9. HCN has an initial molarity of 0.50 M, with a K_a value of 3.7×10^{-8}. Calculate its pH at equilibrium. (Hint: This is an ICE problem.)

(1) $HCN + H_2O \rightleftharpoons H_3O^+ + CN^-$

	HCN		H_3O^+	CN^-
I	0.50 M		0	0
ΔC	− x		+ x	+ x
E	0.50		1.4×10^{-4}	1.4×10^{-4}

(1)

$K_a = \frac{[H_3O^+][CN^-]}{[HCN]}$ (1)

$3.7 \times 10^{-8} = \frac{[x][x]}{[0.50]}$ (1)

$[H^+] = 1.36 \times 10^{-4}$ mol/L (1)

pH = 3.87 (1)

10. Ethylamine ($C_2H_5NH_3$) is a weak Bonsted-Lowry base. If it has an initial molarity of 0.024 M and a K_b of 5.6×10^{-4}, calculate its pH at equilibrium. (Hint: This is an ICE Problem.)

$C_2H_5NH_3 + H_2O \rightleftharpoons C_2H_5NH_4^+ + OH^-$ (1)

	$C_2H_5NH_3$		$C_2H_5NH_4^+$	OH^-
I	0.024 M		0	0
ΔC	− x		+ x	+ x
E	0.024 M		3.67×10^{-3} M	3.67×10^{-3} M

(1)

$K_b = \frac{[C_2H_5NH_4^+][OH^-]}{[C_2H_5NH_3]}$ (1)

$5.6 \times 10^{-4} = \frac{[x][x]}{[0.024]} = 3.67 \times 10^{-3}\ M = [OH^-]$ (1)

(1)

pOH = − log (3.67×10^{-3} M)
= 2.44

pH = 14 − 2.44
= 11.56 (1)

11. A chemist adds 0.75 moles of NH_3 to enough water to make 0.50 liters of solution. K_b of ammonia is 1.8×10^{-5}. Determine the pH of this solution at equilibrium. (Hint: This is an ICE problem.)

$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$ (1)

	NH_3		NH_4^+	OH^-
I	1.5 mol/L		0	0
ΔC	− x		+ x	+ x
E	1.5 mol/L		5.2×10^{-3}	5.2×10^{-3}

(1)

$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$ (1)

$1.8 \times 10^{-5} = \frac{[x][x]}{[1.5\ mol/L]} = 5.2 \times 10^{-3}$ mol/L = $[OH^-]$

(1) (1)

pOH = − log(5.2×10^{-3})
= 2.28

pH = 14 − 2.28
= 11.72 (1)

12. Hydrazine, N_2H_4, has been used as a rocket fuel. Like ammonia, it is a Bronsted base. A 0.15 M solution has a pH of 10.70. What is the K_b for hydrazine?

$N_2H_4 + H_2O \rightleftharpoons N_2H_5^+ + OH^-$ (1)

	N_2H_4		$N_2H_5^+$	OH^-
I	0.15 mol/L		0	0
ΔC	− x		+ x	+ x
E	0.15 mol/L		5.0×10^{-4}	5.0×10^{-4} (1)

pH = 10.70
pOH = 3.30
$[OH^-] = 5.0 \times 10^{-4}$
(1)

$K_b = \frac{[5.0 \times 10^{-4}][5.0 \times 10^{-4}]}{[0.15]}$ (1)

$= 1.7 \times 10^{-6}$ (1)

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