Chemical Reactions in Water

Ron Robertson

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Properties of Compounds in Water

Electrolytes and nonelectrolytes

- Water soluble compounds form what we call aqueous solutions.
- These solutions can be electrolytes (electricity conducting) or nonelectrolytes (does not conduct electricity).
- Electrolytes conduct electricity because of the presence of ions. Electrolytes can be strong or weak depending on how many ions are produced. The classes of electrolytes are acids, bases and salts. Electrolytes can be ionic or covalent compounds dissolved in water.

Acids, Bases and Salts

- Acids dissolve in water to give H^+ ions. These ions attach to water molecules and form the H_3O^+ species called the ***hydronium*** ion. Many acid formulas look like H^+'s attached to anions. Nonmetal oxides dissolve in water to form acid solutions and form the basis of the acid rain problem.
- Bases dissolve in water to give OH^- ions. Many base formulas look like metal ions bonded to hydroxide ions. Some bases like ammonia do not follow this rule. Many metal oxides form basic solutions.
- Salts can be thought of as combinations of acids and bases. Salt is a general chemical term and refers to many compounds, not just table salt.
- HCl, HNO_3, H_2SO_4, and $HClO_4$ are *strong* acids. Most of the hydroxides of column 1 and 2 are strong bases. Most other acids and bases are *weak* electrolytes which do not break apart completely and are in *reversible* reactions.

Aqueous chemistry

- Water *soluble* ionic compounds are strong electrolytes. (See a table of solubility rules and know how to use.)
- These compounds actually do not exist in water as formula units (like NaCl). They exist as ions (like Na^+ and Cl^-). The notation NaCl (aq) means that NaCl is dissolved in water and exists as Na^+ and Cl^- surrounded by water molecules. This process is called *hydration*. The notation for a solid material that does not dissolve in water is (s) and the notation for a gas is (g).
- Nonelectrolytes do not break apart into ions in solution, they exist as whole molecules or formula units.

One way to classify reactions

Decomposition (complex Þ simple)

- Binary compound Þ elements
- Carbonates Þ carbon dioxide + oxide
- Chlorates Þ chloride + oxygen gas

Combination or Synthesis (simple Þ complex)

- 2 elements Þ binary compound
- metal + oxygen Þ metal oxide
- nonmetal + oxygen Þ nometal oxide
- metal or nonmetal + halogen Þ halides

Single displacement (exchange reaction) [many are aqueous]

(element + compound Þ element + compound)

Double displacement (exchange reaction) [many are aqueous]

(compound + compound Þ compound + compound)

- acid + base
- precipitation
- gas forming

Combustion

- Hydrocarbon (H, C) + oxygen Þ carbon dioxide + water
- H, C, O compound+ oxygen Þ carbon dioxide + water

Many reactions do not fit into these categories perfectly.

Writing Reactions

There are three ways to write an equation for a reaction.

- Molecular Equation – overall reaction stoichiometry but not necessarily the actual forms
- Complete Ionic Equation – shows what the reactants and products actually look like in solution; all strong electrolytes are written as ions
- Net Ionic Equation – the short form of the complete ionic equation; ions that are on both reactant and product side do not appear (spectator ions)

Another way to classify!

Intro

- Another way to classify reactions is to determine whether electrons have been transferred in the reaction. Those reactions in which electrons are transferred are called oxidation/reduction reactions (also called redox).
- *Oxidation* - loss of electrons, oxidation number increases
- *Reduction* - gain of electrons, oxidation number decreases
- Remember that LEO goes GER
(Loss of Electrons Oxidation, Gain of Electrons Reduction)

- The general idea is that the number of electrons lost by oxidation are gained by reduction. It is impossible to have oxidation without reduction.
- As a general rule: *combination* (there are exceptions), *decomposition* (there are exceptions), and *single displacement* reactions are redox; *double displacement* reactions are not. Simply assign oxidation numbers to determine if the reaction is redox. If the oxidation number of any element changes from reactant to product side, the reaction is redox.
- As listed above, most double displacement reactions can be classified as acid/base or precipitation. Acids plus bases yield salts and usually water. Precipitation reactions yield an insoluble salt.
- A special type of redox reaction is a *disproportionation* reaction in which one element is both oxidized and reduced.

The agents of oxidation and reduction

- The substance which is oxidized is the agent by which reduction can take place; it loses the electrons necessary for the reduction. Thus the substance that is oxidized is called the reducing agent.

- The substance which is reduced is the agent by which oxidation can take place; it gains the electrons that are lost by the oxidation. Thus the substance that is reduced is called the oxidizing agent.

Common oxidizing and reducing agents

- Oxidizing agents like to gain electrons and be reduced. Oxygen, the halogens, nitric acid, dichromate and permanganate are all common oxidizing agents.

- Reducing agents like to lose electrons and be oxidized. Hydrogen, metals such as sodium, potassium, iron or aluminum, and carbon are common reducing agents.

Balancing Redox reactions (optional material)

Concept --must balance both mass and charge

Steps to balancing

1) Assign oxidation numbers to recognize as redox

2) Break into half cells, oxidation and reduction

3) Mass balance elements that are being oxidized and reduced
4) Electron balance each half cell
5) Balance electrons lost in the oxidation half cell with electrons gained in the reduction half cell
6) Balance charge in each half cell if needed

   Neutral conditions - nothing needed

   Acid - H^+, Base - OH^-
7) Final mass balance with water (if needed)
8) Add half cells together

   Examples: Neutral conditions

   $Cr^{+2} + I_2$ Þ $Cr^{+3} + I^-$

Acid conditions

$Cu + (NO_3)^{-1}$ Þ $Cu^{+2} + NO_2$

Basic conditions

$Cr + (ClO_4)^{-1}$ Þ $Cr(OH)_3 + (ClO_3)^{-1}$

Preparing Solutions

In a solution a solute (lesser) is dissolved in a solvent (greater).

Concentration units

1. Percentage (weight/weight, weight/volume, volume/volume)

2. **Molarity (M)** $= \dfrac{\textbf{moles of solute}}{\textbf{liters of solution}}$

Example What is the Molarity of NaOH for a solution made by dissolving 20.0 g of NaOH in a total solution volume of 100. ml?

What is the Molarity of the sodium ion?

What is the total ion concentration?

How many liters of a 0.100 M solution of NaOH can be prepared from 20.0 g of NaOH?

How would you prepare 2.00 l of 0.500 M NaOH?

Dilutions

Major idea - the moles of solute in a diluted solution are the same as in the concentrated solution.

$M_cV_c = M_dV_d$

where M represents Molarity and V represents Volume

Note: The same idea could apply to concentrated and diluted percent solutions instead of Molarity. In general the amount of solute in the concentrated solution equals the amount of solute in the diluted solution.

Example

A 0.100 M solution of NaOH is to be prepared from a concentrated stock solution of 5.00 M. Describe the preparation of 5.00 l of this solution.

Titrations

A *volumetric* method for determining quantitative information about a solution by reacting it with another solution. Acid-base and redox titrations are common types.

Any titration problem can be solved as a stoichiometry problem

Example

How many ml of 0.100 M sulfuric acid will it take to neutralize 500. ml of 0.500 M NaOH?

- Write the balanced equation
- Identify what you have and what you want
- Convert what you have to moles
- Use the molar ratio in the balanced equation as a conversion factor
- Convert back to whatever units you need