Honors Chemistry Chapter 3 Atoms the Building Blocks of Matter

Section 1: The Atom: From Philosophical to Theory, pgs. 67-71
Objectives:
1. **Explain the law of conservation of mass, the law of definite proportions, and the law of multiple proportions.**
2. **Summarize the five essential of Dalton's atomic theory.**
3. **Explain the relationship between Dalton's atomic theory and the law of conservation of mass, the law of definite proportions, and the law of multiple proportions.**

Vocabulary: Define the following.
1. **law of conservation of mass--**

2. **law of definite proportions--**

3. **law of multiple proportions--**

The term atom comes from the Greek philosopher Democritus who coined the word "atomos" meaning indivisible.
In wasn't until the 1700s that the definition of an element as a substance that could not be broken down by ordinary chemical means was finally accepted. By the end of that century an emphasis in all the sciences turned from qualitative descriptions of the physical world to studies that emphasized a quantitative approach to understanding nature. This time period was an age of scientific enlightenment from which modern science has evolved.
The 3 laws of basic chemistry and physics that emerged from this time period are:
1. **The law of conservation of mass As defined in the vocabulary above, the mass of substances before and after a chemical reaction must be the same. In other words the mass of the substances that react must equal the mass of the substances produced.**
2. **The law of definite proportions. This means that compounds produced in a chemical reaction must contain the same elements in the same proportions as measured by mass. For example no matter what the size of a sample of water the ratio of the mass of hydrogen to mass of oxygen is always 1 to 8 in water.**
3. **The law of multiple proportions. It was discovered that sometimes the same elements can form more than one type of compound. For instance hydrogen and oxygen can form water, H_2O and can also form hydrogen peroxide, H_2O_2. If two or more**

compounds are formed from these elements than the ratio of the mass of these elements must be in a ratio of small whole numbers. In the above example the early scientists found that if samples of water and hydrogen peroxide both contain 2.0 grams of hydrogen, and if the mass of oxygen in water was 16.0 grams then the mass of oxygen in hydrogen peroxide was 32.0 grams. This is an exact ratio of 2:1. It was these early calculations based on mass only that eventually led chemists to find the exact formula for water, H_2O and hydrogen peroxide, H_2O_2.

Dalton's Atomic Theory

In 1809 John Dalton took these three laws and expanded them into the atomic theory which has stood the test of time except for minor alterations as new facts were discovered.

Summary of the 5 statements of Dalton's atomic theory.

1. All matter is composed of atoms.
2. Atoms of a given element are identical in size, mass, and other properties. Latter revised in the modern atomic theory with the discovery of isotopes.
3. Atoms cannot be subdivided, created, or destroyed. Today we know that atoms can be subdivided into smaller particles called subatomic particles, (electrons etc.), and we now know that mass can be changed into energy in nuclear reactions.
4. Atoms combine in simple whole-number ratios forming new chemical compounds.
5. In chemical reactions, atoms are rearranged.

The Modern Atomic Theory

The modern atomic theory incorporates most of Dalton's theory. It has been modified to include isotopes, subatomic particles, and the loss of mass in nuclear reactions. However, the concepts of all matter being composed of atoms, and that atoms from one element differ from atoms of another element, as well as the conservation of mass in chemical reactions remains.

Three compounds containing potassium and oxygen are compared. Analysis shows that for each 1.00g of O, the compounds have 1.22g, 2.44g, and 4.89g of K respectively. How does this data support the law of multiple proportions?

Answer:

A chemical reaction combines 2.5 grams of one substance with 3.2 grams of another substance to produce 5.7 grams of a product. What law does this support and why?

(2)Answers:

Section 2 The Structure of the Atom pgs. 72-76.

Objectives:

1. **Summarize the observed properties of cathode rays that led to the discovery of the electron.**
2. **Summarize the experiment carried out by Rutherford that led to the discovery of the nucleus.**
3. **List the properties of protons, neutrons, and electrons.**
4. **Define atom.**

Vocabulary: Define the following.

1. **atom--**

2. **nuclear forces--**

Discovery of the Atom

The atom is composed of three subatomic particles called protons, neutrons, and electrons. The atom is divided into 2 regions the nucleus and the electron cloud.
The electron was discovered first by using a cathode ray tube, which is a tube containing a gas at low pressure. The negative end of the tube is called the cathode, and it is from this end that electrons will enter the tube and pass through the tube to the positive end called the anode.
Studies showed that cathode rays (streams of electrons) were deflected by a magnetic field and that they were deflected away from a negatively charged object. The physicist J.J. Thomson is credited with calling these particles of the cathode ray electrons.
Another physicist Robert A. Millikan discovered the mass of the electron.
These experiments resulted in the understanding that atoms are electrically neutral, contain a positive charge to offset the negative charge of the electrons as well as the prediction that there must be other more massive subatomic particles since the electron has such little mass.
J.J. Thomson developed a model (long since disproved) called the "plum pudding model" of atomic structure.
Another scientist, Ernest Rutherford, bombarded gold foil with alpha particles and found that most of the particles went straight through the foil but a few were deflected back off the foil. He concluded that the atom was mostly empty space surrounding a massive center he called the nucleus.
Neils Bohr, a student of Rutherford, developed the solar system model of the atom in which the electrons surround the nucleus much like the planet surround the sun.

The three basic subatomic particles are <u>protons</u>, <u>neutrons</u>, and <u>electrons</u>. Masses are expressed in atomic mass units, <u>(amu</u>).

Particle	Location	Charge	Mass
Proton	Nucleus	+	1 amu
Neutron	" " "	0	1 amu
Electron	Outside the nucleus	−	0 amu

Normally particles of the same charge repel, but when protons are very close to each other or to neutrons there is a strong attraction between these particles called <u>nuclear forces.</u>

The modern model of the atom is that the atom consists of a central nucleus composed of protons and neutrons surrounded by a region of electrons called the electron cloud. Because the electrons can be at a very far distance from the nucleus most of the volume of the atom is empty space.

Answer the following questions:

1. **Define each of the following.**

   Electron--

   Nucleus--

   Proton--

   Neutron--

2. **What conclusions were reached by each of the following scientists.**

   Thomson--

   Millikan--

Rutherford--

Bohr--

3. What comprises a cathode ray?

4. What deflects cathode rays?

5. What force holds the nucleus together?

Why must this force in the nucleus exist?

<u>Section 3 Counting Atoms pgs. 77-82 only</u>.
<u>Objectives</u>:
1 Explain what isotopes are.
2. Define atomic number and mass number, and describe how they apply to isotopes.
3. Given the identity of a nuclide, determine its number of protons, neutrons, and electrons.
<u>Vocabulary</u>: Define the following.

1. atomic number--

2. isotope--

3. mass number--

4. nuclide--

5. atomic mass unit--

6. average atomic mass--

Atomic Number

The atomic number (Z) is the number of protons in an atom's nucleus. ***The atomic number determines the identity of the element. No two atoms have the same atomic number. An atom is electrically neutral; meaning the number of protons and electrons are equal. An atom can gain or lose electrons but not protons in chemical reactions.

Ions

Ions are formed when an atom gains or loses electrons. If the atom gains electrons it becomes a negative ion. If the atom loses electrons it becomes a positive ion.
Examples:
If a sodium loses one electron it would become a sodium ion written Na^{+}. If a calcium atom loses two electrons it becomes a calcium ion written Ca^{2+}. If an fluorine atom gains one electron it would be written as F^{-}. If an oxygen atom gains two electrons it would be written as O^{2-}.

Isotopes

Isotopes are atoms of an element that have the same number of protons but a different number of neutrons. These isotopes of the same element are chemically the same but have different masses.
When chemists are referring to isotopes they sometimes use a number after the chemical's name, called the mass number. For example two of the isotopes of carbon are carbon-12 and carbon-14; the 12 and 14 are the mass numbers. This method is referred to as hyphen notation. Another way of writing the mass number of these two isotopes, is a form called the nuclear symbol form, this method gives more information such as $^{12}_{6}C$ and $^{14}_{6}C$. Here the 12 and 14 superscripts refer to the mass number and the subscript 6 refers to the atomic number. ****Mass numbers are not found in the periodic table.

The most common and simplest element is hydrogen. Hydrogen has atomic number 1. This means hydrogen has 1 proton and 1 electron. Hydrogen the element also has three isotopes. Since these isotopes of hydrogen are so well known, they also have chemical "nicknames" that are shown in

the following enclosed in parentheses. The three isotopes of hydrogen are as follows.

Hydrogen-1, or $^{1}_{1}H$ (protium); Hydrogen-2 or $^{2}_{1}H$ (deuterium); Hydro-

gen-3, or $^{3}_{1}H$ (tritium).
Problem: For the three isotopes of hydrogen give the number of protons, electrons, and neutrons for each.
Answer____________
Problem: How many protons, neutrons, and electrons are in uranium-235, $^{235}_{92}U$?
Answer: ___________________

Relative Atomic Masses
Since atoms are so small a relative scale of atomic mass was chosen by which all other atoms could be compared. The atom chosen was carbon-12, and was arbitrarily assigned the mass of 12 amu (atomic mass unit).
Weighted and Unweighted Averages
To understand how atomic mass is calculated for each element it is necessary to understand the difference between a weighted average and an unweighted average. Students are most familiar with unweighted averages. An unweighted average means that one number is no more important than the other in obtaining the average (they carry the same weight). For instance the average of 6 and 10 is 8. In an unweighted average the numbers are added 10 + 6 = 16, and then divided by the total of the numbers being averaged (2). 16 ÷ 2 = 8.
The assumption is that the 6 and the 10 occur in the same percentage, in this case 50%.
A weighted average on the other hand means that the numbers do not occur in an equal percentage. For example, if the number 10 occurs 20% and the number 6 occurs 80% then the weighted average is 6.8.

What most students learned in their early math courses, the unweighted average, was actually a shortcut method. Look at this problem using a percentage method and you will see the difference.

Unweighted

6 + 10 = 16 16 ÷ 2 = 8 or
6 x 50% = 3, 10 x 50% = 5, 3 + 5 = 8

Weighted

6 x 80% = 4.8, 10 x 20% = 2, 4.8 + 2 = 6.8

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The relative percentages of isotopes of a particular element are constant on earth. For example the element chlorine has 2 isotopes, chlorine-35 and chlorine-37. These isotopes do not occur in the same percentages. Chlorine-35 occurs about 75% if the

time and chlorine-37 occurs about 25% of the time. The weighted average of these isotopes is 35.45 amu. This weighted average is called the atomic mass. This is the atomic mass that is found in the periodic table.

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The three terms: atomic mass, mass number, and atomic number often cause confusion. Students must memorize how these terms compare. A summary follows.

Atomic Number

The number of protons in an atom; this is listed in the periodic table.

Atomic Mass

The weighted average of all of the isotopes of that element; this is also listed in the periodic table.

Mass Number

The number of protons and neutrons of a particular isotope of an element; this is not listed in the periodic table.

Determine the number of protons, neutrons, and electrons, in each of the following isotopes.

1. silicon-28

2. iron-56

3. carbon-12

4. carbon-14

4 Determine the relative atomic mass to two decimal places of

a. potassium

b. hydrogen

c. oxygen

d. chlorine

5. Write the hyphen notation for an atom with a mass number of 63 and atomic number 29.

Write the nuclear symbol form for the above isotope.

6. There are three isotopes of argon that occur in nature. Calculate the average atomic mass of argon to 2 decimal places, given the following.
argon-36, mass 35.97 amu occurs 0.337%, argon-38, mass 37.96 amu occurs 0.0063%, and argon-40, mass 39.96 amu occurs 99.600%.