2.1   Definitions and terminology

solution   =

homogeneous $\Rightarrow$

solvent $\leftrightarrow$

solutes $\leftrightarrow$

A solution could be solid, liquid, or gas.

e.g.   stainless steel

vinegar

air

Aqueous solutions have water as solvent.

Concentration and solubility

The composition of a solution is specified by giving the

concentration of its solutes.

molar concentration   =
     (C)

molar solubility   =
     (S)

The concentration of solute in solution is limited by the

solubility of the solute (e.g. NaCl has a solubility of

approximately 6 mol/L in water).

C < S $\Rightarrow$

C = S $\Rightarrow$

C > S $\Rightarrow$

Example:(using concentrations in calculations)

A "stock" solution of HCl(*aq*) is 36.0% HCl by mass. The density of the solution is 1.18 g/mL.

(a)   What is the molarity of HCl in this stock solution?

(b)   What volume of stock solution is required to make 0.750 L of 0.250 mol L^{-1} HCl(*aq*)?

(c)   What volume of 0.250 mol L^{-1} HCl(*aq*) is required to neutralize 0.341 grams of Na_2CO_3. The neutralization reaction is

$$Na_2CO_3(aq) \; + \; 2\,HCl(aq) \; \rightarrow \; CO_2(g) \; + \; H_2O(l) \; + \; 2\,NaCl(aq)$$

2.2   Aqueous solutions

Water is an inexpensive and versatile solvent. Most ionic
and molecular compounds dissolve in water, to some extent
at least.

ionic compound   ↔

molecular compound   ↔

 When an ionic compound is dissolves in water, ions break
away from the solid and become hydrated:

A similar process occurs when a molecular compound
dissolves in water, except that neutral molecules, not ions,
become hydrated.

Some molecular compounds (e.g. acids and bases) react
with water to produce ions in solution.

So, we see that ionic compounds and some molecular
compounds produce ions when dissolved in water. Such
compounds are called electrolytes. Electrolytes are
classified as strong or weak electrolytes, depending on the
extent to which ions are produced.

2.3   Reactions in Solution

A.   Precipitation reactions

In order to understand precipitation reactions, you must learn

the solubility rules. These rules allow us to predict wne (or

if) a precipitation reaction will occur.

Solubility Rules for Ionic Compounds

1. <u>Salts of the alkali metals</u> are <u>soluble</u>. **(Note: The alkali metals are in group 1.)**
   e.g. If M = Li, Na or K, then MX, M_2X, M_3X, etc. are soluble regardless of what X is.

2. <u>Ammonium (NH_4^+)</u> <u>salts</u> are <u>soluble</u>.

   e.g. NH_4X, (NH_4)$_2$X, (NH_4)$_3$X, etc. are soluble regardless of what X is.

3. <u>Nitrates</u> NO_3^- are <u>soluble</u>.

   e.g. MNO_3, $M(NO_3)_2$, $M(NO_3)_3$, etc. are soluble regardless of what M is.

4. <u>Halides</u> **i.e. chlorides (Cl⁻), bromides (Br⁻) and iodides (I⁻) are** <u>soluble</u> **except for the halides of lead (Pb^{2+}), mercury (Hg$^+$ and Hg_2^{2+}) and silver (Ag$^+$).**
   e.g.   If X = Cl, Br or I, then MX, MX_2, MX_3, etc. are soluble unless M = Pb, Hg or Ag.

5. <u>Sulfates</u> **(SO_4^{2-}) are** <u>soluble</u> **except for the sulfates of calcium, strontium, barium, silver mercury and lead.**

   e.g. M_2SO_4, MSO_4, $M_2(SO_4)_3$, etc. are soluble unless M is from group 2 (the alkaline earths) or M = Pb, Hg or Ag.

6. <u>Carbonates</u> **(CO_3^{2-}),** <u>phosphates</u> **(PO_4^{3-}) and** <u>sulfides</u> **(S^{2-}) are** <u>insoluble</u> **except for**
   (i) the carbonates/phosphates/sulfides of the alkalis (because of Rule 1), and
   (ii) ammonium carbonate/phosphate/sulfide (because of Rule 2).

7. <u>Hydroxides</u> **(OH⁻) are** <u>insoluble</u> **or** <u>slightly soluble</u> **except for the hydroxides of the alkalis (because of Rule 1).**
   Note: The hydroxides of group 2 (the alkaline earth metals) are slightly soluble. Virtually all other hydroxides are insoluble

Example: **(solubility rules; molecular and net ionic equations)**

Aqueous solutions of $Bi(NO_3)_3$ and Na_2S are mixed. Does a reaction occur? If so, what are the molecular and net ionic equations for the reaction?

Example: In each of the following cases, write the net ionic equation for the reaction if a reaction occurs.

$Na_3PO_4(aq)$ + $AlCl_3(aq)$ $\rightarrow$

$Al_2(SO_4)_3(aq)$ + $BaCl_2(aq)$ $\rightarrow$

$(NH_4)_2CO_3(aq)$ + $Pb(NO_3)_2(aq)$ $\rightarrow$

Gravimetric analysis

Precipitation reactions are important because they are used to carry out gravimetric analysis. In gravimetric analysis, the amount of substance in a sample is deduced from the mass of precipitate that is produced.

Example: A 1.2500-g sample of impure $MgSO_4$ was dissolved and then treated with excess barium chloride solution, $BaCl_2(aq)$. The dried barium sulfate precipitate weighed 0.9165 g. What is the percentage by mass of $MgSO_4$ in the original mixture?
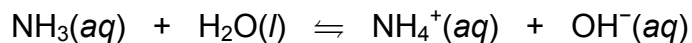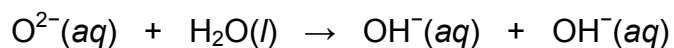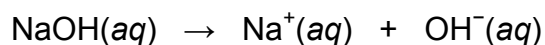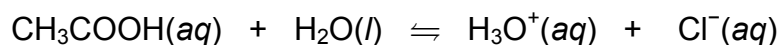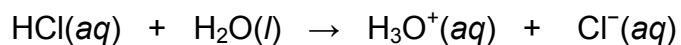
B.   Acid-base reactions

Chemists use a variety of definitions for the terms acid and base. We shall focus on the Bronsted-Lowry definitions:

acid   =

base =

Example: Identify the acids and bases in the following reactions.

$HCl(aq)$ + $H_2O(l)$ $\rightarrow$ $H_3O^+(aq)$ + $Cl^-(aq)$

$CH_3COOH(aq)$ + $H_2O(l)$ $\rightleftharpoons$ $H_3O^+(aq)$ + $Cl^-(aq)$

$NaOH(aq)$ $\rightarrow$ $Na^+(aq)$ + $OH^-(aq)$

$O^{2-}(aq)$ + $H_2O(l)$ $\rightarrow$ $OH^-(aq)$ + $OH^-(aq)$

$NH_3(aq)$ + $H_2O(l)$ $\rightleftharpoons$ $NH_4^+(aq)$ + $OH^-(aq)$

Memorize the names and formulas of the common strong acids and strong bases.

Common strong acids

Common strong bases

Neutralization reactions

In a neutralization reaction, an acid and base react to form a salt
and water. If either the acid or the base is "strong", then the
neutralization reaction goes essentially to completion:

If we know the neutralization reaction goes to completion,
then we can use that reaction in the laboratory to carry out
an acid-base titration. (A titration is an important method of
analysis.)

In a titration, titrant is added from a burette, in controlled amounts,
to a flask or beaker containing another solution. . The
concentration of one of the solutions must be known.

Titrant is added until the endpoint is observed. The purpose of the
titration is to determine the concentration of the other solution.

Example: **(acid-base titration)**

In an experiment, a 10.00 mL portion of phosphoric acid solution, $H_3PO_4(aq)$, was diluted to 40.00 mL. If it took 25.54 mL of 2.111 mol L^{-1} KOH(*aq*) to neutralize all of the H_3PO_4, what is the molarity of the original H_3PO_4 solution?

C. Oxidation-reduction reactions

The electrons are "hidden" in the balanced chemical
equation for a reaction. Therefore, it is not always easy to
recognize an electron transfer reaction.

Chemists have devised a way of keeping track of electrons:
they use an accepted set of rules to assign oxidation states
to the all of the atoms involved in the reaction.

Rules for assigning oxidation states

1. The oxidation state is zero for an atom in elemental form.

2. The sum of oxidation states must equal the total charge.

3. In their compounds, the group 1 metals always have an oxidation state of +1 and the group 2 metals always have an oxidation state of +2.

4. In its compounds, F always has an oxidation state of −1.

5. In its compounds, H normally has an oxidation state of +1, except when H combines with a group 1 or group 2 metal.

6. In its compounds, O normally has an oxidation state of −2, except in peroxides which have O bonded to O.

Formally, the oxidation state of an atom in a compound is equal to the charge the atom would have if the bonding electrons between each pair of atoms were given to the atom having the greatest electronegativity. (Electronegativity is a concept we will encounter later on this term.) Sometimes, we need to resort to this "formal" definition when assigning oxidation states because the rules given above don't cover all possible cases. (e.g. CN^-)

Example: Assign oxidation states to the atoms in each of the following cases.

(a) C_2H_5OH (b) $Cr_2O_7^{2-}$ (c) MnO_4^-

(d) I_3^- (e) $NaNO_3$

In an oxidation-reduction reaction, the oxidation states of one or more atoms (or ions) changes. For example, consider the reaction between $Zn(s)$ and $CuCl_2(aq)$ to give $ZnCl_2(aq)$ and $Cu(s)$:

In the reaction above, the oxidation state of Zn increases and the oxidation state of Cu decreases. We say that Zn has been oxidized (from Zn^0 to Zn^{2+}) and Cu has been reduced (from Cu^{2+} to Cu^0). In general, oxidation is associated with the loss of electrons and reduction is associated with the gain of electrons.

Every electron transfer reaction can be written as the sum of an oxidation process and a reduction process, each of which is referred to as a "half-reaction". For the oxidation of Zn by Cu^{2+}, the half-reactions are:

Not all oxidation-reduction reactions are as simple as the one just considered. This is especially true for oxidation-reduction reactions that occur in acidic or basic solutions. Such reactions are difficult to balance by inspection and we must use a systematic approach to balance them. The following approach is recommended. (Our method is different from the method described in the text. We want you to learn and use our method because our method emphasizes oxidation states. The method in the text ignores oxidation states!)

Completing and balancing redox reactions that occur in aqueous solution

1. Assign oxidation states to each element in the reaction and identify the species being oxidized and reduced.

2. Write separate half-reactions for the oxidation and reduction processes.

3. Balance the separate half-reactions:
    a) first with respect to the element being oxidized or reduced, and
    b) then by adding electrons to one side or the other to account for the number of electrons produced (oxidation) or consumed (reduction).

4. Combine the half-reactions algebraically so that the total number of electrons cancels out.

5. Balance the net charge by either adding OH^- to one side of the equation (for basic solutions) or H^+ to the other (for acidic solutions).

6. Balance the O and H atoms by adding H_2O. Check that the final equation is balanced with respect to each type of atom and with respect to charge.

Note: The balanced equation for any reaction, including a redox reaction, comes from nature's demand to conserve both charge and mass throughout all chemical changes. For electron transfer reactions, mass and charge balance considerations are intertwined.

Example: Complete and balance the equation below,
given that the reaction occurs in acidic, aqueous solution.

$$C_2H_5OH(aq) \ + \ Cr_2O_7^{2-}(aq) \ \longrightarrow \ C_2H_4O_2(aq) \ + \ Cr^{3+}(aq)$$
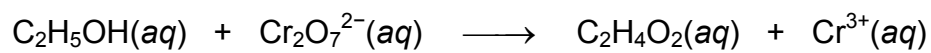
Example: Complete and balance the equation below, given that the reaction occurs in aqueous, basic solution.

$$Cl_2(aq) \longrightarrow ClO_3^-(aq) + Cl^-(aq)$$