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**IDX G9 (SUBJECT) (LEVEL) STUDY GUIDE ISSUE 5**

**By Hayley and Allison**

7.3 Sizes of Atoms and Ions

Sizes of Atoms:

- The bonding atomic radius is defined as one-half of the distance between covalently bond nuclei.

- Across a period, from left to right, smaller radius

- Atoms’ nuclei gain more proton, increase the effective nuclear charge, which exerts a stronger pull on electrons in the same principle energy level.

- Down a group, larger radius

- Atoms have more principle energy level, which larger orbitals means that the valence electrons are further from the nucleus.

Sizes of Ions:

- Ionic size depends on:

1. nuclear charge
2. number of electron
3. orbitals in which electrons reside

- Cations (positive charge ions) are smaller than their parent atoms

- Outermost electron removed & repulsion reduced

- Antions (negative charge ions) are larger than their parent atoms

- Electron added & repulsion increase

- Ionic size increase down group

- In isoelectronic series, ions with same number of electrons, ionic size decreases with an increasing nuclear charge.

7.4 Ionization

Ionization Energy:

- Amount of energy required to remove an electron (valence electron first) from the ground state of a gaseous state of a gaseous atom or ion (akways positive). (Minimum energy required to remove an electron)

- First ionization energy, second ionization energy, etc.

- Requires more energy to remove each successive electron

First ionization energy: energy required to move 1st electron

- Down a group, less ionization energy

- Zeff issame, valence electrons further away from nucleus

- Across period, ionization energy increase (harder to remove electron)

- Zeff increase

- Exceptions: discontinueties

- Group 5A to 6A (double occupied orbitals)

- Group 2A to 3A (electron taken from s orbital which is more attracted to nucleus)

When all valence electrons have been removed (it takes from noble gas core), the ionization energy takes a quantum leap.

7.5 Electron Affinity

Electron affinity:

- Energy change when adding an electron to gaseous atom.

- The more negative of the electron affinity, the greater attraction to the added electron

- Electron affinity are more exothermic across period and down a group (easier to gain electron, more negative)

- Exception:

* Group 2A: Added electron goes to p orbital, takes more energy going to higher energy level
* Group 5A: Has no empty orbital, extra electron goes to occupied electron orbitals, which creates repulsion

Electronegativity:

- The ability of an atom to attract electrons when the atom is in a compound

* Pauling scale: 0.7 - 4.0

- Down a group, decreases

* Valence shell are getting less strongly held by the nucleus. Harder for the atom to attract one electron.

- Across period, increases

* Attractions from the nucleus is getting greater. Getting easier for atom to attract one electron.

7.6 Metal, Nonmetals, and Metalloids

- Across preiod, metallic characters decrease

- Down a group, metallic characters increase

Metals:

- Elements that are generally shiny, malleable, and ductile

- Good conducters

- Metals tend to form anions

- Compounds for between metals and nonmetals tend to be ionic

- Metal oxides tend to be basic

Nonmetals:

- Occupies the upper right side of periodic table

- Generally gases or brittle, dull-looking solids

- Poor conducter

- Subtances containing only nonmetals are molecular compounds

- Most nonmetal oxides are acidic

Metalloids:

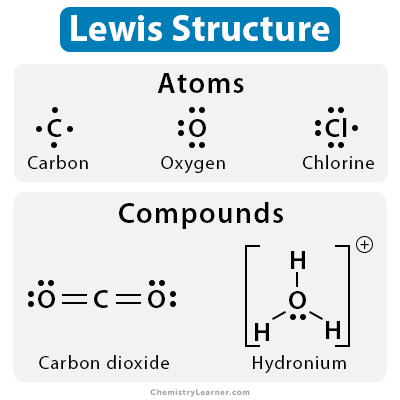
- Elememts with physical & chemical properties of both metal and nonmetal

- fairly poor conducter

Chapter 8: Basic Concepts of Chemical Bonding

* Coulomb’s law:

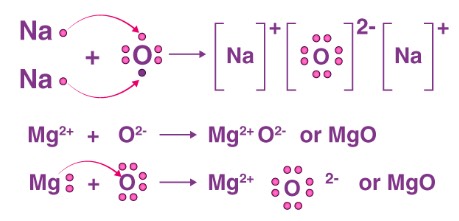
formation of chemical bond — Chapter 8: Basic Concepts of Chemical Bonding

* Coulomb’s law:
  + formation of chemical bond —> more stable —> less potential energy
* chemical bond: the electrostatic attraction force that holds two atoms together

8.1 Lewis Symbols and the Octet Rule

* octet rule: all electrons tend to have 8 electrons in the outermost shell (stable octet)
  + attained through losing, gaining, or sharing electrons
* lewis symbol: an element’s chemical symbol with a dot for each valence electron

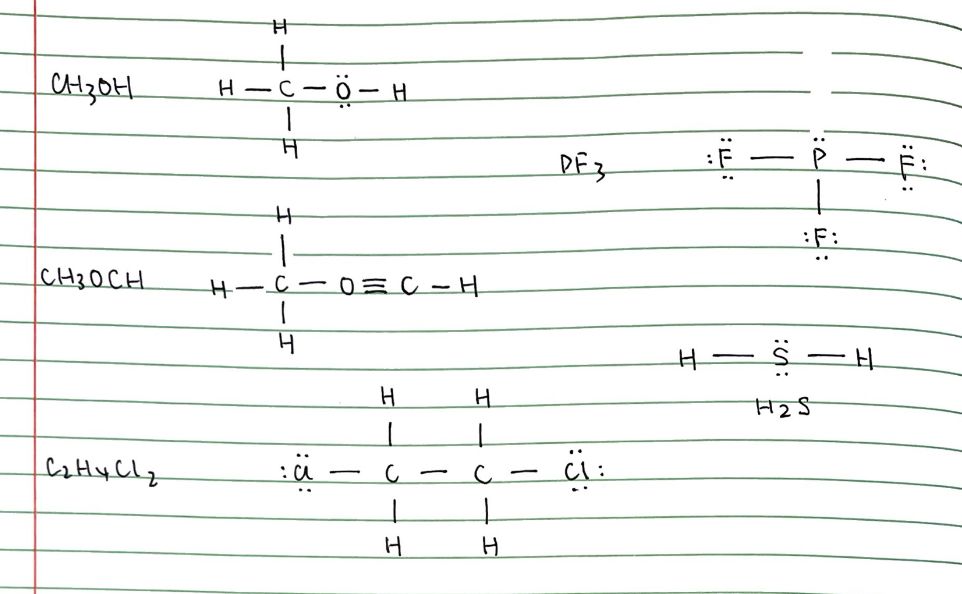
8.2 Ionic Bonds

* ionic bond: attraction between cations and anions
  + usually metals and nonmetals, i.e. NaCl (Na+Cl-)
* elements with low electronegativity (metals, H) tend to form cations
  + lose electrons because number valence electrons is a lot less than 8
  + usually exothermic
* elements with high electronegativity (nonmetals) tend to form anions
  + charge of monoatomic anion = |8 - number of electrons|
  + exothermic process
* representative elements (Mg, Al, etc.): charge of ion = number of valence electrons (group number)
* transition elements (Cr, Fe): various elements with different charges, reference the packet
* 1B to 4A periods 4-6: pseudo-noble gas notation, relatively stable
* structure of ionic compounds
  + do not exist as discrete units but as collections of charged ions in patterns
  + formula unit: lowest whole number ration of ions in an ionic compound (chemical formula)
* properties of ionic compounds
  + high melting point and hardness, brittle
  + do not conduct electricity as solids, but conduct well when molten or aqueous
* lattice energy: energy required to completely separate one mole of a solid ionic compound into its gaseous ions
  + used to measure strength of ionic bond
  + governed by ionic charges and radius

8.3 Covalent Bonding

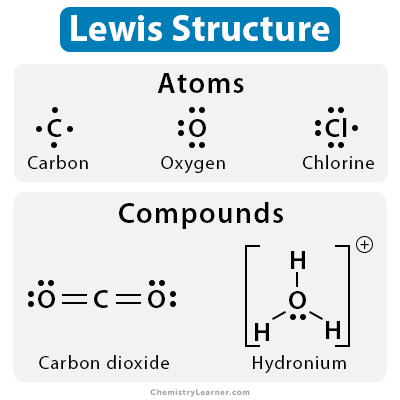
* covalent bond: the electrostatic attraction force between sharing electrons and nucleus
  + formed by sharing of valence electrons
  + occurs between atoms that have high electronegativity
* if electronegativity difference < 2.0, covalent bonding occurs
* single covalent bonds: H2, N2, O2, etc. diatomic elements
* triple > double > single bonds for strength
* single > double > triple bonds for length
* bond disassociation energy: energy needed to break 1 mole of bonds
  + correlated with strength
  + H2(g) —> 2H(g)
  + breaking of bond is an endothermic process
* atoms form longer bonds going down a group

|  | **Sigma Bonds** | **Pi Bonds** |
| --- | --- | --- |
| **orientation** | linear and head-on | parallel and sideways |
| **components (orbitals)** | s-s, s-p, p-p | p-p |
| **dense electron area** | along internuclear axis | above and below internuclear axis |
| **rotation** | yes | no |
| **strength** | stronger | weaker |
| **bonds** | 1 in single, double, and triple bonds | 0 in single bonds, 1 in double, 2 in triple |

examples:

electron dot structure: only dots

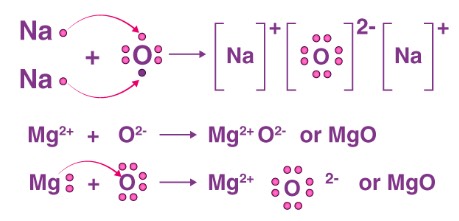
structural formula: bonds as lines and dots

* + > more stable —> less potential energy
* chemical bond: the electrostatic attraction force that holds two atoms together

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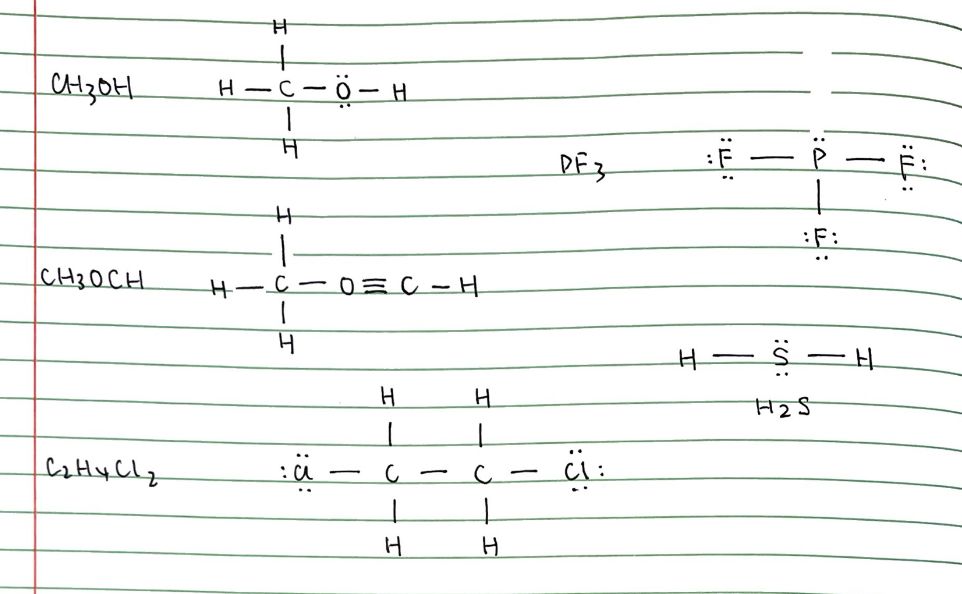
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