Atomic Structure & Atomic Mass Calculations Name: KEY Period:

An atom is made up of protons, neutrons and electrons.

- The atomic number is equal to the number of protons. The mass number is equal to the number of protons plus neutrons.
- In a neutral atom the number of protons are equal to the number of electrons.
- The charge on a ion indicates an imbalance between protons and electrons.
- Extra electrons produces a negative charge, the loss of electrons, a positive charge.

Example

Mass number → $^{15}_{7}N^{3-}$ ← charge

Atomic number

7 protons
8 neutrons (15-7)
10 electrons (7+3)

Complete the following chart.

protons — decimal from PT — whole # = p⁺ + n°

Element/ion	Atomic Number	Atomic mass	Mass number	protons +	neutrons	electrons −
H	1	1.008	1	1	0	1
H^+	1	1.008	1	1	0	0
$^{12}_{6}C$	6	12.01	12	6	6	6
$^{7}_{3}Li^+$	3	6.94	7	3	4	2
$^{35}_{17}Cl^-$	17	35.45	35	17	18	18
$^{39}_{19}K$	19	39.10	39	19	20	19
$^{24}_{12}Mg^{2+}$	12	24.31	24	12	12	10
As^{3-}	33	74.92	75	33	42	36
Ag	47	107.87	108	47	61	47
Ag^+	47	107.87	108	47	61	46
S^{2-}	16	32.07	32	16	16	18
U	92	238.03	238	92	146	92

Elements come in a variety of isotopes, meaning they are made up of atoms with the same atomic number but different atomic masses. These atoms differ in their number of neutrons. The average atomic mass is the weighted average of all the isotopes of the element.

Example: A sample of silver is 51.35% ^{107}Ag and 48.65% ^{108}Ag. What is the average atomic mass?

Answer:
0.5135 x 107 = 54.9445
0.4865 x 108 = 52.542
total = 107.4865 amu = average atomic mass
= 107.5 amu to sig fig.

Try out the practice problems on the back of this piece of paper.

Calculate the atomic mass for the following elements using the <u>weighted average method.</u> Circle your answer.

1. 69.17% ^{63}Cu and 30.83% ^{65}Cu.

   $\left.\begin{array}{l} 0.6917 \times 63 = 43.5771 \\ 0.3083 \times 65 = 20.0395 \end{array}\right\}$ 63.6166 amu = $\boxed{63.62\text{ amu}}$

2. 92.21 % ^{28}Si, 4.70% ^{29}Si, and 3.09% ^{30}Si.

   $\left.\begin{array}{l} 0.9221 \times 28 = 25.8188 \\ 0.0470 \times 29 = 1.363 \\ 0.0309 \times 30 = 0.927 \end{array}\right\}$ 28.1088 amu = $\boxed{28.1\text{ amu}}$

3. 7.42% ^{6}Li and 92.58% ^{7}Li.

   $\left.\begin{array}{l} 0.0742 \times 6 = 0.4452 \\ 0.9258 \times 7 = 6.4806 \end{array}\right\}$ 6.9258 amu = $\boxed{6.93\text{ amu}}$

4. 75.53% ^{35}Cl and 24.47% ^{37}Cl.

   $\left.\begin{array}{l} 0.7553 \times 35 = 26.4355 \\ 0.2447 \times 37 = 9.0539 \end{array}\right\}$ 35.4894 amu = $\boxed{35.49\text{ amu}}$

5. 4.35% ^{50}Cr, 83.79% ^{52}Cr, 9.50% ^{53}Cr and 2.36% ^{54}Cr.

   $\left.\begin{array}{l} 0.0435 \times 50 = 2.175 \\ 0.8379 \times 52 = 43.5708 \\ 0.0236 \times 54 = 1.2744 \\ 0.0950 \times 53 = 5.035 \end{array}\right\}$ 52.0552 amu = $\boxed{52.1\text{ amu}}$

Answer the following questions while reading pages 114-117 in your textbook.

1. What tool has been used since the 1920's to determine the tiny masses of individual atoms?

2. These small numbers are often inconvenient to work with. What is more useful? What is used as a standard for comparison

3. How is an **atomic mass unit** defined?

4. What does the mass of any one atom mainly depend on?

5. How do most elements occur in nature? ______________________________

6. What does the atomic mass of hydrogen (1.0079 amu) take into account?

7. How much of naturally occurring chlorine is chlorine-35? How much is chlorine-37? Why is the average closer to 35?

8. Define **atomic mass**.

9. Do practice problem 22.

10. Finish the sentence that begins, "To calculate the atomic mass ______________________________
______________________________". What is the resulting sum?

11. Answer the "Checkpoint" question at the bottom of page 116.

12. Do practice problem 24.