# Class 6.3 Acids and Bases

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### Hydrolysis of Salts made from Weak Acids or Weak Bases

- When salts like NaCl or KNO<sub>3</sub> are dissolved in water, the pH is not much affected. But when NaCN or NH<sub>4</sub>Cl are dissolved, the pH <u>is</u> changed. Why?
- CN<sup>-</sup> is a base and NH<sub>4</sub><sup>+</sup> is an acid:

$$CN^{-}(aq) + H_2O \rightleftharpoons HCN(aq) + OH^{-}(aq)$$

 $NH_4^+(aq) + H_2O \rightleftharpoons NH_3(aq) + H_3O^+(aq)$ 

## Hydrolysis of Salts - Example

■ For NH<sub>3</sub>, pK<sub>b</sub> = 4.74. What is the pH of a 10<sup>-2</sup> M solution of NH<sub>4</sub>Cl?

TABLE 10.3 Acid name	SOLVENT	LEVELING	Y STRENGTH Name	
Strong acid Very		y weak base		
hydroiodic acid	HI	1-	iodide ion	
perchloric acid	HCIO <sub>4</sub>	CIO <sub>4</sub> =	perchlorate ion	
hydrobromic acid	LIDe	Dv -	hromide ion	
hydrochloric acid	No Acids Stronger than H <sub>3</sub> O <sup>+</sup> loride ion			
sulfuric acid	NO ACIGO OII	onger man m	drogen sulfate ion	
chloric acid	HCIO <sub>3</sub>	CIO3 -	chlorate ion	
nitric acid	HNO <sub>3</sub>	NO3-	nitrate ion	
hydronium ion	H <sub>3</sub> O+	H <sub>2</sub> O		
hydrogen sulfate ion	HSO <sub>4</sub> -	5042-	sulfate ion	
hydrofluoric acid	HF	F-	fluoride ion	
nitrous acid	HNO <sub>2</sub>	NO <sub>2</sub> =	nitrite ion	
acetic acid	CH3COOH	CH <sub>3</sub> CO <sub>2</sub> -	acetate ion	
carbonic acid	H <sub>2</sub> CO <sub>3</sub>	HCO3 =	hydrogen carbonate ior	
hydrosulfuric acid	H <sub>2</sub> S	HS-	hydrogen sulfide ion	
ammonium ion	NH <sub>4</sub> +	NH <sub>3</sub>	ammonia	
hydrocyanic acid	HCN	CN-	cyanide ion	
hydrogen carbonate ion		CO3 <sup>2</sup> -	carbonate ion	
methylammonium ion	CH <sub>3</sub> NH <sub>3</sub> +	CH <sub>3</sub> NH <sub>2</sub>	methylamine	
Hater -		OH-	hydroxide ion ———	
ammonia	NH3	NH <sub>2</sub> =	amide ion	
hydrogen			hydride ion	
	No Bases Str	ronger than Ol		
hydroxide ion	OH	0-	oxide ion	
Very weak acid		S	Strong base	

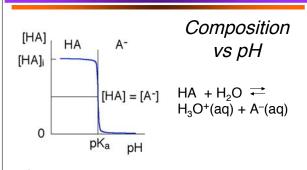
## Composition vs pH

$$HA + H_2O \Rightarrow H_3O^+(aq) + A^-(aq)$$

$$K_a = \frac{[H_3O^+][A^-]}{[HA]};$$

when HA is half deprotonated,

$$[HA] = [A^-] \Rightarrow pH = pK_a$$



■ Composition changes dramatically when pH is close to the pK<sub>a</sub> of the acid.

#### Polyprotic Acids

As we have seen, several common acids can potentially donate more than one proton. Sulfuric (H<sub>2</sub>SO<sub>4</sub>) and phosphoric (H<sub>3</sub>PO<sub>4</sub>) acids are particularly prominent examples:

$$H_2SO_4$$
:  $pK_{a1} < 0$ ;  $pK_{a1} = 1.92$ 

$$H_3PO_4$$
:  $pK_{a1} = 2.12$ ;  $pK_{a2} = 7.21$ ;  $pK_{a3} = 12.68$ 

What are the concentrations of all species in a 0.1 M phosphoric acid solution?

#### Features of Polyprotic Acids

 $H_3PO_4$  is a representative example:  $pK_{a1} = 2.12$ ;  $pK_{a2} = 7.21$ ;  $pK_{a3} = 12.68$ 

At what pH are the concentrations of  $H_3PO_4$  and  $H_2PO_4^-$  equal?

At that pH, what the concentrations of  $HPO_4^{2^-}$  and  $PO_4^{3^-}$ ?

Tough type to answer directly: What is the pH of a  $H_2PO_4^-$  or  $HPO_4^{2^-}$  salt solution? (e.g.,  $K_2HPO_4$  or  $K_2PO_4$  solution)

Easier: what's the pH when  $[H_2PO_4^-] = [PO_4^3]$ ?

$$K_{a2} = \frac{[{\rm H_3O^+}][{\rm HPO_4}^{2^-}]}{[{\rm H_2PO_4}^{-1}]}; \ K_{a3} = \frac{[{\rm H_3O^+}][{\rm PO_4}^{3^-}]}{[{\rm HPO_4}^{2^-}]}$$

when 
$$[H_2PO_4^-] = [PO_4^{3-}]$$
.

$$\frac{[{\rm H_3O^+}][{\rm HPO_4^{2^-}}]}{K_{a2}} = K_{a3} \frac{[{\rm HPO_4^{2^-}}]}{[{\rm H_3O^+}]} \ \Rightarrow \ {\rm p}H = \ \frac{1}{2} \big({\rm p}K_{a2} + {\rm p}K_{a3}\big)$$

But at this pH,  $[HPO_4^{2-}] >> [H_2PO_4^{-1}] = [PO_4^{3-}]$ 

