Chemistry 20 - Unit 2 - pH Calculations

Name: ____

The pH scale, the standard measurement of acidity was developed by the head of Carlsberg Laboratory's Chemical Department in 1909. Dr Søren Sørensen (1868-1939) developed the pH scale during his pioneering research into proteins, amino acids and enzymes - the basis of today's protein chemistry. Basically meaning 'the power of hydrogen', the scale provides a simple and universal measurement of the amount of hydrogen ions in a solution, which affects its acidity and how it reacts chemically.

The concentration of hydrogen ions is determined by measuring the current generated in an electrochemical cell when the ions migrate to oppositely charged electrodes. Sørensen used a negative logarithm of the hydrogen concentration to create a scale from 0-14, where a pH of less than 7 is an acid, 7 is neutral and higher than 7 is an alkali. So water has a pH of 7, lemon juice 2.4 and bleach 12.5. The pH of beer is 4.5. The applications of the pH scale have been countless, ranging from foodstuffs and cosmetics to chemicals and pharmaceuticals. Just about every liquid has had its pH measured at some time to determine how it will react and interact with living organisms.

pH stands for potency power of hydrogen and is the logarithmic scale (usually from 0 to 14) detailing how acidic something is. Acids have a pH between 0-7 whereas bases have a pH between 7-14.

There are multiple important formulas used for pH calculations. These formulas are not given to you in chemistry 20 for final exams, so make sure you understand and remember them all.

pH is calculated using the HYDRONIUM ion concentration - so your first step is to ionize the acidic substance - then use the formulas!

$$pH = -log[H_3O^+]$$

$$[H_3O^+] = 10^{-pH}$$

$[H_3O^+]$ = Concentration.... (molarity).... mol/L

Borrowed from Ms. Mogck of www.msmogcksclassroom.com

Example: What is the pH of HCl with a concentration of 8.93x10-3M?

Step 1: show the ionization equation for HCl in water.

$$HCl + H_2O \rightarrow H_3O^+ + Cl^- \quad \text{(Strong Acid)}$$

Step 2: use the ratio to determine the concentration of hydronium

$$[HCl] = [H_3O^+]$$

$$[H_3O^+] = 8.93 \times 10^{-3} M$$

Step 3: plug into the formula and calculate

$$pH = -\log[H_3O^+] = -\log(8.93 \times 10^{-3}) = 2.\underbrace{049}$$

Step 4: check sig figs

(remember sig figs of pH are only the decimal points**)**

decimal points are significate

$$3sf = 3\ \underline{decimals}$$

Example 2: What is the pH of a solution of hydrofluoric acid created by adding 0.500g of solid into 2.00L water?

Step 1: ionization equation

$$1\,HF + H_2O \rightarrow 1\,H_3O^+ + F^-$$

$$[HF] = [H_3O^+]$$

Step 2: calculate the concentration of hydronium

$$[H_3O^+] = \frac{n}{V} = \frac{0.02499}{2.00L} = 0.01249$$

$$n = 0.500g \times \frac{1mol}{20.01g}$$

Step 3: plug in and check sig figs

$$pH = -\log(0.01249) = 1.903$$

Example 3: What is the concentration of a monoprotic acid (meaning 1 proton or hydrogen) with a pH of 12.22?

$$[H_3O^+] = 10^{-pH}$$

$$= 10^{-12.22}$$

$$= 6.0 \times 10^{-13} \frac{mol}{L}$$

pH - PRACTICE

Name: ______________________________

1. Calculate the pH for the following acidic solutions:
   a. 0.00758M solution of HNO_3

$HNO_3 + H_2O \rightarrow H_3O^+ + NO_3^-$

$[HNO_3] = [H_3O^+] = 0.00758M$

$pH = -\log(0.00758M)$

$\boxed{pH = 2.120}$

   b. 0.00129 solution of HCl

$HCl + H_2O \rightarrow H_3O^+ + Cl^-$

$[HCl] = [H_3O^+] = 0.00129M$

$pH = -\log(0.00129M) = \boxed{2.889}$

   c. 0.100g of HBr in 5.00L solution

$[H_3O^+] = [HBr]$ $\quad n_{HBr} = 0.100g \times \frac{1mol}{80.91g} = 0.00124mol$

$[HBr] = \frac{0.00124mol}{5.00L} = 2.47\times10^{-4}M$

$pH = -\log(2.47\times10^{-7}M) = \boxed{3.607}$

   d. 0.0075g of HI in 3.50L of solution

$[H_3O^+] = [HI]$ $\quad n_{HI} = 0.0075g \times \frac{1mol}{127.91g} = 5.9\times10^{-5}mol$

$[HI] = \frac{5.9\times10^{-5}mol}{3.50L} = 1.7\times10^{-5}M$

$pH = -\log(1.7\times10^{-5}M)$

$\boxed{= 4.78}$

pH - PRACTICE

Name: ______________________________

1. Calculate the pH for the following acidic solutions:
   a. 0.00758M solution of HNO_3
   b. 0.00129 solution of HCl
   c. 0.100g of HBr in 5.00L solution
   d. 0.0075g of HI in 3.50L of solution

2. Calculate the $[H_3O^+]$ for the following solutions:
   a. 2.39g of H_2S in 4.00L of water

$H_2S + 2H_2O \rightarrow 2H_3O^+ + S^{2-}$

$n_{H_2S} = 2.39g \times \frac{1 mol}{34.09g} = 0.0701 mol$

$n_{H_3O} = 2 n_{H_2S} = 2 \times 0.0701 mol = 0.1402 mol$

   b. an acid with a pH of 6.90

$[acid] = 10^{-pH}$

$= 10^{-6.90}$

$\boxed{= 1.3 \times 10^{-7} M}$

   c. an acid with a pH of -1.285

$[acid] = 10^{-(-1.285)}$

$\boxed{= 19.3 M}$

   d. a base with a pH of 12.475

$[base] = 10^{-pH}$

$= 10^{-12.475}$

$\boxed{= 3.35 \times 10^{-13} M}$

   e. Predict what pOH would be a representative of?

potency power of hydroxide (base).

2. Calculate the $[H_3O^+]$ for the following solutions:
   a. 2.39g of H_2S in 4.00L of water

$[H_3O^+] = \frac{0.1402 mol}{4.00L}$

$\boxed{= 0.0351 M}$

   b. an acid with a pH of 6.90

   c. an acid with a pH of -1.285

   d. a base with a pH of 12.475

   e. Predict what pOH would be a representative of?