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CHAPTER 2

ATOMS AND THE ATOMIC THEORY

PRACTICE EXAMPLES

- 1A** (E) The total mass must be the same before and after reaction.
mass before reaction = 0.382 g magnesium + 2.652 g nitrogen = 3.034 g
mass after reaction = magnesium nitride mass + 2.505 g nitrogen = 3.034 g
magnesium nitride mass = 3.034 g – 2.505 g = 0.529 g magnesium nitride
- 1B** (E) Again, the total mass is the same before and after the reaction.
mass before reaction = 7.12 g magnesium + 1.80 g bromine = 8.92 g
mass after reaction = 2.07 g magnesium bromide + magnesium mass = 8.92 g
magnesium mass = 8.92 g – 2.07 g = 6.85 g magnesium
- 2A** (M) In Example 2-2 we are told that 0.500 g MgO contains 0.301 g of Mg. With this information, we can determine the mass of magnesium needed to form 2.000 g of magnesium oxide.

$$\text{mass of Mg} = 2.000 \text{ g MgO} \times \frac{0.301 \text{ g Mg}}{0.500 \text{ g MgO}} = 1.20 \text{ g Mg}$$

The remainder of the 2.00 g of magnesium oxide is the mass of oxygen.
mass of oxygen = 2.00 g magnesium oxide – 1.20 g magnesium = 0.80 g oxygen

- 2B** (M) In Example 2-2, we see that a 0.500 g sample of MgO has 0.301 g Mg, hence, it must have 0.199 g O₂. From this we see that if we have equal masses of Mg and O₂, the oxygen is in excess. First we find out how many grams of oxygen reacts with 10.00 g of Mg.

$$\text{mass}_{\text{oxygen}} = 10.00 \text{ g Mg} \times \frac{0.199 \text{ g O}_2}{0.301 \text{ g Mg}} = 6.61 \text{ g O}_2 \text{ (used up)}$$

Hence, 10.00 g – 6.61 g = 3.39 g O₂ unreacted. Mg is the limiting reactant.

$$\text{MgO(s) mass} = \text{mass Mg} + \text{Mass O}_2 = 10.00 \text{ g} + 6.61 \text{ g} = 16.61 \text{ g MgO.}$$

There are only two substances present, 16.61 g of MgO (product) and 3.39 g of unreacted O₂

- 3A** (E) Silver has 47 protons. If the isotope in question has 62 neutrons, then it has a mass number of 109. This can be represented as $^{109}_{47}\text{Ag}$.

- 3B** (E) Tin has 50 electrons and 50 protons when neutral, while a neutral cadmium atom has 48 electrons. This means that we are dealing with Sn²⁺. We do not know how many neutrons tin has, so there can be more than one answer. For instance,

$^{116}_{50}\text{Sn}^{2+}$, $^{117}_{50}\text{Sn}^{2+}$, $^{118}_{50}\text{Sn}^{2+}$, $^{119}_{50}\text{Sn}^{2+}$, and $^{120}_{50}\text{Sn}^{2+}$ are all possible answers.

- 4A** (E) The ratio of the masses of ^{202}Hg and ^{12}C is: $\frac{^{202}\text{Hg}}{^{12}\text{C}} = \frac{201.97062 \text{ u}}{12 \text{ u}} = 16.830885$

4B (E) Atomic mass is $12 \text{ u} \times 13.16034 = 157.9241 \text{ u}$. The isotope is ${}^{158}_{64}\text{Gd}$. Using an atomic mass of 15.9949 u for ${}^{16}\text{O}$, the mass of ${}^{158}_{64}\text{Gd}$ relative to ${}^{16}\text{O}$ is

$$\text{relative mass to oxygen-16} = \frac{157.9241 \text{ u}}{15.9949 \text{ u}} = 9.87340$$

5A (E) The atomic mass of boron is 10.81, which is closer to 11.0093054 than to 10.0129370. Thus, boron-11 is the isotope that is present in greater abundance.

5B (E) The average atomic mass of indium is 114.82, and one isotope is known to be ${}^{113}\text{In}$. Since the weighted-average atomic mass is almost 115, the second isotope must be larger than both In-113 and In-114. Clearly, then, the second isotope must be In-115 (${}^{115}\text{In}$). Since the average atomic mass of indium is closest to the mass of the second isotope, In-115, then ${}^{115}\text{In}$ is the more abundant isotope.

6A (M) Weighted-average atomic mass of Si =

$$\begin{aligned} (27.9769265325 \text{ u} \times 0.9223) &\rightarrow 25.80 \text{ u} \\ (28.976494700 \text{ u} \times 0.04685) &\rightarrow 1.358 \text{ u} \\ (29.973377017 \text{ u} \times 0.03092) &\rightarrow 0.9268 \text{ u} \\ &\underline{\hspace{1.5cm}} \\ &28.0848 \text{ u} \end{aligned}$$

We should report the weighted-average atomic mass of Si as 28.08 u.

6B (M) We let x be the fractional abundance of lithium-6.

$$6.941 \text{ u} = [x \times 6.01512 \text{ u}] + [(1-x) \times 7.01600 \text{ u}] = x \times 6.01512 \text{ u} + 7.01600 \text{ u} - x \times 7.01600 \text{ u}$$

$$6.941 \text{ u} - 7.01600 \text{ u} = x \times 6.01512 \text{ u} - x \times 7.01600 \text{ u} = -x \times 1.00088 \text{ u}$$

$$x = \frac{6.941 \text{ u} - 7.01600 \text{ u}}{-1.00088 \text{ u}} = 0.075 \text{ Percent isotopic abundances: } 7.5\% \text{ lithium-6, } 92.5\% \text{ lithium-7}$$

7A (M) We assume that atoms lose or gain relatively few electrons to become ions. Thus, elements that will form cations will be on the left-hand side of the periodic table, while elements that will form anions will be on the right-hand side. The number of electrons “lost” when a cation forms is usually equal to the last digit of the periodic group number; the number of electrons added when an anion forms is typically eight minus the last digit of the group number.

Li is in group 1(1A); it should form a cation by losing one electron: Li^+ .

S is in group 6(6A); it should form an anion by adding two electrons: S^{2-} .

Ra is in group 2(2A); it should form a cation by losing two electrons: Ra^{2+} .

F and I are both group 17(7A); they should form anions by gaining an electron: F^- and I^- .

Al is in group 13(3A); it should form a cation by losing three electrons: Al^{3+} .

7B (M) Main-group elements are in the “A” families, while transition elements are in the “B” families. Metals, nonmetals, metalloids, and noble gases are color coded in the periodic table inside the front cover of the textbook.

Na is a main-group metal in group 1(1A).

Re is a transition metal in group 7(7B).

S is a main-group nonmetal in group 16(6A).

I is a main-group nonmetal in group 17(7A).

Kr is a nonmetal in group 18(8A).

Mg is a main-group metal in group 2(2A).

U is an inner transition metal, an actinide.

Si is a main-group metalloid in group 14(4A).

B is a metalloid in group 13(3A).

Al is a main-group metal in group 13(3A).

As is a main-group metalloid in group 15(5A).

H is a main-group nonmetal in group 1(1A). (H is believed to be a metal at extremely high pressures.)

8A (E) This is similar to Practice Examples 2-8A and 2-8B.

$$\text{Cu mass} = 2.35 \times 10^{24} \text{ Cu atoms} \times \frac{1 \text{ mol Cu}}{6.022 \times 10^{23} \text{ atoms}} \times \frac{63.546 \text{ g Cu}}{1 \text{ mol Cu}} = 248 \text{ g Cu}$$

8B (M) Of all lead atoms, 24.1% are lead-206, or 241 ^{206}Pb atoms in every 1000 lead atoms. First we need to convert a 22.6 gram sample of lead into moles of lead (below) and then, by using Avogadro’s constant, and the percent isotopic abundance, we can determine the number of ^{206}Pb atoms.

$$n_{\text{Pb}} = 22.6 \text{ g Pb} \times \frac{1 \text{ mol Pb}}{207.2 \text{ g Pb}} = 0.109 \text{ mol Pb}$$

$$^{206}\text{Pb atoms} = 0.109 \text{ mol Pb} \times \frac{6.022 \times 10^{23} \text{ Pb atoms}}{1 \text{ mol Pb}} \times \frac{241 \text{ }^{206}\text{Pb atoms}}{1000 \text{ Pb atoms}} = 1.58 \times 10^{22} \text{ }^{206}\text{Pb atoms}$$

9A (M) Both the density and the molar mass of Pb serve as conversion factors.

$$\text{atoms of Pb} = 0.105 \text{ cm}^3 \text{ Pb} \times \frac{11.34 \text{ g}}{1 \text{ cm}^3} \times \frac{1 \text{ mol Pb}}{207.2 \text{ g}} \times \frac{6.022 \times 10^{23} \text{ Pb atoms}}{1 \text{ mol Pb}} = 3.46 \times 10^{21} \text{ Pb atoms}$$

9B (M) First we find the number of rhenium atoms in 0.100 mg of the element.

$$0.100 \text{ mg} \times \frac{1 \text{ g}}{1000 \text{ mg}} \times \frac{1 \text{ mol Re}}{186.21 \text{ g Re}} \times \frac{6.022 \times 10^{23} \text{ Re atoms}}{1 \text{ mol Re}} = 3.23 \times 10^{17} \text{ Re atoms}$$

$$\% \text{ abundance } ^{187}\text{Re} = \frac{2.02 \times 10^{17} \text{ atoms } ^{187}\text{Re}}{3.23 \times 10^{17} \text{ Re atoms}} \times 100\% = 62.5\%$$

INTEGRATIVE EXAMPLE

A. (M)*Stepwise approach:*

First, determine the total number of Cu atoms in the crystal by determining the volume of the crystal and calculating the mass of Cu from density. Then we can determine the amount of ^{63}Cu by noting its relative abundance

$$\text{Volume of crystal} = (25 \text{ nm})^3 \times \frac{1 \text{ cm}^3}{(1 \times 10^7 \text{ nm})^3} = 1.5625 \times 10^{-17} \text{ cm}^3$$

$$\text{Mass of Cu in crystal} = d \cdot V = 8.92 \text{ g/cm}^3 \times 1.5625 \times 10^{-17} = 1.3938 \times 10^{-16} \text{ g}$$

of Cu atoms =

$$1.3938 \times 10^{-16} \text{ g Cu} \times \frac{1 \text{ mol Cu}}{63.546 \text{ g Cu}} \times \frac{6.022 \times 10^{23} \text{ Cu atoms}}{1 \text{ mol Cu}} = 1.3208 \times 10^6 \text{ Cu atoms}$$

Therefore, the number of ^{63}Cu atoms, assuming 69.17% abundance, is 9.14×10^5 atoms.

Conversion pathway approach:

$$\begin{aligned} & \frac{8.92 \text{ g Cu}}{1 \text{ cm}^3} \times \frac{1 \text{ cm}^3}{(1 \times 10^7 \text{ nm})^3} \times \frac{(25 \text{ nm})^3}{\text{crystal}} \times \frac{1 \text{ mol Cu}}{63.546 \text{ g Cu}} \times \frac{6.022 \times 10^{23} \text{ Cu atoms}}{1 \text{ mol Cu}} \\ & \times \frac{69.17 \text{ atoms of } ^{63}\text{Cu}}{100 \text{ Cu atoms}} = 9.14 \times 10^5 \text{ atoms of } ^{63}\text{Cu} \end{aligned}$$

B. (M)*Stepwise approach:*

Calculate the mass of Fe in a serving of cereal, determine mass of ^{58}Fe in that amount of cereal, and determine how many servings of cereal are needed to reach 58 g of ^{58}Fe .

$$\text{Amount of Fe in a serving of cereal} = 18 \text{ mg} \times 0.45 = 8.1 \text{ mg Fe per serving}$$

First calculate the amount of Fe

$$0.0081 \text{ g Fe} \times \frac{1 \text{ mol Fe}}{55.845 \text{ g Fe}} = 1.45 \times 10^{-4} \text{ mol Fe}$$

Then calculate ^{58}Fe amount:

$$1.45 \times 10^{-4} \text{ mol Fe} \times \frac{0.282 \text{ mol } ^{58}\text{Fe}}{100 \text{ mol Fe}} = 4.090 \times 10^{-7} \text{ mol } ^{58}\text{Fe}$$

Converting mol of ^{58}Fe to # of servings:

$$\frac{4.090 \times 10^{-7} \text{ mol } ^{58}\text{Fe}}{1 \text{ serving}} \times \frac{57.9333 \text{ g } ^{58}\text{Fe}}{1 \text{ mol } ^{58}\text{Fe}} = 2.37 \times 10^{-5} \text{ g } ^{58}\text{Fe per serving}$$

$$\text{Total \# of servings} = 58 \text{ g total} / 2.37 \times 10^{-5} \text{ per serving} = 2.4477 \times 10^6 \text{ servings}$$

Conversion Pathway Approach:

The number of servings of dry cereal to ingest 58 g of ^{58}Fe =

$$58.0 \text{ g } ^{58}\text{Fe} \times \frac{1 \text{ mol } ^{58}\text{Fe}}{57.9333 \text{ g } ^{58}\text{Fe}} \times \frac{100 \text{ mol Fe}}{0.282 \text{ mol } ^{58}\text{Fe}} \times \frac{58.845 \text{ g Fe}}{1 \text{ mol Fe}} \times \frac{1 \text{ cereal serving}}{0.018 \text{ g Fe} \times 0.45}$$

$$= 2.4477 \times 10^6 \text{ servings}$$

$$2.44477 \times 10^6 \text{ servings} \times \frac{1 \text{ year}}{365 \text{ servings}} = 6706 \text{ years}$$

EXERCISES

Law of Conservation of Mass

- 1.** (E) The observations cited do not necessarily violate the law of conservation of mass. The oxide formed when iron rusts is a solid and remains with the solid iron, increasing the mass of the solid by an amount equal to the mass of the oxygen that has combined. The oxide formed when a match burns is a gas and will not remain with the solid product (the ash); the mass of the ash thus is less than that of the match. We would have to collect all reactants and all products and weigh them to determine if the law of conservation of mass is obeyed or violated.
- 2.** (E) The magnesium that is burned in air combines with some of the oxygen in the air and this oxygen (which, of course, was not weighed when the magnesium metal was weighed) adds its mass to the mass of the magnesium, making the magnesium oxide product weigh more than did the original magnesium. When this same reaction is carried out in a flashbulb, the oxygen (in fact, some excess oxygen) that will combine with the magnesium is already present in the bulb before the reaction. Consequently, the product contains no unweighed oxygen.
- 3.** (E) By the law of conservation of mass, all of the magnesium initially present and all of the oxygen that reacted are present in the product. Thus, the mass of oxygen that has reacted is obtained by difference. mass of oxygen = 0.674 g MgO – 0.406 g Mg = 0.268 g oxygen
- 4.** (E) Reaction: $2 \text{ K(s)} + \text{Cl}_2(\text{g}) \rightarrow 2 \text{ KCl(s)}$
Mass of Cl_2 reacted = 8.178 g – 6.867 g = 1.311 g $\text{Cl}_2(\text{g})$

$$m_{\text{KCl}} = 1.311 \text{ g } \text{Cl}_2 \times \frac{1 \text{ mol } \text{Cl}_2}{70.90 \text{ g } \text{Cl}_2} \times \frac{2 \text{ mol } \text{KCl}}{1 \text{ mol } \text{Cl}_2} \times \frac{74.55 \text{ g } \text{KCl}}{1 \text{ mol } \text{KCl}} = 2.757 \text{ g } \text{KCl}$$

- 5. (M)** We need to compare the mass before reaction (initial) with that after reaction (final) to answer this question.

$$\text{initial mass} = 10.500 \text{ g calcium hydroxide} + 11.125 \text{ g ammonium chloride} = 21.625 \text{ g}$$

$$\text{final mass} = 14.336 \text{ g solid residue} + (69.605 - 62.316) \text{ g of gases} = 21.625 \text{ g}$$

These data support the law of conservation of mass. Note that the gain in the mass of water is equal to the mass of gas absorbed by the water.

- 6. (M)** We compute the mass of the reactants and compare that with the mass of the products to answer this question.

$$\text{reactant mass} = \text{mass of calcium carbonate} + \text{mass of hydrochloric acid solution}$$

$$= 10.00 \text{ g calcium carbonate} + 100.0 \text{ mL soln} \times \frac{1.148 \text{ g}}{1 \text{ mL soln}}$$

$$= 10.00 \text{ g calcium carbonate} + 114.8 \text{ g solution} = 124.8 \text{ g reactants}$$

$$\text{product mass} = \text{mass of solution} + \text{mass of carbon dioxide}$$

$$= 120.40 \text{ g soln} + 2.22 \text{ L gas} \times \frac{1.9769 \text{ g}}{1 \text{ L gas}}$$

$$= 120.40 \text{ g soln} + 4.39 \text{ g carbon dioxide}$$

$$= 124.79 \text{ g products} \left(\begin{array}{l} \text{The same mass within experimental error,} \\ \text{thus, the law of conservation of mass obeyed.} \end{array} \right)$$

Law of Constant Composition

- 7. (E)**

(a) Ratio of O: MgO by mass = $\frac{(0.755 - 0.455) \text{ g}}{0.755 \text{ g}} = 0.397$

(b) Ratio of O: Mg in MgO by mass = $\frac{0.300 \text{ g}}{0.455 \text{ g}} = 0.659$

(c) Percent magnesium by mass = $\frac{0.455 \text{ g Mg}}{0.755 \text{ g MgO}} \times 100\% = 60.3\%$

- 8. (M)**

- (a) We can determine that carbon dioxide has a fixed composition by finding the % C in each sample. (In the calculations below, the abbreviation “cmpd” is short for compound.)

$$\%C = \frac{3.62 \text{ g C}}{13.26 \text{ g cmpd}} \times 100\% = 27.3\% \text{ C} \quad \%C = \frac{5.91 \text{ g C}}{21.66 \text{ g cmpd}} \times 100\% = 27.3\% \text{ C}$$

$$\%C = \frac{7.07 \text{ g C}}{25.91 \text{ g cmpd}} \times 100\% = 27.3\% \text{ C}$$

Since all three samples have the same percent of carbon, these data do establish that carbon dioxide has a fixed composition.

- (b) Carbon dioxide contains only carbon and oxygen. As determined in (a), carbon dioxide is 27.3% C by mass. The percent of oxygen in carbon dioxide is obtained by difference.
 $\%O = 100.0\% - (27.3\%C) = 72.7\%O$

- 9. (M)** In the first experiment, 2.18 g of sodium produces 5.54 g of sodium chloride. In the second experiment, 2.10 g of chlorine produces 3.46 g of sodium chloride. The amount of sodium contained in this second sample of sodium chloride is given by

$$\text{mass of sodium} = 3.46 \text{ g sodium chloride} - 2.10 \text{ g chlorine} = 1.36 \text{ g sodium.}$$

We now have sufficient information to determine the % Na in each of the samples of sodium chloride.

$$\%Na = \frac{2.18 \text{ g Na}}{5.54 \text{ g cmpd}} \times 100\% = 39.4\% Na \quad \%Na = \frac{1.36 \text{ g Na}}{3.46 \text{ g cmpd}} \times 100\% = 39.3\% Na$$

Thus, the two samples of sodium chloride have the same composition. Recognize that, based on significant figures, each percent has an uncertainty of $\pm 0.1\%$.

- 10. (E)** If the two samples of water have the same % H, the law of constant composition is demonstrated. Notice that, in the second experiment, the mass of the compound is equal to the sum of the masses of the elements produced from it.

$$\%H = \frac{3.06 \text{ g H}}{27.35 \text{ g H}_2\text{O}} \times 100\% = 11.2\% H \quad \%H = \frac{1.45 \text{ g H}}{(1.45 + 11.51) \text{ g H}_2\text{O}} \times 100\% = 11.2\% H$$

Thus, the results are consistent with the law of constant composition.

- 11. (E)** The mass of sulfur (0.312 g) needed to produce 0.623 g sulfur dioxide provides the information required for the conversion factor.

$$\text{sulfur mass} = 0.842 \text{ g sulfur dioxide} \times \frac{0.312 \text{ g sulfur}}{0.623 \text{ g sulfur dioxide}} = 0.422 \text{ g sulfur}$$

- 12. (M)**

- (a) From the first experiment we see that 1.16 g of compound is produced per gram of Hg. These masses enable us to determine the mass of compound produced from 1.50 g Hg.

$$\text{mass of cmpd} = 1.50 \text{ g Hg} \times \frac{1.16 \text{ g cmpd}}{1.00 \text{ g Hg}} = 1.74 \text{ g cmpd}$$

- (b) Since the compound weighs 0.24 g more than the mass of mercury (1.50 g) that was used, 0.24 g of sulfur must have reacted. Thus, the unreacted sulfur has a mass of 0.76 g (= 1.00 g initially present – 0.24 g reacted).

Law of Multiple Proportions

- 13. (M)** By dividing the mass of the oxygen per gram of sulfur in the second sulfur-oxygen compound (compound 2) by the mass of oxygen per gram of sulfur in the first sulfur-oxygen compound (compound 1), we obtain the ratio (shown to the right):

$$\frac{\frac{1.497 \text{ g of O}}{1.000 \text{ g of S}} \text{ (cmpd 2)}}{\frac{0.998 \text{ g of O}}{1.000 \text{ g of S}} \text{ (cmpd 1)}} = \frac{1.500}{1}$$

To get the simplest whole number ratio we need to multiply both the numerator and the denominator by 2. This gives the simple whole number ratio 3/2. In other words, for a given mass of sulfur, the mass of oxygen in the second compound (SO_3) relative to the mass of oxygen in the first compound (SO_2) is in a ratio of 3:2. These results are entirely consistent with the Law of Multiple Proportions because the same two elements, sulfur and oxygen in this case, have reacted together to give two different compounds that have masses of oxygen that are in the ratio of small positive integers for a fixed amount of sulfur.

- 14. (M)** This question is similar to question 13 in that two elements, phosphorus and chlorine in this case, have combined to give two different compounds. This time, however, different masses have been used for both of the elements in the second compound. To see if the Law of Multiple Proportions is being followed, the mass of one of the two elements must be set to the same value in both compounds. This can be achieved by dividing the masses of both phosphorus and chlorine in reaction 2 by 2.500:

$$\text{“normalized” mass of phosphorus} = \frac{2.500 \text{ g phosphorus}}{2.500} = 1.000 \text{ g of phosphorus}$$

$$\text{“normalized” mass of chlorine} = \frac{14.308 \text{ g chlorine}}{2.500} = 5.723 \text{ g of chlorine}$$

Now the mass of phosphorus for both reactions is fixed at 1.000 g. Next, we will divide each amount of chlorine by the fixed mass of phosphorus with which they are combined. This gives

$$\frac{\frac{3.433 \text{ g of Cl}}{1.000 \text{ g P}} \text{ (cmpd 1)}}{\frac{5.723 \text{ g of Cl}}{1.000 \text{ g P}} \text{ (cmpd 2)}} = 0.600 = 6 : 10 \text{ or } 3 : 5$$

- 15. (M)**
(a) First of all we need to fix the mass of nitrogen in all three compounds to some common value, for example, 1.000 g. This can be accomplished by multiplying the masses of hydrogen and nitrogen in compound A by 2 and the amount of hydrogen and nitrogen in compound C by 4/3 (1.333):

$$\begin{aligned} \text{Cmpd A} \quad \text{“normalized” mass of nitrogen} &= 0.500 \text{ g N} \times 2 = 1.000 \text{ g N} \\ &\text{“normalized” mass of hydrogen} = 0.108 \text{ g H} \times 2 = 0.216 \text{ g H} \end{aligned}$$

$$\begin{aligned} \text{Cmpd C} \quad \text{“normalized” mass of nitrogen} &= 0.750 \text{ g N} \times 1.333 = 1.000 \text{ g N} \\ &\text{“normalized” mass of hydrogen} = 0.108 \text{ g H} \times 1.333 = 0.144 \text{ g H} \end{aligned}$$

Next, we divide the mass of hydrogen in each compound by the smallest mass of hydrogen, namely, 0.0720 g. This gives 3.000 for compound A, 1.000 for compound B, and 2.000 for compound C. The ratio of the amounts of hydrogen in the three compounds is 3 (cmpd A) : 1 (cmpd B) : 2 (cmpd C)

These results are consistent with the Law of Multiple Proportions because the masses of hydrogen in the three compounds end up in a ratio of small whole numbers when the mass of nitrogen in all three compounds is normalized to a simple value (1.000 g here).

- (b) The text states that compound B is N_2H_2 . This means that, based on the relative amounts of hydrogen calculated in part (a), compound A might be N_2H_6 and compound C, N_2H_4 . Actually, compound A is NH_3 , but we have no way of knowing this from the data. Note that the H:N ratios in NH_3 and N_2H_6 are the same, 3H:1N.

16. (M)

- (a) As with the previous problem, one of the two elements must have the same mass in all of the compounds. This can be most readily achieved by setting the mass of iodine in all four compounds to 1.000 g. With this approach we only need to manipulate the data for compounds B and C. To normalize the amount of iodine in compound B to 1.000 g, we need to multiply the masses of both iodine and fluorine by 2. To accomplish the analogous normalization of compound C, we must multiply by $4/3$ (1.333).

$$\begin{aligned} \text{Cmpd B: "normalized" mass of iodine} &= 0.500 \text{ g I} \times 2 &= 1.000 \text{ g I} \\ \text{"normalized" mass of fluorine} &= 0.2246 \text{ g F} \times 2 &= 0.4492 \text{ g F} \end{aligned}$$

$$\begin{aligned} \text{Cmpd C: "normalized" mass of iodine} &= 0.750 \text{ g I} \times 1.333 = 1.000 \text{ g I} \\ \text{"normalized" mass of fluorine} &= 0.5614 \text{ g F} \times 1.333 = 0.7485 \text{ g F} \end{aligned}$$

Next we divide the mass of fluorine in each compound by the smallest mass of fluorine, namely, 0.1497 g. This gives 1.000 for compound A, 3.001 for compound B, 5.000 for compound C, and 7.001 for compound D. The ratios of the amounts of fluorine in the four compounds A : B : C : D is 1 : 3 : 5 : 7. These results are consistent with the law of multiple proportions because for a fixed amount of iodine (1.000 g), the masses of fluorine in the four compounds are in the ratio of small whole numbers.

- (b) As with the preceding problem, we can figure out the empirical formulas for the four iodine-fluorine containing compounds from the ratios of the amounts of fluorine that were determined in 16(a): Cmpd A: IF Cmpd B: IF_3 Cmpd C: IF_5 Cmpd D: IF_7

- 17. (M)** One oxide of copper has about 20% oxygen by mass. If we assume a 100 gram sample, then ~ 20 grams of the sample is oxygen (~1.25 moles) and 80 grams is copper (~1.26 moles). This would give an empirical formula of CuO (copper(II) oxide). The second oxide has less oxygen by mass, hence the empirical formula must have less oxygen or more copper (Cu:O ratio greater than 1). If we keep whole number ratios of atoms, a plausible formula would be Cu_2O (copper(I) oxide), where the mass percent oxygen is $\approx 11\%$.

- 18. (M)** Assuming the intermediate is “half-way” between CO (oxygen-carbon mass ratio = 16:12 or 1.333) and CO₂ (oxygen-carbon mass ratio = 32:12 or 2.6667), then the oxygen-carbon ratio would be 2:1, or O:C = 24:12. This mass ratio gives a mole ratio of O:C = 1.5:1. Empirical formulas are simple whole number ratios of elements; hence, a formula of C₃O₂ must be the correct empirical formula for this carbon oxide. (Note: C₃O₂ is called tricarbon dioxide or carbon suboxide).

Fundamental Charges and Mass-to-Charge Ratios

- 19. (M)** We can calculate the charge on each drop, express each in terms of 10⁻¹⁹ C, and finally express each in terms of $e = 1.6 \times 10^{-19}$ C.

$$\begin{aligned} \text{drop 1:} & \quad 1.28 \times 10^{-18} & = 12.8 \times 10^{-19} \text{ C} & = 8e \\ \text{drops 2 \& 3:} & \quad 1.28 \times 10^{-18} \div 2 = 0.640 \times 10^{-18} \text{ C} & = 6.40 \times 10^{-19} \text{ C} & = 4e \\ \text{drop 4:} & \quad 1.28 \times 10^{-18} \div 8 = 0.160 \times 10^{-18} \text{ C} & = 1.60 \times 10^{-19} \text{ C} & = 1e \\ \text{drop 5:} & \quad 1.28 \times 10^{-18} \times 4 = 5.12 \times 10^{-18} \text{ C} & = 51.2 \times 10^{-19} \text{ C} & = 32e \end{aligned}$$

We see that these values are consistent with the charge that Millikan found for that of the electron, and he could have inferred the correct charge from these data, since they are all multiples of e .

- 20. (M)** We calculate each drop's charge, express each in terms of 10⁻¹⁹ C, and then, express each in terms of $e = 1.6 \times 10^{-19}$ C.

$$\begin{aligned} \text{drop 1:} & \quad 6.41 \times 10^{-19} \text{ C} & = 6.41 \times 10^{-19} \text{ C} & = 4e \\ \text{drop 2:} & \quad 6.41 \times 10^{-19} \div 2 & = 3.21 \times 10^{-19} \text{ C} & = 3.21 \times 10^{-19} \text{ C} & = 2e \\ \text{drop 3:} & \quad 6.41 \times 10^{-19} \times 2 & = 1.28 \times 10^{-18} \text{ C} & = 12.8 \times 10^{-19} \text{ C} & = 8e \\ \text{drop 4:} & \quad 1.44 \times 10^{-18} & = 14.4 \times 10^{-19} \text{ C} & & = 9e \\ \text{drop 5:} & \quad 1.44 \times 10^{-18} \div 3 & = 4.8 \times 10^{-19} \text{ C} & = 4.8 \times 10^{-19} \text{ C} & = 3e \end{aligned}$$

We see that these values are consistent with the charge that Millikan found for that of the electron. He could have inferred the correct charge from these values, since they are all multiples of e , and have no other common factor.

- 21. (M)**
(a) Determine the ratio of the mass of a hydrogen atom to that of an electron. We use the mass of a proton plus that of an electron for the mass of a hydrogen atom.

$$\begin{aligned} \frac{\text{mass of proton} + \text{mass of electron}}{\text{mass of electron}} &= \frac{1.0073 \text{ u} + 0.00055 \text{ u}}{0.00055 \text{ u}} = 1.8 \times 10^3 \\ \text{or} \quad \frac{\text{mass of electron}}{\text{mass of proton} + \text{mass of electron}} &= \frac{1}{1.8 \times 10^3} = 5.6 \times 10^{-4} \end{aligned}$$

- (b) The only two mass-to-charge ratios that we can determine from the data in Table 2-1 are those for the proton (a hydrogen ion, H^+) and the electron.

$$\text{For the proton : } \frac{\text{mass}}{\text{charge}} = \frac{1.673 \times 10^{-24} \text{ g}}{1.602 \times 10^{-19} \text{ C}} = 1.044 \times 10^{-5} \text{ g/C}$$

$$\text{For the electron : } \frac{\text{mass}}{\text{charge}} = \frac{9.109 \times 10^{-28} \text{ g}}{1.602 \times 10^{-19} \text{ C}} = 5.686 \times 10^{-9} \text{ g/C}$$

The hydrogen ion is the lightest positive ion available. We see that the mass-to-charge ratio for a positive particle is considerably larger than that for an electron.

22. (M) We do not have the precise isotopic masses for the two ions. The values of the mass-to-charge ratios are only approximate. This is because some of the mass is converted to energy (binding energy), that holds all of the positively charged protons in the nucleus together. Consequently, we have used a three-significant-figure mass for a nucleon, rather than the more precisely known proton and neutron masses. (Recall that the term “nucleon” refers to a nuclear particle— either a proton or a neutron.)

$${}^{127}\text{I}^- \quad \frac{m}{e} = \frac{127 \text{ nucleons}}{1 \text{ electron}} \times \frac{1 \text{ electron}}{1.602 \times 10^{-19} \text{ C}} \times \frac{1.67 \times 10^{-24} \text{ g}}{1 \text{ nucleon}} = 1.32 \times 10^{-3} \text{ g/C} \quad (7.55 \times 10^2 \text{ C/g})$$

$${}^{32}\text{S}^{2-} \quad \frac{m}{e} = \frac{32 \text{ nucleons}}{2 \text{ electrons}} \times \frac{1 \text{ electron}}{1.602 \times 10^{-19} \text{ C}} \times \frac{1.67 \times 10^{-24} \text{ g}}{1 \text{ nucleon}} = 1.67 \times 10^{-4} \text{ g/C} \quad (6.00 \times 10^3 \text{ C/g})$$

Atomic Number, Mass Number, and Isotopes

23. (E) (a) cobalt-60 ${}^{60}_{27}\text{Co}$ (b) phosphorus-32 ${}^{32}_{15}\text{P}$ (c) iron-59 ${}^{59}_{26}\text{Fe}$ (d) radium-226 ${}^{226}_{88}\text{Ra}$

24. (E) The nucleus of ${}^{202}_{80}\text{Hg}$ contains 80 protons and $(202 - 80) = 122$ neutrons.

Thus, the percent of nucleons that are neutrons is given by

$$\% \text{ neutrons} = \frac{122 \text{ neutrons}}{202 \text{ nucleons}} \times 100 = 60.4\% \text{ neutrons}$$

25.	Name	Symbol	Number of Protons	Number of Electrons	Number of Neutrons	Mass Number
(E)	Sodium	${}^{23}_{11}\text{Na}$	11	11	12	23
	Silicon	${}^{28}_{14}\text{Si}$	14	14 ^a	14	28
	Rubidium	${}^{85}_{37}\text{Rb}$	37	37 ^a	48	85
	Potassium	${}^{40}_{19}\text{K}$	19	19	21	40
	Arsenic ^a	${}^{75}_{33}\text{As}$	33 ^a	33	42	75
	Neon	${}^{20}_{10}\text{Ne}^{2+}$	10	8	10	20
	Bromine ^b	${}^{80}_{35}\text{Br}$	35	35	45	80
	Lead ^b	${}^{208}_{82}\text{Pb}$	82	82	126	208

^a This result assumes that a neutral atom is involved.

^b Insufficient data. Does not characterize a specific nuclide; several possibilities exist. The minimum information needed is the atomic number (or some way to obtain it, such as from the name or the symbol of the element involved), the number of electrons (or some way to obtain it, such as the charge on the species), and the mass number (or the number of neutrons).

- 26. (E)**
- (a) Since all of these species are neutral atoms, the numbers of electrons are the atomic numbers, the subscript numbers. The symbols must be arranged in order of increasing value of these subscripts. ${}_{18}^{40}\text{Ar} < {}_{19}^{39}\text{K} < {}_{27}^{58}\text{Co} < {}_{29}^{59}\text{Cu} < {}_{48}^{120}\text{Cd} < {}_{50}^{112}\text{Sn} < {}_{52}^{122}\text{Te}$
- (b) The number of neutrons is given by the difference between the mass number and the atomic number, $A-Z$. This is the difference between superscripted and subscripted values and is provided (in parentheses) after each element in the following list.
 ${}_{19}^{39}\text{K}(20) < {}_{18}^{40}\text{Ar}(22) < {}_{29}^{59}\text{Cu}(30) < {}_{27}^{58}\text{Co}(31) < {}_{50}^{112}\text{Sn}(62) < {}_{52}^{122}\text{Te}(70) < {}_{48}^{120}\text{Cd}(72)$
- (c) Here the nuclides are arranged by increasing mass number, given by the superscripts.
 ${}_{19}^{39}\text{K} < {}_{18}^{40}\text{Ar} < {}_{27}^{58}\text{Co} < {}_{29}^{59}\text{Cu} < {}_{50}^{112}\text{Sn} < {}_{48}^{120}\text{Cd} < {}_{52}^{122}\text{Te}$
- 27. (E)**
- (a) A ${}^{108}\text{Pd}$ atom has 46 protons and 46 electrons. The atom described is neutral, hence, the number of electrons must equal the number of protons.
- Since there are 108 nucleons in the nucleus, the number of neutrons is 62 (= 108 nucleons – 46 protons).
- (b) The ratio of the two masses is determined as follows: $\frac{{}^{108}\text{Pd}}{{}^{12}\text{C}} = \frac{107.90389\text{ u}}{12\text{ u}} = 8.9919908$
- 28. (E)**
- (a) The atomic number of Ra is 88 and equals the number of protons in the nucleus. The ion's charge is +2 and, thus, there are two more protons than electrons: no. protons = no. electrons + 2 = 88; no. electrons = 88 – 2 = 86. The mass number (228) is the sum of the atomic number and the number of neutrons: 228 = 88 + no. neutrons; Hence, the number of neutrons = 228 – 88 = 140 neutrons.
- (b) The mass of ${}^{16}\text{O}$ is 15.9949 u. $\text{ratio} = \frac{\text{mass of isotope}}{\text{mass of } {}^{16}\text{O}} = \frac{228.030\text{ u}}{15.9949\text{ u}} = 14.2564$
- 29. (E)** The mass of ${}^{16}\text{O}$ is 15.9949 u. $\text{isotopic mass} = 15.9949\text{ u} \times 6.68374 = 106.936\text{ u}$

- 30. (E)** The mass of ^{16}O is 15.9949 u.
 mass of heavier isotope = $15.9949\text{ u} \times 7.1838 = 114.904\text{ u} = \text{mass of } ^{115}\text{In}$
 mass of lighter isotope = $\frac{114.904\text{ u}}{1.0177} = 112.91\text{ u} = \text{mass of } ^{113}\text{In}$
- 31. (E)** Each isotopic mass must be divided by the isotopic mass of ^{12}C , 12 u, an exact number.
(a) $^{35}\text{Cl} \div ^{12}\text{C} = 34.96885\text{ u} \div 12\text{ u} = 2.914071$
(b) $^{26}\text{Mg} \div ^{12}\text{C} = 25.98259\text{ u} \div 12\text{ u} = 2.165216$
(c) $^{222}\text{Rn} \div ^{12}\text{C} = 222.0175\text{ u} \div 12\text{ u} = 18.50146$
- 32. (M)** We need to work through the mass ratios in sequence to determine the mass of ^{81}Br .
 mass of $^{19}\text{F} = \text{mass of } ^{12}\text{C} \times 1.5832 = 12\text{ u} \times 1.5832 = 18.998\text{ u}$
 mass of $^{35}\text{Cl} = \text{mass of } ^{19}\text{F} \times 1.8406 = 18.998\text{ u} \times 1.8406 = 34.968\text{ u}$
 mass of $^{81}\text{Br} = \text{mass of } ^{35}\text{Cl} \times 2.3140 = 34.968\text{ u} \times 2.3140 = 80.916\text{ u}$

- 33. (E)** First, we determine the number of protons, neutrons, and electrons in each species.

species:	$^{24}_{12}\text{Mg}^{2+}$	$^{47}_{24}\text{Cr}$	$^{60}_{27}\text{Co}^{3+}$	$^{35}_{17}\text{Cl}^{-}$	$^{124}_{50}\text{Sn}^{2+}$	$^{226}_{90}\text{Th}$	$^{90}_{38}\text{Sr}$
protons:	12	24	27	17	50	90	38
neutrons:	12	23	33	18	74	136	52
electrons:	10	24	24	18	48	90	38

- (a)** The numbers of neutrons and electrons are equal for $^{35}_{17}\text{Cl}^{-}$.
- (b)** $^{60}_{27}\text{Co}^{3+}$ has protons (27), neutrons (33), and electrons (24) in the ratio 9:11:8.
- (c)** The species $^{124}_{50}\text{Sn}^{2+}$ has a number of neutrons (74) equal to its number of protons (50) plus one-half its number of electrons ($48 \div 2 = 24$).
- 34. (a)** Atoms with equal numbers of protons and neutrons will have mass numbers that are approximately twice the size of their atomic numbers. The following species are approximately suitable (with numbers of protons and neutrons in parentheses).
 $^{24}_{12}\text{Mg}^{2+}$ (12 p⁺, 12 n), $^{47}_{24}\text{Cr}$ (24 p⁺, 23 n), $^{60}_{27}\text{Co}^{3+}$ (27 p⁺, 33 n), and $^{35}_{17}\text{Cl}^{-}$ (17 p⁺, 18 n).
 Of these four nuclides, only $^{24}_{12}\text{Mg}^{2+}$ has equal numbers of protons and neutrons.
- (b)** A species in which protons have more than 50% of the mass must have a mass number smaller than twice the atomic number. Of these species, only in $^{47}_{24}\text{Cr}$ is more than 50% of the mass contributed by the protons.
- (c)** A species with about 50% more neutrons than protons will have a mass number that is at least 2.5 times greater than the atomic number. Only $^{226}_{90}\text{Th}$ has >50% more neutrons than protons (number of neutrons = $226 - 90 = 136$; number of protons = 90; $90 \times 1.5 = 135$, less than the total number of neutrons). $^{124}_{50}\text{Sn}$ is close, having 74 neutrons, and 50 protons, so $50 \times 1.5 = 75$, slightly less than 50%.

35. (E)

If we let n represent the number of neutrons and p represent the number of protons, then $p + 4 = n$. The mass number is the sum of the number of protons and the number of neutrons: $p + n = 44$. Substitution of $n = p + 4$ yields $p + p + 4 = 44$. From this relation, we see $p = 20$. Reference to the periodic table indicates that 20 is the atomic number of the element calcium.

36. (M) We will use the same type of strategy and the same notation as we used previously in Equation 35 to come up with the answer.

$$n = p + 1$$

There is one more neutron than the number of protons.

$$n + p = 9 \times 3 = 27$$

The mass number equals nine times the ion's charge of 3+.

Substitute the first relationship into the second, and solve for p .

$$27 = (p + 1) + p = 2p + 1$$

$$p = \frac{27 - 1}{2} = 13$$

Thus this is the 3+ action of the isotope Al-27 $\rightarrow {}_{13}^{27}\text{Al}^{3+}$.

37. (M) The number of protons is the same as the atomic number for iodine, which is 53. There is one more electron than the number of protons because there is a negative charge on the ion. Therefore the number of electrons is 54. The number of neutrons is equal to 70, mass number minus atomic number.**38. (E)** For iodine, $Z = 53$, and so it has 53 protons. Because the overall charge on the ion is 1-, there are 54 electrons in a single ion of iodine-131. The number of neutrons is $131 - 53 = 78$.**39. (E)** For americium, $Z = 95$. There are 95 protons, 95 electrons, and $241 - 95 = 146$ neutrons in a single atom of americium-241.**40. (E)** For cobalt, $Z = 27$. There are 27 protons, 27 electrons, and $60 - 27 = 33$ neutrons in a single atom of cobalt-60.**Atomic Mass Units, Atomic Masses****41. (E)** There are no chlorine atoms that have a mass of 35.45 u. The masses of individual chlorine atoms are close to integers and this mass is about midway between two integers. It is an average atomic mass, the result of averaging two (or more) isotopic masses, each weighted by its isotopic abundance.

42. (E) It is exceedingly unlikely that another nuclide would have an exact integral mass. The mass of carbon-12 is *defined* as precisely 12 u. Each nuclidic mass is close to integral, but none that we have encountered in this chapter are precisely integral. The reason is that each nuclide is composed of protons, neutrons, and electrons, none of which have integral masses, and there is a small quantity of the mass of each nucleon (nuclear particle) lost in the binding energy holding the nuclides together. It would be highly unlikely that all of these contributions would add up to a precisely integral mass.

43. (E) To determine the weighted-average atomic mass, we use the following expression:
 average atomic mass = \sum (isotopic mass \times fractional isotopic abundance)
 Each of the three percents given is converted to a fractional abundance by dividing it by 100.
 Mg atomic mass = $(23.985042 \text{ u} \times 0.7899) + (24.985837 \text{ u} \times 0.1000) + (25.982593 \text{ u} \times 0.1101)$
 $= 18.95 \text{ u} + 2.499 \text{ u} + 2.861 \text{ u} = 24.31 \text{ u}$

44. (E) To determine the average atomic mass, we use the following expression:
 average atomic mass = \sum (isotopic mass \times fractional isotopic abundance)
 Each of the three percents given is converted to a fractional abundance by dividing it by 100.
 Cr atomic mass = $(49.9461 \times 0.0435) + (51.9405 \times 0.8379) + (52.9407 \times 0.0950) + (53.9389 \times 0.0236)$
 $= 2.17 \text{ u} + 43.52 \text{ u} + 5.03 \text{ u} + 1.27 \text{ u} = 51.99 \text{ u}$
 If all digits are carried and then the answer is rounded at the end, the answer is 52.00 u.

45. (E) We will use the expression to determine the weighted-average atomic mass.
 $107.868 \text{ u} = (106.905092 \text{ u} \times 0.5184) + ({}^{109}\text{Ag} \times 0.4816) = 55.42 \text{ u} + 0.4816 {}^{109}\text{Ag}$
 $107.868 \text{ u} - 55.42 \text{ u} = 0.4816 {}^{109}\text{Ag} = 52.45 \text{ u} \quad {}^{109}\text{Ag} = \frac{52.45 \text{ u}}{0.4816} = 108.9 \text{ u}$

46. (M) The percent abundance of the two isotopes must add to 100.00%, since there are only two naturally occurring isotopes of gallium. Thus, we can determine the percent abundance of the second isotope by using difference:

$$\text{second isotope} = 100.00\% - 60.11\% = 39.89\%$$

From the periodic table, we see that the atomic mass of gallium is 69.723 u. We use this value in the expression to determine the weighted-average atomic mass, along with the isotopic mass of ${}^{69}\text{Ga}$ and the fractional abundances of the two isotopes (the percent abundances divided by 100).

$$69.723 \text{ u} = (0.6011 \times 68.925581 \text{ u}) + (0.3989 \times \text{other isotope}) = 41.43 \text{ u} + (0.3989 \times \text{other isotope})$$

$$\text{other isotope} = \frac{69.723 - 41.43 \text{ u}}{0.3989} = 70.92 \text{ u} = \text{mass of } {}^{71}\text{Ga}, \text{ the other isotope}$$

- 47. (M)** Since the three percent abundances total 100%, the percent abundance of ^{40}K is found by difference. $\% ^{40}\text{K} = 100.0000\% - 93.2581\% - 6.7302\% = 0.0117\%$

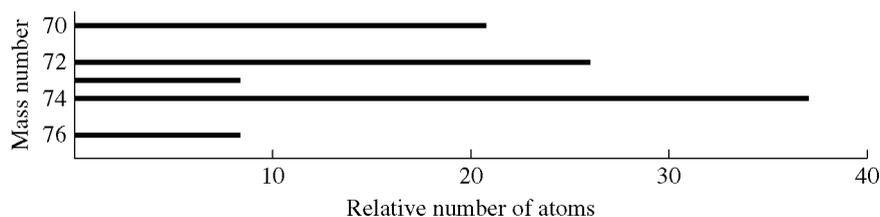
Then the expression for the weighted-average atomic mass is used, with the percent abundances converted to fractional abundances by dividing by 100. Note that the average atomic mass of potassium is 39.0983 u.

$$\begin{aligned} 39.0983 \text{ u} &= (0.932581 \times 38.963707 \text{ u}) + (0.000117 \times 39.963999 \text{ u}) + (0.067302 \times \text{mass of } ^{41}\text{K}) \\ &= 36.3368 \text{ u} + 0.00468 \text{ u} + (0.067302 \times \text{mass of } ^{41}\text{K}) \\ \text{mass of } ^{41}\text{K} &= \frac{39.0983 \text{ u} - (36.3368 \text{ u} + 0.00468 \text{ u})}{0.067302} = 40.962 \text{ u} \end{aligned}$$

- 48. (M)** We use the expression for determining the weighted-average atomic mass, where x represents the fractional abundance of ^{10}B and $(1-x)$ the fractional abundance of ^{11}B
- $$10.81 \text{ u} = (10.012937 \text{ u} \times x) + [11.009305 \times (1-x)] = 10.012937x + 11.009305 - 11.009305x$$
- $$10.81 - 11.009305 = -0.20 = 10.012937x - 11.009305x = -0.996368x$$
- $$x = \frac{0.20}{0.996368} = 0.20$$
- $\therefore 20\% ^{10}\text{B}$ and $(100.0 - 20) = 80\% ^{11}\text{B}$

Mass Spectrometry

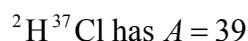
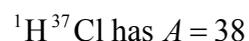
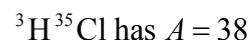
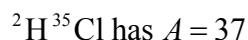
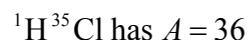
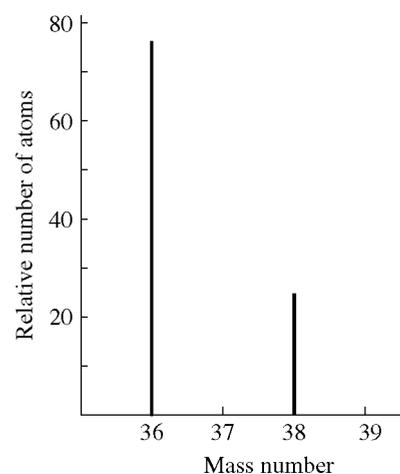
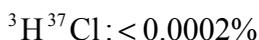
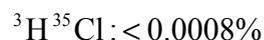
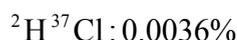
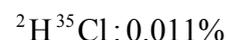
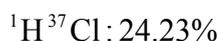
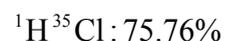
- 49. (M)**
(a)



- (b)** As before, we multiply each isotopic mass by its fractional abundance, after which, we sum these products to obtain the (average) atomic mass for the element.
- $$(0.205 \times 70) + (0.274 \times 72) + (0.078 \times 73) + (0.365 \times 74) + (0.078 \times 76)$$
- $$= 14 + 20 + 5.7 + 27 + 5.9 = 72.6 = \text{average atomic mass of germanium}$$
- The result is only approximately correct because the isotopic masses are given to only two significant figures. Thus, only a two-significant-figure result can be quoted.

50. (M)**(a)** Six unique HCl molecules are possible (called isotopomers):

The mass numbers of the six different possible types of molecules are obtained by summing the mass numbers of the two atoms in each molecule:

**(b)** The most abundant molecule contains the isotope for each element that is most abundant. It is ${}^1\text{H}^{35}\text{Cl}$. The second most abundant molecule is ${}^1\text{H}^{37}\text{Cl}$. The relative abundance of each type of molecule is determined by multiplying together the fractional abundances of the two isotopes present. Relative abundances of the molecules are as follows.**The Periodic Table****51. (E)****(a)** Ge is in group 14 and in the fourth period.**(b)** Other elements in group 16(6A) are similar to S: O, Se, and Te. Most of the elements in the periodic table are unlike S, but particularly metals such as Na, K, and Rb.**(c)** The alkali metal (group 1), in the fifth period is Rb.**(d)** The halogen (group 17) in the sixth period is At.**52. (E)****(a)** Au is in group 11 and in the sixth period.**(b)** Ar, $Z = 18$, is a noble gas. Xe is a noble gas with atomic number (54) greater than 50.**(c)** If an element forms a stable anion with charge $2-$, it is in group 16.**(d)** If an element forms a stable cation with charge $3+$, it is in group 13.

- 53. (E)** If the seventh period of the periodic table is 32 members long, it will be the same length as the sixth period. Elements in the same family (vertical group), will have atomic numbers 32 units higher. The noble gas following radon will have atomic number = $86 + 32 = 118$. The alkali metal following francium will have atomic number = $87 + 32 = 119$.
- 54. (M)** There are several interchanges: Ar/K, Co/Ni, Te/I, Th/Pa, U/Np, Pu/Am, Sg/Bh. The reverse order is necessary because the periodic table lists elements in order of increasing atomic number (protons in the nucleus) and not in order of increasing atomic masses.

The Avogadro Constant and the Mole

55. (E)

$$(a) \text{ atoms of Fe} = 15.8 \text{ mol Fe} \times \frac{6.022 \times 10^{23} \text{ atoms Fe}}{1 \text{ mol Fe}} = 9.51 \times 10^{24} \text{ atoms Fe}$$

$$(b) \text{ atoms of Ag} = 0.000467 \text{ mol Ag} \times \frac{6.022 \times 10^{23} \text{ atoms Ag}}{1 \text{ mol Ag}} = 2.81 \times 10^{20} \text{ atoms Ag}$$

$$(c) \text{ atoms of Na} = 8.5 \times 10^{-11} \text{ mol Na} \times \frac{6.022 \times 10^{23} \text{ atoms Na}}{1 \text{ mol Na}} = 5.1 \times 10^{13} \text{ atoms Na}$$

- 56. (E)** Since the molar mass of nitrogen is 14.0 g/mol, 25.0 g N is almost two moles (1.79 mol N), while 6.022×10^{23} Ni atoms is about one mole, and 52.0 g Cr (52.00 g/mol Cr) is also almost one mole. Finally, 10.0 cm³ Fe (55.85 g/mol Fe) has a mass of about 79 g, and contains about 1.4 moles of atoms. Thus, 25.0 g N contains the greatest number of atoms.

Note: Even if you take nitrogen as N₂, the answer is the same.

57. (E)

$$(a) \text{ moles of Zn} = 415.0 \text{ g Zn} \times \frac{1 \text{ mol Zn}}{65.38 \text{ g Zn}} = 6.347 \text{ mol Zn}$$

$$(b) \text{ \# of Cr atoms} = 147,400 \text{ g Cr} \times \frac{1 \text{ mol Cr}}{51.996 \text{ g Cr}} \times \frac{6.022 \times 10^{23} \text{ atoms Cr}}{1 \text{ mol Cr}}$$

$$= 1.707 \times 10^{27} \text{ atoms Cr}$$

$$(c) \text{ mass Au} = 1.0 \times 10^{12} \text{ atoms Au} \times \frac{1 \text{ mol Au}}{6.022 \times 10^{23} \text{ atoms Au}} \times \frac{196.97 \text{ g Au}}{1 \text{ mol Au}} = 3.3 \times 10^{-10} \text{ g Au}$$

$$(d) \text{ mass of F atom} = 1 \text{ atom F} \times \frac{1 \text{ mol F}}{6.022 \times 10^{23} \text{ atoms F}} \times \frac{18.998 \text{ g F}}{1 \text{ mol F}} = 3.1547 \times 10^{-23} \text{ g F}$$

For exactly 1 F atom, the number of sig figs in the answer is determined by the least precise number in the calculation, namely the mass of fluorine.

58. (E)

$$\begin{aligned} \text{(a) number Kr atoms} &= 5.25 \text{ mg Kr} \times \frac{1 \text{ g Kr}}{1000 \text{ mg Kr}} \times \frac{1 \text{ mol Kr}}{83.80 \text{ g Kr}} \times \frac{6.022 \times 10^{23} \text{ atoms Kr}}{1 \text{ mol Kr}} \\ &= 3.77 \times 10^{19} \text{ atoms Kr} \end{aligned}$$

(b) Molar mass is defined as the mass per mole of substance. mass = 2.09 g.

This calculation requires that the number of moles be determined.

$$\text{moles} = 2.80 \times 10^{22} \text{ atoms} \times \frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ atoms}} = 0.0465 \text{ mol}$$

$$\text{molar mass} = \frac{\text{mass}}{\text{moles}} = \frac{2.09 \text{ g}}{0.0465 \text{ mol}} = 44.9 \text{ g/mol} \quad \text{The element is Sc, scandium.}$$

$$\text{(c) mass P} = 44.75 \text{ g Mg} \times \frac{1 \text{ mol Mg}}{24.305 \text{ g Mg}} \times \frac{1 \text{ mol P}}{1 \text{ mol Mg}} \times \frac{30.9738 \text{ g P}}{1 \text{ mol P}} = 57.03 \text{ g P}$$

Note: The same answer is obtained if you assume phosphorus is P₄ instead of P.

59. (E) Determine the mass of Cu in the jewelry, then convert to moles and finally to the number of atoms. If sterling silver is 92.5% by mass Ag, it is 100 – 92.5 = 7.5% by mass Cu.

Conversion pathway approach:

$$\begin{aligned} \text{number of Cu atoms} &= 33.24 \text{ g sterling} \times \frac{7.5 \text{ g Cu}}{100.0 \text{ g sterling}} \times \frac{1 \text{ mol Cu}}{63.546 \text{ g Cu}} \times \frac{6.022 \times 10^{23} \text{ atoms Cu}}{1 \text{ mol Cu}} \\ &= 2.4 \times 10^{22} \text{ Cu atoms} \end{aligned}$$

Stepwise approach:

$$33.24 \text{ g sterling} \times \frac{7.5 \text{ g Cu}}{100.0 \text{ g sterling}} = 2.493 \text{ g Cu}$$

$$2.493 \text{ g Cu} \times \frac{1 \text{ mol Cu}}{63.546 \text{ g Cu}} = 0.03923 \text{ mol Cu}$$

$$0.03923 \text{ mol Cu} \times \frac{6.022 \times 10^{23} \text{ atoms Cu}}{1 \text{ mol Cu}} = 2.4 \times 10^{22} \text{ Cu atoms}$$

60. (E) We first need to determine the amount in moles of each metal.

$$\text{amount of Pb} = 50.0 \text{ cm}^3 \text{ solder} \times \frac{9.4 \text{ g solder}}{1 \text{ cm}^3} \times \frac{67 \text{ g Pb}}{100 \text{ g solder}} \times \frac{1 \text{ mol Pb}}{207.2 \text{ g Pb}} = 1.5 \text{ mol Pb}$$

$$\text{amount of Sn} = 50.0 \text{ cm}^3 \text{ solder} \times \frac{9.4 \text{ g solder}}{1 \text{ cm}^3} \times \frac{33 \text{ g Sn}}{100 \text{ g solder}} \times \frac{1 \text{ mol Sn}}{118.7 \text{ g Sn}} = 1.3 \text{ mol Sn}$$

$$\text{total atoms} = (1.5 \text{ mol Pb} + 1.3 \text{ mol Sn}) \times \frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol}} = 1.7 \times 10^{24} \text{ atoms}$$

- 61. (E)** We first need to determine the number of Pb atoms of all types in 215 mg of Pb, and then use the percent isotopic abundance to determine the number of ^{204}Pb atoms present.

$$\begin{aligned} {}^{204}\text{Pb atoms} &= 215 \text{ mg Pb} \times \frac{1 \text{ g}}{1000 \text{ mg}} \times \frac{1 \text{ mol Pb}}{207.2 \text{ g Pb}} \times \frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol Pb}} \times \frac{14 \text{ }^{204}\text{Pb atoms}}{1000 \text{ Pb atoms}} \\ &= 8.7 \times 10^{18} \text{ atoms } ^{204}\text{Pb} \end{aligned}$$

- 62. (E)**

$$\begin{aligned} \text{mass of alloy} &= 7.25 \times 10^{23} \text{ Cd atoms} \times \frac{1 \text{ mol Cd}}{6.022 \times 10^{23} \text{ Cd atoms}} \times \frac{112.4 \text{ g Cd}}{1 \text{ mol Cd}} \times \frac{100.0 \text{ g alloy}}{8.0 \text{ g Cd}} \\ &= 1.7 \times 10^3 \text{ g alloy} \end{aligned}$$

- 63. (E)** We will use the average atomic mass of lead, 207.2 g/mol, to answer this question.

$$\text{(a)} \quad \frac{30 \mu\text{g Pb}}{1 \text{ dL}} \times \frac{1 \text{ dL}}{0.1 \text{ L}} \times \frac{1 \text{ g Pb}}{10^6 \mu\text{g Pb}} \times \frac{1 \text{ mol Pb}}{207.2 \text{ g}} = 1.45 \times 10^{-6} \text{ mol Pb / L}$$

$$\text{(b)} \quad \frac{1.45 \times 10^{-6} \text{ mol Pb}}{\text{L}} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol}} = 8.7 \times 10^{14} \text{ Pb atoms / mL}$$

- 64. (M)** The concentration of Pb in air provides the principal conversion factor. Other conversion factors are needed to convert to and from its units, beginning with the 0.500 L volume, and ending with the number of atoms.

$$\begin{aligned} \text{no. Pb atoms} &= 0.500 \text{ L} \times \frac{1 \text{ m}^3}{1000 \text{ L}} \times \frac{3.11 \mu\text{g Pb}}{1 \text{ m}^3} \times \frac{1 \text{ g Pb}}{10^6 \mu\text{g Pb}} \times \frac{1 \text{ mol Pb}}{207.2 \text{ g Pb}} \times \frac{6.022 \times 10^{23} \text{ Pb atoms}}{1 \text{ mol Pb}} \\ &= 4.52 \times 10^{12} \text{ Pb atoms} \end{aligned}$$

- 65. (M)** To answer this question, we simply need to calculate the ratio of the mass (in grams) of each sample to its molar mass. Whichever elemental sample gives the largest ratio will be the one that has the greatest number of atoms.

$$\text{(a)} \quad \text{Iron sample: } 10 \text{ cm} \times 10 \text{ cm} \times 10 \text{ cm} \times 7.86 \text{ g cm}^{-3} = 7860 \text{ g Fe}$$

$$7860 \text{ g Fe} \times \frac{1 \text{ mol Fe}}{55.845 \text{ g Fe}} = 141 \text{ mol of Fe atoms}$$

$$\text{(b)} \quad \text{Hydrogen sample: } \frac{1.00 \times 10^3 \text{ g H}_2}{2 \times (1.008 \text{ g H})} \times 1 \text{ mol H} = 496 \text{ mol of H}_2 \text{ molecules} =$$

$$992 \text{ mol of H atoms}$$

$$\text{(c)} \quad \text{Sulfur sample: } \frac{2.00 \times 10^4 \text{ g S}}{32.06 \text{ g S}} \times 1 \text{ mol S} = 624 \text{ mol of S atoms}$$

$$\text{(d)} \quad \text{Mercury sample: } 76 \text{ lb Hg} \times \frac{454 \text{ g Hg}}{1 \text{ lb Hg}} \times \frac{1 \text{ mol Hg}}{200.6 \text{ g Hg}} = 172 \text{ mol of Hg atoms}$$

Clearly, then, it is the 1.00 kg sample of hydrogen that contains the greatest number of atoms.

66. (M)
- (a) 23 g Na = 1 mol with a density $\sim 1 \text{ g/cm}^3$. 1 mole = 23 g, so volume of 25.5 mol $\sim 600 \text{ cm}^3$.
 - (b) Liquid bromine occupies 725 mL or 725 cm^3 (given).
 - (c) 1.25×10^{25} atoms Cr is ~ 20 moles. At $\sim 50 \text{ g/mol}$, this represents approximately 1000 g. Given the density of 9.4 g/cm^3 , this represents about 100 cm^3 of volume.
 - (d) 2150 g solder at 9.4 g/cm^3 represents approximately 200 cm^3 .

From this we can see that the liquid bromine would occupy the largest volume.

INTEGRATIVE AND ADVANCED EXERCISES

67. (M)

$$(a) \text{ total mass}(40^\circ\text{C}) = 2.50 \text{ g} + 100.0 \text{ mL} \times \frac{0.9922 \text{ g}}{1 \text{ mL}} = 2.50 \text{ g} + 99.22 \text{ g} = 101.72 \text{ g}$$

$$\text{mass of solution}(20^\circ\text{C}) = 100 \text{ mL} \times \frac{1.0085 \text{ g}}{1 \text{ mL}} = 100.85 \text{ g}$$

$$\begin{aligned} \text{solid crystallized} &= \text{total mass}(40^\circ\text{C}) - \text{solution mass}(20^\circ\text{C}) \\ &= 101.72 \text{ g} - 100.85 \text{ g} = 0.87 \text{ g} \end{aligned}$$

- (b) The answer cannot be more precise because both the initial mass and the subtraction only allows the reporting of masses to $\pm 0.01 \text{ g}$. For a more precise answer, more significant figures would be required for the initial mass (2.50 g) and the densities of the water and solution.

68. (M) We now recognize that the values of 24.3 u and 35.3 u for the masses of Mg and Cl represent a weighted-average that considers the mass and abundance of each isotope. The experimental mass of each isotope is very close to the natural number and therefore very close to an integer multiple of the mass of ^1H , thus supporting Prout's hypothesis.

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.

$$\begin{aligned} \text{Total mass} &= \left(9 \text{ protons} \times \frac{1.0073 \text{ u}}{1 \text{ proton}} \right) + \left(10 \text{ neutrons} \times \frac{1.0087 \text{ u}}{1 \text{ neutron}} \right) + \left(9 \text{ electrons} \times \frac{0.0005486 \text{ u}}{1 \text{ electron}} \right) \\ &= 9.0657 \text{ u} + 10.087 \text{ u} + 0.004937 \text{ u} = 19.158 \text{ u} \end{aligned}$$

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.

70. (M)

$$\text{volume of nucleus(single proton)} = \frac{4}{3} \pi r^3 = 1.3333 \times 3.14159 \times (0.5 \times 10^{-13} \text{ cm})^3 = 5 \times 10^{-40} \text{ cm}^3$$

$$\text{density} = \frac{1.673 \times 10^{-24} \text{ g}}{5 \times 10^{-40} \text{ cm}^3} = 3 \times 10^{15} \text{ g/cm}^3$$

- 71. (M)** This method of establishing the Avogadro constant uses fundamental constants.

$$1.000 \text{ g } ^{12}\text{C} \times \frac{1 \text{ mol } ^{12}\text{C}}{12 \text{ g } ^{12}\text{C}} \times \frac{6.022 \times 10^{23} \text{ } ^{12}\text{C atoms}}{1 \text{ mol } ^{12}\text{C}} \times \frac{12 \text{ u}}{1 \text{ } ^{12}\text{C atom}} = 6.022 \times 10^{23} \text{ u}$$

- 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. $E + 2 = 1.100 E$

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

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.

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

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

- 73. (E)** Because the net ionic charge (+2) is one-tenth of its the nuclear charge, the nuclear charge is +20. This is also the atomic number of the nuclide, which means the element is calcium. The number of electrons is 20 for a neutral calcium atom, but only 18 for a calcium ion with a net 2+ charge. Four more than 18 is 22 neutrons. The ion described is $^{42}_{20}\text{Ca}^{2+}$.

- 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.

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

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

$$E = 2.50 Z = 2.50 \times 84 = 210 \quad \text{The isotope E is } ^{210}\text{Po}.$$

75. As a result of the redefinition, all masses will decrease by a factor of $35.00000/35.453 = 0.98722$.

(a) atomic mass of He $4.00260 \times 0.98722 = 3.9515$
 atomic mass of Na $22.9898 \times 0.98722 = 22.696$
 atomic mass of I $126.905 \times 0.98722 = 125.28$

- (b) These three elements have nearly integral atomic masses based on C-12 because these three elements and C-12 all consist mainly of one stable isotope, rather than a mixture of two or more stable isotopes, with each being present in significant amounts (10% or more), as is the case with chlorine.

76. (M) To solve this question, represent the fractional abundance of ^{14}N by x and that of ^{15}N by $(1 - x)$. Then use the expression for determining average atomic mass.

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

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

$$x = \frac{0.9934}{0.9970} \times 100\% = 99.64\% = \text{percent isotopic abundance of } ^{14}\text{N}.$$

Thus, 0.36% = percent isotopic abundance of ^{15}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.

$$^{196}\text{Hg}: 195.9658 \text{ u} \times 0.00146 = 0.286 \text{ u}$$

$$^{198}\text{Hg}: 197.9668 \text{ u} \times 0.1002 = 19.84 \text{ u} \quad ^{199}\text{Hg}: 198.9683 \text{ u} \times 0.1684 = 33.51 \text{ u}$$

$$^{200}\text{Hg}: 199.9683 \text{ u} \times 0.2313 = 46.25 \text{ u} \quad ^{201}\text{Hg}: 200.9703 \text{ u} \times 0.1322 = 26.57 \text{ u}$$

$$^{202}\text{Hg}: 201.9706 \text{ u} \times 0.2980 = 60.19 \text{ u} \quad ^{204}\text{Hg}: 203.9735 \text{ u} \times 0.0685 = 14.0 \text{ u}$$

$$\text{Atomic weight} = 0.286 \text{ u} + 19.84 \text{ u} + 33.51 \text{ u} + 46.25 \text{ u} + 26.57 \text{ u} + 60.19 \text{ u} + 14.0 \text{ u} = 200.6 \text{ u}$$

78. (D) The sum of the percent abundances of the two minor isotopes equals $100.00\% - 84.68\% = 15.32\%$. Thus, we denote the fractional abundance of ^{73}Ge as x , and the other as $(0.1532 - x)$. These fractions are then used in the expression for average atomic mass.

$$\text{atomic mass} = 72.64 \text{ u} = (69.92425 \text{ u} \times 0.2085) + (71.92208 \text{ u} \times 0.2754) + (73.92118 \text{ u} \times 0.3629) \\ + (72.92346 \text{ u} \times x) + [75.92140 \text{ u} \times (0.1532 - x)]$$

$$72.64 \text{ u} = (14.5792 \text{ u}) + (19.8073 \text{ u}) + (26.8260 \text{ u}) + (72.92346 \text{ u} \times x) + [11.6312 \text{ u} - 75.92140 \text{ u} \times x]$$

$$-0.2037 = -2.99794x \quad \text{Hence: } x = 0.068.$$

^{73}Ge has 6.8% isotopic abundance and ^{76}Ge has 8.5% isotopic abundance.

From the calculations, we can see that the number of significant figures drops from four in the percent isotopic abundances supplied to only two significant figures owing to the imprecision in the supplied values of the percent isotopic abundance.

79. (D) First, it must be understood that because we 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:

- Assume percent abundance of $^{84}\text{Kr} \sim 55\%$ as a start (somewhat more than 50)
- Let percent abundance of $^{82}\text{Kr} = x\%$; percent abundance $^{83}\text{Kr} \sim ^{82}\text{Kr} = x\%$
- $^{86}\text{Kr} = 1.50(\text{percent abundance of } ^{82}\text{Kr}) = 1.50(x\%)$
- $^{80}\text{Kr} = 0.196(\text{percent abundance of } ^{82}\text{Kr}) = 0.196(x\%)$
- $^{78}\text{Kr} = 0.030(\text{percent abundance of } ^{82}\text{Kr}) = 0.030(x\%)$

$$100\% = \%^{78}\text{Kr} + \%^{80}\text{Kr} + \%^{82}\text{Kr} + \%^{83}\text{Kr} + \%^{84}\text{Kr} + \%^{86}\text{Kr}$$

$$100\% = 0.030(x\%) + 0.196(x\%) + x\% + x\% + \%^{84}\text{Kr} + 1.50(x\%)$$

$$100\% = 3.726(x\%) + \%^{84}\text{Kr}$$

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:

$$^{78}\text{Kr}: 0.03 \times 12.1 = 0.363\%; \quad ^{80}\text{Kr}: 0.196 \times 12.1 = 2.37\%; \quad ^{83}\text{Kr}: \text{same as } ^{82}\text{Kr}; \quad ^{86}\text{Kr}: 1.5 \times 12.1 = 18.15\%.$$

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

As stated above, the problem here is the inaccuracy of the percent abundance for ^{84}Kr , which is crudely estimated to be $\sim 55\%$. 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.

$$^{35}\text{Cl } ^{79}\text{Br mass} = 34.9689 \text{ u} + 78.9183 \text{ u} = 113.8872 \text{ u}$$

$$^{35}\text{Cl } ^{81}\text{Br mass} = 34.9689 \text{ u} + 80.9163 \text{ u} = 115.8852 \text{ u}$$

$$^{37}\text{Cl } ^{79}\text{Br mass} = 36.9658 \text{ u} + 78.9183 \text{ u} = 115.8841 \text{ u}$$

$$^{37}\text{Cl } ^{81}\text{Br mass} = 36.9658 \text{ u} + 80.9163 \text{ u} = 117.8821 \text{ u}$$

Each molecule has a different intensity pattern (relative number of molecules), based on the isotopic 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.

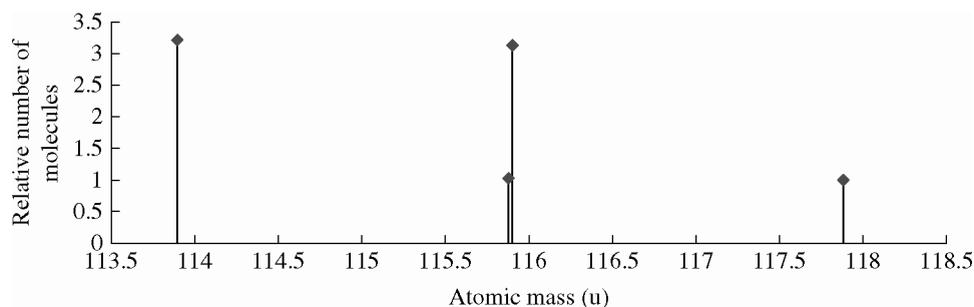
$$^{35}\text{Cl } ^{79}\text{Br Intensity} = (0.7577) \times (0.5069) = 0.3841 \div 0.1195 = 3.214$$

$$^{35}\text{Cl } ^{81}\text{Br Intensity} = (0.7577) \times (0.4931) = 0.3736 \div 0.1195 = 3.127$$

$$^{37}\text{Cl } ^{79}\text{Br Intensity} = (0.2423) \times (0.5069) = 0.1228 \div 0.1195 = 1.028$$

$$^{37}\text{Cl } ^{81}\text{Br Intensity} = (0.2423) \times (0.4931) = 0.1195 \div 0.1195 = 1.000$$

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.

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

$$\text{The radius is } 0.08118 \text{ cm} \times 1/2 = 0.04059 \text{ cm}$$

$$\text{The volume of Cu (cm}^3\text{)} = \pi r^2 h = 3.1416 \times (0.04059 \text{ cm})^2 \times 150 \text{ cm} = 0.7764 \text{ cm}^3$$

$$\text{So, the mass of Cu} = 0 \text{ cm}^3 \times \frac{8.92 \text{ g Cu}}{1 \text{ cm}^3} = 6.93 \text{ g Cu}$$

$$\text{The number of moles of Cu} = 6.93 \text{ g Cu} \times \frac{1 \text{ mol Cu}}{63.546 \text{ g Cu}} = 0.109 \text{ mol Cu}$$

$$\text{Cu atoms in the wire} = 0.109 \text{ mol Cu} \times \frac{6.022 \times 10^{23} \text{ atoms Cu}}{1 \text{ mol Cu}} = 6.56 \times 10^{22} \text{ atoms}$$

82. (D) volume = $l \times w \times h - \pi r^2 h$

$$\begin{aligned} &= (15.0 \text{ cm} \times 12.5 \text{ cm} \times 0.300 \text{ cm}) - (3.1416) \times \left(\frac{2.50 \text{ cm}}{2} \right)^2 \times (0.300 \text{ cm}) \\ &= (56.25 \text{ cm} - 1.47 \text{ cm}) = 54.8 \text{ cm}^3 \end{aligned}$$

$$\text{mass of object} = 54.8 \text{ cm}^3 \times \frac{8.80 \text{ g}}{1 \text{ cm}^3} = 482 \text{ g Monel metal}$$

Then determine the number of silicon atoms in this quantity of alloy.

$$482 \text{ g Monel metal} \times \frac{2.2 \times 10^{-4} \text{ g Si}}{1.000 \text{ g metal}} \times \frac{1 \text{ mol Si}}{28.05 \text{ g Si}} \times \frac{6.022 \times 10^{23} \text{ Si atoms}}{1 \text{ mol Si}} = 2.3 \times 10^{21} \text{ Si atoms}$$

Finally, determine the number of ^{30}Si atoms in this quantity of silicon.

$$\text{number of } ^{30}\text{Si atoms} = (2.3 \times 10^{21} \text{ Si atoms}) \times \left(\frac{3.10 \text{ } ^{30}\text{Si atoms}}{100 \text{ Si atoms}} \right) = 7.1 \times 10^{19} \text{ } ^{30}\text{Si}$$

83. (M) The percent isotopic abundance of deuterium of 0.015% means that, in a sample of 100,000 H atoms, only 15 ^2H are present.

$$\text{mass H}_2 = 2.50 \times 10^{21} \text{ atoms } ^2\text{H} \times \frac{100,000 \text{ H atoms}}{15 \text{ atoms } ^2\text{H}} \times \frac{1 \text{ mol H}}{6.022 \times 10^{23} \text{ H atoms}} \times \frac{1 \text{ mol H}_2}{2 \text{ mol H}} \times \frac{2.016 \text{ g H}_2}{1 \text{ mol H}_2}$$

mass $\text{H}_2 = 27.9 \text{ g H}_2$ is required such that we have 2.50×10^{21} atoms ^2H .

84. (M) The numbers sum to 21 (= 10 + 6 + 5). Thus, in one mole of the alloy there is $\frac{10}{21}$ mol Bi, $\frac{6}{21}$ mol Pb, and $\frac{5}{21}$ 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.

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

85. (M) The atom ratios are of course, the same as the mole ratios. We first determine the mass of alloy that contains 5.00 mol Ag, 4.00 mol Cu, and 1.00 mol Zn.

$$\begin{aligned} \text{mass} &= 5.00 \text{ mol Ag} \times \frac{107.87 \text{ g Ag}}{1 \text{ mol Ag}} + 4.00 \text{ mol Cu} \times \frac{63.546 \text{ g Cu}}{1 \text{ mol Cu}} + 1.00 \text{ mol Zn} \times \frac{65.38 \text{ g Zn}}{1 \text{ mol Zn}} \\ &= 539.4 \text{ g Ag} + 254.2 \text{ g Cu} + 65.38 \text{ g Zn} = 859 \text{ g alloy} \end{aligned}$$

Then, for 1.00 kg of the alloy, (1000 g), we need 1000/859 moles. (859 g is the alloy's "molar mass.")

$$\text{mass of Ag} = 1000 \text{ g alloy} \times \frac{539 \text{ g Ag}}{859 \text{ g alloy}} = 627 \text{ g Ag}$$

$$\text{mass of Cu} = 1000 \text{ g alloy} \times \frac{254 \text{ g Cu}}{859 \text{ g alloy}} = 296 \text{ g Cu}$$

$$\text{mass of Zn} = 1000 \text{ g alloy} \times \frac{65.4 \text{ g Zn}}{859 \text{ g alloy}} = 76.1 \text{ g Zn}$$

86. (M) The relative masses of Sn and Pb are 207.2 g Pb (assume one mole of Pb) to $(2.73 \times 118.71 \text{ g/mol Sn}) = 324 \text{ g Sn}$. Then the mass of cadmium, on the same scale, is $207.2/1.78 = 116 \text{ g Cd}$.

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

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

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

$$\text{Calculate the mass of zinc:} \quad 125 \text{ cm}^3 \times 7.13 \text{ g/cm}^3 = 891 \text{ g}$$

$$\text{Calculate the mass of iodine:} \quad 125 \text{ cm}^3 \times 4.93 \text{ g/cm}^3 = 616 \text{ g}$$

$$\text{Calculate the mass of zinc iodide:} \quad 164 \text{ cm}^3 \times 4.74 \text{ g/cm}^3 = 777 \text{ g}$$

$$\text{Calculate the mass of zinc unreacted:} \quad (891 + 616 - 777) \text{ g} = 730 \text{ g}$$

$$\text{Calculate the volume of zinc unreacted:} \quad 730 \text{ g} \times 1 \text{ cm}^3 / 7.13 \text{ g} = 102 \text{ cm}^3$$

88. (D) First, calculate the total number of Ag atoms in a 1 cm^3 crystal:

$$10.5 \text{ g Ag} \times \frac{1 \text{ mol Ag}}{107.87 \text{ g Ag}} \times \frac{6.02 \times 10^{23} \text{ Ag atoms}}{1 \text{ mol Ag}} = 5.86 \times 10^{22} \text{ atoms of Ag}$$

The actual volume taken up by the Ag atoms (considering that there is 26% empty space in the crystal) is: $1 \text{ cm}^3 \times 0.74 = 0.74 \text{ cm}^3$

$$\text{Therefore, the volume of each atom is:} \quad \frac{0.74 \text{ cm}^3}{5.86 \times 10^{22} \text{ atoms}} = 1.263 \times 10^{-23} \text{ cm}^3$$

$$\text{Volume of a sphere is expressed as } V = 1.263 \times 10^{-23} \text{ cm}^3 = (4/3) \pi r^3$$

Solving for r , we get $1.44 \times 10^{-8} \text{ cm}$ or 144 pm.

FEATURE PROBLEMS

- 89. (M)** The product mass differs from that of the reactants by $(5.62 - 2.50 =)$ 3.12 grains. In order to determine the percent gain in mass, we need to convert the reactant mass to grains.

$$13 \text{ onces} \times \frac{8 \text{ gros}}{1 \text{ once}} = 104 \text{ gros} \times (104 + 2) \text{ gros} \times \frac{72 \text{ grains}}{1 \text{ gros}} = 7632 \text{ grains}$$

$$\% \text{ mass increase} = \frac{3.12 \text{ grains increase}}{(7632 + 2.50) \text{ grains original}} \times 100\% = 0.0409\% \text{ mass increase}$$

The sensitivity of Lavoisier's balance can be as little as 0.01 grain, which seems to be the limit of the readability of the balance; alternatively, it can be as large as 3.12 grains, which assumes that all of the error in the experiment is due to the (in)sensitivity of the balance. Let us convert 0.01 grains to a mass in grams.

$$\text{minimum error} = 0.01 \text{ gr} \times \frac{1 \text{ gros}}{72 \text{ gr}} \times \frac{1 \text{ once}}{8 \text{ gros}} \times \frac{1 \text{ livre}}{16 \text{ once}} \times \frac{30.59 \text{ g}}{1 \text{ livre}} = 3 \times 10^{-5} \text{ g} = 0.03 \text{ mg}$$

$$\text{maximum error} = 3.12 \text{ gr} \times \frac{3 \times 10^{-5} \text{ g}}{0.01 \text{ gr}} = 9 \times 10^{-3} \text{ g} = 9 \text{ mg}$$

The maximum error is close to that of a common modern laboratory balance, which has a sensitivity of 1 mg. The minimum error is approximated by a good quality analytical balance. Thus we conclude that Lavoisier's results conform closely to the law of conservation of mass.

- 90. (D)** One way to determine the common factor of which all 13 numbers are multiples is to first divide all of them by the smallest number in the set. The ratios thus obtained may be either integers or rational numbers whose decimal equivalents are easy to recognize.

Obs.	1	2	3	4	5	6	7	8	9	10	11	12	13
Quan.	19.6	24.60	29.62	34.47	39.38	44.42	49.41	53.91	59.12	63.68	68.65	78.34	83.22
Ratio	1.00	1.251	1.507	1.753	2.003	2.259	2.513	2.742	3.007	3.239	3.492	3.984	4.233
Mult.	4.00	5.005	6.026	7.013	8.012	9.038	10.05	10.97	12.03	12.96	13.97	15.94	16.93
Int.	4	5	6	7	8	9	10	11	12	13	14	16	17

The row labeled "Mult." is obtained by multiplying the row "Ratio" by 4.000. In the row labeled "Int." we give the integer closest to each of these multipliers. It is obvious that each of the 13 measurements is exceedingly close to a common quantity multiplied by an integer.

- 91. (M)** In a 60-year-old chemistry textbook, the atomic mass for oxygen would be exactly 16 u because chemists assigned precisely 16 u as the atomic mass of the naturally occurring mixture of oxygen isotopes. This value is slightly higher than the value of 15.9994 in modern chemistry textbooks. Thus, we would expect all other atomic masses to be slightly higher as well in the older textbooks.
- 92. (D)** We begin with the amount of reparations and obtain the volume in cubic kilometers with a series of conversion factors.

Conversion pathway approach:

$$\begin{aligned}
 V &= \$28.8 \times 10^9 \times \frac{1 \text{ troy oz Au}}{\$21.25} \times \frac{31.103 \text{ g Au}}{1 \text{ troy oz Au}} \times \frac{1 \text{ mol Au}}{196.97 \text{ g Au}} \times \frac{6.022 \times 10^{23} \text{ atoms Au}}{1 \text{ mol Au}} \\
 &\times \frac{1 \text{ ton seawater}}{4.67 \times 10^{17} \text{ Au atoms}} \times \frac{2000 \text{ lb seawater}}{1 \text{ ton seawater}} \times \frac{453.6 \text{ g seawater}}{1 \text{ lb seawater}} \times \frac{1 \text{ cm}^3 \text{ seawater}}{1.03 \text{ g seawater}} \\
 &\times \left(\frac{1 \text{ m}}{100 \text{ cm}} \times \frac{1 \text{ km}}{1000 \text{ m}} \right)^3 = 2.43 \times 10^5 \text{ km}^3
 \end{aligned}$$

Stepwise approach:

$$\begin{aligned}
 &\$28.8 \times 10^9 \times \frac{1 \text{ troy oz Au}}{\$21.25} \times \frac{31.103 \text{ g Au}}{1 \text{ troy oz Au}} = 4.22 \times 10^{10} \text{ g Au} \\
 &4.22 \times 10^{10} \text{ g Au} \times \frac{1 \text{ mol Au}}{196.97 \text{ g Au}} \times \frac{6.022 \times 10^{23} \text{ atoms Au}}{1 \text{ mol Au}} = 1.29 \times 10^{32} \text{ atoms Au} \\
 &1.29 \times 10^{32} \text{ atoms Au} \times \frac{1 \text{ ton seawater}}{4.67 \times 10^{17} \text{ Au atoms}} \times \frac{2000 \text{ lb seawater}}{1 \text{ ton seawater}} = 5.52 \times 10^{17} \text{ lb seawater} \\
 &5.52 \times 10^{17} \text{ lb seawater} \times \frac{453.6 \text{ g seawater}}{1 \text{ lb seawater}} \times \frac{1 \text{ cm}^3 \text{ seawater}}{1.03 \text{ g seawater}} = 2.43 \times 10^{20} \text{ cm}^3 \text{ seawater} \\
 &2.43 \times 10^{20} \text{ cm}^3 \text{ seawater} \times \left(\frac{1 \text{ m}}{100 \text{ cm}} \times \frac{1 \text{ km}}{1000 \text{ m}} \right)^3 = 2.43 \times 10^5 \text{ km}^3
 \end{aligned}$$

- 93. (D)** We start by using the percent isotopic abundances for ^{87}Rb and ^{85}Rb along with the data in the “spiked” mass spectrum to find the total mass of Rb in the sample. Then, we calculate the Rb content in the rock sample in ppm by mass by dividing the mass of Rb by the total mass of the rock sample, and then multiplying the result by 10^6 to convert to ppm.

$$^{87}\text{Rb} = 27.83\% \text{ isotopic abundance} \quad ^{85}\text{Rb} = 72.17\% \text{ isotopic abundance}$$

$$\text{Therefore, } \frac{^{87}\text{Rb}(\text{isotopic})}{^{85}\text{Rb}(\text{isotopic})} = \frac{27.83\%}{72.17\%} = 0.3856$$

For the ^{87}Rb (spiked) sample, the ^{87}Rb peak in the mass spectrum is 1.12 times as tall as the ^{85}Rb peak. Thus, for this sample $\frac{^{87}\text{Rb}(\text{isotopic}) + ^{87}\text{Rb}(\text{spiked})}{^{85}\text{Rb}(\text{isotopic})} = 1.12$

Using this relationship, we can now find the masses of both ^{85}Rb and ^{87}Rb in the sample.

$$\text{So, } \frac{{}^{87}\text{Rb(isotopic)}}{{}^{85}\text{Rb(isotopic)}} = 0.3856; \quad {}^{85}\text{Rb(isotopic)} = \frac{{}^{87}\text{Rb(isotopic)}}{0.3856}$$

$${}^{87}\text{Rb(isotopic)} + {}^{87}\text{Rb(spiked)} = \frac{1.12 \times {}^{87}\text{Rb(isotopic)}}{0.3856} = 2.905 {}^{87}\text{Rb(isotopic)}$$

$${}^{87}\text{Rb(spiked)} = 1.905 {}^{87}\text{Rb(isotopic)}$$

$$\text{and } \frac{{}^{87}\text{Rb(isotopic)} + {}^{87}\text{Rb(spiked)}}{{}^{85}\text{Rb(isotopic)}} = \frac{{}^{87}\text{Rb(isotopic)} + {}^{87}\text{Rb(spiked)}}{{}^{87}\text{Rb(isotopic)}} = 1.12$$

Since the mass of ${}^{87}\text{Rb(spiked)}$ is equal to $29.45 \mu\text{g}$, the mass of ${}^{87}\text{Rb(isotopic)}$ must be $\frac{29.45 \mu\text{g}}{1.905} = 15.46 \mu\text{g}$ of ${}^{87}\text{Rb(isotopic)}$

$$\text{So, the mass of } {}^{85}\text{Rb(isotopic)} = \frac{15.46 \mu\text{g of } {}^{87}\text{Rb(isotopic)}}{0.3856} = 40.09 \mu\text{g of } {}^{85}\text{Rb(isotopic)}$$

Therefore, the total mass of Rb in the sample = $15.46 \mu\text{g}$ of ${}^{87}\text{Rb(isotopic)}$ + $40.09 \mu\text{g}$ of ${}^{85}\text{Rb(isotopic)}$ = $55.55 \mu\text{g}$ of Rb. Convert to grams:

$$= 55.55 \mu\text{g of Rb} \times \frac{1 \text{ g Rb}}{1 \times 10^6 \mu\text{g Rb}} = 5.555 \times 10^{-5} \text{ g Rb}$$

$$\text{Rb content (ppm)} = \frac{5.555 \times 10^{-5} \text{ g Rb}}{0.350 \text{ g of rock}} \times 10^6 = 159 \text{ ppm Rb}$$

SELF-ASSESSMENT EXERCISES

94. (E)

(a) ${}^A_Z\text{E}$: The element "E" with the atomic number of Z (i.e., Z protons in the nucleus) and atomic mass of A (i.e., total of protons and neutrons equals A).

(b) β particle: An electron produced as a result of the decay of a neutron

(c) Isotope: Nuclei that share the same atomic number but have different atomic masses

(d) ${}^{16}\text{O}$: An oxygen nucleus containing 8 neutrons

(e) Molar mass: Mass of one mole of a substance

- 95. (E)**
- (a) The total mass of substances present after the chemical reaction is the same as the total mass of substances before the chemical reaction. More universally, mass is neither created nor destroyed, but converts from one form to another.
 - (b) Rutherford's model of the atom postulates the existence of positively charged fundamental particles at the nucleus of the atom.
 - (c) An average value used to express the atomic mass of an element by taking into account the atomic masses and relative abundances of all the naturally occurring isotopes of the element.
 - (d) A spectrum showing the mass/charge ratio of various atoms in a matrix
- 96. (E)**
- (a) Cathode rays are beams of electrons being generated from a negatively charged surface (cathode) moving to a positively charged surface (anode). X-rays are high energy photons, which are typically generated when the high energy beam of electrons impinges on the anode.
 - (b) Protons and neutrons are both fundamental particles that make up an atom's nucleus. Protons are positively charged and neutrons have no charge.
 - (c) Nuclear charge is determined by the numbers of protons in the nucleus. Ionic charge is determined by the difference between the number of protons and electrons in the atom.
 - (d) Periods are horizontal rows in the periodic table, while groups are vertical columns.
 - (e) Metals are generally characterized by their malleability, ductility, and ability to conduct electricity and heat well. Nonmetals are generally brittle and nonconductive.
 - (f) The Avogadro constant is the number of elementary entities (atoms, molecules, etc.) in one mole of a substance. A mole is the amount of a substance that contains the same number of elementary entities as there are atoms in exactly 12 g of pure carbon-12.
- 97. (E)** $\left(\frac{1}{5} \times 10.013 \text{ u}\right) + \left(\frac{4}{5} \times 11.009 \text{ u}\right) = 10.810 \text{ u}$. Therefore, the element is boron.
- 98. (E)** The answer is (b). If all of the zinc reacts and the total amount of the product (zinc sulfide) is 14.9 g, then 4.9 g of S must have reacted with zinc. Therefore, 3.1 g of S remain.
- 99. (E)** The answer is (d). It should be remembered that atoms combine in ratios of whole numbers. Therefore:
- (a) $16 \text{ g O} \times (1 \text{ mol O}/16 \text{ g O}) = 1 \text{ mol O}$, and $85.5 \text{ g Rb} \times (1 \text{ mol Rb}/85.5 \text{ g Rb}) = 1 \text{ mol Rb}$. Therefore, the O: Rb ratio is 1:1.
 - (b) Same calculations as above give an O: Rb ratio of 0.5:0.5, or 1:1.
 - (c) Same type calculation gives an O: Rb ratio of 2:1.

Because all of the above combine in O and Rb in whole number ratios, they are all at least theoretically possible.

- 100. (E)** If the compound contains 46.7% X, then the percent of oxygen would be $100.0\% - 46.7 = 53.3\%$. Assuming the sample is 100.0 g, there would be 53.3 g of oxygen present.

$$53.3 \text{ g O} \left(\frac{1 \text{ mol O}}{16.00 \text{ g O}} \right) = 3.331 \text{ mol O}$$

The formula of the compound is XO; therefore, X and O are in a 1:1 mole ratio.

$$\frac{46.7 \text{ g X}}{3.331 \text{ mol X}} = 14.0 \text{ g/mol X}$$

- 101. (E)** The answer is (c). Cathode rays are beams of electrons, and as such have identical properties to β particles, although they may not have the same energy.
- 102. (E)** The answer is (a), that the greatest portion of the mass of an atom is concentrated in a small but positively charged nucleus.
- 103. (E)** The answer is (d). A hydrogen atom has one proton and one electron, so its charge is zero. A neutron has the same charge as a proton, but is neutral. Since most of the mass of the atom is at the nucleus, a neutron has nearly the same mass as a hydrogen atom.
- 104. (E)** All the choices in this question are fundamental particles (e).
- 105. (E)** Dalton (d) is correct.
- 106. (E)** There are no particles that have the same mass as the hydrogen atom and a negative charge. The correct answer is (e).
- 107. (E)** ${}_{17}^{35}\text{Cl}^+$
- 108. (E)** The answer is (d), calcium, because they are in the same group.
- 109. (E)** (a) Group 18, (b) Group 17, (c) Group 13 and Group 1 (d) Group 18
- 110. (E)** (d) and (f)
- 111. (E)** (c), because it is not close to being a whole number
- 112. (M)** The answer is (d). Even with the mass scale being redefined based on ${}^{84}\text{Xe}$, the mass ratio between ${}^{12}\text{C}$ and ${}^{84}\text{Xe}$ will remain the same. Using ${}^{12}\text{C}$ as the original mass scale, the mass ratio of ${}^{12}\text{C} : {}^{84}\text{Xe}$ is $12 \text{ u} / 83.9115 \text{ u} = 0.1430$. Therefore, redefining the mass scale by assigning the exact mass of 84 u to ${}^{84}\text{Xe}$, the relative mass of ${}^{12}\text{C}$ becomes $84 \times 0.14301 = 12.0127 \text{ u}$.

113. (M) The answer is (b)

$$5.585 \text{ kg Fe} \times \frac{1000 \text{ g}}{1 \text{ kg}} \times \frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}} = 1.000 \times 10^{-2} \text{ mol Fe}$$

$$600.6 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 50.01 \text{ mol C}$$

Therefore, 100 moles of Fe has twice as many atoms as 50 moles of C.

114. (E) (d) is correct.

$$91.84 \text{ g} \times \frac{\text{mol}}{47.867 \text{ g}} \times \frac{6.022 \times 10^{23} \text{ atoms}}{\text{mol}} \times \frac{22 \text{ electrons}}{\text{Ti atom}} = 2.542 \times 10^{25} \text{ electrons}$$

115. (M)

$$2.327 \text{ g Fe} \times \frac{1 \text{ mol Fe}}{55.8452 \text{ g Fe}} = 0.0417 \text{ mol Fe}$$

$$1.000 \text{ g O} \times \frac{1 \text{ mol O}}{15.999 \text{ g O}} = 0.0625 \text{ mol O}$$

Dividing the two mole values to obtain the mole ratio, we get: $0.0625/0.0417 = 1.50$. That is, 1.50 moles (or atoms) of O per 1 mole of Fe, or 3 moles of O per 2 moles of Fe (Fe_2O_3).

Performing the above calculations for a compound with 2.618 g of Fe to 1.000 g of O yields 0.0469 mol of Fe and 0.0625 mol of O, or a mole ratio of 1.333, or a 4:3 ratio (Fe_3O_4).

116. (D) The weighted-average atomic mass of Sr is expressed as follows:

$$\begin{aligned} \text{atomic mass of Sr} &= 87.62 \text{ amu} = 83.9134(0.0056) + 85.9093x \\ &\quad + 86.9089[1 - (0.0056 + 0.8258 + x)] + 87.9056(0.8258) \end{aligned}$$

Rearrange the above equation and solve for x , which is 0.095 or 9.5%, which is the relative abundance of ^{86}Sr . Therefore, the relative abundance of ^{87}Sr is 0.0735 or 7.3%.

The reason for the imprecision is the low number of significant figures for ^{84}Sr .

117. (M) This problem lends itself well to the conversion pathway:

$$\begin{aligned} &\frac{0.15 \text{ mg Au}}{1 \text{ ton seawater}} \times \frac{1 \text{ ton seawater}}{1000 \text{ kg}} \times \frac{1 \text{ kg}}{1000 \text{ mg}} \times \frac{1.03 \text{ g seawater}}{1 \text{ mL seawater}} \times \frac{250 \text{ mL}}{\text{sample}} \\ &\times \frac{1 \text{ g Au}}{1000 \text{ mg Au}} \times \frac{1 \text{ mol Au}}{196.97 \text{ g Au}} \times \frac{6.02 \times 10^{23} \text{ atoms}}{1 \text{ mol Au}} = 1.2 \times 10^{14} \text{ atoms of Au} \end{aligned}$$

118. (M) In sections 2-7 and 2-8, the simplest concept is the concept of mole. Mole is defined by the number of atoms in 12 g of ^{12}C . Other topics emanate from this basic concept. Molar mass (weight) is defined in terms of moles, as is mole ratios. Percent isotopic abundance is another topic defined directly by the concept of moles.