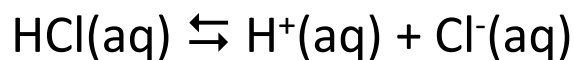


# Grade 12 Chemistry

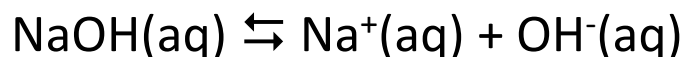
Chemical Systems & Equilibrium  
Class 12

## Arrhenius Theory

- Acids – substances that dissociate in water to produce  $\text{H}^+$



- Bases – substances that dissociate in water to produce  $\text{OH}^-$



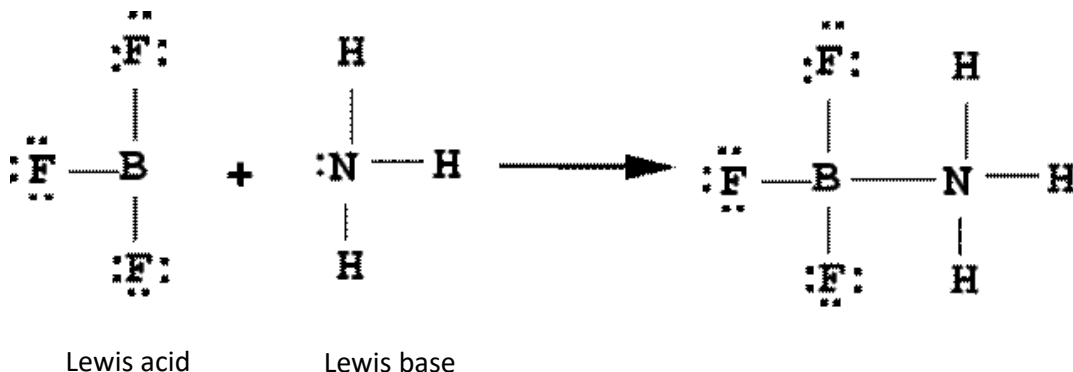
- Limitations:
  - Restricted to aqueous solutions
  - Does not account for bases with no  $\text{OH}^-$  groups

# Bronsted-Lowry Theory

- Acids – substances that can transfer  $H^+$ ; proton-donors
- Bases – substances that can accept  $H^+$ ; proton-acceptors
- Proton = refers to the nucleus of the hydrogen atom ( $H^+$ ); the  $H^+$  that has been removed from the acid molecule

## Lewis Acids and Bases

- Acids – electron-pair acceptors
- Bases – electron-pair donors

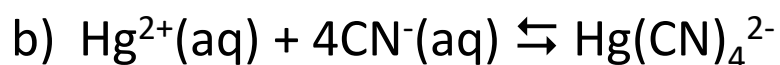




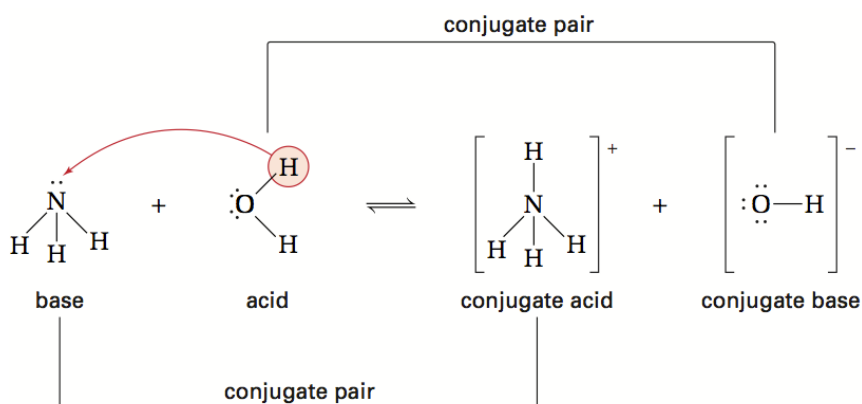
## Checkpoint



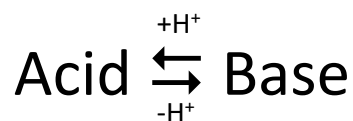
Identify the Lewis acid and Lewis base in each of the following reactions:



## Conjugate Acid-Base Pairs



Chemical species whose formula only differs by one proton

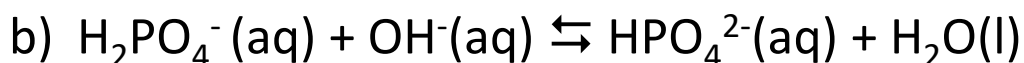




## Checkpoint



Identify the conjugate acid-base pair in each reaction:



## Strong Acids and Strong Bases

### Strong Acids



- A strong acid is a substance that can completely dissociate in water
- Ex: HI, HBr, HCl,  $\text{HClO}_4$ ,  $\text{H}_2\text{SO}_4$ , and  $\text{HNO}_3$

I Bring Cans of Cola So No

increasing acid strength →

increasing electronegativity

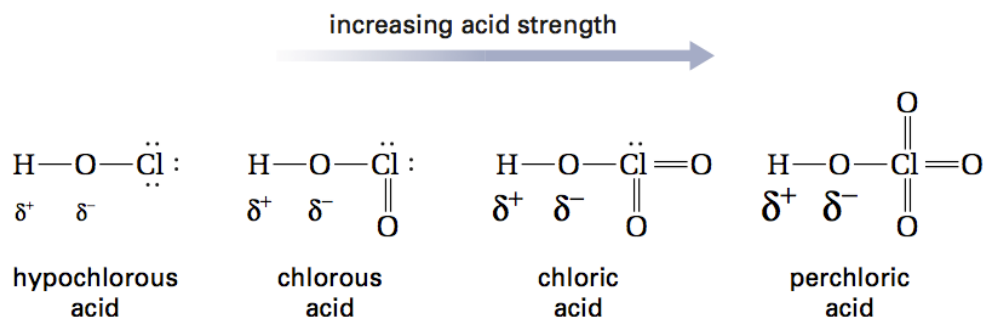
					18(VIIIA)
13 (IIIA)	14 (IVA)	15 (VA)	16 (VIA)	17 (VIIA)	
	CH <sub>4</sub>	NH <sub>3</sub>	H <sub>2</sub> O	HF	
			H <sub>2</sub> S	HCl	
			H <sub>2</sub> Se	HBr	
			H <sub>2</sub> Te	HI	

↓ decreasing bond strength

↑ increasing acid strength

- Across a Period: EN increases drawing electrons away from H; water can pull H away
- Down a Group: Bond strength decreases; easier to pull away H
- Monoprotic – donate 1 H<sup>+</sup>
- Polyprotic – donate more than 1 H<sup>+</sup>

- **Oxoacids** – increase in strength with increasing oxygen atoms since O is very electronegative



## Strong Bases

	1A (1)	2A (2)
1		
2	Li <sub>2</sub> O	BeO
3	Na <sub>2</sub> O	MgO
4	K <sub>2</sub> O	CaO
5	Rb <sub>2</sub> O	SrO
6	Cs <sub>2</sub> O	BaO
7	Fr <sub>2</sub> O	RaO

- All oxides and hydroxides of Group 1 metals (ex: NaOH, KOH) and Group 2 metals below Beryllium (ex: Ca(OH)<sub>2</sub> and Ba(OH)<sub>2</sub>)
- BeO is an exception. It is a weak base because it is a relatively small atom, which means the bond to O is strong and is not easily broken by polar water molecules

## Calculations Involving Strong Acids and Strong Bases

- Since strong acids completely dissociate in water, we can assume that the concentration of the acid is equal to the concentration of H<sub>3</sub>O<sup>+</sup>
- Since strong bases completely dissociate in water, we can assume that the concentration of the base is equal to the concentration of OH<sup>-</sup>



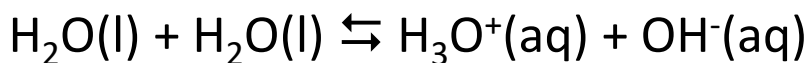
## Checkpoint



During an experiment, a student pours 25.0mL of 1.40M  $\text{HNO}_3(\text{aq})$  into a beaker that contains 15.0mL of 2.00M  $\text{NaOH}(\text{aq})$ . Is the resulting solution acidic or basic?

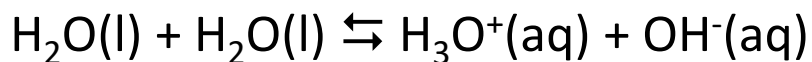
What is the concentration of the ion that causes the solution to be acidic or basic?

## Water



- At 25°C, only 2 water molecules in 1 billion dissociate
- In neutral water at 25°C, the  $[\text{H}_3\text{O}^+] = [\text{OH}^-]$  which is  $1.0 \times 10^{-7} \text{ mol/L}$

## Ion Product Constant of Water, $K_w$



$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$$

$$K_w = (1.0 \times 10^{-7})(1.0 \times 10^{-7})$$

$$K_w = 1.0 \times 10^{-14}$$

- If  $[\text{H}_3\text{O}^+] > [\text{OH}^-]$ , acidic at 25°C
- If  $[\text{H}_3\text{O}^+] = [\text{OH}^-]$ , neutral
- If  $[\text{H}_3\text{O}^+] < [\text{OH}^-]$ , basic



### Checkpoint



Find  $[\text{H}_3\text{O}^+]$  and  $[\text{OH}^-]$  in each solution:

- 2.5 M of nitric acid
- 0.16 M of barium hydroxide



# pH and pOH

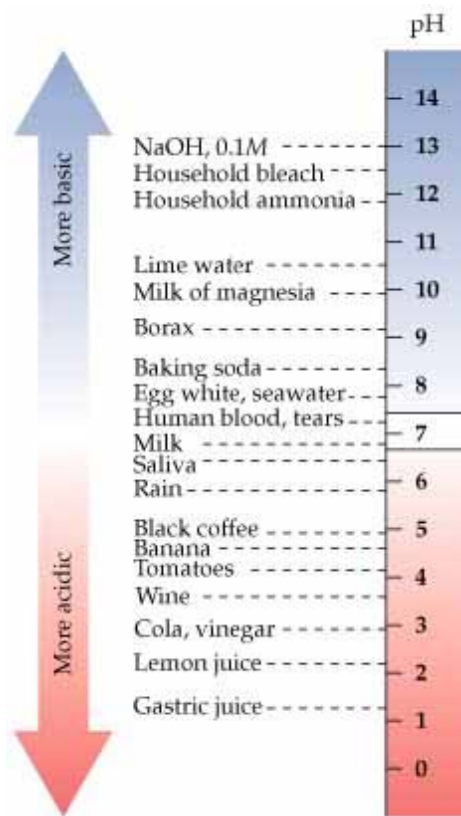
- It is more convenient to express the  $[H_3O^+]$  and  $[OH^-]$  on a logarithmic scale
- pH = power of hydrogen
- pOH = power of hydroxide

$$pH = -\log[H_3O^+] \quad pOH = -\log[OH^-]$$

$$[H_3O^+] = 10^{-pH} \quad [OH^-] = 10^{-pOH}$$

- Practical range of the pH scale = 0 to 14

$$pH + pOH = 14$$



- pH > 7 is basic
- pH = 7 is neutral
- pH < 7 is acidic
- A solution that is more concentrated than 1.0M would give a negative pH
  - Ex: a 10M solution of HCl fully dissociates therefore the pH = -1



## Checkpoint



If the pH of urine is outside the normal range of values, this can indicate medical problems. Suppose that the pH of a urine sample was measured to be at 5.53 at 25°C. Calculate the pOH,  $[H_3O^+]$  and  $[OH^-]$  for the sample.

## Weak Acids and Bases

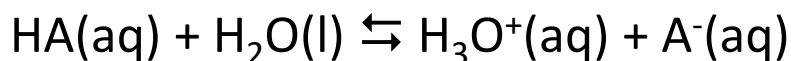
### Weak Acids

- Weak acids do not completely dissociate in water so  $[H_3O^+]$  depends on the initial concentration of the acid and the concentration of the conjugate base
- Ex: Acetic Acid, Aspirin, Vitamins



# Acid Dissociation Constant

- Strength of an acid is directly related to how much the products are favored over the reactants



$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$$

- If  $K_a > 1$ , products are favored, strong acid
- If  $K_a < 1$ , reactants are favored, weak acid

**Table 8.2** Some Acid Dissociation Constants for Weak Acids at 25°C

Acid	Formula	Acid dissociation constant, $K_a$
acetic acid	$\text{CH}_3\text{COOH}$	$1.8 \times 10^{-5}$
chlorous acid	$\text{HClO}_2$	$1.1 \times 10^{-2}$
formic acid	$\text{HCOOH}$	$1.8 \times 10^{-4}$
hydrocyanic acid	$\text{HCN}$	$6.2 \times 10^{-10}$
hydrofluoric acid	$\text{HF}$	$6.6 \times 10^{-4}$
hydrogen oxide (water)	$\text{H}_2\text{O}$	$1.0 \times 10^{-14}$
lactic acid	$\text{CH}_3\text{CHOHCOOH}$	$1.4 \times 10^{-4}$
nitrous acid	$\text{HNO}_2$	$7.2 \times 10^{-4}$
phenol	$\text{C}_6\text{H}_5\text{OH}$	$1.3 \times 10^{-10}$



## Checkpoint



Of the following acids, which one would dissociate to the greatest extent in water?

- a) HCN ( $K_a=6.2 \times 10^{-10}$ )    c) HClO ( $K_a=2.9 \times 10^{-8}$ )  
b) HCNO ( $K_a=3.3 \times 10^{-4}$ )    d) HBrO ( $K_a=2.2 \times 10^{-9}$ )

## Percent Dissociation

- Another useful measure of the strength of a weak acid is the percent dissociation

$$\text{Percent Dissociation} = \frac{[\text{H}^+]}{[\text{HA}]_{\text{initial}}} \times 100\%$$

# Aqueous Equilibria

- The  $K_a$  can be used to calculate the equilibrium concentrations and the pH in a solution of weak acid
- Write the chemical equation and set up an ICE table
  - Let  $x$  represent the change in concentration
  - If  $[HA]/K_a > 500$ ,  $x$  can be ignored
  - If  $[HA]/K_a < 500$ , use quadratic formula



## Checkpoint



Propanoic acid,  $\text{CH}_3\text{CH}_2\text{COOH}$ , is a weak monoprotic acid that is used to inhibit mould formation in bread. A student prepared a 0.10M solution of propanoic acid and found that the pH was 2.96. What is the  $K_a$  for propanoic acid?

What percent of its molecules were dissociated in the solution?



## Checkpoint



Phosphoric acid,  $\text{H}_3\text{PO}_4$ , is one of the world's most important industrial chemicals. It is mainly used to manufacture phosphate fertilizers. It is also the ingredient that gives cola drinks their tart, biting taste. Calculate the pH,  $[\text{H}_2\text{PO}_4^-]$ , and  $[\text{HPO}_4^{2-}]$  of a 3.5M aqueous solution of  $\text{H}_3\text{PO}_4(\text{aq})$ .

### Weak Bases

- Weak bases do not completely dissociate in water so  $[\text{OH}^-]$  depends on the initial concentration of the base and the concentration of the conjugate acid
- Ex: Caffeine, piperidine in black pepper



Piperidine



# Base Dissociation Constant

- Strength of the base is directly related to how much the products are favored over the reactants



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

- If  $K_b > 1$ , products favored, strong base
- If  $K_b < 1$ , reactants favored, weak base

**Table 8.3** Some Base Dissociation Constants at 25°C

Base	Formula	Base dissociation constant, $K_b$
ethylenediamine	$NH_2CH_2CH_2NH_2$	$5.2 \times 10^{-4}$
dimethylamine	$(CH_3)_2NH$	$5.1 \times 10^{-4}$
methylamine	$CH_3NH_2$	$4.4 \times 10^{-4}$
trimethylamine	$(CH_3)_3N$	$6.5 \times 10^{-5}$
ammonia	$NH_3$	$1.8 \times 10^{-5}$
hydrazine	$N_2H_4$	$1.7 \times 10^{-6}$
pyridine	$C_5H_5N$	$1.4 \times 10^{-9}$
aniline	$C_6H_5NH_2$	$4.2 \times 10^{-10}$
urea	$NH_2CONH_2$	$1.5 \times 10^{-14}$



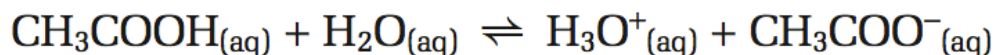
## Checkpoint



The characteristic taste of tonic water is due to the addition of quinine. Quinine is a naturally occurring compound that is also used to treat malaria. The base dissociation constant,  $K_b$  for quinine is  $3.3 \times 10^{-6}$ . Calculate  $[\text{OH}^-]$  and the pH of a  $1.7 \times 10^{-3} \text{ M}$  solution of quinine.

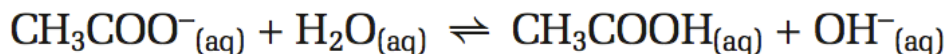
## Relationship between $K_a$ and $K_b$

### Acetic Acid



$$K_a = \frac{[\text{CH}_3\text{COO}^-][\text{H}_3\text{O}^+]}{[\text{CH}_3\text{COOH}]}$$

### Acetate Ion (Conjugate Base)



$$K_b = \frac{[\text{CH}_3\text{COOH}][\text{OH}^-]}{[\text{CH}_3\text{COO}^-]}$$



$$\begin{aligned}
 K_a K_b &= \frac{[\text{H}_3\text{O}^+][\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]} \times \frac{[\text{CH}_3\text{COOH}][\text{OH}^-]}{[\text{CH}_3\text{COO}^-]} \\
 &= [\text{H}_3\text{O}^+][\text{OH}^-] \\
 &= K_w
 \end{aligned}$$

$$K_a \times K_b = K_w$$

- As the strength of an acid increases (larger  $K_a$ ), the strength of its conjugate base decreases (smaller  $K_b$ )
- Strong Acid  $\rightarrow$  Weak Conjugate Base
- Strong Base  $\rightarrow$  Weak Conjugate Acid