Calculating the # of neutrons use the formula: Atomic Mass – Atomic Number

C^{12}_{6} $12 - 6p^+ = 6n^0$

Kernel : (symbol of element) represents the nucleus and all electrons except the valence electrons (valence electrons go around the symbol in an electron dot diagram) includes p^+, n^0 (nucleus) & ALL NON valence e^-

1. How are the *atomic number* and the *number of protons* related to each other?

   At # identifies the # of p^+

2. How do the *number of protons, number of neutrons,* and the *mass number* relate to each other?

   mass # = p^+ PLUS n^0 in an atom

3. What determines the identity of an atom?

   At # = protons!

What do we know now? Take the Edmodo Quizlet basics of Atomic Structure

ISOTOPES ~ elements of same atom w/ same atomic # BUT different atomic mass due to diff # of n^0

2 types of Isotope notations

$^{12}_{6}$Carbon C-12 ← atomic mass $^{12}_{6}C$ 12 ← atomic mass, 6 ← atomic #

Fill in the chart using your knowledge of chemistry and Isotopes and average atomic masses.

isotopes of Uranium

Isotope Notation	Protons	Neutrons	Electrons
$^{238}_{92}U$	92	146	92
$^{23}_{11}Na$	11	12	11
$^{235}_{92}U$	92	143	92

Determining Average isotope mass

(Mass of isotope #1 x % of isotope #1) + (Mass of isotope #2 x % of isotope #2) + (Mass of isotope #3 x % of isotope #3) …

TRY THIS :

Problem #1: Boron has 2 isotopes, B-10 and B-11. The % abundance of B-10 is 19.78% and the % abundance for B-11 is 80.22%. What is the average atomic mass of boron?

Avg = 10(.1978) + 11(.8022)

Problem #2: Nitrogen

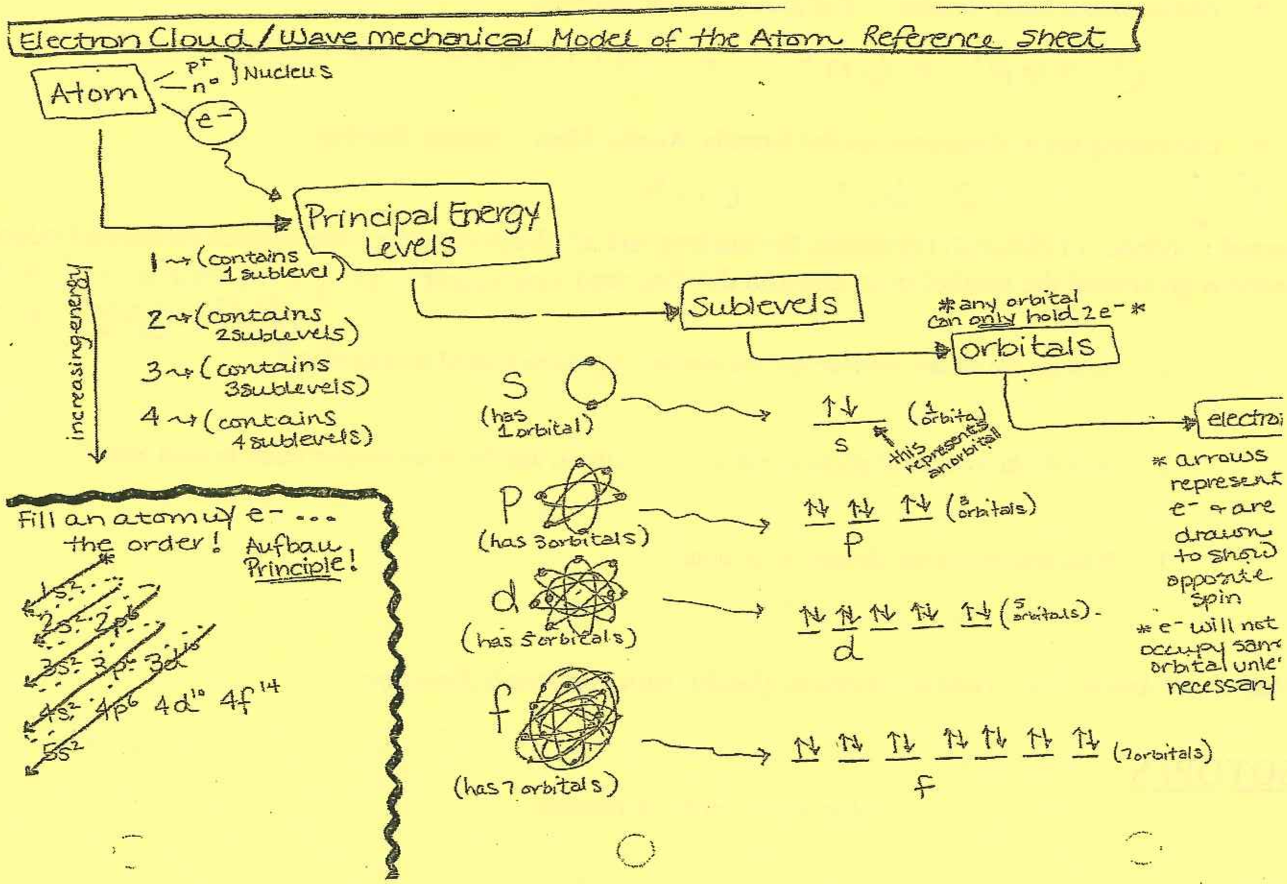
mass number	exact weight	percent abundance
14	14.003074	99.63
15	15.000108	0.37

Avg = 14.003074(.9963) + 15.000108(.0037)
13.95 + .0555

Problem #3: Chlorine		
mass number	exact weight	percent abundance
35	34.968852	75.77
37	36.965903	24.23

34.968852(.7577) + 36.965903(.2423)
26.49

 14.0



Arrangement of Electrons Outside the Nucleus

- **Electron Cloud Model/ Wave mechanical model**

Principal energy level (PEL)- All sublevels with the same coefficient (big number in front)

Big # (remember to include all in PEL)

Sublevels- Any letter after a coefficient (represented by **s,p,d,f**)-representing the area of probability of locating an electron.

Letter s, p, d, f

→ orbitals hold NOMORE 2e-

Orbital –a region in an atom of electron most probable location. (represented by dashed lines)

dashed Lines representing the orbitals in s, p, d, f

Electrons ~ charged particles that occupy orbitals in pairs with opposite spins (represented by up and down arrows) according to ***HUNDS RULE*** (e- single occupy orbitals BEFORE doubling up)

Coefficient represents **PEL→2 s** ^{3}← superscript represents the # of electrons

sublevel

How do we fill atoms with electrons?

- ***Electron Configuration:*** *refer to periodic Table in the CRT*

Example Al 2-8-3

Example N 2-5

Example Br 2-8-18-7

- **Valence electrons** valence electrons in an atom ~ electrons in the outermost shell
  - **octet rule:** stable valence electron configuration of eight electrons

Electron dot diagrams (Lewis Dot Diagram) ~ represent ONLY valence electrons surrounding a symbol (used for short hand in bonding)

$\ddot{\text{Al}}\cdot$	$\cdot\dot{\text{N}}\cdot$ (one dot below)	$\cdot\ddot{\text{Br}}\colon$ (two dots below)

- ***Extended Electron Configuration:*** (Representative e- configuration)

aufbau principle ~ electrons occupy the quantum state with lowest possible energy and electrons continue to fill in order from lowest to highest energy level.

$5s^2$ 9

$4s^2$ 6 $4p^6$ 8 $4d^{10}$ $4f^{14}$

$3s^2$ 4 $3p^6$ 5 $3d^{10}$ 7

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

$1s^2$ 1

* use 3 highlighters on per table to outline s block p block d block

Example Al (13e⁻) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$

Example N (7e-) $1s^2\ 2s^2\ 2p^3$

Example Br (35e⁻) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^5$

- ***Noble gas Electron Configuration:***

Example Al [Ne]-3
(2-8)

Example N [He]-5
(2)

Example Br [Ar]-10-7
2-8-8

- ***Orbital Notation of electrons***

Hunds rule ~ every orbital in a sub shell is singly occupied with electrons before any orbital is doubly occupied. all electrons in singly occupied orbitals have the same spin. (arrow faces same way)

Principal Energy Level	sublevels	orbitals	# of electrons
1	s ↑↓	1	2
2	s ↑↓ p ↑↓ ↑↓ ↑↓	1,3	2,6
3	s _ p _ _ _ d _ _ _ _ _	1,3,5	2,6,10
4	s _ p _ _ _ d _ _ _ _ _ f _ _ _ _ _ _ _	1,3,5,7	2,6,10,14

Example Al
$1s^2 2s^2 2p^6 3s^2 3p^1$
↑↓ (1s) ↑↓ (2s) ↑↓ ↑↓ ↑↓ (2p) ↑↓ (3s) ↑ _ _ (3p)

Example N
$1s^2 2s^2 2p^3$
↑↓ (1s) ↑↓ (2s) ↑ ↑ ↑ (2p)

Example Br
$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$

(↑↓ 1s) (↑↓ 2s) (↑↓ ↑↓ ↑↓ 2p) (↑↓ 3s) (↑↓ ↑↓ ↑↓ 3p) (↑↓ 4s) (↑↓ ↑↓ ↑↓ ↑↓ ↑↓ 3d) (↑↓ ↑↓ ↑ 4p)

Orbital Diagrams ~ tips & tricks

- Remember the sublevels ***S**mart **P**eople **D**on't **F**ail*
- Remember how to fill the orbitals of a particular sublevel; put an ↑ in each orbital before placing the second electron - up up up (↑↑↑)… then down down down (↓↓↓) …
- Look for the words…. *Half filled, occupied and completely filled*
- Electrons travel in energy levels around the nucleus.
- Electrons fill shells from lowest to highest energy.

Example#1: Phosphorous (P)

15p+ $\boxed{15e^-}$ 16n°

e- configuration

2-8-5

representative

$\boxed{\text{extended } e^- \text{ configuration}}$

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

$\boxed{\text{Orbital notation}}$

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

$\boxed{\text{Noble gas configuration}}$

$Ne - 3s^2 3p^3$

Example #2: Potassium (K)

19p+ $\boxed{19e^-}$ 20n°

e- configuration

2-8-8-2

$\boxed{\text{extended } e^- \text{ configuration}}$

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

*NOTE the skip due to PEL overlap

$\boxed{\text{orbital notation}}$

↑↓ 1s ↑↓ 2s ↑↓ ↑↓ ↑↓ 2p ↑↓ 3s ↑↓ ↑↓ ↑↓ 3p ↑↓ 4s

*Note the skip of 3d due to overlap!

$\boxed{\text{Noble gas configuration}}$

$Ar - 1s^2$

Electron Movement within the atom:

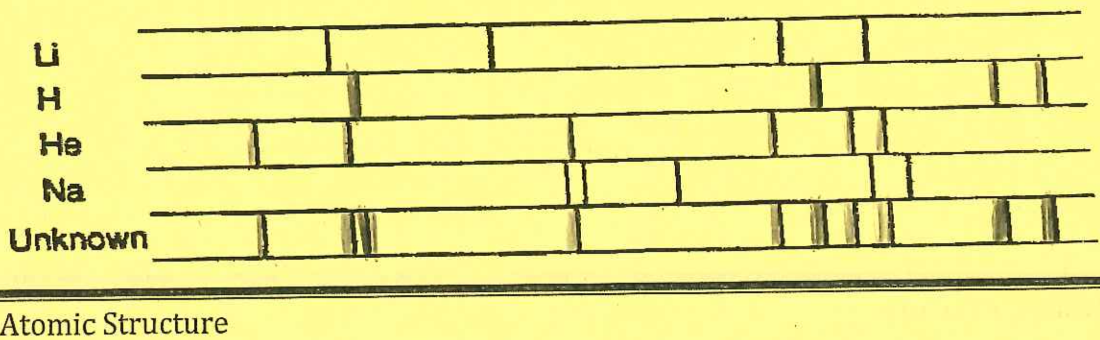
Quantum/Quanta ~ quantity of energy

Ground State: the lowest energy state for e-

Excited State: e- jump to higher energy

Spectral lines: are created when e- move from higher energy state back to ground state (lower energy state)

- **Absorption of energy causes electrons to jump to higher levels**
- **Emission of energy occurs when electrons fall back to their original location. Create colored light**



IONS

Ionization energy : amount of energy needed to remove the most loosely bound electron from a neutral atom.
Electronegativity: a measure of the attraction of a nucleus for valence electrons

- *Atomic Radius → size of any atom*
- *Ionic radii → size of any ion*
  - **Metal atoms** tend to lose electrons, become ***positive ions*** (cations) with a ***smaller radius*** than its atom (left side of steps on periodic table have low # of valence electrons…NOT close to 8!)
  - **Non-metals** tend to gain electrons, become ***negative ions*** (anions) with a ***larger radius*** than its atom (right side of steps on periodic table have high # of valence electrons…close to 8!)

Ions and Subatomic Particles

a) What is the name of the + charged ion cation
b) What is the name of the - charged ion anion
c) Metals tend to lose electrons & become + charged ions.
d) Non Metals tend to gain electrons & become − charged ions
e) Ions of metals are typically Smaller than their atoms
f) Ions of non metals are typically Larger than their atoms

e) *Complete the following table.*

Ion Symbol	Protons	Electrons	Ion Charge
i. O^{-2}	8	10	−2
ii. K^{+1}	19	18	+1
iii. Ba^{+2}	56	54	+2
iv. Fe^{3+}	26	23	+3
v. Fe^{2+}	26	24	+2
vi. F^{1-}	9	10	−1

isoelectronic - atoms & ions w/ same # of e- and same electron configurations.

pauli exclusion principle - e- enter orbitals w/ opposite spins (repel each other)