high school forum

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SECONDARY SCHOOL CHEMISTRY

Editors Note: Some of the popular high school chemistry texts contain discussions and homework problems that clearly suggest that the *only* reason that atomic weights deviate from whole numbers is the existence of isotopes. Furthermore, the editor has found that few undergraduates or beginning graduate students in chemistry appear to understand that matters are not this simple. The misconception is apparently common. For this reason the following article was solicited. Papers dealing with other misconceptions commonly taught in introductory courses would be welcomed.

Binding Energy and Atomic Weight Calculations

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There are several related questions surrounding the concept of atomic weights, as this concept is presented in several introductory chemistry textbooks, which appear to be a source of some confusion. For example

1) The "Handbook of Chemistry and Physics" gives the rest mass of the electron, proton and neutron as 5.48597×10^{-4}, 1.00727663, and 1.0086654 amu, respectively. However, the weight of a $^{12}_{6}C$ atom is defined as exactly 12.000 . . . amu, or less than the sum of the weights of its parts.

2) The average atomic weight of carbon is said to be larger than 12 amu, i.e., 12.001 amu, due to the presence of the $^{13}_{6}C$ and $^{14}_{6}C$ isotopes. However, oxygen exhibits an average atomic weight of 15.9994 amu, although the only isotopes of abundance are $^{16}_{8}O$, $^{17}_{8}O$ and $^{18}_{8}O$

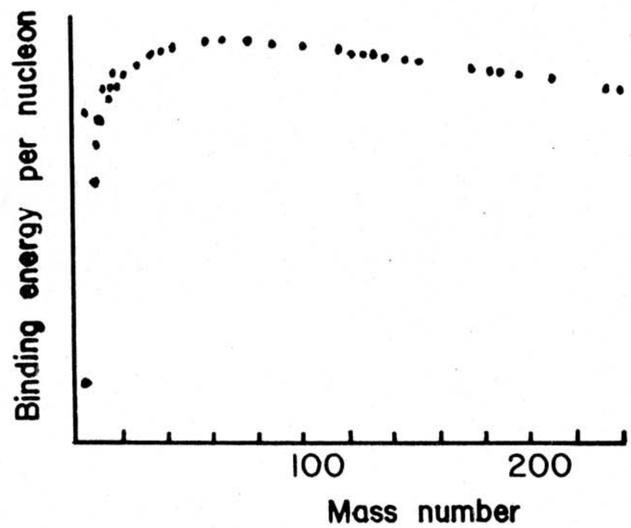
3) When a student is asked to calculate the relative abundances of the two isotopes of bromine, $^{79}_{35}Br$ and $^{81}_{35}Br$, on the basis of the assumption that the average atomic weight reflects these abundances, i.e.

$$(79.904 \text{ amu}) \cdot 100 = (79 \text{ amu}) \cdot X + 81 \text{ (amu)} \cdot Y$$
$$X + Y = 100$$

the values obtained ($X = 54.80\%$, $Y = 45.20\%$) differ significantly from the relative abundances found in the "Handbook" ($X = 50.54\%$, $Y = 49.56\%$).

This apparent confusion stems from the fact that the weight of an atom is not equal to the sum of the weights of its parts. For example, the predicted weight of the $^{12}_{6}C$ isotope ($6^1_1p + 6^1_0n + 12^0_{-1}e$) would be 12.10224 amu. The difference between the observed atomic weight and the predicted atomic weight is said to be the *mass defect* for $^{12}_{6}C$. In relativistic terms ($E = mc^2$) we argue that 0.10224 amu of mass is lost during the formation of the $^{12}_{6}C$ atom, and this corresponds to the release of 2.193×10^9 kcal/mole. This energy is said to represent the *binding energy* of the $^{12}_{6}C$ nucleus, or the energy released when this nucleus is formed via condensation of the nucleons. The binding energy is thus a reflection of the stability of the nucleus, and empirically is equal to 2.145×10^{10} kcal/mole per amu of mass that is lost. The suggestion that 652 million kcal/mole of energy are released during the thermonuclear fusion reactions leading to ^4_2He is readily understood by noting that the mass defect for ^4_2He is 0.03038 amu or 6.517×10^8 kcal/mole.

It is interesting to note that the mass defect for $^{12}_{6}C$ is approximately 3.37 times as large as the mass defect for ^4_2He. It appears that the mass defect or binding energy increases with increasing atomic number at a slightly faster rate than the increase in atomic weight. If we examine the mass defect or binding energy per nucleon, we note that this quantity increases with increasing atomic number (Z) until a maximum is observed at about $^{56}_{26}Fe$. Thus, fusion of the lighter nuclei for which Z is less than 26 liberates energy, whereas fission of the heavier nuclei with Z greater than 26 leads to the liberation of energy. A simplified representation of the binding energy dependence upon atomic number is shown in the figure.



We may now specifically address the questions raised earlier. We see that the weight of the $^{12}_{6}C$ atom is less than the sum of the weights of the particles from which this atom is composed due to the liberation of significant amounts of energy during its formation. The average atomic weight of oxygen is less than 16 amu since the true atomic weight of the $^{16}_{8}O$ isotope is 15.99491 amu, and the natural abundances of $^{17}_{8}O$ and $^{18}_{8}O$ total less than 0.25%. Finally, accurate relative abundances of $^{79}_{35}Br$ and $^{81}_{35}Br$ can be readily calculated on the basis of the true atomic weights of these isotopes.

$$(79.904 \text{ amu}) \cdot 100 = (78.9183 \text{ amu}) \cdot X + (80.9163 \text{ amu}) \cdot Y$$
$$X + Y = 100$$

The Thermometer Needs Renaming

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Beginning chemistry students often have difficulty appreciating the difference between *heat content*, an extensive property of matter, and *temperature*, an intensive property. In an effort to help, teachers point out that oceans contain tremendous amounts of heat, but have only modest temperatures. Yet, the fact remains that when a pot of coffee is