Chemistry H

Exercise 3.3a(H)
Isotopic Formulas & Atomic Weights

Name: ______________________

Date:________________ Per:_________

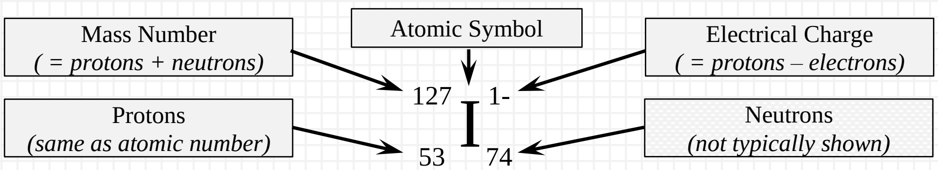
Atoms of an element are not necessarily identical to one another. While atoms of a specific element will always have a specific number of protons (which give the element its atomic number), they may vary in the number of electrons or number of neutrons they possess.

- When atoms of a particular element differ in the number of electrons they have (i.e., **ions**), it causes them to have different electrical charges (which are determined by comparing negative electrons to positive protons).
- When atoms of particular element differ in the number of neutrons they have (i.e., **isotopes**), it causes them to have different masses and different nuclear stability.

DIRECTIONS: Complete the following table.

Atomic Structure	Protons (p^+)	Neutrons (n^o)	Electrons (e^-)
Location in Atom			
Charge			
Mass (u)			

In order to show the exact makeup of a particular atom a complete atomic symbol may be written to describe its composition.



DIRECTIONS: Complete the following table using the information provided. Unless otherwise indicated, assume the atom is neutral.

Atom/Ion	Mass Number	Protons (p^+)	Neutrons (n^o)	Electrons (e^-)	Complete Symbol
Ca^{2+}	40	20	20	18	${}^{40}_{20}Ca^{2+}$
Cl^-	35	17	18	18	${}^{35}_{17}Cl^-$
C^{4+}	14	6	8	2	${}^{14}_{6}C^{4+}$
S^{2-}	32	16	16	18	${}^{32}_{16}S^{2-}$
Cs^+	133	55	78	54	${}^{133}_{55}Cs^+$
Zn^{2+}	65	30	35	28	${}^{65}_{30}Zn^{2+}$
B^{3+}	10	5	5	2	${}^{10}_{5}B^{3+}$
I^-	127	53	74	54	${}^{127}_{53}I^-$
H^-	3	1	2	2	${}^{3}_{1}H^-$
Fe^{3+}	56	26	30	23	${}^{56}_{26}Fe^{3+}$
N	14	7	7	7	${}^{14}_{7}N$
Na^+	23	11	12	10	${}^{23}_{11}Na^+$
H^+	2	1	1	0	${}^{2}_{1}H^+$
Kr	84	36	48	36	${}^{84}_{36}Kr$
H^+	1	1	0	0	${}^{1}_{1}H^+$

Answers: 2) 32.06 amu, sulfur 3) 207 amu 4) 1.0 amu 5) 48 amu 7) ${}^{151}Eu$ = 48%, ${}^{153}Eu$ = 52%

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Since elements are a blend of isotopes and each isotope has a unique mass, the atomic masses on the periodic table are **weighted averages** of all isotopes of each element. For example, three isotopes of hydrogen exist in nature (Hydrogen-1, Hydrogen-2, and Hydrogen-3). The mass of hydrogen shown on the periodic table is very close to 1*u* because almost all hydrogen is Hydrogen-1. Its high abundance has the most impact on the average mass of hydrogen.

$$\textbf{Average Atomic Mass } (A_r) = (\text{abundance}_a \times \text{mass}_a) + (\text{abundance}_b \times \text{mass}_b) + \ldots$$

Example: Chlorine has 2 naturally occurring isotopes. Chlorine–35 represents 75.78% of all chlorine and has a mass of 34.969*u*. Chlorine–37 represents 24.22% of all chlorine and has a mass of 36.966*u*.

$$\text{Mass}_{Cl} = (0.7578 \times 34.969u) + (0.2422 \times 36.966u) = 35.45u$$

DIRECTIONS: Complete the following in the space provided.

1. Boron exists in two isotopes, boron-10 and boron-11. Based on the atomic mass, which isotope should be more abundant?

2. Calculate the average atomic mass of an element if 95.00% of its atoms have a mass of 31.972*u*, 0.76% has a mass of 32.971*u* and 4.22% have a mass of 33.967*u*. What element is this?

$$\text{Average Atomic Mass} = (\text{abundance}_a \times \text{mass}_a) + (\text{abundance}_b \times \text{mass}_b)\ldots$$

$$A_r = (0.9500 \times 31.972u) + (0.0076 \times 32.971u) + (0.0422 \times 33.967u) = 32.0\underline{5}6u \Rightarrow \boxed{32.06u}$$

3. The four isotopes of lead are shown below, each with its percent by mass abundance and the composition of its nucleus. Using the following data, first calculate the approximate atomic mass (mass number) of each isotope. Then calculate the average atomic mass of lead.

	Isotope A	Isotope B	Isotope C	Isotope D
Protons	82	82	82	82
Neutrons	122	124	125	126
Mass	204	206	207	208
Abundance	1.37%	26.26%	20.82%	51.55%

$$A_r = (0.0137 \times 204u) + (0.2626 \times 206u) + (0.2082 \times 207u) + (0.5155 \times 208u) = 20\underline{7}.174u \Rightarrow \boxed{207u}$$

4. Hydrogen is 99% ^{1}H, 0.8% ^{2}H, and 0.2% ^{3}H. Calculate its average atomic mass.

$$A_r = (0.99 \times 1u) + (0.008 \times 2u) + (0.002 \times 3u) = 1.\underline{0}12u \Rightarrow \boxed{1.0u}$$

5. Titanium has five common isotopes: ^{46}Ti (8.0%), ^{47}Ti (7.8%), ^{48}Ti (73.4%), ^{49}Ti (5.5%), ^{50}Ti (5.3%). What is the average atomic mass of titanium?

$$A_r = (0.080 \times 46u) + (0.078 \times 47u) + (0.734 \times 48u) + (0.055 \times 49u) + (0.053 \times 50u) = 4\underline{7}.88u \Rightarrow \boxed{48u}$$

6. There are three isotopes of silicon. They have mass numbers of 28, 29 and 30. The average atomic mass of silicon is 28.086*u*. What does this say about the relative abundances of the three isotopes?

7. The element europium exists in nature as two isotopes: ^{151}Eu has a mass of 150.9196*u* and ^{153}Eu has a mass of 152.9209*u*. The average atomic mass of europium is 151.96. Calculate the relative abundance of the two europium isotopes.

$$\text{Average Atomic Mass} = (\text{abundance}_a \times \text{mass}_a) + (\text{abundance}_b \times \text{mass}_b)$$

The sum of abundances of the isotopes must be equal to 1, $\therefore$ $\textit{abundance}_a = 1 - \textit{abundance}_b$

$$\text{Average Atomic Mass} = (1 - \text{abundance}_b \times \text{mass}_a) + (\text{abundance}_b \times \text{mass}_b)$$

Answers: 2) 32.06 amu, sulfur 3) 207 amu 4) 1.0 amu 5) 48 amu 7) ^{151}Eu = 48%, ^{153}Eu = 52%

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$$151.96u = ((1-x) \times 150.9196u) + ((x) \times 152.9209u)$$

$$151.96u = 150.9196u - 150.9196ux + 152.9209ux$$
$$-150.9196u + 151.96u = 150.9196u - 150.9196ux + 152.9209ux - 150.9196u$$
$$1.0\underline{4}1u = 150.9196ux + 152.9209ux$$
$$1.0\underline{4}1u = 2.001\underline{3}ux$$
$$1.0\underline{4}1u/2.001\underline{3}u = 2.001\underline{3}ux/2.001\underline{3}u$$
$$0.52\underline{0}1 = x = \textit{abundance of } {}^{153}Eu\ (52\%)$$
$$1 - 0.52\underline{0}1 = 0.4799 = \textit{abundance of } {}^{151}Eu\ (48\%)$$

Answers: 2) 32.06 amu, sulfur 3) 207 amu 4) 1.0 amu 5) 48 amu 7) ^{151}Eu = 48%, ^{153}Eu = 52%