Regents Chemistry

Chemical Bonding Part I

Introduction

A chemical bond forms when there is a simultaneous attraction for electrons between two different nuclei to form a compound that is in a lower energy state, resulting in the release of energy. Hence, Bond formation is exothermic. Chemically speaking, which are the most stable elements and what are their electron configurations?

It appears that 8 electrons in the outermost level is a condition of lowest energy – maximum stability.

Our model of bonding describes atoms forming bonds by losing, gaining, or sharing electrons so as to have their valence s and p orbitals complete. Atoms react to attain an isoelectronic configuration with the closest Noble gas – an s^2p^6 configuration or an "octet" of electrons (with exception of helium – which attains an s^2 "duet").

A measure of an atom's attraction for electrons when bonding is termed electronegativity (EN). Electronegativity values are based on a scale developed by the chemist Linus Pauling in which F, the most electronegative element, is assigned a value of 4.0. Recall metals have ________ electronegativities and nonmetals have ________ electronegativities.

The electronegativity difference (Δ) between two atoms can be used to determine the character of the bond between the atoms:

If $\Delta EN > 1.7$, the bond is predominantly ionic
If $\Delta EN < 1.7$, the bond is predominantly covalent (except in metal hydrides which are ionic)

In general, the greater the difference in electronegativity, the stronger the bond.

Types of Chemical Bonds

I. IONIC

-
-
-

Example: potassium chloride

K Cl

electron transfer

after electron transfer

K^+ Cl^-

Lewis structure:

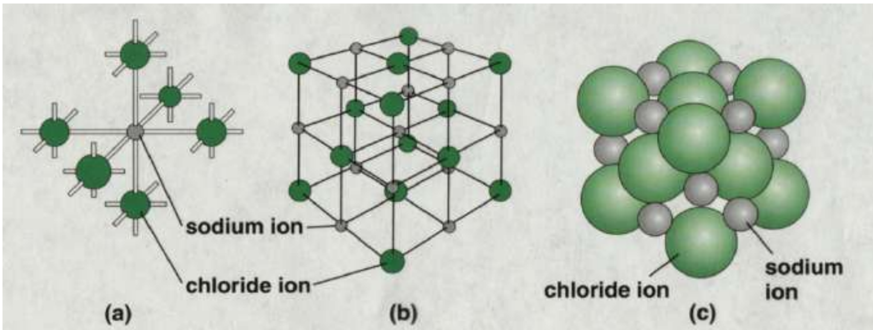
The chlorine atom has a much higher electronegativity (3.5) than potassium (0.8). Our model of bonding views the chlorine as "taking" an electron from potassium rather than "sharing" it. As a result,

oxidation:	$K \rightarrow K^+ + e^-$		reduction:	$Cl + e^- \rightarrow Cl^-$	
protons	19+	19+		17+	17+
electrons	19-	18-		17-	18-
net charge	0	+1 cation		0	-1 anion

In general, ionic bonds are most commonly formed between metals and nonmetals.

Characteristics of Ionic Solids

1. **Crystalline structure** – the ions are held together in fixed positions by strong electrostatic forces in a crystal lattice. All cations (+) are attracted to anions (-) surrounding them; all anions are attracted to cations surrounding them.



(a) A portion of a sodium chloride crystal

(b) A larger portion of the crystal

(c) Another representation of the crystal

2. **Nonconductors** of heat and electricity **in the solid state**. This is explained by immobility of ions and electrons – all free electrons are held in fixed orbitals. On continued heating electrostatic forces are partially disrupted, the crystal lattice is destroyed and ions are relatively free to move. Thus they become **conductors in the molten state**. When dissolved in water, the crystal lattice is destroyed and ions become hydrated and free to move – hence they are also **conductors in aqueous phase.**
3. Soluble in water.
4. High melting and high boiling points.

M^+ X^- M^+ X^-

X^- M^+ X^- M^+

M^+ X^- M^+ X^-

X^- M^+ X^- M^+

II. COVALENT

•

•

•

a) **Nonpolar Covalent** – electrons are shared equally between two nuclei; ΔEN is less than 0.3

<u>single covalent</u> – share one pair of electrons

example: hydrogen, half-filled s orbital of one H overlaps half-filled s orbital of another H

Example: fluorine, half-filled p orbital of one F overlaps half-filled P orbital of another F

<u>double covalent</u> – share two pairs of electrons

example: oxygen, two half-filled orbitals od one O overlap two half-filled p orbital of another O

<u>triple covalent</u> – share three pairs of electrons
example: nitrogen, three half-filled p orbitals of one N overlap three half-filled orbitals of another N

All diatomic molecules involve nonpolar covalent bonding between atoms.

b) **Polar Covalent** – electrons are shared unequally; the more electronegative element "pulls" the shared electrons closer to it resulting in a partial charge on each atom. Molecules having such charges are called **dipoles**. A polar bond between atoms will occur if $\Delta EN > .3$ but <1.7.
example:

hydrogen chloride, half-filled s obribtal of H overlaps half-filled p orbital of Cl

c) Coordinate Covalent – a covalent bond between two atoms in which the shared pair of electrons is contributed by only one of the atoms – typically to a hydrogen ion (i.e. a bare proton which is an empty s orbital). Since coordinate covalent bonding has the effect of adding an ion to a neutral atom, it commonly results in the formation of molecular ions.

example: ammonium ion,

$H = 1s^1$ $\quad$ $H^+ = 1s$

NH_3 $\quad$ $H^+ \rightarrow NH_4^+$
(ammonium ion)

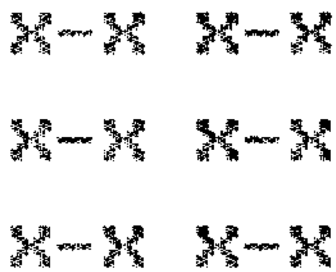
Example: hydronium ion

$H = 1s^1$ $\quad$ $H^+ = 1s$

H_2O $\quad$ $H^+ \rightarrow H_3O^+$
(hydronium ion)

Properties of Molecular Solids

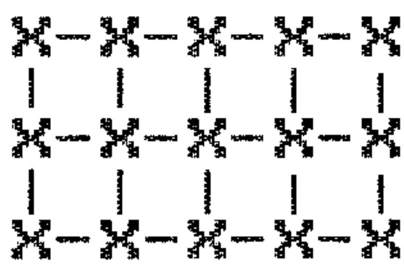
In these solids, individual molecules are "linked" to other molecules by weak intermolecular forces of attraction.



1. Soft
2. Poor conductors of heat and electricity
3. Relatively low melting points

Properties of Network Solids (Macromolecules)

In network solids covalent bonds link atom to atom throughout the solid (i.e. interatomic covalent bonding). There are no discrete covalently bonded molecules (and hence no weak intermolecular forces of attraction linking them).



1. Hard
2. Poor conductors
3. High melting points

examples: Si, SiO_2 (quartz), C (carbon) and Ge and three allotropes of carbon:

graphite – a 2-dimensional network solid; within layers atoms bonded strongly in trigonal planar sheets with weaker bonds linking the sheets

diamond – a 3-dimensional network solid in which each carbon atom shares its four valence electrons with other carbon atoms in a tetrahedral structure

buckminsterfullerene – where the 60 C atoms are bonded to each other in a "soccer ball" type structure

III METALLIC

- All metallic elements (except Hg) are solids at room temperature and pressure and hence exhibit a crystalline structure
- The kernels of the metal atoms are arranged in a fixed crystal pattern
- The valence electrons are free to move between the kernels; these electrons are said top be delocalized and are shared by the entire crystal (i.e. "sea of mobile electrons")