Simple Step-by-Step Guides to Solving Chemistry Problems

Relative Atomic Mass & Isotopes

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RELATIVE ATOMIC MASS & ISOTOPES

Relative Atomic Mass (A_r)

Chemical elements are the building blocks from which everything is constructed, from specks of dust to mobile phones and from flora and fauna to the clothes we wear. There are over 100 known elements. An **element** is a pure substance that cannot be chemically broken down. The smallest unit of an element is the **atom**. Different atoms have different masses. The mass of an atom is so small that it is more convenient to compare atom masses, rather than refer to their actual mass. The standard for this relative scale is an atom of carbon-12, which has a relative atomic mass (A_r) of 12.

The table below lists the relative atomic mass (A_r) values of some common elements.

Selected Relative Atomic Mass Values

Element	Approximate Relative Atomic Mass (A_r)
Hydrogen	1
Carbon	12
Nitrogen	14
Oxygen	16
Sodium	23
Magnesium	24
Silicon	28
Calcium	40
Bromine	80

These relative atomic mass values tell us for example that sodium atoms (A_r = 23) are 23 times heavier than hydrogen atoms (A_r = 1), two atoms of neon (A_r = 20) have the same mass as one atom of calcium (A_r = 40) and that three oxygen atoms (A_r = 16) weigh the same as two magnesium atoms (A_r = 24).

Relative atomic masses are listed in the periodic table.

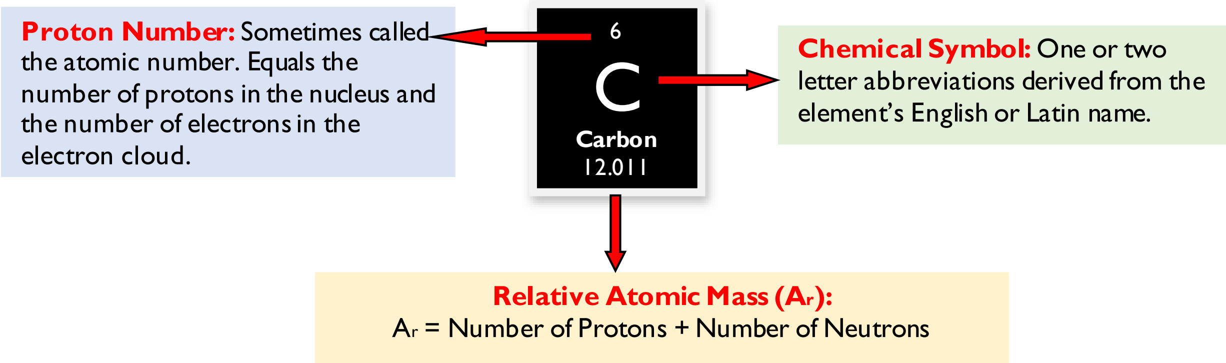
Isotopes

Elements are defined by their proton (atomic) number. An atom with 7 protons is always nitrogen (N), an atom with 20 protons is always calcium and an atom with 70 protons must therefore always be gold (Au). Isotopes are atoms that have the same number of protons, but have different numbers of neutrons. For example, carbon has three naturally occurring isotopes, often referred to as simply carbon-12, carbon-13 and carbon-14, with relative atomic masses of 12, 13 and 14, respectively. Since carbon has a proton number of 6, the isotopes contain 6, 7 and 8 neutrons, respectively.

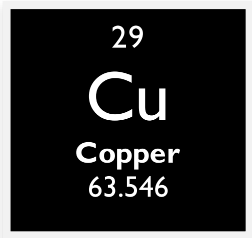
Three isotopes of carbon

Isotope	No. of protons	No of neutrons	No of electrons	Relative Atomic Mass (A_r)
$^{12}C_6$	6	6	6	12
$^{13}C_6$	6	7	6	13
$^{14}C_6$	6	8	6	14

A typical periodic table information box for the element carbon is given below:



The relative atomic masses listed in the periodic table are an average of the masses of the different naturally occurring isotopes. The rounded value gives you the mass of the most abundant isotope. For example, copper occurs naturally as Cu-63 and Cu-65. Given that the average relative atomic mass of copper is 63.546, we can conclude that Cu-63 is the more abundant isotope since the average value is closer to 63 than to 65.



Calculating Average Relative Atomic Mass

Chlorine has a relative atomic mass of 35.5, which is the average of the masses of its two naturally occurring isotopes. This is calculated by working out the relative abundance of each isotope. For example, in any sample of chlorine, 25% will be Cl-37 and 75% Cl-35. The relative atomic mass is therefore calculated using the equation:

(% of isotope 1/100 × mass of isotope 1) + (% of isotope 2/100 × mass of isotope 2)

So, in the case of chlorine:

Average relative atomic mass = (75/100 x 35) + (25/100 × 37) = 26.25 + 9.25 = **35.5**

Example 1: Boron has three naturally occurring isotopes: 9% boron-10 and 80.1% boron-11. Calculate the relative atomic mass of boron.

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Answer:

Average relative atomic mass = (19.9/100 x 10) + (80.1/100 x 11)
= 1.99 + 8.811
= 10.801

Example 2: Magnesium has three isotopes, Given that the natural abundances are Mg-24 (78.70%), Mg-25 (10.13%), and Mg-26 (11.17%) calculate the relative atomic mass of magnesium.

Answer:

Average relative atomic mass = (78.7/100 x 24) + (10.13/100 x 25) + (11.17/100 x 26)
= 18.888 + 2.533 + 2.904
= **24.33**

This problem can also be reversed, as in having to calculate the isotopic abundances when given the atomic mass and isotopic masses.

Example 3: *Nitrogen is made up of two isotopes, ^{14}N and ^{15}N. Given nitrogen's atomic weight of 14.007, what is the percentage abundance of each isotope?*

Answer:

Let x = percentage abundance of ^{14}N
Therefore, percentage abundance of ^{15}N = 100- x
(x/100 x 14) + (100-x/100 x 15) = 14.007
Multiply throughout by 100
14x + 15(100 - x) = 1400.7
14x + 1500 – 15 x = 1400.7
x = 1500 – 1400.7 = 99.3
Thus % abundance ^{14}N = 99.3%, ^{15}N = **0.7%**

Example 4: The relative atomic mass of copper is 63.55. Copper has two naturally occurring isotopes, Cu-63 and Cu-65. Determine the natural abundance (%) of each isotope.

Answer:

Let x = percentage abundance of ^{63}Cu
Therefore, percentage abundance of ^{65}Cu = 100- x
(x/100 x 63) + (100-x/100 x 65) = 63.55
Multiply throughout by 100
63x + 65(100 - x) = 6355
63x + 6500 – 65 x = 6355
2x = 6500 – 6355
x = 72.5

Thus % abundance **Cu-63 = 72.5% and Cu-65 = 27.5%.**

Practice Problems

a) The natural abundance for gallium isotopes is ^{69}Ga 60.11% and ^{71}Ga 39.89%. Calculate the relative atomic mass of gallium.

b) Neon has three isotopes, ^{20}Ne, ^{21}Ne and ^{22}Ne. The percentage of each in order is 90.48%, 0.27% and 9.25%. Calculate the relative atomic mass of neon.

c) Element Y exists as two naturally occurring isotopes, Y-51 and Y-52. Given that Y-51 has a natural abundance of 18%, calculate the relative atomic mass of element Y.

d) Lithium has two naturally occurring isotopes, Li-6 and Li-7. Determine the percentage abundance of each isotope, given that lithium has a relative atomic mass of 6.94.

e) Bromine has a relative atomic mass of 79.90. There are two known isotopes of bromine, with relative atomic masses of 79 and 81. Determine the percentage abundance of each isotope.

Answers are given on the following page.

Practice Problem Answers:

a) The natural abundance for gallium isotopes is 69Ga 60.11% and 71Ga 39.89%. Calculate the relative atomic mass of gallium.

Average A_r = (60.11/100 x 69) + (39.89/100 x 71)
= 41.476 + 28.399
= 69.798

b) Neon has three isotopes, 20Ne, 21Ne and 22Ne. The percentage of each in order is 90.48%, 0.27% and 9.25%. Calculate the relative atomic mass of neon.

Average A_r = (90.48/100 x 20) + (0.27/100 x 21) + (9.25/100 x 22)
= 18.096 + 0.057 + 2.035
= 20.188

c) Element Y exists as two naturally occurring isotopes, Y-51 and Y-52. Given that Y-51 has a natural abundance of 18%, calculate the relative atomic mass of element Y.

Natural abundance Y-52 = 100 – 18 = 28%

Average A_r = (18/100 x 51) + (82/100 x 52)
= 9.18 + 42.64
= 51.82

d) Lithium has two naturally occurring isotopes, Li-6 and Li-7. Determine the percentage abundance of each isotope, given that lithium has a relative atomic mass of 6.94.

Let x = percentage abundance of Li-6 = x
Therefore, percentage abundance of Li-7 = 100- x
(x/100 x 6) + (100-x/100 x 7) = 6.94
Multiply throughout by 100
6x + 700 - 7x = 694
-x = -6
x = 72.5

Thus % abundance **Li-6 = 6% and Li-7 = 94%.**

e) Bromine has a relative atomic mass of 79.90. There are two known isotopes of bromine, with relative atomic masses of 79 and 81. Determine the percentage abundance of each isotope.

Let x = percentage abundance of Br-79 = x
Therefore, percentage abundance of Br-80 = 100- x
(x/100 x 79) + (100-x/100 x 81) = 79.90
Multiply throughout by 100
79x + 8100 - 81x = 7990
110 = 2x
x = 55

Thus % abundance **Br-79 = 55 and Br-81 = 45%**