Chemical Bonds

Prentice Hall *Physical Science* – Chapter 6

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Bonding

- Why are Noble Gases inert?
- Stable electron configuration - the highest occupied energy level of the noble gas atom is filled (eight valence electrons in the case of all noble gases except He, which has 2).
- Other elements tend to react to achieve the same stable electron configuration as the Noble Gases.

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Electron Dot Diagrams

- An *electron dot diagram* is a model of an atom in which each dot represents a valence electron, and the element symbol represents the nucleus and all other electrons of the atom.
- Examples:

Na•   •Mg•   •Ċ•   :F̤:•

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Ionic Bonds – Transfer of e^-

- Atoms of elements that do not have a complete set of valence electrons tend to react with other atoms to form stable electron configurations.
- Some elements do this by a *transfer of electrons*.

$$Na\cdot + \cdot\ddot{Cl}: \longrightarrow Na^{+} :\ddot{Cl}:^{-}$$

- When an atom gains or loses electrons, the number of protons and electrons is no longer equal. It now has a net positive or negative charge and is an *ion*.

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Ionic Bonds – Formation of Ions

Cations = Positive

- If an ion has more protons than electrons, it has a net *positive* charge, and is called a *cation*.
- Represented by a superscript plus sign to the right of the element symbol. Example: Na^+
- Named after their elements. For example, Na^+ is called the *sodium* ion.

Anions = Negative Charge

- If an ion has more electrons than protons, it has a net *negative* charge, and is called an *anion*.
- Represented by a superscript minus sign to the right of the element symbol. Example: Cl^-
- Named by using part of the element name plus the suffix *–ide*. For example, the Cl^- ion is called the *chloride* ion.

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Formation of Ionic Bonds

- *OPPOSITES ATTRACT*: Once ions are formed from the transfer of electrons, the positive and negative ions are attracted to each other and form a *chemical bond.*
- A *chemical bond* is the force that holds atoms or ions together as a unit.
- An *ionic bond* is the force that holds cations and anions together.

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Ionization Energy

- For an atom to lose an electron, the electron must gain enough energy to overcome its attraction to the positive nucleus.
- The amount of energy required to remove an electron is called *ionization energy*. It varies from element to element.
- The lower the ionization energy (IE), the easier it is to remove an electron from an atom.

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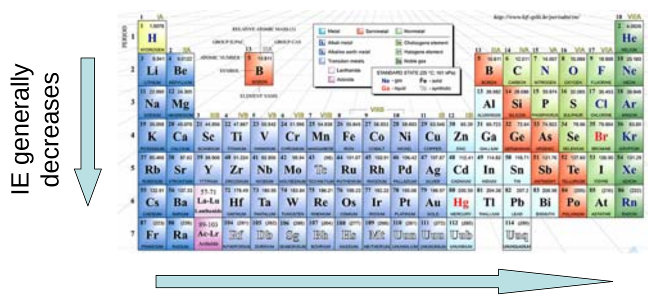
Ionization Energy



IE generally increases

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Ionic Compounds

- Ionic compounds (contain ionic bonds) can be represented by chemical formulas
- Chemical formula:
  - Shows elements in a compound with element symbols
  - Gives ratios of atoms or ions in the compound via subscripts
  - If only one atom of an element is present, no subscript is required
- Examples: NaCl, $MgCl_2$

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Properties of Ionic Compounds

- Tend to form crystal lattices (repeating patterns of ions resulting from attractions between positive and negative ions)
- Properties of ionic compounds can be explained by strong attractions among ions in a crystal lattice
- High melting point (a lot of energy requried to break attractions between ions)
- Solid state is brittle (pushing like ions together causes them to repel in a "rebound effect")
- Solid has poor conductivity (ions can't move), but liquid state has high conductivity (ions can move)

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Covalent Bonds

- ***Covalent Bond***
  - Chemical bond in which two atoms *share* a pair of valence e^-
  - Attractions between shared e^- and the protons in each nucleus hold the atoms together
- ***Molecule***: neutral group of atoms joined by one or more covalent bonds
- Multiple covalent bonds: sometimes 2 atoms can share more than one pair of electrons
  - 2 shared pairs: double bond (4 e^- total are shared)
  - 3 shared pairs: triple bond (6 e^- total are shared)

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Covalent Bonds (cont'd)

- Representation (Example: hydrogen gas, or H_2)
  - Electron dot diagram
    - H:H
  - Structural formula (dash represents a pair of shared e^-)
    - H – H
  - For multiple bonds, multiple lines between two atoms would be used
    - Example: CO_2 is O=C=O (there are two double bonds)

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Polar vs. Nonpolar Bonds (cont'd)

- Unequal sharing of e⁻:
  - Some atoms have a greater attraction for electrons
  - ***Electronegativity***: the ability of an atom to attract electrons to itself *in a covalent bond*
- Polar Covalent Bond
  - Electrons shared UNEQUALLY between 2 atoms
  - The two atoms have DIFFERENT electronegativities
- Nonpolar Covalent Bond
  - Electrons shared EQUALLY between 2 atoms
  - The two atoms have SIMILAR electronegativities

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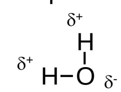
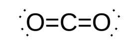
Polar Molecules

- Polar molecules have partial positive and negative charges at opposite ends (like a magnet has north and south poles)
  - Partial negative charge: δ^-
  - Partial positive charge: δ^+
- Polarity of a molecule is determined by
  - Type of covalent bonds (polar or nonpolar)
  - Shape of molecule



Water (H_2O) – Polar Molecule



Carbon Dioxide (CO_2) – Nonpolar molecule

(Both have POLAR covalent bonds, but shape makes one polar and the other nonpolar.)

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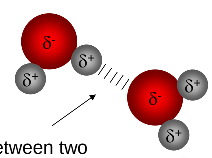
Attractions Between Molecules

- All molecules have forces of attraction between them
- Attractions between polar molecules, like water, are stronger than those between nonpolar molecules
- In water, these attractions are called hydrogen bonds (partial positive hydrogen ends attract partial negative oxygen ends)
- *This explains many of water's strange and important properties:*
  - High surface tension
  - High adhesion and cohesion
  - Lower density in solid state (ice floats!)



Hydrogen bond between two water molecules

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Metallic Bonding

- Metals achieve stable e- configurations by donating their valence electrons to a pool shared by all the metal atoms
- ***Metallic bond***: attraction between metal cations and shared e⁻ around it
- Metal cations form a lattice held in place by metallic bonds
- Electrons are free to move among the atoms – this explains some properties of metals:
  - Conductivity
  - Malleability
  - Ductility

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