Lesson Title	Totally Isotopic!
Length of Lesson	1 hour 15 min
Created By	Chris Ruhs
Subject	Chemistry
Grade Level	10-12th Grade
State Standards	Chemistry 1: 1d,f; 2c,d,e; 3c
DOK Level	DOK 3
DOK Application	Use concepts to solve non-routine problems, investigate, assess.
National Standards	9-12: A (Inquiry); B (Physical Science)
Graduate Research Element	Isotopes are used for quantification and comparison of environmental samples in GC-MS applications.

Student Learning Goal:

MS 9-12th Grade:

1 (d) Apply the language of chemistry appropriately including terms such as element, atom, compound, and molecule. *Student-teacher discussion will revolve heavily around the terms "isotope," "atom," "proton," and "neutron".* (f) Relate symbols to names of common chemical elements. *When discussing isotopes and using the periodic table, students will relate symbols to element names.* 2 (c) Apply the definition of mass, length, volume, time, density, temperature and pressure. *Students understand how mass relates to subatomic particles and is used in determining atomic weight.* (d) Use scientific notation in chemical calculations. *Students will calculate atomic weights using scientific notation.* (e) Round values to the proper significant digits. *Students will calculate atomic weights and round values to proper significant digits.* 3 (c) Determine the number of protons, electrons, or neutrons in an element when given the atomic number and the atomic mass of the element, or vice versa. *Students will need to know the number of protons and neutrons of an element when calculating atomic masses.*

National Science Education Standards of Content 9-12:

A: Inquiry: Use appropriate tools and techniques to gather, analyze, and interpret data; think critically and logically to make the relationships between evidence and explanations. Use technology and mathematics to improve investigations and communications. *Students will "discover" the average atomic masses found on the periodic table.*

B: Physical Science: Structure of atoms. Structure and properties of matter. *Students will understand that isotopes have a slightly different structure, different numbers of subatomic particles, and sometimes different properties within the same element.*

Materials Needed (supplies, hand-outs, resources):

Hand-out on fractional averaging, hand-out directing students through a dynamic online periodic table with problem sets to calculate average atomic weights given natural abundances found in nature, computer lab, computers that can access and run http://www.ptable.com/.

Lesson Performance Task/Assessment:
Formative: opening discussion on isotopes, introduction to averaging technique, and closing teacher-guided, student-lead summary discussion.

Summative: completed hand-outs on the averaging technique and the four elements to calculate.

Lesson Relevance to Performance Task and Students:
Students will be able to transfer their previously formed ability to calculate averages into an alternative averaging technique that uses percentages or decimals. This transfer is easily accomplished by showing students simple averaging problems using both the technique they already have mastered (adding each number together and dividing by the total of numbers) and the new technique where each number is multiplied by 1/the total of numbers. For example, the average of 4 and 6 can be calculated to equal 5 by adding 4 + 6 and dividing by the total of numbers, which is 2 numbers. Therefore, (4 + 6) / 2 = 5. In the new technique, the average of 4 and 6 can be calculated to equal 5 by multiplying each number by 1/the total of numbers, which in this case would be ½, and adding the products together. Therefore, 4 x ½ + 6 x ½ = 5. The instructor can then build upon this approach by assigning a larger total of numbers to average, and then attempt to change from using equal fractions to using unequal fractions. For example, if 3 students made a 90 on their homework, but a fourth student made an 80, what is the average grade? Using the new technique, 75% of students made a 90 while 25% of students made an 80. Therefore, the average can be calculated (using unequal fractions) as 90 x 0.75 + 80 x 0.25 = 87.5. The instructor can show the students the same answer using the old method again, where (90 + 90 + 90 + 80) / 4 = 87.5. Once the students are convinced that this new technique works and have become confident in using it, students can proceed to the computer labs so that they can apply this new skill to the topic of isotopic ratios in nature.

Once at the computer labs, students will be guided through using the dynamic online periodic table found at http://www.ptable.com/ and given a hand-out which will ask them to determine the average natural abundance of 4 different elements of their choosing, one of which should be the element assigned to them from the "3 elementary particles lesson plan." This activity will allow them to calculate the average atomic weight of 4 different elements of their choosing, and then compare their answer to the number given as the actual average natural abundance. Hopefully, students will have the experience of discovering a useful scientific number that agrees with what other scientists have previously reported as well as giving them a complete understanding of what that number actually means and how it can be used.

Anticipatory Set/Capture Interest:
Why do isotopes weigh different amounts? The number of neutrons varies between atoms of the same element, forming isotopes of different weights. Short discussion on how isotopes are written out (two forms). How much does Carbon-12 weigh? 12 grams/mole. What about Carbon-13? 13 grams/mole. What about Carbon-14? 14 grams/mole. Well, which of these isotopes do we work with in chemistry? (This question is intentionally confusing and will either lead the students to naively choose Carbon-12 or intelligently assert that the isotopes are all mixed up in nature—if they naively choose Carbon-12, ask the students where the other isotopes exist). They're all mixed up?! That's not good! It would take forever to count out atoms, but we need to know how many atoms we have so we can set up chemical reactions properly. If we needed exactly $6.022x10^{23}$ atoms (1 mole of atoms) of Carbon-12, then we would simply have to weigh out 12 grams and we would have exactly $6.022x10^{23}$ atoms of Carbon-12. But our Carbon atoms are not just the Carbon-12 variety—our atoms have some Carbon-12, some Carbon-13, and some Carbon-14. We can't just weigh out 12 grams of the mixture of different Carbon atoms to get $6.022x10^{23}$ atoms of Carbon, because some of those atoms weigh more than 12 grams/mole, and our answer of 12 grams would be wrong. So what do we do? (A long pause here is ok). Wouldn't it be nice if someone could tell us that we have mostly Carbon-12 and just a little bit of Carbon-13 and Carbon-14? That way we could adjust our answer and weigh out the right amount. Or even better, wouldn't it be great if someone just gave us an average weight that would represent all the different isotopes of Carbon in our sample? That way we could weigh out that average number of grams/mole in order to get 1 mole. For instance, if we somehow knew that half of our Carbon atoms weighed 12 grams/mole and the other half weighed 14 grams/mole, we would say that the average weight of Carbon atoms would be….? 13 grams/mole. So, all we would have to do is weight out 13 grams of Carbon atoms—half of them weigh 12 grams/mole, the other half weigh 14 grams/mole, so, on average, 13 grams of these mixed isotopes would give us $6.022x10^{23}$ atoms, or 1 mole.

Guided Practice:
Students will be given a hand-out, guided through a discussion on averages, and exposed to an alternative technique for calculating averages using percents and/or decimals. Once this new averaging technique is mastered, students will be taken to the computer labs, given a second hand-out, and guided through the use of http://www.ptable.com/ to answer questions on a the hand-out.

Independent Practice:
Students apply the averaging technique mastered in the first part of this lesson to an inquiry problem using a hand-out and http://www.ptable.com/

Remediation and/or Enrichment:
Remediation:
Individual IEP

Enrichment:
This lesson plan would flow naturally into a discussion on nuclear decay, radiation, half-lives, the use of radioactive isotopes, etc.

Check(s) for Understanding:
How did we come to know the atomic masses of the elements?
Would the average atomic mass change on a different planet?
Would the period table made for Mars have different values for some elements?
Completed hand-outs, ongoing dialogue with students, and observation of student activity.

Closure:
A student-lead, teacher-guided summary discussion will bring closure to the lesson plan.

Possible Alternate Subject Integrations:
Physics could use this lesson plan as an introduction to radioactivity.
Math could use this lesson plan for the averaging techniques found in it and to get math students thinking about math applications.

Teacher Notes:
Inside the nucleus of an atom, protons and neutrons are arranged for stability—this stable arrangement occurs naturally and represents the lowest possible energy state for the system. However, some atoms are arranged in such a way that they are "barely stable," so that a very small fluctuation within the nucleus could trigger a chain reaction leading to a reconfiguration of the nucleons and release of energy either through a photon or through a high energy electron or alpha particle. This phenomenon happens spontaneously with radioactive isotopes and is the essence of radioactivity.
In the case of an atomic bomb, high energy particles are released by the spontaneous decay of radioactive heavy metals. The released particles then bombard neighboring atoms, imparting enough energy to them so that they also decay and release even more high energy particles, which then bomard atoms neighboring them, resulting in a chain reaction until all the radioactive atoms decay within a very small time span. This sudden release of vast amounts of energy is what makes the explosion, with $E=mc^2$ governing the amount of energy released.
It should be noted that "$E=mc^2$" does not mean that mass is converted into energy, but that mass and energy are in fact the same concept. When an atomic bomb explodes there is indeed a measureable loss of mass, but this mass was not *converted* into energy, it *is* the energy. The energy that left the system is itself the mass that was lost. Energy and mass are on and the same. It takes a lot of energy to be able to measure a mass change because the speed of light squared is such a large number.