Chem 103, Section F0F
Unit V - Chemical Reactions and Chemical Properties
Lecture 14

• Acid-base reactions
• Oxidation-reduction reactions
• Elements in Redox reactions
• Reversibility of reactions

Lecture 14 - Reactions, con’d
Reading in Silberberg
• Chapter 4, Section 4
  - Acid-Base Reactions
• Chapter 4, Section 5
  - Oxidation-Reduction (Redox) Reactions
• Chapter 4, Section 6
  - Elements in Redox Reactions
• Chapter 4, Section 7
  - Reaction Reversibility and the Equilibrium State

Lecture 14 - Introduction

Last lecture we looked at precipitation reaction
• Today we will look at two other reaction types that typically occur in water:
  - Acid-base reactions
  - Oxidation-reduction reactions

Lecture 14 - Acid-Base Reactions

In acid-base reactions, water is not only the solvent, but also participates as a reactant or product in the reaction.

In an acid-base reaction, an acid reacts with bases
• Since the one (the acid) counteracts the other (the base), this is often referred to an a **neutralization** reaction.

Lecture 14 - Acid-Base Reactions

There are numerous definitions of acids and bases, but in this course we will focus on only a couple of these.

• Operational definition:
  - **An acid**, when added to a solution, causes the pH of the solution to go down.
  - **A base**, when added to a solution, causes the pH of the solution to go up.

In lab we defined **pH** as: 

\[ \text{pH} = -\log([H^+]) \]

We also discussed that \([H^+][OH^-] = 10^{-14} \text{ M}^2\)

<table>
<thead>
<tr>
<th>Conditions</th>
<th>([H^+])</th>
<th>([OH^-])</th>
<th>pH</th>
</tr>
</thead>
<tbody>
<tr>
<td>neutral</td>
<td>(10^{-7}) M</td>
<td>(10^{-7}) M</td>
<td>(\text{pH} = 7)</td>
</tr>
<tr>
<td>(\text{[H]})</td>
<td>([H^+] = [OH^-])</td>
<td></td>
<td></td>
</tr>
<tr>
<td>acidic</td>
<td>(&gt; 10^{-7}) M</td>
<td>(&lt; 10^{-7}) M</td>
<td>(\text{pH} &lt; 7)</td>
</tr>
<tr>
<td>basic</td>
<td>(&lt; 10^{-7}) M</td>
<td>(&gt; 10^{-7}) M</td>
<td>(\text{pH} &gt; 7)</td>
</tr>
<tr>
<td></td>
<td>([H^+] &lt; [OH^-])</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Lecture 14 - Acid-Base Reactions

Another definition of acids and bases is looks at their ability to produce H\(^+\) and OH\(^-\) ions when dissolved in water:

- **Arrhenius definition:**
  - An acid is a substance that produces H\(^+\) ions when dissolved in water.
  \[ HX + H_2O \rightleftharpoons H^+(aq) + X^-(aq) \]
  - A base is a substance that produces OH\(^-\) ions when dissolved in water.
  \[ MOH + H_2O \rightleftharpoons M^+(aq) + OH^-(aq) \]

- Acids are different than the typically ionic compound.
  - They are covalent molecules that can behave like ionic compounds by releasing a H\(^+\) in water.
    - For example:
      \[ HCl \rightarrow H^+(aq) + Cl^-(aq) \]
  - The difference between acids and ionic compounds, is that the un-ionized form of an acid is also soluble in water.

Lecture 14 - Acid-Base Reactions

In lab we saw that when ionic compounds dissolve in water they are strong electrolytes.

- This is not the case for acids because they do not have to ionize to be soluble in water.
  - The degree to which an acid ionizes in water is a measure of its acid-base strength.
    - Strong acids dissociate completely into ions when dissolved in water and are strong electrolytes.
    - Weak acids dissociate only a little when dissolved in water and are weak electrolytes.

The list of strong acids is a fairly small one.

<table>
<thead>
<tr>
<th>Table 4.2 Strong and Weak Acids and Bases</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Acids</strong></td>
</tr>
<tr>
<td>Strong Hydrochloric acid, HCl</td>
</tr>
<tr>
<td>Hydrobromic acid, HBr</td>
</tr>
<tr>
<td>Nitric acid, HNO(_3)</td>
</tr>
<tr>
<td>Sulfuric acid, H(_2)SO(_4)</td>
</tr>
<tr>
<td>Acetic acid, CH(_3)COOH</td>
</tr>
</tbody>
</table>

Weak bases, such as ammonia, do not contain OH\(^-\), but rather produce them by reacting with water:

\[ NH_3(aq) + H_2O(l) \rightarrow NH_4^+(aq) + OH^-(aq) \]
\[ H\_C\_H\_3 + H\_2O \rightarrow H\_C\_H\_2\_N\^\_H + OH^- \]
Lecture 14 - Acid-Base Reactions

The reaction of a strong acid with a strong base produces a neutral salt solution:

Like precipitation reactions, acid-base reactions are double displacement reactions.
- Instead of producing a precipitate from ions, they produce a molecular compound (H₂O) from ions (H⁺, OH⁻).

Acid-base titration
- In lab we saw that there is a sudden increase in the pH when a stoichiometric amount of base has been added to an acid.
  - The rise comes at the end point or equivalence point in the titration, and can be used to determine an unknown concentration for an acid or a base.

Acid-base reactions can also be considered as proton-transfer reactions.
- This leads to another definition of acids and bases
  - Brønsted-Lowry definition:
    - An acid, donates a proton to a base.
    - A base, accepts a proton from an acid.

For example
- HCl is an acid because it donates a proton to water
  - HCl is the acid
  - H₂O is the base
- NH₃ is a base because it accepts a proton from water
  - NH₃ is the base
  - H₂O is the acid
Lecture 14 - Acid-Base Reactions

The chemical equation for the acid-base reaction that involves a weak acid works a little differently.

- This is because the acid only dissociates a small amount on its own in water.

\[ \text{CH}_3\text{COOH}(aq) + \text{NaOH}(aq) \rightarrow \text{H}_2\text{O}(l) + \text{NaCH}_3\text{COO}(aq) \]

- When the base is added, the proton is transferred directly from the acid to the base (OH\(^-\)).

\[ \text{CH}_3\text{COOH}(aq) + \text{OH}^-(aq) \rightarrow \text{CH}_3\text{COO}^-(aq) + \text{H}_2\text{O}(l) \]

\[ \text{CH}_3\text{COO}^-(aq) + \text{H}_2\text{O}(l) \rightarrow \text{CH}_3\text{COOH}(aq) + \text{OH}^-(aq) \]

Lecture 14 - Clicker Question 1

How many moles of chloride ions are present in 500 mL of a 0.250 M solution of magnesium chloride?

A) 0.500 mol Cl\(^-\)
B) 0.125 mol Cl\(^-\)
C) 0.250 mol Cl\(^-\)
D) 0.0625 mol Cl\(^-\)

Lecture 14 - Clicker Question 2

How many moles of hydronium (hydrogen) ions are present in 500 mL of a 0.250 M solution of nitric acid?

A) 0.500 mol H\(_3\)O\(^+\)
B) 0.125 mol H\(_3\)O\(^+\)
C) 0.250 mol H\(_3\)O\(^+\)
D) 0.0625 mol H\(_3\)O\(^+\)

Lecture 14 - Clicker Question 3

What is the pH of a 0.250 M solution of nitric acid?

A) 0.60
B) 1.0
C) 0.25
D) 2.5

Lecture 14 - Acid-Base Reactions

Weak acids dissociate only partially into ions when dissolved in water.

- This makes it more difficult to predict the hydrogen ion concentration, or pH, for a weak acid.
  - You will learn how to do this when you go on to Chem 104.
Lecture 14 - Acid-Base Reactions

Gas-Forming Reactions
- Acid-base reactions that involve carbonates and bicarbonates as the base, lead to the formation of carbonic acid (H$_2$CO$_3$)
- Carbonic acid is unstable, and spontaneously decomposes to form CO$_2$ and H$_2$O:

$$2\text{H}_2\text{O} + \text{K}_2\text{CO}_3 \rightarrow 2\text{KCl} + \text{H}_2\text{CO}_3$$

Lecture 14 - Oxidation-Reduction Reactions

There are a wide range of important reactions that come under classification oxidation-reduction (redox) reactions:
- The formation of compounds from the elements
- The formation of elements from compounds
- Combustion reactions
- Reactions that generate electricity in batteries
- Biochemical reactions that extract energy from the foods we eat.

Redox reactions do not have to occur in an aqueous solution.

Lecture 14 - Oxidation-Reduction Reactions

Like acid-base reactions, oxidation-reduction (redox) reactions are viewed as transfer reactions.
- In this case, however, it is electrons instead of protons that are being transferred.

Redox reactions involving elements
- If the transfer is complete, ionic compounds are formed.
- If the transfer is not complete, covalent compounds with polar covalent bonds are formed.

Lecture 14 - Oxidation-Reduction Reactions

Examples of a redox reaction that lead to formation of ionic and covalent compounds:

A) Formation of an ionic compound
B) Formation of a covalent compound

Lecture 14 - Oxidation-Reduction Reactions

Some definitions:
- **Oxidation** - the loss of electrons
- **Reduction** - the gain of electrons

Oxidation cannot occur without reduction, and vice versa
- **Oxidation agent** - Oxygen is the oxidation agent because it oxidizes (takes electrons from) the magnesium.
- **Reduction agent** - Magnesium is the reduction agent because it reduces (gives electrons to) the magnesium.

Lecture 14 - Clicker Question 4

Is the following a redox reaction?

$$\text{NH}_3 (aq) + \text{HCl} (aq) \rightarrow \text{NH}_4\text{Cl} (aq)$$

A) Yes
B) No
In the following redox reaction, does the sulfuric acid (H$_2$SO$_4$) act as an oxidizing agent or a reduction agent?

$$4 \text{H}^+(aq) + \text{SO}_4^{2-}(aq) + 2 \text{NaI(s)} \rightarrow 2 \text{Na}^+(aq) + \text{I}_2(aq) + \text{SO}_2(g) + 2 \text{H}_2\text{O(l)}$$

A) Oxidizing agent  
B) Reducing agent

Oxidation numbers can be used to as a helpful “bookkeeping” device to follow the electrons around in a redox reaction.

- For binary ionic compounds, the oxidation number is equal to the charge on each ion.
- For covalent molecules, the bonding electrons are given to one of the elements.
  - There is a set of rules that are used to determine which elements get the electrons.

In a reaction

- Oxidation is evidenced by an increase in the oxidation number.
- Reduction is evidenced by a reduction in the oxidation number.

You can find the highest and lowest possible oxidation number for most main-group elements from their locations on the periodic table:

Give the oxidation number of the sulfur in the following:

A) SOCl$_2$  
B) H$_2$S$_2$  
C) SO$_4^{2-}$  
D) SO$_2$  

Oxidation numbers can be used to as a helpful “bookkeeping” device to follow the electrons around in a redox reaction.
In balancing a redox reaction, make sure that the number of electrons lost by the reducing agent is equal to the number of electrons gained by the oxidizing agent.

- **Step 1**: Assign oxidation numbers to all elements in the reaction.
- **Step 2**: From the changes in oxidation numbers, identify the oxidized and reduced species.
- **Step 3**: Compute the number of electrons lost in the oxidation and gained in the reduction from the oxidation number changes.
- **Step 4**: Multiply one or both of these numbers by appropriate factors to make the electrons lost equal to the electrons gained, and use the factors as balancing coefficients
- **Step 5**: Complete the balancing by inspection, adding states of

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**Lecture 14 - Question 7**

Use oxidation numbers to balance the following reaction:

$$\text{HNO}_3(\text{aq}) + \text{K}_2\text{CrO}_4(\text{aq}) + \text{Fe(NO}_3\text{)_2(aq)} \rightarrow \text{KNO}_3(\text{aq}) + \text{Fe(NO}_3\text{)_3(aq)} + \text{Cr(NO}_3\text{)_3(aq)} + \text{H}_2\text{O(l)}$$

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**Lecture 14 - Elements in Redox Reactions**

When ever elements appear on one side of an equation as free elements, and on the other side combined as part of a compound, they have undergone a redox reaction.

Classifications of reactions involving elements:

- **Combination reactions**
  $$\text{X} + \text{Y} \rightarrow \text{Z}$$

- **Decomposition reactions**
  $$\text{Z} \rightarrow \text{X} + \text{Y}$$

- **Displacement reactions**
  $$\text{X} + \text{YZ} \rightarrow \text{XZ} + \text{Y}$$

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**Lecture 14 - Elements in Redox Reactions**

Combination reactions
- A metal combines with a nonmetal to form an ionic compound
  - Aluminum plus oxygen form aluminum oxide
- Two nonmetals combine to form a covalent compound
  - Nitrogen plus hydrogen form ammonia
- A compound combines with an element to form a second compound.
  - Carbon monoxide combines with oxygen to form carbon dioxide

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**Lecture 14 - Elements in Redox Reactions**

Decomposition reactions
- Decomposition into elements
  - Decomposition of mercury(II) oxide into mercury and oxygen.

- This is the reaction that lead Lavoisier to the discover of oxygen.
Lecture 14 - Elements in Redox Reactions

The activity series for metals participating in displacement reactions:

- U
- K
- Ca
- Ba
- Sr
- Na
- Mg
- Zn
- Sn
- Pb
- Hg
- Ag
- Au

- Can displace H₂ from acid
- Can displace H₂ from water
- Can displace H₂ from NaCl

Combustion reactions
- Are redox reactions
- Do not fall neatly into one of the classifications based on the number of reactants and produces
  - Combustion of butane to form CO₂ and H₂O

Lecture 14 - Reaction Reversibility and Equilibrium State

Not all reactions go to completion
- They instead reach a dynamic equilibrium
  - Involving both a forward and a reverse reaction.
- The reactions we saw involving are weak acids and bases are examples.