Spring 2022

Exam 1 Review

TEXAS
The University of Texas at Austin

Goals for today:

- Review acid and base fundamentals
- Acids and bases presented as chemical reactions
- Quantifying acids and bases
- As seen in the lab: titrations
- Common mistakes
- How to study and exam mindset

Tips

- Structure your notes and studying around the learning outcomes.
  - Class notes, chembook, homework
- Even for calculation problems, try to think about them conceptually.
- Practice, practice, practice

Common Mistakes

- Not appropriately converting calculations to what is asked in the question
- Stoichiometry counts, use the coefficients of your compound and given measurements properly.

Let's start with water:

- Water exhibits unique properties due to its polarity and ability to make hydrogen bonds.
  - It can dissolve other polar substances or ionic compounds by breaking the solutes intermolecular bonds and create new H_2O - solute IMF bonds.
- At room temperature, H_2O spontaneously has moments of dissociation:

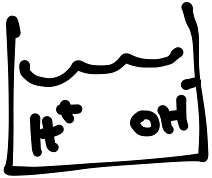
$$H_2O_{(l)} \leftrightarrows H^+_{(aq)} + OH^-_{(aq)}$$

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

Therefore, the H^+ and OH^- concentrations are equal to 1.0×10^{-7} for a neutral solution at room temperature.

- Water also acts as an important solvent as it can dissolve the majority of salts and polar compounds (remember how like dissolves like)
  - When we start adding "stuff" to water, we call this acid-base chemistry.

Acid Base Mindset



- The study of acids and bases revolves around understanding the chemical environment of aqueous solutions associated with hydronium and hydroxide concentrations.
- How do we talk about acids and bases?
  - Quantitatively: ← 0.0001 M
    - The standard units of measurement for acids and bases are [H+], [OH-], pH, and pOH. pH = 3.6
    - pH = -log[H+] pOH = -log[OH-]
  - Qualitatively:
    - For a <u>neutral</u> solution, pH = 7
    - <u>Acidic</u> solutions have pH < 7
    - <u>Basic</u> solutions have pH > 7

Acid: ∗Leaves a proton∗
Base:

pH Calculations

1. strong acid

$$HCl_{(aq)} \rightarrow \underline{H^+_{(aq)}} + Cl^-_{(aq)}$$

$$\underline{pH = -\log(C_A)}$$

2. strong base

$$NaOH_{(aq)} \rightarrow Na^+_{(aq)} + \underline{OH^-_{(aq)}}$$

$$Ca(OH)_2 \rightarrow Ca^{2+} + \underline{2OH^-}_{(aq)}$$

$$pOH = -\log(C_B)$$

$$pH = 14 - pOH$$

$$\underline{pOH = -\log(2 \times C_B)}$$

3. weak acid.

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

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

$$K_a = \frac{x^2}{C_A - x}$$

	HA	⇌ H⁻	+ A
R	HA	H^-	A
I	C_A	0	0
C	$-x$	$+x$	$+x$
E	$C_A - x$	x	x

$$x = [H^+]$$

$$pH = -\log(x)$$

4. weak base

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

$$K_b = \frac{[OH^-][BH^+]}{[B]} = \frac{x^2}{C_B - x}$$

$$x = [OH^-]$$

Quantifying Weak Acids and Bases

- Important Reminder: K_a will get you $[H^+]$, K_b will get you $[OH^-]$. Therefore, K_a corresponds to a weak acid reaction and K_b corresponds to a weak base reaction.

$HA \rightleftharpoons H^+ + A^-$ ⟶ $$K_a = \frac{[H^+][A^-]}{[HA]}$$ ⟶ $$pH = -\log[H^+]$$

$[H^+] = \sqrt{K_a \cdot C_A}$

(Solve using approximation or quadratic)

$B \rightleftharpoons BH^+ + OH^-$ ⟶ $$K_b = \frac{[OH^-][BH^+]}{[B]}$$ ⟶ $$pOH = -\log[OH^-]$$

ignore x when $\frac{C_A}{K_A} \geq 1000$

pH

Weak acid/base

- It is far to say that most questions will use the approximations:

$$[H^+] = \sqrt{C_{HA} \cdot K_a} \qquad [OH^-] = \sqrt{C_B \cdot K_b}$$

- Don't forget that at any time you can convert between different terms:

$$K_w = 1 \cdot 10^{-14} = [H^+][OH^-]$$

$$K_w = K_a K_b$$

$$\frac{K_w}{K_a} = K_b \qquad \frac{K_w}{K_b} = K_a$$

$$HA + H_2O \rightleftharpoons H_3O^+ + A^-$$

Weak Acid Calculations - Percent Ionization

- Percent ionization is an important way of depicting electrolyte strength.
  - K_a is, of course, the more formal way of depicting acid/base strength because % ionization depends on the initial concentration while K_a can represent any concentration of acid.
  - The higher a Ka value, the more the acid dissociates. (Same for K_b but it is base dissociation).

Example: Consider a weak acid, HA, with a % dissociation equal to 10% at a 0.1 M concentration. What is K_a ?

$$[H^+] = (\%\ \text{ionization})(C_{HA})$$

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

HA 0.09

Example: What is the pH of a 0.12 M solution of urea ($(NH_2)_2CO$)? ← B

$K_b = 1.5 \times 10^{-14}$

	$B + H_2O \rightleftharpoons$	BH^+	$+ OH^-$
R	B	BH^+	OH^-
I	0.12	0	0
C	$-x$	$+x$	$+x$
E	$0.12-x$	x	x

K

$K_b = \frac{x^2}{0.12-x}$

$1.5\times10^{-14} = \frac{x^2}{0.12}$

$\sqrt{(1.5\times10^{-14})\times0.12} = x$

$4.24..\times10^{-8} = x$

$pOH = -\log(4.24\times10^{-8})$

$pOH = 7.37$

$pH = 14 - 7.37$

$100 \times \frac{4.24\times10^{-8}}{0.12} = \%$ dissociation

Neutralization Reactions

- The product of a neutralization reaction is a salt. In acid/base chemistry, your salt can be neutral, acidic, or basic depending on the reaction.

1. GENERIC REACTION (very helpful):

   Acid(aq) + Base(aq) $\rightleftharpoons$ Salt(aq) + Water(l)

2. Strong Acid, Strong Base: results in a neutral salt

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

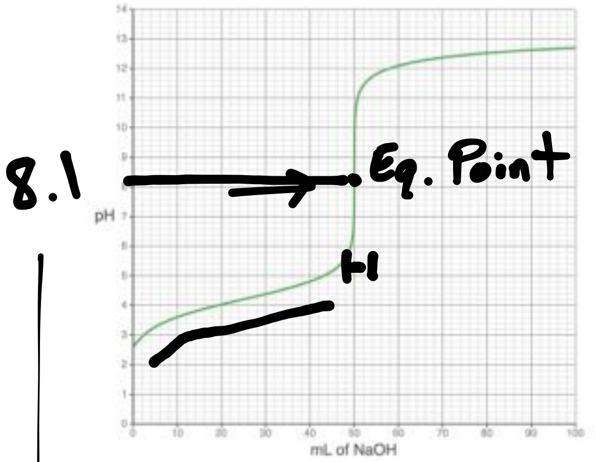
3. Strong Acid to weak base: results in an acidic salt

   $HCl(aq) + NH_3(aq) \rightleftharpoons NH_4Cl(aq) + H_2O(l)$*

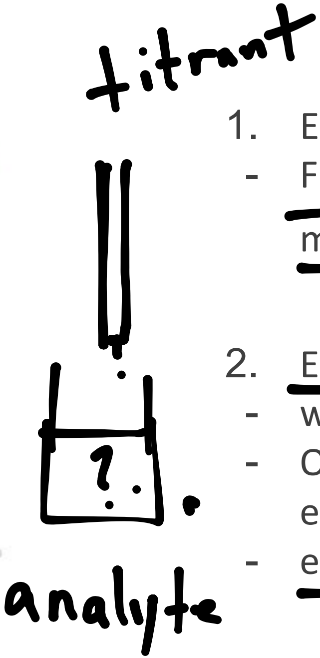
4. Strong base to weak acid: results in a basic salt

   $CH_3COOH(aq) + NaOH(aq) \rightleftharpoons NaCH_3COO(aq) + H_2O(l)$

Titrations



Weak acid + Strong base
Eq. Point > 7



1. Equivalence Point
   - Full neutralization:

     moles titrant = moles analyte

2. End point
   - where your indicator changes color
   - OBSERVATIONAL depiction of equivalence point
   - end point ≠ equivalence point

Using a titration to find an unknown variable:

- During the experiment you add increments of acid/base (titrant) to a solution of base/acid (analyte) until the perfect stoichiometric match is made:
  - $mol_{analye} = mole_{titrant}$
  - $M_{acid}V_{acid} = M_{base}V_{base}$

$M = \frac{mol}{L}$

$M_{H^+}V_{H^+} = M_{OH^-}V_{OH^-}$

0.1M

Example: It takes 28.90 mL of HCl to completely neutralize 50.00 mL of NaOH. What is the NaOH concentration in the initial solution (analyte)?

$M_aV_a = M_bV_b$

$(0.1M)(0.0289L) = x(0.05L)$

$x = 0.0578M$

How to prepare for the exam in the next two nights:

(Did not get to everything in review so :)

* Making the cheat sheet
* Rewrite Notes + Redo HW
  ↳ Given, Solution?
  0.7M HCl pH?
* 0 : chembook, lecture notes, hw, extra practice
  optional : gchem → acid/base → Additional doc

Questions?

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