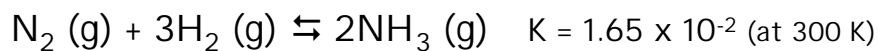


March 1, 2002

- ✓ Old Exam #2 questions posted
- ✓ Quiz today!

1

Example using K_c



➤ Calculate the equilibrium concentrations of the gases if we start with:

$$[\text{N}_2] = 0.500 \text{ mol/L} \quad n_{\text{N}_2} = 2.0 \text{ L (0.500 mol/L)} = 1.00 \text{ mol}$$

$$[\text{H}_2] = 1.500 \text{ mol/L} \quad n_{\text{H}_2} = 2.0 \text{ L (1.500 mol/L)} = 3.00 \text{ mol}$$

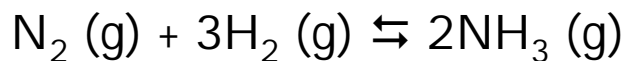
$$\text{Total Volume} = 2.0 \text{ L}$$

Converting K to K_c :

$$K_c = K_p / (RT)^{\Delta n} = \frac{1.65 \times 10^{-2}}{[(0.08206)(300)]^{-2}} = 10.0 = K_c$$

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ICE Time!



Initial mol 1.00 mol 3.00 mol 0 mol

Change -x -3x +2x

Equilib mol 1.00 - x 3.00 - 3x 2x

Equilib conc.	$\frac{1.00 - x}{2.0}$	$\frac{3.00 - 3x}{2.0}$	$\frac{2x}{2.0}$
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3

Solving for x

$$K_c = \frac{[\text{NH}_3]^2}{[\text{N}_2] [\text{H}_2]^3} = 10.0$$

Substituting and collecting terms:

$$\frac{x^2}{(1.00 - x)^4} = 16.875$$

Reduces to the following quadratic:

$$4.108 x^2 - 9.216x + 4.108 = 0$$

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Finally, the Answers

From the quadratic: $x = \cancel{1.63}$ and 0.6143

Substituting:

Too big!

$$[\text{NH}_3] = 2x/2.0 = 2(0.6143)/2.0 = \mathbf{0.61 \text{ mol/L}}$$

$$[\text{H}_2] = \frac{3.00 - 3(0.6143)}{2.0} = \mathbf{0.58 \text{ mol/L}}$$

$$[\text{N}_2] = \frac{1.00 - 0.6143}{2.0} = \mathbf{0.19 \text{ mol/L}}$$

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Acid/Base Chemistry

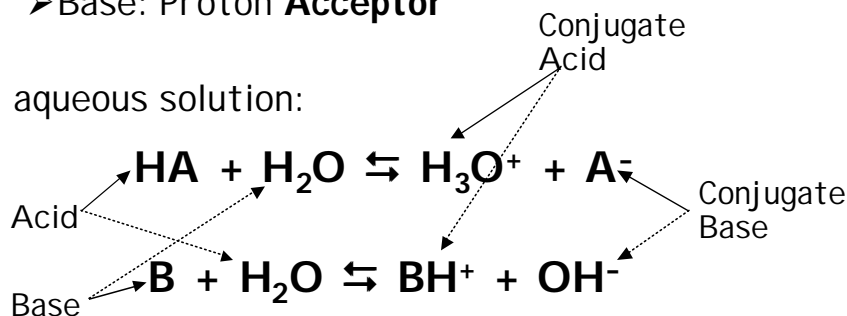
Chem 36
Spring 2002

Definitions

✓ Bronsted-Lowry Model

- Acid: Proton *Donor*
- Base: Proton **Acceptor**

In aqueous solution:



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Water

- Water can act *both* as an acid and as a base
 - *amphoteric*
- Water can react with *itself*
 - Auto-ionization or self-dissociation:



$$K = \frac{[H_3O^+][OH^-]}{1} = [H_3O^+][OH^-] = K_w$$

1.0 x 10⁻¹⁴ at 25 °C

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Water Equilibria

In pure water at 25 °C:

$$[\text{H}_3\text{O}^+] = [\text{OH}^-] = x$$

Substituting:

$$K_w = x^2 = 1.0 \times 10^{-14}$$

$$x = \underline{1.0 \times 10^{-7} \text{ M}} = [\text{H}_3\text{O}^+] = [\text{OH}^-]$$

Express small numbers as *logs*:

$$\text{pH} = -\text{Log}[\text{H}_3\text{O}^+]$$

2 sig figs

So, for pure water at 25 °C:

$$\text{pH} = -\text{Log}(1.0 \times 10^{-7}) = \underline{7.00}$$

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