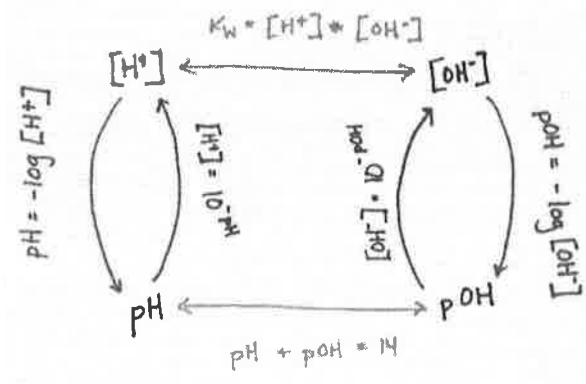
Titration Lab Practical
Review Packet



1) Explain why H^+ ions don't actually exist in an aqueous solution.

As soon as H+ is donated from an acid, it quickly pairs with a water molecule, forming H_3O^+. Therefore H^+ & H_3O^+ mean the same thing.

$$H_3O^+ \xleftarrow{+H_2O} H^+ \leftarrow HCl \rightarrow Cl^-$$

2) Describe what it means for a solution to be acidic, basic, and neutral using the terms hydromium (H_3O^+), hydroxide (OH^-) and pH.

Acidic: High $[H_3O^+]$, low $[OH^-]$, low pH (below 7)

Basic: Low $[H_3O^+]$, high $[OH^-]$, high pH (above 7)

Neutral: Equal $[H_3O^+]$ & $[OH^-]$, pH = 7

3) How can you identify acids and bases from their chemical formula?
- Acids have "H" listed first in the formula (HCl, HF, H_2SO_4)
- Bases have "OH" listed last in the formula ($NaOH$, KOH, $Ca(OH)_2$)

4) Name the following acids: HF, H_2SO_4, HNO_3

HF = hydrofluoric acid
H_2SO_4 = sulfuric acid
HNO_3 = nitric acid

5) Write the formulas for the following acids: hydrobromic acid, acetic acid, phosphoric acid → Balance your charges!

HBr $HC_2H_3O_2$ H_3PO_4

6) Show you understand "dissociation" by:

a. Showing how HBr and $Ca(OH)_2$ would dissociate in water

H^+ Br^- Ca^{2+} $2OH^-$

b. Describing the difference between strong acids and weak acids.

- Strong acids completely dissociate (every molecule splits) and therefore donate a lot of H+ into solution
- Weak acids only partially dissociate (1 in every 100 molecules may split) and therefore donate less H+ into solution.
- More H+ in solution = lower pH, so strong acids will have lower pH values than weak acids.

7) Determine the $[H_3O^+]$ and $[OH^-]$ for the following solutions:

a. 0.45M sulfuric acid $= 0.45\ M\ H_2SO_4 \left(\frac{2\ H^+}{1\ H_2SO_4}\right) = \boxed{0.90\ M\ H^+}$

$$[H^+][OH^-] = 1\times10^{-14}$$

$$[OH^-] = \frac{1\times10^{-14}}{0.90} = \boxed{1.1\times10^{-14}\ M\ OH^-}$$

b. 0.76M potassium hydroxide $= 0.76\ M\ KOH \left(\frac{1\ OH^-}{1\ KOH}\right) = \boxed{0.76\ M\ OH^-}$

$$[H^+][OH^-] = 1\times10^{-14}$$

$$[H^+] = \frac{1\times10^{-14}}{0.76} = \boxed{1.3\times10^{-14}\ M\ H^+}$$

c. 3.2 g of barium hydroxide dissolved in water to make a 300.0mL solution

$$3.2\ g\ Ba(OH)_2 \left(\frac{1\ mol\ Ba(OH)_2}{171.35\ g}\right)\left(\frac{2\ mol\ OH^-}{1\ mol\ Ba(OH)_2}\right) = 0.037\ mol\ OH^-$$

$$[OH^-] = \frac{mol}{L} = \frac{0.037\ mol\ OH^-}{0.300\ L} = \boxed{0.12\ M\ OH^-}$$

$$[H^+] = \frac{1\times10^{-14}}{0.12} = \boxed{8.0\times10^{-14}\ M\ H^+}$$

8) Determine the pH & pOH of the following solutions:

a. 0.075M hydrochloric acid $= 0.075\,M\ HCl \left(\frac{1\ H^+}{1\ HCl}\right) = 0.075\ M\ H^+$

$$pH = -\log[H^+] = -\log(0.075) = \boxed{1.1}$$

$$pOH = 14 - pH = 14 - 1.1 = \boxed{12.9}$$

b. 0.135M calcium hydroxide $= 0.135\ M\ Ca(OH)_2 \left(\frac{2\ OH^-}{1\ Ca(OH)_2}\right) = 0.270\ M\ OH^-$

$$pOH = -\log[OH^-] = -\log(0.270) = \boxed{0.57}$$

$$pH = 14 - pOH = 14 - 0.57 = \boxed{13.4}$$

c. 2.89g of sodium hydroxide dissolved to make a 250.0mL solution

$$2.89\,g\ NaOH \left(\frac{1\ mol\ NaOH}{39.99\ g}\right)\left(\frac{1\ mol\ OH^-}{1\ mol\ NaOH}\right) = 0.0723\ mol\ OH^-$$

$$[OH^-] = \frac{0.0723\ mol}{0.2500\ L} = 0.289\ M\ OH^-$$

$$pOH = -\log[OH^-] = -\log(0.289) = \boxed{0.539}$$

$$pH = 14 - pOH = 14 - 0.539 = \boxed{13.5}$$

9) Complete the following table using your pH equations:

$[H_3O^+]$	$[OH^-]$	pH	pOH	Acid/Base/Neutral
1.2×10^{-7}	8.3×10^{-8}	6.9	7.1	Neutral
3.1×10^{-5}	3.2×10^{-10}	4.5	9.5	Acid
5.0×10^{-3}	2.0×10^{-12}	2.3	11.7	Acid
1.3×10^{-11}	7.9×10^{-4}	10.9	3.1	Base

10) Write a complete equation for the following reactions.

a. Hydrochloric acid reacts with calcium hydroxide

$$2HCl\ (aq) + Ca(OH)_2\ (aq) \longrightarrow CaCl_2\ (aq) + 2H_2O\ (l)$$

b. Nitric acid reacts with sodium carbonate.

$$2HNO_3\ (aq) + Na_2CO_3\ (s) \longrightarrow 2NaNO_3\ (aq) + H_2O\ (l) + CO_2\ (g)$$

$$H_2SO_4\ (aq) + 2NaOH\ (aq) \longrightarrow 2Na_2SO_4\ (aq) + 2H_2O\ (l)$$

11) A reaction is performed between 150.0mL of 0.055M sulfuric acid and 275.0mL of 0.090M sodium hydroxide.

150.0mL $\downarrow$ 0.1500 L

275.0mL $\downarrow$ 0.2750 L

a. Which solution, the acid or the base, was completely neutralized? Support with a calculation showing the moles of H^+ and OH^- combined.

$$0.1500\ L\ H_2SO_4 \left(\frac{0.055\ mol\ H_2SO_4}{1\ L}\right)\left(\frac{2\ mol\ H^+}{1\ mol\ H_2SO_4}\right) = 0.0165\ mol\ H^+$$

H^+ & OH^- combine 1:1, therefore OH^- is excess

$$0.2750\ L\ NaOH \left(\frac{0.090\ mol\ NaOH}{1\ L}\right)\left(\frac{1\ mol\ OH^-}{1\ mol\ NaOH}\right) = \boxed{0.0248\ mol\ OH^-}$$

Excess base

b. How much (moles) excess ion would be left over?

0.0248 mol OH^- (have)

− 0.0165 mol OH^- (used to neutralize H^+)

0.0083 mol OH^- (left over)

c. Would the resulting mixture be acidic or basic? Support with a calculation of the pH of the mixed solution.*

Basic (extra base)

$$[OH^-] = \frac{mol}{L} = \frac{0.0083\ mol\ OH^-}{(0.15L + 0.275L)}$$

* use total (combined) volume

$$= 0.0195\ M\ OH^-$$

$$pOH = -\log[OH^-] = -\log(0.0195) = 1.7$$

$$pH = 14 - pOH = 14 - 1.7 = \boxed{12.3}$$

pH is very basic

12) A tanker truck carrying 43000.0L of 1.00M hydrochloric acid crashes and spills its contents into Lake Kegonsa.

a. If the total volume of the lake after the spill is 6.70×10^7 L, what would be the pH of the lake after the spill?

$pH = -\log[H^+]$, so need to solve for $[H^+]$ in the lake

$$43000\ L\left(\frac{1.00\ mol\ HCl}{1\ L}\right)\left(\frac{1\ mol\ H^+}{1\ mol\ HCl}\right) = 43{,}000\ mol\ H^+$$

$$[H^+] = \frac{mol\ H^+}{L} = \frac{43{,}000\ mol\ H^+}{6.70\times10^7\ L} = 6.41\times10^{-4}\ M\ H^+$$

$$pH = -\log[H^+] = -\log(6.41\times10^{-4}) = \boxed{3.19}$$

b. Theoretically, how many grams of NaOH would you dump into the lake from a helicopter to perfectly neutralize the lake? @ neutral $OH^- = H^+$, want to add 43,000 mol OH^-

$$43{,}000\ mol\ OH^-\left(\frac{1\ mol\ NaOH}{1\ mol\ OH^-}\right)\left(\frac{39.99\ g}{1\ mol\ NaOH}\right) = \boxed{1.72\times10^6\ g\ NaOH}$$

c. If solid sodium hydroxide isn't available, how many liters of a 3.0M NaOH solution should be poured into the lake to neutralize it?

$$43{,}000\ mol\ OH^-\left(\frac{1\ mol\ NaOH}{1\ mol\ OH^-}\right)\left(\frac{1\ L}{3.0\ mol\ NaOH}\right) = \boxed{14{,}300\ L\ of\ NaOH}$$

d. The Crisis Response team panics, does sloppy math, doesn't pay attention to sig figs, and drops 2×10^6g of sodium hydroxide into the lake. What is the pH after trying to neutralize the spill?

① $$2\times10^6\ g\ NaOH\left(\frac{1\ mol\ NaOH}{39.99\ g}\right)\left(\frac{1\ mol\ OH^-}{1\ mol\ NaOH}\right) = 50{,}013\ mol\ OH^-$$

② 50,013 mol OH^- (have)
− 43,000 mol OH^- (used to neutralize H^+)
7,013 mol OH^- (left-over)

③ $$[OH^-] = \frac{7{,}013\ mol\ OH^-}{6.70\times10^7\ L} = 1.05\times10^{-4}\ M\ OH^-$$

④ $pOH = -\log(1.05\times10^{-4}) = 3.98$

⑤ $pH = 14 - 3.98 = \boxed{10.02}$

13) Explain the role of indicators while comparing and contrasting phenolphthalein and universal indicator.

Indicators provide information on the pH of a solution via color changes.

phenol: clear = acid, pink = base. Useful because shows equivalence point where $\underset{mol}{H^+} = \underset{mol}{OH^-}$ when color flips.

Universal: cool colors = basic, warm = acidic. Useful b/c it gives an estimated pH value from 1-14.

14) Write a procedure for determining the concentration of a $Ca(OH)_2$ solution using titration.

1) Rinse & fill buret with an acid of known concentration
2) Pipette the desired volume of unknown base into a clean, dry beaker
3) Add 2-3 drops of phenolphthalein indicator to the base (unknown)
4) Slowly add acid (known) to the base (unknown) until the pink color permanently fades to clear, stir the solution the entire time.
5) Determine the volume of known acid required to neutralize the unknown base.
6) Calculate the [base].

15) A calcium hydroxide solution of unknown concentration is titrated with 0.500M hydrochloric acid. If the neutralization point is reached after delivering 44.2mL of acid, what is the concentration of the base?

Forgot base volume, use 10.0 mL

$$\underset{\text{acid (HCl)}}{0.0442\text{ L}}\left(\frac{0.500\text{ mol HCl}}{1\text{ L}}\right)\left(\frac{1\text{ mol H}^+}{1\text{ mol HCl}}\right) = 0.0221\text{ mol H}^+$$

@ neutral $\underset{mol}{H^+} = \underset{mol}{OH^-}$

$$0.0221\text{ mol OH}^-\left(\frac{1\text{ mol Ca(OH)}_2}{2\text{ mol OH}^-}\right) = 0.01105\text{ mol Ca(OH)}_2$$

$$[Ca(OH)_2] = \frac{mol}{L} = \frac{0.01105\text{ mol}}{0.0100\text{ L}} = \boxed{1.11\text{ M Ca(OH)}_2}$$

16) Explain the effect the following lab errors would have on the calculated concentration for the base in the previous example.

a. The stopcock is not closed right at the point of neutralization, allowing a couple more drops of acid to drip into the base.

Added more acid than needed → volume of acid you record is more than actually needed to get to neutral → will make your calculation for the unknown base a stronger concentration than what it really is.

b. Some water is left in the buret before you put the acid in it to use in the titration.

Dilutes the known → will need to use more acid volume than you would have to deliver 0.0221 mol H^+ → makes calc. of base stronger concentration than what it is.

c. Some water is left in the beaker you put your unknown base in.

No effect, you know the volume of base you pipetted and therefore delivered the correct # of moles of base to the beaker to be neutralized. Water won't add/remove base