Atomic Concepts

Brief History of Atomic Structure

Name	Date	Theory	Experiment
Democritus	400 BC	atomos – smallest particle	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)
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 circular orbits; a drop in energy level causes a quantum of energy to be emitted	Calculated the energy levels in atoms based on the energy of the bright lines in the spectra of elements
Erwin Schrödinger	1926 AD	wave mechanical model – since e^- have wave properties, wrote a wave function equation that predicts the most probable electron position	e^- move in fuzzy orbitals (most probable position of an e^- with non-distinct e^- clouds)

Dalton's Atomic Theory

1. Matter is composed of extremely small particles called atoms
2. Atoms are indivisible and indestructible
3. Atoms of a given element are identical in size, mass, and chemical properties
4. Atoms of a specific element are different from those of another element
5. Different atoms combine in simple whole-number ratios to form compounds
6. In a chemical reaction, atoms are separated, combined, or rearranged

Atomic structure

Atoms are made of three particles

1. protons (p+) } nucleus
2. neutrons (n^0)
3. electrons (e^-)

Isotopes: atoms of the same element with different numbers of neutrons

Examples: hydrogen isotopes are protium, deuterium, and tritium

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

Full notation: protium = ${}^{1}_{1}H$, deuterium = ${}^{2}_{1}H$, tritium = ${}^{3}_{1}H$

Calculating the number of subatomic particles

$$\text{mass number} \ \ {}^{19}_{\ 9}\mathrm{F}$$
$$\text{atomic number}$$

Mass number	= number of p^+ + number of n^0	(top number in a nuclear symbol)
Atomic number	= number of p^+	(bottom number in a nuclear symbol)
Number of n^0	= mass number – atomic number	

Examples: (for neutral atoms and ions)

${}^{3}_{1}H$	$n^0 = 2$	$p^+ = 1$	$e^- = 1$
${}^{7}_{3}Li^{1+}$	$n^0 = 4$	$p^+ = 3$	$e^- = 2$
${}^{18}_{8}O^{2-}$	$n^0 = 10$	$p^+ = 8$	$e^- = 10$

Weighted average atomic masses

The atomic mass unit (amu) is defined as $\frac{1}{12}$ of a carbon – 12 nuclide

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

Example: 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

Electron configurations

Electron configurations on the Periodic Table will show the number of electrons in each energy level (similar to the Bohr planetary model)

Examples for neutral atoms:		Examples for ions	
Na	2 - 8 - 1	Na^{1+}	2 - 8 - 0
O	2 - 6	O^{2-}	2 - 8
Ar	2 - 8 - 8	Noble gases do not form ions easily	

Atoms with only 1 or 2 valence electrons (metals) tend to lose electrons
Atoms with 7 valence electrons (nonmetals) tend to gain electrons (oxygen gains $2e^-$)
Atoms with 8 valence electrons (noble gases) tend not to form ions (He = $2e^-$)

Valence electrons

Electrons in the outermost electron shell (or energy level)

Find the valence using electron configurations from the Periodic Table

Examples for neutral atoms:

Na	2-8-1	1 valence electron
Br	2-8-18-7	7 valence electrons

LED Structures

Atomic symbols of elements are used to denote the kernel (non-valence shell + nucleus)

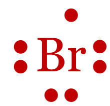
Dots are used to show valence electrons

Find the valence using electron configurations from the Periodic Table

Examples for neutral atoms:

Na 2-8-1 1 valence electron Na•

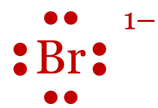
Br 2-8-18-7 7 valence electrons



Examples for ions: (adjust electron count using selected oxidations and show the charges)

Na 2-8-1 loses 1 e^- Na^{1+}

Br 2-8-18-7 gains 1 e^-



Spectra and the Bohr model (or simplified model)

Ground state atom: all the electrons are in the lowest electron shells (lowest energy) possible

Electrons are stable closer to the nucleus because the negatively charged electrons are attracted to the positively charged protons in the nucleus

Electron configurations from the Periodic Table are ground state

Electrons that absorb or gain the exact amount of energy between electron shells can jump up (be promoted) to a higher energy level (electron shell) and enter an excited state

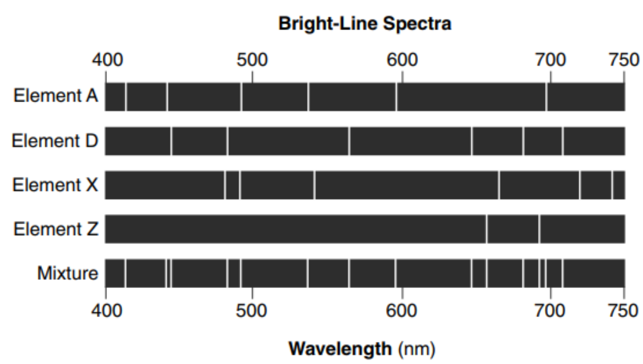
Excited state atom: at least one electron has been promoted to a higher energy level

Excited state electrons are unstable and will relax or drop to a lower electron shell by emitting energy (equal to the gap between electron shells)

The light emitted is quantized and results in the characteristic bright-line spectra

Spectra can be used like fingerprints to identify elements

Example: The diagram to the right represents the bright-line spectra of four elements and a bright-line spectrum produced by a mixture of three of these elements.



Which element is *not* in the mixture?

Answer: X (look between 700 and 750 nm)

Which electron configuration represents a sodium atom in the excited state?

(1) 2-8-1 (2) 2-8-2 (3) 2-7-2 (4) 2-7-1

Answer: (3) [because (1) is ground state, (2) is a negative ion, and (3) is a positive ion]