Calculating the pH of weak bases:

Problem: Dimethylamine $(CH_3)_2NH$, a key intermediate in detergent manufacture, has a Kb of 5.9×10^{-4}. What is the pH of a 1.5 mol L^{-1} solution of dimethylamine?

Plan: We know; The intial concentration of the base (it is 1.5 mol L^{-1})
The Kb of the base (it is 5.9×10^{-4})

We want to calculate the pH of the solution.
The reaction between dimethylamine and water is:

$(CH_3)_2NH + H_2O = (CH_3)_2NH_2^+ + OH^-$

We can write the Kb expression as: $Kb = [(CH_3)_2NH_2^+][OH^-] / [(CH_3)_2NH]$

$[OH\text{-}]$ from base = $[(CH_3)_2NH_2^+] = [OH^-]$

Because Kb is small, we can assume that the amount of amine reacting is also small (it's a weak base), so;

$[(CH_3)_2NH]_{initial} - [(CH_3)_2NH]_{reacting} = [(CH_3)_2NH] \approx [(CH_3)_2NH]_{initial}$

Solution to this Problem:

- **Set up a change table:**

Concentration (mol L^{-1})	$(CH_3)_2NH$ (aq)	H_2O (l)	$(CH_3)_2NH_2^+$ (aq)	OH^- (aq)
Initial	1.5	-	0	0
Change	- x	-	+ x	+ x
Equilibrium	1.5 - x	-	x	x

Making the assumption;

Kb is small, so $[(CH_3)_2NH]_{initial} \approx [(CH_3)_2NH]$; thus, 1.5 mol L^{-1} – x ≈ 1.5 mol L^{-1}

Substituting this into the Kb expression and solving for x;

$Kb = [(CH_3)_2NH_2^+][OH^-] / [(CH_3)_2NH] = 5.9 \times 10^{-4} \approx x^2 / 1.5$ mol L^{-1}

$X = [OH\text{-}] \approx 3.0 \times 10^{-2}$

Calculating the pH:

$[H^+] = Kw / [OH^-] = 1.0 \times 10^{-14} / 3.0 \times 10^{-2} = 3.3 \times 10^{-13}$ mol L^{-1}

$pH = -Log_{10} [H^+]$
$= -Log_{10} [3.3 \times 10^{-13}$ mol $L^{-1}]$

pH = 12.48

Class Practice Problems

a.) Pyridine (C_5H_5N) plays a major role in organic syntheses as a solvent and also as a base. It has a pKb of 8.77. What is the pH of a 0.10 mol L^{-1} pyridine solution?

b.) Codeine ($C_{18}H_{21}NO_3$) is a narcotic pain reliever that forms a salt with HCl (codeine hydrochloride). What is the pH of 0.050 mol L^{-1} codeine hydrochloride? The pKb of codeine is 5.80.