HISTORY OF THE ATOM

Atoms are the fundamental unit of elements. They are the smallest particle that retains the chemical identity of the element. You could say that atoms of the same element are the same, atoms of different elements are different. The atomic model has developed slowly over time through the discoveries of various scientists.

In 1803, a British school teacher, **John Dalton** expressed that matter could be explained in terms of atoms. Dalton proposed an "Atomic Theory of Matter".

Dalton's Atomic Theory of Matter

1. Each element is composed of extremely small particles called **atoms**.
2. All atoms of a given element are identical, but they differ from those of any other element.
3. Atoms are neither created or destroyed in any chemical reaction.
4. A given compound always has the same relative number and kinds of atoms.
   (Law of Constant Composition – A given compound always retains the same elements in the same proportions by mass.)

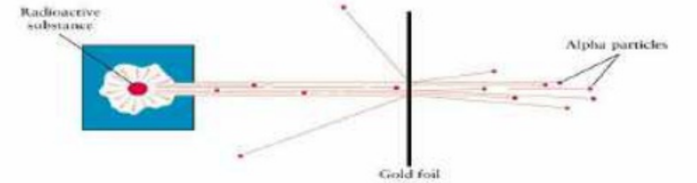
Dalton's theory was a great starting place but we have since learned that many of his postulates are incorrect:

1. Atoms have subatomic particles.
2. Elemental atoms differ: isotopes.
3. Atoms can change from one element to another during nuclear reactions.

Faraday is credited with proposing that atoms contained particles that have electrical charge. **J. J. Thompson** used his studies with cathode rays to prove the existence of charged subatomic particles. A cathode tube creates a cathode ray that is emitted from a negatively charged electrode (cathode) and travels to the positively charged electrode (anode). Thompson found that the stream could move a paddle wheel and could also be deflected by a magnet. Thompson named the negatively charged particles electrons.

Millikan discovered that electrons are extremely light particles. He charged oil droplets and allowed them to fall between charged plates, through this experiment he found that electrons are 2000 times lighter than an atom of hydrogen. (A hydrogen atom contains 1 neutron, 1 proton and 1 electron.

Rutherford studied radioactive materials: alpha (**α**) particles which have a 2+ charge, beta (**β**) particles which are high speed electrons and gamma (**δ**) radiation which is similar to x-rays and is not composed of particles. From his information Rutherford deduced that atoms were more complex than Dalton had thought. Rutherford used a stream of α particles and focused them on a thin piece of gold foil. (See illustration on next page.) He found that most of the particles passed through, however, some of the particles were deflected. From this he deduced that the atom must be made up of mostly space with a tight *nucleus* that has a positive charge, with electrons orbiting it as if it were a miniature solar system.



Bohr felt that Rutherford's model of an atom did not explain why an atom is stable. Bohr proposed that the electrons were in different energy levels orbiting at different levels from the nucleus. When an electron is at its lowest energy it is in the lowest energy level (orbit) – **ground state**. At times electrons can be given additional energy, when this happens electrons will jump to higher energy levels – **excited state**. When an atom is in the excited state it is unstable, it will not remain excited. It will eventually *emit* this energy and return to its ground state. Energy levels are numbered 1,2,3… and correspond to shells *k, l, m*…

ATOMIC STRUCTURE

Atoms are made up of protons, neutrons and electrons; it is now believed that there are additional particles inside these particles called quarks, gluons, mesons, muons and the list goes on. It has also been found that electrons do not have a defined orbit but are in an electron cloud.

Electrons carry a negative charge, as proved by Thompson, and protons carry a positive charge. However, the proton is not the only particle found in Rutherford's positive nucleus, neutrons are located in the nucleus also but are neutral. An individual atom is neutral; the number of protons and electrons are the same. The number of protons in an atom is denoted by the **atomic number**. The **atomic weight** is the average weight of all naturally occurring isotopes. Weight and mass are expressed in **atomic mass units** which is an arbitrary standard based on carbon-12; 1 amu = 1/12 the mass of a carbon-12 atom = 1.66×10^{-24}g. The **atomic mass** is the number of protons and neutrons (the atomic weight rounded to the nearest whole number.) Although the mass of a neutron is slightly larger than that of the proton we often round them both to 1 amu (atomic mass unit) and the electron to 0 amu.

Complete the following PNE chart using your knowledge of atomic structure and the periodic table.

Complete the following chart.

Element	Symbol	Atomic Number	Mass Number	No. of Protons	No. of Electrons	No. of Neutrons
Oxygen			16		8	
	Na	11				12
		14		14		14
Hydrogen		1	1	1		
		7		7		7

An **ion** is an atom that has lost or gained electrons. We describe ions based on the new net charge created by the change in electrons; for example when oxygen gains 2 additional electrons we now call it O^{2-}. This can be determined by the formula: charge = number of protons – number of electrons

Isotopes are atoms of the same element that have different numbers of neutrons. We describe ions using both the atomic weight and the number of neutrons. For carbon-12 we would write $^{12}_{6}C$ indicating that element 6 has 12 neutrons. For carbon-13 we would write $^{13}_{6}C$.

Complete the following chart based on your knowledge of atoms, ions and isotopes.

Isotope	Symbol	Atomic Number	Mass Number	Number of Neutrons
Aluminum-28				15
	$^{14}_{6}C$			
Chlorine-37		17		

To determine the average atomic mass of an element percent of each of its various isotopes need to be taken into account.

The formula for calculating ave. atomic mass = $\dfrac{(\text{mass}_a \times \%_a) + (\text{mass}_b \times \%_b) + (\text{mass}_c \times \%_c) + \ldots}{100}$

Calculate the ave. atomic mass of the isotopes of carbon: carbon-12 98.9% and carbon-13 1.1%

Name: ______________________________ **Block:** ___ **Date:** ______________

Homework: Atoms – History and Models

Completion: Fill in each blank with a completed term or short answer.

1. Atoms of each element are different from the atoms of all other elements.
2. The law of definite proportions states that a given compound always retains the same elements in the same proportions by mass.
3. Which scientist theorized that the atom was indivisible? Dalton
4. Electrons have a 1– charge, protons have a 1+ charge, and neutrons are neutral.
5. Which to particles are found in the nucleus? protons and neutrons
6. The atom's nucleus was discovered by Rutherford and his gold foil experiement.
7. Dalton was incorrect when he said that all atoms of the same element are the same. We know that atoms of the same element can be different. When the atoms have different amounts of electrons they are called ions.
8. When atoms of the same element have different numbers of neutrons they are called isotopes.
9. The number of protons in the nucleus of the atom is the atomic number for that element.
10. The number of protons and e⁻ must be equal for an atom to be neutral.
11. Thompson discovered the negatively charged particles and named them electrons.
12. What does "amu" stand for? atomic mass unit
13. Bohr proposed that electrons actually exist in discrete energy levels.
14. The lowest energy level that an electron can exist in is called ground state.
15. When an electron drops from the excited state energy is emitted.
16. List the relative masses of the electron, proton and neutron.
    Electron = 0 amu Proton = 1 amu Neutron = 1 amu
17. The ave atomic weight (mass) of an element is the weighted average of the masses of the isotopes of that element.
18. For the isotope $^{54}_{24}Cr$ the 24 represents the atomic # and 54 represents the mass #.
19. mass # is the atomic weight rounded to the nearest whole number.
20. The number of neutrons in the nucleus can be determined by subtracting the # protons (atomic #) from the atomic mass.

# of protons   same #   #p = #e if neutral   m# - p# = # of n

Complete the following chart.

Element	Symbol	Atomic Number	Mass Number	No. of Protons	No. of Electrons	No. of Neutrons
Carbon	C	6	12	6	6	6
potassium	K	19	40	19	19	21
magnesium	Mg	12	24	12	12	12
Helium	He	2	4	2	2	2
Boron	B	5	11	5	5	6

Complete the following chart.

Isotope mass	Symbol	Atomic Number	Mass Number	Number of Neutrons
Nitrogen-15	$^{15}_{7}N$	7	15	8
Neon-22	$^{22}_{10}Ne$	10	22	12
Beryllium-9	$^{9}_{4}Be$	4	9	5

Calculate the average atomic ~~weight~~ mass for each of the following.

1.
→ $^{54}_{24}Cr$ x 2.36%

→ $^{53}_{24}Cr$ 9.50%

$\boxed{\rightarrow {}^{52}_{24}Cr \quad 83.79\%}$

→ $^{50}_{24}Cr$ 4.35%

$$\frac{(54 \times 2.36\%) + (53 \times 9.50\%) + (52 \times 83.79\%) + (50 \times 4.35\%)}{100}$$

ave. atomic mass = $\boxed{52.06}$

2.
$^{20}_{10}Ne$ 90.51%

$^{21}_{10}Ne$ 0.27%

$^{22}_{10}Ne$ 9.22%

$$\frac{(20 \times 90.15\%) + (21 \times 0.27\%) + (22 \times 9.22\%)}{100}$$

ave. atomic mass = $\boxed{20.19}$

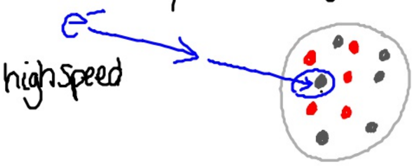
$$\text{ave. atomic mass} = \frac{(mass_a \times \%) + (mass_b \times \%) + (mass_c \times \%)}{100}$$

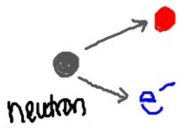
10/31/14 Types of Atomic Rxns

The nucleus is held together by a strong nuclear force. The neutrons also help buffer the repulsion of all the like charges of the protons. After element 83 there isn't enough neutrons or a strong enough force to hold the nucleus together + the elements are considered unstable.

Name	Identity	Charge	Pene. Ability	Written As
α	2P+2N	2+	Low	$\boxed{{}^{4}_{2}\alpha^{2+}}$
β	high speed e^-	1-	medium	$\boxed{{}^{0}_{-1}\beta^{1-}}$
γ	high energy (not a particle)	~	high	

Beta decay





splits into a new p^+ and e^-

Changes the ID of the element by 1 proton

Examples:

$${}^{226}_{88}Ra \longrightarrow {}^{4}_{2}\alpha^{2+} + {}^{222}_{86}Rn$$

(alpha particle lost)

$${}^{131}_{53}I \longrightarrow {}^{0}_{-1}\beta^{1-} + {}^{131}_{54}Xe$$

Practice

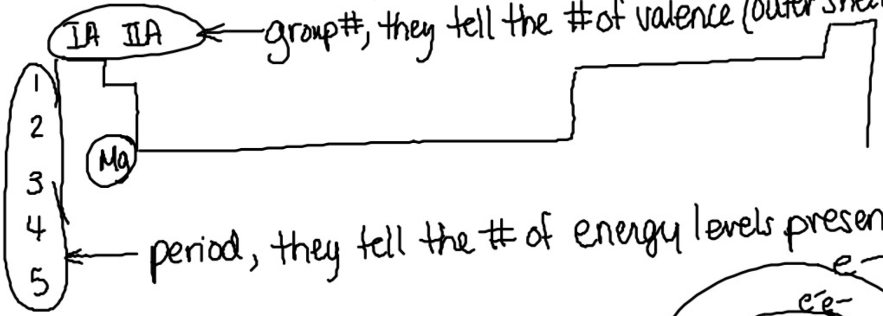
① $${}^{242}_{94}Pu \longrightarrow {}^{4}_{2}\alpha^{2+} + {}^{238}_{92}U$$

② $${}^{231}_{91}Pa \longrightarrow {}^{4}_{2}\alpha^{2+} + {}^{227}_{89}Ac$$

③ $${}^{223}_{87}Fr \longrightarrow {}^{0}_{-1}\beta^{1-} + {}^{223}_{88}Ra$$

GrpA Ions (want 0 or 8 v.e⁻)

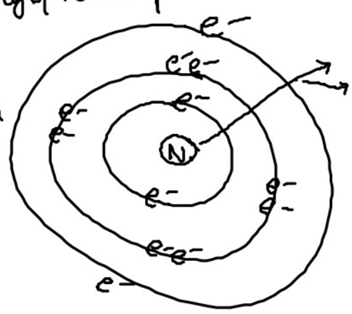
IA	IIA	B	IIIA	IVA	VA	VIA	VIIA	VIIIA
1vale⁻	2v.e⁻	SKIP	3ve⁻	4ve⁻	5ve⁻	6ve⁻	7ve⁻	8ve⁻
1+	2+		3+	4+ 4−	3−	2−	1−	☺



group#, they tell the # of valence (outer shell) e⁻

period, they tell the # of energy levels present

if Mg is in grp 2 and pd. 3 then



Mg-24

p = 12

N = 12

Phosphorous P^{3-}