MOLES:

1 mole: 6.02×10^{23} things
↳ Avogadro's number

molar mass: mass of 1 mole of a substance
varies because the size of atoms / molecules vary
units = g/mol

$$\text{Grams} \xrightarrow{\times \frac{1 \text{ mole}}{\text{molar mass (g)}}} \text{moles} \xrightarrow{\times \frac{6.02 \times 10^{23} \text{ particles}}{1 \text{ mole}}} \text{particles}$$

$$\text{Grams} \xleftarrow{\times \frac{\text{molar mass (g)}}{1 \text{ mol}}} \text{moles} \xleftarrow{\times \frac{1 \text{ mol}}{6.02 \times 10^{23} \text{ particles}}} \text{particles}$$

Particles: atoms, molecules, formula unit, ions

atom ex: Fe, Ca
molecule ex: H_2O CO_2
formula unit ex: NaCl
ion: Mg^{2+}

FEATURES OF THE PERIODIC TABLE

octet rule: atoms are most stable when their valence shell are filled with 8 electrons

Groups:
- group 1: Alkali metals
- group 2: Alkaline Earth metals

} s orbital

- middle box (d orbital) = Transition metals
- group 13: Boron family
- group 14: carbon family
- group 15: nitrogen family
- group 16: Chalogens
- group 17: Halogens
- group 18: Noble gases

} p orbital

Outer Transition Metals
↳ d block

Inner Transition Metals
↳ f block

non-metals and metalloids are located in the p orbital box except for Hydrogen

non-metals: Hydrogen, Helium, carbon, Nitrogen, Oxygen, Fluorine, Neon, Phosphorus, sulfur, chlorine, Argon, Selenium, Bromine, Krypton, Iodine, Xenon, Radon, Oganesson (18)

metalloids: Boron, Silicon, Germanium, Arsenic, Antimony, Tellurium, Polonium, Astatine (8)

metals: everything else (92)

PROPERTIES OF ELEMENTS

metals:
- high melting/boiling temp.
- malleable (bendable)
- shiny and hard
- ductile (can be wired)
- good conductors of heat/electricity

non-metals:
- brittle if solid
- poor conductors
- not malleable
- dull (most = gas)
- low boiling/melting

metalloids:
- mix of metals/non-metals
- semi-conductive
  ↳ can conduct/insulate

COULOMB'S LAW

$$F_e = \frac{kq_1q_2}{r^2}$$

F_e = force k: constant (ignore) q_1 / q_2 = charges r: distance
Think of q_2 as the ENC and r as the atomic radius

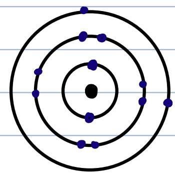
Coulomb's law: compares the product of charge on 2 particles
compares the distance between particles

[-] → ← [+] "opposites attract"
← [+] [+] →

The larger the product of two oppositely charged particles, the stronger the force of attraction
The larger the product of two same charged particles the stronger the force of repulsion

* negative fe = attraction positive fe = repulsion *
* larger = magnitude / absolute value, not number-line *

Bohr Model:



The positively charged nucleus tries to pull the negatively charged electrons to the center (opposites attract)

But the negative electrons repel each other in the outer energy levels (likes repel)

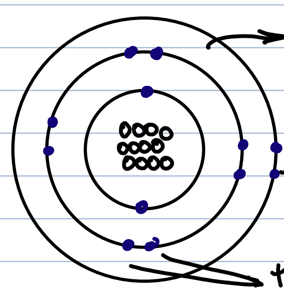
EFFECTIVE NUCLEAR CHARGE

ENC / Zeff: The attractive force that an electron "feels" from the positively charged nucleus
↳ the nuclear charge - sheilding / inner-shell electrons
p^+ - non-valence electrons

outer shell electrons are sheilded from the nucleus by inner shell electrons because like charges repel

Electrons in the same energy level don't sheild each other to a great extent

Bohr Model: Magnesium: (12)



→ p^+ - inner-shell electrons:
12 - 10 = 2+

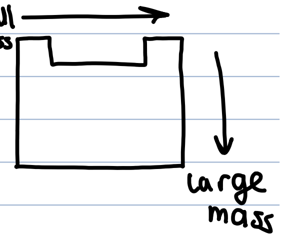
both of these electrons feel a charge of +2 from the nucleus

these electrons feel a charge of +10 from the nucleus → 12 - 2 = +10

electron sheilding: electrons are all negative, so e^- in the inner energy shell "sheild" e^- in the outer shells from feeling the full attraction of the nucleus because of electron sheilding

connection to coulomb: force and charge are directly proportional (if ENC increases, force increases too) higher ENC means there is a higher force of attraction

ATOMIC MASS TREND

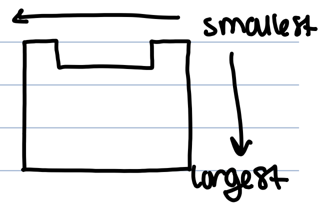


Atomic mass: tends to increase across the period
↳ more protons and neutrons are being added across period
protons ~ 1 amu neutrons ~ 1 amu
electrons ~ 0 amu, so they are negligible

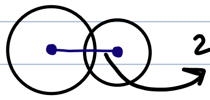
Atomic mass tend to increase down a group
↳ Atoms get more heavier as there is a large jump in # of protons/neutrons

Exceptions: Ar and K — Ar (18) = 39.95 K (19) = 39.10
Co and Ni — Co (27) = 58.93 Ni (28) = 58.69
Te and I — Te (52) = 127.60 I (53) = 126.90
weighted averages
↳ has to do with isotopes and relative abundance

ATOMIC RADIUS TREND



Atomic Radius: 1/2 the distance between the nucleus of identical atoms that are bonded together.

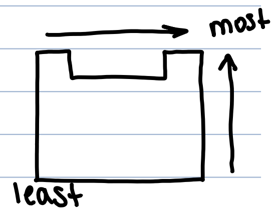

2 times atomic radius (nucleus = center)

Atomic radius tends to decrease across the period
↳ caused by increasing effective nuclear charge (pulls electrons closer, shrinking radius)

Atomic radius tends to increase down a group
↳ the addition of energy levels increases the electron's distance from the nucleus and the size of the atom
electron-electron repulsion "plumps" up the atom

Connection to Coulomb: force of attraction from the nucleus for outermost electrons decreases the further the electrons are from the nucleus
↳ radius and force are inversely proportional (if radius increases, force of attraction decreases)

IONIZATION ENERGY TREND



Ionization Energy: the energy required to remove one electron from neutral atom forms an ion

$A + \text{energy} \rightarrow A^{1+} + e^-$
$Li\,(1s^2 2s^1) + \text{energy} \rightarrow Li^+\,(1s^2) + e^-$

IE tends to increase across the period
↳ caused by increase in ENC → more ENC = stronger valence electrons, which means that atom needs more energy to remove an electron

Noble gases have the highest first ionization energies (highest IE)

IE tends to decrease down a group
↳ Outermost electrons are farthest from nucleus
electron-electron repulsion forces increase, making it easier to remove an electron

Connection to Coulomb: relates to ENC/Atomic radius relationship with the equation
higher ENC → higher force of attraction → higher IE
greater Atomic radius → lower force of attraction → lower IE

high IE → hard to remove electron   low IE → easy to remove electron

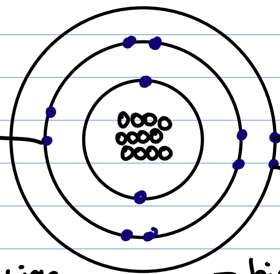
Sucessive Ionization Energies: 1st IE = energy needed to remove 1 electron
2nd IE = energy needed to remove 2 electrons ...

Each sucessive electron removed from an ion feels a stronger pull from the nucleus because there are the same number of protons pulling on fewer electrons

It is harder to remove an electron when it feels a stronger pull from the nucleus

A large jump in ionization energy occurs when removing an electron from an ion that assumes a noble gas configuration

Bohr Model: Magnesium: (12)



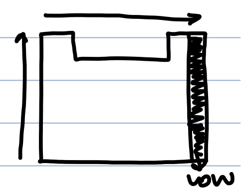
→ energy needed to remove this electron = 738 kJ/mol (1st IE)

↳ 2nd IE = 1450 kJ/mol

After the outer shell is gone, the next ionization energy = 7730 kJ/mol → big jump cause this electron's ENC = (12-2) = 10+
very high ENC, so hard to remove electron
Also, smaller Atomic radius, making it harder

ELECTRON AFFINITY TREND

Electron Affinity: the change in energy that a neutral atom undergoes when an electron is acquired (ability to attract electron)



$A + e^- \rightarrow A^{1-} + \text{energy}$

$Cl\ ([Ne]\ 3s^2 3p^5) + e^- \rightarrow Cl^{1-}\ ([Ne]\ 3s^2 3p^6) + \text{energy}$

negative energy value means that energy is being given off (exothermic)
the more negative the value:
↳ more willing an atom is to accept an electron   * opposite of IE *
more stable the anion is (negative ion)

Positive EA would mean that an atom is forced to gain an electron
↳ energy is absorbed (endothermic) $A + e^- + \text{energy} \rightarrow A^-$

EA tends to increase across the period (energy is more negative)
↳ higher ENC and smaller atomic radius makes it easier to accept an electron
non-metals want to gain electron(s) to have full octet shell

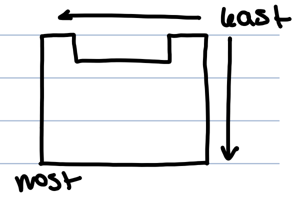
EA tends to decrease down a group (becomes more positive)
↳ Atomic radius increases, making force of attraction between protons and valence electrons weaker
makes it more difficult to add an electron

Exceptions: Noble gases have a low EA because they are stable on their own, and to add another electron, they would have to add an extra energy level, giving it a weak shell

Group 2A has a slightly smaller EA because an electron would have to go in the p orbital, and there would be more electron-electron repulsion

Group 5A has a low EA because it would force the electron in the p orbitals to pair, increasing electron-electron repulsion

REACTIVITY OF METALS AND NON-METALS

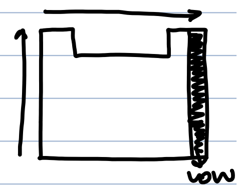


metals tend to react in order to loose their few valence electrons

Reactivity of metals: to determine reactivity, look at the metal's IE

metals toward the bottom left are more reactive
↳ Alkali metals > Alkaline Earth > transition
large metals > small metals

IE and reactivity for metals are inversely proportional
↳ high IE for metals : low reactivity



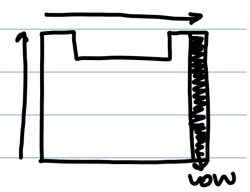
non-metals tend to react in order to gain electrons for valence shell

Reactivity of non-metals: to determine reactivity, look at EA

non-metals towards the upper right corner are the most reactive, excluding the noble gases

EA and Reactivity are directly proportional
↳ high EA for non-metals: high reactivity

ELECTRONEGATIVITY



electronegativity: the ability of an atom to attract a bonding pair of electrons in a covalently bonded compound

Electronegativity increases across period
↳ increasing ENC and decreasing radius

★ follows same rules as EA ★

Electronegativity decrease down group
↳ increasing atomic radius

★ Electronegativity is a property of compounds, not singular atoms ★

in many chemical reactions the electrons involved in bonding are closer to one atom than another

↳ because one atom has higher reactivity

Electronegativity makes a compound Polar : electrons are not evenly shared

H • • F → hydrogen and fluorine share 2 electrons

H ••F → Flourine attracts electrons close to itself because it has a higher EA and is willing to accept more electrons

$H^{\delta+}$ ••$F^{\delta-}$ → makes hydrogen partial positive and Fluorine partial negative

↳ so... Electronegativity of F > Electronegativity of H

Electronegativity follows same rules as EA because EA measures how willing an element is to accept electron

high EA = high electronegativity

★ Noble gases have low EN because they have a low reactivity / low EA ★

IONIC RADIUS

ion: an atom / group of bonded atoms that have positive/negative charges

cations: positively charged ion (lose electrons)

anions: negatively charged ion (gain electrons)

Ionic Radius of Cations: cations are smaller than original neutral atoms
↳ loss of energy level
overall positive charge, puls electrons tighter toward nucleus
higher ENC
larger the positive charge, smaller cation

Ionic Radius of Anions: Anions are larger than original neutral atoms
↳ overall negative charge results in greater repulsion between the electron
electron cloud "swells" or spreads out to minimize repulsion
↳ move far away enough to not move away further
larger the negative charge, larger the anion

★ even though ENC is same for neutral atom/ anion, radius = little larger ★

Isoelectronic: ions that have same number of electrons
Ex. O^{2-}, F^-, Ne, Na^+, Mg^{2+} = 10 electrons

when comparing isoelectronic cations and anions the anions will generally have a larger radius (more protons = smaller radius)

when comparing ions of similar charge, size trend : same as atomic radius