NOTES: Finding weights of the elements

1. Finding formulas

Dalton and others – experimented to find amount of elements in compounds

Law of conservation of mass: matter can neither be created nor destroyed.

- Analyzed copper carbonate and got
  - 1 atom of Copper,
  - 1 atom of Carbon and
  - 3 atoms of Oxygen

- Formula: $CuCO_3$.

- The same result every time no matter where $CuCO_3$ cam from!

- Conclusion: Law of definite proportions
  - *A compound contains the same elements in exactly the same proportions by mass regardless of the size of the sample or source of the compound.*

2. Finding weights of atoms

<u>Vocabulary</u>: Atomic Mass = Atomic Mass = Mass Number
*<u>Units</u>: g/mol or **amu** (atomic mass unit)*

- Atoms too small to weigh on scale

- Another method: Compare weights
  - Set mass of Hyrdogen (H) = **1 atomic mass unit (amu).**
  - (b/c H is smallest atom)

  - Compare everything else to H.
    - Example: O weighs 16 times more than H
    - ∴ O's mass is 16 amu
  - For many years, H used for comparison

  - 1800's O used for comparison b/c it combined with more elements

  - Isotopes confused things

- 1961 chemists and physicists agreed to use Carbon-12
- Now an **amu** = 1/12 of the weight of one Carbon-12 atom.
- One **amu** is less than one-billionth of a gram.

➢ **amu** ≈ mass of proton ≈ mass of neutron

Particle	Symbol	Relative electric charge	Mass number	Relative Mass (amu)	Actual mass (Kg)
Electron	e^-	-1	0	0.0005486	$9.109*10^{-31}$
Proton	p^+	+1	1	1.007276	$1.673*10^{-27}$
Neutron	n^0	0	1	1.008665	$1.675*10^{-27}$

➢ e^- too small to count in mass of atom

This is why we use ***<u>mass</u>*** *<u>number</u> to figure out number of protons and neutrons!*

3. Isotopes

Remember from last time:

➢ Before neutrons were discovered there were unsolved problems

1. Radioactivity: where is all that energy coming from?
2. Mass of atom: why doesn't the mass of electrons + protons = mass of atom?
3. Elements found with the same properties but different masses…how could that be?
   a. Soddy named these "isotopes," Greek for "same place" b/c they belonged in same spot on P.T. (periodic table)

➢ 1932 James Chadwick used radiation to eject a particle that had zero charge!

- He'd detected the neutron!

➢ Heisenberg said: nucleus = protons + neutrons

Example:
Carbon

- Z = 6
- A = 12

∴

- N has $6p^+$, $6e^-$, **6**n^0

- We call this C-12 or Carbon-12

Sometimes

- C has $6p^+$, $6e^-$, **8**n^0
- Z = 8
- A = 14
- We call this C-14 or Carbon-14

When we write the number after the element name we are specifying an isotope.

Definition of Isotope:

1. atoms of the same element that have different masses
2. one of two or more atoms having the same number of protons but different numbers of neutrons

- Mass number = number of p^+ and n^0.

Example:

- Mass number of C-12
  - $6p^+ + 6n^0 = 12$ amu
- Mass number of C-13
  - $6p^+ + 7n^0 = 13$ amu
- Mass number of C-14
  - $6p^+ + 8n^0 = 14$ amu

- How many neutrons does O-16 have?
  - How about:
    - O-17?
    - O-15?
  - How about:
    - U-235?
    - U-238?

4. Relative/Average Atomic Mass Number

- On P.T. (periodic table) mass number is decimal
- It is the **average** mass of an atom
- It is a **weighted average** b/c there is more of one isotope than another on earth

Analogy:

- Grades have three components
  - homework, 10%

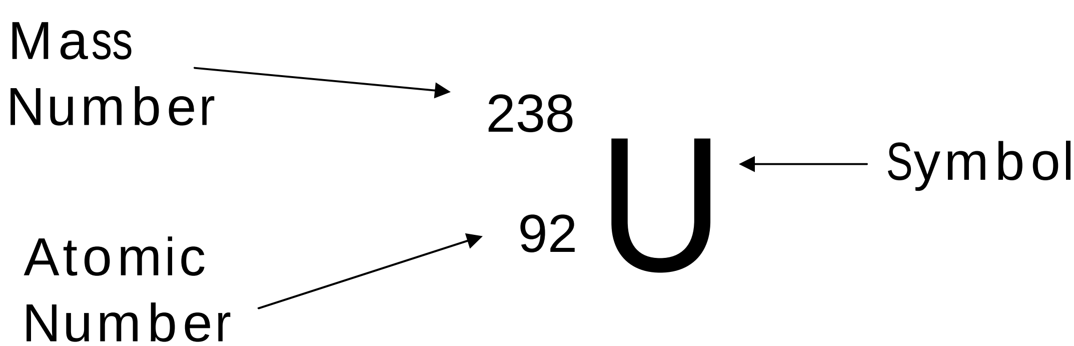
- labs, 30%
- Tests, 60%

- Oxygen has 3 isotopes
  - O-16, *Abundance* 99.76%,
  - O-17, *Abundance* 0.038%
  - O-18, *Abundance* 0.2%

- The effect of homework on your grade is less than the effect of tests on your grade because of the weighting

The effect of the mass of O-16 on the average mass of an oxygen atom is more than the effect of O-18 because there is more O-16 on earth.

Isotope Notation:



Try writing
- U-235,
- O-15,
- H,
- Deuterium (H w/ 1 n^0),
- Tritium (H w/ 2 n^0)