Slide 1

The End of Equilibrium!

(well, for us!)

1

Slide 2

K_{sp}

What is the solubility of $FeCO_3$?

Solubility = MAXIMUM amount of a compound that can dissolve in water.

This is actually an equilibrium.

2

Slide 3

Equilibrium problems involve 3 parts:

1. Balanced equation
2. "K-equation"
3. ICE chart

What is the balanced equation for dissolving something?

3

Slide 4

$FeCO_{3\,(s)} \leftrightarrow Fe^{2+}_{(aq)} + CO_3^{2-}{}_{(aq)}$

What is the "K-equation"?

$K = [Fe^{2+}][CO_3^{2-}]$

The "K" is the PRODUCT of the SOLUBLE ions. Hence, this reaction is called a "solubility product".

$K_{sp} = [Fe^{2+}][CO_3^{2-}]$

$K_{sp}(FeCO_3) = 3.07 \times 10^{-11}$

4

Slide 5

What is the solubility of $FeCO_3$?

$K_{sp}(FeCO_3) = 3.07 \times 10^{-11}$

	$FeCO_{3\,(s)}$ ↔	$Fe^{2+}_{(aq)}$ +	$CO_3^{2-}{}_{(aq)}$
I	S	0	0
C	-x	+x	+x
E	0.000000000001	x	x

$K_{sp} = 3.07 \times 10^{-11} = [Fe^{2+}][CO_3^{2-}] = x*x$

$x = SQRT(3.07 \times 10^{-11}) = 5.54 \times 10^{-6}$ M

5

Slide 6

Clicker question

What is the solubility of $Ba_3(PO_4)_2$ at 298K?

$K_{sp}(Ba_3(PO_4)_2) = 6 \times 10^{-39}$

A. 8×10^{-20} M
B. 2×10^{-8} M
C. 3×10^{-9} M
D. 9×10^{-9} M
E. 3×10^{-20} M

6

Slide 7

$Ba_3(PO_4)_2$ (s)↔ 3 Ba^{2+} (aq) + 2 PO_4^{3-}(aq)

I	S	0	0
C	-x	+3x	+2x
E	-	3x	2x

$K_{sp} = 6\text{x}10^{-39} = (3x)^3(2x)^2 = 27x^3 * 4x^2$

$5.56\text{x}10^{-41} = x^5$

$x = 8.89\text{x}10^{-9}$ M = $9\text{x}10^{-9}$ M

7

Slide 8

More common units for solubility…

…are g/L.

If you wanted g/L

$$9\times10^{-9}M = \frac{9\times10^{-9} mol Ba_3(PO_4)_2}{L}\frac{602g}{1mol} = 5\times10^{-6} g/L$$

8

Slide 9

Precipitation Reaction

The reverse reaction:

Solubility Product:

$Ba_3(PO_4)_2$ (s)↔ 3 Ba^{2+} (aq) + 2 PO_4^{3-}(aq)

Precipitation:

3 Ba^{2+} (aq) + 2 PO_4^{3-}(aq) ↔ $Ba_3(PO_4)_2$ (s)

It’s just K “upside down”

9

Slide 10

How do you know if something precipitates?

$Ba_3(PO_4)_2$ (s)↔ 3 Ba^{2+} (aq) + 2 PO_4^{3-}(aq)
K_{sp} = $6x10^{-39}$

What is the K_{sp}?
It's the limit on the amount of ions in solution.
$K_{sp} = [Ba^{2+}]^3[PO_4^{3-}]^2$

Remember our old friend "Q"?

10

Slide 11

What's Q?

Q is just the concentrations of products and reactants when you are NOT at equilibrium.

$K_{sp} = [Ba^{2+}]^3[PO_4^{3-}]^2 = 6x10^{-39}$
$Q = [Ba^{2+}]^3[PO_4^{3-}]^2$ = any other number

11

Slide 12

Q is less than K means…

1. You are NOT at equilibrium.
2. You could dissolve more solid: the products (dissolved ions) are too small.

12

Slide 13

Q is more than K means…

1. You are NOT at equilibrium.
2. You have TOO MANY products (dissolved ions). They can't stay dissolved, they need to precipitate out!

Slide 14

A little precipitation question:

500 mL of 0.100 M $Fe(NO_3)_3$ is mixed with 250 mL of 0.100 M KOH. What, if anything, precipitates from the solution? What mass of precipitate is formed?

Slide 15

What COULD form…?

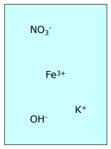
$KOH(s) \rightarrow K^+ (aq) + OH^- (aq)$

$Fe(NO_3)_3 (s) \rightarrow Fe^{3+} (aq) + 3 NO_3^- (aq)$

A beaker of KOH and $Fe(NO_3)_3$ has neither KOH nor $Fe(NO_3)_3$, it's all ions!

Slide 16

The 1st Rule of Chemistry…



Opposites attract!

Positive ions like negative ions.

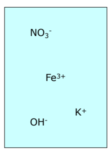
Negative ions like positive ions.

Postive ions hate positive ions.

Negative ions hate negative ions.

Slide 17

Only possible products are…



KOH or KNO_3

$Fe(OH)_3$ or $Fe(NO_3)_3$

We know that KOH and $Fe(NO_3)_3$ don't form…that's what we started with.

What about KNO_3 and $Fe(OH)_3$?

Slide 18

How many "Ks" are there in the beaker?

A. None of the below
B. 3
C. 2
D. 4
E. 5

Slide 19

What about KNO_3 and $Fe(OH)_3$?

They are both possible products of the reaction. Could they both form? Which one forms first? Do they form together? How would you know?

K_{sp}

Slide 20

When you have 2 possible reactions…

BIGGEST K wins!

Or, in this case, SMALLEST K_{sp}

$K_{precipitation} = 1/K_{sp}$

Small K_{sp} means big $K_{precipitation}$.

Slide 21

Precipitation is just the reverse of dissolution.

KNO_3 (s) ↔ K^+ (aq) + NO_3^- (aq)

K_{sp} (KNO_3) = HUGE (K+ salts are very soluble and nitrates are very soluble)

$Fe(OH)_3$ (s) ↔ Fe^{3+} (aq) + 3 OH^- (aq)

K_{sp} ($Fe(OH)_3$)=2.79×10^{-39}

Slide 22

So the only reaction to consider is…

$Fe(OH)_3$ (s) ↔ Fe^{3+} (aq) + 3 OH^-(aq)

K_{sp} ($Fe(OH)_3$)=2.79x10^{-39}

All equilibrium problems have 3 parts…yada yada yada…

22

Slide 23

K_{sp} ($Fe(OH)_3$)=2.79x10^{-39}

$Fe(OH)_3$ (s) ↔ Fe^{3+} (aq) + 3 OH^-(aq)

I

C

E

K_{sp} = 2.79x10^{-39} = $[Fe^{3+}][OH^-]^3$

What do we know?

23

Slide 24

Don't forget the dilution

500 mL of 0.100 M $Fe(NO_3)_3$ is mixed with 250 mL of 0.100 M KOH.

So…

Dilution is the solution!

0.100 M x 0.500 L = 0.05 mol/0.750 L = 0.0667 M

0.100 M x 0.250 L = 0.025 mol/0.750 L = 0.0333 M

24

Slide 25

K_{sp} $(Fe(OH)_3)$=2.79x10^{-39}

	$Fe(OH)_3$ (s) ↔	Fe^{3+} (aq) +	3 OH^-(aq)
I	0	0.067	0.033
C	+x	-x	-3x
E	x	0.067-x	0.033-3x

K_{sp} = 2.79x10^{-39} = [0.067-x][0.033-3x]^3

This is an algebraic mess BUT…K is really small.

Slide 26

K is really small…

…which means… $Fe(OH)_3$ is not very soluble.

So, x is going to be huge! We can use that to our advantage.

We can mathematically precipitate out ALL of the $Fe(OH)_3$ and then redissolve it!

Slide 27

What is product limiting?

	$Fe(OH)_3$ (s) ↔	Fe^{3+} (aq) +	3 OH^-(aq)
I	-	0.067	0.033
C	-	-x	-3x
E	-	0.067-x	0.033-3x

0.067-x = 0
X = 0.067

0.033 – 3x = 0
X=0.011

The hydroxide runs out first!

Slide 28

What is product limiting?

$Fe(OH)_3$ (s) ↔ Fe^{3+} (aq) + 3 OH^-(aq)

I	-	0.067	0.033
C	-	-0.011	-3(0.011)
E	-	0.056	0

0.067-x = 0
X = 0.067

0.033 – 3x = 0
X=0.011

The hydroxide runs out first!

28

Slide 29

Double your ICE, double your pleasure!

$Fe(OH)_3$ (s) ↔ Fe^{3+} (aq) + 3 OH^-(aq)

I	-	0.067	0.033
C	+0.011	-0.011	-3*(0.011)
I	0.011	0.056	0
C	-x	+x	+3x
E	0.011-x	0.056+x	3x

K_{sp} = 2.79x10^{-39} = [0.056+x][3x]^3

Look how much simpler that is. Even better, let's try and solve it the easy way!

29

Slide 30

Double your ICE, double your pleasure!

K_{sp} = 2.79x10^{-39} = [0.056+x][3x]^3

Assume x<<0.056!

2.79x10^{-39} = [0.056][3x]^3 = 0.056*27x^3

1.8452x10^{-39} = x^3
1.23x10^{-13} = x!

Pretty good assumption.

30

Slide 31

Double your ICE, double your pleasure!

	$Fe(OH)_3$ (s) ↔	Fe^{3+} (aq) +	3 OH^-(aq)
I	-	0.067	0.033
C	+0.011	-0.011	-3*(0.011)
I	0.011	0.056	0
C	-1.23×10^{-13}	$+1.23\times10^{-13}$	$+3(1.23\times10^{-13})$
E	0.011	0.056	3.68×10^{-13}

0.011 M $Fe(OH)_3$ precipitate

0.011 M $Fe(OH)_3$*0.75 L * 106.9 g/mol = 0.9 g $Fe(OH)_3$

Slide 32

Neat trick, huh?

Actually, that is another little trick in your ICE arsenal…

We know what to do when x is small. Now, if we suspect x is large, we can try this little trick.

In fact, you could always forcibly do a reaction to change the initial condition. After all, in the end the equilibrium will decide where it finishes.

Slide 33

Consider the following problem:

I have 1 Liter of a solution that is 2.3×10^{-2} M in $Fe(NO_3)_2$ and 1.5×10^{-2} M in $Mg(NO_3)_2$. I want to separate the two metals by adding K_2CO_3.

The 1st issue really is: can I do it?

If their solubilities are too close, they won't effectively separate. It all comes down to K_{sp}.

Slide 34

I've got a whole bunch of ions

I have 1 Liter of a solution that is 2.3×10^{-2} M in $Fe(NO_3)_2$ and 1.5×10^{-2} M in $Mg(NO_3)_2$. I want to separate the two metals by adding K_2CO_3.

$$Fe(NO_3)_2\ (s) \leftrightarrow Fe^{2+}\ (aq) + 2\ NO_3^-(aq)$$
$$Mg(NO_3)_2\ (s) \leftrightarrow Mg^{2+}\ (aq) + 2\ NO_3^-(aq)$$
$$K_2CO_3(s) \leftrightarrow 2\ K^+(aq) + CO_3^{2-}(aq)$$

I could look up the K_{sp} for these…but I almost never care about the things I start with…

34

Slide 35

I've got a whole bunch of ions

It's the new stuff I could make with the ions…

$FeCO_3$ $MgCO_3$ and maybe KNO_3

$$Fe(NO_3)_2\ (s) \leftrightarrow Fe^{2+}\ (aq) + 2\ NO_3^-(aq)$$
$$Mg(NO_3)_2\ (s) \leftrightarrow Mg^{2+}\ (aq) + 2\ NO_3^-(aq)$$
$$K_2CO_3(s) \leftrightarrow 2\ K^+(aq) + CO_3^{2-}(aq)$$

35

Slide 36

I've got a whole bunch of ions

$FeCO_3$ (s) $\leftrightarrow Fe^{2+}$ (aq) + CO_3^{2-} (aq)

$MgCO_3$ (s) $\leftrightarrow Mg^{2+}$ (aq) + CO_3^{2-} (aq)

$$KNO_3(s) \leftrightarrow K^+(aq) + NO_3^-(aq)$$

36

Slide 37

Given the following Ksp values, which salt is LEAST soluble in water?

$K_{sp}[FeCO_3] = 3.07\times10^{-11}$
$K_{sp}[MgCO_3] = 6.82\times10^{-6}$
$K_{sp}[KNO_3] = 1.26\times10^{2}$

37

Slide 38

Be a little careful about stoichiometry…

Imagine the following chloride salts:

$K_{sp}(XCl_2) = 12$
$K_{sp}(YCl) = 9$

Which salt is MORE SOLUBLE?

A. XCl_2
B. YCl
C. I need more information

38

Slide 39

$$K_{sp} = 12 = [X][Cl]^2 = (x)(2x)^2$$
$$12 = 4x^3$$
$$x = \sqrt[3]{\frac{12}{4}} = 1.44$$
$$K_{sp} = 9 = [X][Cl] = (x)(x)$$
$$9 = x^2$$
$$x = 2\sqrt{9} = 3$$

39

Slide 40

I should also be a little careful about concentration…

If the concentrations are very different, even the bigger K_{sp} may not matter.

It's really a question of Q! (Q_{sp})

40

Slide 41

The question is when $Q > K_{sp}$. But before it can exceed K_{sp}, Q must equal K_{sp}, so

$$Q = K_{sp} = 3.07 \times 10^{-11} = [Fe^{2+}][CO_3^{2-}]$$
$$= [2.3 \times 10^{-2}][CO_3^{2-}]$$
$$[CO_3^{2-}] = \frac{3.07 \times 10^{-11}}{2.3 \times 10^{-2}} = 1.33 \times 10^{-9}$$

As soon as my carbonate is 1 tiny bit above that, precipitation starts.

41

Slide 42

$$Q = K_{sp} = 6.82 \times 10^{-6} = [Mg^{2+}][CO_3^{2-}]$$
$$= [1.5 \times 10^{-2}][CO_3^{2-}]$$
$$[CO_3^{2-}] = \frac{6.82 \times 10^{-6}}{1.5 \times 10^{-2}} = 4.55 \times 10^{-4}$$

As soon as my carbonate is 1 tiny bit above that, precipitation starts.

But that is waaaay higher than the carbonate required for the iron to precipitate!

42

Slide 43

The effectiveness of the separation…

…depends on how much of the ion is left in solution when the second (or third or fourth) cation starts to precipitate.

Once K_{sp} is in play, it stays in play…

43

Slide 44

Puzzle 1. How much of the first cation is still in solution when the second cation begins to precipitate?

A. 4.55×10^{-4} M
B. 6.75×10^{-8} M
C. 1.5×10^{-2} M
D. 1.3×10^{-2} M
E. 2.3×10^{-2} M

44

Slide 45

$$Q = K_{sp} = 6.82 \times 10^{-6} = [Mg^{2+}][CO_3^{2-}]$$
$$= [1.5 \times 10^{-2}][CO_3^{2-}]$$
$$[CO_3^{2-}] = \frac{6.82 \times 10^{-6}}{1.5 \times 10^{-2}} = 4.55 \times 10^{-4}$$

$$K_{sp} = 3.07 \times 10^{-11} = [Fe^{2+}][CO_3^{2-}]$$
$$= [Fe^{2+}][4.55 \times 10^{-4} M]$$

$$[Fe^{2+}] = 6.75 \times 10^{-8}\ M$$

45

Slide 46

Puzzle 2. What is the total mass of K_2CO_3 that must be added to get the second cation to begin to precipitate?

A. 0.0629 g
B. 3.179 g
C. 3.24 g
D. 0.13 g
E. 1.12 g

46

Slide 47

The carbonate is actually in TWO places: solution AND THE $FeCO_3$!!

How much is in solution?

Well, the carbonate is 4.55×10^{-4}M in a 1 L solution:

$$1\,L\,\frac{4.55\times10^{-4}mol}{L}\,\frac{138.2\,g\,K_2CO_3}{mol} = 0.0629\,g\,K_2CO_3$$

47

Slide 48

We started with 2.3×10^{-2} M Fe^{2+}

When the magnesium starts to precipitate, the iron concentration is down to: $6.75\times10^{-8}\,M$

How much precipitated?

2.3×10^{-2} M originally – 6.75×10^{-8} M left = 2.29999325×10^{-2} M precipitate.

$$1\,L\,\frac{2.3\times10^{-2}mol\,Fe^{2+}}{L}\,\frac{1\,mol\,CO_3^{2-}}{1\,mol\,Fe^{2+}}\,\frac{1\,mol\,K_2CO_3}{1\,mol\,CO_3^{2-}}\,\frac{138.2\,g\,K_2CO_3}{mol} = 3.179\,g\,g\,K_2CO_3$$

Total mass = 3.1786 g +0.0629 g=3.24 g

48

Slide 49

Clicker Questions

What affect would adding acid have on the solubility of $Ca(OH)_2$?

A. Increase the solubility
B. Decrease the solubility
C. Have no effect on the solubility.

49

Slide 50

The solubility of $FeBr_2$ in pure water is 1.2 g/100 mL at 298 K. Would you expect the solubility of $FeBr_2$ to be ________ in 0.100 M NaBr at 298 K?

A. higher
B. lower
C. the same

50

Slide 51

LeChatelier's principle

If you stress an equilibrium, the equilibrium shift to respond to the stress.

$Ca(OH)_2$ (s) → Ca^{2+} (aq) + 2 OH^- (aq)

$K_{sp} = [Ca^{2+}][OH^-]^2$
"Lowering the pH" means increasing $[H_3O^+]$ which will neutralize some of the [OH-]. If you decrease the amount of hydroxide the reaction needs to make more to keep it at equilibrium.

51

Slide 52

REACTIONS ARE STUPID!

They don't know where the Ca^{2+} or the OH^- come from.

52

Slide 53

LeChatelier's Principle

Adding NaBr gives you a second source of Br-. So the reaction needs to make less to get to equilibrium.

$NaBr(s) \rightarrow Na^+ (aq) + Br^- (aq)$

$K_{sp}=[Na^+][Br^-]$

Still stupid!

53

Slide 54

Sample question

If you have 500.0 mL of a solution that is 0.022 M in Fe^{2+} and 0.014 M in Mg^{2+} and add 10.00 mL of 0.100 M K_2CO_3. What is left in solution after the precipitation?

$K_{sp}(FeCO_3) = 3.07 \times 10^{-11}$

$K_{sp}(MgCO_3) = 6.82 \times 10^{-6}$

54

Slide 55

If you have 500.0 mL of a solution that is 0.022 M in Fe^{2+} and 0.014 M in Mg^{2+} and add 10.00 mL of 0.100 M K_2CO_3. What is left in solution after the precipitation?

Dilution:

$$0.022\ M \times \frac{500\ mL}{510\ mL} = 0.0216\ M\ Fe^{2+}$$

$$0.014\ M \times \frac{500\ mL}{510\ mL} = 0.0137\ M\ Mg^{2+}$$

$$0.100\ M \times \frac{10\ mL}{510\ mL} = 0.00196\ M\ CO_3$$

1

Slide 56

$K_{sp}(FeCO_3) = 3.07\text{x}10^{-11}$
$K_{sp}(MgCO_3) = 6.82\text{x}10^{-6}$

I expect the $FeCO_3$ to precipitate first

56

Slide 57

$FeCO_3(s) \leftrightarrow Fe^{2+}(aq) + CO_3^{2-}(aq)$

		0.0216		0.00196
		-x		-x
		0.0216-x		0.00196-x

$3.07\text{x}10^{-11} = (0.0216\text{-x})(0.00196\text{-x})$
Assume x is LARGE!

3

Slide 58

$FeCO_3(s) \leftrightarrow Fe^{2+}(aq) + CO_3^{2-}(aq)$

0		**0.0216**		**0.00196**
+0.00196		-0.00196		-0.00196
0.00196		0.0196		0
-x		+x		+x
0.00196-x		0.0196+x		x

$3.07 \times 10^{-11} = (0.0196+x)(x)$

Assume $x << 0.0196$

$3.07 \times 10^{-11} = 0.0196x$

$X = 1.56 \times 10^{-9}$

Slide 59

$FeCO_3(s) \leftrightarrow Fe^{2+}(aq) + CO_3^{2-}(aq)$

0		**0.0216**		**0.00196**
+0.00196"M"		-0.00196		-0.00196
0.00196 "M"		0.0196		0
-1.56×10^{-9}		$+1.56 \times 10^{-9}$		$+1.56 \times 10^{-9}$
0.00196 "M"		0.0196		1.56×10^{-9}

Slide 60

Check $MgCO_3$ equilibrium

After the $FeCO_3$ precipitates, the CO_3^{2-} concentration is only 1.56×10^{-9}

$K_{sp}(MgCO_3) = 6.82 \times 10^{-6}$

$Q_{sp} = (1.56 \times 10^{-9})(0.0137\ M) = 2.137 \times 10^{-11}$

Q << K, so no $MgCO_3$ precipitates!

Slide 61

In terms of the solid…

$FeCO_3(s) \leftrightarrow Fe^{2+}(aq) + CO_3^{2-}(aq)$

0		0.0216		0.00196
+0.00196"M"		-0.00196		-0.00196
0.00196 "M"		0.0196		0
-1.56×10^{-9}		$+1.56\times10^{-9}$		$+1.56\times10^{-9}$
0.00196 "M"		0.0196		1.56×10^{-9}

Effectively "0.00196 M" $FeCO_3$ precipitated.

This is not a real concentration – it's not dissolved anymore!

1

Slide 62

In terms of the solid:

$$0.00196\ M\ FeCO_3 = \frac{0.00196\ mol\ FeCO_3}{L\ solution} 0.510\ L$$

$$= 9.996 \times 10^{-4} mol \frac{115.86\ g}{mol} = 0.1158\ g$$

So 0.1158 g of solid $FeCO_3$ have precipitated in the beaker!

2

Slide 63

Let's take it a little farther.

Suppose we add another 10.00 mL of 0.100 M K_2CO_3?

Well, we do the same thing all over again but we are starting with less Fe^{2+} and there's some initial CO_3^{2-} still floating around.

The reaction starts where the previous one left off!

3

Slide 64

Here's where we left off…

$FeCO_3(s) \leftrightarrow Fe^{2+}(aq) + CO_3^{2-}(aq)$

0		0.0216		0.00196
+0.00196"M"		-0.00196		-0.00196
0.00196 "M"		0.0196		0
-1.56×10^{-9}		$+1.56 \times 10^{-9}$		$+1.56 \times 10^{-9}$
0.00196 "M"		0.0196		1.56×10^{-9}

So… And we still have 0.1158 g of $FeCO_3$ in the bottom of the beaker!

4

Slide 65

We need to do the DILUTION!

We add another 10.00 mL of 0.100 M K_2CO_3

Dilution:

$$0.0196\, M \times \frac{510\ mL}{520\ mL} = 0.01922\, M\ Fe^{2+}$$

$$0.0137\, M \times \frac{510\ mL}{520\ mL} = 0.01344\, M\ Mg^{2+}$$

$$1.56 \times 10^{-9} M + 0.100\, M \times \frac{10\ mL}{520\ mL} 0.001923\, M\ CO_3$$

If you want to keep the $FeCO_3$ in the ICE chart:

$$0.00196\, M \times \frac{510\ mL}{520\ mL} = 0.001922\, M\ FeCO_3$$

5

Slide 66

Here's where we start…

$FeCO_3(s) \leftrightarrow Fe^{2+}(aq) + CO_3^{2-}(aq)$

0.001922		0.01922		0.001923
		-x		-x
		0.01922-x		0.001923-x

Same problem as before – x is BIG!

Same solution as before – precipitate it all and then let it redissolve!

6

Slide 67

CO_3^{2-} is thelimiting reactant

$FeCO_3(s) \leftrightarrow Fe^{2+}(aq) + CO_3^{2-}(aq)$

0.001922	0.01922	0.001923
+0.001923	-0.001923	-0.001923
0.003845	0.01730	0
-x	+x	+x
0.003845-x	0.01730+x	X

$3.07 \times 10^{-11} = (0.01730+x)(x)$

Assume x <<0.0173

$3.07 \times 10^{-11} = 0.01730x$

$x = 1.775 \times 10^{-9} M$

Slide 68

So, we have made more $FeCO_3$

$FeCO_3(s) \leftrightarrow Fe^{2+}(aq) + CO_3^{2-}(aq)$

0.001922	0.01922	0.001923
+0.001923	-0.001923	-0.001923
0.003845	0.01730	0
$-1.775 \times 10^{-9} M$	$+1.775 \times 10^{-9} M$	$+1.775 \times 10^{-9} M$
0.003845 "M"	0.01730	$1.775 \times 10^{-9} M$

Again, it's sort of a fake Molarity for the solid.

Slide 69

In terms of the solid:

$$0.003845\ M\ FeCO_3 = \frac{0.003845\ mol\ FeCO_3}{L\ solution} 0.520\ L$$

$$= 1.9994 \times 10^{-3} mol \frac{115.86\ g}{mol} = 0.2317\ g$$

So 0.2317 g of solid $FeCO_3$ have precipitated in the beaker!

Slide 70

Check $MgCO_3$ equilibrium

After the $FeCO_3$ precipitates, the CO_3^{2-} concentration is only $1.56\text{x}10^{-9}$

$K_{sp}(MgCO_3) = 6.82\text{x}10^{-6}$

$Q_{sp} = (1.775\text{x}10^{-9})(0.01344\ M) = 2.386\text{x}10^{-10}$

Q << K, so no $MgCO_3$ precipitates!

1

Slide 71

If I compare my two results

After 1st addition:

- **0.1158 g $FeCO_3$**
- **0.0196 M Fe^{2+}**
- **$1.56\text{x}10^{-9}$ M CO_3^{2-}**
- **0.0137 M Mg^{2+}**

After 2nd addition:

- **0.2317 g $FeCO_3$**
- **0.01730 M Fe^{2+}**
- **$1.775\text{x}10^{-9}$ M CO_3^{2-}**
- **0.01344 M Mg^{2+}**

Almost all the CO_3^{2-} precipitates each time.

I've doubled the amount of $FeCO_3$.

The Fe^{2+} is dropping and the CO_3^{2-} is gradually increasing.

The Mg^{2+} is being diluted but the actual amount dissolved isn't changing

You could keep doing this until virtually all the $FeCO_3$ is gone EXCEPT there is also Mg^{2+} in solution.

2

Slide 72

What happens to the Mg^{2+}?

The CO_3^{2-} is creeping upwards. Eventually, you'll get to the point at which it precipitates!

When…? Check K_{sp}!!!

(Note: There is a dilution issue, so we'll use the undiluted values to get a ballpark figure.)

$K_{sp}(MgCO_3) = 6.82\text{x}10^{-6} = [Mg^{2+}][CO_3^{2-}]$

3

Slide 73

What happens to the Mg^{2+}?

At all times, BOTH equilibria are in play IF the concentrations are high enough. At first only the $FeCO_3$ equilibrium is an issue because there's not enough CO_3^{2-} to exceed the $MgCO_3$ K_{sp}. Once it kicks in, however, BOTH equilibria must be satisfied all the time. So, when is this…?

$K_{sp}(MgCO_3) = 6.82x10^{-6} = [Mg^{2+}][CO_3^{2-}]$

$6.82x10^{-6} = [0.014\ M][CO_3^{2-}]$

$[CO_3^{2-}] = 4.87x10^{-4}$ M

So, until the CO_3^{2-} concentration has increased to about $5x10^{-4}$ M, the Mg^{2+} won't do anything!

Slide 74

If I compare my two results

After 1st addition:

- **0.1158 g $FeCO_3$**
- **0.0196 M Fe^{2+}**
- **$1.56x10^{-9}$ M CO_3^{2-}**
- **0.0137 M Mg^{2+}**

After 2nd addition:

- **0.2317 g $FeCO_3$**
- **0.01730 M Fe^{2+}**
- **$1.775x10^{-9}$ M CO_3^{2-}**
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Almost all the CO_3^{2-} precipitates each time.

I've doubled the amount of $FeCO_3$.

The Fe^{2+} is dropping and the CO_3^{2-} is gradually increasing.

The Mg^{2+} is being diluted but the actual amount dissolved isn't changing

You could keep doing this until virtually all the $FeCO_3$ is gone EXCEPT there is also Mg^{2+} in solution.

Slide 75

When…?

Depends on the $FeCO_3$ equilibrium!

When the Fe^{2+} concentration has dropped enough to allow the CO_3^{2-} concentration to get high enough!

$K_{sp}(FeCO_3) = 3.07x10^{-11} = [Fe^{2+}][4.87x10^{-4}]$

$[Fe^{2+}] = 6.304x10^{-8}$ M

So, we need to precipitate enough Fe^{2+} to drop the initiall 0.22 M Fe^{2+} down to $6.304x10^{-8}$ M (ignoring

Slide 76

In terms of amounts:

$$\frac{0.22\ mol\ Fe^{2+}}{L} 0.500L = 0.11\ mol \frac{55.85\ g}{mol} = 6.14\ g\ Fe^{2+}$$

$$\frac{6.30 \times 10^{-8}\ mol\ Fe^{2+}}{L} 0.500L = 3.15 \times 10^{-8}\ mol \frac{55.85\ g}{mol}$$
$$= 1.76 \times 10^{-6}\ g\ Fe^{2+}$$

$$\frac{1.76 \times 10^{-6} g\ Fe^{2+} left}{6.14\ g\ Fe^{2+} initial} \times 100 = 2.87 \times 10^{-5}\%\ left$$

$$100\% - 2.87 \times 10^{-5}\% left = 99.99997\%\ precipitated!$$

Pretty good!

7

Slide 77

Once the Mg^{2+} starts precipitating…

There is still a wee bit of Fe^{2+} left that will co-precipitate with the Mg^{2+}.

There ain't a ton, so you are precipitating mostly pure Mg^{2+} in this case. But that depends on how different the K_{sp} is. If they were closer together, you might have more left.

Only the first compound to precipitate will precipitate in pure form!

8