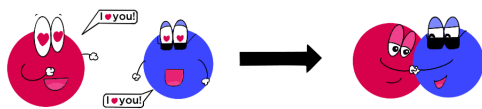
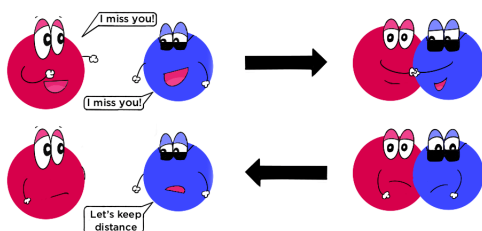


Chemistry Unit 6 Notes

Differences Between Non-reversible and Reversible Reactions

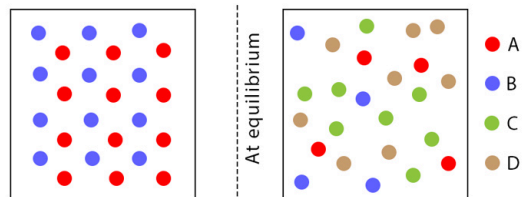
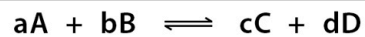


In non-reversible reactions, reactants convert to products and cannot convert back to the reactants



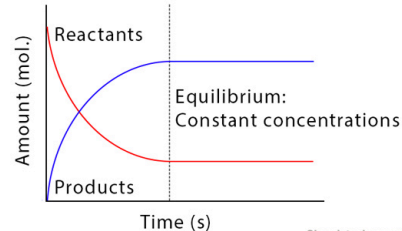
In reversible reactions, the forward reaction and backward reaction occur simultaneously

Chemical Equilibrium with Constant Concentrations



Initial concentration of the reactants

Concentrations of reactants and products at equilibrium



ChemistryLearner.com

CHARACTERISTICS OF EQUILIBRIUM

1. Reaction must take place in a closed system- no reactant or product can enter or leave
2. Temperature must remain constant.
3. All reactants and products present are in constant dynamic motion.

Le Chatelier's Principle

Disturbing a system at dynamic equilibrium shifts the equilibrium in the direction that counteracts the change.



Concentration

↑ reactant concentration
↑ favors product formation

↑ product concentration
↑ favors reactant formation

Temperature

↑ temperature
❄️ favors endothermic reaction

↓ temperature
🔥 favors exothermic reaction

Pressure

↑ pressure
🔗 favors side with fewer molecules

↓ pressure
🔗 favors side with more molecules

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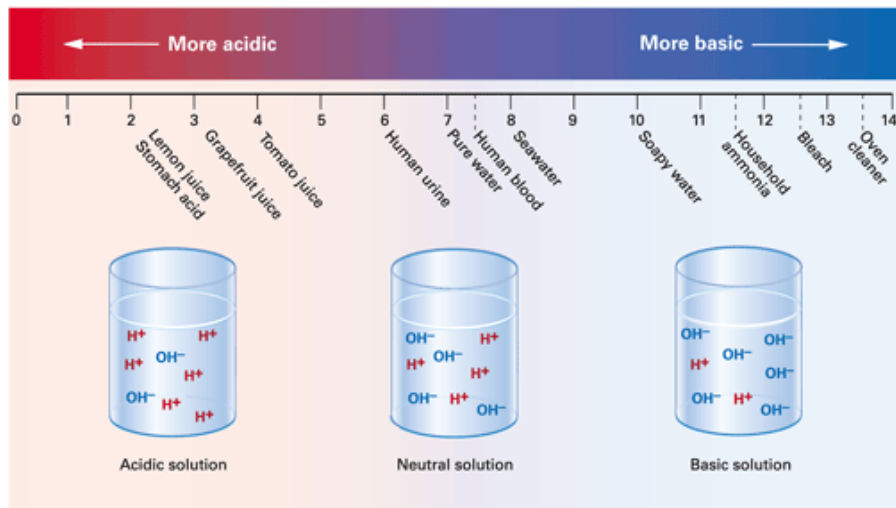
Properties

ACIDS

- electrolyte
- sour taste
- sticky feel
- turn litmus red
- react with bases to form water and a salt (ionic compound)

BASES

- electrolyte
- bitter taste
- slippery feel
- turn litmus blue
- react with acids to form water and a salt (ionic compound)



Arrhenius Acids and Bases



Acid - forms H^+ in water



Base - forms OH^- in water

Limitations: Some bases (like NH_3 or Na_2CO_3) don't have an OH group, but DO produce hydroxide ions in solution. According to this model, though, they would not be considered bases.

Bronsted Lowry Acid and Base

A Bronsted Lowry acid donates protons, while a Bronsted Lowry base accepts protons.

donates proton

Acid

accepts proton

Base

\rightleftharpoons

Conjugate base

+

Conjugate acid

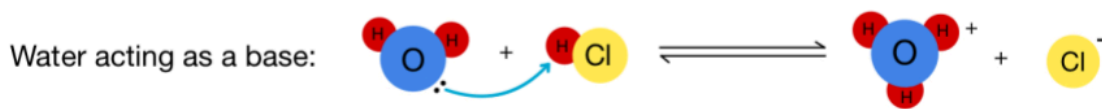
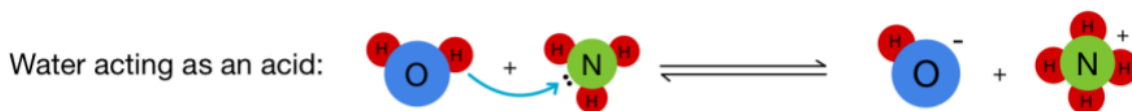
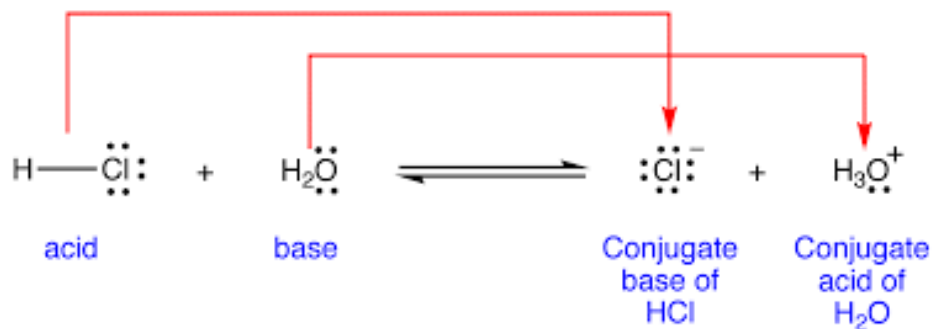
$$\text{H}_2\text{O}(l) + \text{NH}_3(aq) \rightleftharpoons \text{OH}^-(aq) + \text{NH}_4^+(aq)$$

Remember, a proton is a hydrogen or H^+ .

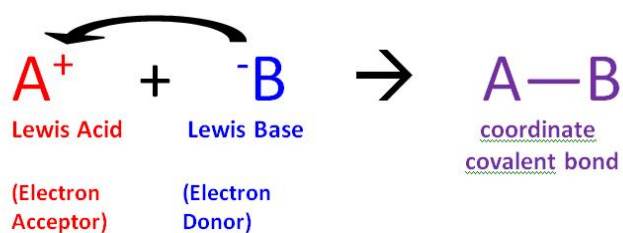
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Conjugate acid- formed when base accepts hydrogen ion

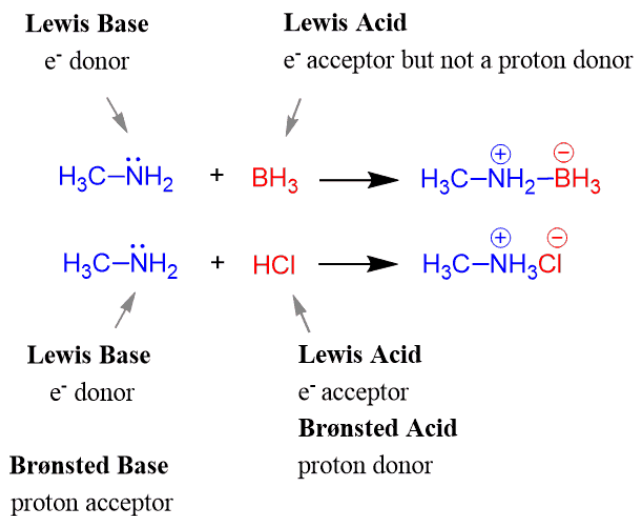
Conjugate base- formed when acid donates hydrogen ion



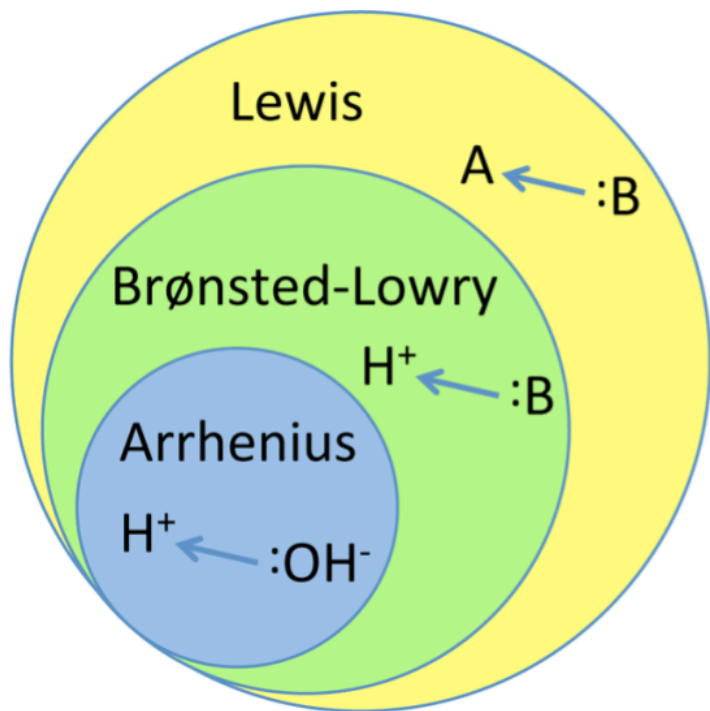
Disadvantages of Bronsted Lowry: can't explain acidic or basic nature of compounds with no tendency to gain or lose hydrogen ion like CO_2 .



Any Base is a Lewis Base and Any Acid is a Lewis Acid



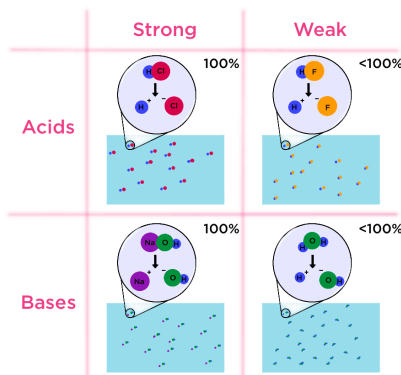
Not every Base is a Brønsted Base and Any Acid is a Brønsted Acid



Summary of 3 Models of Acids & Bases

Model	Definition of Acid	Definition of Base
Arrhenius	H ⁺ producer	OH ⁻ producer
Brønsted-Lowry	H ⁺ donor	H ⁺ acceptor
Lewis	Electron-pair acceptor	Electron-pair donor

What are Strong and Weak Acids/Bases



Strong Acids/bases break apart 100% while weak acids/bases do not break apart 100% in aqueous solution

Brønsted-Lowry Acids and Bases

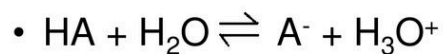
Relative Strengths of Acids and Bases

	ACID	BASE	
100 percent ionized in H ₂ O	Strong	HCl	Cl ⁻
		H ₂ SO ₄	HSO ₄ ⁻
		HNO ₃	NO ₃ ⁻
	H ⁺ (aq)	H ₂ O	Negligible
Acid strength increases ↑	Weak	HSO ₄ ⁻	SO ₄ ²⁻
		H ₃ PO ₄	H ₂ PO ₄ ⁻
		HF	F ⁻
		HC ₂ H ₃ O ₂	C ₂ H ₃ O ₂ ⁻
		H ₂ CO ₃	HCO ₃ ⁻
		H ₂ S	HS ⁻
		H ₂ PO ₄ ⁻	HPO ₄ ²⁻
		NH ₄ ⁺	NH ₃
		HCO ₃ ⁻	CO ₃ ²⁻
		HPO ₄ ²⁻	PO ₄ ³⁻
		H ₂ O	OH ⁻
Negligible		HS ⁻	S ²⁻
		OH ⁻	O ₂ ⁻
		H ₂	H ⁻
			Strong

- The *stronger* the **acid**, the *weaker* the conjugate **base**.
- The *stronger* the **base**, the *weaker* the conjugate **acid**.
- H⁺ is the strongest acid that can exist in equilibrium in aqueous solution.
- OH⁻ is the strongest base that can exist in equilibrium in aqueous solution.

Acid Ionization Constant

- **Acid Ionization Constant** (K_a): the equilibrium constant for the ionization reaction of an acid with water

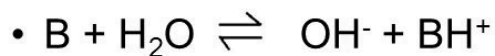


$$K_a = \frac{[H^+][A^-]}{[HA]}$$

- Large K_a = Strong acid
- Small K_a = Weak acid

Base Ionization Constant

- **Base Ionization Constant** (K_b): the equilibrium constant for the ionization reaction of a base with water



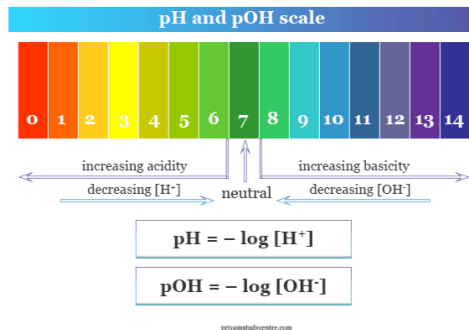
$$K_b = \frac{[BH^+][OH^-]}{[B]}$$

- Large K_b = Strong base
- Small K_b = Weak base

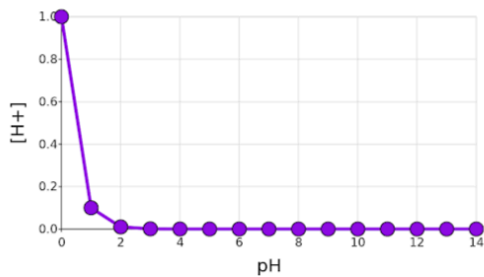
$$K_w = [H^+][OH^-]$$

$$= (1 \times 10^{-7})(1 \times 10^{-7})$$

$$= 1 \times 10^{-14}$$



$$pH + pOH = 14.00$$



pH = -log[H⁺]


0	10 ⁰	1
1	10 ⁻¹	0.1
2	10 ⁻²	0.01
3	10 ⁻³	0.001
4	10 ⁻⁴	0.0001
5	10 ⁻⁵	0.00001
6	10 ⁻⁶	0.000001
7	10 ⁻⁷	0.0000001
8	10 ⁻⁸	0.00000001
9	10 ⁻⁹	0.000000001
10	10 ⁻¹⁰	0.0000000001
11	10 ⁻¹¹	0.00000000001
12	10 ⁻¹²	0.000000000001
13	10 ⁻¹³	0.0000000000001
14	10 ⁻¹⁴	0.00000000000001

$$K_a = \frac{[H^+][A^-]}{[HA]}$$

Neutralization Reaction

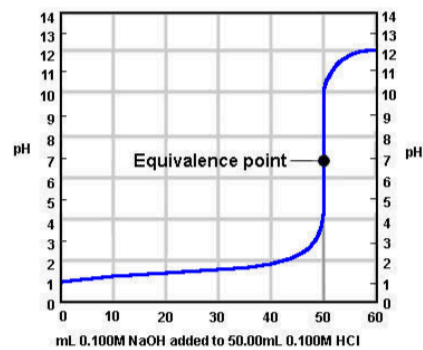
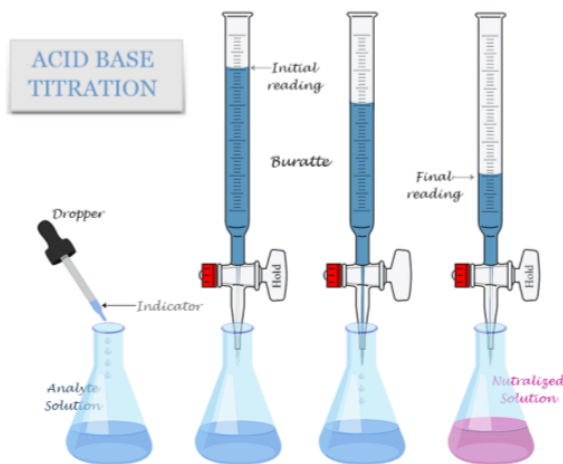
A neutralization reaction is a chemical reaction between an acid and a base that forms a salt and water.

Acid + Base → Salt + Water
 $\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O}$



$\text{H Cl} + \text{Na OH} \rightarrow \text{Na Cl} + \text{H OH}$

Double Replacement



BUFFERS

1. Definition

- Buffers are solutions which can resist changes in pH when acid or alkali is added.

2. Composition of a buffer

Buffers are 2 types:

- Mixtures of weak acids with their salt with a strong base or
- Mixtures of weak bases with their salt with a strong acid.

- A few examples are given below:

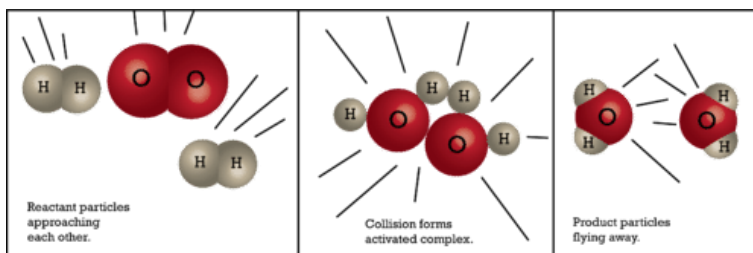
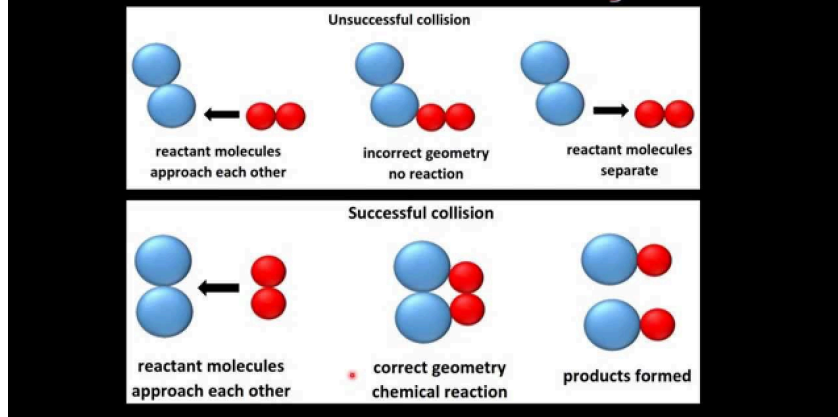
- Bicarbonate buffer
- Acetate buffer
- Phosphate buffer.

Buffer Capacity

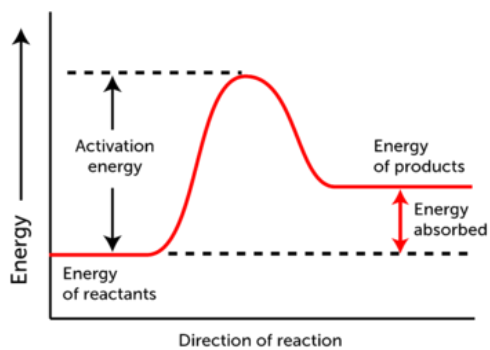
$$\text{REACTION RATE} = \frac{\text{CHANGE IN MASS OF REACTANT OR PRODUCT}}{\text{TIME}}$$

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Tutorials for IB Chemistry

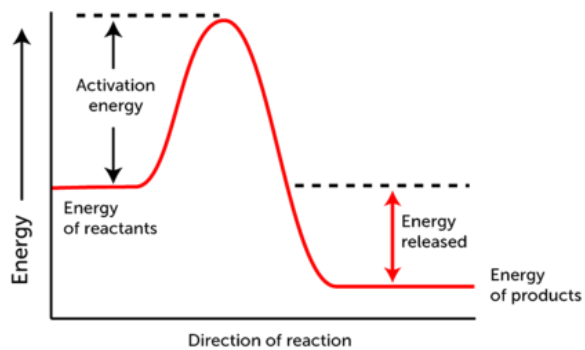
Collision theory

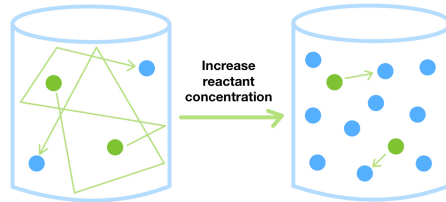
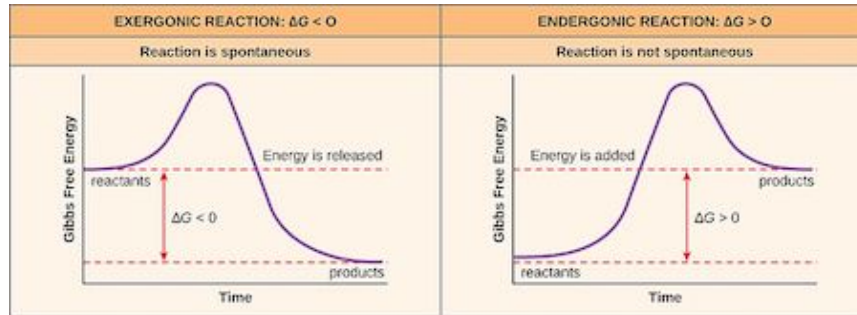


Endothermic Reaction



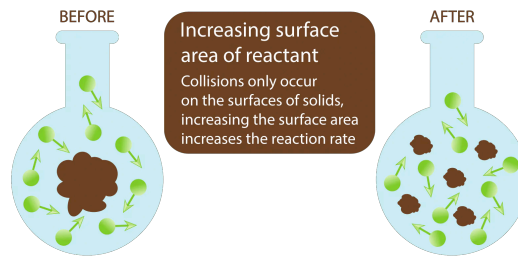
Exothermic Reaction





In this reaction, because the concentration of the reactants is small, it's hard for the green molecules to make collisions with the blue molecules

A higher concentration of blue molecules makes it more likely for the green molecules to make collisions, thus causing the reaction to move faster!



SOLIDS ONLY

