Elise Bonin

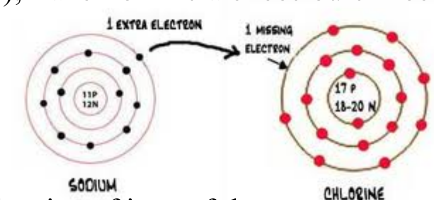
Chapter 8: Chemical Bonding

Sec. 1: Chemical Bonds, Lewis Symbols, & The Octet Rule

- Chemical bond – between atoms or ions that are strongly attached to one another
- Bonds formed by outermost electrons of the atoms, valence electrons, interacting
- Metallic bond – bind atoms in metals (i.e. copper, iron, aluminum)
- Octet rule – atoms tend to gain, lose, or share electrons until they are surrounded by eight valence electrons

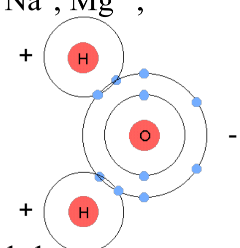
Sec. 2: Ionic Bonding  Na(s) + ½ Cl(g) → NaCl(s)

- Ionic bond – electrostatic forces that exist between ions of opposite charge, formed by transfer of one or more electrons from one atom to another, form three-dimensional lattice made of charged particles
- Stability of ions b/c strong electrostatic attractions between one ion and surrounding oppositely charged ions
- Lattice energy – energy required to completely separate an ionic lattice into its gaseous ions; a measure of the magnitude of interactions between ions, increases w/increasing charge of ions & decreasing radaii
- Potential energy of 2 interacting charged particles: $E=k(Q_1Q_2)/d$; Q_1&Q_2 are charges on particles, d is distance between centers, k is 8.99×10^9 J-m/C^2
- Elements, besides transition metals, usually gain/lose electrons so as to achieve noble-gas configuration (most commonly metals form cations, nonmetals form anions); transition metals lose outermost s electrons before d
- Electron configuration of ions = that of the neutral atoms plus or minus number of electrons lost/gained
    - Na^+ $1s^22s^22p^6$ = [Ne]



Sec. 3: Sizes of Ions

- Cations smaller, anions larger than neutral atoms; generally the size of ions of the same charge increases going down the column of periodic table
- Isoelectronic series (of ions) – ions which possess same number of electrons (i.e. O^{2-}, F^-, Na^+, Mg^{2+}, Al^{3+}); size decreases as nuclear charge increases b/c electrons more attracted to nucleus

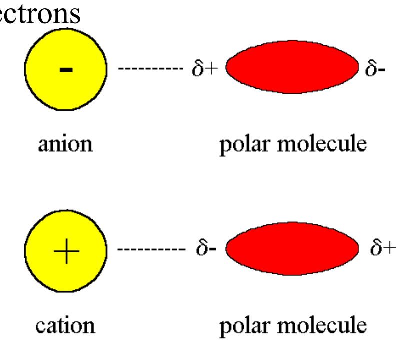


Sec. 4: Covalent Bonding

- Covalent bond – sharing of elections between two atoms, usually both nonmetallic
    - Single bond – sharing of one pair of electrons constitutes a single covalent bond;
    - Multiple bonding: double bond – 2 pairs; triple bond – 3 pairs, etc.
- Lewis symbol – chemical symbol for element plus dot for each valence electron
- Lewis structures – representative diagram with shared electrons shown as a line, unshared shown as dots, to show how valence electrons form bonds and remaining unshared electrons

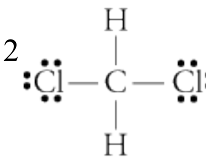
Sec. 5: Bond Polarity & Electronegativity

- Bond polarity – measure of how equally electrons shared between 2 atoms in chemical bond
    - Non-polar covalent bond – electrons shared equally between two atoms
    - Polar covalent bond – one of the atoms exerts a greater attraction for bonding electrons
- Electronegativity – measure of ability of an atom in a molecule to attract electrons to itself; ranges from 0.7 to 4.0; used to determine polarity by finding difference in electronegativity (the greater difference means the more polar the bond)
    - Increases left to right, decreases going down
    - Fluorine is most electronegative so greatest ability to attract electrons
- Polar molecule – molecule whose centers of positive and negative charge don't coincide
    - Dipole – two electrical charges of equal magnitude but opposite sign are separated by a distance
    - Dipole moment – quantitative measure of magnitude of dipole; increases w/increasing amount of charge separated and increasing distance between charges

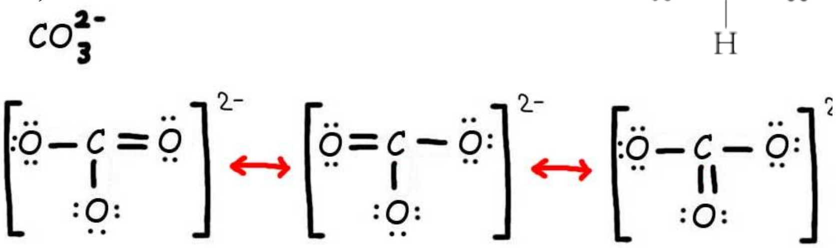


Sec. 6: Drawing Lewis Structures

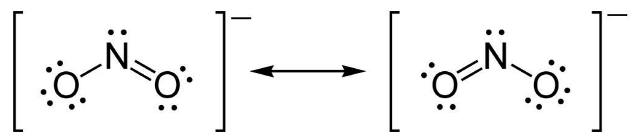
- Formal charge – charge that an atom in a molecule would have if all atoms had same electronegativity
    - To calculate: number of valence electrons minus electrons assigned in Lewis structure (atoms are assigned all unshared electrons, ½ bonded electrons)
- Steps to draw Lewis structure: 1. Sum valence electrons from all atoms 2. Write symbols for atoms to show which atoms connected to which and show attachment w/single bond (–) 3. Complete octets of atoms bonded to central atom 4. Place leftover electrons on central atom 5. If not enough electrons for central atom to have full octet, use multiple bonds
- EX: Lewis Structure for CH_2Cl_2: Valence electrons: 4+2(1)+2(7)=20; Carbon surrounded by 2 hydrogens and 2 chlorines; 4 bonds=8 electrons; 20-8=12 unshared electrons



Sec. 7: Resonance Structures

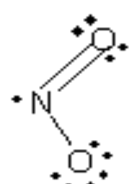


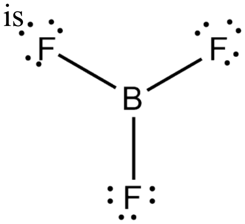
- Used when single Lewis structure does not accurately represent molecule/ion b/c it is envisioned as a blend of multiple structures
- Resonance structures – equivalent Lewis structures except for placement of electrons
- EX: Draw resonance structures for NO_2^-
  Valence electrons: 5+ 2(6) + 1 =17
  2 single bonds leaves too many unshared electrons;
  2 double bonds leaves to few so 1 single and 1 double
  Each has complete octet

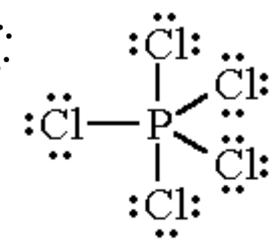


Sec. 8: Exceptions to the Octet Rule

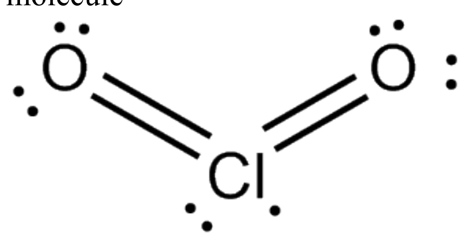
- Molecules w/odd number of electrons where complete pairing is impossible (i.e. NO_2)
- Molecules w/less than octet (usually boron or beryllium) (i.e. BF_3)
- Molecules w/more than octet (i.e. PCl_5)
- Consider the formal charge when choosing which structure is most likely and should be used to represent the molecule







- EX: Draw the Lewis Structure for ClO_2
  Valence electrons 7+2(6)=19
  Cl surrounded by 2 O atoms: w/2 single bonds formal charge=-1 for O, +2 for Cl; w/2 double bonds f.c. = 0 for O, 0 for Cl so double bonds is best structure



Sec. 9: Strengths of Covalent Bonds

- Bond enthalpy – strength of covalent bond measure by molar enthalpy change, ΔH, for breaking of particular bond in a mole of gaseous substance; used to estimate enthalpy change during chemical reactions where bonds are broken and new ones formed
- Enthalpy of reaction: ΔH=Σ(bond enthalpies of bonds broken) – Σ (bond enthalpies of bonds formed); if ΔH>0 endothermic, if ΔH <0 exothermic
- Strengths increase w/ number of electron pairs shared between 2 atoms
- Bond length – distance between nuclei of bonded atoms; decreases as number of bonds between atoms increases
- EX: Estimate ΔH for H−N−N−H(g) →N≡N(g) + 2H−H(g)
  (each N also bonded to an H: | | H H)

  Must break 4 H−N bonds and 1 N−N bond; Must form 1 N≡N bond and 2 H−H
  ΔH= 4D(H−N) + D(N−N) – D(N≡N) – 2D(H−H)
  =4(391kJ) + (163kJ) – (941kJ) – 2(436kJ) = 1727kJ – 1813kJ
  = –86kJ