Understanding Isotopes Application Questions Answer Page

1. A naturally occurring sample of potassium contains 93.12% of the isotope potassium-39 and 6.88% of the isotope potassium-41. Calculate the average atomic mass for this sample.

   ^{39}K, 0.9312; ^{41}K, 0.0688
   Average atomic mass = (mass ^{39}K x fraction) + (mass ^{41}K x fraction)
   = (39 x 0.9312) + (41 x 0.0688)
   = 36.3168 + 2.8208
   = 39.1376

   The average atomic mass of potassium is 39.1376 u.

2. Calculate the average atomic mass of magnesium if the abundances making up a naturally occurring sample are: magnesium-24 (78.70%), magnesium-25 (10.13%), magnesium-26 (11.17%)

   ^{24}Mg, $fraction_1$ = 0.787; ^{25}Mg, $fraction_2$ = 0.1013; ^{26}Mg, $fraction_3$ = 0.1117
   Average atomic mass = (mass ^{24}Mg x $fraction_1$) + (mass ^{25}Mg x $fraction_2$) + (mass ^{26}Mg x $fraction_3$)
   = (24 x 0.787) + (25 x 0.1013) + (26 x 0.1117)
   = 18.888 + 2.5325 + 2.9042
   = 24.3247

   The average atomic mass of magnesium is 24.3247 u.

3. Naturally occurring uranium is composed of three major isotopes: uranium-238 (99.28%), uranium-235 (0.71%) and uranium-234 (0.0054%). Calculate the average atomic mass of uranium.

   ^{238}U, $fraction_1$ = 0.9928; ^{235}U, $fraction_2$ = 0.0071; ^{234}U, $fraction_3$ = 0.000054
   Average atomic mass = (mass ^{238}U x $fraction_1$) + (mass ^{235}U x $fraction_2$) + (mass ^{234}U x $fraction_3$)
   = (238 x 0.9928) + (235 x 0.0071) + (234 x 0.000054)
   = 236.2864 + 1.6685 + 0.0126
   = 237.9675

   The average atomic mass of uranium is 237.9675 u.

4. The average atomic mass of copper is 63.545 u (unified atomic mass). It is made up of two isotopes: copper-63 (atomic mass 62.930) and copper-65 (atomic mass 64.928). What must be the relative abundance (in %) of each of these isotopes in naturally occurring samples of copper?

Average atomic mass = 63.545; ^{63}Cu = 62.930, $fraction_1 = fraction_1$; ^{65}Cu = 64.928, $fraction_2 = (1 - fraction_1)$
Average atomic mass = (mass ^{63}Cu x $fraction_1$) + (mass ^{65}Cu x $fraction_2$)

$$63.545 = (62.930 \times fraction_1) + (64.928 \times [1 - fraction_1])$$
$$63.545 = 62.93 \times fraction_1 + 64.928 - 64.928 \times fraction_1$$
$$63.545 - 64.928 = 62.93 \times fraction_1 - 64.928 \times fraction_1$$
$$-1.383 = -1.998 \times fraction_1$$
$$fraction_1 = 0.6922 \text{ or } 69.22\% \text{ of } ^{63}Cu$$
$$fraction_2 = (1 - fraction_1) = (1 - 0.6922) = 0.3078 \text{ or } 30.78\% \text{ of } ^{65}Cu$$

The relative abundance of copper-63 is 69.22% and the relative abundance of copper-65 is 30.78%.

5. The average atomic mass of carbon is 12.011 u. It is made up of two stable isotopes: carbon-12 and carbon-13. What must be the relative abundance (in %) of each of these isotopes in naturally occurring samples of carbon?

Average atomic mass = 12.011; ^{12}C = 12, $fraction_1 = fraction_1$; ^{13}C = 13, $fraction_2 = (1 - fraction_1)$
Average atomic mass = (mass ^{12}C x $fraction_1$) + (mass ^{13}C x $fraction_2$)

$$12.011 = (12 \times fraction_1) + (13 \times [1 - fraction_1])$$
$$12.011 = 12 \times fraction_1 + 13 - 13 \times fraction_1$$
$$12.011 - 13 = 12 \times fraction_1 - 13 \times fraction_1$$
$$-0.989 = -fraction_1$$
$$fraction_1 = 0.989 \text{ or } 98.9\% \text{ of carbon-12}$$
$$fraction_2 = (1 - fraction_1) = (1 - 0.989) = 0.011 \text{ or } 1.1\% \text{ of carbon-13}$$

The relative abundance of carbon-12 is 98.9% and the relative abundance of carbon-13 is 1.1%.