Chapter from Chemistry 3

Chemical bonds. Chemical reactions. Chemical equation

Chemical Bonds. Chemical reactions. Chemical equations.
Balanced Chemical Equations. Stoichiometry
Nomenclature of inorganic compounds.

Ionic bond

The base of ionic bond is the transfer of electrons and the formation of ions, cations - positive charged and anions - negatively charged particles. The typical model is structure of NaCl. Positively charged Na (Na^+) is attracted by negatively charged Cl (Cl^-). Na^+ results from electroneutral atom (parent atom) of sodium by loss of 1 electron, Cl^- by admission of 1 electron. Ionic bonds are the bonds of oppositely charged ions. They are very strong, and crystals of ionic compounds are rigid up to high temperatures, after which the crystals melt. The force of attraction that Na^+ has for a negative charge radiates equally in all directions, like light from a light bulb. Many common substances are ionic compounds (recall from your study at the Secondary school that chemical and physical properties of neutral atoms and ions of these atoms are different, for example Na and Na^+). Every fluid in every living thing, whether plant or animal, contains dissolved ions. At the molecular level of life, the shapes and sizes of ions and molecules are as important as anything else about them. Sodium ions and chloride ions have quite different sizes. In general, when a cation forms from a metal atom, the remaining electron cloud shrinks because the positive nuclear charge now numerically exceeds the negative charge of the electron cloud. The excess positive charge in the cation pulls the electron cloud more tightly inward, so metal ions are always smaller than their parent atoms. The sodium ion has radius 95pm (picometer=
10^{-12}meter) only about half as large as the sodium atom (186pm). When an anion forms from an atom, one or more electrons are added to the atom´s electron cloud. Monoatomic anions are thus always larger than their parent atoms. Cl^- ion has the radius 181 pm, the radius of Cl atom is 99 pm.
The Octet Rule (Noble Gas Rule) : the atoms of the reactive representative elements tend to undergo those chemical reactions that most directly give them electron configuration of the nearest noble gas.

Note :
Ionic bonds also stabilize tertiary structure of proteins.
Electron transfer reactions are examples of redox reactions - (see later)

Covalent bond

The nature has ways besides electron transfers to develop net forces of attraction, chemical bonds, between atoms. It is usually in formation of compounds from nonmetals. Several million compounds, probably over 90% of all compounds, involve only nonmetal atoms. The nonmetals in groups IVA and VA almost never become negative ions, because they would have to gain too many electrons to get outer octets. They form the molecule (molecular compound), small electrically neutral particle consisting of at least two nuclei and enough electrons to make the whole system neutral. The principle of covalent bond is the sharing of two electrons (electron pair), each from another atom. Typical covalent bond is between two same atoms (H_2). The covalent bonds in larger molecules are like in H_2 . They involve the electron density of a pair of electrons that becomes concentrated between two nuclei. The nuclei are attracted into this region and held there at very small distance apart.

When atoms have unequal nuclear charges, the electron cloud of the shared pair is shifted toward the nucleus with the larger charge. This shift means uneven division of charge and the partial charge (δ) results from it. We can see it on the molecule HF. The fluorine nucleus of the H-F molecule pulls some electron density of the shared pair toward itself. Electron cloud is thicker than necessary at F in H-F, so this end has a partial negative charge δ- . We can not say precisely what the sizes of the fractions represented by δ+ and δ- are. However the molecule as a whole is electricaly neutral (sum of δ+ and δ- is zero). This type of bonds we call **polar covalent bond**. Because there are two partial opposite charges in H-F, this molecule is sometimes said to have an **electrical dipole**. Molecules that are electrically polar can stick to each other just like magnets. This is why neutral molecules are able to adhere to each other. The relative ability of an element´s atoms to draw away electron density of a covalent bond is called the **electronegativity** of the element. Fluorine has the highest electronegativity of all of the elements. Oxygen atom is the second one. Metals have the lowest electronegativity.

Whether larger molecules are polar in an overall sense depends not just on the presence of polar bonds but also on the geometry of the molecule. Examples CO_2 and H_2O.

Noncovalent bonds (weak molecular interactions)

The weak molecular interactions are represented by types of bonds with small bond energy (compared with covalent bonds). These types of bonds we divide into the electrostatic interactions, van der Waals forces, hydrogen bonds and hydrophobic interactions.

Electrostatic interaction present attractive forces among oppositively charged parts of molecules.

Van der Waals interaction : interaction dipole-dipole, dipole-induced dipole, London dispersion forces (interaction induced dipole-temporary dipole).

The **hydrogen bond**, a special type of noncovalent bond, is very important for biochemistry. In hydrogen bonds, hydrogen atoms of -OH, -NH or -SH groups (hydrogen bond donors) interact with free electrons of acceptor atoms (for example O, N, or S). The bonding energies of such bonds (10 - 40 $kJ.mol^{-1}$) are much lower than those of covalent bonds (more than 400 $kJ.mol^{-1}$). However, as hydrogen bonds can be very numerous in macromolecules, they play the key role in the stabilization of these molecules.

Hydrophobic interactions is the way of nonpolar molecules or nonpolar parts of molecules to connect in water solutions and hereby to decrease

Illustration : Ozone in the Stratosphere and Ozone in Smog.

Covalent bonds are broken by effect of the convenient light absorption. Ozone (triatomic molecule O_3) in stratosphere (10 to 50 km) is extremely reactive molecule which usually attack most of the molecules in the living system. Ozone in stratosphere prevents to UV light (photons with higher energy than visible light, especially its part UV-B radiation) to damage the cells of skin and which can cause even cancer of skin. Almost all incoming UV-B radiation is absorbed in stratosphere by reaction and :

$$O_2 + \text{UV radiation} \rightarrow 2\,O$$

This reaction then launches a series of reactions called stratospheric ozone cycle. Ozone layer is the part of stratosphere where ozone cycle occurs. Next reaction is the reaction of reactive atomic oxygen with molecule of oxygen or N_2 at the surface of a neutral particle M, which absorbs some energy from the collision O and O_2 . New ozone molecule is then broken to molecule of oxygen and again atomic oxygen.

$$O + O_2 + M \rightarrow O_3 + M + \text{heat}$$

$$O_3 + \text{UV energy} \rightarrow O_2 + O$$

Net effect of this reaction is

UV energy $\rightarrow$ heat

There are substances which can effect this ozone cycle by breaking off it (reduction of ozone layer in stratosphere). Chemical substances can present the great risk in this respect (example chlorofluorocarbons – used as the nonflammable fluids in air conditioners, refrigerators and propellants or paints in cowsheds).
As ozone in stratosphere is useful for life on earth, ozone in smog (lower atmosphere) is very dangerous especially for people. Most of living tissues more exactly double bonds of carbon chains can be attacked by ozone. The chief source of oxygen atoms for formation of ozone, is the breakup of nitrogen dioxide (air pollutant) :

$$NO_2 + \text{solar energy} \rightarrow NO + O$$

The NO_2 can be formed from nitrogen monooxide which is produced inside motor vehicle because of high pressure. NO then reacts with cold O_2 in the air :

$$2\,NO + O_2 \rightarrow 2\,NO_2$$

Nitrogen dioxide has reddish-brown colour and we can find it in environment of heavy traffic in the time of strong sun-light.

Types of Chemicala reactions

1. **Combination reactions** : 2 or more substances react to produce one substance

$S + O_2 \rightarrow SO_2$
$CaO + SO_3 \rightarrow CaSO_4$
$NH_3 + HCl \rightarrow NH_4Cl$
$CaO + H_2O \rightarrow Ca(OH)_2$

2. **Decomposition reactions :** 1 substance forms 2 or more substance :

$PbCO_3 \rightarrow PbO + CO_2$
$2H_2O_2 \rightarrow 2H_2O + O_2$
$2KClO_3 \rightarrow 2 KCl + 3O_2$

3. **Single replacement reactions** : 1 element replaces another one in a compound

a) **Metal-Metal (hydrogen)**
Activities series of metals must be used :

Li, K, Ba, Ca, Na, Mg, Al, Zn, Fe, Cd, Ni, Sn, Pb, H, Cu, Hg, Ag, Pt, Au

Any metal is able to replace only the folowing one !

$Zn + NiCl_2 \rightarrow ZnCl_2 + Ni$
$Pb + 2HCl \rightarrow PbCl_2 + H_2$
$Cu + MgCl_2 \rightarrow 0$
$Ag + Na_2SO_4 \rightarrow 0$
$Zn + 2AgNO_3 \rightarrow Zn(NO_3)_2 + 2Ag$

b) **Nonmetal-Nonmetal : →F, Cl, Br, I**

$Cl_2 + 2NaBr \rightarrow 2NaCl + Br_2$
$Br_2 + 2NaI \rightarrow 2NaBr + I_2$

4. **Double Replacement reactions** : Exchange of the cations between two compounds.
Precipitate or gas slightly ionized compound is formel.

<u>Rules for solubilities of ionic compounds</u> **:**

1. Most of alkali metals salts and NH_4^+ -salts are soluble
2. All nitrates, chlorates, and perchlorates are soluble
3. All chlorides, bromides, and iodides are soluble except for those of Ag^+, Hg^+, Pb^{2+}
4. All sulfates are soluble except for those of Sr^{2+}, Ba^{2+}, Hg^+, Pb^{2+}
5. All sulfides are insoluble except for those of alkali metals, alkaline earth metals and NH_4^+
6. All carbonates, phosphates and chromates are insoluble except for those of alkali metals, alkaline earth metals and NH_4^+

7. All hydroxides are insoluble except for those of alkali metals and Ca^{2+}, Sr^{2+}, Ba^{2+} (moderately soluble).

$$AgNO_3 + NaCl \rightarrow \underline{AgCl} + NaNO_3$$
$$Ni(NO_3)_2 + 2NaOH \rightarrow \underline{Ni(OH)_2} + 2NaNO_2$$
$$\underline{CaCO_3} + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$$
$$HCl + NaOH \rightarrow NaCl + \underline{H_2O}$$
$$NaCl + KOH \rightarrow 0$$