# SCIENCE

## AP* Chemistry

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## Student Packet

# AP* CHEMISTRY

## Acids, Bases, and Salts

*$K_a$, $K_b$, and $K_{sp}$*

## Student Packet

# ACIDS, BASES, AND SALTS

$K_a$, $K_b$, and $K_{sp}$

## What I Absolutely Have to Know to Survive the AP Exam

The following might indicate the question deals with acids and bases and salts:
pH, pOH, $[H_3O^+]$, $[OH^-]$, strong and weak, salt hydrolysis, solubility product, $K_a$, $K_b$, $K_{sp}$, acid or base dissociation constant, percent ionized, Bronsted-Lowry, Lewis, Arrhenius, hydronium ion, etc…

## Acids and Bases: *By Definition!*

Arrhenius

- Acids: Hydrogen ion, $H^+$, donors
- Bases: Hydroxide ion, $OH^-$ donors

$$HCl \rightarrow H^+ + Cl^-$$
$$NaOH \rightarrow Na^+ + OH^-$$

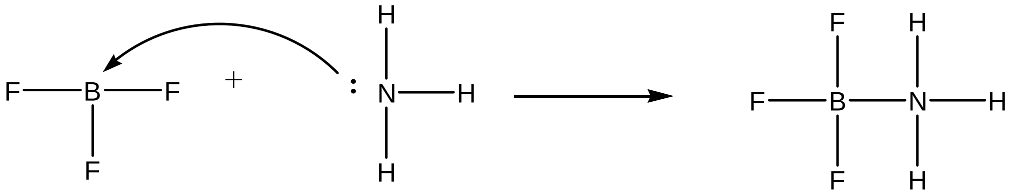
Bronsted-Lowry

- Acids: proton donors ( lose $H^+$)
- Bases: proton acceptors (gain $H^+$)

$$NH_4^+ + H_2O \rightleftharpoons H_3O^+ + NH_3$$
$$HNO_3 + H_2O \leftrightarrow H_3O^+ + NO_3^-$$

Lewis

- Acids: electron pair acceptors
- Bases: electron pair donors
- Explains all traditional acids/bases + a host of coordination compounds. $BF_3$ is famous!



## Hydrogen Ion, Hydronium Ion, Which is it!!!?

**Hydronium** – $H^+$ riding piggy-back on a water molecule; water is polar and the + charge of the naked proton is greatly attracted to “Mickey's chin” (i.e. the oxygen atom)

- $H_3O^+$ “Anthony”
- $H^+$ “Tony”
- Often used interchangeably in problems; if $H_3O^+$ is used be sure water is in the equation!

## Bronsted–Lowry and Conjugate Acids and Bases: *What a Pair!*

Acid and conjugate base pairs differ by the presence of one $H^+$ ion.

$$HC_2H_3O_2 + H_2O \rightarrow H_3O^+ + C_2H_3O_2^-$$

- $HC_2H_3O_2$ is the acid; thus $C_2H_3O_2^-$ is its conjugate base (what remains after the $H^+$ has been donated to the $H_2O$ molecule)
- $H_2O$ behaves as a base in this reaction. The hydronium ion is its conjugate acid (what is formed after the $H_2O$ accepts the $H^+$ ion)

$$NH_3 + H_2O \rightarrow NH_4^+ + OH^-$$

- $NH_3$ is the base; thus $NH_4^+$ is it's conjugate acid (what is formed after the $NH_3$ accepts the $H^+$ ion)
- $H_2O$ behaves as an acid in this reaction. The hydroxide ion is its conjugate base (what remains after the $H^+$ has been donated to $NH_3$)

Understanding conjugate acid/base pairs is very important in understanding acid-base chemistry; this concepts allows for the understanding of many complex situations (buffers, titrations, etc…)

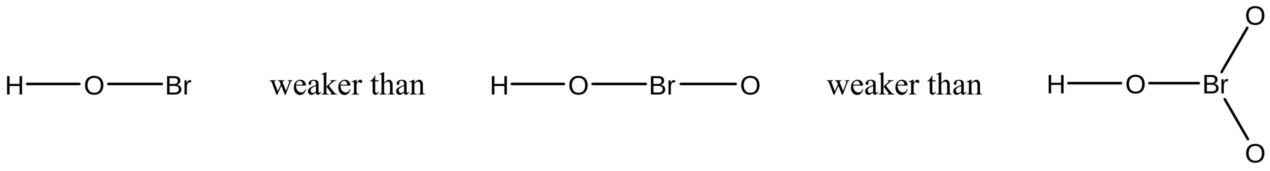
Important Notes…
Amphiprotic/amphoteric--molecules or ions that can behave as EITHER acids or bases; water, some anions of weak acids, etc… fit this bill.
Monoprotic – acids donating one $H^+$
Diprotic – acids donating two $H^+$
Polyprotic – acids donating 3+ $H^+$

Regardless, always remember: **ACIDS ONLY DONATE ONE PROTON AT A TIME!!!**

## Ionization: *That's What it's All About!*

Relative Strengths
A strong acid or base ionizes completely in aqueous solution (100% ionized [or very darn close to it!])

  - The equilibrium position lies far, far to the right (products)…
  - Since a strong acid/base dissociates into the ions, the concentration of the $H^+/OH^-$ ion is equal to the original concentration of the acid/base respectively.
  - They are strong electrolytes.
- **Do Not** confuse concentration (*M* or mol/L) with strength!
  - Strong Acids
    - Hydrohalic acids: HCl, HBr, HI; Nitric: $HNO_3$; Sulfuric: $H_2SO_4$; Perchloric: $HClO_4$
  - Oxyacids! More oxygen atoms present, the stronger the acid WITHIN that group. The $H^+$ that is "donated" is bonded to an oxygen atom. The oxygen atoms are highly electronegative and are pulling the bonded pair of electrons AWAY from the site where the $H^+$ is bonded, "polarizing it", which makes it easier [i.e. requires ***less energy***] to remove – thus the stronger the acid!



- Strong Bases
  - Group IA and IIA (1 and 2) metal hydroxides; be cautious as the pour solubility of $Be(OH)_2$ and $Mg(OH)_2$ limits the effectiveness of these 2 strong bases.
  - IT'S A 2 for 1 SALE!! (with the Group 2 (IIA) ions), i.e. 0.10*M* $Ca(OH)_2$ is 0.20*M* $OH^-$

## Ionization: *That's What it's All About! con't.*

***Where the fun begins…***

A weak acid or base does not completely ionize (usually < 10%)

  - They are weak electrolytes.
  - The equilibrium position lies far to the left (reactants)…
  - The $[H^+]$ is *less than* the acid concentration – thus to calculate this amount and the resulting pH you must return to the world of EQUILIBRIUM Chemistry!
- The vast majority of acid/bases are weak. Remember, ionization not concentration!!!!
- Acids and Bases ionize one proton (or $H^+$) at a time!
- For weak acid reactions: $HA + H_2O \rightarrow H_3O^+ + A^-$

$$K_a = \frac{[H_3O^+][A^-]}{[HA]} = <1$$

- For weak base reactions: $B + H_2O \rightarrow HB^+ + OH^-$

$$K_b = \frac{[HB^+][OH^-]}{[B]} = <1$$

- $K_a \times K_b = 1\times10^{-14}$ **Very important when dealing with SALT pH**

- MONUMENTAL CONCEPT…
  - If $[HA]_o$ is $1000\times K_a$ or greater, you can use the initial concentration of the acid/base as the equilibrium concentration –WITHOUT subtracting the amount of weak acid ionized (it is just mathematically insignificant). Hallelujah!!! This just eliminated the Quadratic Equation (which is never needed on the AP exam).

- MUCHO time saver!!! Since acids ionize 1 $H^+$ at a time the $[H_3O^+] = [A^-]$ thus all you need to know is…

$$K_a = \frac{[x][x]}{[M_o]} \text{where x} = [H_3O^+] \quad \text{or} \quad K_b = \frac{[x][x]}{[M_o]} \text{where x} = [OH^-]$$

- Percent Ionization
  - Often will be asked to determine how ionized the weak acid or base is…
  - $\% = \frac{[x]}{[M_o]}\times100 \quad \text{where x} = [H_3O^+] \text{or} [OH^-]$

## Salts and pH: *It's all about Hydrolysis*

A salt is the PRODUCT of an acid base reaction; how the salt affects the pH of the solution depends on whether or not the salt ions will hydrolyze water to any significant extent…

- Cation: $M^+$ hydrolyzes $H_2O$ to produce MOH and $H_3O^+$ ions
  - $M^+ + H_2O \rightarrow H_3O^+ + MOH$
    - If the MOH is a strong base this reaction WILL NOT occur (strong bases ionize 100%). Therefore NO additional hydronium ion is formed in the solution and this salt ion has NO effect on the pH of the solution!
    - If the MOH is a weak base this reaction WILL occur; and the additional hydronium ion will lower the pH of the solution!!
- Anion: $A^-$ hydrolyzes $H_2O$ to produce HA and $OH^-$ ions
  - $A^- + H_2O \rightarrow HA + OH^-$
    - If the HA produced is a strong acid this reaction WILL NOT occur (strong acids ionize 100%). Therefore NO additional hydroxide ion is formed in the solution and this salt ion has NO effect on the pH of the solution!
      - If the HA is a weak acid this reaction WILL occur; and the additional hydroxide ion will increase the pH of the solution!
- So… an easy way to tell, you ask…?
  - Ask yourself, which acid and which base reacted….were they strong or weak?? Strong wins!
    1. A salt such as $NaNO_3$ gives a neutral solution—SA ($HNO_3$) & SB (NaOH)
       – Neither salt can hydrolyze water because it would result in the formation of a strong acid and base
    2. $K_2S$ should be basic—SB (KOH) & WA ($H_2S$)
       – Only the weak can effectively hydrolyze water so the $S^{2-}$ ion reacts with water to produce $H_2S$ and $OH^-$ ions; therefore the salt solution is basic…
    3. $FeCl_2$ should be weakly acidic—since WB ($Fe(OH)_2$) & SA (HCl)
       – Only the weak can effectively hydrolyze water so the $Fe^{2+}$ ion reacts with water to produce $Fe(OH)_2$ and $H_3O^+$ ions; therefore the salt solution is slightly acidic…

Calculating the pH of a salt solution

- A salt in solution is either behaving as a weak acid or a weak base.
- Remember, this means they do not ionize much. So treat them just like any other weak acid or base.
- The MAJOR DIFFERENCE; no $K_a$ or $K_b$ value will be provided. Its conjugate acid or base $K$ will be given instead.
    - Example, the salt $NaC_2H_3O_2$; the $K_a$ for $HC_2H_3O_2$ will be provided. The $K_b$ has to be calculated before any pH calculations can be attempted.
      - $K_a \times K_b = 1 \times 10^{-14}$ Very important!
    - Then solve below to find the pH…

$$K_b = \frac{[x][x]}{[M_o]} \text{ where } x = [OH^-]$$

## Solubility Product: *It's $K_{sp}$!*

When solids are first added to water no salt ions are present. As dissolution occurs, the ions dissolve in the water until equilibrium is established and a saturated solution is formed. The extent to which a substance dissolves in the solvent is the solubility. The equilibrium constant, $K_{sp}$, is the product of all ions in solution.

$$PbI_2(s) \rightleftharpoons Pb^{2+}(aq) + 2\,I^{-}(aq) \quad K_{sp} = [Pb^{2+}][I^{-}]^2$$

- Note: solids don't appear in the $K$ expression
- Be aware of the coefficients
    - the 1:2 ratio is important when determining concentrations of the ions present
    - the 2 results in an expression of $[I^{-}]^2$

A precipitate is an insoluble compound formed when two soluble solutions are mixed. To determine whether a precipitate will form the reaction quotient will need to be calculated.

- $Q < K$ unsaturated solution is formed as the reaction has not reach the equilibrium position; no ppt
- $Q = K$ a saturated solutions is made; no ppt
- $Q > K$ a supersaturated solution is formed the concentration of ions are greater than allowed; a ppt will form

## Acids, Bases and Salts Cheat Sheet

| **Relationships** | |
|---|---|
| Equilibrium Expression<br>$K_a = \dfrac{[H_3O^+][A^-]}{[HA]} = <1$<br>$K_a = \dfrac{[x][x]}{[M_o]}$ where $x = [H_3O^+]$ | $K_b = \dfrac{[HB^+][OH^-]}{[B]} = <1$<br>$K_b = \dfrac{[x][x]}{[M_o]}$ where $x = [OH^-]$ |
| $\% = \dfrac{[x]}{[M_o]} \times 100$ where $x = [H_3O^+]$ or $[OH^-]$ | |
| $K_a \times K_b = 1\times10^{-14}$ | $pH = -\log[H^+]$ |
| $pH + pOH = 14$ | $pOH = -\log[OH^-]$ |
| $K_w = [H_3O^+][OH^-] = 1.0\times10^{-14}$ @ 25°C | Acidic $pH < 7$ $[H_3O^+] > [OH^-]$ |
| Neutral $pH = 7$ $[H_3O^+] = [OH^-]$ | Basic $pH > 7$ $[H_3O^+] < [OH^-]$ |
| Conjugate acids and bases | $K_{sp} = [M^+]^x[A^-]^y$ $MA(s) \rightleftharpoons xM^+(aq) + yA^-(aq)$ |
| **Connections** | |
| Equilibrium | Buffers and Titrations |
| Precipitation and Qualitative Analysis | Bonding and Lewis Structures – justify oxyacid strengths |
| **Potential Pitfalls** | |
| Coefficients in $K_{sp}$ problems | $K_a$ or $K_b$ with salt pH |
| Weak acids and bases – be sure you know $K_a$ or $K_b$; solving a $K_b$ problem gives $[OH^-]$ thus you are finding the pOH! | Weak is about IONIZATION not CONCENTRATION |

## Free Response

$$NH_3(aq) + H_2O(\ell) \rightleftharpoons NH_4^+(aq) + OH^-(aq) \qquad K_b = 1.80 \times 10^{-5}$$

1. Ammonia reacts with water as indicated in the reaction above.

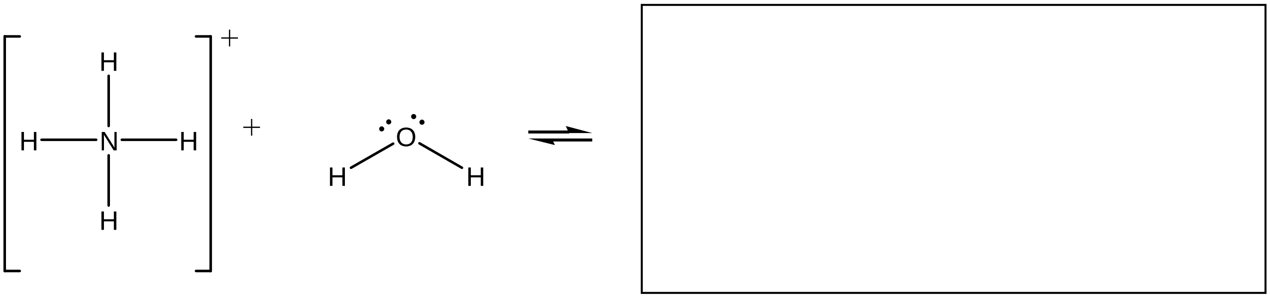
   (a) Write the equilibrium constant expression for the reaction represented above.

   (b) Calculate the pH of a 0.150 *M* solution of $NH_3$

   (c) Determine the percent ionization of the weak base $NH_3$.

   (d) Calculate the hydronium ion, $H_3O^+$, concentration in the above solution. Be sure to include units with your answer.

Ammonium ions react with water to form a solution with a pH of 4.827.



   (e) Complete the reaction above by drawing the *complete* Lewis structures for *both* products of the reaction.

   (f) Determine the number of moles ammonium ions in 250 mL of the above solution.

## Free Response

2. Hydrocyanic acid, HCN, is a weak acid. The acid dissociation constant for HCN is $6.2 \text{ x } 10^{-10}$. The equilibrium expression for the reaction is shown below.

$$K_a = \frac{[H_3O^+][CN^-]}{[HCN]} = 6.2\times10^{-10}$$

(a) Write the equation for the dissociation of HCN in water.

(b) Calculate the hydronium ion, $H_3O^+$ concentration of a 0.250 *M* solution of HCN.

When lithium cyanide, LiCN, is dissolved in water according to the equation below the resulting solution is basic.

$$LiCN(s) \rightarrow Li^+(aq) + CN^-(aq)$$

(c) Write a net ionic equation that illustrates why the resulting solution is basic.

(d) Calculate the equilibrium constant for reaction in part c.

The overall dissociation reaction for carbonic acid, $H_2CO_3$, is shown below.

$$H_2CO_3 \rightleftharpoons CO_3^{2-} + 2H^+ \qquad K_{eq} = 2.4\times10^{-17}$$

(e) Write a balanced chemical equation showing both the first and second dissociation reactions for carbonic acid, $H_2CO_3$.

(f) The dissociation constant for the second dissociation, $K_{a_2} = 5.6\times10^{-11}$. Calculate the equilibrium constant for the first dissociation, $K_{a_1}$.

## Free Response

3. Solid potassium lactate, $KC_3H_5O_3$, is dissolved in enough water to make a 0.25 *M* solution. The solution has a pH of 12.25 at 25°C.

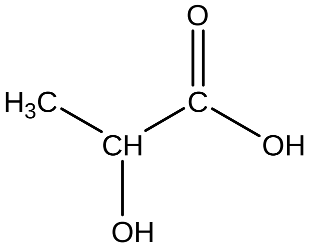
   (a) Explain why the solution produced by dissolving $KC_3H_5O_3$ is basic. Use an equation as a part of your justification.

   (b) Calculate the hydroxide ion concentration of the solution described above.

   (c) Calculate the value of the equilibrium constant for the reaction in part (a)

   (d) Calculate the acid dissociation constant, $K_a$, for lactic acid, $HC_3H_5O_3$.

   (e) The structural formula for lactic acid is shown below. Circle the hydrogen lost when $HC_3H_5O_3$ behaves as an acid.



   (f) A student needs to dilute a 0.25 *M* solution of $KC_3H_5O_3$ to make 50.0 mL of 0.10 *M* potassium lactate solution. Describe the steps needed to prepare the new solution. Be sure to include appropriate lab equipment.

## Free Response

4. Answer the following questions about compounds of strontium.

In a saturated solution of strontium phosphate, $Sr_3(PO_4)_2$, the concentration of $Sr^{2+}$ is $7.5\times10^{-7}$ *M* at 25°C.

(a) Write the balanced equation for the dissolution of solid $Sr_3(PO_4)_2$ in water.

(b) Calculate the value of the solubility product constant.

(c) Will a precipitate form when 20.0 mL of 0.0015 M strontium acetate and 30.0 mL of 0.0050 M lithium phosphate are mixed? Justify your answer with appropriate calculations.

The $K_{sp}$ for strontium fluoride, $SrF_2$, is $7.9\times10^{-10}$ at 25°C.

(d) Calculate the molar solubility of a saturated solution of strontium fluoride.

(e) A small sample of the saturated solution of $SrF_2$ was allowed to evaporate. Would the concentration of $Sr^{2+}$ ions in the solution increase, decrease or remain the same? Justify your answer.

(f) At 35°C, the amount of strontium fluoride present at the bottom of the flask increases. Is this dissolution process endothermic or exothermic? Justify your answer.

## Multiple Choice

<u>Questions 1–4</u> refer to the following answer choices.  All solutions are 1.0 *M*.

(A) HCl
(B) LiCl
(C) $HClO_4$
(D) $NaC_2H_3O_2$
(E) $BF_3$

1. Can be used as a Lewis acid

2. Would form a basic solution when dissolved in water

3. Forms the solution with the lowest pH

4. Can produce a buffer with a pH < 7 when combined with $HC_2H_3O_2$

5. Which of the following can act as an amphoteric species?

I. $SO_4^{2-}$
II. $HC_2O_4^{-}$
III. $NH_4^{+}$

(A) I only
(B) II only
(C) I and II only
(D) II and III only
(E) I, II, and III

6. Estimate the pH of a $1.0\times10^{-5}$ M solution of the weak base aniline, $C_6H_5NH_2$.  $K_b = 4.0 \times 10^{-10}$

(A) 11
(B) Between 8 and 11
(C) 8
(D) Between 3 and 8
(E) 3

7. All the following statements are true about the reaction below EXCEPT:

$$650\ \text{kJ} + HCO_3^-(aq) + HC_2O_4^-(aq) \rightleftharpoons CO_3^{2-}(aq) + H_2C_2O_4(aq) \qquad K_c < 1$$

(A) $HC_2O_4^-$ is a weaker base than $CO_3^{2-}$.
(B) $HCO_3^-$ and $H_2C_2O_4$ are both acting as acids in the reaction.
(C) $HCO_3^-$ is an amphoteric species.
(D) The reaction lies far to the right (favors products).
(E) Lowering the temperature of the solution will decrease the value of $K$.

8. A solution of a monoprotic weak acid, HA, has a pH of 5.00. Calculate the acid dissociation constant for the weak acid if the solution has a molar concentration of 0.50 $M$.

(A) $2.0 \times 10^{-3}$
(B) $1.0 \times 10^{-5}$
(C) $2.0 \times 10^{-6}$
(D) $2.0 \times 10^{-9}$
(E) $2.0 \times 10^{-10}$

9. The solubility product constant of $Ag_2SO_4$ is $1.2 \times 10^{-5}$ at 25°C. Which expression can be used to calculate the solubility for$Ag_2SO_4$?

(A) $\left[\dfrac{1.2 \times 10^{-5}}{2}\right]^{\frac{1}{3}}$

(B) $\left[1.2 \times 10^{-5}\right]^{\frac{1}{2}}$

(C) $\left[1.2 \times 10^{-5}\right]^{\frac{1}{3}}$

(D) $\left[\dfrac{1.2 \times 10^{-5}}{4}\right]^{\frac{1}{2}}$

(E) $\left[\dfrac{(1.2 \times 10^{-5})}{4}\right]^{1/3}$

10. At a certain temperature, a saturated solution of silver phosphate has a $[Ag^+]$ of $3.0 \times 10^{-5}$ M. What is the solubility product constant, $K_{sp}$, for silver phosphate, $Ag_3PO_4$?

(A) $2.7 \times 10^{-21}$
(B) $2.7 \times 10^{-20}$
(C) $2.7 \times 10^{-19}$
(D) $8.1 \times 10^{-10}$
(E) $9.0 \times 10^{-10}$

# AP* CHEMISTRY

# Buffers and Titrations

## *As Easy As It Ever Gets*

## Student Packet

# BUFFERS AND TITRATIONS

*As Easy As It Ever Gets*

## What I Absolutely Have to Know to Survive the AP Exam

The following might indicate the question deals with buffers and/or titrations:
buffer solution, common ion, conjugate base/acid, equivalence point, ½ equivalence point, end-point, weak acid/base "reacted with a strong..", titrant, buret, analyte, etc…

## Buffers: *All about those "common ions"*

IS: A buffer is a solution of a weak acid or base and its salt [which is its conjugate].

DOES: A buffer resists a change in pH.

HOW IT WORKS: Since a buffer consists of both an acid or base and its conjugate, which differ by an $H^+$, both a weak acid and a weak base are present in all buffer solutions.

In order to understand buffers and buffer problems we must be proficient with the following…

- MUST know the conjugate A/B concept to be successful at buffer problems. This means understanding that HA has $A^-$ as it's conjugate and that NaA is not only a salt, but that any soluble salt releases the common ion, $A^-$
- $pH = -\log [H^+]$
- $K_a \times K_b = K_w$
- Any titration involving a weak A/B, from the first drop to the last, before equivalence, is a buffer problem. If HA is titrated with NaOH, as soon as the first drop splashes into the container, I've added $A^-$ ions to the solution and the presence of both HA and $A^-$ constitutes a buffer solution.

## Buffers: *Calculating pH*

Solving Buffer Problems: the one and only, handy dandy, all I'm ever going to need equation…

$$\left[H_3O^+\right] = K_a \frac{[\text{Acid}]}{[\text{Base}]}$$

Where:
$K_a$ – is the acid dissociation constant
[Acid] – is the concentration of the substance behaving as the **weak** acid in the buffered solution

- Either the weak acid or the salt of the weak base

[Base] – is the concentration of the substance behaving as the **weak** base in the buffered solution

- Either the weak base or the salt of the weak acid

To Solve:

- MUST recognize who's who in the buffered and then plug into the equation
- If the buffer is a weak acid, HA, and its conjugate salt, $A^-$ then HA is the acid and $A^-$ is the base.
- If the buffer is a weak base, WB, and its conjugate salt, $B^+$, then WB is the base and $B^+$ is the acid.
  - You still use the $K_a$; thus you must calculate it first using the $K_b$ provided for the weak base
- If concentrations are given plug and chug
- If concentrations of separate solutions are given with volumes and then the two are added together you must recalculate the "new concentrations" due to dilution.

## Buffers: *Calculating pH Con't*

Example:

Calculate the pH of a solution containing 0.75 *M* lactic acid, $HC_3H_5O_3$ ($K_a = 1.4\times10^{-4}$), and 0.25 *M* sodium lactate, $NaC_3H_5O_3$.

- Lactic acid is the weak ACID and lactate ion is its conjugate BASE

$$\left[H_3O^+\right] = K_a\frac{[\text{Acid}]}{[\text{Base}]} = (1.4\times10^{-4})\frac{[0.75]}{[0.25]} = 4.2\times10^{-4}$$

$$\text{pH} = -\log(4.2\times10^{-4}) = 3.38$$

Example:

A buffered solution contains 0.250 *M* $NH_3$ ($K_b = 1.80\times10^{-5}$) and 0.400 *M* $NH_4Cl$. Calculate the pH of this solution.

- $NH_3$ is the weak BASE and $NH_4^+$ is its conjugate ACID
- Calculate $K_a$!

$$K_a \times K_b = 1.00\times10^{-14}$$

$$K_a = \frac{1.00\times10^{-14}}{1.8\times10^{-5}} = 5.56\times10^{-10}$$

$$\left[H_3O^+\right] = K_a\frac{[\text{Acid}]}{[\text{Base}]} = (5.56\times10^{-10})\frac{[0.400]}{[0.250]} = 8.89\times10^{-10}$$

$$\text{pH} = -\log(8.89\times10^{-10}) = 9.051$$

Example:

Calculate the pH of a solution prepared by mixing 20.0 mL of 0.300 *M* acetic acid, $HC_2H_3O_2$, with 20.0 mL of 0.350 *M* $NaC_2H_3O_2$. $K_a$ for acetic acid is $1.80\times10^{-5}$

- Acetic acid is the weak ACID and acetate ion is its conjugate BASE
- Recalculate molarities OR better yet just calculate the number of moles (volumes cancel)….

$$20.0\,\text{mL}\times0.300M = 6.00\,\text{m}M\ \text{acid}$$

$$20.0\,\text{mL}\times0.350M = 7.00\,\text{m}M\ \text{conjugate base}$$

$$\left[H_3O^+\right] = K_a\frac{[\text{Acid}]}{[\text{Base}]} = (1.80\times10^{-5})\frac{[6.00]}{[7.00]} = 1.54\times10^{-5}$$

$$\text{pH} = -\log(1.54\times10^{-5}) = 4.812$$

## Buffer Attack: *Who Defends?*

The purpose of a buffer solution is to resist large swings in pH. If a small amount of strong acid is added to the buffer, there is a base component ready and waiting to neutralize the "invader". Same goes for the addition of a small amount of strong base, there is an acid component ready and waiting to neutralize it…

If strong acid is added to the buffer, simply ***add*** the amount of acid added to the numerator (acid) AND subtract the same quantity from the denominator (base). Why? If strong acid is added to a buffer solution the weak base component is self-sacrificing and neutralizes the acid – thus the amount present is reduced by the amount of strong acid present in the invasion. The reaction between the weak base component of the buffer and the strong acid eliminates the strong acid and produces more weak acid.

Strong Acid Attack: $HCl + A^- \rightarrow HA + Cl^-$

Example:
Calculate the pH of a solution containing 0.75 *M* lactic acid, $HC_3H_5O_3$ ($K_a = 1.4\times10^{-4}$), and 0.25 *M* sodium lactate, $NaC_3H_5O_3$ after the addition of 0.10 *M* HCl.

$$\left[H_3O^+\right] = K_a\frac{[\text{Acid}]}{[\text{Base}]} = (1.4\times10^{-4})\frac{[0.75+0.10]}{[0.25\text{-}0.10]} = 7.9\times10^{-4}$$
$$\text{pH} = -\log(7.9\times10^{-4}) = 3.10$$

If strong base is added, simply add the base to the denominator and subtract from the numerator. Just the exact opposite occurs; the acid component reacts with the base and converts the strong base to more weak base.

Strong Base Attack: $NaOH + HA \rightarrow H_2O + Na^+ + A^-$

Example:
What is the pH of the final solution after the addition of 125 mL of 0.500 *M* NaOH to 250 mL of a solution that is 0.550 M acetic acid, $HC_2H_3O_2$ and 0.450 *M* sodium acetate, $NaC_2H_3O_2$?

- Acetic acid is the weak ACID and acetate ion is its conjugate BASE; NaOH is the INVADER
- Recalculate molarities OR better yet just calculate the number of moles (volumes cancel)….

$$125\,\text{mL}\times0.500M = 62.5\,\text{m}M\text{ NaOH "the invader"}$$
$$250\text{ mL}\times0.550M = 137.5\,\text{m}M\text{ HA "the weak acid"}$$
$$250\text{ mL}\times0.450M = 112.5\,\text{m}M\ A^-\text{ "the conjugate base"}$$
$$\left[H_3O^+\right] = K_a\frac{[\text{Acid}]}{[\text{Base}]}$$
$$\left[H_3O^+\right] = (1.80\times10^{-5})\frac{[137.5-62.5]}{[112.5+62.5]} = 7.71\times10^{-6}$$
$$\text{pH} = -\log(7.71\times10^{-6}) = 5.113$$

CAUTION:
If the amount of invader exceeds the amount present in the buffer then the invader has overrun the buffer!

- In this case calculate how much strong "invader" remains after all of the buffer is used
- Recalculate the molar concentration of the remaining "invader"; this is a strong acid or base. Calculate the pH as with any strong acid or base. $\text{pH} = -\log[H^+]$

## Preparing a Buffer: *It's all about $K_a$ and the Ratio!*

If asked to describe how to prepare a buffer of a certain pH… think $\left[H_3O^+\right] = K_a \frac{[\text{Acid}]}{[\text{Base}]}$

- Use 0.10 M to 1.0 M solutions of reagents & choose an acid whose $K_a$ is near the $[H_3O^+]$ concentration we want. Or its $pK_a$ should be as close to the pH desired as possible. Adjust the ratio of $\frac{[\text{Acid}]}{[\text{Base}]}$ to fine tune the pH.
- It is the ***relative*** # of moles (not *M*) of acid/CB or base/CA that is important since they are in the same solution and share the same solution volume.
  - This allows companies to make concentrated versions of buffer and let the customer dilute--this will not affect the # of moles present--just the dilution of those moles. So on the EXAM, never fall for the trick of "what happens to the pH of the buffer solution after the addition of 250 mL of water. Yes the [Acid] and [Base] are diluted but the $\frac{[\text{Acid}]}{[\text{Base}]}$ ratio is the same and that is what matters in a buffer!

IMPORTANT NOTE:

- When equal concentrations of Acid and Base are present [which occurs at the ½ equivalence point of a titration] the ratio of acid to base equals ONE and therefore, the pH = $pK_a$.

## Acid-Base Titrations: *It Depends on Where You Are!*

There are 3 types of titrations you will be dealing with.

1. Strong Acid and Strong Base
   - Net ionic reaction: $H^+ + OH^- \rightarrow H_2O$
   - Only when the acid AND base are both strong is the pH at the equivalence point 7. [Any other conditions and you get to do a nasty equilibrium problem.] It's really a stoichiometry problem with a limiting reactant. The "excess" is responsible for the pH
   - Just deal with what is left over. Be sure you are in moles if additive volumes are given and recalculate the *M* after doing the stoichiometry work
   - Use $\text{pH} = -\log [H^+]$ to determine the pH of the solution.

2. Weak Acid and Strong Base
   - The equivalence point > pH 7
   - 4 ZONES OF INTEREST ALONG A TITRATION CURVE for a Weak Acid and Strong Base Titration

   ZONE 1: The pH before the titration begins.
   - This is simply a weak acid problem.

$$K_a = \frac{x^2}{M} \quad \text{where } x = [H_3O^+]$$

   ZONE 2: The pH during the titration but before the equivalence point.
   - Once the titration begins the weak acid has reacted with the strong base to produce salt and water. The salt is the conjugate base of the weak acid. This is a BUFFER problem!!!

$$HA + OH^- \rightarrow H_2O + A^-$$

   - Remember that the amount of strong base added = amount of weak acid reacted = amount of salt made
   - Plug in the amount of weak acid left and the amount of salt made into $\left[H_3O^+\right] = K_a \frac{[\text{Acid}]}{[\text{Base}]}$

   CAUTION! TIME SAVER: At the ½ equivalence point [Weak Acid] = [Conjuagte Base formed] thus $[H_3O^+] = K_a$

## Acid-Base Titrations: *It Depends on Where You Are!* *Con't*

ZONE 3: The pH at the Equivalence Point

- At the equivalence point of the titration ALL of the WEAK ACID has been reacted with the STRONG BASE – the only species left are $H_2O$ and the Conjugate Base of the Weak Acid (the SALT).
- The pH is based only on the salts hydrolysis properties The *M* of the salt at equivalence point must be calculated.
- Remember that the amount of strong base added = amount of weak acid reacted = amount of salt made
- Thus the molarity of the salt in solution is:

$$M = \frac{\text{moles of salt formed}}{\text{total L of solution at the equivalence pt.}}$$

- Must use the $K_a$ of the weak acid to calculate the $K_b$ of the salt and then use

$$K_b = \frac{x^2}{M} \quad \text{where } x = [\text{OH}^-]\text{; find the pOH then convert to pH}$$

ZONE 4: Beyond the Equivalence Point

- It's all about the excess!
- Calculate the amount of excess strong base added beyond the equivalence point and then recalculate its *M*. This is a strong base so $\text{pOH} = -\log[\text{OH}^-]$

3. Weak Base and Strong Acid
   - Just like the Weak Acid/Strong Base above, just flip flop the process…
   - The equivalence point < pH 7

IMPORTANT NOTES:

There is a distinction between the equivalence point and the end point. The end point is when the indicator changes color. If you've made a careful choice of indicator, the equivalence point (the number of moles of acid = number of moles of base) and the end point of the titration will be achieved at the same time.

- When choosing an indicator, determine the pH at the equivalence pt. Then, pick the indicator that has a $pK_a$ close to the pH at the equivalence point.

# Buffers and Titrations Cheat Sheet

## Relationships

$$\left[H_3O^+\right] = K_a \frac{[\text{Acid}]}{[\text{Base}]}$$

$$K_a = \frac{x^2}{M} \quad \text{where } x = [H_3O^+]$$

$$K_b = \frac{x^2}{M} \quad \text{where } x = [OH^-]$$

$$K_a \times K_b = 1\times10^{-14}$$

$pH = -\log[H^+]$

$pOH = -\log[OH^-]$

## <u>4 ZONES OF INTEREST</u> ALONG A TITRATION CURVE for a Weak Acid and Strong Base

1. Initial pH – simply a weak acid/weak base problem $K_a = \frac{x^2}{M}$ where $x = [H_3O^+]$ or $K_b = \frac{x^2}{M}$ where $x = [OH^-]$
2. During the titration – BUFFER, BUFFER, IT'S A BUFFER! $\left[H_3O^+\right] = K_a \frac{[\text{Acid}]}{[\text{Base}]}$ where the weak and its conjugate salt are what we need to know about.
3. Equivalence Pt. – only salt and water left; the salt is either a weak base (if a weak acid was titrated) or a weak acid (if a weak base was titrated); work it as such: $K_a = \frac{x^2}{M}$ where $x = [H_3O^+]$ or $K_b = \frac{x^2}{M}$ where $x = [OH^-]$ BUT you have to convert $K_a$ to $K_b$ or vise versa!!!
4. Beyond the Equivalence Pt. – This is a stoichiometry problem; find out how much STRONG (acid or base) is left after equivalence, recalculate its molarity (remember the volume increased during the titration), and use $pH = -\log[H^+]$ or $pOH = -\log[OH^-]$

Be aware of where the volume of STRONG added in the titration is ½ way to equivalence point!

- Here $\left[H_3O^+\right] = K_a$ because the $\frac{[\text{Acid}]}{[\text{Base}]}$ is 1 (same amount of weak acid (base) and its conjugate are present. HUGE time saver and easy way to find the $K_a$ or $K_b$ of the weak acid or base.
- In a titration problem ALWAYS know where the equivalence point is!!!!!!

BUFFERS: Only need to use the equation: $\left[H_3O^+\right] = K_a \frac{[\text{Acid}]}{[\text{Base}]}$; read carefully and know who is the weak acid (base) and who is the conjugate salt (it plays the other role!).

Don't be tricked – NEVER select a strong acid or base as a part of a buffer solution; it HAS TO BE WEAK

When a buffered solution is "attacked" remember:

- If the attacker is a strong acid; the base component of the buffer "sacrifices" and reacts with the invading acid, converting it to water and more weak acid – thus resisting large decreases in the pH
- If the attacker is a strong base; the acid component of the buffer "sacrifices" and reacts with the invading base, converting it to water and more weak base – thus resisting large increases in the pH

## Connections

| | |
|---|---|
| Equilibrium | Stoichiometry |

## Potential Pitfalls and Things to Watch For

| | |
|---|---|
| Knowing when you need $K_a$ or $K_b$ | Recognizing the weak organic bases |
| BE AWARE of when a weak acid (base) is present with its conjugate salt – HELLO! YOU ARE IN THE BUFFER ZONE!!!! | |

## Free Response

$$NH_3(aq) + H_2O(\ell) \rightleftharpoons NH_4^+(aq) + OH^-(aq) \qquad K_b = 1.80 \times 10^{-5}$$

1. Ammonia reacts with water as indicated in the reaction above.

   (a) Write the equilibrium constant expression for the reaction represented above.

   (b) Calculate the pH of a 0.150 $M$ solution of $NH_3$

A 100 mL sample of the 0.150 $M$ solution of $NH_3$ is titrated with 0.100 $M$ HCl

   (c) Determine the volume of HCl required to reach the equivalence point.

   (d) Calculate the pH of the solution at the equivalence point of the titration.

To the above solution (at the equivalence point) 250 mL of 0.060 $M$ $Pb(NO_3)_2$ are added. After mixing, the following concentrations were determined: $[Pb^{2+}] = 0.030M$; $[NO_3^-] = 0.060M$; $[Cl^-] = 0.030M$

   (e) Will a precipitate form? Justify you answer. The $K_{sp}$ for the insoluble salt in question is $1.6 \times 10^{-5}$

**Free Response**

$$CH_3NH_2 + H_2O \rightleftharpoons CH_3NH_3^+ + OH^- \qquad K_b = 4.40 \times 10^{-4}$$

2. Methylamine, $CH_3NH_2$, a weak base, is added to a flask to form 1.50 L of solution. The reaction shown above occurs at 25°C.

   (a) Write the expression for the base dissociate constant, $K_b$.

   (b) The pH of the solution was determined to be 11.822. Calculate the molar concentration of methylamine in this solution.

In a different experiment, 150.0 mL of a 0.150 *M* solution of methylamine was prepared. To this solution, 0.120 mole of methylammonium chloride was added.

   (c) Calculate the pH after the addition of the methylammonium chloride, $CH_3NH_3Cl$.

   (d) To the 150 mL solution describe in part (c) 100 mL of 0.950 *M* HCl was added. Calculate the resulting pH of the solution.

   (e) Draw the Lewis structure for

      i. the weak base methylamine, $CH_3NH_2$.

      ii. methylammonium ion, $CH_3NH_3^+$.

## Free Response

3. A student performed a titration by dissolving a sample of a weak, monoprotic acid, HA, in 15.00 mL of water. The solution was then titrated with 0.100 *M* NaOH(*aq*). The equivalence point was reached after 20.00 mL of the NaOH solution had been added. The student's data from the titration are shown in the table below.

| **Volume of 0.100 *M* NaOH Added** | **pH** |
|---|---|
| 0.00 mL | 2.22 |
| 5.00 mL | |
| 10.00 mL | 3.44 |
| 15.00 mL | 3.92 |
| 20.00 mL | 8.13 |

(a) Calculate the number of moles of unknown acid originally present in the solution.

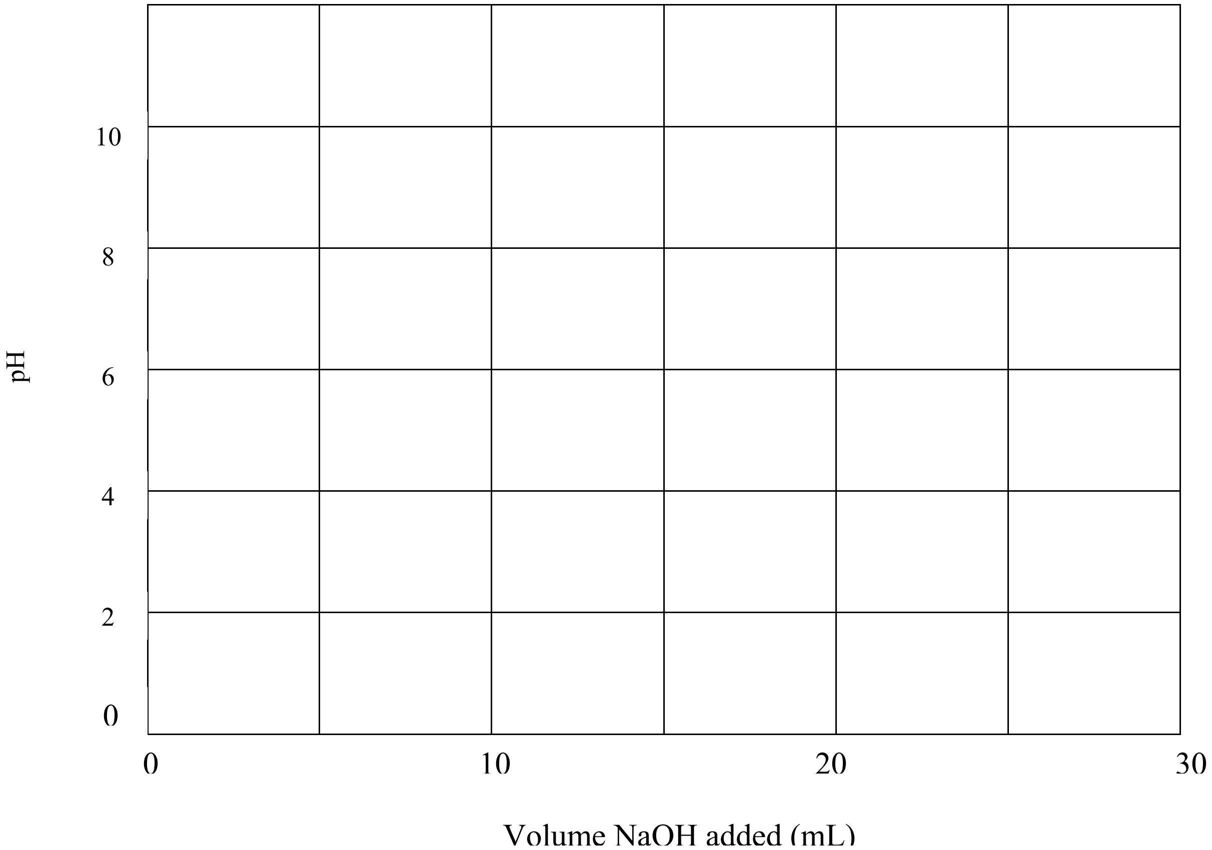
(b) Using the information provided above, calculate the value of the acid dissociation constant, $K_a$ , for HA.

(c) Determine the pH of the solution after 5.00 mL of 0.100 *M* NaOH had been added, and record the value in the student's data table.

(d) From the table below, select the most appropriate indicator for the titration. Justify your choice

| **Indicator** | **pH** |
|---|---|
| Methyl Red | 5.5 |
| Bromothymol Blue | 7.1 |
| Phenolphthalein | 8.7 |

(e) Using the axes provided below, sketch the titration curve that would result if the student had used 0.200 *M* NaOH instead of the 0.100 *M* used above, to perform the titration. The equivalence point must be clearly marked.



## Free Response

$$\text{HA}\,(aq) + \text{H}_2\text{O}(\ell) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{A}^-\,(aq) \qquad K_a = 4.5\times10^{-5}$$

4. An unknown monoprotic acid HA, is a weak acid that ionizes according to the reaction above. A 50 mL sample of the 0.020 *M* unknown weak acid solution described above was titrated with 0.050 *M* NaOH

   (a) Write a balanced net ionic equation showing the reaction between the weak acid, HA and NaOH.

   (b) For the titration, calculate the

      (i) equivalence point of the titration

      (ii) pH of the solution in the titration flask after 5.0 mL of 0.050 *M* NaOH were added

      (iii) pH of the solution in the titration flask after 20.0 mL of 0.050 *M* NaOH were added

      (iv) pH of the solution in the titration flask after 30.0 mL of 0.050 *M* NaOH were added

## Multiple Choice

1. If 50.0 mL of 0.25 *M* $NH_3$ is titrated with 0.50 *M* HCl in a flask. Which of the following are present in the flask at the equivalence point?

    I. ammonia
    II. hydrochloric acid
    III. ammonium ions

    (A) II only
    (B) III only
    (C) I and II only
    (D) II and III only
    (E) I, II, and III

2. A student titrates 25 mL of 0.50 *M* $NH_3$ with 0.25 *M* HCl. Which of the following would be true about the equivalence point of the titration?

    (A) 25 mL of HCl is needed to reach the equivalence point.
    (B) At the equivalence point, only the hydrochloric acid remains in the solution.
    (C) At the equivalence point, the $[OH^-]$ is equal to the base ionization constant.
    (D) There is no change in the pH of the solution after the equivalence point is passed.
    (E) The pH at the equivalence point is less than 7.

3. Which pair of 1.0 *M* solutions below would form a buffer solution with a pH less than 6? (assume 1:1 ratios of each component)

    (A) $CH_3NH_2$ and $CH_3NH_3Cl$
    (B) HCl and NaOH
    (C) $HC_2H_3O_2$ and $NaC_2H_3O_2$
    (D) HF and $HC_2H_3O_2$
    (E) $NH_3$ and $CH_3NH_2$

4. What is the pH of a buffer that is 0.30 *M* $NH_3$ and 0.20 *M* $NH_4Cl$? $K_b$ for $NH_3 = 1.8 \times 10^{-5}$

    (A) 4.92
    (B) 9.08
    (C) 10.98
    (D) 11.37
    (E) 13.24

5. When used in a titration which ion below turns purple when the endpoint is reached?

   (A) $I^-$
   (B) $Cl^-$
   (C) $Cr_2O_7^{2-}$
   (D) $CrO_4^{2-}$
   (E) $MnO_4^-$

6. 200.0 mL of 0.25 *M* $HC_2H_3O_2$ is titrated with 0.95 *M* $Ba(OH)_2$. What volume of barium hydroxide is needed to reach the equivalence point? $K_a$ acetic acid = $1.8 \times 10^{-5}$

   (A) 26 mL
   (B) 40 mL
   (C) 53 mL
   (D) $1.0 \times 10^2$ mL
   (E) $2.0 \times 10^2$ mL

Questions 7 and 8 refer to the following titration data.

In order to reach the equivalence point (pH = 5.36), it takes 50.0 mL of 0.10 *M* HCl to titrate 100.0 mL of a weak base. During the same titration, the halfway point (pH = 9.25) is reached with the addition of only 25.0 mL of 0.10 *M* HCl.

7. What is the $K_b$ of the weak base?

   (A) $6 \times 10^{-10}$
   (B) $2 \times 10^{-8}$
   (C) $4 \times 10^{-6}$
   (D) $2 \times 10^{-5}$
   (E) $5 \times 10^{-2}$

8. Which indicator would be best to use in this titration?

   (A) Phenolphthlein, $K_a = 5.0 \times 10^{-10}$
   (B) Phenol red, $K_a = 1.3 \times 10^{-8}$
   (C) Litmus, $K_a = 1.0 \times 10^{-7}$
   (D) Bromophenol Blue, $K_a = 7.9 \times 10^{-6}$
   (E) Thymol Blue, $K_a = 1.0 \times 10^{-2}$

# AP* CHEMISTRY

# The Lab Based Question
# Techniques, Types, and Tricks

## Student Packet

# THE LAB BASED QUESTION

Techniques, Types, and Tricks

## What I Absolutely Have to Know to Survive the AP Exam

If asked to do the following, it might indicate the question deals with laboratory questions:
Design an experiment; list measurements needed; show setup of calculations needed, use sample data to do calculations; interpret or draw graphs; explain the affect of error on the calculated value; use qualitative observations…

## Parts of a Lab Question

You may not have done that exact experiment but you can use the knowledge gained doing other experiments to help account for certain observations.

You may be asked to describe how to do an experiment or to design an experiment.

- Materials: If they give you a list of equipment, don't think you have to use all of it.
- Procedure: Be sure to include important techniques like rinsing the buret with the solution before a titration or heating to constant mass.
- Data needed: The data needed are values that can be measured like initial and final temperatures. Writing all the mathematical equations needed to do the calculations will help you determine what data is needed.
- Calculations: A calculation is using what was measured like temperature change. Show the set up of the mathematical equations required for the calculations. Use sample data when appropriate.
- Graphs: Be sure you label the axes and other important points on your graph.
- Error Analysis: State whether the quantity will be too high, too low, or no change. Use equations to help you determine what change will occur and to support your answer.

## Common Lab Procedure: Calorimetry

Calorimetry is used to determine the heat released or absorbed in a chemical reaction. A calorimeter can determine the heat of a solution reaction at constant pressure.

Techniques:

- Use a double Styrofoam cup with a plastic top and hole for the thermometer
- Determine the change in temperature accurately
- Measure solution volumes precisely
- Start with a dry calorimeter

Information to know about calorimetry:

- Heat capacity (C) = the amount of heat needed to raise the temperature of an object by one degree Celsius or Kelvin, J/°C or J/K.
- The heat capacity of 1 mol of a substance is called its molar heat capacity (Joules per mole per degree) J/mol·°C or J/mol·K.
- Specific heat, c, also known as specific heat capacity, is defined as the amount of heat necessary to raise the temperature of 1.00 g of a substance by one degree. Units are (joules per gram per degree), J/g·°C or J/g·K. You often use the specific heat capacity in analyzing gathered data then convert to molar heat capacity.

Assumptions often made during calorimetry: (Be able to answer error analysis questions about each assumption below)

- The density of dilute solutions is the same for water. D= 1.0 g/mL
- The specific heat of the solutions is the same as that for water. c = 4.184 J/g°C
- The solutions react in their stochiometric amounts.
- There is no loss of heat to the surroundings.

Equations:

- $q = mc\Delta T$ (c = 4.184 J/g°C)
- $q = \Delta H = mc\Delta T$ , at constant pressure
- Use the density of the solution to convert from volume to mass for $q = mc\Delta T$

## Common Lab Procedure: Titration

A titration is a laboratory procedure for quantitative analysis. In a titration two reagents are mixed, one with a known concentration & known volume [or a solid with a known mass] and one with an unknown concentration. There is some way to indicate when the two reagents have reacted completely (typically an indicator), and at the end of the titration the unknown solution's concentration can be calculated since you have accurately determined the volume of that solution required to complete the reaction.

Terms to know:

- Titrant – A solution of known concentration; it is often standardized
- Standardized solution – A solution in which the exact concentration is known.
- Indicator - A weak acid or base used in a titration to indicate the endpoint has been reached.
- Equivalence point – moles of acid = moles of base; point at which enough titrant has been added to completely react with the solution being analyzed.
- End point – the point at which the indicator changes color; important to pick an indicator with a pKa very close to the pH at the equivalence point.

Techniques:

Preparing the Buret:

1. Rinse a clean buret with distilled water and then the titrant (the solution that will be added to the flask).
2. Allow the titrant to drain through the buret so that the tip gets rinsed with titrant as well.
3. Discard the rinse solution. Fill it with the titrant. Remove air bubbles from the tip of the buret by draining several milliliters of titrant.

Preparing the Sample:

4. Pipet the desired volume of the solution to be analyzed into an Erlenmeyer flask. Record the exact volume. If the sample is a solid, weigh the desired mass, add the solid to an Erlenmeyer flask, and dissolve it in distilled water (the amount of water does matter since it doesn't change the moles of the solid). Be sure to record the exact mass of sample used.
5. Add the titrant to the flask until the equivalence point is reached. Calculate the volume of titrant added.
6. Change in color of a chemical indicator is usually used to signal the endpoint of the titration. The endpoint for this titration is reached when you reach a pale color that persists for several seconds.

**Measured Data** Required:

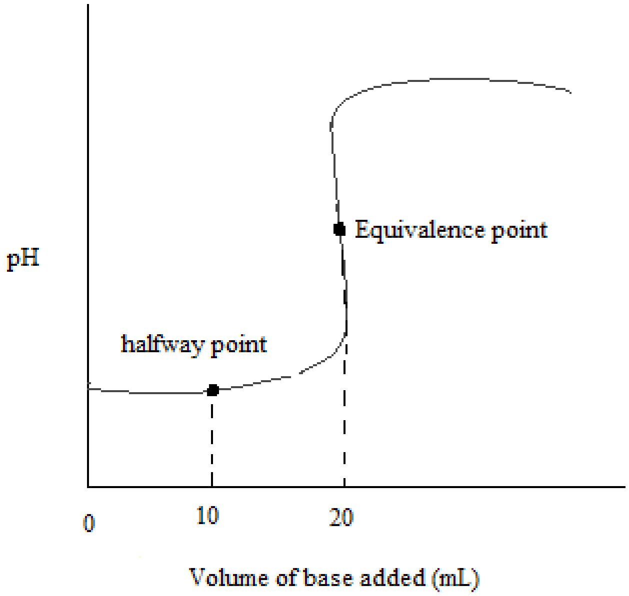
**moles** titrant = **moles** of substance analyzed @ equivalence point

- mass of DRY substance analyzed OR accurately measured volume of solution analyzed [measure with a pipet OR buret ]
- *initial* volume of titrant (substance of known molarity) and *final* volume of titrant (required to reach end point)
- Molarity of titrant

**Calculation** Hints:

- Substance analyzed is solution/liquid:
  - $M_1V_1 = M_2V_2$ @ equivalence point
  - Volume of titrant used to reach end point [difference between final and initial volumes]
  - $M_{\text{titrant}}V_{\text{of titrant added}}$ = moles of titrant = moles of unknown
  - Molarity of unknown = $\text{Molarity of unknown} = \dfrac{\text{moles of unknown}}{\text{Liters of unknown}}$
- Substance analyzed is solid:
  - Same as process as above
  - $M_{\text{titrant}}V_{\text{of titrant added}}$ = moles of titrant = moles of unknown
  - $\text{Molecular weight of the unknown} = \dfrac{\text{mass of solid dissolved}}{\text{moles of unknown}}$

**Graphs:**



## Common Lab Procedure: Gravimetric Analysis

One method for determining the amount of a given substance in solution is to form a precipitate that includes the substance. The precipitate is then filtered and dried. This process is called gravimetric analysis.

Techniques/Procedure:

1. Weigh sample
2. Form precipitate
3. Filter precipitate (A buchner funnel and aspirator can be used)
4. Dry precipitate (Be sure to dry to constant mass)
5. Weigh precipitate

For example if we wanted to determine the amount of chloride ions present in a given solid, we would weigh the solid sample, dissolve the sample in water, add an excess of silver nitrate solution to form the precipitate silver chloride. This precipitant would be filtered, and dried to constant mass. From the mass of silver chloride formed, we can determine the moles of silver chloride and the moles of chloride ion in the original sample.

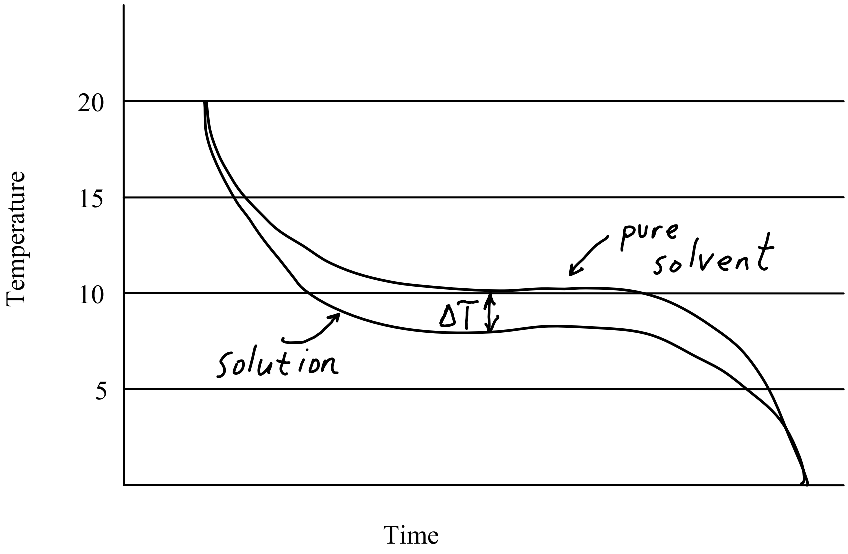
## Common Lab Procedure: Determining Molar Mass

$$\text{molar mass} = \frac{\text{mass of sample in grams}}{\text{moles of sample}}$$

- Organize answer around calculations—paying special attention to what quantities are **measured** versus **calculated**!

A. Freezing point depression:

- $\Delta Tf = mik_f$ (nonelectrolyte $i$ = 1 (most common); $i$ is the van't Hoff factor and is equal to the # of ions released);
- $m = \frac{\text{moles of solute}}{\text{kg of solvent}}$ or $m = \frac{\frac{\text{mass solute}}{\text{molar mass solute}}}{\text{kg solvent}}$
- **Measured DATA**:
    - FP of pure solvent
    - final FP of solution
    - kg of solvent
    - grams of solute
- **Constants** needed:
    - $k_f$ = given OR determined by data (for $H_2O$ = 1.86)
- **Calculations** required:
    - $\Delta T$ = freezing point pure solvent – freezing point of solution
    - use $\Delta Tf = mik_f$ to solve for molality
    - use molality equation to solve for molar mass
- **Graphs**:



B. Titration Data

- **moles** titrant = **moles** of substance analyzed @ equivalence point
- **Measured DATA** required:
  - mass of substance
  - *initial* volume of titrant (substance of known molarity) and *final* volume of titrant (required to reach end point)
  - Molarity of titrant
- **Calculations** required:
  - Molarity of titrant
  - Substance analyzed is solid:
  - $M_{titrant}V_{titrant}$ = moles titrant
  - $\text{molar mass} = \dfrac{\text{g of solid analzyed}}{\text{moles of titrant used}}$

c. Vaporization of a Volatile Liquid

- $PV = nRT$ used to determine moles
- **Measured DATA** Required:
  - Pressure = atmospheric pressure unless collected over water
  - *initial* mass of flask
  - *final* mass of flask
  - Temperature of boiling water—don't assume 100°C
  - Volume of gas = fill flask with water and measure the volume of water in a graduated cylinder
- **Constants** needed:
  - If collected over water, the water vapor pressure at the experimental temperature.
- **Calculations** Required:
  - If the gas was collected over water $P_{vapor} = P_{atmospheric/baraometric} - P_{water\ vapor\ at\ certain\ temperature}$
  - mass of sample = *final* mass of flask [includes vapors] – *initial* mass of flask

## Common Lab Procedure: Colorimetric or spectrophotometric analysis

Colorimetric analysis is a quantitative analysis of a solution using color based on Beer's Law. Colorimetric analysis can be used to determine the concentration of an unknown solution, the rate constant of a reaction, the order of a reaction, etc.

Beer's Law is an expression than can be used to determine how much light passes through the solution. It also shows that concentration and absorbance are directly related.

$$A = \varepsilon bc$$

- A = absorbance (measured with a colorimeter)
- ε = molar absorptivity (how much light will be absorbed by 1 cm of a 1 M solution of the chemical)
- b = path length of the cuvette in cm
- c = concentration in molarity

# Lab Based Cheat Sheet

## Relationships

Calorimetry $q = \Delta H = mc\Delta T$ at constant pressure

- c = specific heat
- $\Delta t$ = change is temperature
- m = mass
- q = heat

Beer's Law $A = \varepsilon bc$

- A = absorbance
- $\varepsilon$ = molar absorptivity
- b = path length of the cuvette in cm
- c = concentration in molarity

Freezing point depression $\Delta T_f = mik_f$

- *i* = 1 for nonelectrolytes (most common)
- $\Delta T$ = freezing point pure solvent – freezing point of solution
- $k_f$ = molal freezing point constant
- $m = \dfrac{\text{mass solute}}{\dfrac{\text{molar mass solute}}{\text{kg solvent}}}$

Ideal Gas Law $PV = nRT$

- V = volume in liters
- n = moles
- T = temperature in Kelvin
- P = pressure

If collected using water displacement:

$$P_{\text{vapor}} = P_{\text{atmospheric/barometric}} - P_{\text{water vapor at certain temperature}}$$

$$\text{molar mass} = \frac{\text{mass of sample in grams}}{\text{moles of sample}}$$

@ Equivalence point $M_1V_1 = M_2V_2$

$$\%\text{error} = \frac{\text{accepted - experimental}}{\text{accepted}} \times 100$$

## Connections

Electrochemistry: draw diagram of galvanic or electrolytic cell, qualitative observations that can be made at the cathode and anode, etc.

Stoichiometry: empirical formula of a compound, percent of an element in a compound, etc.

Thermodynamics: heat of reaction, molar heat capacity, etc.

Acids and Bases: standardizing a solution, drawing a titration curve, etc.

Qualitative Analysis: identifying a compound based on observations and tests

## Potential Pitfalls

Be aware there are quantities you **measure** [such an initial temperature and final temperature or initial pressure and final pressure, etc.] and terms you **calculate** *using what you measured* such as Δ T or Δ P

Be sure to include important steps in the procedure like "heat to constant mass," "rinse the buret with distilled water and then the ***solution*** before titrating", "dissolve the solid in about 100 mL of water and then add water to the 500 mL mark on the volumetric flask," etc.

Writing the mathematical equations may help determine how an error affects the results. Use the equations to justify your error analysis.

If given a laboratory situation you have not specifically done, use the observations and the concepts you learned in your labs throughout the year to reason your way through the lab question.

## Free Response

1. Hydrochloric acid (about 5 mL) is placed at the bottom of the test tube and then carefully filled with distilled water so not to disturb the HCl. A piece of magnesium is placed on top of the distilled water and then the test tube is inverted into a beaker with distilled water. The magnesium is allowed to react completely. The gas is collected in the test tube at room temperature. Before the test tube with the gas is removed, the water levels inside and outside the test tube are the same. The chemicals and lab equipment used are listed below.

Hydrochloric acid
Table of water vapor pressures
Distilled water
Barometer
Graduated cylinder
600 mL beaker
Balance
Strip of magnesium metal
Test tube
Thermometer

(a) Write the balanced reaction that occurs when hydrochloric acid reacts with magnesium.

(b) What measurements are needed to calculate the molar volume of the gas at STP?

(c) Show the setup for the calculations needed to determine:

(i) the moles of gas produced

(ii) the molar volume of the gas at STP

(d) What test can be done to prove which gas is produced?

(e) What is the purpose of making the water level inside the test tube equal to the water level in the beaker?

(f) The vapor pressure of water is not used in the calculation. How will the molar volume of the gas at STP be affected? (higher, lower or the same) Explain.

## Free Response

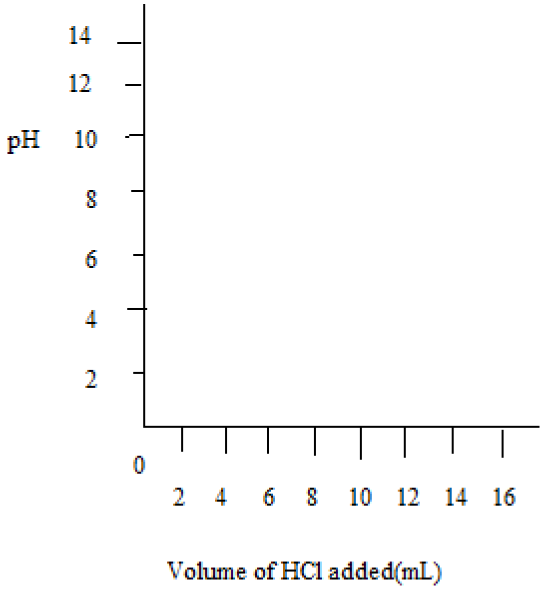
2. A solution of calcium hydroxide is titrated with hydrochloric acid.

   (a) Describe the steps needed to make 50.0 mL of a 0.050 *M* solution from a 2.00 *M* solution of hydrochloric acid using a dropper, 5.0 mL pipet, 50.0 mL volumetric flask, and distilled water.

   (b) Write the balanced chemical equation for the dissociation of $Ca(OH)_2(s)$ in aqueous solution, and write the equilibrium-constant expression for the dissolving $Ca(OH)_2$.

A 15.0 mL saturated solution of $Ca(OH)_2$ is titrated with a standardized 0.050 *M* solution of hydrochloric acid. It takes 10.50 mL of hydrochloric acid to reach the equivalence point where the pH is 8.50.

   (c) Sketch the titration curve that shows the pH change as the volume of hydrochloric acid added increases from 0 to 16.0 mL. Label the equivalence point.



   (d) Calculate the concentration of $[OH]^-$ from the titration.

   (e) Calculate the $K_{sp}$ for calcium hydroxide.

## Free Response

3. A molecular compound is analyzed and found to be 76% C, 12.6 % H, and 11.2 % O.

    (a) What is the empirical formula of the molecular compound?

The freezing point of the pure solvent, lauric acid, is found to be 44.0 °C. A nonvolatile molecular compound is mixed with lauric acid. The freezing point of a mixture of 8.11 grams of lauric acid and 1.12 grams of the organic compound is found to be 42.1 °C. The freezing point molal constant for stearic acid is 4.9 °C/m, and the freezing point molal constant for lauric acid is 3.9 °C/m.

(b) Calculate the molality of the mixture of lauric and and the unknown compound.

(c) Calculate the molar mass of the unknown compound.

(d) Determine the molecular formula of the unknown compound.

(e) If the same experiment is conducted but with an ionic compound instead of a molecular compound as the solute, how would the molar mass of the compound be affected? (higher, lower or no change) Explain.

## Multiple Choice

1. Which ion below would be colorless when in solution?

(A) $Zn^{2+}$
(B) $Fe^{3+}$
(C) $Cu^{2+}$
(D) $Mn^{2+}$
(E) $Ni^{2+}$

2. When solid LiCl is heated in a flame, what is the color of the flame?

(A) Yellow
(B) Orange
(C) Red
(D) Purple
(E) Green

3. A student heats a test tube containing a volatile liquid in a warm water bath and then condenses the gas using a cold water bath. The data obtained is used to calculate the molar mass of the volatile liquid. Which measurements below are needed to determine the molar mass of the volatile liquid?

I. Temperature of the gas
II. Mass of the condensate
III. Barometric pressure
IV. Vapor pressure of water

(A) II only
(B) I and III only
(C) II and III only
(D) I, II, and III only
(E) I, II, III, and IV

4. 10.00 mL of acetic acid is titrated with 0.100 *M* sodium hydroxide. If an indicator is used to signal the endpoint of the titration, which indicator below is best to use?

(A) Methyl red, $K_a = 2.0 \times 10^{-4}$
(B) Litmus, $K_a = 3.2 \times 10^{-7}$
(C) Thymolphthalein, $K_a = 2.4 \times 10^{-10}$
(D) Crystal violet, $K_a = 7.9 \times 10^{-2}$
(E) Bromothymol blue, $K_a = 5.6 \times 10^{-8}$

5. Which solution would produce a precipitate with the addition of hydrochloric acid?

(A) $Ca(NO_3)_2$
(B) $[Ag(NH_3)_2]NO_3$
(C) $Fe(C_2H_3O_2)_2$
(D) $AlF_3$
(E) $Cr(ClO_4)_2$

6. The concentration of a solution is determined using a spectrophotometer and Beer's Law.

Beer's Law: $A = \varepsilon bc$

Which of the following statements are true about Beer's Law?

I. The absorbance is directly related to the concentration.
II. The $\varepsilon$ in Beer's Law is the molar absorptivity constant.
III. As the absorbance of a solution increases, the percent transmittance decreases.
IV. The path length and the percent transmittance are indirectly related.

(A) I only
(B) I and III only
(C) I and II only
(D) I, II, and III only
(E) I, II, III, and IV

7. An unknown gas in a vial is collected in a graduated tube over water.

| | |
|---|---|
| Temperature of water | 20.0ºC |
| Atmospheric pressure | 744.8 mm Hg |
| Mass of vial before gas collected | 25.63 g |
| Mass of vial after gas collected | 25.20 g |
| Volume of the gas collected | 125.0 mL |
| Water vapor pressure at 20ºC | 17.54 mm Hg |

Using the data above, which expression below can be used to calculate the molar mass of the gas?

(A) $\dfrac{(25.63 - 25.20)\times(0.0821)\times(293)\times(760)\times(1000)}{(125)\times(744.8)}$

(B) $\dfrac{(25.63 - 25.20)\times(0.0821)\times(293)\times(1000)}{(125)\times(744.8 - 17.54)}$

(C) $\dfrac{(25.63 - 25.20)\times(125)\times(293)\times(760)}{(0.0821)\times(744.8 - 17.54)\times(1000)}$

(D) $\dfrac{(25.63 - 25.20)\times(0.0821)\times(293)\times(760)\times(1000)}{(0.125)\times(744.8 - 17.54)}$

(E) $\dfrac{(25.63 - 25.20)\times(0.0821)\times(293)\times(1000)}{(125)\times(17.54)}$

8. A student separates a mixture using distillation. Distillation can be used to separate the mixture because the two substances have different

(A) melting points.
(B) reactivities.
(C) densities .
(D) boiling points.
(E) viscosities.

9. Which of the following is true about the oxidation-reduction reaction below?

$$2\,I^- + Cl_2 \rightarrow I_2 + 2\,Cl^-$$

(A) $I^-$ is being reduced by $Cl_2$.
(B) The reducing agent is the $Cl_2$.
(C) The solution will be purple after the reaction.
(D) The half reaction that occurs at the anode is $Cl_2 + 2e^- \rightarrow 2Cl^-$.
(E) The voltage will increase if additional $I_2$ is added.

10. Which mixture of substances below will not react?

(A) Solutions of HI and $CaCO_3$
(B) Solutions of $(NH_4)_2SO_4$ and KOH
(C) Solutions of $SnCl_2$ and $KMnO_4$
(D) Solution of CaO and $H_2O$
(E) Solution of $CaCl_2$ and $Br_2$ gas

# AP* CHEMISTRY

## Net Ionic Equations
## Notes, Hints, and Practice

### Student Packet

# Net Ionic Equations

Notes, Hints, and Practice

## General Ideas Regarding AP Equations

1. Know your solubility rules inside and out.
2. Reactions are net ionic so no spectator ions are to be included.
3. All reactions happen so “no reactions” are not possible.
4. Equations must be balanced.
5. Do not include states of matter.
6. Strong acids and bases need to be written in ion form. There are a couple exceptions to this but no worries.

## Double Replacement Reactions

When trying to decide reaction type, try to determine if it’s a double replacement FIRST. If you can form any of the following substances, you know you have a double replacement reaction:

**A precipitate (an insoluble solid)**
**Water**
**A gas**
**A weak acid or a weak base**

Reminder: Double replacement reactions are NOT redox reactions. Keep your oxidation states the same.

Note: Ammonium chloride, sulfurous acid, and carbonic acid are unstable.
$NH_4OH \rightarrow NH_3 + H_2O$
$H_2SO_3 \rightarrow SO_2 + H_2O$
$H_2CO_3 \rightarrow CO_2 + H_2O$
To receive full credit, make sure to decompose those unstable compounds totally as indicated in the reactions above.

## Redox Reactions

When faced with a redox reaction consider the following:

1. Remember that something must go up in oxidation state/charge and something must go down.
2. Never include group 1 metal ions in the reaction unless included in a “solid”.
3. Do not include the nitrate ion unless it is in the form of nitric acid reacting with a metal.
4. “Go To Metals First!” Meaning…..try and make redox happen between metals and metal ions first. If you can, then that is where the redox happens. There are a few exceptions to this but none you need to worry about.
5. Polyatomic ions will decompose to gases. If you can’t make redox occur between metals and a polyatomic is present, a gas that comes from that ion will be part of the redox coupling.

A couple must know rules:
$MnO_4^-$ in an acidic solution produces $Mn^{2+}$
$Cr_2O_7^{2-}$ in an acidic solution produces $Cr^{3+}$

## Single Replacement Reactions

Single Replacement reactions are just redox. The same rules apply.
These are actually easy to identify due to the presence of a single element trying to displace another from a compound.

Note: Be very careful when there is a metal being added to water. The most common student mistake is to produce a metal oxide and hydrogen gas. This is WRONG!!! When adding a reactive metal to water you produce a metal HYDROXIDE (base) and hydrogen gas.

## Combustion Reactions

You know you have a combustion reaction when something is burned in air or burned in oxygen.
Combustion reactions are also REDOX equations so be sure to look for oxidation number changes.
No memorization is required for these equations. Oxygen will just smack into all the elements in present with oxygen always going down in oxidation state from a zero to a minus 2.

## Addition Reactions

Some addition reactions are easy to identify. Those easy ones include two elements that are coming together to form one product. These are also another type of REDOX reaction where oxidation state changes occur.

However one of the most common addition reactions is the strangest one which includes an acidic or basic anhydride.
OXIDES ARE WEIRD!
Non metal oxides are "fake acids" – acids without water. When you add water to one of these you make an acid.
Metal oxides are "fake bases" – bases without water. When you add water to one of these you make a base.
Say to yourself **"NO acid MO base"** (NO is nonmetal oxide and MO is metal oxide)

## Neutralization Reaction

Simply put a neutralization reaction is a specialized double replacement reaction where an acid and a base make a salt and water. These are NOT redox reactions so keep your oxidation numbers the same.

Be careful about addition a fake acid to a fake base (nonmetal oxide to a metal oxide). Since water is not present you will produce just the salt! This makes it an addition reaction as well.

Also if you're adding a fake acid (nonmetal oxide) to a real base or a fake base (metal oxide) to a real acid you will still produce water as a product. One of the ions from the salt will be a component derived from the metal or non metal oxide involved.

Here are a few odd ones that might be helpful to memorize remember; OXIDES ARE WEIRD!
Metallic oxides plus $CO_2$ make metallic carbonates
Metallic oxides plus $SO_2$ make metallic sulfites

## Decomposition Reactions

Decomposition reactions are EASY to recognize. One reactant is decomposing to multiple products. Simple compounds containing just two types of elements easily split into two separate elements. (Be carefully not to produce ions in the decomposition unless water is present)

Some of you more difficult decomposition reactions contain polyatomics and can get a bit harder. Rather than trying to memorize a billion rules, just remember: **Polyatomics can decompose to gases.** Knowing that will get you through many of these problems.

Remember those unstable compounds from double replacement? They are decomposition reactions as well!
$NH_4OH \rightarrow NH_3 + H_2O$
$H_2SO_3 \rightarrow SO_2 + H_2O$
$H_2CO_3 \rightarrow CO_2 + H_2O$

One last one you should know:
Hydrogen peroxide decomposes to make water and oxygen

## Anhydride Reactions

Anhydride means "without water" so you are adding water as a reactant in these problems. Pretty easy to identify!

Just as we saw in addition reactions:
Non metal oxides are "fake acids" – acids without water. When you add water to one of these you make an acid.
Metal oxides are "fake bases" – bases without water. When you add water to one of these you make a base.
Say to yourself **"NO acid MO base"** (NO is nonmetal oxide and MO is metal oxide)
OXIDES ARE WEIRD!

A couple additional ones you'll want to learn:
Metallic hydrides plus water produce metallic hydroxides plus hydrogen gas (same as a metal plus water right?)
Phosphorus halides plus water produce a hydrohalic acid plus a ternary acid of phosphorus

## Complex Ion Reactions

How do you know if you're doing a complex ion problem? You need to have a metal in the presence of a very high concentration of a ligand.

Metal: "Fish Metals" Ag, Cu, Cd, Zn and Al (These metals form a diving fish on the periodic table…check it out! You'll need to use your imagination because it is a stretch!) Theoretically any transition metal can complex but these are your most common.

Ligand: **High concentration** of $OH^-$ (hydroxy) or $NH_3$(amine) or $SCN^-$ (thiocyano). Again, there are other possible ligands but these are the most common.

The number of the ligands that attach to the metal in the complex is usually double that of the most common oxidation state of that metal OR using the fish mentioned above, the nose of the fish ($Ag^+$) can attach two ligands and the rest of the fish can attach 4 ligands.

The overall charge of the complex is based on the combination of charges from both the metal and the ligand (if the ligand is even charged).

## Organic Nomenclature Review

**Alkanes**

Alkanes only contain single bonds (sigma bonding) and have the general formula $C_nH_{2n+2}$.
The prefixes for alkenes that you need to know in combination with the suffix –ANE are listed for below:

| | | | | |
|---|---|---|---|---|
| 1-carbon | methane | 4-carbons butane | 7-carbons heptanes | 10-carbons decane |
| 2-carbons | ethane | 5-carbons pentane | 8-carbons octane | |
| 3-carbons | propane | 6-carbons hexane | 9-carbons nonane | |

**Cycloalkanes**

These hydrocarbons exist in ring structures with general formula $C_nH_{2n}$. You name them just as alkanes but add the prefix **CYCLO-**

**Alkenes**

These are compounds containing at least one carbon to carbon double bond with the general formula $C_nH_{2n}$. They use the same prefixes as alkanes but end in an –ENE suffix.

**Alkynes**

These are compounds containing at least one carbon to carbon triple bond with the general formula $C_nH_{2n-2}$. They use the same prefixes as alkenes but end in a –YNE suffix.

**Nomenclature of Alkanes, Alkenes, and Alkynes:**

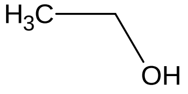
1. Use generic prefix according to the number of carbons then add the correct suffix.
2. Select the longest continuous carbon chain to dictate the number of carbons for the correct prefix "parent chain".
3. Any branch off of the parent chain is called a substituent group and it is named according to the number of carbon atoms it contains with a "-yl" ending rather than "-ane".
4. Specify the position of the branch on the parent by putting a number in front with the lowest #s possible.
5. For multiple substituents that are identical, use di-, tri-, tetra- etc, and repeat the numbers.
6. Numbers are separated by commas and from letters by dashes (1,1-dichloroethane).
7. Arrange substituents by name, alphabetically, regardless of numbers or complexity.
8. For alkenes and alkynes put a number right before the parent chain name, specifying position of the double or triple bond. Select numbering which always gives the lowest value.

### Functional Groups and Naming:

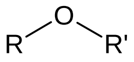
**Alcohols** R—OH (hydroxyl group)

- Alcohols are names by the alkyl prefix + "-ol"

Ex) ethanol

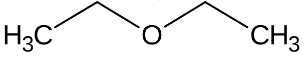


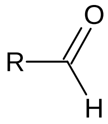
**Ethers**



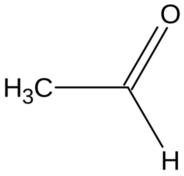
- Ethers are named by listing substituents with a "-yl" ending + "ether"

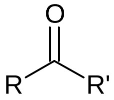
Ex) diethyl ether



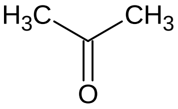
**Aldehydes**  (terminal carbonyl, -CHO)

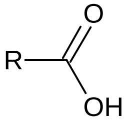
- Aldehydes are named by the alkyl prefix + "-al" ending

Ex) ethanal 

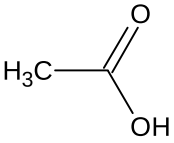
**Ketones**  (carbonyl group)

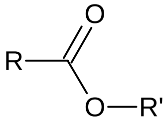
- Ketones are named by alkyl prefix + "-one" ending

Ex) propanone 

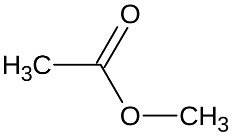
**Organic acids**  (-COOH)

- Organic Acids are named by alkyl prefix + "-oic acid" ending

Ex) ethanoic acid 

**Esters** 

- Esters are named from the alcohol name with a "-yl" ending, then the acid name with an "-oate" ending

Ex) methyl ethanoate 

# Last Minute Attack Strategies

These probably won't get you a 15 on Question 4, BUT they will help you if you are still having trouble with that **and** they WILL HELP YOU BEAT THE NATIONAL AVERAGE. The national average on this question is usually between 7—8. If you can consistently score 11—12, that puts you WAY ahead of the game!

First, the minimum knowledge required to survive this question…

MUST KNOW:

1. Big Mamma: All nitrates are **soluble**.
2. Big Daddy: All IA metals and ammonium salts are **soluble**.
3. Halides ($Cl^-$, $Br^-$, $I^-$): All are **soluble** except silver, mercury or lead.
4. Strong acids: **hydrochloric, hydrobromic, hydroiodic, nitric, perchloric, sulfuric**—WRITE THESE DISSOCIATED *except* concentrated sulfuric, it really is 97% $H_2SO_4$ and 3% water so the molecules don't dissociate completely.
5. Strong bases: hydroxides [and oxides] of IA and IIA* metals—write these bases dissociated.

WRITE ALL WEAK ACIDS AND BASES AS MOLECULES—be on the look out for $BF_3$ and its cousins $BCl_3$, etc. They are classic Lewis acids and when reacting with ammonia (a classic weak Lewis base), the product is $F_3BNH_3$ (just smash everything together) since nitrogen donated its unshared electron pair to boron in an act of extreme generosity and formed a coordinate covalent bond. Lewis **A**cids **A**ccept an electron pair.

* the little guys in the IIA's have solubility issues, write Be and Mg UNdissociated—calcium can go either way, the big guys are soluble. HF is not a strong acid since it's the little guy as well in the halogen series. The little guys make a stronger bond with H and do not dissociate in water. Also remember that the IA metals are named the alkali metals and the IIA's are the alkaline earth metals. What does "alkaline" mean? BASIC, so put them in water as metals, they dissolve and you make $OH^-$. Put IA metals in water and KABOOM. KABOOM = formation of explosive hydrogen gas, $H_2$.

## Attack Strategies

Before trying to figure out the “answers”, ***scan the words*** on each reaction for question 4 and do the following: [don’t write any products until you’ve done all SIX things!]

1. Cross out the word ~~***nitrate*** an~~y time it appears on the page.

2. Circle any word that implies solid or gas. (powdered, turnings, chunk, vapor, etc.)

3. Cross out any IA~~ metal t~~hat you see ***UNLESS*** it is associated with a circled solid or gas word.

4. Underline **halides** then ask yourself if silver, mercury or lead is present—if not you can cross the halide off as well such as with hydro~~**chloric**~~ acid. The $H^+$ is the reacting species. (Bring the halide back as a reacting ion IF you need to oxidize something halide$^{-1}$ →halogen$_2$.)

5. Circle “burned in air” or “combines with oxygen” or anything that implies combustion and celebrate!

6. Circle the word ***concentrated***. Get very excited if you see ***excess concentrated***. It means you have entered the land of complex coordinated ions (excess is not necessary, but often appears). Sounds scary, but VERY easy. LOSE THE FEAR!

**Now**, WRITE THE REACTANT SETS FOR ALL OF THE ONES YOU’VE MARKED. Spend 8 minutes writing products using those solubility rules and strong acid-base guidelines listed on the other side of this page. To get the easy three points involved with step SIX above do the following:

Write the reactants. On the product side, open a set of brackets [ ]. Put the metal ion in the brackets first then open a set of parentheses [M ion( )]. Next put a subscript on the parentheses that is twice the charge on the metal—I’m not proud of this, but it will earn credit. For a +2 metal it becomes $[M^{2+}(\ )_4]$. Finally, plop the ligand inside the parentheses and do the math to get the charge. If the ligand is ammonia or water, the ligand is neutral, so our example carries a +2 overall charge, $[M^{2+}(\ )_4]^{2+}$, if the ligand is hydroxide or a halide, which are both negative one, then our example becomes $[M^{2+}(OH)_4]^{2-}$. Other ligands are possible, like $SCN^-$, the thiocyanate ion and other polyatomic ions you should recognize.

Additional knowledge that contributes to survival:

- metal oxides + water → bases (ask yourself strong or weak? Dissociate the strong)
- nonmetal oxides + water → acids (ask yourself strong or weak? Dissociate the strong)
- metal carbonate heated → $CO_2$ + metal oxide
- Redox, “acidified”? $H^+$ is a reactant and water is a product.
- React a metal with oxygen → metal oxide
- React a nonmetal with oxygen [combustion] → make oxides of the nonmetal(s), NOT always $CO_2$ & $H_2O$!

## Free Response Examples

1. Sodium sulfide is added to a zinc nitrate solution.

   (i)

   (ii) What are the spectator ions in this reaction?

2. Magnesium ribbon is added to an iron(III) chloride solution.

   (i)

   (ii) What will be observed if a magnet is brought near the reaction vessel?

3. Lithium solid is burned in air.

   (i)

   (ii) What is the change in oxidation number for lithium?

4. Solutions of tin(II) chloride and iron(III) sulfate are mixed

   (i)

   (ii) What is the reducing agent? Explain.

5. Solid aluminum hydroxide is added to a highly concentrated sodium hydroxide solution.

   (i)

   (ii) What is the name of the product for this reaction?

6. Solid calcium oxide is heated with sulfur trioxide gas.

   (i)

   (ii) What state of matter is the product?

7. Solutions of potassium fluoride and hydrochloric acid are mixed.

   (i)

   (ii) Would the resulting solution conduct electricity? Explain.

8. Aluminum metal is added to a solution of copper(II) chloride.

   (i)

   (ii) What color changes will be observed during this reaction? Why?

9. Water is added to solid sodium hydride.

   (i)

   (ii) Describe a test to positively identify the gaseous product.

10. Ethanol is burned in oxygen completely.

    (i)

    (ii) Is the reaction endothermic or exothermic?

11. Hydrochloric acid is added to a potassium carbonate solution.

    (i)

    (ii) Describe a test to positively identify the gaseous product.

12. Iron (III) ions react with iodide ions.

    (i)

    (ii) How many electrons are transferred in this reaction?

13. Aluminum pellets are added to a solution of silver nitrate.

    (i)

    (ii) What element is oxidized?

14. Solid lithium hydride is mixed with water.

    (i)

    (ii) If phenolphthalein is added to the final solution, what is the color of the solution?

15. Chlorine gas is bubbled into a solution of potassium iodide.

    (i)

    (ii) What color is the solid product formed?

16. Solid sodium is added to water.

    (i)

    (ii) Explain how inserting a glowing splint into the reaction test tube can positively identify the gas produced.

17. Solid potassium oxide is added to water.

    (i)

    (ii) Would the pH of the final solution be greater than 7, less than 7, or equal to 7?

18. A piece of solid bismuth is exposed to a large volume of oxygen.

    (i)

    (ii) Which substance is the oxidizing agent?

19. A concentrated solution of ammonia is added to a solution of zinc iodide.

    i)

    ii) name the ion produced

20. Magnesium turnings are burned in nitrogen gas.

    (i)

    (ii) Write a balanced half reaction for the substance being oxidized.

21. A solution of potassium hydroxide is added to a solution of potassium dihydrogen phosphate until the same number of moles of each compound has been added.

    (i)

    (ii) Is the resulting solution acidic, basic or neutral? Explain.

22. Solutions of silver nitrate and sodium chromate are mixed.

    (i)

    (ii) What are the spectator ions in this reaction?

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23. The gases boron trifluoride and ammonia are combined

    (i)

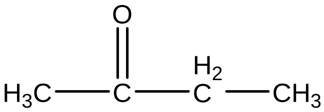
    (ii) Which reactant is the Lewis acid?

24. Solid calcium oxide and solid tetraphosphorus decaoxide are mixed and heated.

    (i)

    (ii) Is the product soluble in water? Explain.

## Multiple Choice



$$H_3C-\overset{\overset{\displaystyle O}{\|}}{C}-\overset{H_2}{C}-CH_3$$

1) The compound above is classified as a/an
   - (A) Ester
   - (B) Ether
   - (C) Ketone
   - (D) Carboxylic acid
   - (E) Aldehyde

2) A yellow precipitate forms when a solution of KI is mixed with which of the following ions?
   - (A) $SO_4^{2-}(aq)$
   - (B) $Zn^{2+}(aq)$
   - (C) $Pb^{2+}(aq)$
   - (D) $Na^{1+}(aq)$
   - (E) $CrO_4^{2-}(aq)$

**Questions 3-6** refer to the reactions represented below.

(A) A piece of iron is added to a solution of iron(III) sulfate

(B) Solid zinc sulfide is heated in an excess of oxygen

(C) Manganese(II) nitrate solution is mixed with a sodium hydroxide solution

(D) A suspension of zinc hydroxide is treated with concentrated sodium hydroxide solution

(E) Solid lithium hydride is added to water

3) A reaction that produces a complex ions

4) A precipitation reaction

5) A combustion reaction

6) A reaction in which the same element undergoes both oxidation and reduction

**<u>Questions 7–10</u>** refer to the following solids.

(A) $KMnO_4$
(B) $NiSO_4$
(C) $CuSO_4$
(D) $NaCl$
(E) $FeSCN^{+2}$

6. Makes a purple solution dissolved in water
7. Is white and very soluble in water
8. Makes a blue solution when dissolved in water
9. Makes a green solution when dissolved in water
10. Makes an orange coordination complex