

## Atomic and Molecular Mass

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## Isotopes

- Nuclei that have the same number of protons but a different number of neutrons
- Have identical chemical properties (all have same number of protons and thus electrons) but different physical properties
- The existence of isotopes is evidence for the existence of neutrons inside the nucleus

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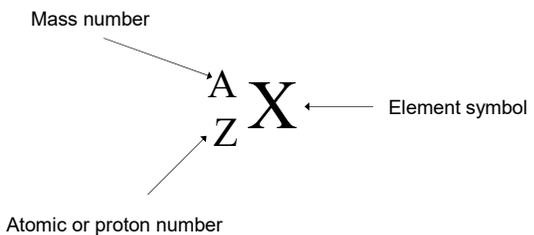
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## Examples

${}^1_1\text{H}$ ,  ${}^2_1\text{H}$ ,  ${}^3_1\text{H}$  (hydrogen)

${}^{238}_{92}\text{U}$ ,  ${}^{235}_{92}\text{U}$  (uranium)

Sometimes we drop the atomic number

${}^{24}\text{Mg}$ ,  ${}^{25}\text{Mg}$ ,  ${}^{26}\text{Mg}$  (magnesium)

We can also represent the isotope like this:

Sodium-23, Sodium-24 (sodium)

Carbon-12, Carbon-14 (carbon)

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## Relative Atomic Mass ( $A_r$ )

- Atoms are small and therefore measuring mass in kilograms or grams would give extremely small numbers.
- Instead the mass of an atom is compared with that of an atom of carbon-12.
- The relative atomic mass of carbon-12 is taken to be 12.

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## Average Atomic Mass

- If the relative atomic mass is based on carbon, why doesn't carbon have a mass of exactly 12?
  - Carbon has more than one isotope
  - The mass quoted on the periodic table is really the average atomic mass of all atoms of carbon
  - The amount of each isotope of an element is known as the percent abundance

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## Calculating Average Atomic Mass

- The average atomic mass is the weighted average of all of the isotopes of the element

$$\begin{aligned} &(\text{mass of isotope 1})(\text{percent abundance of isotope 1}) + \\ &(\text{mass of isotope 2})(\text{percent abundance of isotope 2}) + \\ &(\text{mass of isotope 3})(\text{percent abundance of isotope 3}) + \dots \\ &= \text{average atomic mass} \end{aligned}$$

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## Example

- Potassium (K)

Isotope	Percent Abundance	Atomic Mass (u)
K-39	93.2581	38.963707
K-40	0.0117	39.963998
K-41	6.7302	40.961826

$$\begin{aligned} &(38.963707)(0.932581) \\ &(39.963998)(0.000117) \\ &+ (40.961826)(0.067302) \\ &39.098301 \text{ u} \end{aligned}$$

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