March 1, 2002

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

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Example using K_c

N_2 (g) + $3H_2$ (g) ⇆ $2NH_3$ (g) K = 1.65×10^{-2} (at 300 K)

➢Calculate the ***equilibrium concentrations*** of the gases if we start with:

$[N_2]$ = 0.500 mol/L n_{N2} = 2.0 L (0.500 mol/L) = 1.00 mol

$[H_2]$ = 1.500 mol/L n_{H2} = 2.0 L (1.500 mol/L) = 3.00 mol

Total Volume = 2.0 L

Converting K to K_c:

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

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ICE Time!

$$N_2\,(g) + 3H_2\,(g) \leftrightarrows 2NH_3\,(g)$$

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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Solving for x

$$K_c = \frac{[NH_3]^2}{[N_2]\,[H_2]^3} = 10.0$$

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**

Too big!

Substituting:

$[NH_3] = 2x/2.0 = 2(0.6143)/2.0 =$ ***0.61 mol/L***

$[H_2] = \frac{3.00 - 3(0.6143)}{2.0} =$ ***0.58 mol/L***

$[N_2] = \frac{1.00 - 0.6143}{2.0} =$ ***0.19 mol/L***

Acid/Base Chemistry

Chem 36

Spring 2002

Definitions

✓ **Bronsted-Lowry Model**

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

In aqueous solution:

Conjugate Acid

Acid

$$\mathbf{HA + H_2O \leftrightharpoons H_3O^+ + A^-}$$

Conjugate Base

Base

$$\mathbf{B + H_2O \leftrightharpoons BH^+ + OH^-}$$

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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:

$$\mathbf{H_2O\ (l) + H_2O\ (l) \leftrightharpoons H_3O^+\ (aq) + OH^-\ (aq)}$$

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

1.0×10^{-14} at 25 °C

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Water Equilibria

In pure water at 25 ^{o}C:

$$[H_3O^+] = [OH^-] = x$$

Substituting:

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

$$x = \underline{\boldsymbol{1.0 \times 10^{-7}}\ M} = [H_3O^+] = [OH^-]$$

Express small numbers as *logs:*

$$\boxed{\mathbf{pH = -Log[H_3O^+]}}$$

2 sig figs

So, for pure water at 25 ^{o}C:

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

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