Atomic Concepts

Unit 2

Atomic Concepts

Early Ideas about Matter

Name	Date	Theory	Experiment
Democritus	400 BC	atomos – smallest particle	none, philosophical argument
Aristotle	350 BC	hyle – continuous matter	none, philosophical argument
John Dalton*	1803 AD	billiard ball – indivisible, smallest particle	laws of: conservation of matter definite proportion multiple proportion
JJ Thomson*	1897 AD	plum pudding – electrons in positive pudding $(^{mass} / _{charge})$ ratio e^-	cathode ray tube / e^- beam (a Crookes tube)
Robert Millikan	1910 AD	discovered the charge of the electron	oil drop experiment (gravity, e^- charge, and charged plates)
Ernest Rutherford*	1911 AD	nuclear model – small, dense nucleus with e^- surrounding empty space	α – particle / gold foil
Niels Bohr*	1913 AD	planetary model – e^- travel in energy levels (quantum mechanics)	calculated the energy of e^- shells that matched the spectrum of hydrogen
Erwin Schrödinger	1926 AD	quantum mechanical model – an orbital is the most probable e^- location	calculated the shape of e^- orbitals assuming e^- travel in waves
James Chadwick	1932 AD	discovered the neutron	α – particle / beryllium

*Names required by the NYS Regents Chemistry Common Core

Dalton's Atomic Theory (billiard ball theory) drawn from deductions based on the law of conservation of matter, law of definite proportions, and law of multiple proportions

- matter is composed of extremely small particles called atoms
- atoms are indivisible and indestructible
- atoms of a given element are identical in size, mass, and chemical properties
- atoms of a specific element are different from those of another element
- different atoms combine in simple whole-number ratios to form compounds
- in chemical reactions, atoms are separated, combined, or rearranged

Thomson's Atomic Theory (plum pudding model) drawn from experiments using a CRT

- negative particles (electrons) can be pulled out of atoms (atoms <u>not</u> indestructible)
- the rest of the atom thought to be like a positively charged mass (pudding)

Rutherford's Atomic Theory (nuclear model) drawn from gold-foil experiments

- atom is mostly empty space (most α particles pass through undisturbed)
- must be a small, dense, positively charged nucleus (causing a few large deflections)
- electrons must exist outside the nucleus surrounding the empty space

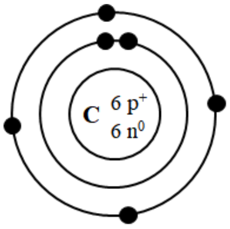
Go on to the next page

Atomic Concepts

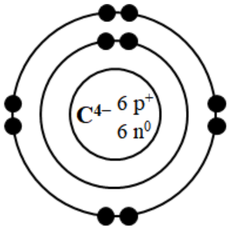
Unit 2

Bohr's Atomic Theory (planetary model) drawn from spectral calculations

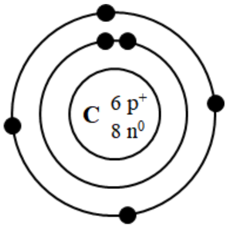
- a circle represents the nucleus – element symbol, #p^{+}, and #n^{o}
- electrons travel in evenly spaced electron shells (energy levels)
    - 1st level holds up to 2 e^{-}
    - 2nd level holds up to 8 e^{-}
    - 3rd level holds up to 18 e^{-}
- for ions, write the charge on the element symbol and change the #e^{-}
- for nuclides of an isotope, write the correct #n^{o}



carbon atom



carbon ion



carbon isotope

Schrödinger's Atomic Theory (quantum mechanical model) drawn from calculations of wave-like electrons

- electrons travel in orbitals (regions of most probable electron location)

Modern Atomic Structure

Atoms are made of three particles

1. protons (p+)
2. neutrons (n^{o})
3. electrons (e^{-})

} nucleus – contains most of the mass

enclose mostly empty space

Subatomic Particle Summary

Particle	General Symbol	Full Notation	Mass	Charge	Location
electron	e^{-}	${}^{0}_{-1}e$	$\frac{1}{1836}$ u	−1	outside nucleus
proton	p^{+}	${}^{1}_{1}H$ or ${}^{1}_{1}p$	1 u	+1	in the nucleus
neutron	n^{o}	${}^{1}_{0}n$	1 u	0	in the nucleus

Electron Configuration

- records the number of electrons in each electron shell (or energy level)
    - 1st level holds up to 2 e^{-}
    - 2nd level holds up to 8 e^{-}
    - 3rd level holds up to 18 e^{-}
- see the Periodic Table of the Elements

Examples:

lithium	Li 2–1
argon	Ar 2–8–8
mercury	Hg 2–8–18–32–18–2

Atomic Concepts

Unit 2

Ground state: all e^- are in the lowest possible electron shells (lowest energy levels)

- NYS Reference Tables for Chemistry shows ground state e^- configurations

Excited state: at least one e^- has absorbed (gained) energy to move to a higher shell

- e^- must absorb (gain) energy to move to higher shells (becomes an excited state e^-)
- excited states are unstable
- e^- will lose energy to move to lower shells closer to the nucleus (the e^- relaxes)
- as e^- relax, they emit (give off) energy (light) of a specific frequency (color) which can be used to identify elements

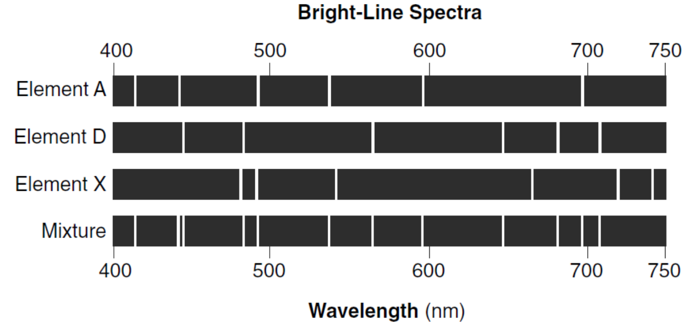
Examples:	ground state	excited states
lithium	2–1	1–2 or 1–1–1
argon	2–8–8	2–8–7–1 or 1–8–8–1

Quantum: the exact (discrete) amount of energy required to jump (or fall) from one energy level to another energy level

- the closer to the nucleus, the larger the quantum of energy
- the farther from the nucleus, the smaller the quantum of energy
- moving away from the nucleus, e^- absorb a quantum of energy
- moving toward the nucleus, e^- release (emit) a quantum of energy

Memorize the sequence this way: When an excited state e^- returns to a lower level, a quantum of energy is emitted as light of a specific frequency producing a bright-line spectrum which can be used to identify an element

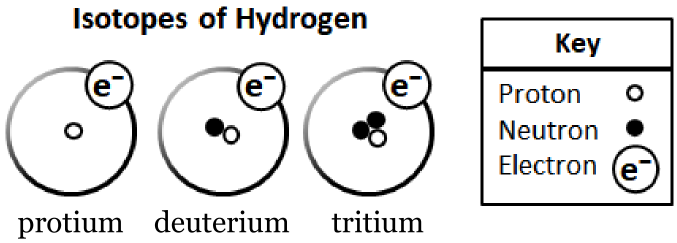
Example: The diagram below shows the bright-line spectra of three elements and a mixture that contains two of these elements. Which elements are in the mixture?



Elements A and D are in the mixture. (The lines at 660 nm and above 700 nm for element X do not appear in the mixture.)

Isotope: atoms that contain the same number of protons (are the same element) but a different number of neutrons are called isotopes of that element

There are three isotopes of hydrogen:



Notations for writing isotopes

Hyphen notation: protium = hydrogen-1, deuterium = hydrogen-2, tritium = hydrogen-3

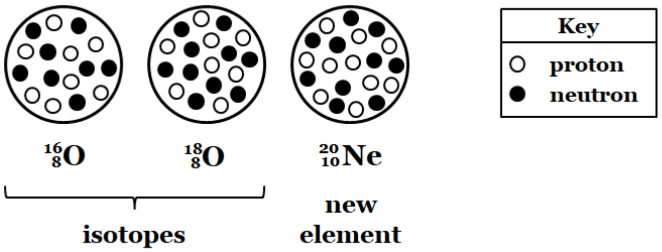
H–1 H–2 H–3

Full nuclear symbol: protium = $^{1}_{1}H$ deuterium = $^{2}_{1}H$ tritium = $^{3}_{1}H$

Atomic Concepts

Unit 2

Nuclear diagrams:



Calculating nucleons in nuclides (easiest using nuclear symbols)

Mass number
Atomic number $\;{}^{19}_{9}\mathrm{F}$

Mass number = number of p^+ + number of n^o (top number, always whole number)
Atomic number = number of p^+ (bottom number)
Number of n^o = mass number – atomic number (top number – bottom number)

Examples:

neutral atoms				ions			
${}^{3}_{1}\mathrm{H}$	$n^o = 2$	$p^+ = 1$	$e^- = 1$	${}^{7}_{3}\mathrm{Li}^+$	$n^o = 4$	$p^+ = 3$	$e^- = 2$
${}^{7}_{3}\mathrm{Li}$	$n^o = 4$	$p^+ = 3$	$e^- = 3$	${}^{16}_{8}\mathrm{O}^{2-}$	$n^o = 8$	$p^+ = 8$	$e^- = 10$

Calculating Averages

Straight average (unweighted): Find the average score for three tests: 80, 85, and 90

$$\frac{80 + 85 + 90}{3} = \frac{255}{3} = 85$$

Weighted average: Find the average score for three tests: 80 worth 20% of your grade, 85 worth 20% of your grade, and 90 worth 60% of your grade

$$\frac{(80 \times 20\%) + (85 \times 20\%) + (90 \times 60\%)}{100\%} = \frac{(1600 + 1700 + 5400)}{100} = 87$$

Weighted average atomic masses

Requires two numbers:

- the percent natural abundance of each isotope
- the atomic mass of each isotope (if the atomic mass is missing, use the mass number)

Example 1: magnesium has three nuclides: ^{24}Mg at 78.99% and 23.985 041 7 amu, ^{25}Mg at 10.00% and 24.985 836 92 amu, and ^{26}Mg at 11.01% and 25.982 592 93 amu.

^{24}Mg: 23.985 041 7 amu x 0.789 9 = 18.945 784 amu
^{25}Mg: 24.985 836 92 amu x 0.100 0 = 2.498 584 amu
^{26}Mg: 25.982 592 93 amu x 0.110 1 = 2.860 683 amu

24.305 amu

Example 2: magnesium has three nuclides: ^{24}Mg at 78.99%, ^{25}Mg at 10.00%, and ^{26}Mg at 11.01%.

^{24}Mg: 24 amu x 0.789 9 = 18.9576 amu
^{25}Mg: 25 amu x 0.100 0 = 2.5000 amu
^{26}Mg: 26 amu x 0.110 1 = 2.8626 amu

24.32 amu (off by only 0.06%)

Go on to the next page

Atomic Concepts

Unit 2

Atomic masses on the Periodic Table are not whole numbers because they are the weighted average of all the naturally occurring nuclides of that isotope

Atomic masses on the Periodic Table rounded to a whole number gives the most common naturally occurring nuclides of that isotope

Example: Li = 6.941 amu and lithium–7 is the most common nuclide

- lithium–7 = 92.5%
- lithium–6 = 7.5%

Lewis Electron Dot Structures

- show valence electrons only
- atomic symbol represents the kernel of the atom (everything inside the valence shell)
- dots represent the valence electrons in the atom

Rules for Lewis Electron Dot Structures

1. Write the element symbol (represents the kernel of the atom)
2. Use dots to show the valence electrons
   a. Pair the 1st and 2nd electrons together $\mathrm{Li}\cdot$ $\mathrm{Be}:$
   b. 3rd, 4th, and 5th valence electron will be alone on separate sides
      $\overset{\cdot}{\mathrm{B}}:$ $\cdot\overset{\cdot}{\mathrm{C}}:$ $\cdot\underset{\cdot}{\overset{\cdot}{\mathrm{N}}}:$
   c. 6th, 7th, and 8th valence electron will pair with any side with one electron
      $\cdot\underset{\cdot}{\overset{\cdot\cdot}{\mathrm{O}}}:$ $:\underset{\cdot}{\overset{\cdot\cdot}{\mathrm{F}}}:$ $:\underset{\cdot\cdot}{\overset{\cdot\cdot}{\mathrm{Ne}}}:$
3. Remember to change the number of electron dots for ions
   a. Remove one dot for each positive charge
   b. Add one dot for each negative charge
   c. Remember the stable octet rule (atoms are more stable with 8 valence electrons)
      H and He have a maximum of 2 valence electrons
   d. Check the Periodic Table to see what ions can form
      $[:\underset{\cdot\cdot}{\overset{\cdot\cdot}{\mathrm{O}}}:]^{2-}$ Mg^{2+} $:\underset{\cdot\cdot}{\overset{\cdot\cdot}{\mathrm{Ne}}}:$