Today

Titration

determining something about an unknown by reacting it with a known solution

Neutralization (again)

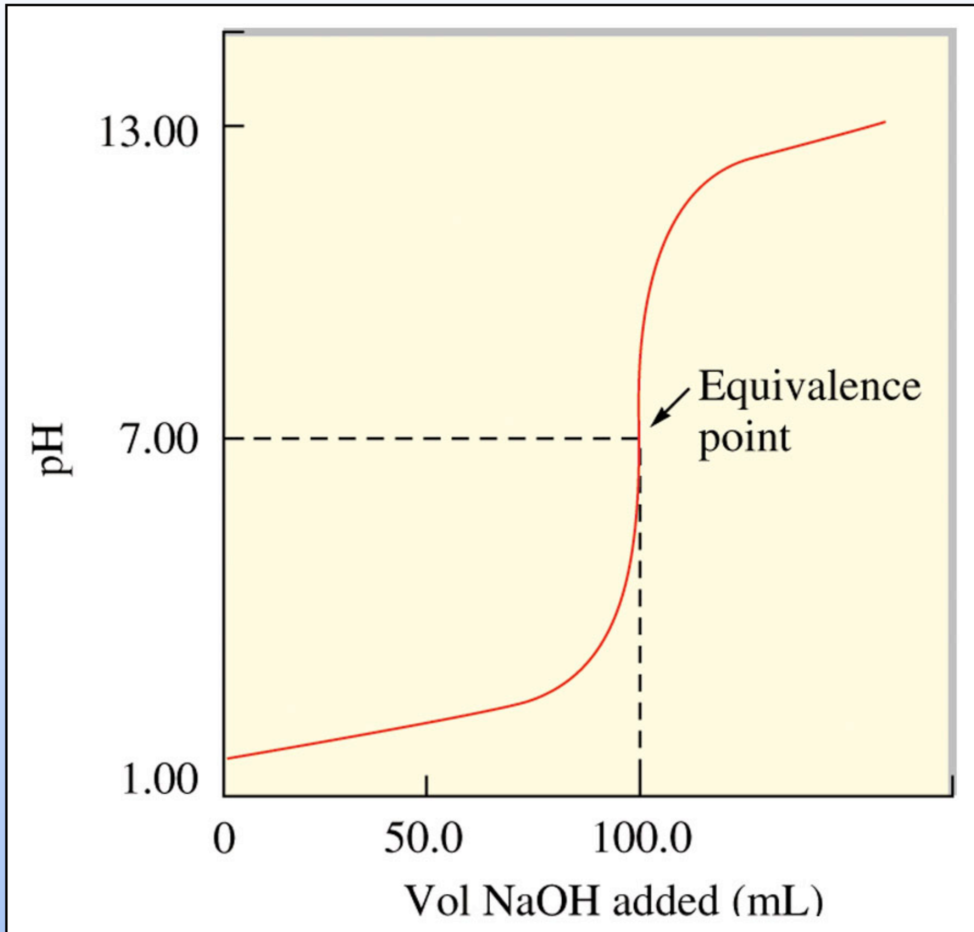
we'll need this to figure out titration

Solubility

The easiest of all the equilibria

we'll need this for polyprotic acids

# Last Time Strong Acid/Strong Base Titration




original solution 50 mL HCl
adding .1 M NaOH
at equivalence point

same number of moles of base
$.1L \times .1M = 0.01$ moles $OH^-$

therefore the solution originally
had 0.01 moles $H^+$

concentration was .2 M

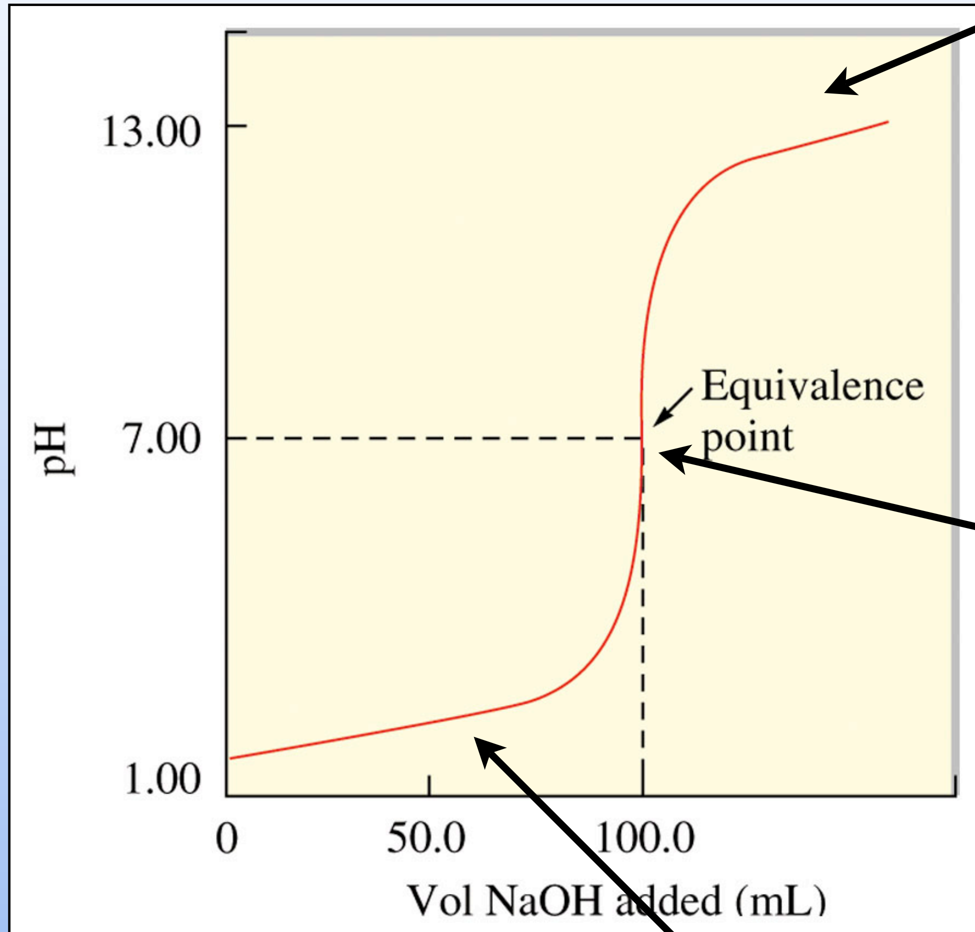
at the equivalence point we have
equal number of moles of acid and base

# Neutralize first
# Then look at the equilibrium

imagine a 100 mL solution with 0.1 moles of HCl
we add .01 moles of NaOH in each titration step (10 mL of 1M)

| Initial | | After Neutralization | | Volume (L) | Equilibrium | |
|---|---|---|---|---|---|---|
| mol $H^+$ | mol $OH^-$ | mol $H^+$ | mol $OH^-$ | | pH | pOH |
| 0.1 | 0.01 | 0.09 | 0.00 | 0.11 | 0.09 | 13.91 |
| 0.09 | 0.01 | 0.08 | 0.00 | 0.12 | 0.18 | 13.82 |
| 0.08 | 0.01 | 0.07 | 0.00 | 0.13 | 0.27 | 13.76 |
| ......... | | | | | | |
| 0.02 | 0.01 | 0.01 | 0.00 | 0.19 | 1.28 | 12.72 |
| 0.01 | 0.01 | 0.00 | 0.00 | 0.20 | 7.00 | 7.00 |
| 0.0 | 0.01 | 0.0 | 0.01 | 0.21 | 12.67 | 1.33 |
| 0.0 | 0.02 | 0.0 | 0.02 | 0.22 | 12.86 | 1.04 |

# Strong Acid/Strong Base Titration




Strong Base

Neutral

Strong Acid

## What volume of a 1 M NaOH will you need to add to 200 mL of a 0.2 M solution of HCl to neutralize it?

A. 10 mL

B. 20 mL

C. 30 mL

D. 40 mL

E. 400 mL

There are .04 moles of $H^+$ .2M x .2L
to neutralize you'll need .04 moles of $OH^-$

For that you'll need .04L of a 1M solution

At the endpoint of your titration you have added 40 mL of a 1M NaOH solution to 200 mL of an unknown HCl solution. What was the concentration of the HCl?

A. 0.1 M

B. 0.2 M

C. 0.4 M

D. 1 M

E. 2 M

At the endpoint there are equal moles $H^+$ and $OH^-$

0.04 L x 1 M = 0.04 moles $OH^-$

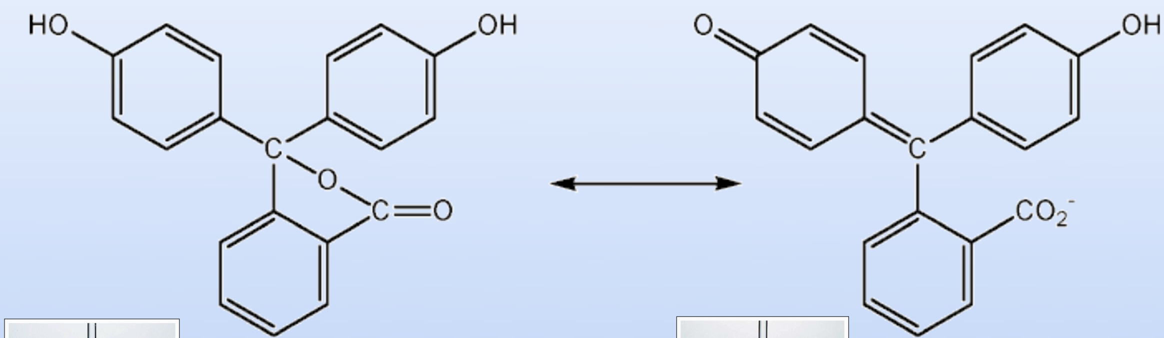
.04 moles $H^+$/0.2L = 0.2 M

# Finding the endpoint (equivalence point)

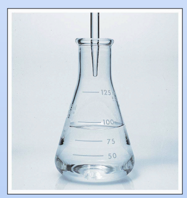
## Indicator dye

Phenolphthalein

pKa = 8.2 $K_a = 6.3 \times 10^{-9}$




amount of indicator is so small it doesn't affect the pH, but the equilibrium of the dye is strongly affected by the pH



Colourless
HA

Pink
$A^-$

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

$$= [H^+] \times \frac{\text{Pink}}{\text{Clear}}$$

$[H^+] > 6.3 \times 10^{-9}$

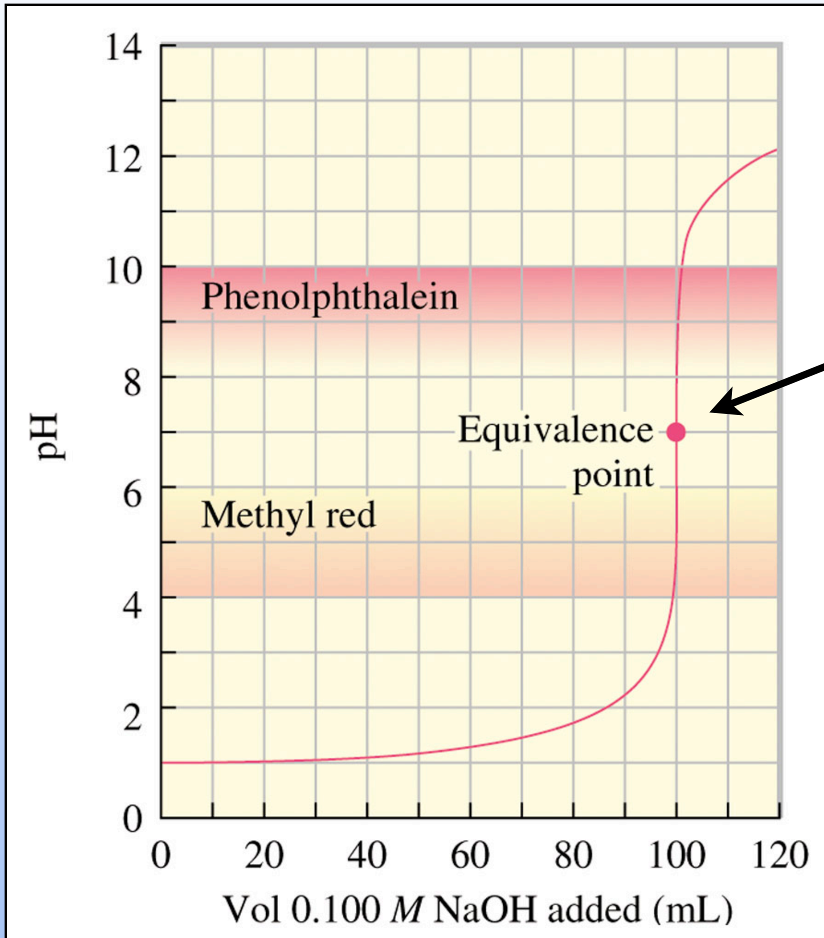
pH < 8.2

$$\frac{\text{Pink}}{\text{Clear}} < 1$$

$[H^+] < 6.3 \times 10^{-9}$

pH>8.2

$$\frac{\text{Pink}}{\text{Clear}} > 1$$




color just barely changing for Phenolpthalein

Bromophenol Blue has a $pK_a$ of around 4. When it is protonated it is green, when it is deprotonated it is blue.

What color would in be in a solution in which the pH was 8?

A. blue

B. green

C. a mix of blue and green

$pK_a = 4 \quad K_a = 10^{-4}$

$pH = 8 \; [H^+] = 10^{-8}$

$[H^+] \ll K_a \qquad [A^-]/[HA] \gg 1$

all deprotonated

blue

# Titration with weak acid/base

All the same
Neutralize first
Then equilibrium

Neutralization reactions

$$H^+(aq) \quad + \quad OH^-(aq) \longleftrightarrow H_2O(l)$$

$$H^+(aq) \quad + \quad A^-(aq) \longleftrightarrow HA(aq)$$

$$H^+(aq) \quad + \quad B(aq) \longleftrightarrow BH^+(aq)$$

$$OH^-(aq) \quad + \quad HA(aq) \longleftrightarrow A^-(aq)$$

$$OH^-(aq) \quad + \quad BH^+(aq) \longleftrightarrow B(aq)$$

I have a 100 mL of a 1 M solution of acetic acid
I add 100 mL of 0.5 M NaOH
What remains in the solution?

A. 0.1 moles

B. 0.1 moles

C. 0.05 mole

D. 0.05 mole

E. 0.1 moles

.1L x 1 M = 0.1 moles acetic acid
.1L x .5M = 0.05 moles of $OH^-$

neutralize $OH^-$ and HA

left with .05 moles of HA
and 0.05 moles of $A^-$

If I have a solution that has 0.05 moles of acetic acid and 0.05 moles of acetate what do I have?

A. strong acid solution

B. weak acid solution

C. buffer

D. weak base

E. strong base

Neutralization of a weak acid or weak base
will yield a buffer
because you generate
the conjugate base or acid

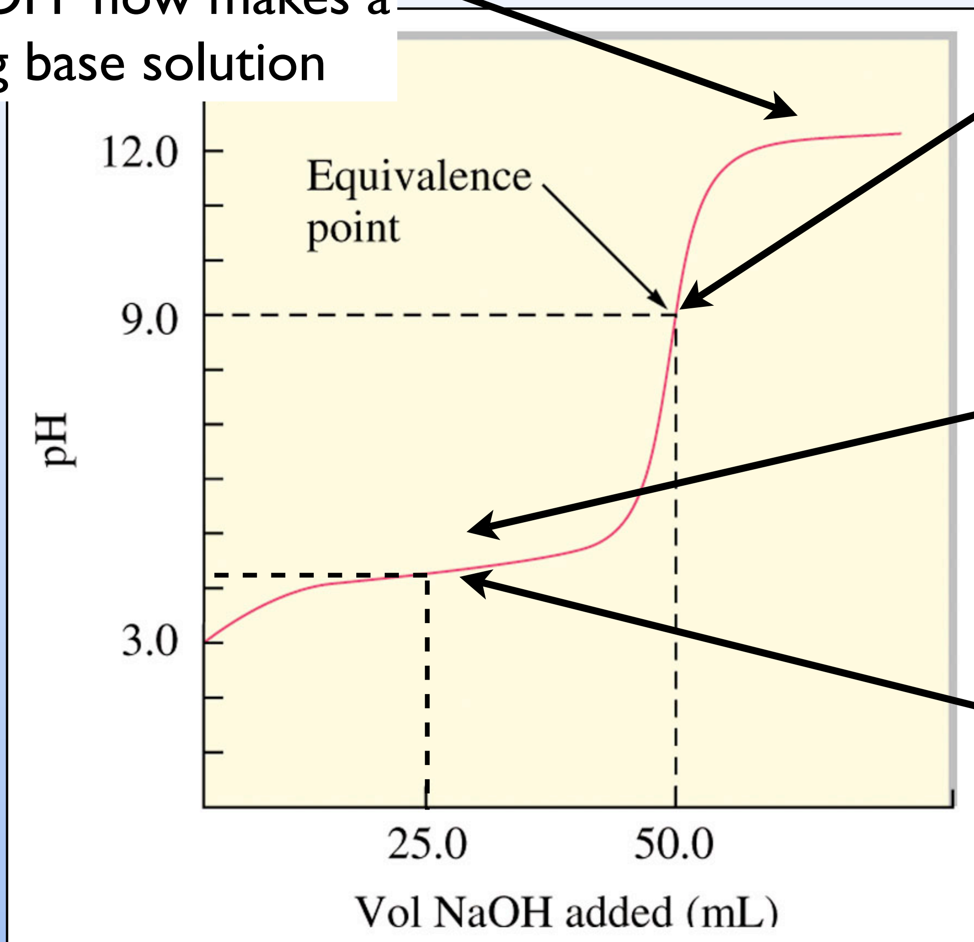
$$H^+(aq) \quad + \quad A^-(aq) \longleftrightarrow HA(aq)$$

$$OH^-(aq) \quad + \quad HA(aq) \longleftrightarrow A^-(aq)$$

Buffer will remain until
you react all of the initial acid or base

# Titrating a weak acid

all HA neutralized at equivalence point
Adding $OH^-$ now makes a strong base solution




Equivalence point
All of HA converted to $A^-$
Now a weak base solution

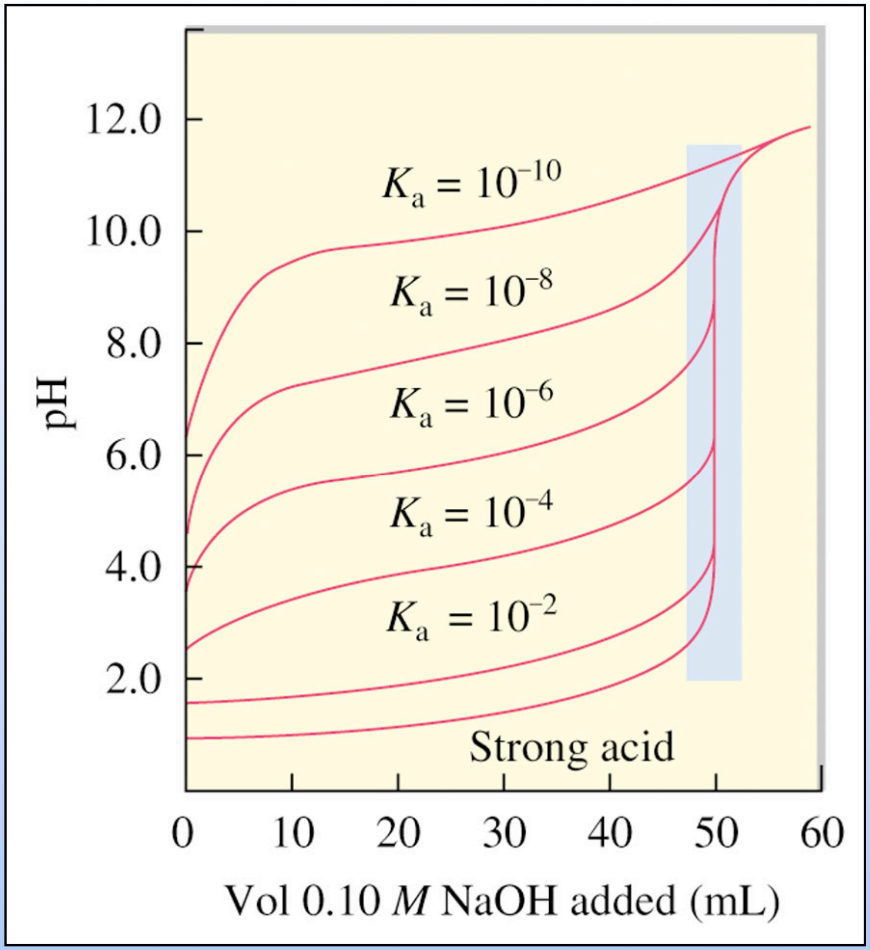
buffer region

half-way to equivalence point

half of HA converted to $A^-$

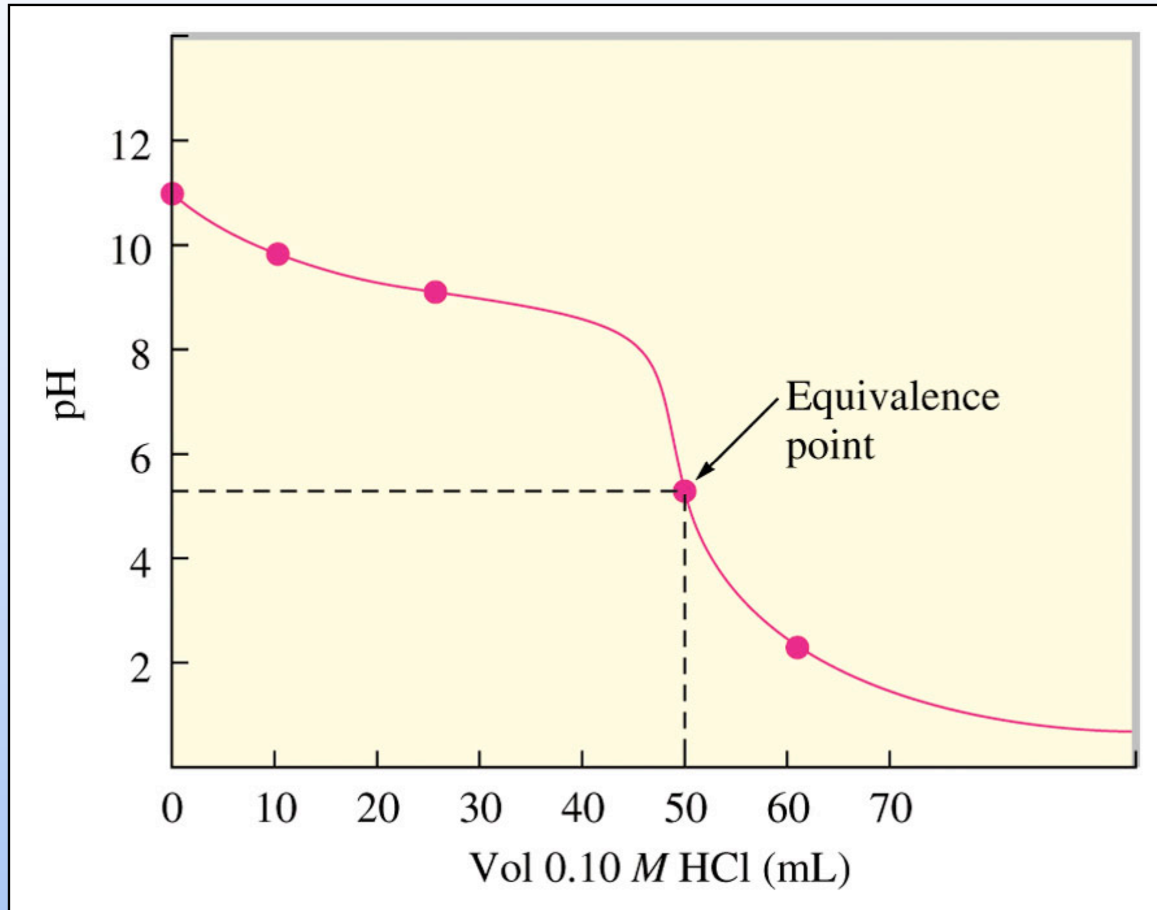
$[HA] = [A^-]$
$[H^+] = K_a$
$pH = pK_a$

# Calc on Doc Cam




All have the same concentration so they all have the same equivalence point (end point)

# Weak base titrated with strong acid



Rolaids® contain about 0.1 g of Magnesium Hydroxide
Why in the world would you ever put such a thing in your mouth?

A. 0.1 g is nothing. I each 10-20 g NaOH daily just for laughs

B. Acids are dangerous by bases as quite safe

C. The saliva in my mouth is acidic enough to "handle it"

D. $Mg(OH)_2$ is not soluble in water

# Solubility Equilibria

$$Mg(OH)_2\,(s) \longleftrightarrow Mg^{2+}(aq) + 2OH^-(aq)$$

$$K_{sp} = [Mg^{2+}][OH^-]^2$$

solubility product

# example