How many atoms are there per gram of an element?

The Model: Average Atomic Mass

(Reference: Section 2.5 in *Silberberg 5th ed.*)

A single atom is extremely small. The typical atom will have a mass of approximately 3×10^{-23} g. The smallest mass that the standard analytical balance can weigh reliably is 0.0001 g, which corresponds to roughly 3 quintillion (*i.e.*, 3,000,000,000,000,000,000) atoms. Therefore we define

$$\mathbf{1\ amu = 1.6606 \times 10^{-24}\ g}$$

to make it convenient to discuss the mass of very small particles in terms of **atomic mass units** (amu) rather than very tiny fractions of what can be weighed out on a balance.

An overwhelming majority of the elements that are encountered in the chemistry lab have two or more naturally-occurring isotopes. If an element has more than one naturally-occurring isotope, then a random sample of the element should be assumed to exist as a mixture of these isotopes that are found in Nature.

Carbon has been found to be 98.89% ^{12}C and 1.11% ^{13}C. Carbon-12 (chosen by the scientific community to define the amu) has an **isotopic mass** of exactly 12 amu (*i.e.*, one ^{12}C atom weighs 12 amu) and that of ^{13}C is 13.0034 amu.

It is assumed that the composition of a sample of an element (in terms of the **percent natural abundances** of each of the element's isotopes) is the same everywhere on Earth. Therefore, in any sample containing carbon (be it a diamond or an organic compound containing carbon in addition to other elements), 98.89% of the carbon atoms in that sample will be C-12 (each weighing 12 amu) and the remaining 1.11% of the carbon atoms will be C-13 (each weighing 13.0034 amu), regardless of where the sample was taken from.

We now determine the average atomic mass of carbon in a manner that is organized and easy to follow (*i.e.*, by tabulating what we know and what we can derive from our given information):

Isotope	Isotopic Mass (amu)	Percent Natural Abundance	Mass Contribution (amu)
^{12}C	12.0000	98.89	$(12.0000 \times 0.9889 =)$ 11.87
^{13}C	13.0034	1.11	$(13.0034 \times 0.0111 =)$ 0.144
			Average Mass (amu) $(11.87 + 0.144 =)$ 12.01

Key Questions

1. If you were able to select one carbon atom at random, what is the mass of that atom most likely to be (in amu)? Why?

2. Yes or No (Circle your answer.): Does any carbon atom anywhere in the Universe have a mass equivalent to the average atomic mass of carbon on Earth? *Briefly explain your reasoning*.

3a. What is the mass in amu of ten sextillion (*i.e.*, 10^{22}) Carbon-12 atoms? *Show your work* and *circle your answer.*

b. What is the mass in grams of ten sextillion Carbon-12 atoms? *Show your work* and *circle your answer.*

c. What is the mass of ten sextillion Carbon-13 atoms (in g)? *Show your work* and *circle your answer.*

4. If a diamond consisting of ten sextillion "randomly-selected" carbon atoms (roughly a "one-karat diamond") were placed on an analytical balance, the balance would read (circle your answer and briefly explain your reasoning below):

   a.) slightly less than 0.1993 g
   b.) 0.1993 g
   c.) slightly more than 0.1993 g
   d.) slightly less than 0.2159 g
   e.) 0.2159 g
   f.) slightly more than 0.2159 g

   Explanation:

Exercises

5. Europium has two naturally-occurring isotopes. Use the data in the table below to determine the average atomic mass of europium to the nearest 0.1 amu. *Hint*: Complete the table similarly to what is shown for carbon in the Model.

Isotope	Isotopic Mass (amu)	Percent Natural Abundance	
^{151}Eu	150.92	47.8	
^{153}Eu	152.92	52.2	
		Average atomic mass (amu)	

6. 20.5% of germanium is Ge-70, which has an isotopic mass of 69.924 amu. 27.4% of germanium is Ge-72, which has an isotopic mass of 71.922 amu. 7.8% of germanium is Ge-73, which has an isotopic mass of 72.923 amu. 36.5% of germanium is Ge-74, which has an isotopic mass of 73.921 amu. The rest of naturally-occurring germanium is Ge-76, which has an isotopic mass of 75.921 amu. Use this data to calculate the average atomic mass of germanium to the nearest 0.01 amu. *Hint*: Complete the table below to help you organize the information given.

7. Copper has two naturally occurring isotopes, ^{63}Cu (isotopic mass 62.9396 amu) and ^{65}Cu (isotopic mass 64.9278 amu). If copper has an atomic mass of 63.546 amu, what is the percent abundance of each element*? Show your work* and *circle your answer*.

Key Questions

8a. Look at the Periodic Table. How do the numbers under the symbols for carbon, europium, and germanium compare to 12.01 amu and the values you determined in Exercises 5 and 6 respectively?

b. What is the number under the symbol of an element on the Periodic Table?

The Model: The Mole

(Reference: Section 3.1 in *Silberberg 5th ed.*)

1 dozen items = 12 items	1 gross of items = 144 items
1 score of items = 20 items	1 **mole** of items = 6.022×10^{23} items

Key Questions

9a. One often buys donuts by the dozen. How many donuts are there in a dozen donuts?

b. Abraham Lincoln started his Gettysburg Address with the words, "Four score and seven years ago . . ." How many years are there in a score of years?

c. Mardi Gras bead necklaces are very often sold "by the gross". How many necklaces of Mardi Gras beads are there in a gross of necklaces?

d. A mole of liquid water occupies a volume of approximately 18 mL. How many water molecules are in a mole of water molecules?

10a. 1 mol of items is 6.022×10^{23} of those items. 1 carbon atom weighs on average 12.01 amu. 1 amu is equivalent to 1.6606×10^{-24} g. *Use the Factor-Label Method* to calculate the mass (in grams) of a randomly-chosen sample of carbon containing 1 mole of carbon atoms. *Circle your answer*.

b. Look at the Periodic Table. How does the number under the symbol for carbon compare to the number of grams you determined in Question 10a?

c. What is the relationship between the units "atomic mass units" and "grams per mole"?