CLASSIFYING MATTER

Pure Substance: matter that cannot be physically seperated (dissolved, magnets, filtered)

mixture: matter than can be physically seperated

element: a pure substance that cannot be chemically decomposed
↳ are box from periodic table

compand: a pure substance that has more than one type of element
↳ Ex. H_2O ; NaCl

Heterogeneous Mixture: a mixture that does not have a uniform composition

Homogeneous Mixture: a mixture that has a uniform composition

DEFINITIONS

law: recognizes a behavior of particles/matter accepted by most scientists
↳ like a theorem

theory: explains the behavior recognized by the laws
↳ like a postulate

Law of Conservation of Matter: Matter cannot be created or destroyed
from Lavoisier who experimented with reactions on scales

Law of Constant Composition: a given compound always has the same elements in the same proportions, no matter the sample size
AKA: Definite Proportions
from Proust: water bottle / swimming pool

Dalton's Atomic Theory:
- All matter is compated of atoms
- Atoms are indivisible/ indestructable* (not true → nuclear fission)
- All atoms of the same element are identical
- They all have the same mass / properties* (not true → isotopes)
- Compounds are formed from 2 or more substances in whole number ratios
- A chemical reaction is a rearrangement of atoms

proven false in later years

Law of Multiple Proportions: If 2 or more different compounds are compated of the same 2 elements, then the ratio of masses of the 2nd element (while holding the 1st element mass constant) is a ratio of whole numbers
Atoms combine in whole numbers

$^{12g}CO\ ^{16g}$
CO_2
$^{12g}\ ^{32g}$

PARTS OF AN ATOM

Protons:	p^+	+1	1 amu	inside the nucleus	subatomic particles
neutron:	n^0	0	1 amu	inside the nucleus	
electron:	e^-	-1	~0 amu	outside the nucleus	

amu : atomic mass unit

Isotopes: elements with the same number of protons (define element) but a different number of neutrons

magnesium	element name
12	atomic number (number of protons)
Mg	element symbol
24.305	average atomic mass (average of ALL isotopes)

* pay attention to capitals / lowercase: CO : carbon monoxide Co : cobalt *

Nuclide Notation: shows specific isotopes isotope mass ← $^{24}_{12}Mg^{2+}$ → charge; Atomic # ← 12; Mg → element; (ionization state)

mass # : $(p^+) + (n^0)$ charge : $(p^+) - (e^-)$

Percent Abundance: $\frac{\#\text{ of one type of isotope}}{\text{total \# of atoms}} \times 100$

Weighted Average: A method of calculating an arithmetic mean where some data carries more weight than others

→ (isotope 1 mass × % abundance) + . . . + (isotope n mass × % abundance) = AAM

* make sure to convert % abundance into decimal format *

ELECTRO MAGNETIC RADIATION (EMR)

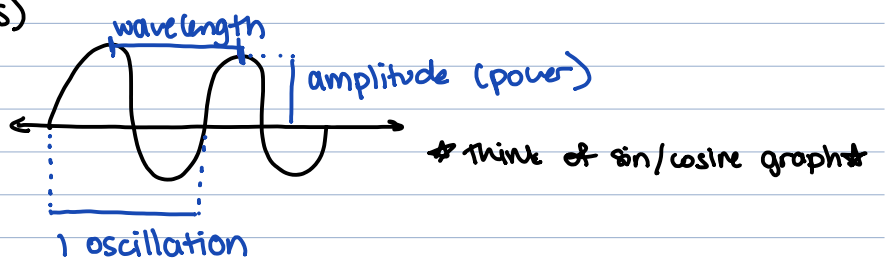
All EMR travels at the same speed: c : speed of light : 3.00×10^8 m/s

Visible light is only a small portion of the EMR spectrum

wavelength (λ) is the distance between 2 neighboring peaks (measured in meters (m))

frequency (ν) is the number of cycles / oscillations of a wave that pass per sec measured in Hertz (1/s)

amplitude : the height of wave



* think of sin/cosine graph *

Wavelength and frequency are indirectly related

low frequency : long wavelength high frequency : short wavelength

A photon : particle of EMR with 1 quantum of energy photon : piece of light

Quantum : piece of energy

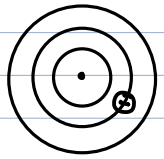
A photon does not have one specific energy value assigned to it, energy depends on the frequency of the wave

$\lambda\nu = c$ wavelength × frequency = speed of light

$E = h\nu$ Planck's constant × frequency = energy Planck's constant : 6.626×10^{-34} Js
Js : Joules × seconds

ELECTRONS AND THEIR QUANTUM STATES

Bohr model: Hydrogen has 1 proton and 1 electron
In the ground state, hydrogen's one electron is in the energy level that is closest to the nucleus
In order to move to higher energy levels (excited state) the electrons must absorb a photon (piece of light)
In order to move to a lower energy level, the electron must release a photon



* energy level closest to nucleus has the lowest energy level *
* electrons always fill a low energy level first *

elements release light when their electrons move from higher energy levels to lower energy levels

The amount of energy released depends on the distance between energy levels

Unique to each element, sometimes called an "atomic fingerprint"
↳ Absorption and Emission Spectra (color of light given off)

light and electrons both behave as particles and waves

Bohr model vs. modern atomic theory (quantum theory):
Bohr model shows that electrons orbited the nucleus in flat, fixed orbitals (shells) in circular paths

modern view atomic electron structure is more like 3D standing waves
↳ more accurate

MODELING ATOMS

ground state: when all the electrons in an atom are assigned to energy levels/ sublevels closest to the nucleus (neutral)

excited state: when electrons are promoted to higher energy levels by absorbing energy

* They have the same number of electrons but some energy levels may have more/less electrons in them *

Atomic orbital: a function describing the location and wave-like behavior of an electron in an atom

energy level: the levels where electrons are located

sublevel: a 3D shape that can be found in a given energy level

orientation: how the sublevel is arranged in space (layout)

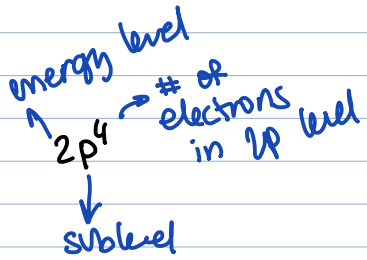
orbital: 3D representation of each orientation (parts of the things that make a sublevel)

s sublevel: each energy level has an s sublevel
spherical shape, only 1 orientation

p sublevel: All energy levels 2 and higher have p sublevels
3 orientations
Dumbell shape

d sublevel: All energy levels 3 and higher have d sublevels
5 orientations

f sublevels: energy levels 4 and higher have f sublevels
7 orientations



each orbital can have a max of 2 electrons

max electrons in s sublevel : 2 → 1 orbital x 2 : 2
max electrons in p sublevel : 6 → 3 orbital x 2 : 6
max electrons in d sublevel : 10 → 5 orbital x 2 : 10
max electrons in f sublevel : 14 → 7 orbital x 2 : 14

Each energy level increases in energy and is further from the nucleus
↳ electron nee s more energy to move to a higher and higher energy level

Each electron in orbital must spin in opposite directions

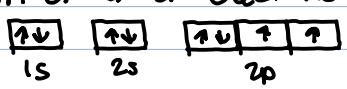
Afbau Principle: an electron occupies the lowest energy orbital that can receive it

Pauli Exclusion Principle: electrons in the same orbital cannot spin the same direction

Hund's Rule: orbitals of equal energy are filled with one electron first before pairing (maximized number of unpaired electrons)

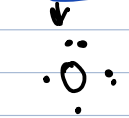
Electron configuration: condensed version of orbital diagram
show # of electrons in each level and sublevel, but not spin of electrons
oxygen: $1s^2\ 2s^2\ 2p^4$

Orbital diagram: shows spin of # of electrons in each level and sublevel
oxygen:



Electron Dot Diagrams: Give us info about valence electrons

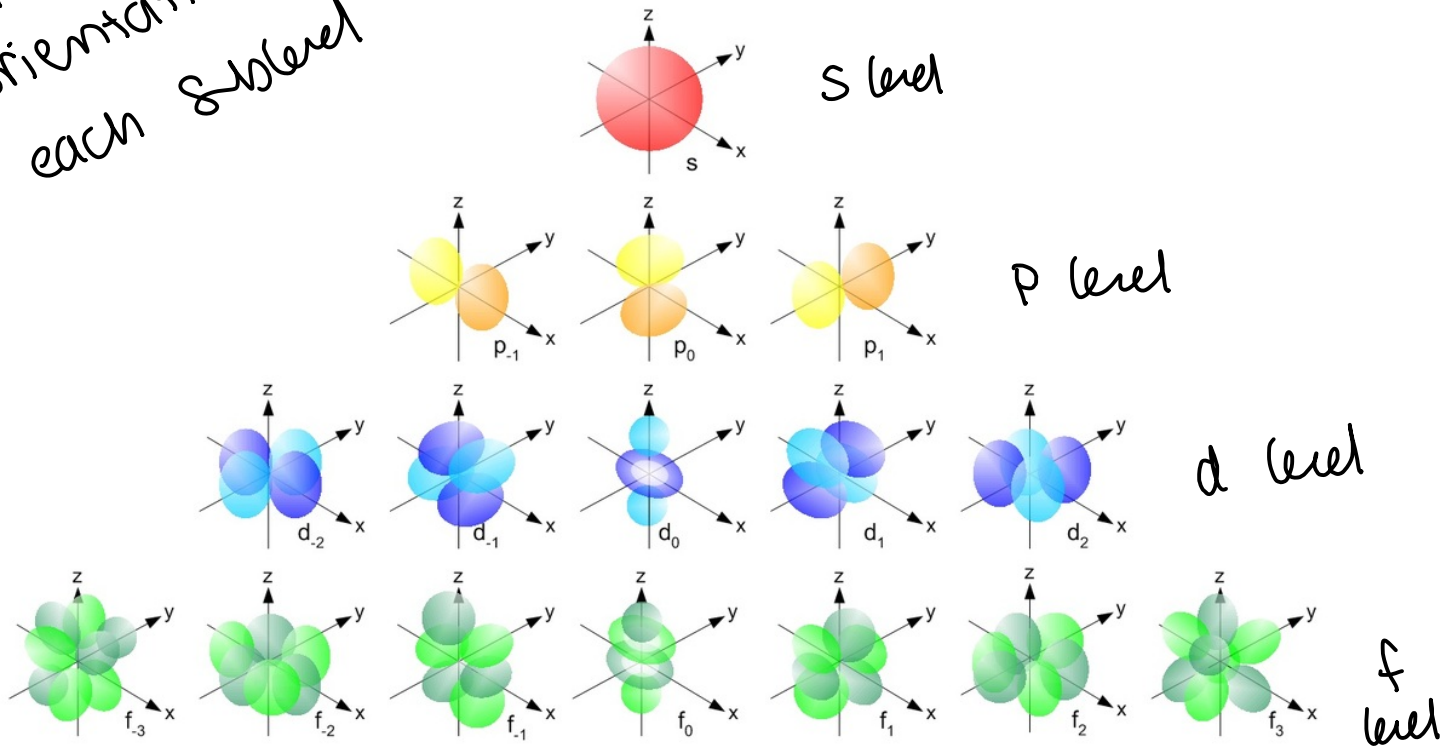
oxygen: $1s^2 2s^2 2p^4$



Valence electrons: electrons in the highest energy s and p sublevels

the different orientations of sublevels below:

Different orientations of each sublevel



IONS

ion: atom that has lost/gained electrons and has a charge

octet rule: atoms will lose, gain, or share electrons in order to have a full valence shell (makes the outermost electron shell more stable)

A shell has full valence shell when the highest energy levels has 8 electrons

Is is the only energy level/sublevel that has full valence, but does not have 8 electrons

atoms will try to lose/gain the least amount of electrons possible
1-3: lose 4: either 5-7: gain

Positive ions: cations Negative ions: anions

Mg^{2+}: $1s^2 2s^2 2p^6$ Mg: $1s^2 2s^2 2p^6 3s^2$ → lost 2 negative electrons

O^{2-}: $1s^2 2s^2 2p^6$ O: $1s^2 2s^2 2p^4$ → added 2 negative electrons