Chapter 4: Phenomena

Phenomena: Many different reactions are known to occur. Scientists wondered if these reactions could be separated into groups based on their properties. Look at the reactions below and divide the reactions into groups of similar reactions. Be able to state what property(s) you used to group them. Can any of the reactions be put into multiple groups? Hint: It might be helpful to print the reactions out so that you can move them around.

a) $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$
b) $NH_3(aq) + HCN(aq) \rightarrow NH_4CN(aq)$
c) $Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$
d) $HC_2H_3O_2(aq) + KOH(aq) \rightarrow KC_2H_3O_2(aq) + H_2O(l)$
e) $3Ca^{2+}(aq) + 2PO_4^{3-}(aq) \rightarrow Ca_3(PO_4)_2(s)$
f) $HBr(aq) + LiOH(aq) \rightarrow LiBr(aq) + H_2O(l)$
g) $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$
h) $2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$
i) $Sr(NO_3)_2(aq) + 2NaOH(aq) \rightarrow Sr(OH)_2(s) + 2NaNO_3(aq)$
j) $Cu^+(aq) + e^- \rightarrow Cu(s)$
k) $H_2SO_3(aq) + H_2O(l) \rightarrow SO_4^{2-}(aq)+ 4H^+(aq)+ 2e^-$
Big Idea: Reactions can be broken down into subgroups. Three types of reactions are precipitation (a solid is formed from 2 aqueous solutions), acid/base (salt and H₂O are produced), and oxidation/reduction (e⁻ are transferred). Many reactions involve species that are in solution. Molarity (mol per liter) is used to describe the concentration of species in a solution.
Electrolytes
Electrolytes

Chapter 4: Types of Chemical Reactions and Solution Stoichiometry
Electrolytes

- **Electrolyte**: A substance that, in solution, is present as ions.
  
  *Examples*: Ionic solids that are soluble in water and acids.

  *Note*: Electrolyte solutions conduct electricity.

- **Strong Electrolyte**: A substance that is fully ionized in solution.

- **Weak Electrolyte**: A substance that is only partially ionized in solution.

- **Nonelectrolyte**: A substance that does not form ions in solution.
  
  *Example*: Molecular compounds that are not acids.
Molarity and Dilutions

Molar Mass

Grams

Moles

Chemical Formula or Equation

Moles

Molarity

N_A

Atoms/Molecules/etc

Volume
Student Question

If the molarity of a solution of calcium chloride is known. What would you have to do to the molarity of the calcium chloride solution to get the molarity of the chloride ions in solution?

a) It is the same
b) Multiply the molarity by 2
c) Divide the molarity by 2
d) Multiply the molarity by 3
e) Not enough information given
The steps to make a solution of known molarity.

1. A known mass of the solute is dispensed into a volumetric flask.
2. Some water is added to dissolve it.
3. Water is added up to the mark on the stem of the flask.
Calculating Amount of Stock Solution Needed

- **Step 1:** Calculate the amount of solute, \( n \), needed in the final solution, \( V_2 \).
  \[
  n = (M_2)(V_2)
  \]

- **Step 2:** Calculate the volume, \( V_1 \), of the initial stock solution of molarity \( M_1 \) that contains \( n \) moles.
  \[
  V_1 = \frac{n}{M_1}
  \]

**Note:** Since the amount of moles is the same in step 1 and 2, the equations can be combined into one equation.

\[
V_1 = \frac{M_1 V_2}{M_1} \quad \text{or} \quad M_1 V_1 = M_2 V_2
\]
Student Question

A solution is prepared by dissolving 0.005736 mol of oxalic acid (C₂H₂O₄) to make 0.1000 L of solution. A 0.01000 L portion is then diluted to 0.2500 L. What is the molarity of the final solution?

a) 2.295×10⁻³ M
b) 2.295 M
c) 5.738×10⁻² M
d) 5.737 M
e) None of the above
Precipitation Reactions

- **Precipitate**: The formation of a solid in a solution during a chemical reaction.

- **Soluble Substance**: A substance that dissolves to a significant extent in a specified solvent.

  **Note**: If no solvent is mentioned, the solvent is assumed to be water.

- **Insoluble Substance**: A substance that does not dissolve significantly in a specified substance.

  **Note**: A substance is considered insoluble if they do not dissolve to more than ~0.1 M.
Precipitation Reactions

1. Most nitrate, $\text{NO}_3^-$, salts are soluble.
2. Most salts of $\text{Na}^+$, $\text{K}^+$, and $\text{NH}_4^+$ are soluble.
3. Most chlorides are soluble. Notable exceptions are $\text{AgCl}$, $\text{PbCl}_2$, and $\text{Hg}_2\text{Cl}_2$.
4. Most sulfate salts are soluble. Notable exceptions are $\text{BaSO}_4$, $\text{SrSO}_4$, $\text{PbSO}_4$, and $\text{CaSO}_4$.
5. Most hydroxide salts are only slightly soluble. The important soluble hydroxides are $\text{NaOH}$, $\text{KOH}$, and $\text{Ca(OH)}_2$ (marginally soluble).
6. Most sulfide ($S^{2-}$), carbonate ($\text{CO}_3^{2-}$), and phosphate ($\text{PO}_4^{3-}$) salts are only slightly soluble.
Precipitation Reactions

Solubility Song
(Sung to the tune of 99 Bottles of Beer on the Wall)

Potassium, sodium, and ammonium salts, Whatever they may be, Can always be depended on for solubility.

When asked about the nitrates The answer is always clear, They each and all are soluble, is all we want to hear.

Most every chloride's soluble At least we've always read Except silver, mercurous mercury And (slightly) chloride of lead.

Rule 1: Most nitrate(NO$_3^-$) salts are soluble

Rule 2: Most salts of Na$^+$, K$^+$, and NH$_4^+$ are soluble

Rule 3: Most chlorides are soluble. Notable exceptions are AgCl, PbCl$_2$, and Hg$_2$Cl$_2$
Precipitation Reactions

Every single sulfate is soluble, tis said, ‘Cept barium and strontium and calcium and lead.

Hydroxides of metals won't dissolve That is, all but three. Potassium, sodium, and ammonium dissolve quite readily.

And then you must remember That you must not "forgit" Calcium, barium, strontium dissolve a little bit.

The carbonates are insoluble, It's lucky that it's so, Or else, our marble buildings would melt away like snow.

**Rule 4:** Most sulfate salts are soluble. Notable exceptions are BaSO₄, SrSO₄, PbSO₄, and CaSO₄.

**Rule 5:** Most hydroxide salts are only slightly soluble. The important soluble hydroxides are NaOH, KOH, and Ca(OH)₂ (marginally soluble)

**Rule 6:** Most sulfide(S²⁻), carbonate(CO₃²⁻), and phosphate (PO₄³⁻) salts are only slightly soluble.
Precipitation Reactions

**Student Question**

How many of the following compounds are soluble in water?

- $\text{Ba}_3(\text{PO}_4)_2$
- $\text{Ba(NO}_3)_2$
- $\text{K}_2\text{CO}_3$
- $\text{Cu(OH)}_2$

a) 1  
b) 2  
c) 3  
d) 4
Precipitation Reactions

- **Step 1**: Determine if a chemical reaction takes place and write a balanced chemical reaction.

  **Note**: If no reaction takes place stop here.

- **Step 2**: Write the complete ionic equation (show all ions separately).

- **Step 3**: Determine the spectator ions (ions that do not participate in the reaction).

- **Step 4**: Write the net ionic equation (only species that participate in the reaction).
Precipitation Reactions

- **Molecular Equation:** Balanced chemical reaction.

  *Example:* \( \text{CoCl}_2(\text{aq}) + \text{Ca(OH)}_2(\text{aq}) \rightarrow \text{Co(OH)}_2(\text{s}) + \text{CaCl}_2(\text{aq}) \)

- **Complete Ionic Equation:** A balanced equation expressed in terms of the cations and anions present.

  *Example:* \( \text{Co}^{2+}(\text{aq}) + 2\text{Cl}^-(\text{aq}) + \text{Ca}^{2+}(\text{aq}) + 2\text{OH}^-(\text{aq}) \rightarrow \text{Co(OH)}_2(\text{s}) + \text{Ca}^{2+}(\text{aq}) + 2\text{Cl}^-(\text{aq}) \)

  **Note:** For substances that do not completely dissociate, they are not broken apart in complete ionic and net ionic equations. This is most often seen in acid base reactions.

  *Example:* \( \text{HCN}(\text{aq}) + \text{Na}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{Na}^+(\text{aq}) + \text{CN}^-(\text{aq}) + \text{H}_2\text{O}(\text{l}) \)

  HCN is a weak acid, therefore, most of the compound does not dissociate. When the complete and net ionic equations are written out, HCN is kept as a unit.
**Precipitation Reactions**

- **Net Ionic Equation**: The equation showing the net change in a chemical reaction, obtained by canceling the spectator ions in a complete ionic equation.

  Example: \( \text{Co}^{2+}(aq) + 2\text{OH}^-(aq) \rightarrow \text{Co(OH)}_2(s) \)

- **Spectator Ions**: Ions that do not play a role in the chemical reaction.

  Examples: \( \text{Ca}^{2+} \) and \( \text{Cl}^- \)
Does a reaction occur when the following substances are mixed? If so, write out the molecular equation, complete ionic equation, and net ionic equation, as well as identify the spectator ions.

- NH₄Cl(aq) and H₂SO₄(aq)

- K₂CO₃(aq) and SnCl₂(aq)
Precipitation Reactions

**Student Question**

What is the final concentration of OH\(^-\) ions in solution if 1 mol of CsNO\(_3\) is mixed with 2 mol of NaOH and the final volume of solution is 1.0 L.

\[
\text{CsNO}_3(\text{aq}) + \text{NaOH}(\text{aq}) \rightarrow \text{CsOH(s)} + \text{NaNO}_3(\text{aq})
\]

a) 0 M  

b) 0.5 M  

c) 1.0 M  

b) 2.0 M  

e) None of the above
Applications of Precipitation Reactions

- **Make Compounds**: Choose starting solutions that form a precipitate of the desired insoluble compound when they are mixed.

- **Qualitative Analysis**: Determine substances present in a sample.

  **Note**: The formation of a precipitate is used to confirm the identity of certain ions.

- **Quantitative Analysis**: Determine the amount of each substance or element present.

  **Note**: This can be done by gravimetric analysis [the amount of substance present is determined by measurements of mass (i.e. measure the amount of precipitate)].
Precipitation Reactions

**Student Question**

How many grams of K₄Fe(CN)₆ are required to precipitate all the Cd²⁺ ions as Cd₂Fe(CN)₆ from 0.00400 L of 0.15 M cadmium chloride solution according to the following equation.

\[ \text{K₄Fe(CN)₆}(aq) + 2\text{CdCl₂}(aq) \rightarrow 4\text{KCl}(aq) + \text{Cd₂Fe(CN)₆}(s) \]

Helpful Information: \( M_{K₄Fe(CN)₆} = 368.37 \frac{g}{mol} \)

- a) 0.11 g
- b) 7.4 g
- c) 2.4 g
- d) 0.78 g
- e) None of the above
Acid Base Reactions

**Arrhenius Acids and Bases**

- **Acid**: A compound that forms hydrogen ($\text{H}^+$) ions in water.
  
  *Example*: HCl(aq) acid
  
  CH$_4$(aq) not an acid because it does not release ($\text{H}^+$) ions in solution.

- **Base**: A compound that produces hydroxide ($\text{OH}^-$) ions in water
  
  *Examples*: NaOH(aq) base
  
  NH$_3$ base because NH$_3$(aq) + H$_2$O(l) → NH$_4^+$ (aq) + OH$^-$ (aq).
In this titration, HCl, analyte, is in the Erlenmeyer flask and NaOH, titrant, is in the burette. Phenolphthalein, which is clear in the presence of an acid and pink in the presence of a base, is added to the HCl. NaOH is added to the HCl and Phenolphthalein until the solution turns pink, which signifies the equivalence point has been reached.
Acid Base Reactions

**Titration Calculations**

- **Step 1:** Write the equation.
- **Step 2:** Find the number of moles of titrant, or “known” substance needed to get to the equivalence point.
  - **Equivalence point:** The stage in a titration when exactly the right volume of solution needed to complete the reaction has been added.
- **Step 3:** Use the mole ratio to find the moles of analyte.
- **Step 4:** Turn moles into M, mass%, etc.
During a titration a 0.02500 L sample of $\text{H}_2\text{SO}_4$ required 24.16 mL of 0.106 M NaOH to reach the equivalence point. What is the initial concentration of the $\text{H}_2\text{SO}_4$?

If 35 mL of 0.45 M HCl(aq) and 25 mL of 1.0 M NaOH(aq) are mixed together will the solution be acidic, basic, or neutral?
Assigning Oxidation Number

1) The oxidation number (ON) of an element uncombined with another element is zero: Na(s), H\(_2\)(g), I(s) ...

2) For monoatomic ions, the charge is the ON: Na\(^+\) ON = +1.

3) The ON’s of elements in group 1 equal 1 (ex. Lithium ON = +1) ON’s of elements in group 2 equal 2 (ex. Magnesium ON = +2).

4) The ON of fluorine is always -1 in compounds.

5) The ON of the other elements in group 7 usually equals -1.

6) The ON of oxygen is usually -2 in compounds. Exceptions are fluorine compounds and peroxide (a compound that contains an O-O single bond).

7) Hydrogen's ON is +1 when combine with non metals and -1 when combined with metals.

8) The sum of the ON’s of all the atoms in a species is equal to its total charge.
Oxidation Reduction Reactions

**Student Question**

What is the oxidation number of chromium in the ionic compound Na$_2$Cr$_2$O$_7$?

a) 2  
b) 6  
c) 7  
d) 12  
e) None of the above
Oxidation Reduction Reactions

Determining Which Element is Oxidized and Which is Reduced

- **Step 1:** Assign oxidation numbers.
- **Step 2:** Use oxidation numbers and ‘OIL RIG’ to identify which element is oxidized and which is reduced.

**Note:** If the question asks which substance is oxidized; instead of giving the element that is oxidized give the compound that the element is in.
Oxidation Reduction Reactions

- **Oxidizing Agent**: A species that removes electrons from a species being oxidized in a redox reaction.

  **Note**: The oxidizing agent contains the species being reduced.

- **Reducing Agent**: The species that supplies electrons to a substance being reduced in a redox reaction.

  **Note**: The reducing agent contains the species being oxidized.
Oxidation Reduction Reactions

Student Question

Identify the oxidizing agent in the reaction

\[ \text{Al}(s) + \text{Fe}_2\text{O}_3(s) \rightarrow \text{Fe}(s) + \text{Al}_2\text{O}_3(s) \]

a) Al(s)
b) Fe\(_2\)O\(_3\)(s)
c) Fe(s)
d) Al\(_2\)O\(_3\)(s)
Balancing Simple Redox Reactions
(no oxygen or hydrogen in the reaction)

- **Step 1**: Write unbalanced half reactions.
- **Step 2**: Balance half reactions (atoms and electrons).
- **Step 3**: Multiply half reactions by an integer so that number of electrons match, then add reactions together.
Oxidation Reduction Reactions

Balancing Redox Reactions in Acidic Conditions

- **Step 1**: Write unbalanced half reactions.
- **Step 2**: Balance half reactions except for O and H.
- **Step 3**: Balance O by using H₂O.
- **Step 4**: Balance H by using H⁺.
- **Step 5**: Balance electrons in each half reaction.
- **Step 6**: Multiply half reactions by an integer so that number of electrons match, then add reactions together.
Balancing Redox Reactions in Basic Conditions

- **Step 1**: Balance the reaction as if it were in acidic conditions.
- **Step 2**: Determine the number of $H^+$ in the balanced equation.
- **Step 3**: Add the same number of $OH^-$ as there are $H^+$ to BOTH sides of the equation.
- **Step 4**: The $H^+$ and $OH^-$ on one side of the reaction will combine and form $H_2O$.
- **Step 5**: Simplify your reaction (combined waters) if necessary.
**Student Question**

Balance the following equation in basic solution and determine the coefficient of OH\(^{-}\) and its location (right or left side) in the equation.

\[
\text{Ce}(s) + \text{PO}_4^{3-}(aq) \rightarrow \text{HPO}_3^{2-}(aq) + \text{Ce(OH)}_3(s)
\]

a) 3, left  
b) 5, left  
c) 5, right  
d) 3, right  
e) none of the above
Big Idea: Reactions can be broken down into subgroups. Three types of reactions are (1) precipitation (a solid is formed from 2 aqueous solutions), (2) acid/base (salt and H₂O are produced), and (3) oxidation/reduction (e⁻ are transferred). Many reactions involve species that are in solution. Molarity (mol per liter) is used to describe the concentration of species in a solution.

Electrolytes
- Be able to identify electrolytes (ionic compounds and acids) and nonelectrolytes (molecular compounds that are not acids). (15,17)

Molarity and Dilutions
- Be able to calculate the molarity of a solution (13,21,26)
  - \[ M = \frac{n}{V} \]
- Be able to determine the molarity of solutions after dilutions (27,28,30)
  - \[ M_1V_1 = M_2V_2 \]
Take Away From Chapter 4

- **Precipitation Reactions**
  - Know the solubility rules and be able to predict when a precipitate forms. (36,40,42)
  - Be able to write the molecular, complete ionic, and net ionic equations for a reaction and identify the spectator ions. (39)
  - Be able to apply stoichiometry to precipitation reactions. (44,45,49,53,55)

- **Acid and Base Reactions**
  - Know that when an acid and a base react they form water and a salt.
  - Be able to use titration data to get the concentration of an acid or base. (61,64,66,67)
  - Be able to apply stoichiometry to acid/base reactions. (71)
Take Away From Chapter 4

- **Oxidation and Reduction Reactions**
  - Be to assign oxidation numbers. (75,77)
  - Be able to identify what is oxidized and what is reduced in a redox reaction (OIL RIG) as well as identify the oxidizing and reducing agents. (78)
  - Be able to balance redox reactions in acid and base solutions. (81,82)
  - Be able to apply stoichiometry to redox reactions. (85,86)

Numbers correspond to end of chapter questions.