

## Chapter 4 - Types of Chemical Reactions and Solution Chemistry

### 4.1 Water, the Common Solvent

- the water molecule is bent with and H-O-H angles of approx. 105 °
- O-H bonds are covalent
- O is slightly negative; H is slightly positive
- polar molecules
- the solvation of salts by H<sub>2</sub>O is called hydration (break up into constituent ions)
- ionic bond is replaced by dipole-ion interaction
- hydrated ions are written with a subscript (aq) in a reaction eq.
- nonionic substances can also dissolve in H<sub>2</sub>O, ex. ethanol
- “like dissolves like”

### 4.2 The Nature of Aqueous Solutions: Strong and Weak Electrolytes

- solute – a substance dissolved in a liquid to form a solution
- solvent – the dissolving medium in a solution
- 1) weak electrolytes - conduct a small amount of current
  - 1a. - weak acids & bases (most common)
  - 1b. - only about 1% ionizes
  - 1c. - weak acids and bases only slightly ionize
- 2) strong electrolytes - conduct current very efficiently
  - 2a. - completely ionize (soluble salts, strong acids, strong bases)
  - 2b. - Arrhenius proposed that acid produce H<sup>+</sup> ions
  - 2c. - (aq) subscript is used
  - 2d. - strong acids completely ionize (H<sub>2</sub>SO<sub>4</sub>, HNO<sub>3</sub>, HCl, HClO<sub>4</sub>)
  - 2e. - strong bases completely ionize (NaOH, KOH)
- 3) nonelectrolytes - don't allow current to flow
  - 3a. - do not produce ions

### 4.3 The Composition of Solutions

$$\text{molarity (M)} = \frac{\text{moles of solute}}{\text{liters of solution}}$$

\*Sample exercise 4.2, 4.3, & 4.4

*Exercise 4.2:*

Calculate the molarity of a solution prepared by dissolving 1.56 g of gaseous HCl in enough water to make 26.8 mL of solution.

Since molarity is a combination of moles and liters, we have to get those two units. We can get moles by multiplying grams by the molar mass. Liters can be found with the mL and

dimensional analysis. Finally, divide these two terms to get molarity.

$$1.56 \text{ g HCl} \times \frac{1 \text{ mol}}{36.46 \text{ g}} = 4.28 \times 10^{-2} \text{ mol}$$

$$26.8 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 2.68 \times 10^{-2} \text{ L}$$

$$\text{Molarity} = \frac{4.28 \times 10^{-2} \text{ mol}}{2.68 \times 10^{-2} \text{ L}} = \mathbf{1.60 \text{ M HCl}}$$

standard solution - solution with known concentration

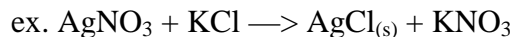
#### 4.4 Types of Chemical Reactions

- 1) Precipitations reactions
- 2) Acid-base reactions
- 3) Oxidation-reduction reactions

#### 4.5 Precipitation Reactions (Double Replacement/Metasthesis)

precipitation reactions - form insoluble solids

\*Note - if no state is specified, any ions can be assumed to be aqueous



Solubility rules: The first ones always override the latter ones!

- 1) Always soluble - Alkali metals,  $\text{C}_2\text{H}_3\text{O}_2^-$ ,  $\text{NO}_3^-$ ,  $\text{NH}_4^+$ ,  $\text{ClO}_3^-$ ,  $\text{ClO}_4^-$
- 2) Usually soluble -  $\text{Cl}^-$ ,  $\text{Br}^-$ ,  $\text{I}^-$  (except  $\text{Ag}^+$ ,  $\text{Pb}^{2+}$ ,  $\text{Hg}_2^{2+}$ , or AP/H)  
 $\text{SO}_4^{2-}$  (except  $\text{Ca}^{2+}$ ,  $\text{Ba}^{2+}$ ,  $\text{Sr}^{2+}$ ,  $\text{Pb}^{2+}$ ,  $\text{Hg}_2^{2+}$ , or CBS/PBS)
- 3) Usually not soluble -  $\text{O}^{2-}$ ,  $\text{OH}^-$  (except  $\text{Ca}^{2+}$ ,  $\text{Ba}^{2+}$ ,  $\text{Sr}^{2+}$ , or CBS)
- 4) Never soluble -  $\text{S}^{2-}$ ,  $\text{CO}_3^{2-}$ ,  $\text{CrO}_4^{2-}$ ,  $\text{PO}_4^{3-}$

*Exercise 4.8a:*

Using the solubility rules in Table 4.1, predict what will happen when the following pairs of solutions are mixed

- a)  $\text{KNO}_3$  and  $\text{BaCl}_2$  - When these two ionic compounds are mixed in double replacement,  $\text{KCl}$  and  $\text{Ba}(\text{NO}_3)_2$ . Both of these are soluble, according to the solubility rules, so both remain aqueous.

## 4.6 Describing Reactions in Solution

molecular equation - reactants and products written in molecular form (balanced)

complete ionic equation - all substances are represented as ions (strong electrolytes)

spectator ions - do not participate in reaction

net ionic equation - includes those species directly involved in the reaction

### *Exercise 4.9b:*

For each of the following reactions, write the molecular equation, the complete ionic equation, and the net ionic equation.

b) Aqueous potassium hydroxide is mixed with aqueous iron(III) nitrate to form precipitate of iron(III) hydroxide and aqueous potassium nitrate.

Molecular equation –  $3\text{KOH} + \text{Fe}(\text{NO}_3)_3 \rightarrow \text{Fe}(\text{OH})_{3(s)} + 3\text{KNO}_3$

Complete ionic equation –  $3\text{K}^+ + 3\text{OH}^- + \text{Fe}^{3+} + 3\text{NO}_3^- \rightarrow \text{Fe}(\text{OH})_{3(s)} + 3\text{K}^+ + 3\text{NO}_3^-$

Net ionic equation -  $3\text{OH}^- + \text{Fe}^{3+} \rightarrow \text{Fe}(\text{OH})_{3(s)}$

## 4.7 Stoichiometry of Precipitation Reactions

### *Exercise 4.11:*

When aqueous solutions of  $\text{Na}_2\text{SO}_4$  and  $\text{Pb}(\text{NO}_3)_2$  are mixed,  $\text{PbSO}_4$  precipitates. Calculate the mass of  $\text{PbSO}_4$  formed when 1.25 L of 0.0500 M  $\text{Pb}(\text{NO}_3)_2$  and 2.00 L of 0.0250 M  $\text{Na}_2\text{SO}_4$  are mixed.

$$1.25 \text{ L} \times \frac{0.0500 \text{ mol Pb}(\text{NO}_3)_2}{1 \text{ L}} \times \frac{1 \text{ mol SO}_4^{2-}}{1 \text{ mol Pb}^{2+}} = 0.0625 \text{ mol SO}_4^{2-}$$

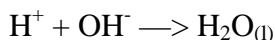
$$2.00 \text{ L} \times \frac{0.0250 \text{ mol Na}_2\text{SO}_4}{1 \text{ L}} = 0.0500 \text{ mol SO}_4^{2-} \text{ (limiting reagent)}$$

$$0.0500 \text{ mol SO}_4^{2-} \times \frac{1 \text{ mol PbSO}_4}{1 \text{ mol SO}_4^{2-}} \times \frac{303.27 \text{ g}}{1 \text{ mol PbSO}_4} = \mathbf{15.2 \text{ g PbSO}_4}$$

## 4.8 Acid-Base Reactions

Bronsted- Lowry Acid - proton donor

Bronsted- Lowry Base - proton acceptor



Acid-Base Titrations -

titration - technique used to determine the amount of a particular substance by the use of a standard solution

titrant - known concentration

analytic - unknown concentration

equivalence point - point where the titrant has completely reacted with the analytic

indicator - substance that changes color near or at the equivalence point

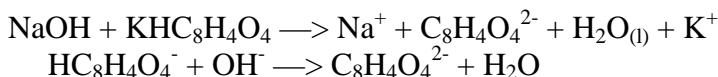
endpoint - point where the indicator color changes

acid-base titrations - an acid or a base is reacted with a strong acid or base

\*sample exercise 4.14 & 4.15

#### *Exercise 4.14:*

A student carries out an experiment to standardize (determine the exact concentration of) a sodium hydroxide solution. To do this, the student weighs out a 1.3009-g sample of potassium hydrogen phthalate ( $\text{KHC}_8\text{H}_4\text{O}_4$ , often abbreviated KHP). KHP (molar mass 304.33 g/mol) has one acidic hydrogen. The student dissolves the KHP in distilled water, adds phenolphthalein as an indicator, and titrates the resulting solution with the sodium hydroxide solution to the phenolphthalein endpoint. The difference between the final and initial buret reading indicates that 41.20 mL of the sodium hydroxide solution is required to react exactly with the 1.3009 g KHP. Calculate the concentration of the sodium hydroxide solution.



$$1.3009 \text{ KHP} \times \frac{1 \text{ mol KHP}}{204.22 \text{ g KHP}} = 6.3701 \times 10^{-3} \text{ mol KHP}$$

$$\frac{6.3701 \times 10^{-3} \text{ mol OH}^-}{41.20 \text{ L}} = \mathbf{0.1546 \text{ M NaOH}}$$

### 4.9 Oxidation-Reduction Reactions

Oxidation-reduction reactions or redox reactions - reactions where one or more electrons are transferred

Oxidation states (oxidation #'s) - provide a way to keep track of electrons in redox reactions

#### *Rules for Assigning Oxidation States*

- 1) elements - oxidation state = 0
- 2) ions - oxidation state = charge
- 3) H, O, & F within covalent compounds - oxidation state equals the charge of an atom if it were in an ionic compound (H=+1, O=-2 [unless peroxide], F= -1)
- 4) The sum of the oxidation states must be equal to zero in an electrically neutral compound

- 5) Oxidation states are written as +n or -n.  
6) It is possible to have non-integer oxidation states ( $\text{Fe}_3\text{O}_4$ ,  $\text{Fe}=\frac{8}{3}$ )

*Example 4.16:*

Assign oxidation states to all atoms in the following

a)  $\text{CO}_2$

**C=+4, O=-2**

b)  $\text{SF}_6$

**S=-6, F=+1**

c)  $\text{NO}_3^-$

**N= +5, O=-2 (5 - [2x3] = -1)**

oxidation - increase in oxidation state (loss of  $e^-$ )

reduction - decrease in oxidation state (gain of  $e^-$ )

oxidizing agent (electron acceptor) - does the oxidizing

reducing agent (electron donor) - does the reducing

\* “LEO (lose electrons, oxidation) the lion says GER (gain electrons, reduction)!!!”

When an atom is oxidized, it's the reducing agent!

When an atom is reduced, it's the oxidizing agent!

#### 4.10 Balancing Oxidation-Reduction Equations

- redox reactions in aqueous solutions can be balanced by using half-reactions (a.k.a. half-reaction method)

##### *Steps for balancing redox reactions in acidic solutions*

- 1) write both the reduction and oxidation half-reactions separately
- 2a) balance elements (except H & O)
  - b) balance oxygen using  $\text{H}_2\text{O}$
  - c) balance hydrogen using  $\text{H}^+$
  - d) balance the charges using  $e^-$ 's
- 3) make number of electrons equal in both half-reactions
- 4) add half-reactions together and cancel out identical species (electrons should cancel out, due to step 3)
- 5) check that elements and charges cancel

##### *Steps for balancing redox reactions in basic solution*

1-5) same as for acidic solutions

6) add  $\text{OH}^-$  ions to both sides of the equation that equal the number of  $\text{H}^+$  ions

7) sides containing both  $\text{H}^+$  and  $\text{OH}^-$  ions will form water (eliminate  $\text{H}_2\text{O}$  molecules)

8) check that elements and charges are balanced

### Balancing Redox Reactions

p. 183 #59, #61, & #63

For solutions, see homework