Chapter 2
Solutions to Selected Integrative and Advanced Exercises

69. **(M)** Each atom of ^{19}F contains 9 protons (1.0073 u each), 10 neutrons (1.0087 u each) and 9 electrons (0.0005486 u each). The mass of each atom should be the sum of the masses of these particles.

$$\text{Total mass} = \left(9\ \text{protons} \times \frac{1.0073\ \text{u}}{1\ \ \text{proton}}\right) + \left(10\ \text{neutrons} \times \frac{1.0087\ \text{u}}{1\ \ \text{proton}}\right) + \left(9\ \text{electrons} \times \frac{0.0005486\ \text{u}}{1\ \ \text{electron}}\right)$$

$$= 9.0657\ \text{u} + 10.087\ \text{u} + 0.004937\ \text{u} = 19.158\ \text{u}$$

This compares with a mass of 18.9984 u given in the periodic table. The difference, 0.160 u per atom, is called the mass defect and represents the energy that holds the nucleus together, the nuclear binding energy. This binding energy is released when 9 protons and 9 neutrons fuse to give a fluorine-19 nucleus.

72. **(M)** Let Z = # of protons, N = # of neutrons, E = # of electrons, and A = # of nucleons = $Z + N$.

(a) $Z + N = 234$ The mass number is 234 and the species is an atom.
$N = 1.600\ Z$ The atom has 60.0% more neutrons than protons.
Next we will substitute the second expression into the first and solve for Z.

$$Z + N = 234 = Z + 1.600\,Z = 2.600\,Z$$

$$Z = \frac{234}{2.600} = 90\ \text{protons}$$

Thus this is an atom of the isotope ^{234}Th.

(b) $Z = E + 2$ The ion has a charge of +2. $Z = 1.100\ E$
There are 10.0% more protons than electrons. By equating these two expressions and solving for E, we can find the number of electrons. $E + 2 = 1.100\,E$

$$2 = 1.100\,E - E = 0.100\,E \qquad E = \frac{2}{0.100} = 20\ \text{electrons} \qquad Z = 20 + 2 = 22,\ \text{(titanium)}.$$

The ion is Ti^{2+}. There is not enough information to determine the mass number.

(c) $Z + N = 110$ The mass number is 110. $Z = E + 2$ The species is a cation with a charge of +2.
$N = 1.25\ E$ Thus, there are 25.0% more neutrons than electrons. By substituting the second and third expressions into the first, we can solve for E, the number of electrons.

$$(E + 2) + 1.25\,E = 110 = 2.25\,E + 2 \qquad 108 = 2.25\,E \qquad E = \frac{108}{2.25} = 48$$

Then $Z = 48 + 2 = 50$, (the element is Sn) $N = 1.25 \times 48 = 60$ Thus, it is $^{110}Sn^{2+}$.

74. **(M)** $A = Z + N = 2.50\,Z$ The mass number is 2.50 times the atomic number
The neutron number of selenium-82 equals 82 – 34 = 48, since Z = 34 for Se. The neutron number of isotope Y also equals 48, which equals 1.33 times the atomic number of isotope Y.

Thus $48 = 1.33 \times Z_Y \quad Z_Y = \dfrac{48}{1.33} = 36$

The mass number of isotope Y = 48 + 36 = 84 = the atomic number of E, and thus, the element is Po. Thus, from the relationship in the first line, the mass number of $\text{E} = 2.50\,Z = 2.50 \times 84 = 210$ The isotope E is ^{210}Po.

76. **(M)** To solve this question, represent the fractional abundance of ^{14}N by x and that of ^{14}N by $(1 - x)$. Then use the expression for determining average atomic mass.

$$14.0067 = 14.0031\,x + 15.0001(1 - x)$$

$$14.0067 - 15.0001 = 14.0031\,x - 15.0001\,x \qquad \text{OR} \qquad -0.9934 = -0.9970\,x$$

$$x = \frac{0.9934}{0.9970} \times 100\% = 99.64\% = \text{percent abundance of } ^{14}\text{N}. \quad \text{Thus, } 0.36\% = \text{percent abundance of } ^{15}\text{N}.$$

77. **(D)** In this case, we will use the expression for determining average atomic mass- the sum of products of nuclidic mass times fractional abundances (from Figure 2-14)- to answer the question.

^{196}Hg: 195.9658 u × 0.00146 = 0.286 u

^{198}Hg: 197.9668 u × 0.1002 = 19.84 u ^{199}Hg: 198.9683 u × 0.1684 = 33.51 u

^{200}Hg: 199.9683 u × 0.2313 = 46.25 u ^{201}Hg: 200.9703 u × 0.1322 = 26.57 u

^{202}Hg: 201.9706 u × 0.2980 = 60.19 u ^{204}Hg: 203.9735 u × 0.0685 = 14.0 u

Atomic weight = 0.286 u + 19.84 u + 33.51 u + 46.25 u + 26.57 u + 60.19 u + 14.0 u = 200.6 u

79. **(D)** First, it must be understood that because we have don't know the exact percent abundance of ^{84}Kr, all the percent abundances for the other isotopes will also be approximate. From the question, we may initially infer the following:
(a) Assume percent abundance of ^{84}Kr ~ 55% as a start (somewhat more than 50)
(b) Let percent abundance of ^{82}Kr = x %; percent abundance ^{83}Kr ~ ^{82}Kr = x %
(c) ^{86}Kr = 1.50(percent abundance of ^{82}Kr) = 1.50(x%)
(d) ^{80}Kr = 0.196(percent abundance of ^{82}Kr) = 0.196(x%)
(e) ^{78}Kr = 0.030(percent abundance of ^{82}Kr) = 0.030(x%)

100% = %^{78}Kr + %^{80}Kr + %^{82}Kr + %^{83}Kr + %^{84}Kr + %^{86}Kr
100% = 0.030(x%) + 0.196(x%) + x% + x% + %^{84}Kr + 1.50(x %)
100% = 3.726(x%) + %^{84}Kr

Assuming percent abundance of ^{84}Kr is 55%, solving for x gives a value of 12.1% for percent abundance of ^{82}Kr, from which the remaining abundances can be calculated based on the above relationships, as shown below:

^{78}Kr: 0.03×12.1 = 0.363%; ^{80}Kr: 0.196×12.1=2.37%; ^{83}Kr: same as ^{82}Kr; ^{86}Kr: 1.5×12.1 = 18.15%.

The weighted-average isotopic mass calculated from the above abundances is as follows:
Weighted-average isotopic mass = 0.030(12.1%)(77.9204 u) + 0.196(12.1%)(79.9164 u) + 12.1%(81.9135 u) + 12.1%(82.9141 u) + 55%(83.9115 u) + 1.50(12.1%)(85.9106 u) = 83.8064 u

As stated above, the problem here is the inaccuracy of the percent abundance for ^{84}Kr, which is crudely estimated to be ~ 55%. If we vary this percentage, we vary the relative abundance of all other isotopes accordingly. Since we know the weighted-average atomic mass of Kr is 83.80, we can try different values for ^{84}Kr abundance and figure out which gives us the closest value to the given weighted-average isotopic mass:

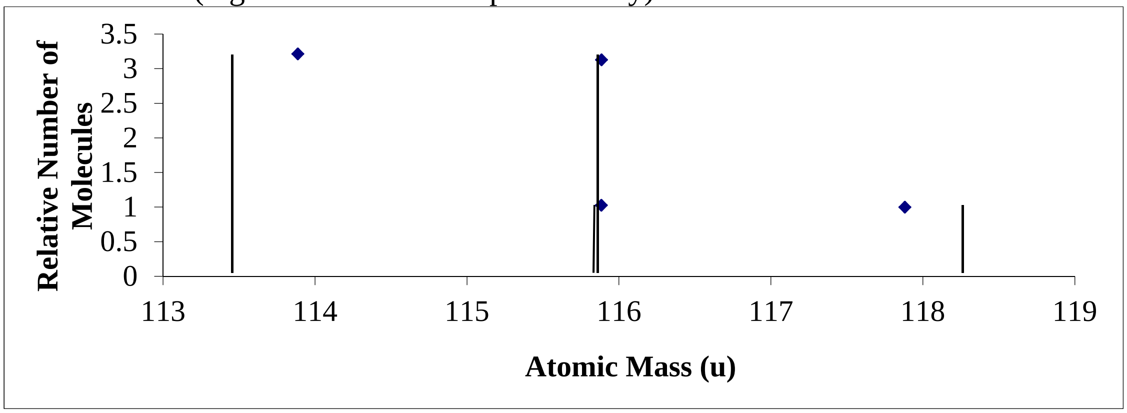
Percent Abundance ^{84}Kr	**Weighted-Average Isotopic Mass**
50 %	83.793
51 %	83.796
52 %	83.799
53 %	83.801
54 %	83.803
55 %	83.806

From this table, we can see that the answer is somewhere between 52% and 53%.

80. **(D)** Four molecules are possible, given below with their calculated molecular masses.
^{35}Cl-^{79}Br mass = 34.9689 u + 78.9183 u = 113.8872 u
^{35}Cl-^{81}Br mass = 34.9689 u + 80.9163 u = 115.8852 u
^{37}Cl-^{79}Br mass = 36.9658 u + 78.9183 u = 115.8841 u
^{37}Cl-^{81}Br mass = 36.9658 u + 80.9163 u = 117.8821 u

Each molecule has a different intensity pattern (relative number of molecules), based on the natural abundance of the isotopes making up each molecule. If we divide all of the values by the lowest ratio, we can get a better idea of the relative ratio of each molecule.
^{35}Cl-^{79}Br Intensity = (0.7577) × (0.5069) = 0.3841 ÷ 0.1195 = 3.214
^{35}Cl-^{81}Br Intensity = (0.7577) × (0.4931) = 0.3736 ÷ 0.1195 = 3.127
^{37}Cl-^{79}Br Intensity = (0.2423) × (0.5069) = 0.1228 ÷ 0.1195 = 1.028
^{37}Cl-^{81}Br Intensity = (0.2423) × (0.4931) = 0.1195 ÷ 0.1195 = 1.000

A plot of intensity versus molecular mass reveals the following pattern under ideal circumstances (high resolution mass spectrometry).



81. **(M)** Let's begin by finding the volume of copper metal.

wire diameter (cm) = 0.03196 in. × $\dfrac{2.54\text{ cm}}{1\text{ in.}}$ = 0.08118 cm

The radius is 0.08118 cm × 1/2 = 0.04059 cm

The volume of Cu(cm^3) = $(0.04059\text{ cm})^2 \times (\pi) \times (100\text{ cm}) = 0.5176\text{ cm}^3$

So, the mass of Cu = $0.5176\text{ cm}^3 \times \dfrac{8.92\text{ g Cu}}{1\text{ cm}^3} = 4.62\text{ g Cu}$

The number of moles of Cu = $4.62\text{ g Cu} \times \dfrac{1\text{ mol Cu}}{63.546\text{ g Cu}} = 0.0727\text{ mol Cu}$

Cu atoms in the wire = $0.0727\text{ mol Cu} \times \dfrac{6.022 \times 10^{23}\text{ atoms Cu}}{1\text{ mol Cu}} = 4.38 \times 10^{22}\text{ atoms}$

84. **(M)** The numbers sum to 21 (= 10 + 6 + 5). Thus, in one mole of the alloy there is $^{10}\!/_{21}$ mol Bi, $^{6}\!/_{21}$ mol Pb, and $^{5}\!/_{21}$ mol Sn. The mass of this mole of material is figured in a similar fashion to computing a weighted-average atomic mass from isotopic masses.

$$\text{mass of alloy} = \left(\frac{10}{21}\text{mol Bi} \times \frac{209.0\text{ g}}{1\text{ mol Bi}}\right) + \left(\frac{6}{21}\text{mol Pb} \times \frac{207.2\text{ g}}{1\text{ mol Pb}}\right) + \left(\frac{5}{21}\text{mol Sn} \times \frac{118.7}{1\text{ mol Sn}}\right)$$
$$= 99.52\text{ g Bi} + 59.20\text{ g Pb} + 28.26\text{ g Sn} = 186.98\text{ g alloy}$$

86. **(M)** The relative masses of Sn and Pb are 207.2 g Pb (assume one mole of Pb) to (2.73 × 118.710 g/mol Sn =) 324 g Sn. Then the mass of cadmium, on the same scale, is 207.2/1.78 = 116 g Cd.

$$\%\text{Sn} = \frac{324\text{ g Sn}}{207.2 + 324 + 116\text{ g alloy}} \times 100\% = \frac{324\text{ g Sn}}{647\text{ g alloy}} \times 100\% = 50.1\%\text{ Sn}$$

$$\%\text{Pb} = \frac{207.2\text{ g Pb}}{647\text{ g alloy}} \times 100\% = 32.0\%\text{ Pb} \qquad \%\text{Cd} = \frac{116\text{ g Cd}}{647\text{ g alloy}} \times 100\% = 17.9\%\text{ Cd}$$

87. **(M)** We need to apply the law of conservation of mass and convert volumes to masses:

Calculate the mass of zinc:	$125\text{ cm}^3 \times 7.13\text{ g/cm}^3 = 891\text{ g}$
Calculate the mass of iodine:	$125\text{ cm}^3 \times 4.93\text{ g/cm}^3 = 616\text{ g}$
Calculate the mass of zinc iodide:	$164\text{ cm}^3 \times 4.74\text{ g/cm}^3 = 777\text{ g}$
Calculate the mass of zinc unreacted:	$(891 + 616 - 777)\text{ g} = 730\text{ g}$
Calculate the volume of zinc unreacted:	$730\text{ g} \times 1\text{cm}^3 / 7.13\text{ g} = 102\text{ cm}^3$