"We are all bold as love." -*Jimi Hendrix*

Chapter 5 Notes - Atomic Structure - Protons, Neutrons and Electrons

Subatomic Particle Table:

Particle	Symbol	Relative Charge	Mass number	Actual mass (amu)	Location
electron	e^-	-1	0	9.109×10^{-31}	cloud
proton	p^+	+1	1	1.673×10^{-27}	nucleus
neutron	n^0	0	1	1.675×10^{-27}	nucleus

Distinguishing Between Atoms

Atom type or element determined by the **proton number** or the **atomic number (Z)**

ex) Z = 13 Z = 29

Al Cu

$13p^+$ $13e^-$

In **neutral atoms**, the proton(+) number equals the electron (-) number.

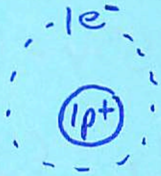
ex) S ☐ ← no charge O

$16p^+$ $16e^-$ sulfur $8p^+$ $8e^-$ oxygen

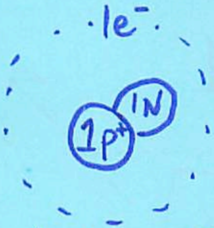
Each element may have a different percentage of **isotopes** – atoms with the same proton number but different **neutron** number

- this causes the mass of the atom to be different
- some isotopes are radioactive - meaning the nucleus is unstable
- this has no effect on the chemical reactivity, charge or identity of the element

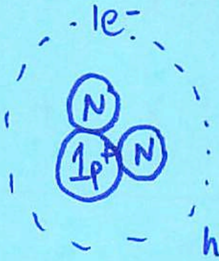
ex) Hydrogen's isotopes:



99.98% abundant



0.015% abundant



trace amounts

mass # → p+N hydrogen-1 hydrogen-2 hydrogen-3

Symbol Notation: $^{1}_{1}H$ $^{2}_{1}H$ $^{3}_{1}H$ "Tritium" ↑ mass #

atomic # → p "Protium" "Deuterium"

Since neutrons are so massive, different isotopes have different masses, even though they are the same element!

mass number = number of protons and neutrons in a single atom

ex) Oxygen's Isotopes

P+N → $^{16}_{8}O$ ← atomic # (p) $^{18}_{8}O$

8P 8P

8N 10N

oxygen-16 oxygen-18

Stable Isotopes

lighter elements 1:1 ratio P:N

heavier elements 1:1.5 ratio P:N

review: atomic # = ________ mass # = ____________

<u>**Average atomic mass**</u> - the weighted average mass derived from all of the naturally occurring isotopes.

Isotopes of elements in nature exist in certain percentages in the universe - **percent abundance**

<u>**Magnesium Isotopes in the Universe**</u>

mass number	**isotope mass**	**(amu)**	**percent abundance**
24 P+N	23.99		78.99 %
25	24.99		10.00
26	25.98		11.01

.78,99

Weighted Average:

$$\underset{\text{mass}}{\underline{23.99\text{ amu}}}\left(\underset{\text{abundance}}{0.\underline{7899}}\right) + 24.99(0.1000) + 25.98(0.1101) =$$

Q: Why aren't AMU masses of each isotope in whole numbers like the mass numbers?
A: An amu is exactly 1/12 of the mass of a carbon-12 atom. This amount is not the same as the mass of a proton or a neutron, but very close. (since carbon 12 has 6 protons, 6 neutrons and 6 electrons)

Challenge Problem:

Cu has two stable isotopes, Copper-63 and Copper-65. If the average atomic mass of Cu is 63.55, find the percent abundance of each of its two isotopes.

$$63(x) + 65(y) = 63.55$$
$$x + y = 1$$
$$\boxed{x = 1-y}$$

$$63(1-y) + 65y = 63.55$$
$$63 - 63y + 65y = 63.55$$
$$63 + 2y = 63.55$$
$$2y = 0.55$$
$$y = 0.275 \sim 0.28$$

Copper-65 = 28% abundant

Copper-63 = 72% abundant