Chapter 3 Section 3
SECTION 3-3 OBJECTIVE 1 Explain what isotopes are.
Isotopes are atoms of the same elements that have different masses. The isotopes of a particular element all have the same number of protons and electrons but different numbers of neutrons.

SECTION 3-3 OBJECTIVE 2 Define atomic number and mass number, and describe how they apply to isotopes.
The Atomic Number \( Z \) of an element is the number of protons in the nucleus of each atom of that element.
The atomic number identifies an element and tells you how many protons it has, and how many electrons it has.
If you want to know which element has atomic number 47 look at the periodic table it is silver, Ag. All silver atoms contain 47 protons in their nuclei. Because atoms are neutral, we know from the atomic number that all silver atoms must also contain 47 electrons.

Most of the elements consist of mixtures of isotopes. Tin (Sn) has 10 stable isotopes, the most of any element.

Identifying an isotope requires knowing both the name (atomic number) of the element and the mass of the isotope.

Isotopes of Hydrogen
Protium, Deuterium, and Tritium
Protium is the most common type of hydrogen has a mass number of 1 (1 proton) and accounts for 99.985% of Hydrogen on earth
Deuterium (accounting for 0.015% of Hydrogen on earth) has a mass number of 2 (1 proton, 1 neutron). Tritium, which is radioactive, exists in very small amounts in nature, but it can be prepared artificially has a mass number of 3 (1 proton, 2 neutrons).

Mass Number: the total number of protons and neutrons in the nucleus of an isotope.
From the periodic table the Mass Number is the atomic mass rounded to the nearest whole number.

SECTION 3-3 OBJECTIVE 3 Given the identity of a nuclide; determine its number of protons, neutrons, and electrons.

Nuclide is a general term for any isotope of any element. The isotopes of hydrogen are unusual in that they have distinct names; Again Hydrogen is an exception!
There are two methods (ways) for specifying (identifying) isotopes.
1. Hyphen Notation, the mass number is written with a hyphen after the name of the element. For Example: uranium isotope 235 is therefore known as Uranium–235 or U–235

2. Nuclear Symbol, written Mass Number \( Z \) Atomic Number symbol Example: 235 \( _{92}^{\text{U}} \) as

The number of neutrons is found by subtracting the atomic number from the mass number.
Example:
mass number - atomic number = number of neutrons
235 (protons + neutrons) - 92 protons = 143 neutrons

Look on page 77 of your text sample problem 3-1
Relative atomic mass: Mass of atoms expressed in grams are very small.
One atomic mass unit, or 1 amu: is exactly 1/12 the mass of one carbon 12 atom
1 amu = 1.660 540 2 x 10\(^{-27}\) kg.
Average Atomic Masses of Elements: the weighed average of the atomic masses of the naturally occurring isotopes of the elements. Most elements occur naturally as mixtures of isotopes. Although isotopes have different masses, they do not differ significantly in their chemical behavior.
SECTION 3-3 OBJECTIVE 4 Define mole in terms of Avogadro’s number, and define molar mass.  
The mole is the SI unit for amount of a substance. One mole is equal to the number of atoms in exactly 12 grams of Carbon -12.  
**Avogadro’s number:** $6.0221376 \times 10^{23}$ the number of particles in exactly one mole of a pure substance.  
**Molar mass:** the mass of one mol of a pure substance. Molar mass is usually written in units of g/mol.

**SAMPLE PROBLEM 1:** How many protons, electrons, and neutrons are there in an atom of chlorine-37?  
**GIVEN:** name and mass of chlorine-37.  
**UNKNOWN:** numbers of protons electrons, and neutrons.  
**PLAN:**  
Atomic number = # of protons = # of electrons  
Mass number = # of neutrons + # of protons  

The mass number of Chlorine-37 is 37. From the periodic table we find that chlorine’s atomic number is 17.  

Mass # - atomic # = 37(protons + neutrons) - 17 protons = 20 neutrons.  
An atom of chlorine-37 contains 17 electrons, 17 protons, and 20 neutrons.

Recall Hyphen Notation: Uranium - 238  
Nuclear symbol $\quad \begin{array}{c} \mathbf{238} \\ \mathbf{92} \end{array}$ U

1. How many protons, electrons, and neutrons are in an atom of bromine-80?  
2. Write the nuclear symbol for carbon-13.  
3. Write the hyphen notation for the element that contains 15 electrons and 15 neutrons.

SECTION 3-3 OBJECTIVE 5 Solve problems involving mass in grams, amount in moles, and number of atoms of an element.  
Calculating **Average Atomic Mass**- (the weighted average of the atomic masses of the naturally occurring isotopes of an element) the average atomic mass of an element depends both on the mass and the relative abundance if each of the elements isotopes.

The atomic mass of any nuclide is determined by comparing it with the mass of the carbon-12 atom. The hydrogen-1 atom has an atomic mass of about 1/12 that of the carbon-12 atom, or about 1 amu. The precise value of the atomic mass of a hydrogen-1 atom is 1.007 825 amu. An oxygen-16 atom has about 16/12 (or 4/3) the mass of a carbon-12 atom. Careful measurements show the atomic mass of oxygen-16 to be 15.994 915 amu. The mass of a magnesium-24 atom is found to be slightly less than twice that of a carbon-12 atom. Its atomic mass is 23.985 042 amu.

The following is an example of how to calculate a weighted average.  
Suppose you have a box containing two sizes of marbles. If 25% of the marbles have masses of 2.00 g each and 75% have masses of 3.00 g each, how is the weighted average calculated?  
You could count the marbles, calculate the total mass of the mixture, and divide by the total number of marbles. If you had 100 marbles, the calculations would be as follows.  
25 marbles x 2.00 g = 50 g  
75 marbles x 3.00 g = 225 g  
Adding these masses gives the total mass of the marbles.  
50 g + 225 g = 275 g

However, that is not the average mass, to get that we would then need to divide by 100 to get the average mass per marble for this mixture.  
Dividing the total mass by 100 gives an average marble mass of 2.75 g.  
A simpler method is to multiply the mass of each marble by the decimal fraction representing its percentage in the mixture. Then add the products (results).  
$25\% = 0.25 \quad 75\% = 0.75$  
$(2.00 \text{ g} \times 0.25) + (3.00 \text{ g} \times 0.75) = 2.75 \text{ g}$
The average atomic mass of an element depends on both the mass and the relative abundance of each of the element's isotopes. Naturally occurring copper consists of 69.17% copper-63, which has an atomic mass of 62.929 599 amu, and 30.83% copper-65, which has an atomic mass of 64.927 793 amu. The average atomic mass of copper can be calculated by multiplying the atomic mass of each isotope by its relative abundance (expressed in decimal form) and adding the results.

\[
0.6917 \times 62.929\text{ 599 amu} + 0.3083 \times 64.927\text{ 793 amu} = 63.55\text{ amu}
\]

The calculated average atomic mass of naturally occurring copper is 63.55 amu. In your textbook, an element's atomic mass is usually rounded to two decimal places before it is used in a calculation. We will do the same – Round the numbers to 2 places after the decimal for atomic mass.

Sample problem 3-2
What is the mass in grams of 3.50 mol of the element Copper?
GIVEN: 3.50 mol Cu
UNKNOWN: mass of Cu in grams

**PLAN:**
Cu in moles X mass of Cu (grams)
Moles Cu x grams Cu/ mol Cu = grams Cu
3.50 mol Cu x 63.55 g Cu/ mol Cu = 222 g Cu

Where did 63.55 g Cu/ mol come from? From the periodic table we find that Cu has an average atomic mass of 63.55 amu which by definition translate to 63.55 g Cu/ mol. So to find the atomic mass from the periodic table is to find the g/ mol for that element. This is used as a conversion factor for g to mol or mol to g calculations.

Sample problem 3-3
A chemist produced 11.9 g of aluminum, Al. How many moles of aluminum were produced?
GIVEN: 11.9g Al
UNKNOWN: amount of Al in moles

Grams Al x moles Al/ grams Al = moles Al
(11.9 grams Al)(1 mol Al / 26.98 g Al) = 0.441 mol Al

This answer is correctly given to three (3) significant figures!

**Sample Problem 3-4**
How many moles of silver are in 3.01 x 10e23 atoms of silver?
GIVEN: 3.01 x 10^{23} atoms of Ag
UNKNOWN: amount of Ag in moles

**PLAN:**
Number of atoms of Ag to amount of Ag in moles:
Ag atoms*1 mol Ag/Avogadro's number of Ag atoms = moles Ag
\[
(3.01\times10^{23}\text{  Ag atoms}) \frac{1\text{ mol Ag}}{6.02\times10^{23}\text{  Ag Atoms}} = 0.500\text{ mol Ag}
\]
Sample Problem 3-5: What is the mass in grams of $1.20 \times 10^8$ atoms of Copper?

**GIVEN:** $1.20 \times 10^8$ atoms of Cu  
**UNKNOWN:** mass of Cu in grams  
**PLAN:**  
Convert (Number of atoms of Cu) to (amount of Cu in moles) then to (mass of Cu in grams)  
(# of Cu atoms) times (1 mole Cu divided by Avogadro’s # of Cu atoms) times the (molar mass of Cu; by definition the grams Cu per mol Cu) to find the mass of Cu in grams.  

$$\frac{\text{# of Cu atoms}}{\text{Avogadro’s # of Cu atoms}} \times \frac{1 \text{ mole Cu}}{\text{1 mol Cu}} \times \frac{\text{grams Cu}}{1 \text{ mole Cu}} = \text{mass of Cu}$$

$$\frac{1.20 \times 10^8 \text{ Cu atoms}}{6.02 \times 10^{23} \text{ Cu atoms}} \times \frac{1 \text{ mol Cu}}{63.55 \text{ g Cu}} = 1.27 \times 10^{-14} \text{ g Cu}$$

Sample Problem 3-6: A certain element exists as three natural isotopes as shown in the following table:

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Mass (AMU)</th>
<th>% natural abundance</th>
<th>Mass number</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>19.99244</td>
<td>90.51</td>
<td>20</td>
</tr>
<tr>
<td>2</td>
<td>20.99395</td>
<td>0.27</td>
<td>21</td>
</tr>
<tr>
<td>3</td>
<td>21.99138</td>
<td>9.22</td>
<td>22</td>
</tr>
</tbody>
</table>

Calculate the average atomic mass of this element.  
**GIVEN:** See Table Above  
**UNKNOWN:** average atomic mass  
**Plan:**  
$$\text{(Percent of Isotope 1 x Mass Isotope 1) + (Percent of Isotope 2 x Mass Isotope 2) + (Percent of Isotope 3 x Mass Isotope 3) = Average Atomic Mass Of This Element}$$

$$\left(0.9051 \times 19.99244 \text{ amu}\right) + \left(0.0027 \times 20.99395 \text{ amu}\right) + \left(0.0922 \times 21.99138 \text{ amu}\right) = 20.17945 \text{ amu}$$