NS F2 (pg 1 of 2) Navigating the Periodic Table and Isotopes

Name________________________Per___

About 70 elements had been discovered by the mid 1800's. Be sure and check out the periodic table on the unit C home page on the web site that gives a colorful representation of when the elements were discovered. Dmitri Mendeleev, a Russian chemistry professor was writing a book for his chemistry students and in his attempt to organize the known elements he noticed that when the elements were arranged in columns and rows, increasing mass from left to right, that the columns were arranged with elements that had similar chemical properties. Mendeleev's original periodic table had gaps in it to represent the elements that had not yet been discovered. He even went so far as to predict the properties of a few of those undiscovered elements. His predictions were remarkably accurate when those elements were finally isolated. Mendeleev's chart did have some irregularities, and he assumed that this may have been caused by errors in the measurements of the masses of the elements. Subsequently, the mass measurements were shown to be remarkably accurate.

Forty years later in 1913 Henry Mosley had determined that the nucleus of each element has a positive charge unique to that element. Mosley decided to rearrange Mendeleev's chart in order of increasing nuclear charge instead of by increasing mass. Mosely's improvement to the chart remains today as the modern periodic chart is arranged in order from left to right by increasing number of protons.

The periodic chart is periodic, meaning that it exhibits repeating patterns and trends of the properties of the elements. These trends can be analyzed both vertically and horizontally. The trends that we will eventually study will be:

1 electron arrangement
2 atomic size
3 ionic size
4 ionization energy
5 electronegativity

It is important to make friends with the periodic chart so that you can navigate around it.

Columns and Rows on the Chart

- Column, group, or family are the three terms used to describe a **vertical** set of elements.
- Row or period is the term used to describe a **horizontal** set of elements.

Metals and Nonmetals

It is very important to recognize the staircase that separates the metal elements from the nonmetals. All elements to the left are metals (which exhibit physical characteristics such as high heat and electrical conductivity, shiny luster, and all are solid at room temp except for mercury), and all elements to the right (and hydrogen as well) are nonmetals (which exhibit characteristic properties that are different from the metals). Many elements along the stairs separating the metals from the nonmetals exhibit both metal and nonmetal physical properties. These elements are sometimes classified as metalloids. Silicon and germanium are two of these important elements.

Representative Elements (group A)

The elements in the columns headed by the numbers 1A – 8A (also 1-2, and 13−18) are known as the representative elements.

Noble or Inert Gases

The last column, group 8A (aka 18) to the far right headed by helium ($_2$He) is known as the Noble or inert gases. They are considered inert because they participate in very few chemical reactions, and are not found in naturally occurring compounds.

Halogens

The elements in group 7A (aka 17) headed by the element fluorine are known as the halogens.

Alkali Metals

The metals in group 1A headed by the metallic element lithium are known as the alkali metals. Note that hydrogen, while at the top of this family, is not an alkali metal hydrogen is a nonmetal

Transition Metals (group B)

The elements in between and including Sc and Zn are known as the transition metals.

Lanthanide and Actinide Series (group C)

The elements in the bottom row of the periodic chart are not very common, are all radioactive, and are known as the actinide series. The row just above it is known as the lanthanide series and is equally uncommon. These elements will not be the focus of this course.

Isotopes on the back →

NS F2 (pg 2 of 2) Isotopes

The average atomic mass (aka average molar mass) is the weighted average of all of the naturally occurring isotopes of an element. So what does this mean???

Two atoms (or ions) with the **same number of protons but different number of neutrons are called isotopes**. For instance, naturally occurring chlorine occurs as two isotopes. ^{37}Cl which has 17 protons and 20 neutrons, and ^{35}Cl which has 17 protons and 18 neutrons. A few elements, such as fluorine, phosphorus, and iodine occur naturally with just one isotope, but most elements are made up of a mixture of isotopes. Another example is element number 22, titanium (Ti), a light and strong metal used in jet engines and in artificial human joints. There are five naturally occurring isotopes of titaium. Each one has 22 protons in its nucleus, but the number of neutrons varies from 24 to 28. Tin has the largest number of naturally occurring stable isotopes, ten.

In chemical reactions, all isotopes of an element behave nearly identically since the number of neutrons does not affect its electron or chemical behavior.

Each isotope of a particular element has a different mass, therefore the mass of one mole of a particular isotope has its own unique value. Since a sample of a chemical element is usually a **mixture of all its naturally occurring stable isotopes**, the atomic mass that is reported in the periodic chart must reflect the masses of all its naturally occurring stable isotopes. Since the isotopes occur in differing proportions, the average mass must proportionately reflect the masses of the different isotopes. Thus the average atomic mass of an element is a weighted average.

Calculating the weighted averages

For example, the analysis of the four isotopes of iron give the following data:

isotope	isotopic molar mass (g/mole)	naturally occurring abundance
^{54}Fe	53.940	5.82 %
^{56}Fe	55.935	91.66 %
^{57}Fe	56.935	2.19 %
^{58}Fe	57.933	0.33 %

To calculate the weighted average atomic mass (which is the average atomic mass given in the periodic chart), multiply the percent of each isotope (in decimal form) by its particular molar mass, then add those quantities together.

$(0.0582 \times 53.940) + (0.9166 \times 55.935) + (0.0219 \times 56.935) + (0.0033 \times 57.933) = 55.85$ g/mole

Note that the average molar mass will usually round to the most abundant isotope. While this is not always true, it is a close enough approximation.

Calculating % abundances

The reverse calculation can also be made, since the average atomic mass is in the periodic chart, it can be used together with information about the mass of each isotope, ^{35}Cl and ^{37}Cl to calculate the percent natural abundance.

Using chlorine and its two isotopes as an example.

Knowing that % always adds up to 100, and when used as a decimal, the two values will always add up to 1. Thus you can name one isotope as x and the other as [1 − x]. It doesn't matter which one is designated as which.

$(x \times 35) + ([1 - x] \times 37) = 35.45$

Distributing and gathering terms will yield a fractional value for x which allows the determination of the two percentages.

$35x + 37 - 37x = 35.45$
$-2x = -1.55$
$x = 0.775$ since x was defined as ^{35}Cl there must be approximately 77.5 % and by subtraction from 100 since there are only two isotopes, there must be approximately 22.5 % of ^{37}Cl.

The calculation above is a close approximation because the actual isotopic molar masses are not really 35 and 37, but rather 34.96885 and 36.9659 respectively.

You would only be asked to do this calculation for a element that has only two naturally occurring isotopes.