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.*

Lecture 14 - Acid-Base Reactions

In lab we defined *pH* as ***pH* = -log([H⁺])** $pH = -\log([H^+])$

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

conditions	$[H^+]$	$[OH^-]$	*pH*
neutral	10^{-7} M	10^{-7}	*pH* = 7 $[H^+] = [OH^-]$
acidic	> 10^{-7} M	< 10^{-7} M	*pH* < 7 $[H^+] > [OH^-]$
basic	< 10^{-7} M	> 10^{-7} M	*pH* > 7 $[H^+] < [OH^-]$

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.*

$$HX \xrightarrow{H_2O} H^+(aq) + X^-(aq)$$

  - *A **base** is a substance that produces* OH^- ions when dissolved in water

$$MOH \xrightarrow{H_2O} M^+(aq) + OH^-(aq)$$

Lecture 14 - Acid-Base Reactions

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:

$$HCl \xrightarrow{H_2O} HCl\ (aq)$$

$$HCl\ (aq) \longrightarrow H^+(aq) + Cl^-(aq)$$

- 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

Though we often write H^+, the H^+ that is released is covalent att

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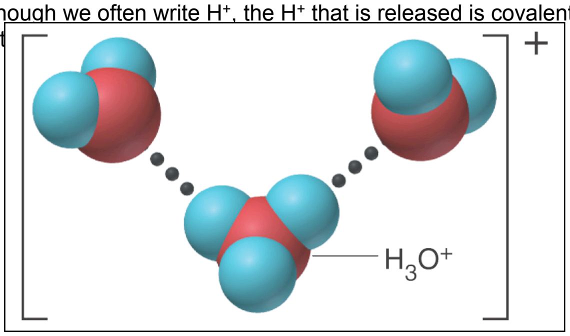

Lecture 14 - Acid-Base Reactions

In … n ionic compoun … the … olytes.

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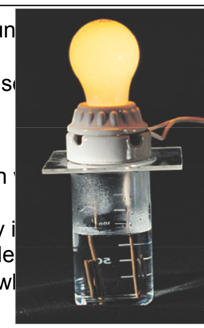


Lecture 14 - Acid-Base Reactions

The list of strong acids is a fairly small one.

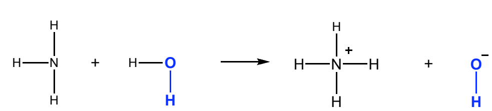
Table 4.2 Strong and Weak Acids and Bases

Acids
Strong
Hydrochloric acid, HCl
Hydrobromic acid, HBr
Hydriodic acid, HI
Nitric acid, HNO_3
Sulfuric acid, H_2SO_4
Perchloric acid, $HClO_4$
Weak
Hydrofluoric acid, HF
Phosphoric acid, H_3PO_4
Acetic acid, CH_3COOH (or $HC_2H_3O_2$)

Lecture 14 - Acid-Base Reactions

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

$$NH_3(aq) + HOH\ (l) \longrightarrow NH_4^+(aq) + OH^-(aq)$$



Lecture 14 - Acid-Base Reactions

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

$$HCl\,(aq) + NaOH\,(aq) \xrightarrow[\text{molecular}]{} NaCl\,(aq) + HOH\,(l)$$

$$H^+\,(aq) + Cl^-\,(aq) + Na^+\,(aq) + OH^-\,(aq) \xrightarrow[\text{total ionic}]{} HOH\,(l) + Na^+\,(aq) + Cl^-\,(aq)$$

$$H^+\,(aq) + OH^-\,(aq) \xrightarrow[\text{net ionic}]{} HOH\,(l)$$

- Like precipitation reactions, acid-bases reactions are double displacement reactions
  - Instead of producing a precipitate from ions, they produce a molecular compound (H_2O) from ions (H^+, OH^-).

Lecture 14 - Acid-Base Reactions

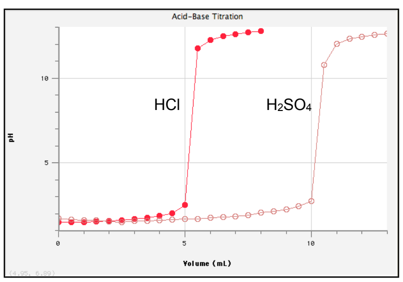
Acid-base titration

- 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.

Lecture 14 - Acid-Base Reactions

Acid-base titration

- The titration of 5 mL 1.0 M HCl and 5 mL 1.0 M H_2SO_4 with 1.0 M NaOH



Lecture 14 - Acid-Base Reactions

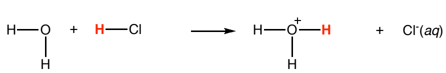
Acid-base reactions can also be considered as proton-transfer reactions.
This leads to another definition of acids and bases

- Brønsted-Lowry definition:
  - *An **acid**, donates a proton to a base.*
  - *A **base**, accepts a proton from an acid.*

Lecture 14 - Acid-Base Reactions

For example

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

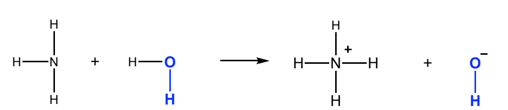


- HCl is an acid because it donates a proton to water
  - HCl is the acid
  - H_2O is the base

Lecture 14 - Acid-Base Reactions

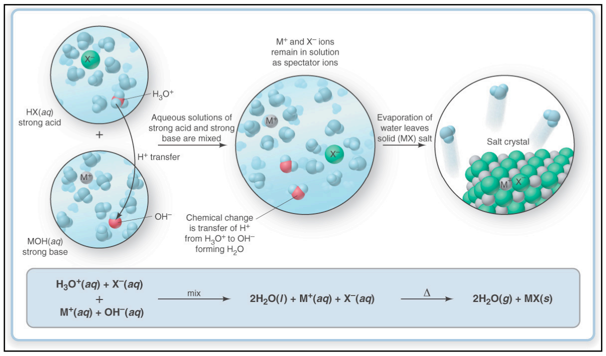
For example

$$NH_3(aq) + HOH\,(l) \longrightarrow NH_4^+(aq) + OH^-(aq)$$



- NH_3 is a base because it accepts a proton from water
  - NH_3 is the base
  - H_2O is the acid

Lecture 14 - Acid-Base Reactions



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_3O^+
B) 0.125 mol H_3O^+
C) 0.250 mol H_3O^+
D) 0.0625 mol H_3O^+

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

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.

$$CH_3COOH\,(aq) + NaOH\,(aq) \longrightarrow HOH\,(l) + NaCH_3COO(aq)$$
molecular

$$CH_3COOH\,(aq) + Na^+\,(aq) + OH^-\,(aq) \longrightarrow HOH\,(l) + Na^+\,(aq) + CH_3COO^-\,(aq)$$
total ionic

$$CH_3COOH\,(aq) + OH^-\,(aq) \longrightarrow HOH\,(l) + CH_3COO^-\,(aq)$$
net ionic

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

Lecture 14 - Acid-Base Reactions

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

$$CH_3COOH\,(aq) \rightleftharpoons H^+ + CH_3COO^-(aq)$$
96% 4%

$$CH_3COOH\,(aq) + HOH\,(aq) \rightleftharpoons H\overset{+}{O}H\,(l) + CH_3COO^-(aq)$$
(with H below the O)
96% 4%

- 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_2CO_3)
- Carbonic acid is unstable, and spontaneously decomposes to form CO_2 and H_2O:

$$2\,HCl\,(aq) + K_2CO_3\,(aq) \longrightarrow 2\,KCl\,(aq) + H_2CO_3\,(aq)$$
$$H_2CO_3\,(aq) \longrightarrow H_2O\,(l) + CO_2\,(g)$$
$$\overline{2\,HCl\,(aq) + K_2CO_3\,(aq) \longrightarrow 2\,KCl\,(aq) + H_2O\,(l) + CO_2\,(g)}$$

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.

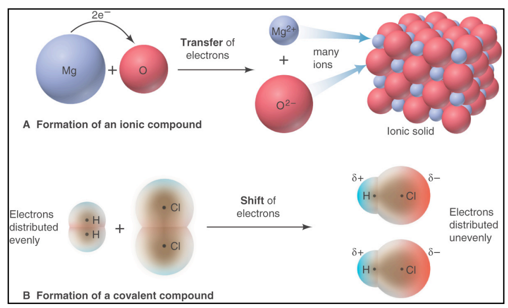
Lecture 14 - Oxidation-Reduction Reactions

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:



Lecture 14 - Oxidation-Reduction Reactions

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

$$Mg \longrightarrow Mg^{2+} \quad 2 \quad 2\,e^-$$

- **Reduction** - the gain of electrons

$$1/2\,O_2 + 2\,e^- \longrightarrow O^{2-}$$

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?

$$NH_3\,(aq) + HCl\,(aq) \longrightarrow NH_4Cl\,(aq)$$

A) Yes

B) No

Lecture 14 - Clicker Question 5

In the following redox reaction, does the sulfuric acid (H_2SO_4) act as an *oxidizing agent* or a *reduction agent*?

$$4\,H^+\,(aq) + SO_4^{-2}\,(aq) + 2\,NaI\,(s) \longrightarrow 2\,Na^+\,(aq) + I_2\,(aq) + SO_2\,(g) + 2\,H_2O\,(l)$$

A) Oxidizing agent

B) Reducing agent

Lecture 14 - Oxidation-Reduction Reactions

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 equation 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.

Lecture 14 - Oxidation-Reduction Reactions

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

Table 4.3 Rules for Assigning an Oxidation Number (O.N.)

General Rules

1. For an atom in its elemental form (Na, O_2, Cl_2, etc.): O.N. = 0
2. For a monatomic ion: O.N. = ion charge
3. The sum of O.N. values for the atoms in a molecule or formula unit of a compound equals zero. The sum of O.N. values for the atoms in a polyatomic ion equals the ion's charge.

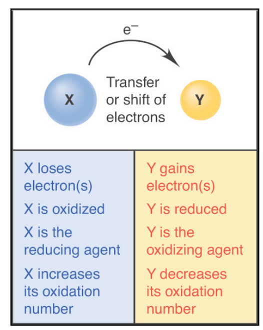
Rules for Specific Atoms or Periodic Table Groups

1. For Group 1A(1):	O.N. = +1 in all compounds
2. For Group 2A(2):	O.N. = +2 in all compounds
3. For hydrogen:	O.N. = +1 in combination with nonmetals
	O.N. = −1 in combination with metals and boron
4. For fluorine:	O.N. = −1 in all compounds
5. For oxygen:	O.N. = −1 in peroxides
	O.N. = −2 in all other compounds (except with F)
6. For Group 7A(17):	O.N. = −1 in combination with metals, nonmetals (except O), and other halogens lower in the group

Lecture 14 - Oxidation-Reduction Reactions

In a reaction

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



Lecture 14 - Question 6

Give the oxidation number of the sulfur in the following:

A) $SOCl_2$

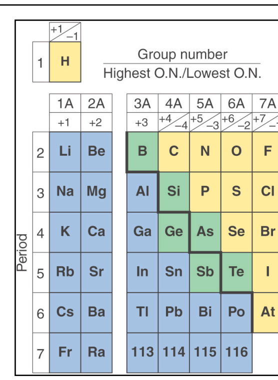
B) H_2S_2

C) SO_4^{2-}

D) SO_2

Lecture 14 - Oxidation-Reduction Reactions

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



Lecture 14 - Oxidation-Reduction Reactions

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 approiate factors to make the electrons lost equal to the electrons gained, and use the factors as balancing coeficients
- *Step 5*: Complete the balancing by inspection, adding states of

Lecture 14 - Question 7

Use oxidation numbers to balance the following reaction:

$$HNO_3\,(aq) + K_2CrO_4\,(aq) + Fe(NO_3)_2\,(aq) \longrightarrow KNO_3\,(aq) + Fe(NO_3)_3\,(aq) + Cr(NO_3)_3\,(aq) + H_2O\,(l)$$

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

$$X + Y \longrightarrow Z$$

- Decomposition reactions

$$Z \longrightarrow X + Y$$

- Displacement reactions

$$X + YZ \longrightarrow XZ + Y$$

Lecture 14 - Elements in Redox Reactions

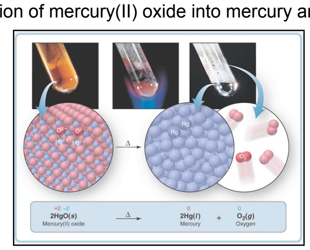
Combination reactions

- 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

Lecture 14 - Elements in Redox Reactions

Decomposition reactions

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

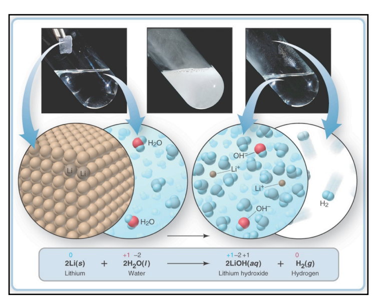


- This is the reaction that lead Lavoisier to the discover of oxygen.

Lecture 14 - Elements in Redox Reactions

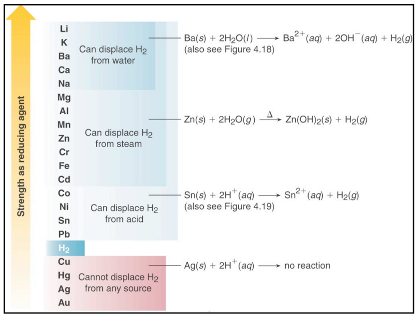
Displacement Reactions

- Double displacement (metathesis) *versus* single displacement reactions



Lecture 14 - Elements in Redox Reactions

The activity series for metals participating in displacement reactions:



Lecture 14 - Elements in Redox Reactions

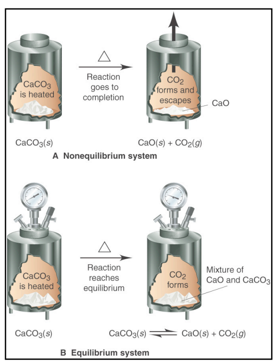
Combustion reactions

- 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 CO_2 and H_2O

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.



Unit VI - Up Next

Lecture 15 - Reactions con’d

- Bonding atoms together to make compounds
- Describing covalent compounds with formulas and names
- The covalent bond
- Bond energies and chemical change

The End