Course Specification
Term 2 (AY 21-22)

Course Name: Chemistry

Course Code: CHM71

Grade: 12 Advanced

Aim: The Grade 12 Chemistry course provides students with a college-level foundation to support future advanced coursework in Chemistry. The course will give students the opportunity to cultivate their understanding of Chemistry through inquiry-based investigations as they explore topics as Matter, Energy & Equilibrium, Oxidation & Reduction Reactions and Organic & Nuclear Chemistry.

Course Outline:

Unit	Number of Periods
Term 2	
Unit 1: Matter, Energy and Equilibrium (Continued)	
• 1.4 – Acids and Bases *(Inspire Chemistry, Unit 3 – Module 17: Lesson 1 to 4)*	12
Unit 2: Oxidation & Reduction Reactions	**18**
• 2.1 – Redox Reactions *(Inspire Chemistry, Unit 4 – Module 18: Lesson 1 to 2)*	8
• 2.2 – Electrochemistry *(Inspire Chemistry, Unit 4 – Module 19: Lesson 1 and 3)*	10

Unit 1: Matter, Energy & Equilibrium (Continued)

1.4 – Acids & Bases

Inspire Chemistry – Module 17 – Lesson 1: Introduction to Acids and Bases

CHM.5.3.04.001.01 List six **general** properties of aqueous acids (taste, color of indicators, reaction with metals, metal carbonates and bases, and electrical conductivity)

1) Aqueous solution of acids has a sour taste

2) Acids change the color of acid-base indicators (Turn blue litmus paper red and methyl orange red)

3) Acids react with metals (above hydrogen in the reactivity series) to produce salt and hydrogen

   Examples:
   $Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$
   $Zn_{(s)} + 2HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$
   $2Al(s) + 3H_2SO_{4(aq)} \rightarrow Al_2(SO_4)_{3(aq)} + 3H_{2(g)}$

4) Acids react with metal carbonates and hydrogen carbonates to produce salt, water and carbon dioxide

   Example:
   $CaCO_{3(s)} + 2\ HCl_{(aq)} \rightarrow CaCl_{2(aq)} + CO_{2(g)} + H_2O_{(l)}$
   $NaHCO_{3(s)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + CO_{2(g)} + H_2O_{(l)}$
   $NaHCO_{3(s)} + CH_3COOH_{(aq)} \rightarrow CH_3COONa_{(aq)} + CO_{2(g)} + H_2O_{(l)}$

5) Acids react with bases to produce salt and water

   Example:
   $NaOH_{(aq)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$

6) Acids conduct electric current. Some acids completely separate into ions in water and are strong electrolytes while others are weak electrolytes

CHM.5.3.04.001.02 List five **general** properties of aqueous bases (taste, color of indicators, how it feels, reaction with acids and electrical conductivity)

1) Aqueous solution of bases has bitter taste

2) Bases change the color of acid-base indicators (Turn red litmus paper blue, phenolphthalein pink and methyl orange yellow)

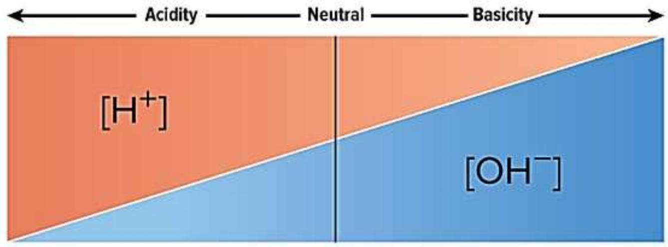
3) Dilute aqueous solutions of bases feel slippery

4) Bases react with acids to produce salt and water

   Example:
   $NaOH_{(aq)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$

5) Bases conduct electric current. Some bases completely separate into ions in water and are strong electrolytes while others are weak electrolytes

CHM.5.3.04.001.03 Differentiate among acidic, basic and neutral solutions (in terms of the relative amounts of hydrogen ions and hydroxide ions)



The relative amounts of the hydrogen ion, H^+, and hydroxide ions, OH^-, present in a solution determine whether an aqueous solution is acidic, basic or neutral.

Acidic Solution	Basic Solution	Neutral Solution
It is a solution that contains more hydrogen ions than hydroxide ions	It is a solution that contains more hydroxide ions than hydrogen ions	It is a solution that contains equal amount of hydrogen ions and hydroxide ions
$[H^+] > [OH^-]$	$[OH^-] > [H^+]$	$[H^+] = [OH^-]$

An acidic solution is a solution that contains more hydrogen ions than hydroxide ions

A basic solution is a solution that contains more hydroxide ions than hydrogen ions

A neutral solution is a solution that contains equal amount of hydrogen ions and hydroxide ions

CHM.5.3.04.001.04 Identify the color change of different indicators (Phenolphthalein, Methyl orange, Litmus paper) in acidic, basic and neutral mediums

Indicator	Acidic Medium	Neutral Medium	Basic Medium
Phenolphthalein	Colorless	Colorless	Pink
Methyl orange	Red	Orange	Yellow
Litmus Paper	Turns blue litmus paper red	No effect on blue or red litmus paper	Turns red litmus paper blue

CHM.5.3.04.001.05 Perform an experiment to investigate the color of different indicators in neutral, acidic and basic solutions

CHM.5.3.04.001.06 Compare between binary acids and oxyacids; while writing the chemical name and chemical formula of some common binary acids and oxyacids

Binary acid is an acid that contains only two elements (hydrogen and one of the more electronegative elements (as F, Cl, Br, I, S)

Examples:

HF	Hydrofluoric acid
HCl	Hydrochloric acid
HBr	Hydrobromic acid
HI	Hydroiodic acid
H_2S	Hydrosulfuric acid
HCN	Hydrocyanic acid

Oxyacid is an acid that is a compound of hydrogen, oxygen and a third element usually a non-metal.

Examples:

HNO_3	Nitric acid
HNO_2	Nitrous acid
H_2SO_4	Sulfuric acid
H_2SO_3	Sulfurous acid
H_3PO_4	Phosphoric acid
H_2CO_3	Carbonic acid
CH_3COOH	Acetic acid
$HClO_4$	Perchloric acid
$HClO_3$	Chloric acid
$HClO_2$	Chlorous acid
HClO	Hypochlorous acid

CHM.5.1.01.012.02 Write the chemical name and chemical formula of some acids commonly used in industry and the laboratory

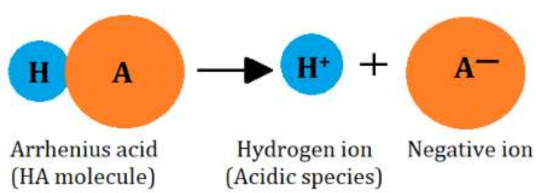
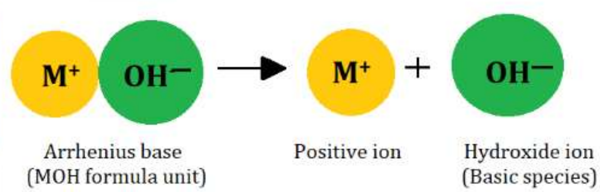
Acid	Use(s)
H_2SO_4	As a dehydrating agent
H_3PO_4	Used to make fertilizers
HNO_3	Used to make rubber, plastics and dyes
HCl	Used in removing impurities from iron and steel (pickling iron and steel)
CH_3COOH	Used in making plastics

CHM.5.1.01.012.03 Write the chemical name and chemical formula of some common bases

Examples:

$NaOH$	Sodium hydroxide
KOH	Potassium hydroxide
$LiOH$	Lithium hydroxide
$Ca(OH)_2$	Calcium hydroxide
$Mg(OH)_2$	Magnesium hydroxide
NH_3	Ammonia
CH_3NH_2	Methylamine
$C_2H_5NH_2$	Ethylamine
$C_6H_5NH_2$	Aniline

CHM.5.3.04.001.07 Use the Arrhenius model to write the conceptual definition of acids and bases (Examples, particulate models, space-filling models and ionization equations are required)

Arrhenius Acid	**Arrhenius Base**
It is a chemical compound that increases the concentration of hydrogen ions, H^+, in an aqueous solution	It is a chemical compound that increases the concentration of hydroxide ions, OH^-, in an aqueous solution
It is a substance that contains hydrogen and ionizes to produce hydrogen ions in an aqueous solution	It is a substance that contains hydroxide group and dissociates to produce a hydroxide ion in an aqueous solution
Particulate model: 	Particulate model: 
Ionization equation: $HCl_{(g)} \xrightarrow{H_2O} H^+_{(aq)} + Cl^-_{(aq)}$	Ionization equation: $NaOH_{(s)} \xrightarrow{H_2O} Na^+_{(aq)} + OH^-_{(aq)}$

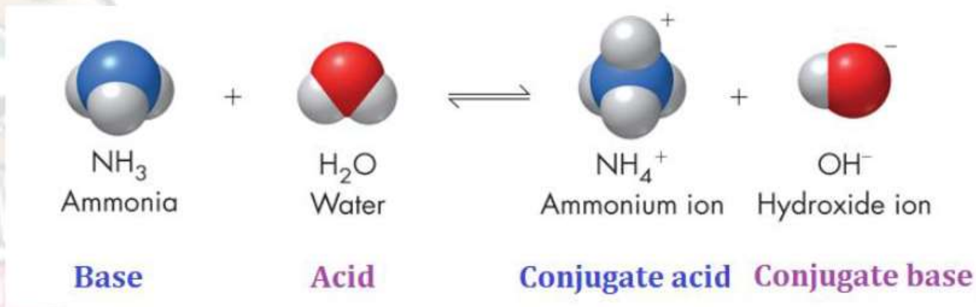
Note:

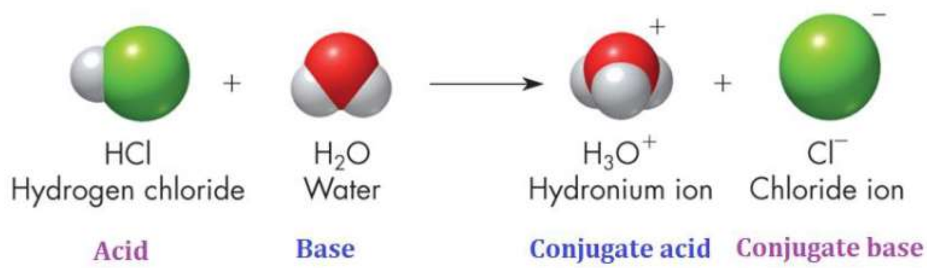
The Arrhenius model has some shortcoming.

Example: NH_3, ammonia, and sodium carbonate, Na_2CO_3, sodium carbonate, are bases, yet they do not contain hydroxide group, but they produce hydroxide ions in solution.

CHM.5.3.04.001.08 Define acids and bases according to Brønsted-Lowry theory, indicating the acid, base, conjugate acid, conjugate base and conjugate acid-base pairs, when chemical equations, formula or space-filling models are given

- Brønsted acid is a proton donor while Brønsted base is proton acceptor
- Conjugate acid is the species produced when a base accepts a proton
- Conjugate base is the species that results when an acid donates a proton
- A conjugate acid-base pair consists of two substances related to each other by the donating and accepting of a single hydrogen ion or proton





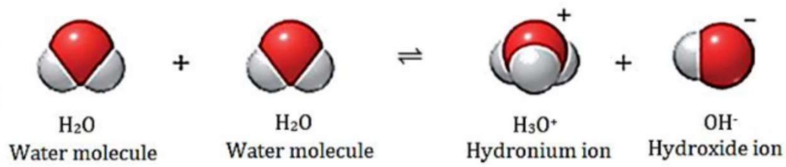
Some Conjugate Acid-Base Pairs	
Acid	**Conjugate Base**
HCl	Cl^-
HNO_3	NO_3^-
H_2SO_4	HSO_4^-
HSO_4^-	SO_4^{2-}
HF	F^-
HCN	CN^-
H_3O^+	H_2O
H_2O	OH^-
CH_3COOH	CH_3COO^-
H_2CO_3	HCO_3^-
HCO_3^-	CO_3^{2-}
NH_4^+	NH_3
H_3PO_4	$H_2PO_4^-$
$H_2PO_4^-$	HPO_4^{2-}
HPO_4^{2-}	PO_4^{3-}

CHM.5.3.04.001.09 Describe the amphoteric behavior of water, H_2O, and ammonia, NH_3 (Using chemical equation, particulate diagram and space-filling models)

Water, H_2O

$$H_2O_{(l)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + OH^-_{(aq)}$$

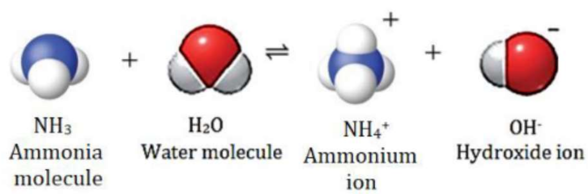
Base	Acid	Conjugate acid	Conjugate base



Ammonia, NH_3

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

Base	Acid	Conjugate acid	Conjugate base



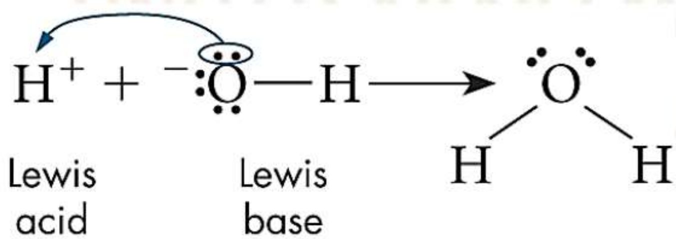
CHM.5.3.04.001.10 Distinguish among monoprotic, diprotic and triprotic (polyprotic) acids using ionization equations, examples and particulate diagrams

Monoprotic Acid HX	**Diprotic Acid H_2X**	**Triprotic or Polyprotic Acid H_3X**
- It is an acid that has one ionizable hydrogen atom. - It is an acid that can donate only one proton per molecule.	- It is an acid that has two ionizable hydrogen atoms. - It is an acid that can donate two protons per molecule.	- It is an acid that has three ionizable hydrogen atoms. - It is an acid that can donate more than two protons per molecule.
Examples: HCl HBr HI HF HNO_3 HNO_2	Examples: H_2SO_4 H_2SO_3 H_2CO_3	Example: H_3PO_4

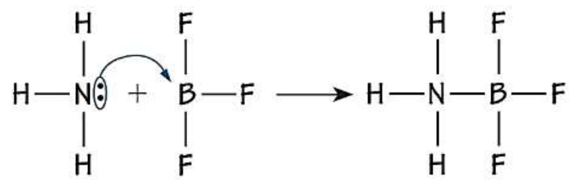
Please refer to the KPI 1.4.13 for the equations

CHM.5.3.04.001.11 Define acids and bases according to Lewis theory

Lewis Acid	Lewis Base
A <u>Lewis acid</u> is an atom, ion or molecule that is an electron-pair acceptor Any compound that has three valence electrons and forms three covalent bonds can act as a Lewis acid	A <u>Lewis base</u> is an atom, ion or molecule that is an electron-pair donor
Examples: BF_3, $AlCl_3$, SO_3	Example: NH_3, F^-, OH^-, O^{2-}



<u>Example:</u>

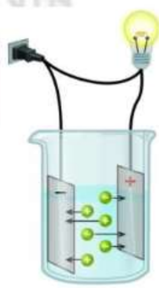
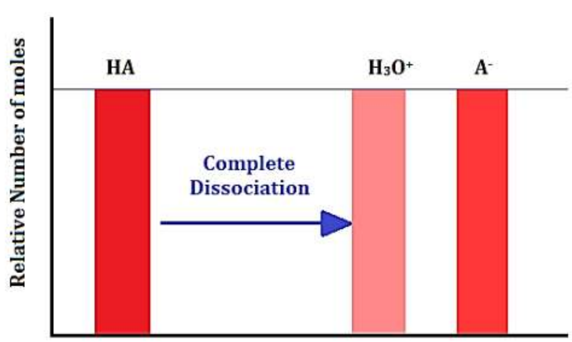
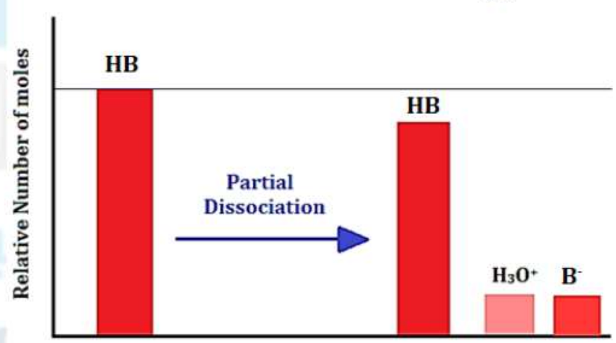


Ammonia has an unshared pair of electrons to donate. It is a Lewis base
The boron atom can accept the donated electrons. It is a Lewis acid

<u>Summary</u>

Acid-Base Definitions		
Type	**Acid**	**Base**
Arrhenius	H^+ producer	OH^- producer
Brønsted-Lowry	H^+ donor	H^+ acceptor
Lewis	Electron-pair acceptor	Electron-pair donor

CHM.5.3.04.003.01 Compare between strong and weak acids (using examples, particulate diagrams and ionization equations)

Strong Acid	**Weak Acid**
A <u>strong acid</u> is an acid that <u>ionizes completely</u> in water. The complete ionization of the strong acid is represented by **⟶.**	A <u>weak acid</u> is an acid that <u>partially ionizes</u> in water. The partial ionization of the weak acid is represented by **⇌.**
It is a strong electrolyte. It is an aqueous solution that conducts an electric current. When the electrical conductivity of a solution of a strong acid is tested, the light bulb glows brightly. 	It is a weak electrolyte. It is an aqueous solution that conducts an electric current. When the electrical conductivity of a solution of a strong acid is tested, the light bulb is dim. 
Examples: HCl, HBr, HI, H_2SO_4, HNO_3, $HClO_4$	Examples: CH_3COOH, HNO_2, H_2CO_3, H_3PO_4
Dissociation of a Strong Acid $HA_{(aq)} + H_2O_{(l)} \longrightarrow H_3O^+_{(aq)} + A^-_{(aq)}$  A solution of a strong acid, HA, mainly contains H_3O^+ & A^-	**Dissociation of a Weak Acid** $HB_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + B^-_{(aq)}$  A solution of a weak acid, HB, mainly contains HB, H_3O^+ & B^-

$HA(aq) + H_2O(l) \rightarrow H_3O^+(aq) + A^-(aq)$ Acid Base Conjugate acid Conjugate base - HA is a strong acid. Its conjugate base, A^-, is weak. - HA is almost 100% ionized because H_2O is a stronger base (in the forward reaction) than is the conjugate base, A^-, (in the reverse reaction). - Hence, the ionization equilibrium position lies almost completely to the right because the base H_2O has a greater attraction for the H^+ ion that does the base A^-.	$HB(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + B^-(aq)$ Acid Base Conjugate acid Conjugate base - HB is a weak acid. Its conjugate base, B^-, is strong. - The conjugate base, B^- (in the reverse reaction) is stronger than the base H_2O (in the forward reaction) and manages to attract the H^+ ion. - The ionization equilibrium position lies almost far to the left because the conjugate base, B^-, has a greater attraction for the H^+ ion more than the base H_2O.
Hydrochloric acid, HCl, completely ionizes in water $HCl \rightarrow H^+ + Cl^-$	Acetic acid, CH_3COOH, partially ionizes in a single step $CH_3COOH \rightleftharpoons CH_3COO^- + H^+$
Nitric acid, HNO_3, completely ionizes in water $HNO_3 \rightarrow H^+ + NO_3^-$	Nitrous acid, HNO_2, partially ionizes in a single step $HNO_2 \rightleftharpoons NO_2^- + H^+$
Perchloric acid, $HClO_4$, completely ionizes in water $HClO_4 \rightarrow H^+ + ClO_4^-$	Carbonic acid, H_2CO_3, partially ionizes in two steps $H_2CO_3 \rightleftharpoons H^+ + HCO_3^-$ $HCO_3^- \rightleftharpoons H^+ + CO_3^{2-}$
	Phosphoric acid, H_3PO_4, partially ionizes in three steps $H_3PO_4 \rightleftharpoons H^+ + H_2PO_4^-$ $H_2PO_4^- \rightleftharpoons H^+ + HPO_4^{2-}$ $HPO_4^{2-} \rightleftharpoons H^+ + PO_4^{3-}$

CHM.5.3.04.006.01 Define acid ionization constant, K_a, while writing the ionization constant expression for different weak acids

Acid ionization constant is the value of the equilibrium constant expression for the ionization of a weak acid.

Generally, $$K_a = \frac{[\text{Conjugate base}][H_3O^+]}{[\text{Weak acid}]}$$

Example: Weak acid HA

$HA_{(aq)} \rightleftharpoons H^+_{(aq)} + A^-_{(aq)}$ or $HA_{(aq)} + H_2O_{(l)} \rightleftharpoons A^-_{(aq)} + H_3O^+_{(aq)}$

$$K_a = \frac{[H^+][A^-]}{[HA]} \quad \text{or} \quad K_a = \frac{[A^-][H_3O^+]}{[HA]}$$

CHM.5.3.04.003.02 Relate the strength of weak acids to the numerical values of K_a

The value of K_a indicates whether reactants or products are favored at equilibrium.
The weaker the acid the smaller is its K_a value.
For polyprotic acids, the K_a value decreases for each successive ionization.

Dissociation Constants of Weak Acids			
Weak Acid			**K_a (at 25°C)**
Hydrofluoric acid	HF	$HF \rightleftharpoons H^+ + F^-$	6.3×10^{-4}
Hydrocyanic acid	HCN	$HCN \rightleftharpoons H^+ + CN^-$	6.2×10^{-10}
Ethanoic (acetic) acid	CH_3COOH	$CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$	1.8×10^{-5}
Methanoic acid	$HCOOH$	$HCOOH \rightleftharpoons H^+ + HCOO^-$	1.8×10^{-4}
Benzoic acid	C_6H_5COOH	$C_6H_5COOH \rightleftharpoons H^+ + C_6H_5COO^-$	6.4×10^{-5}
Carbonic acid	H_2CO_3	$H_2CO_3 \rightleftharpoons H^+ + HCO_3^-$	4.3×10^{-7}
		$HCO_3^- \rightleftharpoons H^+ + CO_3^{2-}$	4.8×10^{-11}
Hydrosulfuric acid	H_2S	$H_2S \rightleftharpoons H^+ + HS^-$	8.9×10^{-8}
		$HS^- \rightleftharpoons H^+ + S^{2-}$	1.0×10^{-19}
Phosphoric acid	H_3PO_4	$H_3PO_4 \rightleftharpoons H^+ + H_2PO_4^-$	7.5×10^{-3}
		$H_2PO_4^- \rightleftharpoons H^+ + HPO_4^{2-}$	6.2×10^{-8}
		$HPO_4^{2-} \rightleftharpoons H^+ + PO_4^{3-}$	4.8×10^{-13}

CHM.5.3.04.003.03 Compare between strong and weak bases (using examples, particulate diagrams and ionization equations)

Strong Base	**Weak Base**
A strong base is a base that ionizes completely into metal ion and hydroxide ion. The complete ionization of the strong base is represented by ⟶.	A weak base is a base that partially ionizes in water. The partial ionization of the weak base is represented by ⇌.
It is a strong electrolyte	It is a weak electrolyte
Examples: $NaOH$, KOH, $LiOH$, $RbOH$, $CsOH$, $Ba(OH)_2$, $Ca(OH)_2$, $Sr(OH)_2$	Examples: NH_3, CH_3NH_2, $C_6H_5NH_2$
Sodium hydroxide, $NaOH$, completely ionizes in water $NaOH \rightarrow Na^+ + OH^-$	Ammonia, NH_3, partially ionizes in water $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$
Barium hydroxide, $Ba(OH)_2$, completely ionizes in water $Ba(OH)_2 \rightarrow Ba^{2+} + 2\ OH^-$	Methylamine, CH_3NH_2, partially ionizes in water $CH_3NH_2 + H_2O \rightleftharpoons CH_3NH_3^+ + OH^-$

CHM.5.3.04.003.04 Identify the relationship between the strength of an acid and its conjugate base and the strength of a base and its conjugate acid

The stronger an acid is, the weaker its conjugate base; the stronger a base is, the weaker its conjugate acid

CHM.5.3.04.006.02 Define base ionization constant, K_b, while writing the ionization constant expression of different weak bases

Base ionization constant is the value of the equilibrium constant expression for the ionization of a weak base

Generally, $\boldsymbol{K_b = \frac{[\text{Conjugate acid}][OH^-]}{[\text{Weak base}]}}$

Example: Weak base NH_3

$NH_{3(aq)} + H_2O_{(l)} \rightleftharpoons NH_4^+{}_{(aq)} + OH^-{}_{(aq)}$

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

CHM.5.3.04.003.05 Relate the strength of weak bases to the numerical values of K_b

The value of K_b indicates whether reactants or products are favored at equilibrium.
The weaker the base the smaller is its K_b value.

Dissociation Constants of Weak Bases			
Weak Base			**K_b (at 25°C)**
Ammonia	NH_3	$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$	1.8×10^{-5}
Methylamine	CH_3NH_2	$CH_3NH_2 + H_2O \rightleftharpoons CH_3NH_3^+ + OH^-$	4.4×10^{-4}
Ethylamine	$C_2H_5NH_2$	$C_2H_5NH_2 + H_2O \rightleftharpoons C_2H_5NH_3^+ + OH^-$	4.3×10^{-4}
Aniline	$C_6H_5NH_2$	$C_6H_5NH_2 + H_2O \rightleftharpoons C_6H_5NH_3^+ + OH^-$	4.3×10^{-10}

Relative Strengths of Common Acids and Bases		
Substance	**Formula**	**Relative Strength**
Hydrochloric acid	HCl	
Nitric acid	HNO_3	**Strong Acids**
Sulfuric acid	H_2SO_4	↑
Phosphoric acid	H_3PO_4	
Ethanoic (acetic) acid	CH_3COOH	**Increasing Strength of Acid**
Carbonic acid	H_2CO_3	
Hypochlorous acid	$HClO$	
		Neutral Solutions
Ammonia	NH_3	
Sodium Silicate	Na_2SiO_3	
Calcium hydroxide	$Ca(OH)_2$	
Sodium hydroxide	$NaOH$	**Increasing Strength of Base**
Potassium hydroxide	KOH	↓ **Strong bases**

CHM.5.3.04.005.01 Define the ion-product constant for water, K_w, while writing its expression and value at 25°C

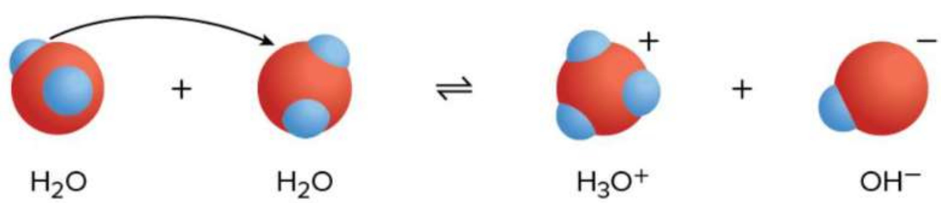
The ion-product constant for water is the value of the equilibrium constant expression for the self-ionization of water

$K_w = [H^+][OH^-]$

Where:
K_w is the ion product constant of water
$[H^+]$ represents the concentration of the hydrogen ion
$[OH^-]$ represents the concentration of the hydroxide ion

In the self-ionization of water, one molecule acts as an acid, and the other acts as a base.

$$H_2O + H_2O \rightleftharpoons H_3O^+ + OH^-$$



At 25°C or 298 K, $[H^+] = [OH^-] = 1.0 \times 10^{-7}$ *M.*

$K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$

Hence, at 298 K, the product of $[H^+]$ and $[OH^-]$ is always 1.0×10^{-14}
This means that if $[H^+]$ increases, the $[OH^-]$ deceases.

According to Le Chatelier's principle, adding extra H^+ ions to water at equilibrium is a stress on the system. The system reacts in a way to reduce the stress and the added H^+ ions react with OH^- ions to form more water molecules. Hence, the $[OH^-]$ decreases.

CHM.5.3.04.007.01 Use K_w to calculate the hydronium ion and hydroxide ion concentration at a given temperature and vice versa

<u>Example:</u>
At 298 K, $[H^+]$ in a cup of coffee is 1.0×10^{-5} *M.*

a) Calculate $[OH^-]$

$K_w = [H^+][OH^-]$

$[OH^-] = \frac{K_W}{[H^+]} = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-5}} = 1.0 \times 10^{-9}\ M$

b) Is the solution acidic, basic or neutral? Justify your answer.

Acidic because $[H^+] > [OH^-]$

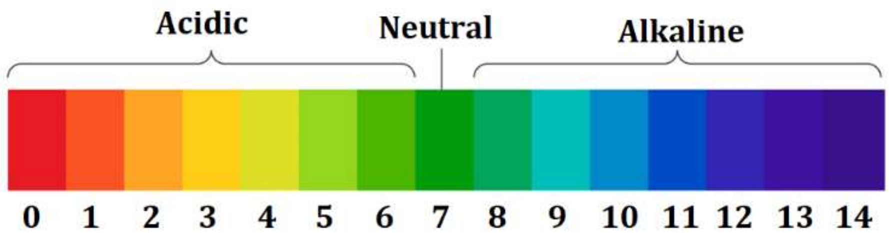
CHM.5.3.04.005.03 Define pH and write its mathematical formula

pH is the negative logarithm of the hydrogen ion concentration

$$pH = -\text{Log}\,[H^+]$$

It is a value (from 0 to 14) that is used to express the acidity or basicity of a system.

CHM.5.3.04.005.04 Know what the pH scale is



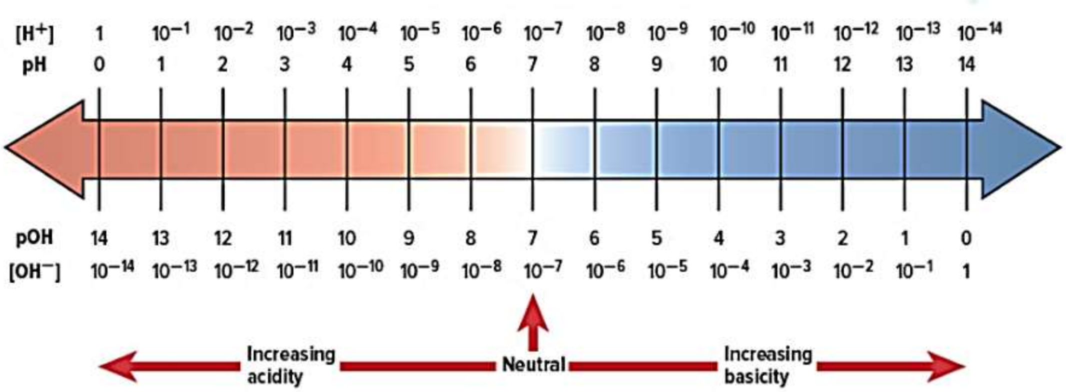
- A pH of 7 is neutral, a pH of less than 7 is acidic, and pH of greater than 7 is basic.
- As the concentration of hydrogen ion increases, the solution becomes more acidic and the pH decreases.
- As the concentration of hydrogen ion decreases, the solution becomes more basic and the pH increases.
- Each whole number on the scale indicates a **tenfold change** in acidity.

  A solution with a pH of 3.0 has ten times the hydrogen ion concentration of a solution with a pH of 4.0.

CHM.5.3.04.005.05 Define pOH and write its mathematical formula

pOH is the negative logarithm of the hydroxide ion concentration

$$pOH = -\text{Log}\,[OH^-]$$



CHM.5.3.04.007.02 Describe the relation between pH and pOH and perform calculations involving this relation

pH + pOH = 14.00 (at 25°C or 298 K)

CHM.5.3.04.006.03 Relate the acidity and basicity of an aqueous solution to the hydronium and hydroxide ion concentration and pH at 25°C or 298 K

Acidic Solution	pH < 7	$[H^+] > [OH^-]$
Basic Solution	pH > 7	$[H^+] < [OH^-]$
Neutral Solution	pH = 7	$[H^+] = [OH^-]$

CHM.5.3.04.007.03 Calculate pH of a solution when the $[H^+]$ or $[OH^-]$ is given and vice versa

Example:
Calculate the pH of a solution with $[H^+] = 0.0055\ M$

$pH = -\text{Log}\ [H^+] = -\text{Log}\ (0.0055) = 2.26$

Example:
Calculate the pH of a solution where $[OH^-] = 8.2 \times 10^{-6}\ M$

$pOH = -\text{Log}\ [OH^-] = -\text{Log}\ (8.2 \times 10^{-6}) = 5.09$

$pH = 14.00 - pOH = 14.00 - 5.09 = 8.91$

Note
To report your answer correctly, the number of significant figures in the concentration equals the number of decimal places in the log function.

CHM.5.3.04.007.04 Calculate pOH when the $[H^+]$ or $[OH^-]$ is given and vice versa

Example:
Calculate the pOH of a solution with $[H^+] = 0.25\ M$

$[OH^-] = \frac{K_W}{[H^+]} = \frac{1.0 \times 10^{-14}}{0.25} = 4.0 \times 10^{-14}\ M$

$pOH = -\text{Log}\ [OH^-] = -\text{Log}\ (4.0 \times 10^{-14}) = 13.40$

(If the pH is to be calculated: $pH = 14 - pOH = 14 - 13.40 = 0.60$)

CHM.5.3.04.007.05 Calculate the pH and pOH from $[OH^-]$

Example:
Calculate the pH and pOH of a solution with $[OH^-] = 4.0 \times 10^{-3}\ M$

$pOH = -\text{Log}\ [OH^-] = -\text{Log}\ (4.0 \times 10^{-3}) = 2.40$

$pH = 14 - pOH = 14.00 - 2.40 = 11.60$

CHM.5.3.04.007.06 Calculate $[H^+]$ and $[OH^-]$ from pH

Example

Calculate $[H^+]$ and $[OH^-]$ for a solution that has a pH of 7.40

$[H^+]$ = antilog (−pH)
$[H^+]$ = antilog (−7.40) = 4.0×10^{-8} M

pOH = 14.00 – pH = 6.60
$[OH^-]$ = antilog (−pOH)
$[OH^-]$ = antilog (−6.60) = 2.5×10^{-7} M

CHM.5.3.04.007.07 Calculate the pH of a strong acid given its concentration

Example:

Calculate the pH of a 0.10 M solution of the strong acid, HCl.

$HCl \rightarrow H^+ + Cl^-$
$[H^+] = [HCl] = 0.10\ M$

Hint: For all strong monoprotic acids, the concentration of the acid is the concentration of H^+ ions.

pH = − Log $[H^+]$ = − Log (0.10) = 1.00

CHM.5.3.04.007.08 Calculate the pH of a strong base given its concentration

Example:

Calculate the pH of a 0.10 M solution of the strong base, NaOH.

$NaOH \rightarrow Na^+ + OH^-$
$[OH^-] = [NaOH] = 0.10\ M$

pOH = − Log $[OH^-]$ = − Log (0.10) = 1.00
pH = 14.00 – 1.00 = 13.00

Example:

Calculate the pH a solution of barium hydroxide, $Ba(OH)_2$, with a concentration of 2.5×10^{-3} M.

$Ba(OH)_2 \rightarrow Ba^{2+} + 2\ OH^-$
1 : 1 : 2

$[OH^-] = 2 \times 2.5 \times 10^{-3} = 5.0 \times 10^{-3}\ M$

$[H^+] = \frac{K_W}{[OH^-]} = \frac{1.00 \times 10^{-14}}{5.0 \times 10^{-3}} = 2.0 \times 10^{-12}\ M$

pH = — Log $[H_3O^+]$

$pH = -\text{Log}\ (2.0 \times 10^{-12}) = 11.70$

CHM.5.3.04.006.04 Calculate the acid dissociation constant, K_a, given acid concentration, $[H^+]$ and pH

Example:
Calculate the K_a of a 0.100 *M* solution of ethanoic acid, CH_3COOH, with $[H^+] = 1.34 \times 10^{-3}$ *M*.

$CH_3COOH \rightleftharpoons CH_3COO^- + H^+$

$[H^+] = [CH_3COO^-] = 1.34 \times 10^{-3}$ *M*

$[CH_3COOH] = 0.100 - 1.34 \times 10^{-3} = 0.0987$ *M*

Concentration	$[CH_3COOH]$	$[H^+]$	$[CH_3COO^-]$
Initial	0.100	0	0
Change	-1.34×10^{-3}	$+ 1.34 \times 10^{-3}$	$+ 1.34 \times 10^{-3}$
Equilibrium	0.0987	1.34×10^{-3}	1.34×10^{-3}

$$K_a = \frac{[H^+][CH_3COO^-]}{[CH_3COOH]} = \frac{(1.34 \times 10^{-3})(1.34 \times 10^{-3})}{0.0987} = 1.82 \times 10^{-5}$$

Example:
The pH of 0.100 M solution of formic acid, HCOOH, is 2.38. Calculate K_a of HCOOH.

$[H^+] = \text{antilog}(-\text{pH}) = 4.2 \times 10^{-3}$ *M*

$HCOOH \rightleftharpoons H^+ + HCOO^-$

$[HCOO^-] = [H^+] = 4.2 \times 10^{-3}$ *M*

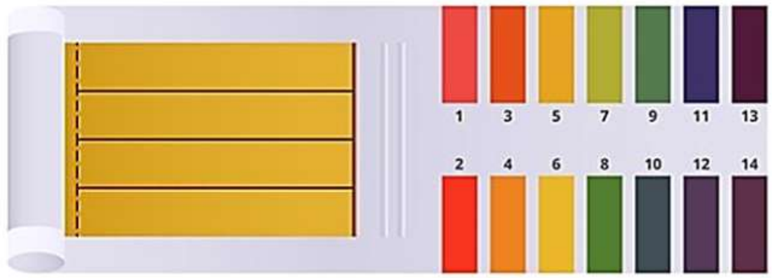
$[HCOOH] = 0.100 - 4.2 \times 10^{-3} = 0.096$ *M*

Concentration	$[HCOOH]$	$[H^+]$	$[HCOO^-]$
Initial	0.100	0	0
Change	-4.2×10^{-3}	$+4.2 \times 10^{-3}$	$+4.2 \times 10^{-3}$
Equilibrium	0.096	4.2×10^{-3}	4.2×10^{-3}

$$K_a = \frac{[H^+][HCOO^-]}{[HCOOH]} = \frac{(4.2 \times 10^{-3})(4.2 \times 10^{-3})}{0.096} = 1.8 \times 10^{-4}$$

CHM.5.3.04.005.07 List different methods used to measure the pH of a solution

1) Indicator Paper
   The strip of pH paper is inserted into an acidic or basic solution. The color of the paper changes.
   The new color is compared with standard pH colors on a chart.



2) pH Meter
   It is a device that digitally measures the pH of a solution.
   When the electrodes are placed in a solution, the meter gives a direct analog or digital readout of the pH value of the solution.

CHM.5.3.04.007.09 Perform an experiment to investigate the pH of different solutions

Inspire Chemistry – Module 17 – Lesson 4: Neutralization Reactions

CHM.5.3.04.004.01 Define neutralization reaction while writing the neutralization equation (Complete ionic and net ionic equations)

It is a reaction in which an acid and a base in an aqueous solution react to produce a salt and water

Neutralization is a double replacement reaction

Example: $HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$

Complete ionic equation: $H^+_{(aq)} + Cl^-_{(aq)} + Na^+_{(aq)} + OH^-_{(aq)} \rightarrow Na^+_{(aq)} + Cl^-_{(aq)} + H_2O_{(l)}$

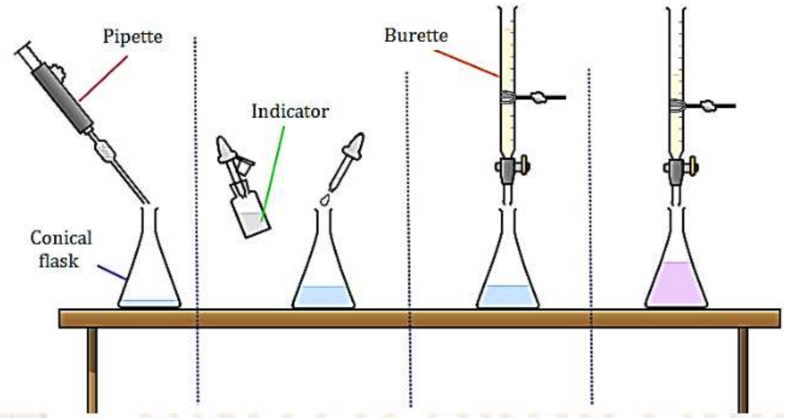
Net ionic equation: $H^+_{(aq)} + OH^-_{(aq)} \rightleftharpoons H_2O_{(l)}$ or $H_3O^+_{(aq)} + OH^-_{(aq)} \rightleftharpoons 2\ H_2O_{(l)}$

CHM.5.3.04.004.02 Define titration, and titrant

Titration is a method of determining the concentration of a solution by reacting a known volume of that solution with a solution of known concentration.

Titrant is the solution of known concentration that is added to the burette.

CHM.5.3.04.004.03 Explain how to carry out an acid-base titration



1) A measured volume of an acidic or basic solution of unknown concentration is placed in a beaker (or conical flask).
2) Few drops of a suitable indicator are added to the solution (or the electrodes of a pH meter are immersed in this solution, and initial solution of the pH is read and recorded)
3) A burette is filled with the titrating solution of known concentration (Titrant)
4) Measured volumes of the titrant is added into the solution in the beaker (or conical flask)
   The process continues till the indicator changes color (end point is reached)
   Or till the equivalence point is reached (as recorded by the pH meter)

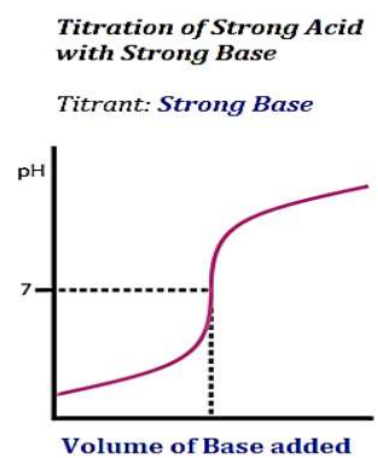
CHM.5.3.04.004.04 Explain the difference between the equivalence point and the end point of titration process

Equivalence point is the point at which moles of hydrogen ions from the acid equal the moles of hydroxide ions from the base

End point is the point at which the indicator that is used in titration changes color

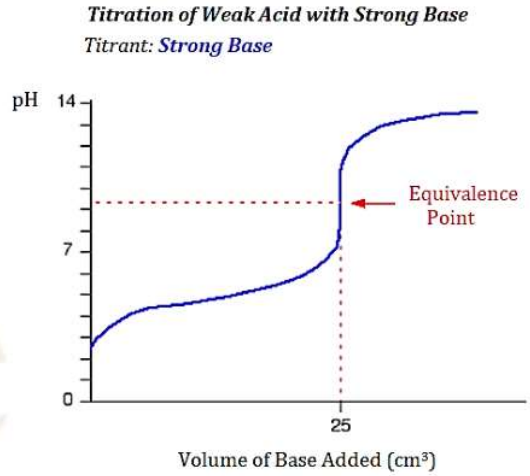
CHM.5.3.04.009.01 Describe the titration curve of a strong acid with a strong base with respect to type of salt formed, pH and nature of solution at equivalence point, indicator used and its color change and volume of titrant needed for changing color of indicator

- Salt formed is neutral
- pH at equivalence point = 7
- Indicator used is phenolphthalein and color changes from colorless to pink



CHM.5.3.04.009.02 Describe the titration curve of a weak acid with a strong base with respect to type of salt formed, pH and nature of solution at equivalence point, indicator used and its color change and volume of titrant needed for changing color of indicator

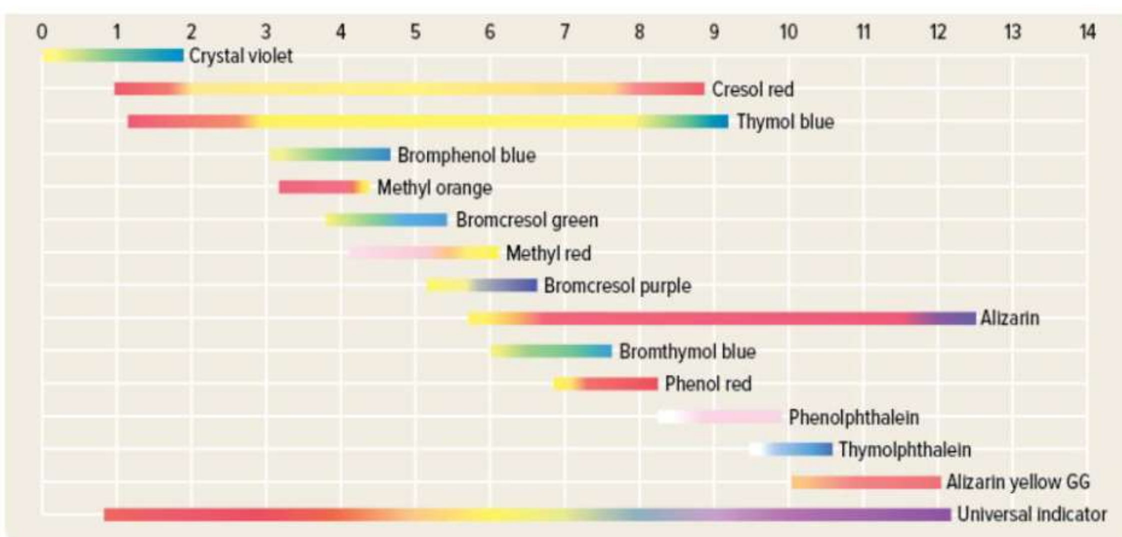
- Salt formed is basic
- pH at equivalence point > 7
- Indicators used are
  - phenolphthalein; color changes from colorless to pink
  - phenol red; color changes from yellow to pink



CHM.5.3.04.004.05 Define acid-base indicator and its function

Acid-base indicator is a compound whose color is sensitive to pH
The color of the indicator changes as the pH of solution changes

Choosing the right indicator for titration is important. The indicator must change color at the equivalence point of the titration (which is not always 7)



CHM.5.3.04.009.03 Perform a titration experiment

CHM.5.3.04.004.06 Calculate the molarity (concentration) and volume of a solution using titration data ($C_aV_a = C_bV_b$)

Example:

A student titrates 50. mL HCl with a 0.45 *M* solution of NaOH. The volume of NaOH needed to reach the equivalence point is 20. mL. Calculate the concentration of HCl.

$NaOH_{(aq)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$

At the equivalent point, number of moles of acid = number of moles of base

$C_aV_a=C_bV_b$

$C_{HCl} \times 50.= 0.45 \times 20.$ $C_{HCl} = 0.18\ M$

Example:

A student titrates 50. mL H_2SO_4 with a 0.55 *M* solution of KOH. The volume of KOH needed to reach the equivalence point is 40. mL. Calculate the concentration of H_2SO_4.

$2KOH_{(aq)} + H_2SO_{4(aq)} \rightarrow K_2SO_{4(aq)} + 2H_2O_{(l)}$

At the equivalent point, number of moles of acid = number of moles of base

$2C_aV_a=C_bV_b$

$2C_a \times 50.= 0.55 \times 40.$ $C_a = 0.22\ M$

Example:

A 25.0 mL of sodium hydroxide solution, NaOH, is titrated with 0.150 *M* of hydrochloric acid solution, HCl, solution. If the concentration of NaOH is 0.0600 *M*, what is the volume of HCl needed at the equivalence point?

At the equivalent point, number of moles of acid = number of moles of base

$C_aV_a = C_bV_b$

$0.150 \times V_{HCl} = 0.0600 \times 25.0$ $V_{HCl} =10.0$ mL

Different Acid-Base Reactions

- **Case 1**

$HCl\ (aq) + NaOH\ (aq) \rightarrow NaCl\ (aq) + H_2O\ (l)$

$1 : 1$

$n_a \quad n_b$

$$\boxed{\begin{aligned} n_a &= n_b \\ C_aV_a &= C_b V_b \end{aligned}}$$

- **Case 2**

$H_2SO_4\ (aq) + 2\ NaOH\ (aq) \rightarrow Na_2SO_4\ (aq) + 2\ H_2O(l)$

$1 : 2$

$n_a \quad n_b$

$$\boxed{\begin{aligned} 2\ n_a &= n_b \\ 2C_aV_a &= C_bV_b \end{aligned}}$$

- **Case 3**

$H_3PO_4\ (aq) + 3\ NaOH\ (aq) \rightarrow Na_3PO_4\ (aq) + 3\ H_2O(l)$

$1 : 3$

$n_a \quad n_b$

$$\boxed{\begin{aligned} 3\ n_a &= n_b \\ 3\ C_aV_a &= C_bV_b \end{aligned}}$$

- **Case 4**

$2\ HCl(aq) + Ca(OH)_2(aq) \rightarrow CaCl_2(aq) + 2H_2O(l)$

$2 : 1$

$n_a \quad n_b$

$$\boxed{\begin{aligned} n_a &= 2\ n_b \\ C_aV_a &= 2\ C_b V_b \end{aligned}}$$

- **Case 5**

$H_2SO_4(aq) + Ca(OH)_2(aq) \rightarrow CaSO_4(s) + 2H_2O\ (l)$

$1 : 1$

$n_a \quad n_b$

$$\boxed{\begin{aligned} n_a &= n_b \\ C_aV_a &= C_b V_b \end{aligned}}$$

- **Steps to solve these type of questions**

1. **Write balanced equation.**
2. **Write relation between acid and base in terms of moles.**
3. **Write given and required to find.**
4. **Find the unknown.**

CHM.5.3.04.022.01 Define salt and salt hydrolysis
CHM.5.3.04.022.02 Identify the type of salt (acidic, basic or neutral) and its constituent acid and base with their strengths

- Salt is an ionic compound made up of a cation from a base and an anion from an acid.
- Salt hydrolysis is the process in which anions of the dissociated salt accept hydrogen ions from water or the cations of the dissociated salt donate hydrogen ions to water.
  (i.e The weak species of the dissociated salt accepts or donates a hydrogen ion)
- Neutral salt results from reaction of strong acid and strong base.
- Acidic salt results from reaction of strong acid and weak base.
- Basic salt results from reaction of weak acid and strong base.

Strong acid + **Strong base** → Neutral Solution

Strong acid + **Weak base** → **Acidic** Solution

Weak acid + **Strong base** → **Basic** Solution

Examples:

1) Sodium fluoride, NaF, is a basic salt

$NaF_{(s)} \xrightarrow{H_2O} Na^+_{(aq)} + F^-_{(aq)}$

NaF is a salt of a strong base, KOH, and a weak acid, HF.

$F^-_{(aq)} + H_2O_{(l)} \rightleftharpoons HF_{(aq)} + OH^-_{(aq)}$

The OH^- makes the solution basic.

2) Ammonium chloride, NH_4Cl, is an acidic salt

$NH_4Cl_{(s)} \xrightarrow{H_2O} NH^+_{4(aq)} + Cl^-_{(aq)}$

NH_4Cl is a salt of a weak base, NH_3, and a strong acid, HCl.

$NH^+_{4(aq)} + H_2O_{(l)} \rightleftharpoons NH_{3(aq)} + H_3O^+_{(aq)}$

The H_3O^+ makes the solution acidic.

3) Sodium nitrate, $NaNO_3$, is a neutral salt

$NaNO_{3(s)} \xrightarrow{H_2O} Na^+_{(aq)} + NO^-_{3(aq)}$

$NaNO_3$ is a salt of a strong base, NaOH, and a strong acid, HNO_3.

No salt hydrolysis occurs because neither Na^+ nor NO_3^- react with water. Hence, it is a neutral solution.

Unit 2: Oxidation & Reduction Reactions

2.1 – Redox Reaction

Inspire Chemistry – Module 18 – Lesson 1: Oxidation & Reduction

CHM.5.3.05.001.01 Distinguish between oxidation and reduction in terms of loss and gain of electrons, oxygen and hydrogen

Oxidation is the loss of electrons, gain of oxygen or loss of hydrogen

Reduction is the gain of electrons, loss of oxygen or gain of hydrogen

CHM.5.3.05.001.02 Define oxidation number

Oxidation number is a number assigned to an atom or ion to indicate the degree of oxidation or reduction

CHM.5.3.05.001.03 Assign oxidation number to atoms, ions and compounds according to a set of rules

- Rules for assigning oxidation number:

- The oxidation state of an element is zero, including all elemental forms of the elements (Ex. N_2, P_4, S_8 and O_3)
- The oxidation state of a monoatomic ion is the same as its charge
- In compounds, fluorine is always assigned an oxidation state of −1
- Oxygen is usually assigned an oxidation state of −2 in its covalent compounds

  Exception: In peroxides, compounds containing the group O_2^{2-}, each oxygen is assigned an oxidation state of −1 (Ex H_2O_2)
- Hydrogen is assigned an oxidation state of +1

  Exception: Metal hydrides, where H has an oxidation state of −1
- The sum of oxidation states must be zero for an electrically neutral compound
- For a polyatomic ion, the sum of the oxidation states must be equal to the charge of the ion

Example:

Determine the **oxidation number** of the underlined element in each of the following species.

Species	**Oxidation Number**
P in $\underline{\mathbf{P}}O_4^{3-}$	+5
Mn in $K\underline{\mathbf{Mn}}O_4$	+7
O in $H_2\underline{\mathbf{O}}_2$	−1
S in $\underline{\mathbf{S}}^{2-}$	−2
H in $\underline{\mathbf{H}}NO_3$	+1
Fe in $\underline{\mathbf{Fe}}(NO_3)_2$	+2
Cl in $\underline{\mathbf{Cl}}_2$	0
Cl in $Na\underline{\mathbf{Cl}}O_4$	+7
P in $Al\underline{\mathbf{P}}O_4$	+5
N in $H\underline{\mathbf{N}}O_2$	+3
N in $\underline{\mathbf{N}}H_4^+$	− 3
As in $\underline{\mathbf{As}}O_4^{3-}$	+ 5

Cr in **<u>Cr</u>**O_4^{2-}	+ 6
N in **<u>N</u>**H_3	− 3
N in **<u>N</u>**$_2H_4$	− 2

CHM.5.3.05.001.04 Distinguish between oxidation and reduction in terms of change in oxidation number

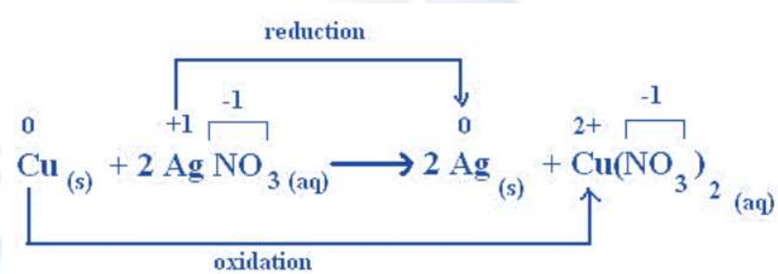
<u>Oxidation</u> involves increase in oxidation number while <u>reduction</u> involves decrease in oxidation number

Process Leading to Oxidation and Reduction	
Oxidation	**Reduction**
Complete loss of electrons (ionic reactions)	Complete gain of electrons (ionic reactions)
Shift of electrons away from an atom in a covalent bond	Shift of electrons towards an atom in a covalent bond
Gain of oxygen	Loss of oxygen
Loss of hydrogen by a covalent compound	Gain of hydrogen by a covalent compound
Increase in oxidation number	Decrease in oxidation number

CHM.5.3.05.001.05 Define redox reaction while explaining what must be conserved in a redox reaction

Redox reaction is a reaction in which oxidation and reduction occur simultaneously
Charge and mass are conserved in a redox reaction

<u>Example:</u>



CHM.5.3.05.001.06 Identify a redox reaction from a given list of reactions while indicating the oxidized and reduced species

The species oxidized has its oxidation number increasing

The species reduced has its oxidation number decreasing

Example:

For each of the following equations identify the species oxidized and the species reduced.

	Species Oxidized	**Species Reduced**
$2\ Br^- + Cl_2 \rightarrow Br_2 + 2\ Cl^-$	Br^-	Cl_2
$2\ Ce + 3\ Cu^{2+} \rightarrow 3\ Cu + 2\ Ce^{3+}$	Ce	Cu^{2+}
$2\ Zn + O_2 \rightarrow 2\ ZnO$	Zn	O_2
$2\ Na + 2\ H^+ \rightarrow 2\ Na^+ + H_2$	Na	H^+
$Zn + 2\ HCl \rightarrow ZnCl_2 + H_2$	Zn	H in HCl

CHM.5.3.05.001.07 Define oxidizing agent and reducing agent in a redox reaction

Oxidizing agent: A substance that has the ability to oxidize another substance where itself undergoes reduction

Reducing agent: A substance that has the ability to reduce another substance where itself undergoes oxidation

CHM.5.3.05.001.08 Identify oxidizing agent and reducing agent in a redox reaction

Example:

$$Cu_{(s)} + 2\ AgNO_{3(aq)} \rightarrow 2\ Ag_{(s)} + Cu(NO_3)_{2(aq)}$$

$\mathbf{Cu \rightarrow Cu^{2+}}$	$\mathbf{Ag^+ \rightarrow Ag}$
- The oxidation number of Cu changes from 0 to +2 - The oxidation number of Cu increases - Cu is oxidized - Cu is the reducing agent	- The oxidation number of Ag changes from +1 to 0 - The oxidation number of Ag decreases - Ag is reduced - $AgNO_3$ or Ag^+ is the oxidizing agent

Example:

For each of the following equations identify the oxidizing agent and reducing agent.

	Oxidizing agent	**Reducing agent**
$Fe_{(s)} + 2\ Ag^+_{(aq)} \rightarrow Fe^{2+}_{(aq)} + 2\ Ag_{(s)}$	Ag^+	Fe
$Mg + I_2 \rightarrow MgI_2$	I_2	Mg

$N_2 + 3\ H_2 \rightarrow 2\ NH_3$	N_2	H_2

CHM.5.3.05.001.09 Define half-reaction, oxidation half-reaction and reduction-half reaction

A half-reaction is one of two parts of a redox reaction

Oxidation half-reaction is the half-reaction that shows the number of electrons lost when a species is oxidized

Reduction half-reaction is the half-reaction that shows the number of electrons gained when a species is reduced

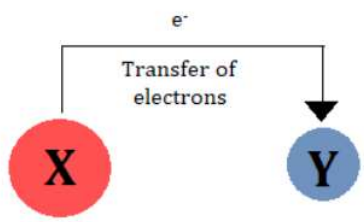
CHM.5.3.05.001.10 Write oxidation-half reaction and reduction-half reaction for a redox reaction

Example:
For the reaction: $\mathbf{Al_{(s)} + HCl_{(aq)} \rightarrow AlCl_{3(aq)} + H_{2(g)}}$, write the oxidation-half reaction and reduction-half reaction for.

Oxidation-half reaction: $Al \rightarrow Al^{3+} + 3\ e^-$

Reduction-half reaction: $2H^+ + 2e^- \rightarrow H_2$

Summary of Redox Reactions



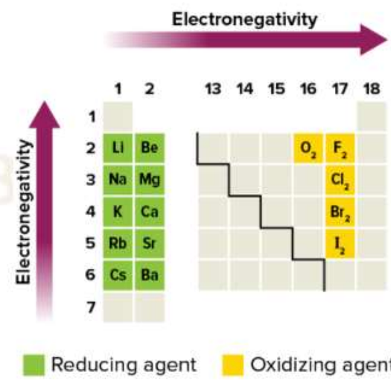
What happens to X?	**What happens to Y?**
$X \rightarrow X^+ + e^-$ - X loses an electron - The oxidation number of X increases - X is oxidized - X is the reducing agent	$Y + e^- \rightarrow Y^-$ - Y gains an electron - The oxidation number of Y decreases - Y is reduced - Y is the oxidizing agent
In general, during Oxidation: - A reactant loses an electron - Oxidation number increases - Reducing agent is oxidized	**In general, during Reduction:** - A reactant gains an electron - Oxidation number decreases - Oxidizing agent is reduced

CHM.5.3.05.003.01 Describe the role of electronegativity in redox reactions

- Electronegativity is the relative ability of an atom to attract electrons in a chemical bond.
- Electronegativity generally increases from left to right across a period and decreases down a group.

- Elements with low electronegativity (Groups 1 and 2) are strong reducing agents.
- Elements with high electronegativity (Groups 17 and Oxygen in Group 16) are strong oxidizing agents.

- In a redox reaction with molecular compounds or polyatomic ions with covalent bonds:
  The less electronegative element is treated as if it has been oxidized by losing electrons to the other element.
  The more electronegative element is treated as if it has been reduced by gaining electrons from the other element.



<u>Example:</u>
$N_{2(g)} + 3H_{2(g)} \rightarrow 2\ NH_{3(g)}$

Nitrogen, N, is more electronegative than hydrogen, H.
Nitrogen, N, is to be reduced, while hydrogen, H, is to be oxidized.

Inspire Chemistry – Module 18 – Lesson 2: Balancing Redox Reactions

CHM.5.3.05.002.01 Describe the steps for balancing redox reactions, in acidic medium, by the half-reaction method

- <u>Steps (In Acidic Medium)</u>

1) Write the unbalanced equation in ionic form
2) Write the two half reactions half-reactions (oxidation and reduction)
3) Balance any element other than oxygen and hydrogen
4) Balance oxygen by adding water
5) Balance hydrogen by adding hydrogen ion (H^+)
6) Balance the charge by adding suitable number of electrons
7) Balance the number of electrons for the two reactions (Number of electrons lost and gained must be equal)
8) Add up the two equations and write the final balanced equation
   Check that charge and atoms are conserved

CHM.5.3.05.002.03 Balance redox reaction, in acidic medium, using half-reaction method

<u>Example:</u>
When copper metal is placed in a dilute solution of nitric acid, bubbles of NO gas are produced. The solution turns blue, indicating that Cu^{2+} is forming. Write a balanced equation, showing all the steps.

Steps:

Step	Equation
1	$Cu_{(s)} + H^+_{(aq)} + NO^-_{3\,(aq)} \rightarrow Cu^{2+}_{(aq)} + NO_{(g)}$
2	$Cu_{(s)} \rightarrow Cu^{2+}_{(aq)}$ $NO^-_{3\,(aq)} \rightarrow NO_{(g)}$
3	$Cu_{(s)} \rightarrow Cu^{2+}_{(aq)}$ $NO^-_{3\,(aq)} \rightarrow NO_{(g)}$ (Elements Cu & N are balanced on the sides of the reaction)
4	$Cu_{(s)} \rightarrow Cu^{2+}_{(aq)}$ $NO^-_{3\,(aq)} \rightarrow NO_{(g)} + \mathbf{2H_2O_{(l)}}$
5	$Cu_{(s)} \rightarrow Cu^{2+}_{(aq)}$ $\mathbf{4H^+_{(aq)}} + NO^-_{3\,(aq)} \rightarrow NO_{(g)} + \mathbf{2H_2O_{(l)}}$
6	$Cu_{(s)} \rightarrow Cu^{2+}_{(aq)} + \mathbf{2e^-}$ $\mathbf{3e^-} + \mathbf{4H^+_{(aq)}} + NO^-_{3\,(aq)} \rightarrow NO_{(g)} + \mathbf{2H_2O_{(l)}}$
7	$\mathbf{3} \times \left(Cu_{(s)} \rightarrow Cu^{2+}_{(aq)} + \mathbf{2e^-}\right)$ $\mathbf{2} \times \left(\mathbf{3e^-} + \mathbf{4H^+_{(aq)}} + NO^-_{3\,(aq)} \rightarrow NO_{(g)} + \mathbf{2H_2O_{(l)}}\right)$ $\mathbf{3}Cu_{(s)} \rightarrow \mathbf{3}Cu^{2+}_{(aq)} + \mathbf{6e^-}$ $\mathbf{6e^-} + \mathbf{8H^+_{(aq)}} + \mathbf{2}NO^-_{3\,(aq)} \rightarrow \mathbf{2}NO_{(g)} + \mathbf{4H_2O_{(l)}}$
8	$\mathbf{3}Cu_{(s)} \rightarrow \mathbf{3}Cu^{2+}_{(aq)} + \cancel{\mathbf{6e^-}}$ $\cancel{\mathbf{6e^-}} + \mathbf{8H^+_{(aq)}} + \mathbf{2}NO^-_{3\,(aq)} \rightarrow \mathbf{2}NO_{(g)} + \mathbf{4H_2O_{(l)}}$ $\mathbf{3}Cu_{(s)} + \mathbf{2}NO^-_{3\,(aq)} + \mathbf{8H^+_{(aq)}} \rightarrow \mathbf{3}Cu^{2+}_{(aq)}\mathbf{2}NO_{(g)} + \mathbf{4H_2O_{(l)}}$

CHM.5.3.05.002.04 Describe the steps for balancing redox reactions, in basic medium, by the half-reaction method

- Steps (In Basic Medium)

1) Write the unbalanced equation in ionic form
2) Write the two half reactions half-reactions (oxidation and reduction)
3) Balance any element other than oxygen and hydrogen
4) Balance oxygen by adding water
5) Balance hydrogen by adding hydrogen ion (H^+)
6) Balance the charge by adding suitable number of electrons
7) Balance the number of electrons for the two reactions (Number of electrons lost and gained must be equal)
8) Add up the two equations
9) **Add to the right and left sides number of hydroxide ions OH^- equal to the H^+, and eliminate water molecules**
10) Add up the two equations and write the final balanced equation
    Check that charge and atoms are conserved

CHM.5.3.05.002.05 Balance redox reaction in basic medium using half-reaction method

Example:

Complete and balance the following equation which takes place in basic aqueous solution

$$P_{4(s)} + H_2O_{(l)} \rightarrow H_2PO^-_{3\,(aq)} + P_2H_{4(g)}$$

Steps:

Step	Equation
1	$P_{4(s)} \rightarrow H_2PO_3^-{}_{(aq)}$ $\quad\quad$ $P_{4(s)} \rightarrow P_2H_{4(g)}$
2	$P_{4(s)} \rightarrow 4H_2PO_3^-{}_{(aq)}$ $\quad\quad$ $P_{4(s)} \rightarrow 2P_2H_{4(g)}$
3	$12H_2O_{(l)} + P_{4(s)} \rightarrow 4H_2PO_3^-{}_{(aq)}$ $\quad\quad$ $P_{4(s)} \rightarrow 2P_2H_{4(g)}$
4	$12H_2O_{(l)} + P_{4(s)} \rightarrow 4H_2PO_3^-{}_{(aq)} + 16H^+_{(aq)}$ $\quad\quad$ $8H^+_{(aq)} + P_{4(s)} \rightarrow 2P_2H_{4(g)}$
5	$12H_2O_{(l)} + P_{4(s)} \rightarrow 4H_2PO_3^-{}_{(aq)} + 16H^+_{(aq)} + 12e^-$ $8e^- + 8H^+_{(aq)} + P_{4(s)} \rightarrow 2P_2H_{4(g)}$
6	$2 \times \left(12H_2O_{(l)} + P_{4(s)} \rightarrow 4H_2PO_3^-{}_{(aq)} + 16H^+_{(aq)} + 12e^-\right)$ $3 \times \left(8e^- + 8H^+_{(aq)} + P_{4(s)} \rightarrow 2P_2H_{4(g)}\right)$ $24H_2O_{(l)} + 2P_{4(s)} \rightarrow 8H_2PO_3^-{}_{(aq)} + 32H^+_{(aq)} + 24e^-$ $24e^- + 24H^+_{(aq)} + 3P_{4(s)} \rightarrow 6P_2H_{4(g)}$
7	$24H_2O_{(l)} + 2P_{4(s)} \rightarrow 8H_2PO_3^-{}_{(aq)} + \overset{\boxed{8H^+_{(aq)}}}{\cancel{32H}^+_{(aq)}} + \cancel{24e^-}$ $\cancel{24e^-} + \cancel{24H}^+_{(aq)} + 3P_{4(s)} \rightarrow 6P_2H_{4(g)}$ $5P_{4(s)} + 24H_2O_{(l)} \rightarrow 6P_2H_{4(g)} + 8H_2PO_3^-{}_{(aq)} + 8H^+_{(aq)}$
8	$8OH^-_{(aq)} + 5P_{4(s)} + \underset{16H_2O_{(l)}}{\cancel{24H_2O_{(l)}}} \rightarrow 6P_2H_{4(g)} + 8H_2PO_3^-{}_{(aq)} + \underset{\cancel{8H_2O_{(l)}}}{\boxed{8H^+_{(aq)} + 8OH^-_{(aq)}}}$ The final equation is: $5P_{4(s)} + 16H_2O_{(l)} + 8OH^-_{(aq)} \rightarrow 6P_2H_{4(g)} + 8H_2PO_3^-{}_{(aq)}$

2.1 – Electrochemistry

Inspire Chemistry – Module 19 – Lesson 1: Voltaic Cells

CHM.5.3.05.004.01 Define electrochemistry

Electrochemistry is the branch of chemistry that deals with electricity related application of oxidation-reduction reactions

It is the study of the interchange of electrical energy and chemical energy

CHM.5.3.05.007.01 Describe an electrochemical cell while specifying its types

An electrochemical cell is an apparatus that uses a redox reaction to produce electrical energy or used electrical energy to cause a chemical reaction

Types of electrochemical cells

- **Voltaic cell (Galvanic cell)** is a type of electrochemical cell that converts chemical energy to electrical energy by a spontaneous redox reaction

- **Electrolytic cell** is a type of electrochemical cell that converts electrical energy to chemical energy by a non-spontaneous redox reaction. It is an electrochemical cell where electrolysis occurs

CHM.5.3.05.007.02 Identify components of a voltaic or galvanic cell (anode, cathode, salt bridge or porous barrier, wires, electrolyte compartments); while explaining the role of each component, when does the reaction start and determining the direction of electron and current flow

- Anode: electrode where oxidation occurs

- Cathode: electrode where reduction occurs

- Salt bridge: It is a pathway to maintain solution neutrality by allowing the passage of ions from one side to another, where anions migrate towards anode and cations migrate towards cathode. It usually contains a conducting neutral soluble solution as KCl, NaCl or $NaNO_3$

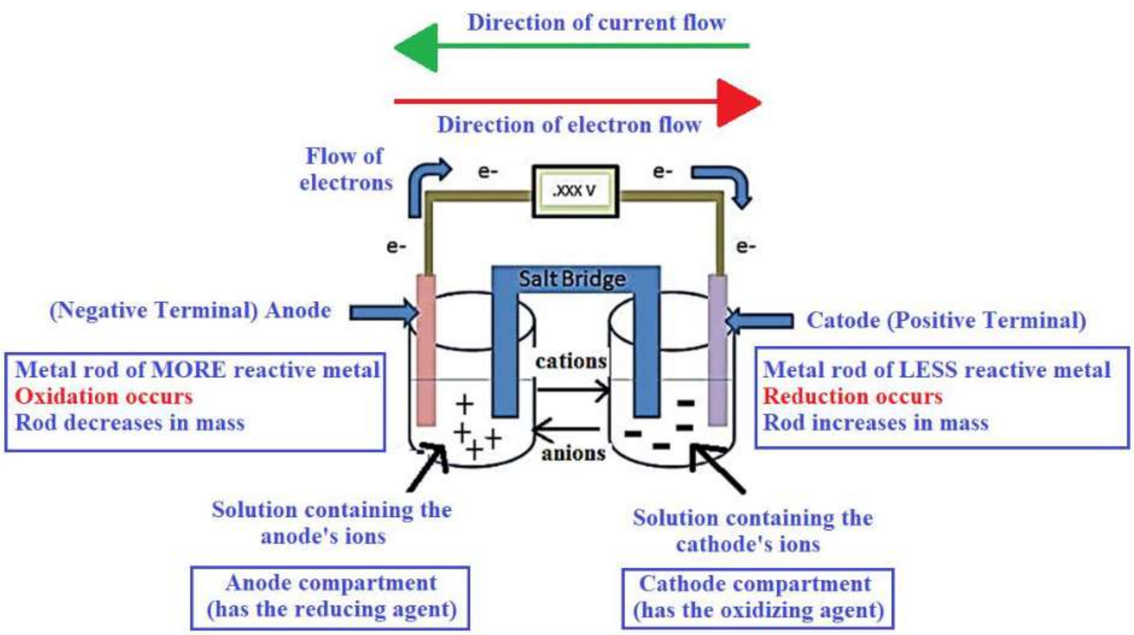
- The spontaneous reaction starts when the connecting metal wire and salt bridge are in place

- Electrons flow through the wire from the oxidation-half reaction (anode) to the reduction-half reaction (cathode) while positive and negative ions move through the salt bridge

- Current flows from cathode to anode

- The flow of electrons through the wire and the flow of ions through the salt bridge make up the electric current

General Set-up of Galvanic Cell



More reactive metal	Less reactive metal
- Acts as anode - It is the negative terminal of the battery - It is the electrode where oxidation occurs - The anode loses weight - The anode compartment contains the reducing agent - Anions move towards the anode (from the salt bridge)	- Acts as cathode - It is the positive terminal of the battery - It is the electrode where reduction occurs - The cathode gains weight - The cathode compartment contains the oxidizing agent - Cations move towards the cathode (from the salt bridge)

CHM.5.3.05.007.03 Write the oxidation and reduction half-reactions occurring at cathode and anode for a voltaic cell

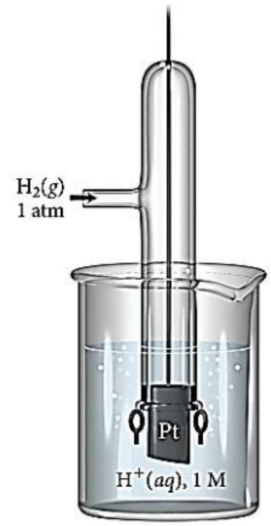
Please refer to the example on page 38

Standard Reduction Potentials

Half-Reaction	E^o(V)	Half-Reaction	E^o(V)
$Li^+ + e^- \rightleftharpoons Li$	−3.0401	$Cu^{2+} + e^- \rightleftharpoons Cu^+$	+0.153
$Ca^{2+} + 2e^- \rightleftharpoons Ca$	−2.868	$Cu^{2+} + 2e^- \rightleftharpoons Cu$	+0.3419
$Na^+ + e^- \rightleftharpoons Na$	−2.71	$O_2 + 2H_2O + 4e^- \rightleftharpoons 4OH^-$	+0.401
$Mg^{2+} + 2e^- \rightleftharpoons Mg$	−2.372	$I_2 + 2e^- \rightleftharpoons 2I^-$	+0.5355
$Be^{2+} + 2e^- \rightleftharpoons Be$	−1.847	$Fe^{3+} + e^- \rightleftharpoons Fe^{2+}$	+0.771
$Al^{3+} + 3e^- \rightleftharpoons Al$	−1.662	$NO_3^- + 2H^+ + e^- \rightleftharpoons NO_2 + H_2O$	+0.775
$Mn^{2+} + 2e^- \rightleftharpoons Mn$	−1.185	$Hg_2^{2+} + 2e^- \rightleftharpoons 2Hg$	+0.7973
$Cr^{2+} + 2e^- \rightleftharpoons Cr$	−0.913	$Ag^+ + e^- \rightleftharpoons Ag$	+0.7996
$2H_2O + 2e^- \rightleftharpoons H_2 + 2OH^-$	−0.8277	$Hg^{2+} + 2e^- \rightleftharpoons Hg$	+0.851
$Zn^{2+} + 2e^- \rightleftharpoons Zn$	−0.7618	$2Hg^{2+} + 2e^- \rightleftharpoons Hg_2^{2+}$	+0.920
$Cr^{3+} + 3e^- \rightleftharpoons Cr$	−0.744	$NO_3^- + 4H^+ + 3e^- \rightleftharpoons NO + 2H_2O$	+0.957
$S + 2e^- \rightleftharpoons S^{2-}$	−0.47627	$Br_2(l) + 2e^- \rightleftharpoons 2Br^-$	+1.066
$Fe^{2+} + 2e^- \rightleftharpoons Fe$	−0.447	$Pt^{2+} + 2e^- \rightleftharpoons Pt$	+1.18
$Cd^{2+} + 2e^- \rightleftharpoons Cd$	−0.4030	$O_2 + 4H^+ + 4e^- \rightleftharpoons 2H_2O$	+1.229
$PbI_2 + 2e^- \rightleftharpoons Pb + 2I^-$	−0.365	$Cl_2 + 2e^- \rightleftharpoons 2Cl^-$	+1.35827
$PbSO_4 + 2e^- \rightleftharpoons Pb + SO_4^{2-}$	−0.3588	$Au^{3+} + 3e^- \rightleftharpoons Au$	+1.498
$Co^{2+} + 2e^- \rightleftharpoons Co$	−0.28	$MnO_4^- + 8H^+ + 5e^- \rightleftharpoons Mn^{2+} + 4H_2O$	+1.507
$Ni^{2+} + 2e^- \rightleftharpoons Ni$	−0.257	$Au^+ + e^- \rightleftharpoons Au$	+1.692
$Sn^{2+} + 2e^- \rightleftharpoons Sn$	−0.1375	$H_2O_2 + 2H^+ + 2e^- \rightleftharpoons 2H_2O$	+1.776
$Pb^{2+} + 2e^- \rightleftharpoons Pb$	−0.1262	$Co^{3+} + e^- \rightleftharpoons Co^{2+}$	+1.92
$Fe^{3+} + 3e^- \rightleftharpoons Fe$	−0.037	$S_2O_8^{2-} + 2e^- \rightleftharpoons 2SO_4^{2-}$	+2.010
$2H^+ + 2e^- \rightleftharpoons H_2$	**0.0000**	$F_2 + 2e^- \rightleftharpoons 2F^-$	+2.866

CHM.5.3.05.004.02 Describe standard hydrogen electrode (SHE), while identifying the importance of its E° value and writing the half-cell reactions of the two possible reactions that could occur at the hydrogen electrode

SHE consists of a small sheet of platinum immersed in a hydrochloric acid solution that has a hydrogen ion concentration of 1*M*.



Hydrogen gas at a pressure of 1 atm is bubbled in and the temperature is maintained at 25°C.

The standard reduction potential (potential of hydrogen electrode) E° is 0.000 V

SHE can act as an oxidation-half reaction or a reduction-half reaction, depending on the half-cell to which it is connected

Oxidation half reaction $H_{2\,(g)} \rightleftharpoons 2\,H^+_{\,(aq)} + 2\,e^-$

Reduction half-reaction $2H^+_{\,(aq)} + 2\,e^- \rightleftharpoons H_{2\,(g)}$

CHM.5.3.05.004.03 Define the reduction potential and standard electrode potential ($E°$)

Reduction potential is the decrease in the voltage that takes place when a positive ion becomes less positive or neutral or when a neutral atom becomes a negative ion.
It is the tendency of a substance to gain electrons

Standard electrode potential is the potential developed by a metal or other material immersed in an electrolyte solution relative to the potential of the hydrogen electrode which is zero

CHM.5.3.05.004.06 Define electrode potential

Electrode potential is the potential difference between an electrode and its solution

Its magnitude measures the tendency for reduction-half reactions to occur

CHM.5.3.05.007.04 Write the cell notation and the overall chemical equation for a redox reaction occurring in a voltaic cell

General Cell notation:

Anode electrode | Anode solution || Cathode solution | Cathode electrode

- **Electrochemical Cell Notation**
- It is a way to indicate an electrochemical cell
- The first part (to the left) is the anode reaction with a single vertical line (|) between.
- The second part (to the right) is the cathode compartment.
- The double vertical line (‖) represents the salt bridge or the porous disk.

Example:

Consider the following cell notation: $Al_{(s)} \mid Al^{3+}_{(aq)} \parallel Cd^{2+}_{(aq)} \mid Cd_{(s)}$

a) Write the oxidation half-reaction.
$Al \rightarrow Al^{3+} + 3e^-$

b) Write the reduction-half reaction.
$Cd^{2+} + 2e^- \rightarrow Cd$

c) Write the overall reaction.
$2\ Al_{(s)} + 3\ Cd^{2+}_{(aq)} \rightarrow 2\ Al^{3+}_{(aq)} + 3\ Cd_{(s)}$

d) Identify the species oxidized and the species reduced.
The species oxidized is Al and the species reduced is Cd^{2+}.

CHM.5.3.05.007.05 Use the half-cell standard reduction potentials to calculate the electrochemical cell standard potential, while determining whether the redox reactions are spontaneous or nonspontaneous

To calculate the electrochemical cell standard potential:

$$E^o_{cell} = E^o_{cathode} - E^o_{anode}$$

Steps:
- All half-reactions are given as reduction half reactions in a standard table
- When a half reaction is reversed, the sign of E° is **reversed**.
- When a half reaction is multiplied by an integer, E° **remains the same**.
- Redox reactions are spontaneous if E^o_{cell} has a positive value and nonspontaneous if E^o_{cell} has a negative value

Example:

Refer to the standard reduction potential given below.
For each cell notation given below, write the balanced chemical equations, calculate the cell potentials and predict whether the reaction is spontaneous or not.

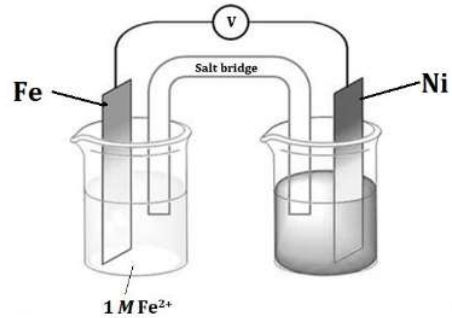
Half-reduction reaction	Standard Reduction Potential
$I_2 + 2\ e^- \rightarrow 2\ I^-$	$E^o = +0.53$ V
$Fe^{3+}_{(aq)} + e^- \rightarrow Fe^{2+}_{(aq)}$	$E^o = +0.77$ V
$Sn^{2+}_{(aq)} + 2\ e^- \rightarrow Sn_{(s)}$	$E^o = -0.14$ V
$Ag^+_{(aq)} + e^- \rightarrow Ag_{(s)}$	$E^o = +0.80$ V
$Zn^{2+}_{(aq)} + 2\ e^- \rightarrow Zn_{(s)}$	$E^o = -0.76$ V
$Cd^{2+}_{(aq)} + 2\ e^- \rightarrow Cd_{(s)}$	$E^o = -0.40$ V

Standard cell notation	**Balanced chemical equation**	**Cell potential**	**Is the reaction spontaneous?**
$I^- \mid I_2 \parallel Fe^{3+} \mid Fe^{2+}$	$2I^- + 2Fe^{3+} \rightarrow I_2 + 2Fe^{2+}$	+ 0.24 V	Yes
$Sn \mid Sn^{2+} \parallel Ag^+ \mid Ag$	$Sn + 2Ag^+ \rightarrow Sn^{2+} + 2Ag$	+ 0.94 V	Yes
$Cd \mid Cd^{2+} \parallel Zn^{2+} \mid Zn$	$Cd + Zn^{2+} \rightarrow Cd^{2+} + Zn$	− 0.36 V	No

Example:

Consider the following galvanic cell that operates spontaneously and the information below to answer questions **a** – **e**.

$Fe^{2+}_{(aq)} + 2\ e^{-} \rightarrow Fe_{(s)}$	$E^{o}\ =\ -0.44\ V$
$Ni^{2+}_{(aq)} + 2\ e^{-} \rightarrow Ni_{(s)}$	$E^{o}\ =\ -0.25\ V$



a) Identify each of the following.

- The anode	Fe
- The cathode	Ni
- Direction of electron flow	Fe to Ni
- Direction of current flow	Ni to Fe
- The electrode that decreases in mass	Fe
- The electrode that increases in mass	Ni
- The half-cell in which oxidation occurs	Fe half cell
- The half-cell in which reduction occurs	Ni half cell
- Direction of migration of anions in the salt bridge	towards Fe half-cell
- Direction of migration of cations the salt bridge	towards Ni half-cell

b) Write the reduction half-reaction.

$Ni^{2+}_{(aq)} + 2e^{-} \rightarrow Ni_{(s)}$

c) Write the oxidation half-reaction.

$Fe_{(s)} \rightarrow Fe^{2+}_{(aq)} + 2e^{-}$

d) Write the overall chemical reaction occurring.

$Fe_{(s)} + Ni^{2+}_{(s)} \rightarrow Ni_{(s)} + Fe^{2+}_{(aq)}$

e) Calculate the E^{o}_{cell}.

$E^{o}_{cell} = E^{o}_{cathode} - E^{o}_{anode} = -0.25 - (-0.44) = +0.19$ V

CHM.5.3.05.008.01 Use the standard reduction potentials to predict if a reaction occurs or not, while identifying the strongest reducing or oxidizing agent and the substance that is easily oxidized or reduced

- A reaction occurs if the E^{o}_{cell} has a positive value (spontaneous) and does not occur if it has a negative value
- Elements that have the most positive reduction potentials are easily reduced and would act as strong oxidizing agents (In general the non-metals)
- Elements that have the least positive reduction potentials are easily oxidized and would act as strong reducing agents (In general the metals)

CHM.5.3.05.017.01 Perform an experiment to investigate voltaic cell

Inspire Chemistry – Module 19 – Lesson 3: Electrolysis

CHM.5.3.05.011.01 Describe how a spontaneous redox reaction of an electrochemical cell can be reversed

An external source of power causes a nonspontaneous redox reaction to occur

CHM.5.3.05.011.02 Define electrolysis, while relating the definition to the concept of spontaneity of redox reactions

Electrolysis is the process of using electrical energy to produce a chemical reaction
The electrolytic process is non-spontaneous

Do Not Forget

a) ***PANIC*** ***Positive Anode Negative Is Cathode***

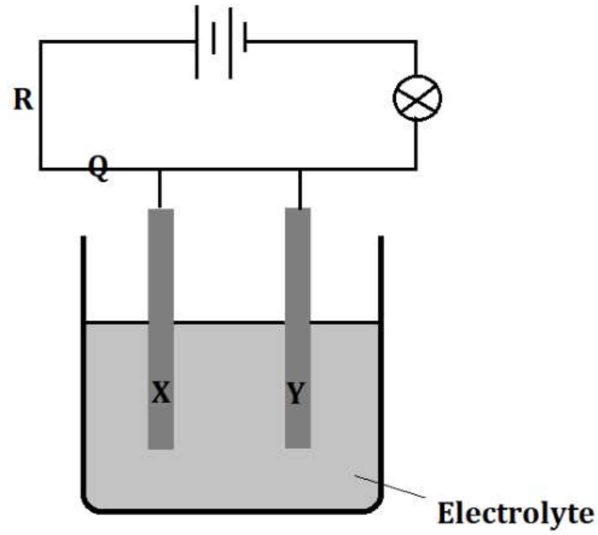
b) ***RED CAT*** ***Reduction at Cathode and oxidation at anode***

- **What happens at the electrodes?**

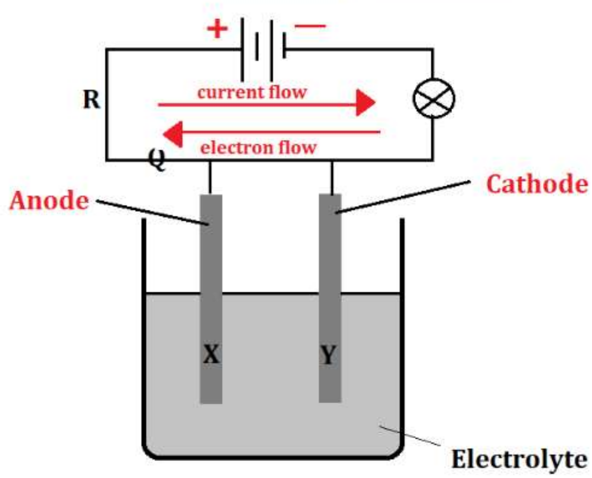
At Anode	At Cathode
- It is the positive electrode - It is the electrode connected to the positive terminal of the battery - It attracts the anions (negative ions) - Oxidation occurs (loss of electrons) - Non-metal ion is discharged (by the loss of electrons) to release a non-metal	- It is the negative electrode - It is the electrode connected to the negative terminal of the battery - It attracts the cations (positive ions) - Reduction occurs (gain of electrons) - Metal ion is discharged (by the gain of electrons) to release a metal

Example:

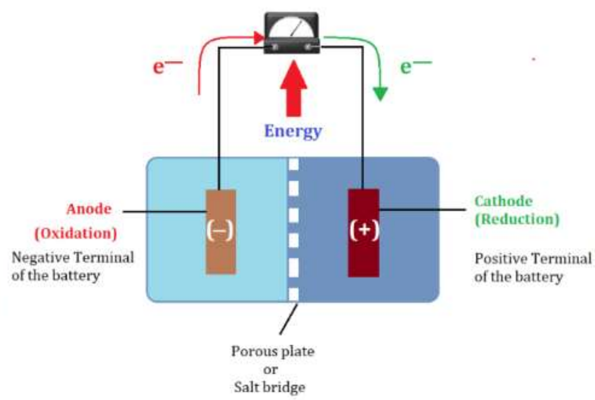
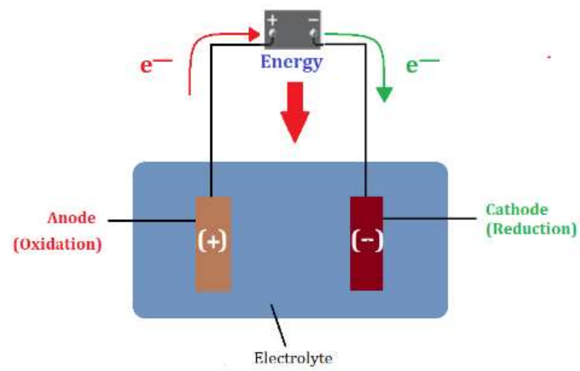
On the diagram below, an electrolyte, MX, was electrolyzed. Show the following:



1) Direction of flow of electric current in external circuit.
2) Direction of flow of electrons
3) Anode
4) Cathode
5) Positive and negative terminals of the battery
6) What is responsible for carrying the current between:
   i. Q and R **free electrons**
   ii. X and Y **free moving ions**



CHM.5.3.05.011.03 Compare between electrolytic cell and voltaic cell in terms of identifying where will reduction and oxidation processes take place, anode, cathode, direction of electron flow and current flow and spontaneity of the reaction occurring

Voltaic Cell	Electrolytic Cell
	
- It is an electrochemical cell used to convert chemical energy into electrical energy	- It is an electrochemical cell used to cause a chemical change through the application of electrical energy
- Energy is released from a spontaneous redox reaction	- Energy is absorbed to drive a non-spontaneous reaction

CHM.5.3.05.018.01 Write half-cell reactions while identifying the products of electrolysis of molten ionic compounds

1. **At anode**
   1) Non-metal is always discharged and the product is always a gas (as O_2, Cl_2, Br_2 or I_2)
   2) If bromine gas is produced, the experiment must be carried out in the fume cupboard because bromine gas produced is highly toxic or poisonous.

2. **At cathode**
   1) Metals are always discharged
   2) All metals are grey in color (except copper, Cu, it is red in color).
   3) All metals sink except Group one metals (Na, K, Li) they float on the top of the molten electrolyte because they have low density.
   3) The molten metal produced stays liquid because the temperature of the electrolyte is above the melting point of the metal.

- Example: The electrolysis of molten lead (II) bromide, $PbBr_2$

  At anode: $2\ Br^-_{(l)} \rightarrow Br_{2\,(g)} + 2\ e^-$

  At cathode: $Pb^{2+}_{(l)} + 2\ e^- \rightarrow Pb_{(l)}$

  Overall reaction: $PbBr_{2\,(l)} \rightarrow Pb^{2+}_{(l)} + 2\ Br^-_{(l)}$

CHM.5.3.05.018.02 Write half-cell reactions while identifying the products of electrolysis aqueous ionic compounds

1. If the solution is **<u>dilute</u>**, H^+ and OH^- are **<u>always</u>** discharged and the solution becomes concentrated.

   - ✓ Reaction at anode: $4\,OH^-_{(aq)} \rightarrow 2\,H_2O_{(l)} + O_{2\,(g)} + 4e^-$
   - ✓ Reaction at cathode: $4\,H^+_{(aq)} + 4\,e^- \rightarrow 2\,H_{2\,(g)}$

2. **Product at Cathode:**

   a) If the metal ion of the electrolyte is high in the reactivity series (as Na^+, K^+, Ca^{2+} or Mg^{2+}) in other words above hydrogen in the reactivity series, the metal is never discharged, and the product at the cathode is hydrogen gas.

   $2\,H_2O_{(l)} + 2\,e^- \rightarrow H_{2\,(g)} + 2\,OH^-$

   OR $2\,H^+_{(aq)} + 2\,e^- \rightarrow H_{2\,(g)}$

   b) If the metal ion of the electrolyte is below hydrogen in the reactivity series (as Cu^{2+} or Ag^+), then the metal is deposited at the cathode.

   $Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$ and $Ag^+_{(aq)} + e^- \rightarrow Ag_{(s)}$

3. **Product at Anode:**

   a) If you have concentrated halide (chloride or bromide or iodide), then the product is the halogen.

If halide is	Product is	Observation
Chloride, Cl^-	Chlorine gas, Cl_2	Yellowish green gas is produced that bleaches blue litmus paper $2\,Cl^-_{(aq)} \rightarrow Cl_{2\,(g)} + 2\,e^-$
Bromide, Br^-	Bromine liquid, Br_2	Reddish brown liquid is produced $2\,Br^-_{(aq)} \rightarrow Br_{2\,(l)} + 2\,e^-$
Iodide, I^-	Iodine solid, I_2	Grey solid is produced $2\,I^-_{(aq)} \rightarrow I_{2\,(s)} + 2\,e^-$

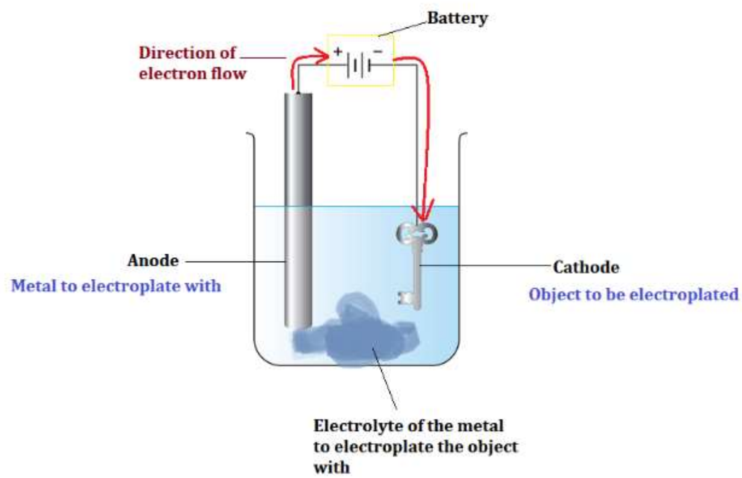
   b) If you do not have a halide, then the product is oxygen gas.

   $4\,OH^-_{(aq)} \rightarrow 2\,H_2O_{(l)} + O_{2\,(g)} + 4e^-$

CHM.5.3.05.011.05 Define electroplating while describing how it works, identifying anode, cathode and electrolyte needed for an electrolytic cell in which a selected metal is to be plated on an object (car or spoon,…etc)

Electroplating is the process of covering an object with metal using electricity. It is an electrolytic process that deposits metal on a surface
The object to be plated is the cathode and the plating metal is the anode

Set-up for Electroplating



Oxidation half reaction (At Anode): $M \rightarrow M^{+} + e^{-}$
Reduction half-reaction (At Cathode): $M^{+} + e^{-} \rightarrow M$

CHM.5.3.05.011.06 Perform an experiment to prepare and operate an electrochemical cell for electroplating