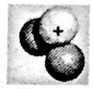
Unit 3: Atomic Theory & Structure

Section 3 - Distinguishing Among Atoms





Introduction

Just as apples come in different varieties, a chemical element can come in different "varieties" called isotopes.

Atomic Number

Atomic number of an element is the number of protons in the nucleus of each atom of that element.

It equals the number of electrons.

Element	Atomic #	# of p^+	# of e^-
Carbon	6	6	6
Phosphorus	15	15	15
Gold	79	79	79

Mass Number

Mass number is the number of protons AND neutrons in the nucleus of an atom.

$$\text{Mass \#} = p^+ + n^0$$

Name	p^+	n^0	e^-	Mass #
Oxygen - 18	8	10	8	18
Arsenic - 75	33	42	33	75
Phosphorus - 31	15	16	15	31

Isotopes

Isotopes are atoms of the same element having different masses due to varying numbers of neutrons.

Isotope	Symbol	Protons	Electrons	Neutrons	Nucleus
Hydrogen-1 (protium)	${}^{1}_{1}H$	1	1	0	
Hydrogen-2 (deuterium)	${}^{2}_{1}H$	1	1	1	
Hydrogen-3 (tritium)	${}^{3}_{1}H$	1	1	2	

Isotopes

Isotopes are atoms of the same element having different masses due to varying numbers of neutrons.

Isotope Names

Hydrogen - 1

Mass # ($p^+ + n^0$)

Isotope Symbols

${}^{1}_{1}H$

Mass # ($p^+ + n^0$)

Atomic # (p^+, e^-)

Mass of Atoms

- One atomic mass unit (amu) is defined as $1/12^{th}$ the mass of a carbon-12 atom.
- One amu is nearly, but not exactly, equal to one proton and one neutron.

Table 4.4 Masses of Subatomic Particles

Particle	Mass (amu)
Electron	0.000549
Proton	1.007276
Neutron	1.008665

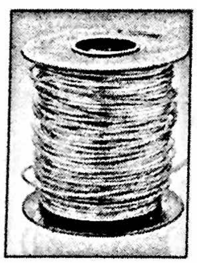
Atomic mass is the **weighted** average mass of all the naturally occurring isotopes of that element.

Isotope	Symbol	Composition of the nucleus	% in nature
Carbon-12	^{12}C	6 protons 6 neutrons	98.89%
Carbon-13	^{13}C	6 protons 7 neutrons	1.11%
Carbon-14	^{14}C	6 protons 8 neutrons	<0.01%

$\boxed{\text{Carbon} = 12.011}$

Using Atomic Mass to Determine the Relative Abundance of Isotopes

The atomic mass of copper is 63.546 amu. Which of copper's two isotopes is more abundant: copper-63 or copper-65?



Copper-63
The average is closer to 63 than it is to 65.

Using Atomic Mass to Determine the Relative Abundance of Isotopes

Boron has two isotopes: boron-10 and boron-11. Which is more abundant, given that the atomic mass of boron is 10.81?

boron-11
10.81 is closer to 11 than it is to 10

Steps to Calculate Atomic Mass HONORS

1) Divide % abundance by 100 to make it a decimal.
2) Multiply the mass of each isotope by the decimal abundance.
3) Add the products from part 2.

Example:

Isotope	Mass (amu)	% Abundance	Decimal abundance
Carbon-12	12.000	98.89	0.9889
Carbon-13	13.003	1.11	0.0111

(12.000)(0.9889) = 11.87
(13.003)(0.0111) = 0.144
$\boxed{12.01 \text{ amu}}$

Calculating Atomic Mass HONORS

Element X has two natural isotopes. The isotope with a mass of 10.012 amu (^{10}X) has a relative abundance of 19.91%. The isotope with a mass of 11.009 amu (^{11}X) has a relative abundance of 80.09%. Calculate the atomic mass of this element.

for ^{10}X: (10.012)(0.1991) = 1.993
For ^{11}X: (11.009)(0.8009) = 8.817
$\boxed{10.810 \text{ amu}}$

ATOMIC STRUCTURE

Name ____________________

An atom is made up of protons and neutrons (both found in the nucleus) and electrons (in the surrounding electron cloud). 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 equals the number of electrons. The charge on an ion indicates an imbalance between protons and electrons. Too many electrons produces a negative charge, too few, a positive charge.

This structure can be written as part of a chemical symbol.

Example:

mass number ↓ *charge* ↙

$^{15}_{7}N^{+3}$

↗ *atomic number*

7 protons

8 neutrons (15 – 7)

4 electrons

+ charge means the atom lost negative electrons

- charge means the atom has gained electrons

Complete the following chart.

Element/ Ion	Atomic Number	Atomic Mass (on the Periodic Table)	(Mass Number	= Protons	+ Neutrons)	Electrons
$^{1}_{1}H$	1	1.01 amu	1	1	0	1
$^{3}_{1}H^{+1}$	1	1.01 amu	3	1	2	0
$^{12}_{6}C$						
$^{7}_{3}Li^{+1}$						
$^{35}_{17}Cl^{-1}$						
$^{39}_{19}K$						
$^{24}_{12}Mg^{2+}$						
$^{80}_{33}As^{3-}$						
$^{107}_{47}Ag$						
$^{108}_{47}Ag^{+1}$						
$^{32}_{16}S^{-2}$						
$^{238}_{92}U$						

ISOTOPES AND AVERAGE ATOMIC MASS

Name ____________________

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 the number of neutrons.

The average atomic mass is the weighted average of all the isotopes of an element.

Example: A sample of cesium is 75% ^{133}Cs, 20% ^{132}Cs and 5% ^{134}Cs. What is its average atomic mass?

Answer:
$$.75 \times 133 = 99.75$$
$$.20 \times 132 = 26.4$$
$$.05 \times 134 = 6.7$$
$$\text{Total} = 132.85 \text{ amu} = \text{average atomic mass}$$

Determine the average atomic mass of the following mixtures of isotopes.

1. 80% ^{127}I, 17% ^{126}I, 3% ^{128}I

$(.80 \times 127) + (.17 \times 126) + (.03 \times 128) =$ ____________

2. 50% ^{197}Au, 50% ^{198}Au

____

3. 15% ^{55}Fe, 85% ^{56}Fe

____

4. 99% ^{1}H, 0.8% ^{2}H, 0.2% ^{3}H

____

5. 95% ^{14}N, 3% ^{15}N, 2% ^{16}N

____

6. 98% ^{12}C, 2% ^{14}C

____

Chemistry Worksheet NAME: ______________________________

Isotope Notation Block: _________

1. Uranium-235 and uranium-238 are considered isotopes of one another. How are uranium-235 similar, and how are they different?

2. Define isotope:

3. The number of protons in an atom is known as the ______________________________ of that atom.

4. The number of ______________________ determines the name of the atom.

5. The mass number of an atom is the number of ___________________________ plus the number of _________________________ in the nucleus of the atom.

6. The isotope notation for nitrogen-15 is as follows:
   a. The number 15 is the _______________________ number.
   b. The number 7 is the _______________________ number.
   c. How many neutrons does nitrogen-15 have? ______________

7. Write the following in isotope notation:

a. zinc-66: _________	d. helium-4: _________	g. silver-108: _________
b. chlorine-35: _________	e. uranium-235: _________	h. thorium-234: _________
c. plutonium-239: _________	f. potassium-40: _________	i. oxygen-16: _________

j. the atom with 12 protons and 12 neutrons _____________

k. the atom with 6 protons and 7 neutrons _____________

l. the atom with 79 protons and 117 neutrons. _____________

m. the phosphorus atom that has 17 neutrons. _____________

n. the copper atom that has 34 neutrons. _____________

o. the iodine atom that has 72 neutrons _____________

Chemistry Worksheet NAME: ______________________________

Isotope Notation Block: ________

Complete the table. Recall that A = mass number, Z = atomic number, and ions have charge!

Name	$^{A}_{Z}\mathrm{Sym}^{+/-}$	Atom or ion?	Atomic Number	Mass Number	# of n°	# of p^{+}	# of e^{-}	Metal, non-metal, or metalloid?
	^{24}Mg	Atom						
	^{30}Si	Atom						
	^{108}Pd	Atom						
		Atom		131		53		
		Atom	25		30			
	$^{35}S^{2-}$							
	$^{112}Cd^{2+}$							
					50	38	36	
			85		125		86	
	$^{4+}$			119	69			
				242		94	89	
		Atom			80	56		
	$^{52}_{24}\ ^{2+}$							

9. The data table is a chemist's record of data about six isotopes.

Isotope	# of Protons	# of Electrons	# of neutrons	Mass (amu)
1	24	24	26	49.946
2	24	24	28	51.941
3	26	26	30	55.999
4	24	24	29	52.941
5	24	24	30	53.939
6	26	26	31	56.969

a) Which of the isotopes listed are the same element? ____________________

How do you know? ____________________

b) Why aren't the masses of the isotopes whole numbers?

____

Upon further research, the chemist determined the percent abundance of each isotope. These are listed in the data table below.

Isotope	Percent Abundance
1	4.35
2	83.80
3	81.32
4	9.50
5	2.35
6	18.68

c) Assume that Isotope 1 is an isotope of element X and that all the isotopes of X are listed in the data table. Determine the atomic mass of X. Show all work.

d) Which isotope of X is most abundant? ____________________

e) Which isotope of X has the greatest effect on the mass of X? Explain why.

____

10. The element copper has naturally occurring isotopes with mass numbers of 63 and 65. The relative abundance and atomic masses are 69.2% for a mass of 62.93 amu and 30.8% for a mass of 64.93 amu. Calculate the average atomic mass of copper.

11. Calculate the average atomic mass of sulfur if 95.00% of all sulfur atoms have a mass of 31.972 amu, 0.76% has a mass of 32.971amu and 4.22% have a mass of 33.967 amu.

12. Calculate the average atomic mass of bromine. One isotope of bromine has an atomic mass of 78.92 amu and a relative abundance of 50.69%. The other major isotope of bromine has an atomic mass of 80.92 amu and a relative abundance of 49.31%.