

Admin:
Return tests

last time:

- 0) Thermal analysis
- 1) liquid solutions
- 2) net ionic equations (review)
- 3) acids

Today:

- 1) Bases
- 2)

Lecture:

Acids and Bases:

0) acids = donate H^+ ions (hydronium ions)
hydronium ions = H_3O^+
(i.e. an H^+ on top of a H_2O molec)

know acid's names: (usually end with "-ic")

Examp les	HCl	HNO ₃	H ₃ PO ₄
Names	Hydrochlo ric acid	Nitric acid	Phospho ric acid
Type	Monoproti c	Monopro tic	triprotic
Strengt h	Strong	Strong	weak

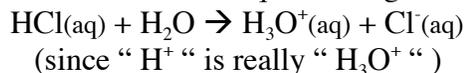
Examples	H ₂ SO ₄	HClO ₄	H ₂ CO ₃
Names	sulfuric acid	perchloric acid	carbonic acid
Type	diprotic	monoprotic	diprotic
Strength	Strong	Strong	weak

Strong acids: example: HCl
is in solution, it dissociates 100 % into
 H^+ ions and Cl^- ions

It is convenient to write:



The more correct equation is given below:



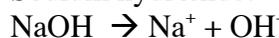
1) **bases** = donate OH^- (*Arrhenius* definition)

=accept H^+ ions (*Bronsted-Lowry* definition)

Examples	NaOH	KOH	$\text{Ca}(\text{OH})_2$
Names	Sodium hydroxide	Potassium hydroxide	Calcium hydroxide
Type	Monobasic	Monobasic	dibasic
Strength	Strong	Strong	weak

Consider a **strong** base:

Sodium hydroxide:



Review Bronsted-Lowry definition of a base: a proton-acceptor!

To illustrate this, consider ammonia, a weak base:



(NH_3 accepts a proton from H_2O and is therefore a Bronsted-Lowry base)

Know difference between **acids** and **bases**.

But know also the difference between

strong

acids and **weak acids**. (It is not a matter of *concentration* but of *degree of dissociation*)

Strong acids dissociate 100%

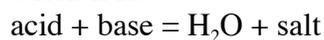
Weak acids don't dissociate 100%

(in fact, usually less than 1% dissociation)

3) Neutralization reactions:

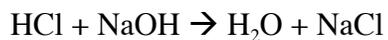
Acids and bases neutralize each other!

Often it is:

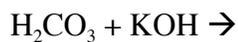


(salt = an soluble ionic compound)

classic example:



Complete the following:



Write down the net ionic equations!

4) **Concentrations:** important to measure concentration quantitatively.

There are several units of concentration but the most important is the *molar*, M.

$$\mathbf{a) \text{ Molarity (M)} = \frac{\text{moles solute}}{\text{liters of solution}}}$$

Or sometimes convenient to use:

$$\mathbf{M} = \frac{\text{mmol solute}}{\text{mL solution}}$$

Example: what is the molarity of a 100 mL solution containing 0.23 moles of acetic acid, HAc ?

$$\begin{aligned} \# \text{ M} &= \frac{0.23 \text{ mol HAc}}{100 \text{ mL} \cdot \frac{1 \text{ L}}{1000 \text{ mL}}} = 2.3 \frac{\text{mol HAc}}{\text{L}} \\ &= 2.3 \text{ M} \end{aligned}$$

Example: how many grams of NaCl
(58.5g/mol) are in 25.0 mLs of a 0.110M
NaCl solution?

$$\begin{aligned} \#g \text{ NaCl} &= 25.0 \text{ mL} \left(\frac{1 \text{ L}}{1000 \text{ mL}} \right) \left(\frac{0.110 \text{ mol NaCl}}{\text{L}} \right) \left(\frac{58.5 \text{ g NaCl}}{\text{mol NaCl}} \right) = \\ &= 0.161 \text{ g NaCl} \end{aligned}$$

b) *Preparing solutions!*

First determine the grams of solute needed

c) *Diluting solutions!!!*

$$M_1 V_1 = M_2 V_2$$

5) Titration

During titration, solutions are added until they JUST completely neutralize the other.

end point or *equivalence point* = point where neutralization is just completed.

Indicator = reagent which changes color to show endpoint is reached. Example:
phenolphthalein (pink = base, colorless=acid)

at the **endpoint**: (“equivalence point”)

moles H^+ donated = # moles H^+ accepted

equivalents H^+ = # equivalents H^+ accepted.

“equivalents” = # moles of H^+ accepted or donated.

$$\text{Molar (M)} = \frac{\text{moles solute}}{\text{liters of solution}} \quad (\text{units: } \frac{\text{mol}}{\text{L}})$$

$$\text{Normal (N)} = \frac{\text{equiv. solute}}{\text{liters of solution}} \quad (\text{units: } \frac{\text{eq.}}{\text{L}})$$

$$0.10 \text{ M HCl} = .1 \text{ mol HCl/L}$$

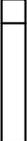
but each HCl produces a H^+

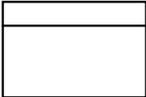
$$\text{so, } .10 \text{ M HCl} = .1 \text{ N HCl}$$

$$0.25 \text{ M H}_2\text{SO}_4 = \text{ ______ } \text{ N H}_2\text{SO}_4$$

6) Titration examples:

example: volumetric analysis of a solution containing acid or base

 ← .100 M HCl
suppose it takes 20 mL to reach end point (e.p.)

 25 mL of NaOH
where $[NaOH]=?$

End point is reached when neutralization is "just completed".

2 ways to solve for [NaOH]:

first way, direct stoichiometry:

$$[\text{NaOH}] = 20.0 \times 10^{-3} \text{ L HCl} \times (.10 \text{ mol HCl/L HCl}) (1 \text{ mol NaOH/mol HCl}) / 25 \times 10^{-3} \text{ L NaOH}$$

$$= 0.080 \text{ M NaOH}$$

second way:

In acid-base titrations, the end point is reached when

#mol H⁺ donated = #mol H⁺ accepted

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

$$(0.10 \text{ M})(20 \text{ mL}) = M_{\text{OH}^-}(25 \text{ mL})$$

$$M_{\text{OH}^-} = .10 \text{ M}(20/25) = .080 \text{ M}$$

How about H₂SO₄ vs NaOH?



Suppose

 ← .100 M H₂SO₄
suppose it takes 20 mL to reach end point (e.p.)

 25 mL of NaOH
[NaOH] = _____ ?

Concentrations:

Molarity (M) = $\frac{\text{mol solute}}{\text{Liter of sol'n}}$ = mmol/mL

Review:

a) *Preparing solutions!*

First determine the grams of solute needed

b) *Diluting solutions!!!*

$$M_1 V_1 = M_2 V_2$$