A SALTY REVIEW

In acid/base chemistry, there are a variety of electrolytes that we will discuss. Similarly to STRONG ACIDS and BASES, SALTS are also **STRONG ELECTROLYTES**. Strong electrolytes **completely dissociate** in solution. When a salt is dissolved it can lead to a variety of different scenarios depending on the species it dissociates into. So, here's a little outline to help! But first, remember to succeed in this unit it is **incredibly important** to be able to **identify** strong versus weak acids and bases.

There are two types of salts that we will discuss:

1. Metal-Nonmetal Ionic Salts

   These salts are identified by a metal bonded to a nonmetal. There is high bond polarity making the bond easy to break when in solution. You may be familiar with some of these already such as NaCl, NaF, KBr, etc. These typically involve Group IA and IIA metals, halogens, and polyatomic ions. When dissolved, these salts can produce an acidic, basic, or neutral solution.

2. Aminium Salts

   These salts are identified by cationic ammonium, NH_4^+, or an ammonium derivative with an anionic nonmetal component, typically a halogen. Similarly, the bond between the cation and anion is highly polarizable thus easily breaks when in aqueous solutions. Remember a positively charged nitrogen compound (has four bonds) is acidic. It cannot accept more hydrogens and has no lone pair to donate, but it can donate a hydrogen thus making it an acid.

**Forewarning this does not cover ALL salt-based questions that may be asked, but the common scenarios. Please make sure to use the homeworks, PLQs, recitation worksheets, lecture activities, etc. when preparing for Exam 3. 😊

Let's discuss some problems!

Scenario 1: Metal-Nonmetal Ionic Salts

If a question pertains to finding the concentration or pH/pOH of a salt you will need to identify if the ion(s) are acidic, basic, or neutral. To begin you should **break apart the salt**. This is because salts are **strong electrolytes**, so they **completely dissociate**. Let's discuss the three different types of ionic salts:

I. Finding the pH of a **basic** salt

For example, say a question asks about the pH of a 0.8 M NaCN solution (K_b = 1.6 x 10^{-5} at 25°C):

1. Dissociate the salt and identify the ions

NaCN (*aq*)	→	Na^+ (*aq*)	+	CN^- (*aq*)
0.8 M		(Negl.)		0.8 M
SALT		SPEC.		**W. BASE**

2. The weak base is the only component that will affect the pH of the solution. Since this is a weak base, we will need to do a **base hydrolysis** in order to determine the pH. Weak bases **partially dissociate** and exist at **equilibrium**, so we need to use an **ICE table**.

	CN^- (*aq*)	+	H_2O (ℓ)	⇋	OH^- (*aq*)	+	HCN (*aq*)
I	0.8 M		\		0		0
C	−x		\		+x		+x
E	0.8−x		\		x		x

We can now solve for x using the given K_b

$$1.6 \times 10^{-5} = x^2 / 0.8$$

$$x = 0.00358 \text{ M} = [OH^-]$$

3. Calculate the pOH, then subtract from 14 to find the pH

$$pOH = -\log(0.00358) = 2.446$$

pH = 14 – 2.446 = **11.55** (This pH makes sense as it is basic)

Practice below by finding the pH of 0.75 M NaF solution (K_a = 6.81 x 10^{-4}):

II. Finding the pH of an acidic salt

For example, say a question asks about the pH of a 1.25 M $NaHSO_4$ solution through the approximation method (K_a = 1.02 x 10^{-2} at 25°C):

1. Dissociate the salt and identify the ions

$NaHSO_4$ (*aq*)	→	Na^+ (*aq*)	+	HSO_4^- (*aq*)
0.4 M		(Negl.)		0.4 M
SALT		SPEC.		**W. ACID**

2. The weak acid is the only component that will affect the pH of the solution. Since this is a weak acid, we will need to do an **acid hydrolysis** in order to determine the pH. Weak acids **partially dissociate** and exist at **equilibrium**, so we need to use an **ICE table**. **HSO_4^- is NOT a weak base as it is the conjugate base of a STRONG ACID (H_2SO_4). It does have an acidic proton so it can still act as a **weak acid**.

	HSO_4^- (*aq*)	+	H_2O (ℓ)	⇋	H^+ (*aq*)	+	SO_4^{2-} (*aq*)
I	1.25 M		\		0		0
C	−x		\		+x		+x
E	1.25−x		\		x		x

We can now solve for x using the given K_a. While this K_a is larger than 10^{-5} the problem specifies to use the approximation method so we will disregard -x.

$$1.02 \times 10^{-2} = x^2 / 1.25$$

$$x = 0.113 \text{ M} = [H^+]$$

3. Calculate the pH

pH = -log(0.113) = **0.95** (This pH makes sense as it is acidic)

Practice below by finding the pH of 0.75 M $KHSeO_4$ solution. H_2SeO_4 is a STRONG ACID (not one of the 7 you need to know). Please use the approximation method (K_a = 1.26 x 10^{-2}):

III. Finding the pH of a neutral salt

For example, say a question asks about the pH of a 0.5 M NaCl solution:

1. Dissociate the salt and identify the ions

$NaCl$ (*aq*)	→	Cl^- (*aq*)	+	Na^+ (*aq*)
0.4 M		(Negl.)		(Negl.)
SALT		SPEC.		**SPEC.**

2. Both ions are negligible as they are the conjugates of a STRONG ACID (HCl) and a STRONG BASE (NaOH). Neither species will affect the pH of the solution, thus the solution is neutral.

$$pH = 7$$

Practice below by identifying the qualitative pH if the salts below were placed in solution:

Salt	Dissociation Product	pH (> 7, < 7, = 7)
KBr		
SrF_2		
$(CH_2CH_3)_2NH_2Cl$		
$BaCO_3$		
$H_6C_6NH_3I$		

For extra practice, try calculating the pH of each solution if the concentration of salt is 0.65 M. Use the data tables in the eBook to find the dissociation constants:

Scenario 2: Aminium Salts

IV. Finding the pH of an **aminium** salt

For example, say a question asks about the pH of a 0.4 M CH_3NH_3Br solution (K_b = 4.37 x 10^{-4} at 25°C):

1. Dissociate the aminium salt and identify the ions

CH_3NH_3Br (*aq*)	→	Br^- (*aq*)	+	$CH_3NH_3^+$ (*aq*)
0.4 M		(Negl.)		0.4 M
SALT		SPEC.		**W. ACID**

2. The weak acid is the only component that will affect the pH of the solution. Since this is a weak acid, we will need to do an **acid hydrolysis** in order to determine the pH. Weak acids **partially dissociate** and exist at **equilibrium**, so we need to use an **ICE table**.

	$CH_3NH_3^+$ (*aq*)	+	H_2O (ℓ)	⇋	H^+ (*aq*)	+	CH_3NH_2 (*aq*)
I	0.4 M		\		0		0
C	−x		\		+x		+x
E	0.4−x		\		x		x

Before solving for x, we need to calculate the K_a as this is an **acid dissociation**:

$$K_a = K_w / K_b = (1 \times 10^{-14}) / (4.37 \times 10^{-4}) = 2.29 \times 10^{-11}$$

We can now solve for x using the calculated K_a

$$2.29 \times 10^{-11} = x^2 / 0.4$$

$$x = 3.03 \times 10^{-6} \text{ M} = [H^+]$$

3. Calculate the pH

pH = -log(3.03 x 10^{-6}) = **5.52** (This pH makes sense as it is acidic)

Practice below by finding the pH of 0.95 M $(CH_3)_3NHBr$ solution (K_b = 6.46 x 10^{-5}):