AgCl: 107.9 + 35.45 = 143.35 g/mol

EXAMPLE: Calculate the grams per liter of silver(i) chloride (AgCl) in a solution that is at equilibrium with solid AgCl.

$$AgCl(s) \rightleftharpoons Ag^+(aq) + Cl^-(aq)\ ;\ K_c = 1.8 \times 10^{-10}$$

$$K_c = \boxed{[Ag^+][Cl^-] = 1.8 \times 10^{-10}}$$

We will define 'x' as the change in concentration of silver ion...

Species	[Initial]	Δ	[Equilibrium]
Ag^+	0	$+x$	x
Cl^-	0	$+x$	x

Each time we make a Ag+ ion, we also make a Cl- ion (1:1 ratio in the equation)

Substitute into the equilibrium expression ...

$$(x)(x) = 1.8 \times 10^{-10}$$

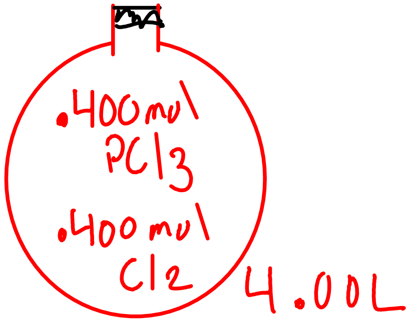
$$x^2 = 1.8 \times 10^{-10},\ \text{so } x = 1.3 \times 10^{-5} = [Ag^+] = [Cl^-]$$

Since every silver ion comes from dissolving 1 formula unit of AgCl, the dissolved AgCl concentration equals the silver ion concentration ...

$$\frac{1.3 \times 10^{-5}\ \text{mol AgCl}}{\text{L}} \times \frac{143.35\ \text{g AgCl}}{\text{mol AgCl}} = 0.0019\ \frac{\text{g AgCl}}{\text{L}}$$

$$PCl_3(g) + Cl_2(g) \rightleftharpoons PCl_5(g) \quad K_c = 49$$

If you add 0.400 moles of each reactant to a 4.00 L reaction vessel, what is the concentration of each species in the equilibrium mixture?



Initial conditions

$$\boxed{K_c = \frac{[PCl_5]}{[PCl_3][Cl_2]} = 49}$$

Let x equal the decrease in concentration of phosphorous trichloride

Species	[Initial]	Δ	[Equilibrium]
PCl_3	$\frac{0.400\ mol}{4.00\ L} = 0.100$	$-x$	$0.100 - x$
Cl_2	$\frac{0.400\ mol}{4.00\ L} = 0.100$	$-x$	$0.100 - x$
PCl_5	0	$+x$	x

$$\frac{[PCl_5]}{[PCl_3][Cl_2]} = \boxed{\frac{x}{(0.100 - x)(0.100 - x)} = 49}$$

To solve the problem, we must solve for 'x' ...

$$\frac{x}{(0.100-x)(0.100-x)} = 49$$

$$\frac{x}{(0.100-x)^2} = 49$$

$$x = 49(0.100-x)^2$$

$$(a-b)^2 = a^2 - 2ab + b^2$$

$$x = 49(0.0100 - 0.200x + x^2)$$

$$x = 0.49 - 9.8x + 49x^2$$

$$0 = 49x^2 - 10.8x + 0.49$$

$$a = 49 \quad b = -10.8 \quad c = 0.49$$

$$x = \frac{+10.8 \pm \sqrt{(-10.8)^2 - 4(49)(0.49)}}{2(49)} = \frac{10.8 \pm \sqrt{20.6}}{98}$$

$$x = \cancel{0.157} \text{ or } 0.0639$$

This value of 'x' gives us negative concentrations at equilibrium for both phosphorus trichloride and chlorine. That's physically impossible (you can't use more of a reactant than you supply), so the 0.0639 value must be the correct one!

The QUADRATIC EQUATION:

$$ax^2 + bx + c = 0$$

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

Each quadratic has two solutions (see the +/- part of the equation), but onlu one of them will be the correct chemical solution.

Species	[Initial]	Δ	[Equilibrium]
PCl_3	$\frac{0.400\ mol}{4.00\ L} = 0.100$	$-x$	$0.100 - x$
Cl_2	$\frac{0.400\ mol}{4.00\ L} = 0.100$	$-x$	$0.100 - x$
PCl_5	0	$+x$	x

$x = 0.0639$ Now, plug 'x' back into the (Equilibrium) column ...

$[PCl_3] = 0.100 - 0.0639 = \boxed{0.036\ M\ PCl_3}$

$[Cl_2] = 0.100 - 0.0639 = \boxed{0.036\ M\ Cl_2}$

$[PCl_5] = 0.0639 = \boxed{0.064\ M\ PCl_5}$

114 An 8.00 L reaction vessel at 3900C is charged with 0.850 mol of nitrogen and oxygen gases. Find the concentration of all species at equilibrium.

$$N_2(g) + O_2(g) \rightleftharpoons 2NO(g) \quad K_c = 0.0123$$

$$K_c = \frac{[NO]^2}{[N_2][O_2]} = 0.0123$$

As before, we'll need to express all these concentrations in terms of one variable...

Species	[Initial]	Δ	[Equil.]
N_2	$\frac{0.850\text{ mol}}{8.00\text{ L}} = 0.10625$	$-x$	$0.10625 - x$
O_2	$\frac{0.850\text{ mol}}{8.00\text{ L}} = 0.10625$	$-x$	$0.10625 - x$
NO	0	$+2x$	$2x$

We let 'x' equal the decrease in nitrogen gas concentration

$$\frac{[NO]^2}{[N_2][O_2]} = \boxed{\frac{(2x)^2}{(0.10625-x)(0.10625-x)} = 0.0123}$$

We need to solve for 'x' to complete this problem