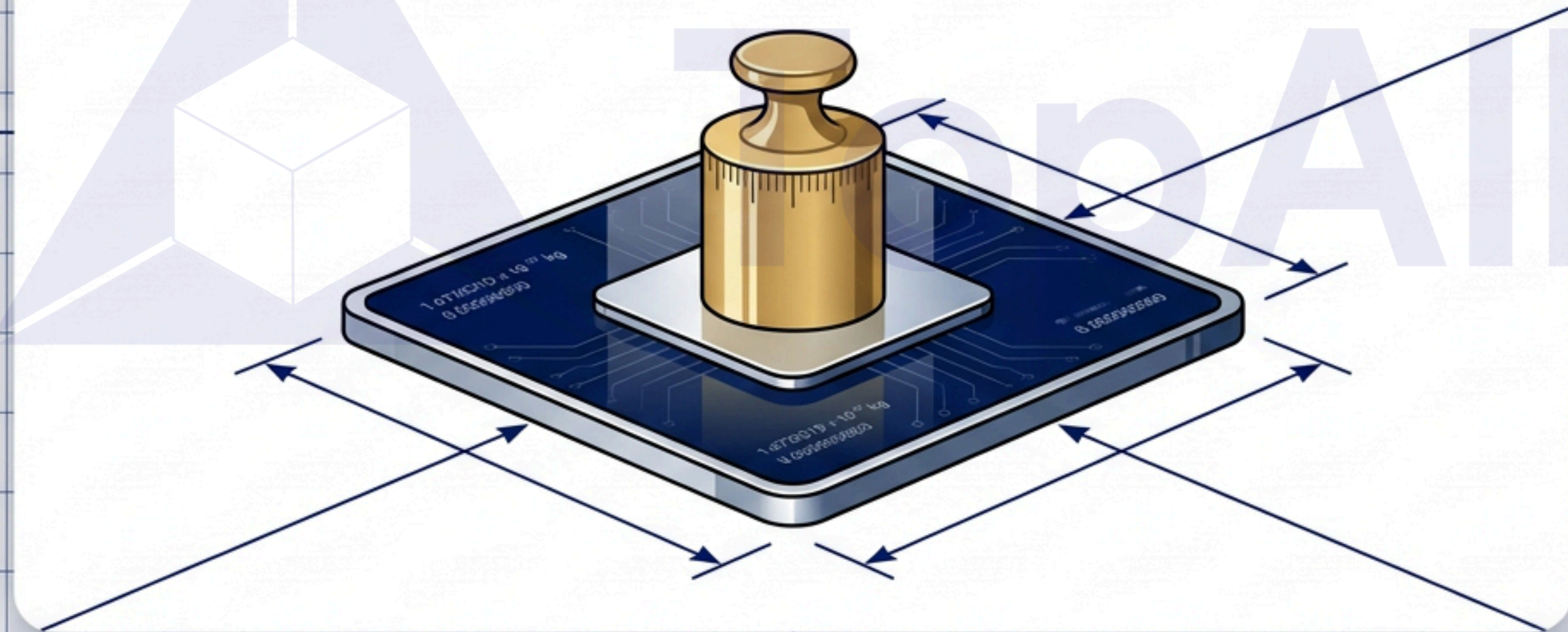


# Weighing the Invisible

The engineering behind atomic and molecular masses

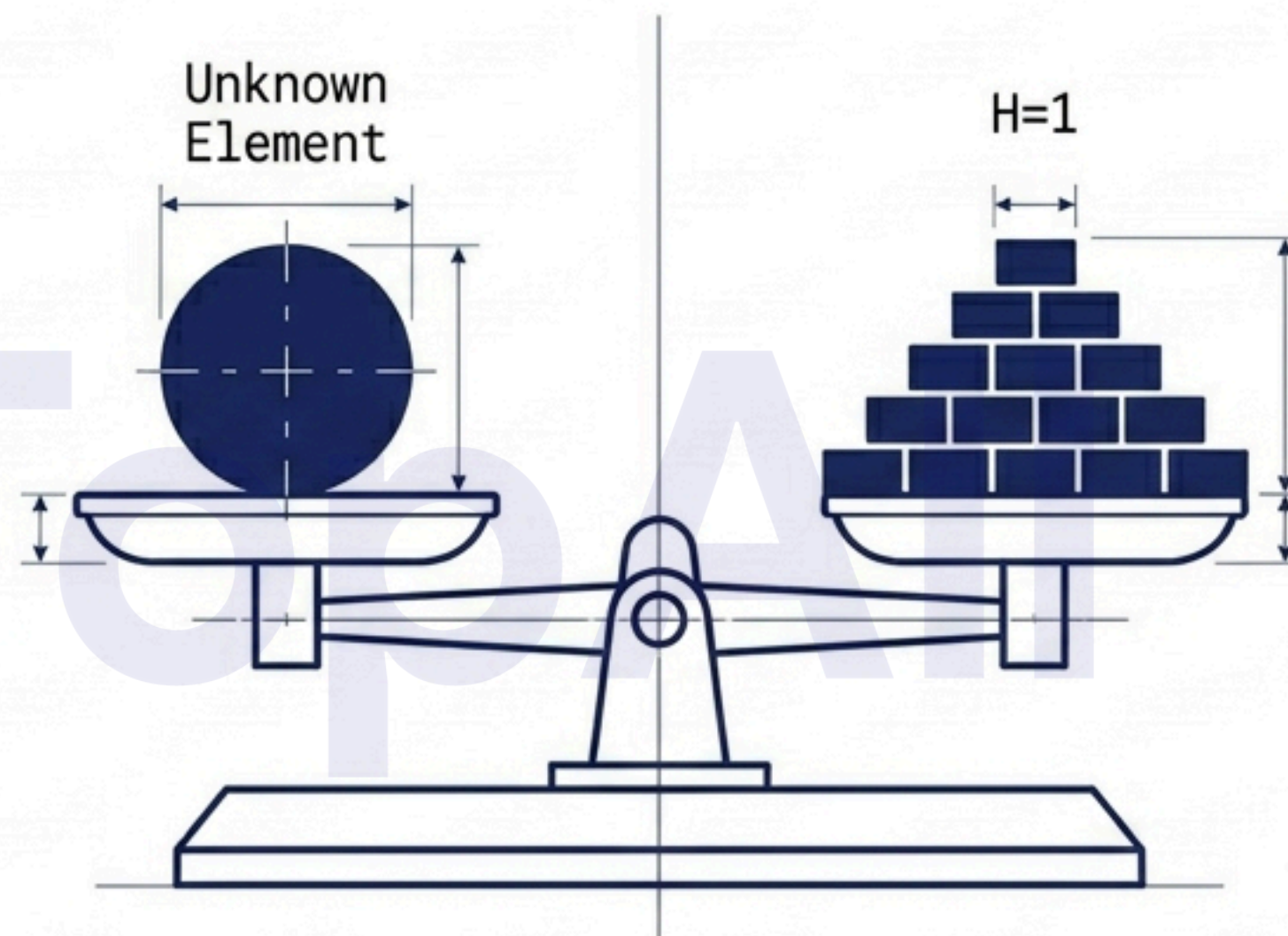




# The Relative Mass Solution

## Relative Measurement

Instead of calculating absolute mass in grams, early scientists assigned Hydrogen an arbitrary mass of '1' and weighed everything else relative to it.



Comparative ratios eliminate the need for absolute units.



# You've Learned the Theory. Let's Test Them Now.

$$E=mc^2$$



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## Key Highlights

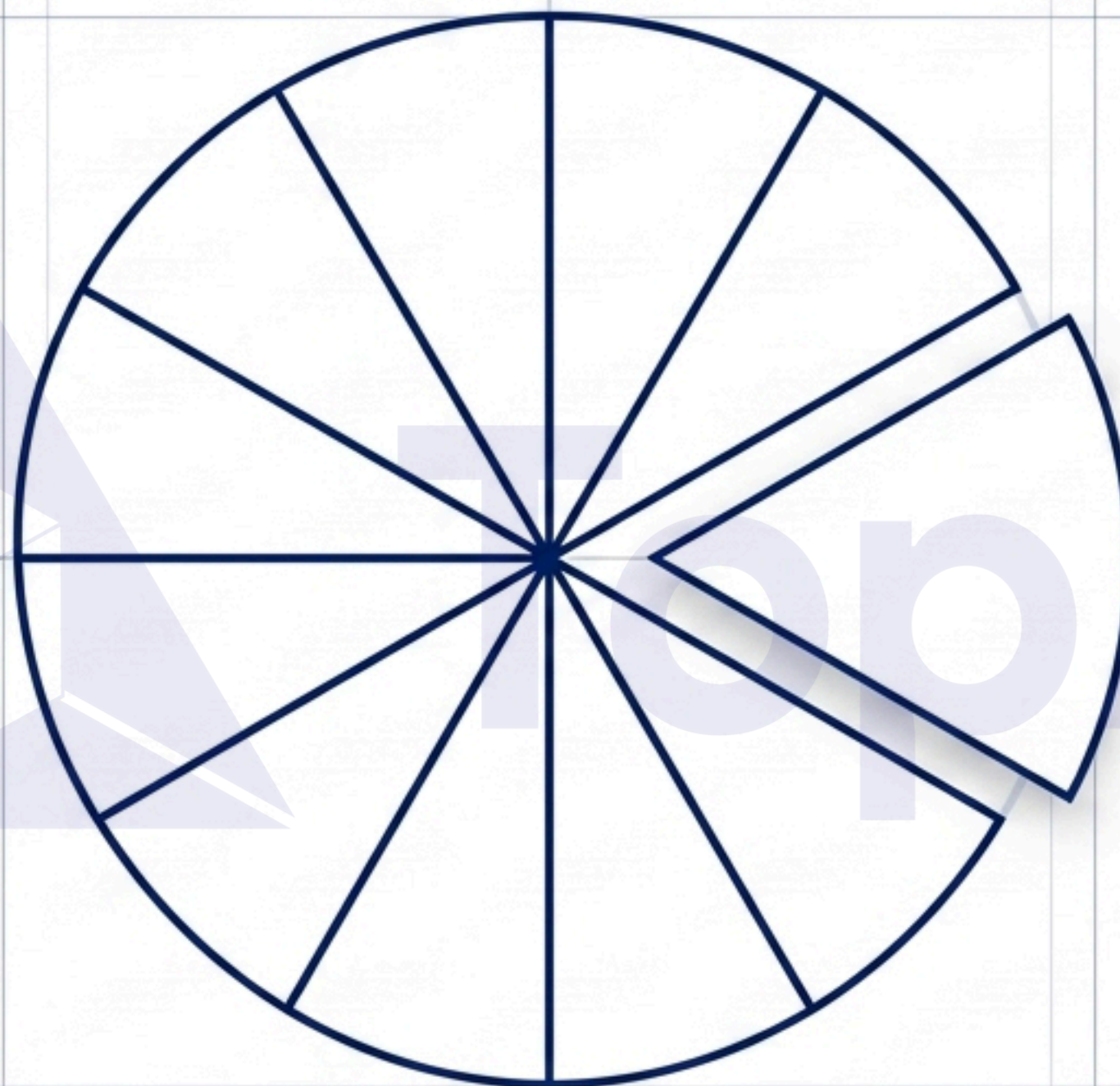
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# The Carbon-12 Standard

Established 1961 as  
the universal anchor.

$$1 \text{ u} = 1.66056 \times 10^{-24} \text{ g}$$

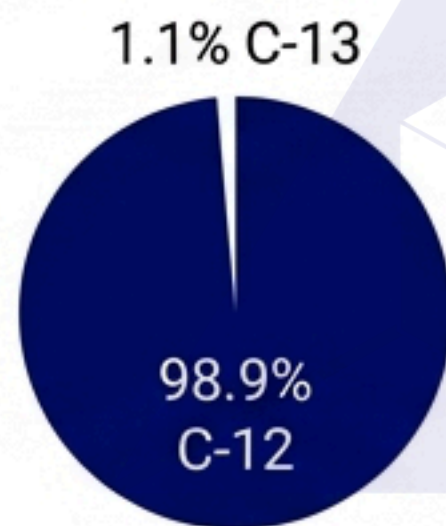


**1 atomic mass  
unit (u) =  
exactly 1/12th  
of Carbon-12**

# Structural Mass Taxonomy

## Average Atomic Mass

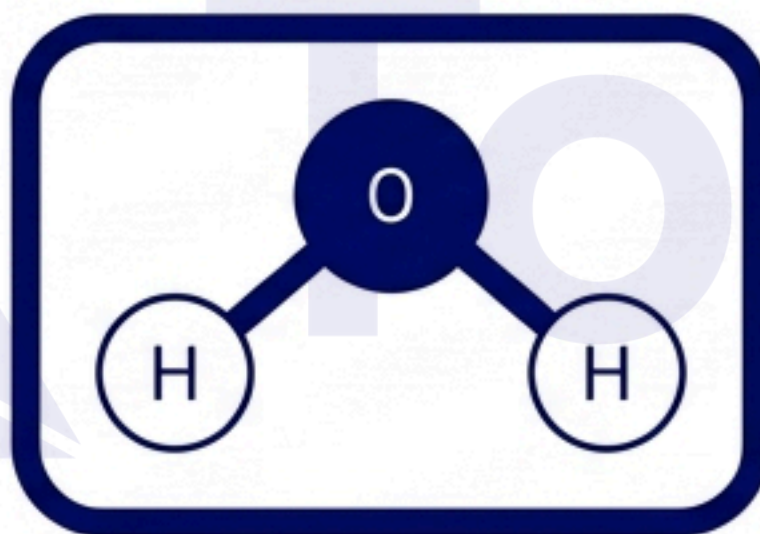
Elements with Isotopes  
(e.g., Carbon)



Calculation:  
Fractional abundance  
sum = 12.011 u

## Molecular Mass

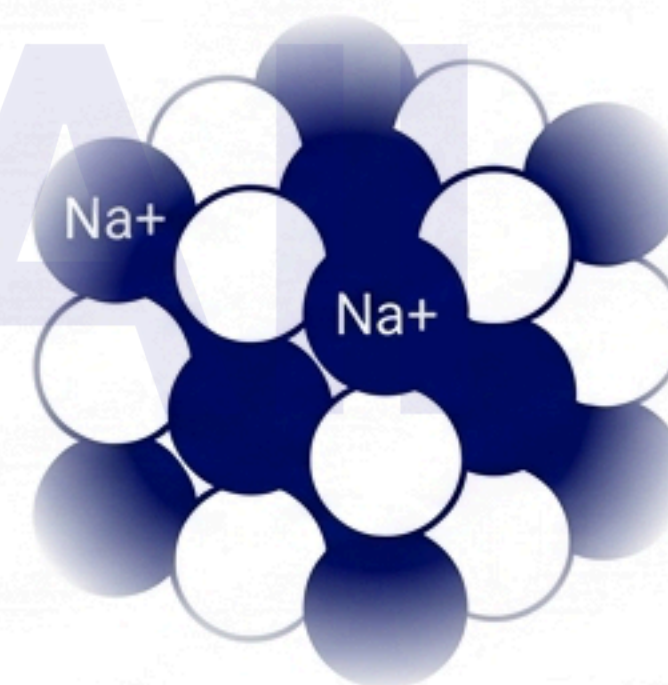
Discrete Molecules  
(e.g., H<sub>2</sub>O)



Calculation:  
 $2(1.008) + 16.00 = 18.02 \text{ u}$

## Formula Mass

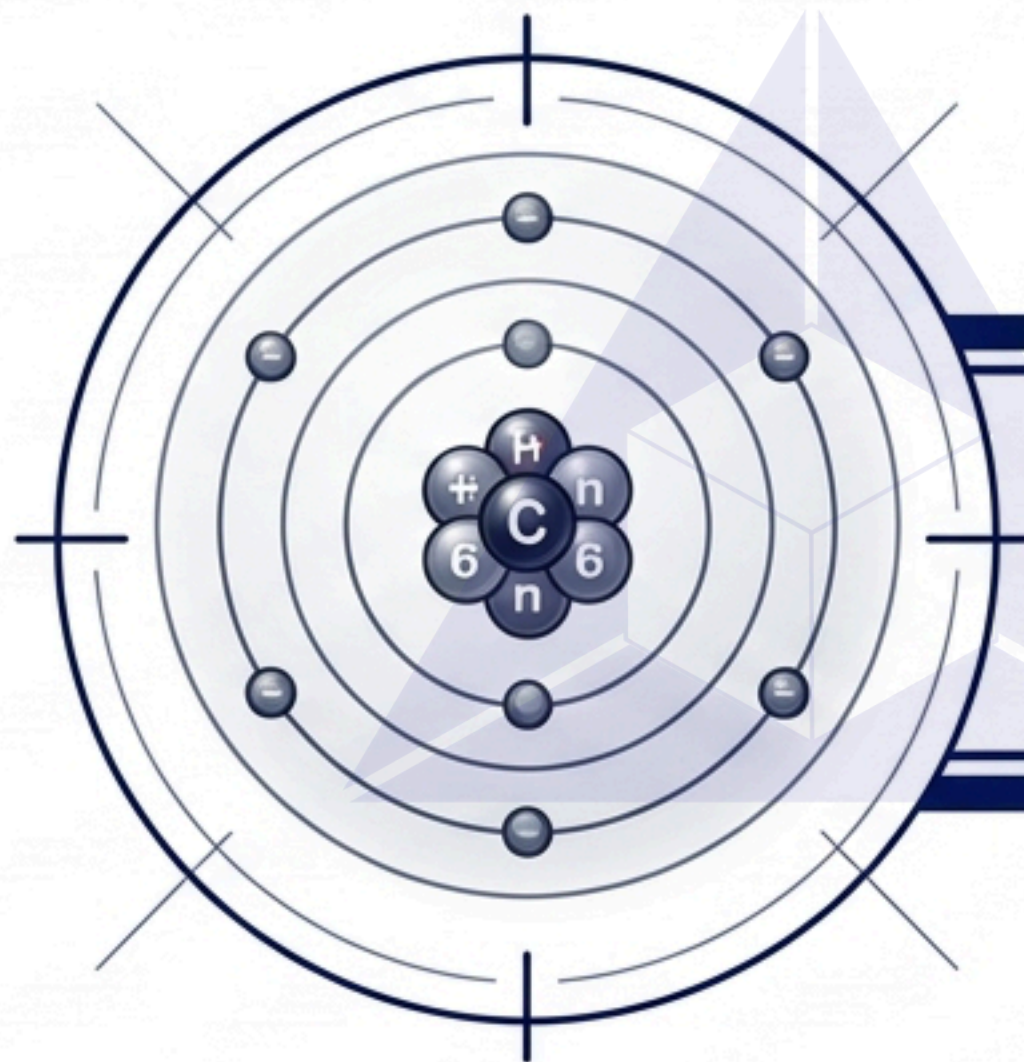
3D Crystal Lattices  
(e.g., NaCl)



Calculation:  
 $23.0 + 35.5 = 58.5 \text{ u}$

# Bridging the Micro and Macro

Key Takeaway: The numerical value remains identical; only the unit scales up.



Mass = 12 u

**× 1 MOLE**  
[Avogadro's Constant:  $6.022 \times 10^{23}$ ]



Mass = 12 grams



# You've Learned the Theory.

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