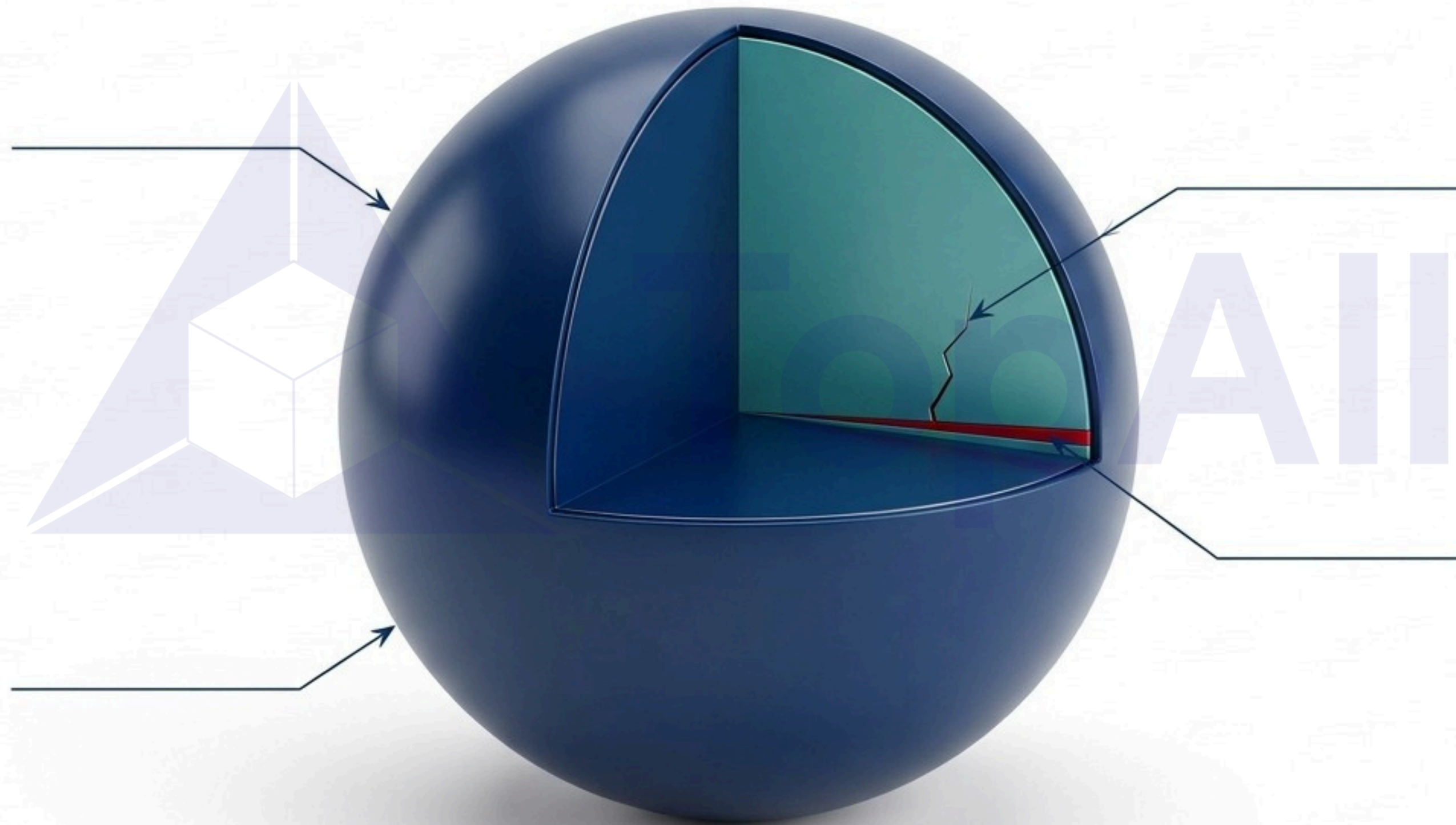


Average Atomic Mass from Isotopes

A Statistical Consequence of Elemental Mixtures (NCERT 1.7.2)



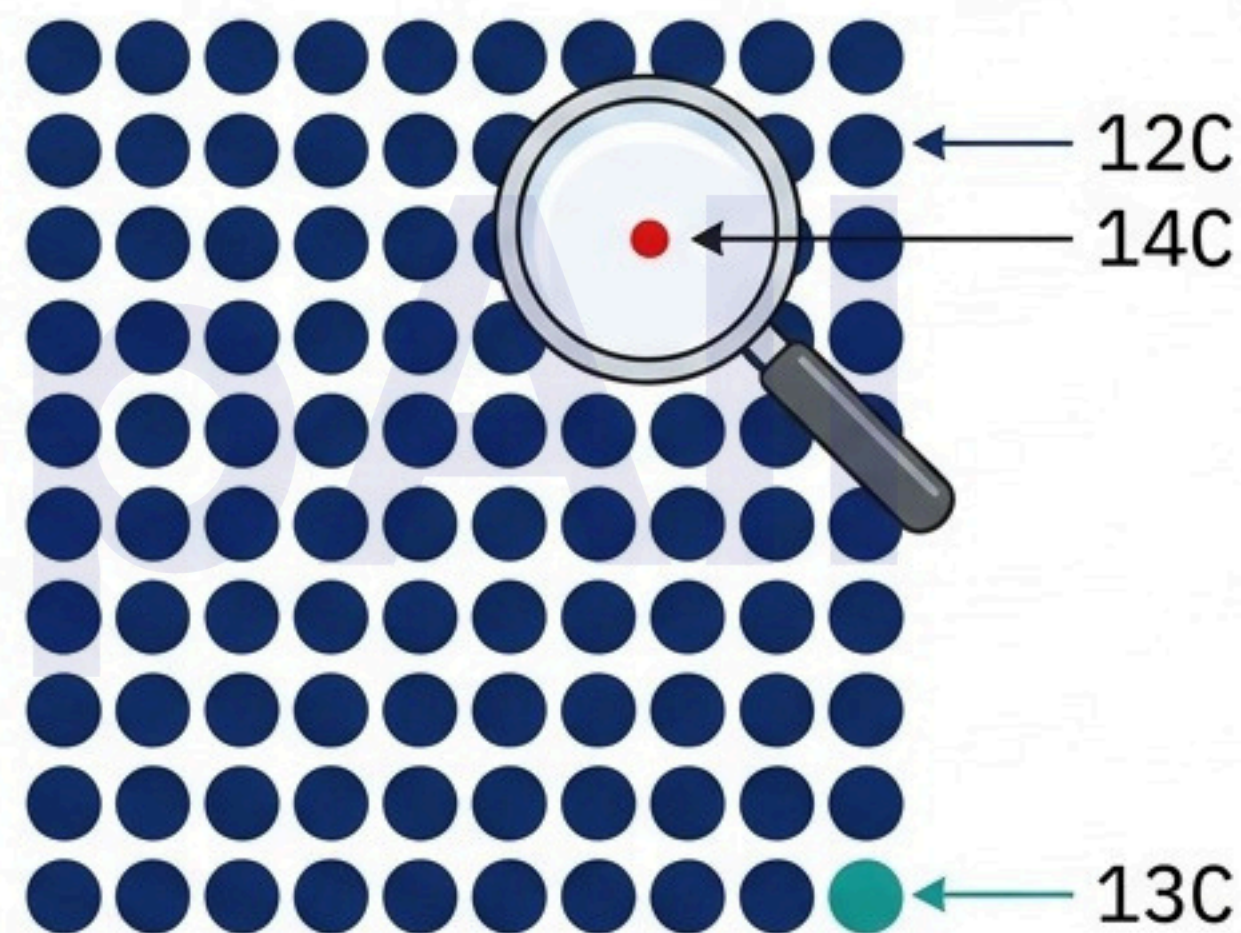
The Core Concept: Elements are Mixtures

The Definition

Isotopes: Atoms of the same element of the element with different masses.

Relative Abundance: The percentage (%) occurrence of each isotope in nature.

The Visual Proof



Because elements exist as mixtures, we cannot use a single integer for their mass.
We must calculate a statistical average.

The Data: Carbon Isotope Distribution

Isotope	Relative Abundance (%)	Atomic Mass (amu)
^{12}C	98.892	12
^{13}C	1.108	13.00335
^{14}C	2×10^{-10}	14.00317



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$$E=mc^2$$



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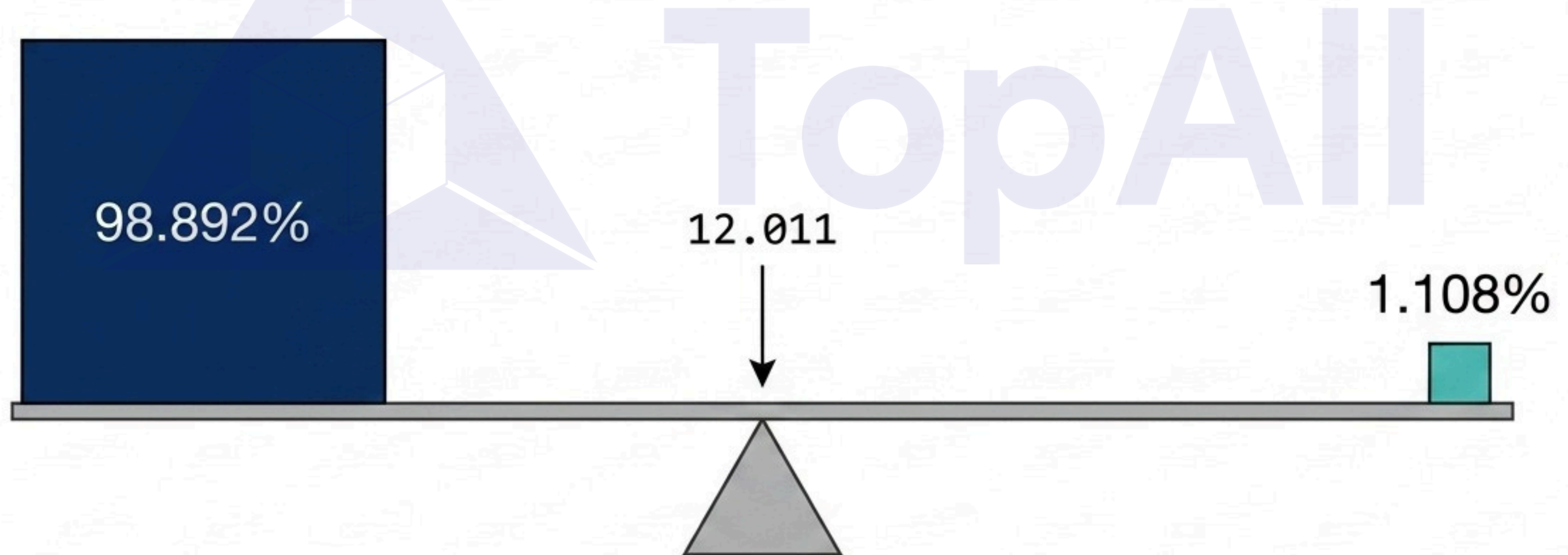
Key Highlights

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The Governing Rule

$$\text{Average atomic mass} = \sum (\text{fractional abundance} \times \text{isotopic mass})$$



JEE Trap Alert

DO NOT plug percentages directly into the formula! Fractional Abundance requires dividing the percentage by 100.


$$\text{Abundance} = 2 \times 10^{-10}$$

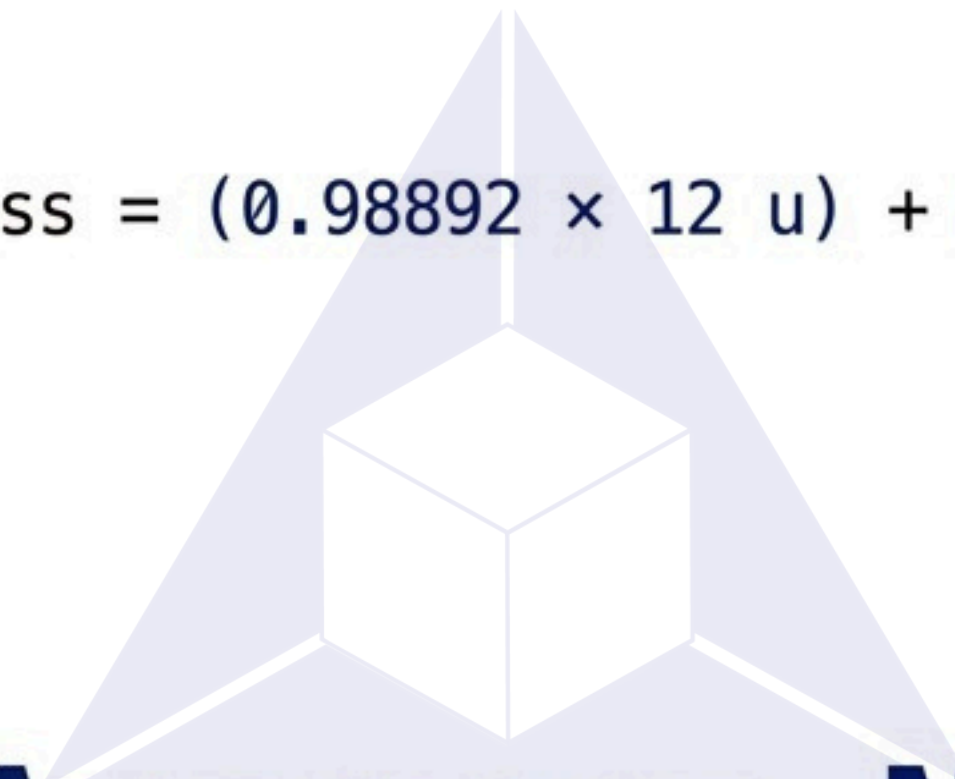
$$\text{Abundance} = \frac{2 \times 10^{-10}}{100} = 2 \times 10^{-12}$$



Losing that 10^{-2} factor will result in a negative marking on numerical integer type questions.

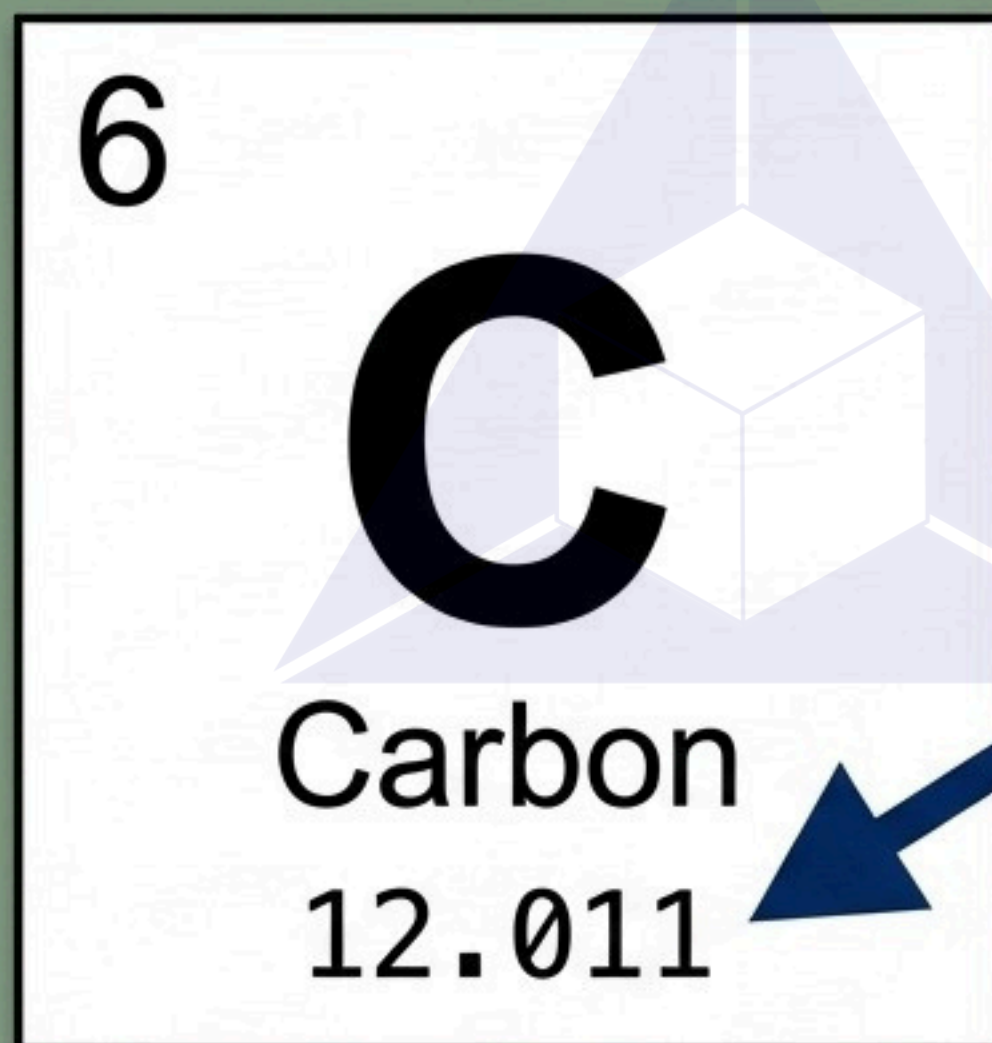
Mini Application: Calculating Carbon's Mass

$$\text{Average Mass} = (0.98892 \times 12 \text{ u}) + (0.01108 \times 13.00335 \text{ u}) + (2 \times 10^{-12} \times 14.00317 \text{ u})$$



Average Mass = 12.011 u

Conceptual Synthesis



6
C
Carbon
12.011

The fractional masses seen on the periodic table are NOT the mass of a single atom. You cannot hold a carbon atom that weighs **12.011** u. The **periodic table displays** the weighted statistical average of the universe's isotope distribution.



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