

## Modern View of Atomic Structure

- Since 1963, following the work of Murray Gell-Mann, physicists have described the structure of the atom in terms of quarks and electrons.
  - T Up quarks have a charge of  $+2/3$  unit charge.
  - T Down quarks have a charge of  $-1/3$  unit charge.
  - T Electrons have a charge of  $-1$  unit charge.
  - T Protons and neutrons are made up of three quarks.
- proton = 2 up + 1 down  
=  $(2)(+2/3) + (-1/3) = +1$
- neutron = 1 up + 2 down  
=  $(+2/3) + (2)(-1/3) = 0$
- L Quarks are not essential to understand general chemistry.

## Subatomic Particles for Chemistry

Particle	Unit Charge	Mass
Proton (p)	1+	$1.6726 \times 10^{-24}$ g
Neutron (n)	0	$1.6749 \times 10^{-24}$ g
Electron (e)	1-	$9.1095 \times 10^{-28}$ g

## Nuclear Parameters

1. All atoms of a given element have the same number of protons, which defines the element's **atomic number**, given the symbol  $Z$ .
2. Together, protons and neutrons are known as **nucleons**.
3. Any atom with a certain number of nucleons is called a **nuclide**.
4. The number of nucleons defines the nuclide's **mass number**,  $A$ :

$A = \text{number of protons} + \text{number of neutrons}$

- K Note that  $A$  is an integer count of the number of nucleons, and *not* a statement of an atom's mass.

## **Isotopes and Isobars**

**Isotopes** of an element have the same atomic number ( $Z$ ) but have different numbers of neutrons and therefore different mass numbers ( $A$ ).

**Isobars** are nuclides of different elements (different  $Z$  values) with the same mass number ( $A$ ).

# Nuclide Notation



X = element's symbol

Z = atomic number = number of protons

A = mass number = number of nucleons

Symbol	p	n	e
$\begin{matrix} 16 \\ 8 \end{matrix} \text{O}$	8	8	8
$\begin{matrix} 17 \\ 8 \end{matrix} \text{O}$	8	9	8
$\begin{matrix} 15 \\ 8 \end{matrix} \text{O}$	8	7	8
$\begin{matrix} 15 \\ 7 \end{matrix} \text{N}$	7	8	7

## Monatomic Ions

Ion = electrically charged atom or molecule

Cation (kat!  $23\text{e}^-$ ) = positively charged ion

Anion (an!  $23\text{e}^-$ ) = negatively charged ion

Symbol	p	n	e
$^{16}_8\text{O}^{2-}$	8	8	10
$^{64}_{30}\text{Zn}^{2+}$	30	34	28
$^{35}_{17}\text{Cl}^-$	17	18	18
$^{39}_{19}\text{K}^+$	19	20	18

Isoelectronic = same number of electrons

## Atomic Mass Units (amu or u)

One atomic mass unit is defined as 1/12 of the mass of a  $^{12}_6\text{C}$  atom.

$$1 \text{ u} = 1.66054 \times 10^{-24} \text{ g}$$

Particle	Mass (u)
proton	1.007277 u
neutron	1.008665 u
electron	0.0005486 u

## Binding Energy

The mass of a nuclide is not simply the sum of the masses of its fundamental particles.

Nuclide	Measured Mass	Calc'd Mass	Difference
${}^1_6\text{C}$	12 u (exactly)	12.098944 u	0.098944 u
${}^{16}_8\text{O}$	15.99491 u	16.13192 <sub>5</sub> u	0.13702 u
${}^{15}_7\text{N}$	15.00011 u	15.12409 <sub>9</sub> u	0.12399 u

- K When atoms are formed from protons, neutrons, and electrons, some mass is converted into energy, called the **binding energy**.
- K The mass equivalent of this energy can be calculated from the difference between the measured mass of the nuclide and the sum of the masses of its subatomic particles, using  $E = mc^2$ .

## Atomic Weights

- L Tabulated values of atomic weights of elements represent the average atomic mass of all isotopes comprising a naturally occurring sample.
- The average atomic mass is a weighted average, according to abundance of each isotope in a typical sample.
  - Unless the element naturally occurs as only one isotope (e.g., F), atomic weights generally *do not* represent the masses of any individual atoms.