Chemistry 12
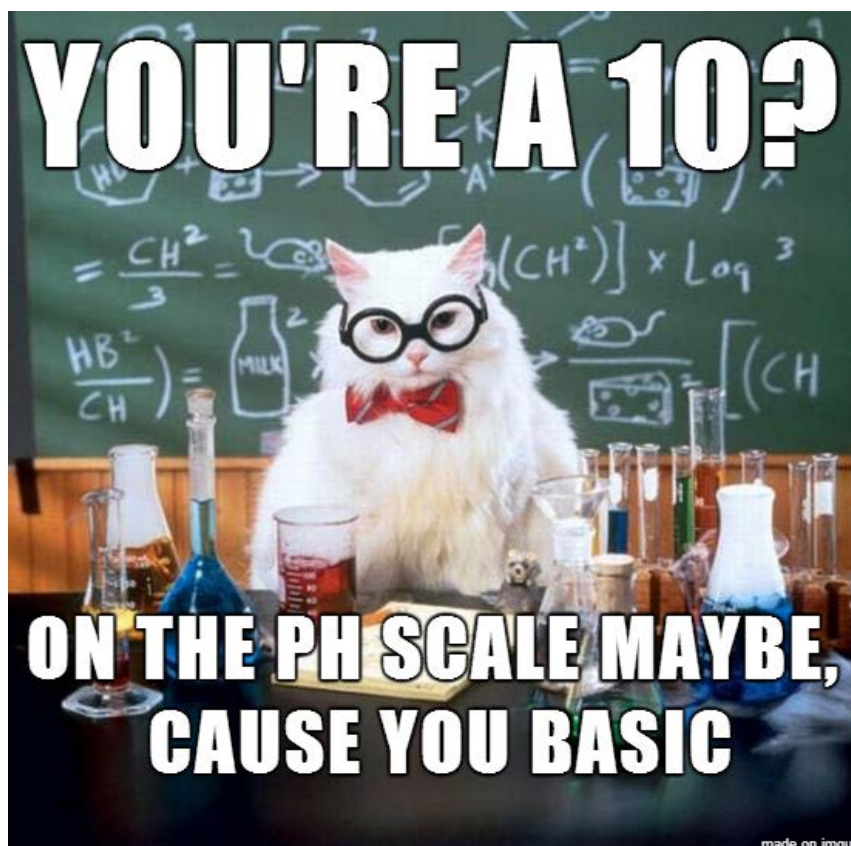## *Unit IV*

Acid/Base I

Name:_____   Block:_____

I) Introduction to Acids and Bases

What is an acid?

What are properties of acids?

1. Acids react with _____.

2. Acids create _____ in solution and therefore _____.

3. Acids react with some _____ to produce H_2 gas.
   e.g.

4. Acid turns litmus paper _____.

5. Acids taste _____.

6. Acids donate _____ to other substances.

What is a base?

What are properties of bases?

1. Bases react with _____.

2. Bases create _____ in solution and therefore _____.

3. Bases feel _____.

4. Bases turn litmus paper _____.

5. Bases taste _____.

6. Bases accept _____ from other substances.

II) Arrhenius Acids and Bases

Svante Arrhenius was a Swedish scientist who lived from 1859-1927. In

1884, he proposed the following definitions for acids and bases.
Arrhenius Acid:

Arrhenius Base:

These definitions stood until 1923, when they were revised by Bronsted (Danish) and Lowry (English), as we will see shortly.

When an acid reacts with a base (one which contains OH^-), what is produced?

e.g. $HCl_{(aq)} + NaOH_{(aq)} \Rightarrow$

     acid   +   base   makes
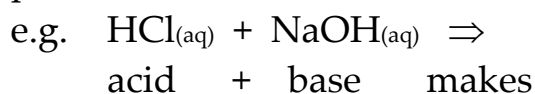
The OH^- acts as the base and takes the H^+ to make _____.
What type of reaction is this?

Acids and bases are both harmful substances, but if they react in stoichiometric amounts (so that there is no excess acid or base left over), the products are not harmful (_____).

What is a salt?

When the salt is produced in an acid-base reaction, depending on what salt is produced, what two results can occur?

Another example: $H_2SO_{4(aq)} + Sr(OH)_{2(aq)} \Rightarrow$

               acid   +   base
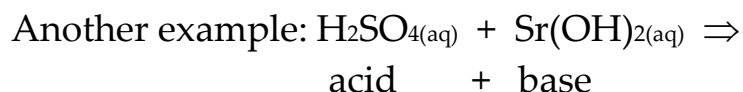
What is a Net Ionic Equation?

What is the net ionic equation of the reaction above?

What is the net ionic equation of the first neutralization equation we looked at (remember, only the things that change are included in the net.

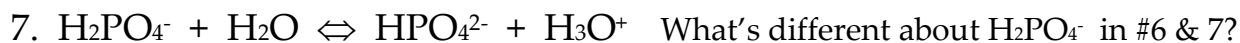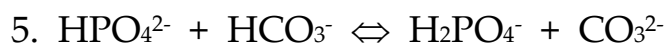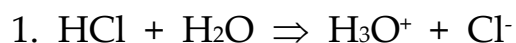**Assignment 1**: Hebden Read p.109-114 and do p.112 #1-4

III) Bronsted-Lowry Acids and Bases Part 1

Bronsted-Lowry definitions from 1923

Bronsted-Lowry Acid:

Bronsted-Lowry Base:

Practice: Label each reactant as an acid or a base depending on if it donates H^+ or accepts H^+. Some of the reactions are 100% and some are at equilibrium. You will learn why soon.

1. $HCl + H_2O \Rightarrow H_3O^+ + Cl^-$

2. $NH_3 + H_2O \Leftrightarrow NH_4^+ + OH^-$ What's different about H_2O in #1 compared to 2?

3. $CO_3^{2-} + H_2O \Leftrightarrow HCO_3^- + OH^-$ Why can CO_3^{2-} only be a base?

4. $HPO_4^{2-} + HBr \Rightarrow H_2PO_4^- + Br^-$

5. $HPO_4^{2-} + HCO_3^- \Leftrightarrow H_2PO_4^- + CO_3^{2-}$

6. $H_2PO_4^- + HF \Leftrightarrow H_3PO_4 + F^-$

7. $H_2PO_4^- + H_2O \Leftrightarrow HPO_4^{2-} + H_3O^+$ What's different about $H_2PO_4^-$ in #6 & 7? Notice from the practice equations that when bases do not contain OH^-, the products are not water and salt. Why is water not a product?

Assignment 2: Read Hebden p.116 & do p.117 #11

IV) Strong and Weak Acids and Bases

Strong acids and bases react to completion (100%) in solution (in water).

Examples
strong acid: HCl + H_2O
          0.10M

strong base: NH_2^- + H_2O
          0.200M

strong 'OH⁻' base: $NaOH$
                1.0M

The OH⁻ produced is now ready to act as a base and accept H⁺.

Weak acids and bases do not react to completion in solution (in water).
They create an equilibrium with reactants heavily favoured.

Examples
1.0M weak acid: HF + H_2O

1.0M weak base: CO_3^{2-} + H_2O

V) H⁺ and H₃O⁺

The most abundant hydrogen atom, by far, is hydrogen-1, which has an atomic weight of 1 amu, which means it must be made up of ____ proton, ____ neutron, and ____ electron.

H⁺ has lost an _____, and thus it is simply a _____, which is what it is commonly called.

When an acid such as HCl is put into solution, what happens?
Simple version: $HCl_{(aq)}$ $\Rightarrow$

The equation suggests that HCl gives up a proton 100% in water. However, in the reaction above, no substance is accepting the proton, which is inaccurate. However, we use this reaction to show what HCl does in the presence of any base. If HCl is in solution and just water is present, HCl donates its proton to water, as shown below:

$$HCl_{(aq)} + H_2O_{(l)} \Rightarrow$$
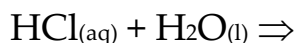
Both equations are commonly used, so H^+ is analogous to H_3O^+ (called *hydronium*).

VI) The Acid/Base Table (in aqueous solution)
Where are the acids on the table, and how are they arranged?

Strong Acids:

Weak Acids:

Notice the table uses the 'simple' version of the two reactions described earlier. Ex. HCl → H^+ + Cl^-
This is because the acids or bases put into water only react with water if it's the only other substance present. If a base different than water is present, it will react with the base. So the reaction above is the 'general form'.

Where are the bases on the acid/base table, and how are they arranged?

Strong Bases:

Weak Bases:

What strong bases are missing from the acid/base table?
How do these types of strong bases behave in water?

How are weak acids/bases different from strong acids/bases, and how does this affect their conductivity?

Which substances listed as bases are not bases at all?

Which substances listed as acids are not acids at all?

Where can water be found on the table?

Water is the weakest of the weak bases, so any other base present in solution will react with any acid put into the solution before water will. Water is the weakest of the weak acids, so any other acid present in solution will react with any base put into the solution before water will.

VI) Bronsted-Lowry Acids and Bases Part 2

Define the following terms:

Monoprotic Acid (_____):

Diprotic Acid (_____):

Polyprotic Acid (_____):

Amphiprotic substance:

How can you tell if a substance is amphiprotic using the acid/base table?

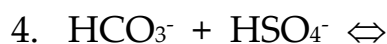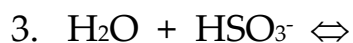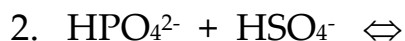List as many amphiprotic substances as you can find on the table.

For one of those substances, give an example of that substance acting as an acid, and an example of it acting as a base:

Notice that all amphiprotic substances (except for H_2O), are polyatomic groups that contain at least one proton and are negatively charged.

In a reaction between two amphiprotic substances in aqueous solution, how can you use your table to find out which substance will act as the acid and which will act as the base?

Practice: Using your table, label each reactant as an acid or base, and determine the products.
Then label each product as an acid or base *(look at the reverse reaction)*.
1. HCO_3^- + $H_2PO_4^-$ $\Longleftrightarrow$

2. HPO_4^{2-} + HSO_4^- $\Longleftrightarrow$

3. H_2O + HSO_3^- $\Longleftrightarrow$

4. HCO_3^- + HSO_4^- $\Longleftrightarrow$

Assignment 3: Hebden p.117 #12 & read p.117 & 118 & do p.119 #13, 14

Conjugate acid/base pairs are particles that directly opposite each other on the table.

Examples: conjugate acid   and its   conjugate base
H_3PO_4 $H_2PO_4^-$

Is HCl / Cl⁻ a conjugate acid/base pair?

What is the difference between a conjugate acid and its conjugate base?

A base has one _____ proton than its conjugate acid, and an acid has one _____ proton than its conjugate base.
Remember to **adjust the charge** when writing a conjugate.
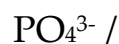Example: Write the conjugate base of NH_4^+ :

Write the conjugate base of CH_3COOH:

Write the conjugate acid of HPO_4^{2-}:

What is the conjugate base of each?
HClO₄ /   H₃BO₃ /
H₂CO₃ /   HC₂O₄⁻ /

What is the conjugate acid of each?
CN⁻ /   H₂PO₃⁻ /
NH₃ /   PO₄³⁻ /

Using your table, complete the following equation and identify the conjugate acid-base pairs.

$HCO_3^- + H_2PO_4^- \Leftrightarrow$

Remember: Strong acids (such as HCl) have a conjugate base (Cl^- - even though it's not a base at all), but they are **not** at equilibrium.
Strong bases such as NaOH have a conjugate acid (Na^+) (even though it's not an acid at all), but they are **not** at equilibrium.

Assignment 4: Complete each equation and identify conjugate acid-base pairs

1. $HNO_3 + H_2O \Rightarrow$
2. $H_2O + HNO_2 \Leftrightarrow$
3. $HIO_3 + NH_3 \Leftrightarrow$
4. $CO_3^{2-} + HF \Leftrightarrow$
5. $HS^- + H_3PO_4 \Leftrightarrow$
6. $HCO_3^- + CN^- \Leftrightarrow$
7. $H_3BO_3 + HO_2^- \Leftrightarrow$
8. $C_2O_4^{2-} + H_2O \Leftrightarrow$
9. $H_2O + H_2SO_3 \Leftrightarrow$
10. Hebden p. 121 #17-19

VIII) Determining Whether Reactants or Products are Favoured in an Acid/Base Reaction

Finish the reaction and label conjugate acids and bases:

$$HCO_3^- + HF \Leftrightarrow$$
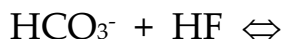
What do acids do that make them acids?

There is a competition between the two acids HF and H_2CO_3 to donate the proton, and this will have repercussions as to what side is favoured. Which of the two is the stronger acid?

So which side of the equilibrium will be favoured?

Will the K_{eq} be greater than or less than 1?

RULE: The side of the reaction with the _____ acid is always favoured.

Assignment 5: State whether reactants or products are favoured.

1. $NH_4^+ + H_2O \Leftrightarrow NH_3 + H_3O^+$

2. $H_2S + NH_3 \Leftrightarrow HS^- + NH_4^+$

3. $H_2PO_4^- + HS^- \Leftrightarrow HPO_4^{2-} + H_2S$

4. $H_2O_2 + SO_3^{2-} \Leftrightarrow HO_2^- + HSO_3^-$

5. $CH_3COOH + PO_4^{3-} \Leftrightarrow CH_3COO^- + HPO_4^{2-}$

6. $H_2PO_4^- + C_2O_4^{2-} \Leftrightarrow HPO_4^{2-} + HC_2O_4^-$

7. $H_2SO_3 + SO_4^{2-} \Leftrightarrow HSO_3^- + HSO_4^-$

IX) Strong, Weak, Concentrated, Dilute

The terms **strong** and **weak** differ from the terms **concentrated** and **dilute**. What is a strong acid, and give an example.

What is a weak acid, and give an example.

What is a concentrated acid, and give an example.

What is a dilute acid, and give an example.

The terms *strong, weak, concentrated*, and *diluted* are used for **bases** as well.
6M KOH is _____ and _____.
0.0001M KOH is _____ and _____.
6M CH₃COOH is _____ and _____.
0.0001M CH₃COOH is _____ and _____.

Notice that a strong acid can be dilute, and a weak acid can be concentrated.

X) Leveling Effect

If you had a 1M solution of each strong acid, which would be the strongest (which would create the greatest [H₃O⁺])?
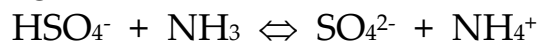
So what is the leveling effect?

What is the strongest acid that actually exists in water?

How does this compare with its position on the acid/base table?

What is the strongest base that actually exists in water?

Practice Questions:
1. Will the K_{eq} be greater or less than 1 for the following equilibrium? Why?
$$HSO_4^- + NH_3 \Leftrightarrow SO_4^{2-} + NH_4^+$$

2. Which acid has the higher $[H_3O^+]$ when reacting with water, HCN or CH_3COOH? Why?

3. Will a reaction occur between NH_2^- and $C_2O_4^{2-}$? Explain why or why not.

4. Write an equation to show the reaction between NH_2^- and water and explain why products are favoured.

Assignment 6: Hebden p. 125-126 #21-27, p. 133 #38-46

XI) The Ionization of Water

Water is amphiprotic, meaning it can act as an _____ in the presence of a base, and a base in the presence of an acid. If two water molecules collide with enough kinetic energy and with correct geometry, what can occur?

$$2H_2O_{(l)} + 59kJ \Leftrightarrow$$
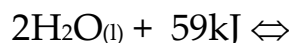
Write a K_{eq} equation for the reaction above:

The K_{eq} for this equation is called K_w, as 'w' stands for _____.

$$K_w = \underline{\hspace{3cm}} = \underline{\hspace{3cm}} \text{ at } 25°C$$

Notice how small the K_w is, meaning _____ are heavily favoured in the above reaction, which suggests…

This small concentration of ions is why water can moderately conduct.
*2 in every 550 million water molecules have an effective collision at 25°C.
Pure water is neutral. What does the term 'neutral' mean?

Since $K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$, and pure water is neutral, then $[H_3O^+] = [OH^-]$, so $[H_3O^+]$ in pure water = _____ and
$[OH^-]$ in pure water = _____
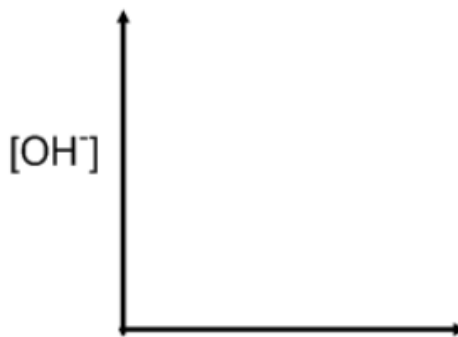$[H_3O^+]$ and $[OH^-]$ must be the same in **pure water** because every reaction between two water molecules produces _____.

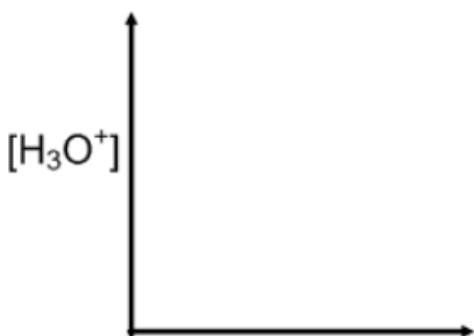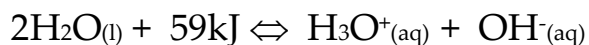If an acid is placed into water, the acid will react with water to produce _____ ions, thereby causing $[H_3O^+]$ to be _____ than $[OH^-]$. In this case, we have an _____ solution. If base is placed in water, more _____ ions will be produced, thereby creating a _____ solution.

When acid or base is placed in water, the concentration of H_3O^+ and OH^- changes, but the K_w remains constant at 1.0×10^{-14} (remember, the only thing that alters K_{eq} is _____).

Therefore, according to the K_w equation $1.0 \times 10^{-14} = [H_3O^+][OH^-]$, if one of the hydronium ion or hydroxide ion concentrations increases, the other must _____.

Investigate these scenarios using the water equilibrium and LeChatelier's Principle:

Acid added to water:

$$2H_2O_{(l)} + 59kJ \Leftrightarrow H_3O^+_{(aq)} + OH^-_{(aq)}$$

[H₃O⁺]

[OH⁻]

Base added to water:

$$2H_2O_{(l)} + 59kJ \Leftrightarrow H_3O^+_{(aq)} + OH^-_{(aq)}$$

[H₃O⁺]

[OH⁻]
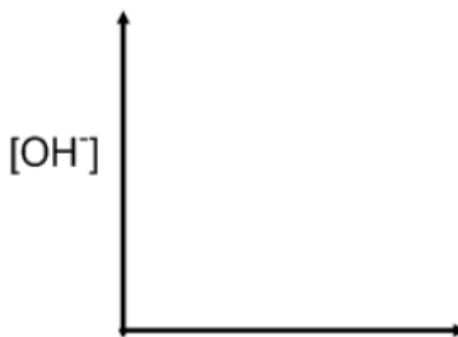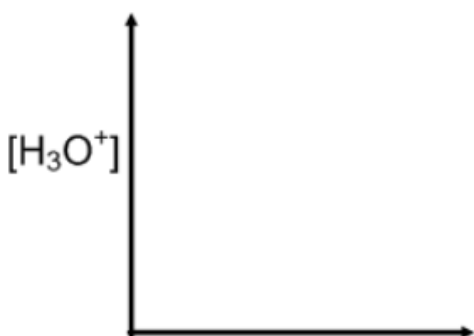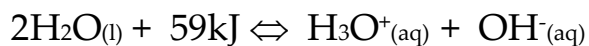
Conclusion:

In acid: $[H_3O^+]$ _____ $1.0 \times 10^{-7}M$ and $[OH^-]$ _____ $1.0 \times 10^{-7}M$

In base: $[H_3O^+]$ _____ $1.0 \times 10^{-7}M$ and $[OH^-]$ _____ $1.0 \times 10^{-7}M$

Practice Questions:

1. Calculate the $[OH^-]$ in a solution in which $[H_3O^+]$ is $1.0 \times 10^{-12}M$. Is the solution neutral, acidic, or basic?

2. Calculate the $[H_3O^+]$ in a solution in which $[OH^-]$ is $1.0 \times 10^{-8}M$. Is the solution acidic, basic, or neutral?

3. What is the $[H_3O^+]$ and $[OH^-]$ in 0.0010M HCl?
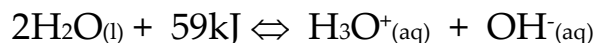
4. What is the $[H_3O^+]$ and $[OH^-]$ in $4.2 \times 10^{-2}M$ $Sr(OH)_2$?

Effect of Temperature on K_w

$$2H_2O_{(l)} + 59kJ \Leftrightarrow H_3O^+_{(aq)} + OH^-_{(aq)}$$

This reaction is _____thermic in the forward direction and

_____ in the reverse direction.

If temperature is increased, in what direction will a shift occur?

How does this affect $[H_3O^+]$ and $[OH^-]$?

How does this affect the value of K_w?

If temperature is decreased, in what direction will a shift occur?
How does this affect $[H_3O^+]$ and $[OH^-]$?
How does this affect K_w?

Example:
If pure water is heated on the stove, explain the effect on $[H_3O^+]$, K_w, and explain if it's acidic, basic, or neutral. $2H_2O_{(l)} + 59kJ \Leftrightarrow H_3O^+_{(aq)} + OH^-_{(aq)}$

Assignment 7: K_w Exercises

1. Calculate the $[OH^-]$ for solutions with the given $[H_3O^+]$. Is each solution acidic, basic, or neutral?
   a. $[H_3O^+] = 1.0 \times 10^{-3}M$
   b. $[H_3O^+] = 2.6 \times 10^{-10}M$
   c. $[H_3O^+] = 8.7 \times 10^{-7}M$
2. Calculate the $[H_3O^+]$ for solutions with the given $[OH^-]$. Is each solution acidic, basic, or neutral?
   a. $[OH^-] = 1.0 \times 10^{-2}M$
   b. $[OH^-] = 3.4 \times 10^{-6}M$
   c. $[OH^-] = 9.2 \times 10^{-9}M$
3. What is the $[H_3O^+]$ and $[OH^-]$ in 0.00345M NaOH?
4. Calculate the $[H_3O^+]$ and $[OH^-]$ in
      a. $2.5 \times 10^{-4}M$ HNO_3
      b. 5.0M HCl

c. 6.00×10^{-3}M Sr(OH)$_2$
5. Hebden p. 127 #28, 29

What does pH stand for?

pH is an indication of the acidity/basicity of a solution.
pH is the negative logarithm of [H$^+$] or [H$_3$O$^+$] in a solution: -log[H$^+$] or –log[H$_3$O$^+$]
For example, take a solution with [H$_3$O$^+$] = 1.0×10^{-7}M. The log is -7, so negative log (the pH) is –(-7), or _____.
What kind of solution has [H$_3$O$^+$] = 1.0×10^{-7}M?

What is the pH of a solution that has [H$_3$O$^+$] = 1.0×10^{-4}M, and is the solution acidic, basic, or neutral?

What is the pH of a solution that has [H$_3$O$^+$] = 1.0×10^{-11}M and is the solution acidic, basic, or neutral?
The pH scale is generally considered to be from 0 to 14 at 25°C.
pH values can sometimes be below 0 (very _____ solutions) or above 14 (very _____ solutions).

Complete the table below (Remember: K$_w$ = [H$_3$O$^+$][OH$^-$] = 1.0×10^{-14})

[H$_3$O$^+$] (M)	1.0×10^{1}	1×10^{0}	10^{-1}	10^{-2}	10^{-3}	10^{-4}	10^{-5}	10^{-6}	10^{-7}	10^{-8}	10^{-9}	10^{-10}	10^{-11}	10^{-12}	10^{-13}	10^{-14}	10^{-15}
pH	-1	0	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15
[OH$^-$]	1.0×10^{-15}	10^{-14}															
acidic, basic, or neutral?																	

Remember that [H$_3$O$^+$] and [OH$^-$] are inversely related (as one goes up, the other goes down). Thus, a high [H$_3$O$^+$] in a solution corresponds to a low [OH$^-$], as their product must always equal the K$_w$ = 1.0×10^{-14} at 25°C.

If pH decreases by 1, what happens to [H$_3$O$^+$]? [OH$^-$]?

If pH increases by 1, what happens to $[H_3O^+]$? $[OH^-]$?

What is the pH of a $1.0 \times 10^{-6}M$ H_3O^+ solution?

Is the pH of a $4.2 \times 10^{-6}M$ H_3O^+ solution greater than 6 or less than 6? How do you know?

pH is defined as -log $[H_3O^+]$, and can be found using a calculator:

How do 'sig figs' work when calculating pH?

Find the pH of each solution below with proper sig figs:

$[H_3O^+](M)$	2.15×10^{-2}	8×10^{-9}	9.334×10^{-5}	5.0×10^{-13}	3.500
pH					
A, B, or N					

$[H_3O^+]$ is calculated from pH by the following:
$$[H_3O^+] = 2^{nd} \log (-pH) \quad *2^{nd} \text{ same as } shift \text{ or } inv \text{ on calc}$$

Find $[H_3O^+]$ for each with proper sig figs:

$[H_3O^+]$ (M)					
pH	7.321	4.56	1.3	13.22	15.6257
A, B, or N					

pOH
What is pOH?

How do you calculate it if you know [OH⁻]?

How do you calculate [OH⁻] if you know the pOH?

If [H_3O^+] = 3.45 x 10⁻⁵M, find pH, [OH⁻], and pOH. Is solution A, B, or N?

If [OH⁻] = 7.2 x 10⁻³M, find pOH, [H_3O^+], and pH. Is solution A, B, or N?

Using the results of the last two examples what relationship exists between pH and pOH at 25°C?

Therefore if pH < 7, then pOH _____ and the solution is _____.
If pH > 7, then pOH _____ and the solution is _____.
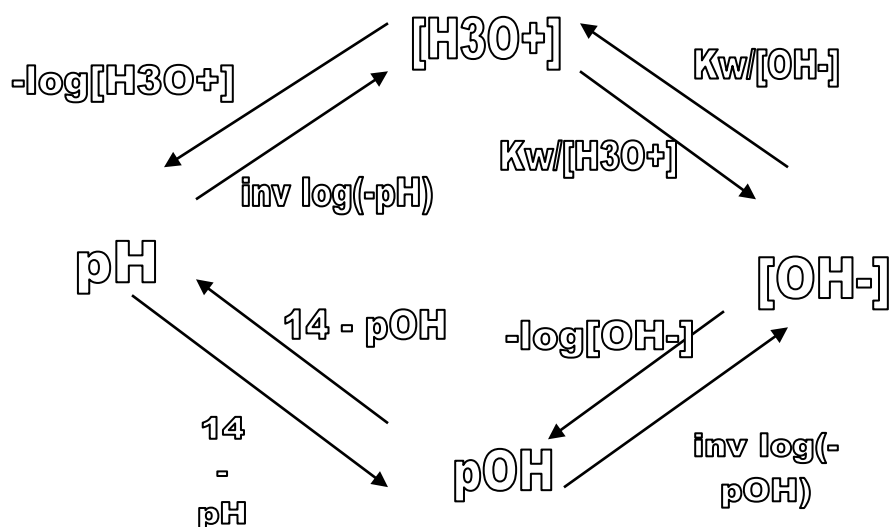
Example: Find the pH, $[H_3O^+]$, and $[OH^-]$ of a solution that has a pOH of 8.421.

Assignment 8: pH/pOH Exercises

1. Find the pH, pOH, and $[OH^-]$ if
   a) 1.0×10^{-5}M H_3O^+.
   b) 2.65×10^{-7}M H_3O^+.
   c) 6.744×10^{-12}M H_3O^+.
2. Find the $[H_3O^+]$, pOH, and $[OH^-]$ if
   a) pH = 2.35
   b) pH = 6.456
   c) pH = 10.76
3. Find the $[OH^-]$, $[H_3O^+]$, and pH if
   a) pOH = 2.34
   b) pOH = 12.59
   c) pOH = 7.10
4. Hebden p.141 #55, 56

pK$_w$

The 'p' of any value is the –log of that value. For example, the pH of $[H_3O^+]$ is $-log[H_3O^+]$, and the pOH of $[OH^-]$ is $-log[OH^-]$. Therefore, how would you calculate the pK$_w$ at 25°C if K$_w$ = 1.0×10^{-14}?

pK$_w$ =

How do pH and pOH relate to each other?

This is because when you multiply powers in math, the shortcut is to add their exponents!

Outside the pH scale (below 0 and above 14)

What is the pH of 1.00M HCl?

Therefore, what would the pH be if [HCl] > 1.00M?

Very concentrated acidic solutions (solutions that have $[H_3O^+]$ > _____M) have pH values less than 0.

Very concentrated basic solutions (solutions that have $[OH^-]$ > _____M; $[H_3O^+]$ < _____M) have pH values greater than 14.

Example: Find the pH and pOH of:

a) 5.0M HNO_3 b) 1.5M $Sr(OH)_2$

Practice Questions:

1. Find the pH of a 0.0020M solution of HNO_3.

2. Calculate the pH of a 0.010M NaOH solution.

3. If the pH is decreased from 5 to 2, what happens to the $[H_3O^+]$ and $[OH^-]$?

4. If pH is increased from 7.2 to 8.9, what happens to the $[H_3O^+]$?

5. Calculate the pH of the final solution if 100.0mL of a strong acid at pH =

4.500 is diluted by adding 50.0mL of water.

6. By how many pH units does the pH change if 80.0mL of 0.0200M HCl is diluted to a final volume of 160.0mL?

7a) Using your acid/base table for assistance, which has lower pH, a 0.01M solution of HF or a 0.01M solution of CH_3COOH? Why?

b) Which of the solutions above will conduct better? Why?

Assignment 9:

1. Hebden p. 139 #49deh, 50ef
2. Calculate the pH, pOH and $[OH^-]$ of a 0.00100M solution of HNO_3.
3. Calculate the pOH, pH, and $[H_3O^+]$ of a 2.34×10^{-4}M solution of $Ca(OH)_2$.
4. If the pH is increased from 1 to 6, how do the $[H_3O^+]$ and $[OH^-]$ change?
5. If the pH decreases from 9.3 to 6.5, how does $[H_3O^+]$ change?
6. What is the pH of the final solution if 35.00mL of a strong acid at pH 3.56 is diluted by adding 100.0mL of water?
7. You have 50.00mL of a 0.00345M solution of $HClO_4$. How does the pH change if you dilute the solution to a final volume of 175.0mL?
8. Hebden p.139 #53 & p. 141 #57
9. You dissolve 0.4g of $Ca(OH)_2$ in 500mL of solution. What is the pH?

Temperature and pH

$$2H_2O_{(l)} + 59kJ \Leftrightarrow H_3O^+_{(aq)} + OH^-_{(aq)}$$
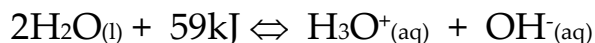
At 25°C: $K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$, so $pK_w = 14.00$

The pH scale is generally thought of from 0-14 because the pK_w is 14. However, this is only the case at 25°C? Why?

If the temperature is increased, what happens to the equilibrium and the resulting K_w? What will happen to the pH scale?

What if the temperature is decreased?

Example:
An increase in temperature to 50°C results in a K_w of 5.48×10^{-14}. Calculate the pH, pOH, $[H_3O^+]$, and $[OH^-]$ in pure water. Is the water acidic, basic, or neutral?

A sample of distilled, pure water has a pH of 7.50. Is the temperature greater than or less than 25°C? Explain.

$$2H_2O_{(l)} + 59kJ \Leftrightarrow H_3O^+_{(aq)} + OH^-_{(aq)}$$
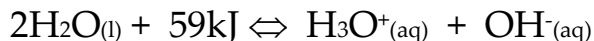
Assignment 10:

1) Hebden p. 139 #51, 52

2) Water at a certain temperature has a K_w of 4.4 x 10^{-15}.

  a) Is the water at a temperature above or below 25°C?

  b) What is the pK_w?

  c) What would the pH scale be at this temperature?

  d) Find the $[H_3O^+]$ and $[OH^-]$.

  e) Find the pH and pOH.

  f) Is water at this temperature acidic, basic, or neutral?

XIII) Mixtures of Strong Acids and Bases

Mixing an acid solution with a basic solution produces a solution that can be _____, _____, or _____ depending on the moles of acid compared to the moles of base mixed. H_3O^+ ions react with OH^- ions to make $2H_2O$ molecules, known as neutralization. But if there are more of one ion than the other, the resulting solution will not be neutral.

Example:

1. Calculate the pH of a solution obtained by adding 50.0mL of 0.10M HCl to 80.0mL of 0.15M NaOH.

2. Calculate the pH of a solution obtained by adding 1.00g of $Ca(OH)_2$ to 650.0mL of 0.800M HCl.

3. What mass of NaOH would have to be added to 500.0mL of 0.100M HCl in order to produce a solution with a pH of 3.200?

4. How many moles of HCl must be added to 40.0 mL of 0.180 M NaOH to produce a solution having a pH of 12.500? (Assume that there is no change in volume when the HCl is added).

XIV) Titrations

A **titration** is a laboratory technique that is most often used to find the concentration (molarity) of a solution. Acid/base titrations are a common type of titration in which a base is used to find an unknown acid concentration, or *visa versa*.

Suppose you are cleaning up the lab and you find a large container labeled 'hydrochloric acid', but the concentration is not given. A _____ can be done to find the unknown concentration.

Titration is a process (procedure and calculations) for determining the concentration of a substance accurately and precisely using a measurable volume of a standardized solution. A _____ is simply a reactant of known concentration.

The standardized solution (or **titrant**) used to find the concentration of hydrochloric acid would be a strong base, such as NaOH solution. The volume of NaOH added to the acid solution would be measured using a skinny tube called a **buret**.
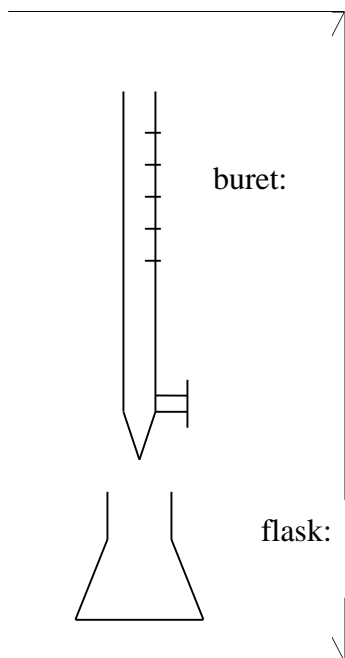
buret:

flask:

The flask contains a measured volume of the solution of unknown concentration, in this case 10.00mL of $HCl_{(aq)}$, and the buret contains a standardized base, in this case 0.10M $NaOH_{(aq)}$.
The standardized base is added from the buret to the flask. The OH⁻ from the buret reacts with the H_3O^+ from the flask to produce water. This continues until the _____ _____ is reached, the point at
which_____.
At this point, what is in the flask?

For a strong acid/strong base titration, such as the example we are investigating, the pH at the equivalence point is _____.

How do you know when the equivalence point has been reached in an acid/base titration?
An _____ is used to help visually determine when the

equivalence point has been reached in an acid/base titration. The indicator changes colour at or very near the equivalence point, signaling an end to the titration. When an indicator changes colour, it's called the _____ , or _____ _____, and this is what signals that the **equivalence point** (_____) has been reached.
The chemistry of indicators will be studied in the Acid/Base II unit.

Since at the equivalence point, moles of OH⁻ = moles of H_3O^+, if we can calculate the moles of OH⁻ that vacated the buret, it will be equal to the moles of H_3O^+ that were originally in the flask (since H_3O^+ and OH⁻ react one to one). The [NaOH] = 0.10M, and we kept track of the volume of NaOH that vacated the buret (using the scale on the buret), therefore we can find the moles of OH⁻ that went into the flask. If we stop the titration at the equivalence point, the moles of OH⁻ that went into the flask will be equal to the moles of H_3O^+ in the flask. Since we know the original volume of H_3O^+ in the **flask**, we can calculate the unknown [HCl] on the next page:

Now, the HCl solution is standardized, as we know the concentration.

Practice Questions:

1. A 10.00mL sample of an unknown concentration of $LiOH_{(aq)}$ is titrated using 23.62mL of 0.150M HNO_3. Determine [LiOH].

2. 37.86mL of 0.250M NaOH was required to neutralize a 20.0mL sample of HF. Calculate the [HF]. *Even though HF is a weak acid and in water it will only dissociate under 5%, in the presence of a strong base such as NaOH, it will react 100%.

3. A 15.0mL sample of unknown [Sr(OH)₂] was titrated using 18.56mL of 0.350M HNO_3. Find [Sr(OH)₂].

4. 50.78mL of 0.020M $Ba(OH)_2$ was required to neutralize a 30.0mL sample of H_2SO_4. Find $[H_2SO_4]$.

Assignment 12: Titration Exercises

1. Find the concentration of an HCl solution if 25.00mL is titrated with 28.46mL of a 0.105M standardized solution of NaOH.
2. You titrated a 30.0mL solution of HNO_3 with 23.75mL of a 0.25M standardized solution of KOH. What is the $[HNO_3]$?
3. A 35.00mL unknown solution of LiOH is titrated with 17.67mL of 0.200M HI. What is the [LiOH]?
4. A 24.00mL sample of H_2SO_4 is titrated with 32.43mL of 0.150M NaOH solution. Find $[H_2SO_4]$.
5. A 40.00mL sample of $Ca(OH)_2$ is titrated with 16.55mL of 0.100M HCl. Find $[Ca(OH)_2]$.
6. A 20.00mL sample of H_3PO_4 is titrated with 25.76 mL of a 0.100M $Ba(OH)_2$ solution. Find $[H_3PO_4]$.

Making Standardized Solutions

How could you make 1.0L of a 0.50M solution of NaOH in the lab? NaOH originates as solid white pellets.

This method, though sound for making many types of solutions, would actually produce an NaOH solution that is slightly less than 0.50M (probably about 0.48M). This is because NaOH pellets actually absorb water, and so the mass of NaOH you measure is not all due to NaOH; some is due to water absorbed onto the pellets. This problem is the case for many acids and bases, which makes it very hard to create an accurate standardized solution from scratch. These acids and bases are **hygroscopic**, meaning that they absorb water.

There are a few acids and bases that are **non-hygroscopic**, meaning they are pure and dry acids or bases and can be used to make solutions with accurate concentrations. Non-hygroscopic acids and bases are known as **primary standards**, and are used to make standardized solutions.

Examples: Primary Standard Base: sodium carbonate (Na_2CO_3)
            Primary Standard Acids: potassium hydrogen phthalate
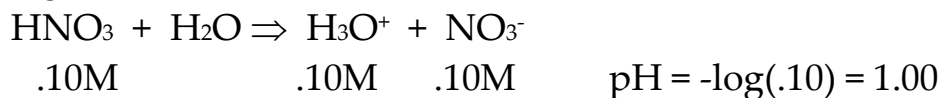   oxalic acid ($H_2C_2O_4$)

Once a primary standard of known concentration is accurately prepared in the lab, it can be used to **standardize** any other acid or base solution by titration. For example, oxalic acid is a primary standard acid, and once an accurate standardized solution of it is prepared (using your method from the top of the page), it can be used to standardize any basic solution by titration. Then, that same basic solution that is now standardized can be used to titrate an unknown concentration of any acid, thereby standardizing that acid solution, and so on.

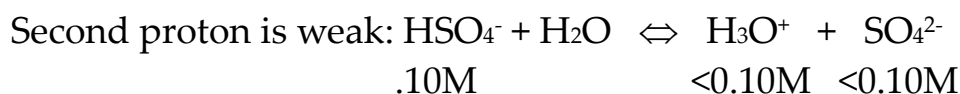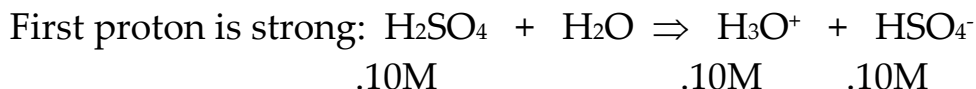**Calculating Unknown Volume by Titration**

What volume of 0.0350M $Ba(OH)_2$ will be required to neutralize 50.0mL of 0.0275M HCl?

Assignment 13: Hebden p.158 #96, 97, 106, p.165 #122

Helpful Equations – A Summary

1. Strong acid in water: 100% dissociation

$HNO_3 + H_2O \Rightarrow H_3O^+ + NO_3^-$

.10M   .10M   .10M   pH = -log(.10) = 1.00

H_2SO_4 in water:

First proton is strong: $H_2SO_4 + H_2O \Rightarrow H_3O^+ + HSO_4^-$

.10M   .10M   .10M

Second proton is weak: $HSO_4^- + H_2O \Leftrightarrow H_3O^+ + SO_4^{2-}$

.10M   <0.10M   <0.10M

2. Weak acid in water: not 100% dissociation

$H_3PO_4 \quad + \quad H_2O \Leftrightarrow \quad H_3O^+ \quad + \quad H_2PO_4^-$

.10M $2.79 \times 10^{-2}M$ $2.79 \times 10^{-2}M$

pH = -log(2.79×10^{-2}) = 1.56
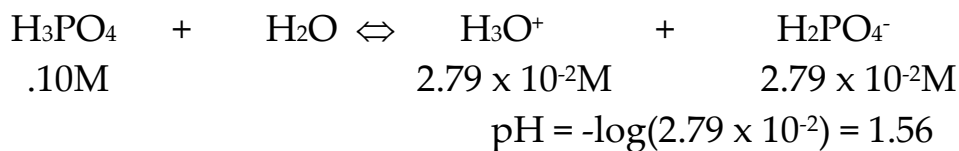
3. Strong base (hydroxide base) in water: 100% dissociation (water not in reaction; just a dissociation)

$Sr(OH)_2 \quad \Rightarrow \quad Sr^{2+} \quad + \quad 2OH^-$

.10M   .10M   .20M

pOH = -log(.20) = 0.70; pH = 13.30
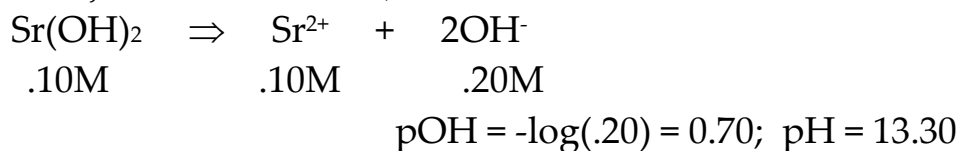
4. Weak base in water: not 100% dissociation

$NH_3 \quad + \quad H_2O \quad \Leftrightarrow \quad NH_4^+ \quad + \quad OH^-$

.10M $1.34 \times 10^{-3}M$ $1.34 \times 10^{-3}M$

pOH = -log(1.34×10^{-3}) = 2.87

so, pH = 11.13
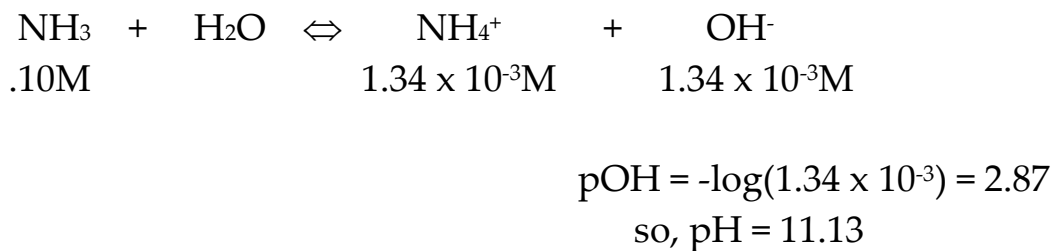
5. ANY reaction that involves a STRONG acid or base goes to 100% completion. So a weak acid with a strong base is 100% due to the strong base. A strong acid with a weak base is 100% due to the strong acid.

Example: $H_3PO_4 + 3NaOH \Rightarrow 3H_2O + Na_3PO_4$

       weak   strong
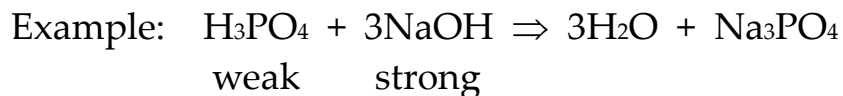
The OH$^-$ ions take all three protons off of each H_3PO_4 molecule, such as in a titration. If H_3PO_4 was merely in water, only one proton would come off at less than 100% like #2 above.

6. Weak acid and weak base:

$NH_4^+ + SO_4^{2-} \Leftrightarrow NH_3 + HSO_4^-$ side with weaker acid is favoured