APES ~ ATOM REVIEW NOTES

I. Subatomic particles

A. *electron* (e^-) negatively charged subatomic particles
- fixed charge of -1
- very light mass (9.11×10^{-28} g)
- orbit the center

B. *proton* (p+) positively charged subatomic particle
(a hydrogen atom stripped of its electron is a "raw proton")
- fixed charge of +1
- same mass as a neutron (1.67×10^{-24} g)
- located in the center of an atom

C. *neutron* (n^0) neutral subatomic particle
- fixed charge of 0
- same mass as a proton (1.67×10^{-24} g)
- located in the nucleus

D. there are hundreds of subatomic particles, but these are the important ones to us

II. *Atomic number*—number of protons in the nucleus of an atom

A. characteristics
1) the atomic number is the unique I.D. number of an element
2) each element only has one atomic number

B. examples
What is the atomic number of the following elements?
O (8) I (53) Cl (17) Au (79)

C. *atomic neutrality*:
1) atoms are electrically neutral

number of protons = number of electrons in an atom

2) examples
How many electrons does Cu have? (29)
How many electrons does Rn have? (86)

III. *Mass number*

MASS NUMBER = PROTONS + NEUTRONS
OF NEUTRONS = MASS NUMBER - ATOMIC NUMBER

A. symbols can be written two ways:

mass number / SYMBOL / atomic number	SYMBOL—atomic number
$^{12}_{6}C$	C-12

B. mass number is the total mass of the nucleus

C. mass number is *not* the decimal number on the periodic table! (that's atomic mass)

D. example
How many p^+, n^0 and e^- are in an atom of S-34?
S = sulfur, which is #16. S has $\boxed{16\ p^+}$ and because # p^+ = # e^-, S has $\boxed{16\ e^-}$.
Mass number = 34. # n^0 = mass # - atomic # = 34-16 = $\boxed{18\ n^0}$.

IV. *Isotopes*

A. atoms of the same element that contain different numbers of neutrons

1) same number of p+
2) different mass numbers
3) different atomic masses
4) in nature, most elements occur as a mix of two or more isotopes

B. examples:

ISOTOPE	MASS #	ATOMIC #	p^+	n^0	e^-
O-16	16	8	8	8	8
O-17	17	8	8	9	8
O-18	18	8	8	10	8

Remember, # n^0 = mass number - atomic number.

V. *Atomic mass*

A. a weighted average based on mass and relative abundance of all naturally occurring isotopes of an element

ATOMIC MASS =
(MASS x RELATIVE ABUNDANCE) of natural isotope #1 +
(MASS x RELATIVE ABUNDANCE) of natural isotope #2 +
(MASS x RELATIVE ABUNDANCE) of natural isotope #3 … etc.

B. unit is amu = atomic mass unit

C. synthetic isotopes (made in lab, not found in nature) are not considered

D. example

Magnesium has three isotopes: Mg-24, Mg-25, and Mg-26:

ISOTOPE	ABUNDANCE	ATOMIC MASS
Mg-24	78.70%	23.985
Mg-25	10.13%	24.986
Mg-26	11.17%	25.983

Using the data, calculate the atomic mass of Mg.

ATOMIC MASS = (MASS x RELATIVE ABUNDANCE)

$(23.985)(0.7870) + (24.986)(0.1013) + (25.983)(0.1117) = \boxed{24.31 \text{ amu}}$