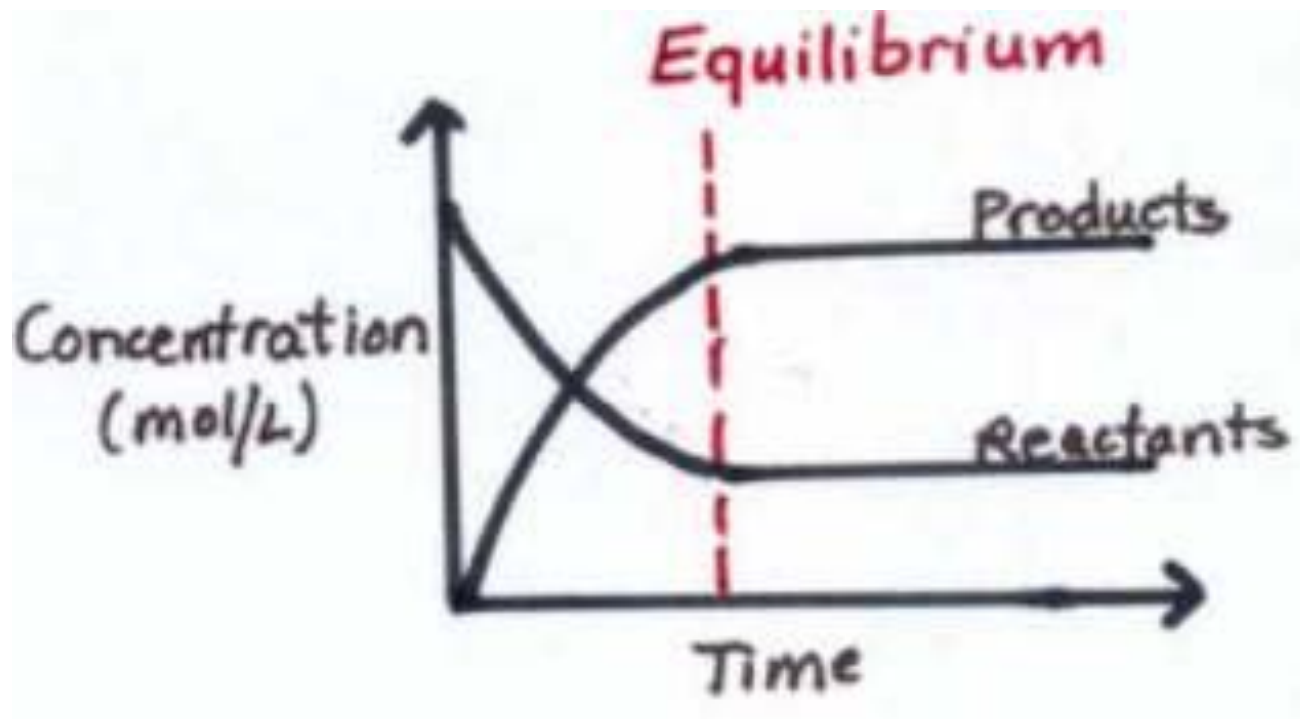


Reaction Quotient Q

Chemical Equilibrium K_{eq}

Solubility Product K_{sp}



★ general reaction $aA + bB \rightleftharpoons cC + dD$

★ reaction quotient $Q = \frac{[C]^c \times [D]^d}{[A]^a \times [B]^b}$

★ in equilibrium, $Q = K_{eq} = \text{constant for given } T$

★ solubility product = $K_{sp} = K_{eq}$ for solid dissolving

Example in class today: how much AgCl will dissolve in 1 L of H₂O? Or, when does this reach a saturated solution?

We are given that $K_{sp} = 1.7 \times 10^{-10}$

AgCl in H₂O dissolves to IONS:



Let $x = [\text{Ag}^+]$ then x also = $[\text{Cl}^-]$ since from the stoichiometric coefficients we know these concentrations are the same

$$K_{sp} = [\text{Ag}^+] \times [\text{Cl}^-] = x^2 = 1.7 \times 10^{-10} \longrightarrow x = 1.3 \times 10^{-5} \text{ M}$$

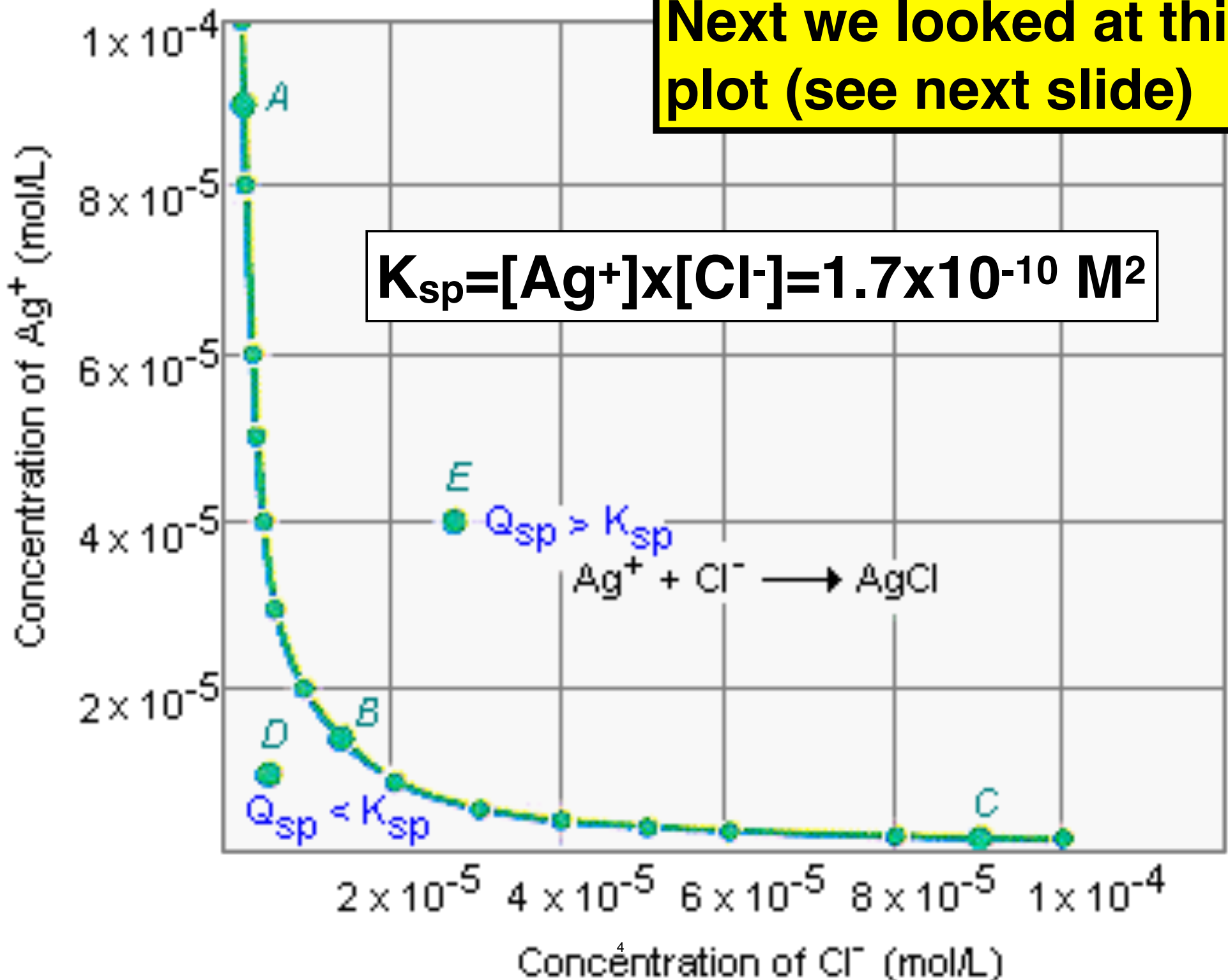
UNITS of K depend on reaction, in this case it must be M²

I.C.E. = Initial, Change, Equilibrium

	[AgCl]	[Ag ⁺]	[Cl ⁻]
I	all solid	0	0
C	-1.3x10 ⁻⁵	+1.3x10 ⁻⁵	+1.3x10 ⁻⁵
E	less solid	+1.3x10 ⁻⁵	+1.3x10 ⁻⁵

Next we looked at this plot (see next slide)

$$K_{sp} = [Ag^+] \times [Cl^-] = 1.7 \times 10^{-10} \text{ M}^2$$



The green curve on this plot shows where the solubility constant remains a constant, so if the concentrations are equal (as in the previous example) then we are at position B. But if the concentration of one species changes then for the system to remain in equilibrium (K_{sp} remains constant) then the concentration of the other species must change too.

If we are at point B, and we simply add more AgCl this is kind of boring: we know it precipitates out since we have already reached saturation.

But what is the effect of adding a DIFFERENT salt to this saturation? If this other salt has a common ion, it can have a huge effect.

For example, suppose we put 0.1 M of NaCl into the solution, and that this salt dissolves fully into Na^+ and Cl^- . This adds 0.1 M of Cl^- ions, and changes the equilibrium.

We can go back to the ICE table, but now we start with AgCl in equilibrium, we add 0.1M to the Cl⁻ column.

Let x=amount of Ag⁺ and Cl⁻ that react to form AgCl (in other words they precipitate). We know they must precipitate because K_{sp} is still the same constant, and we've increased [Cl⁻] from the addition of NaCl. If [Cl⁻] increases then [Ag⁺] decreases to keep K_{sp} constant (moving right on the plot).

	[AgCl]	[Ag ⁺]	[Cl ⁻]
I	all solid	1.3×10^{-5}	1.3×10^{-5}
C	+x precipitates	-x	+0.1-x
E	more solid	$1.3 \times 10^{-5} - x$	$1.3 \times 10^{-5} + 0.1 - x$

$$K_{sp} = (1.3 \times 10^{-5} - x) * (1.3 \times 10^{-5} + 0.1 - x)$$

but $1.3 \times 10^{-5} \ll 0.1$ and $x \ll 0.1$ so $1.3 \times 10^{-5} + 0.1 - x \sim 0.1$

$$\rightarrow K_{sp} = .1(1.3 \times 10^{-5} - x) = 1.7 \times 10^{-10} \quad x = 1.7 \times 10^{-9}$$

COMMON ION EFFECT : solubility repression by 2nd solute

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