Chemistry UNIT 3

Name:
Date:

Chemistry Unit 3

Atomic Theory and structure of an Atom

Definitions

- Model: A familiar idea used to explain unfamilar facts observed in nature.
- Theory: An explanation of observable facts and phenomena
  - To remain valid, models and theories must:
    - Explain all known facts
    - Enable scientists to make correct predictions

History of an Atom

- Democritus
  - Proposed the existence of an atom
  - Word comes from the Greek word atomis which means not to cut or indivisible

- Aristotle
  - Rejected the idea of the atom
  - Said matter could be cut continually

- Dalton’s theory proposed that atoms:
  - Are building blocks of matter
  - Are indivisible
  - Of the same element are identical
  - Of different elements are different
  - Unite in small, whole number ratios to form compounds

- J.J. Thomson
  - Credited with the discovery of electron; a blow to Dalton’s indivisible atom
  - Proposed the plum pudding model of the atom: negatively charged electrons embedded in a ball of positive charge

- Rutherford’s Gold Foil experiment:
  - Aimed alpha particles at gold foil
  - Most passed through
  - A few particles were deflected
  - Some particles bounced back

Rutherford’s Experiment

- Most of the atom is empty space
- Dense positively charged core
- Planetary model

Bohr's Model of the Atom

- Nucleons- particles in the nucleus of atom
  - Protons
  - Neutrons
- Atomic number- number of protons in the nucleus of an atom
- Neutral atom- same number of protons (+) and same number of electrons(-)

Isotopes

- Isotopes- atoms of an element that have different numbers of neutrons
- Hydrogen-1
  - ________ proton and _______ neutrons
- Hydrogen-2
  - ________ proton and _______ neutrons
- Hydrogen-3
  - ________ proton and _______ neutrons

Mass number

- Total number of protons and neutrons in an atom
  - Carbon-14
  - Neon-20

Opening

- What is the difference between C-12 and C-14?

Particle Chart

Particle	Charge	Mass	Location
Proton	Positive	1 amu	nucleus
Neutron	Neutral	1 amu	nucleus
Electron	Negative	0	Electron cloud

Atomic mass

- Average of the masses of all the element’s isotopes

Subatomic particles

- # of protons = atomic number
- # of electrons = atomic number
- # of neutrons = mass number – atomic number

Examples

- Iron Fe-56
- Oxygen-17
- He-4
- Calcium-40

Unit 3 Notes

Bohr's Energy Levels

- Electrons in certain energy levels
- Low energy levels are closer to nucleus
- High energy levels are further from nucleus
- Ground state- all electrons are in lowest energy level possible

Excited Atom

- Atom has absorbed energy
- Excited state is unstable
- Atom soon emits same amount of energy absorbed
- Energy is seen as visible light

Wave Description of Light

- Wavelength (λ): distance between corresponding points on adjacent waves
- Frequency (f): the number of waves passing a given point in a given time
- $c = 3.0 \times 10^8$ m/s, speed of light
- $c = f\lambda$

Ex.

- What is the frequency of light if the wavelength is 6.0×10^{-7}m?

Particle Description of Light

- Energy exists as particles called quanta or photons.
- $E = hf$
- h is a constant number called Planck's constant.

The Modern View of Light

- Light has a dual nature
- Light may behave as a wave
- Light may behave as a stream of particles called quanta or photons.

Spectroscopy

- Spectral lines represent energy releases as electron returns to lower energy state
- Spectral lines identify an element
- Called Bright line spectrum of an element

Orbital

- Region of space where an electron is likely to be found

Quantum Numbers

- n, l, m, s
- Used to describe an electron in an atom

n

- Principle quantum number
- Represents the main energy level of electron
- Is always a whole number
- Max. # of electrons in an energy level is $2n^2$
  - What is the maximum number of electrons that can be in the 5^{th} main energy level?

$2n^2 \quad 2(5^2) = 50$ electrons

l

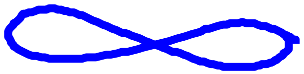
- The 2nd quantum number
- Describes the orbital shape within an energy level
- Number of orbital shapes possible in energy level = n

Orbital shapes

- Designated s, p, d, f
- Level 1: s
- Level 2: s,p
- Level 3: s, p, d
- Level 4: s, p, d, f

S = sphere

p = dumbbell shape



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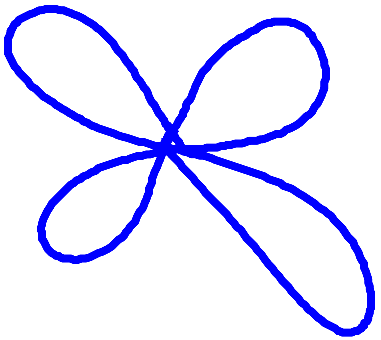

How many electrons can each sublevel hold?

- s = 1 orbital X 2 electrons= 2 electrons
- p = 3 orbitals X 2 electrons = 6 electrons
- d = 5 orbitals X 2 electrons = 10 electrons
- f = 7 orbitals X 2 electrons = 14 electrons

m

- The 3rd quantum number
- Describes orientation of orbital in space
- x, y, z axis

s

- The 4th quantum number
- Describes spin of electron in orbital
- Hund’s Rule- orbitals of equal energy are each occupied by one electron before any orbital is occupied by a second electron
- Pauli Exclusion Principle: No two electrons can have the same four quantum numbers.

Diagonal Rule

- See worksheet