Subatomic Particles

	Charge	Mass	Location
proton	+ 1	1 amu	nucleus
electron	– 1	1/1836	outside nucleus
neutron	0	1 amu	nucleus

(1 **a**tomic **m**ass **u**nit = 1/12 mass C-12 atom)

Aim: How do we find the "pen" for an atom of an element?

1) Referring to the key in the Periodic Table,

~~atomic~~ **mass** number → **12**.~~011~~ The atomic mass rounded off to the nearest whole is called the **mass number**.
(**#p^+ + #n^o**)

C

atomic number → <u>**6**</u>p^+, 6e^- In a **neutral atom**, the **#p^+ = # e^-**
(**#p^+**)

6 n^0

Again, { **mass number – atomic number = #n^o**}

Also, both protons & neutrons makeup the nucleus, but the total charge of the nucleus is due only to the #p^+ because neutrons have no charge. This is called the **nuclear charge.**

Therefore, the nuclear charge of a carbon atom is +6.

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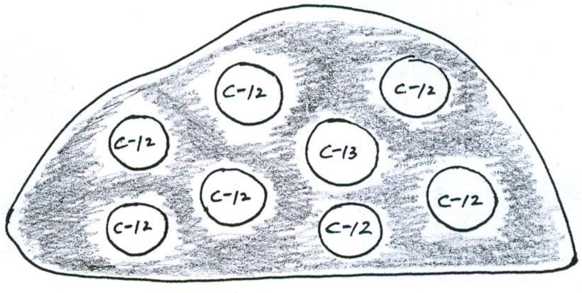
2) Practice finding "pen"

	7 6.~~941~~ Li **3**p^+ 3e^- ----------- 4n^o	**23** 22.~~98977~~ Na **11**p^+ 11e^- ------------ 12n^o	**52** 51.~~996~~ Cr **24**p^+ 24e^- ------------ 28n^o	**84** 83.~~80~~ Kr **36**p^+ 36e^- ----------- 48n^o
nuclear: charge	+3	+11	+24	+36

3) **Isotopes** are atoms of the **same** element that differ in the #n°.



Isotopes of Carbon

12	**13**
C	C
$6p^+$ $6e^-$	$6p^+$ $6e^-$
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6n°	7 n°

In other words, isotopes have the **same atomic number** but **different mass numbers**.

Another example; O-16 & O-18 are the isotopes of oxygen, Can you find **#p+, e^-, n°?**

Isotopes of Oxygen

16	**18**
O	O
$8p^+$ $8e^-$	$8p^+$ $8e^-$
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8n°	10n°

4) The **atomic mass** of an element is the **weighted average mass** of all of its isotopes.

$$\frac{12.00 + 13.00}{2} = 12.5 \neq 12.011$$

average weighted average

To calculate the atomic mass you must factor in the percent natural abundance of each isotope.

Isotope	Mass	Percent Natural Abundance
C-12	12.00	98.93%
C-13	13.00	1.07%

Just move the decimal point 2 places to the left, multiply and, then, add.

Weighted avg atomic mass = (12.00 x 0.9893) + (13.00 x 0.0107) = 12.01 amu

or

Weighted avg atomic mass = $\frac{(12.00 \times 98.93) + (13.00 \times 1.07)}{100}$ = 12.01 amu

Note: 1st multiply, and then divide the total by 100

Isotope	Mass	Percent Natural Abundance
B-10	10.01	19.91%
B-11	11.01	80.09%

atomic mass = (10.01 x 0.1991) + (11.01 x 0.8009) = 10.8109 → 10.81amu

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Last example, refer to backside of today's HW sheet .

24.305

Mg

$12p^{+}$, $12e^{-}$

$12n^{0}$

nuclear : +12
charge

Isotope	Mass	Percent Natural Abundance
Mg-24	23.99	78.99 %
Mg-25	24.99	10.00 %
Mg-26	25.98	11.01 %

a) In terms of *subatomic particles,* state one difference between these three isotopes of magnesium.

Isotopes differ in the number of neutrons.

b) Based on the isotopic masses and the natural abundances shown in the data table, in the space provided below, calculate the average atomic mass of magnesium.

weighted avg atomic mass = 23.99x0.7899 + 24.99x0.1000 + 25.98x0.1101 = 24.309099 → 24.31 amu

c) Based on the % natural abundances, why is the average atomic mass of neon closest to the mass of Mg-24?

Mg-24 has the highest percentage.