pH

- the measure of the acidity or alkalinity

- pH scale often used as a quick reference

  < 7 acid 7 neutral >7 base

  ← →

- Sören Sörenson defined pH mathematically

  $$pH = -\log_{10} [H^+]$$

- the "p" in pH comes from the German word potenz which means "to the power of" or exponents

- pH is therefore the power of the H^+ ion (H_3O^+) concentration measured in powers of 10

- for every change in pH of 1 unit, it means a factor of 10 change in the concentration of H^+ ions

- pH measures the concentration of H^+ ions (H_3O^+) can tell you which acid is more potent given any two acids.

- we can easily calculate the pH of strong acids because they dissociate completely. If we have the concentration of the solution we can use mole ratio to calculate the concentration of the H^+ ions.

Example: Calculate the pH of a 0.5M solution of HCl(aq).

HCl is a strong acid, and dissociates 100%

so

$HCl\ (aq) \rightarrow H^+\ (aq) + Cl^-\ (aq)$

0.5mol/L → 0.5mol/L (we could care less about the chloride ions)

$$\begin{aligned} pH &= -\log [H^+] \\ &= -\log (0.5) \\ &= 0.3 \end{aligned}$$

IMPORTANT NOTE : any numbers to the **left** of the decimal in pH calculation do not count in significant figures for logarithms.

We can also rearrange the pH equation to solve for concentration of ions:

Example: The pH of a solution is 6.05 (note this would only count as 2 sig fig) Determine the concentration of H^+ in solution.

$pH = -\log [H^+]$

so $[H^+] = 10^{-pH}$

$= 10^{-6.05}$ *

$= 8.9 \times 10^{-7}$ M

* Use your second log button

Try these:

1. Calculate the pH of a hydrochloric acid solution with a hydrogen ion concentration of 2.3×10^{-2} mol/L.
2. Calculate the $[H^+]$ in an acid solution whose pH is 6.50.
3. Calculate the pH in a 0.356 mol/L solution of hydroiodic acid

We can calculate pOH the same way as we calculate pH.

$$pOH = -\log [OH^-]$$

and $$[OH^-] = 10^{-pOH}$$

Try these:

1. Calculate the pOH of a solution with a hydroxide ion concentration of $[OH^-] = 0.99$ mol/L

2. Calculate the pOH of a magnesium hydroxide solution with a concentration of 1.25 mol/L.

Self-Ionization of Water

You might think pure water is 100% water molecules but it isn't.   One in about 500 000 000 or so water molecules ionizes into hydronium ions (H_3O^+) and hydroxide ions (OH^-).

$2H_2O\ (l) \rightleftharpoons H_3O^+\ (aq) + OH^-\ (aq)$

We can write the equilibrium constant, K for water!

$K_w = [H_3O^+]\ [OH^-]$ (Of course there is no $[H_2O]$ because it is a liquid)

This value has been measured to be :
$K_w = 1.0 \times 10^{-14}$

Water is neutral, and we can prove this with the K_w and pH calculations.

$K_w = [H_3O^+]\ [OH^-]$
$K_w = [H_3O^+]^2$ (we can say this because of 1:1 ratio)
$1.0 \times 10^{-14} = [H_3O^+]^2$
$[H_3O^+] = 1.0 \times 10^{-7} = [OH^-]$

If you calculate the pH
$pH = -\log [H_3O^+]$
$= -\log (1.0 \times 10^{-7})$
$= 7.00$

A pH of 7 is neutral - neither acidic nor basic.

Now that we have this new expression, K_w, we have a relationship established between the hydronium ion and the hydroxide ion.

We can use the K_w to solve for $[OH^-]$ given $[H_3O^+]$ or vise versa.

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

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

Examples:

Calculate the hydroxide ion concentration of a solution where the concentration of hydronium ion was measured to be 0.57M.

Since $K_w = [H_3O^+] [OH^-]$

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

so $[OH^-] = \frac{1.0 \times 10^{-14}}{0.57}$

$= 1.8 \times 10^{-14}$

Try these:

1. Calculate the hydronium ion concentration in a solution whose hydroxide ion concentration is 2.5 M

2. Calculate the hydroxide ion concentration in a 0.560 mol/L solution of hydrochloric acid.

We can also use the pH and pOH to solve for ion concentration.

This is because of the relationship between pH and pOH.

$$K_w = [H^+] [OH^-]$$
$$1.0 \times 10^{-14} = [H^+] [OH^-]$$
$$-\log (1.0 \times 10^{-14}) = -\log ([H^+] [OH^-])$$
$$-\log (1.0 \times 10^{-14}) = -\log [H^+] + -\log [OH^-]$$
$$14 = pH + pOH$$

This means that pH and pOH will always add up to 14, so if we know one, we subtract to find the other!

Example: What is the pOH of a solution whose pH is 4.25?

$$pH + pOH = 14$$
$$\therefore \quad pOH = 14 - 4.25$$
$$= 9.75$$

More Fun With pH and pOH

We can now use K_w to solve problems with pH, pOH, $[H_3O^+]$ and $[OH^-]$

Example #1 Calculate the pH of a solution if the $[OH^-]$ is 0.0065 mol/L.

What do you need to calculate pH? **$[H_3O^+]$**
How can we calculate $[H_3O^+]$ if we know $[OH^-]$?

We use K_w

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

$$= \frac{1.0 \times 10^{-14}}{0.0065}$$

$$= 1.54 \times 10^{-12} \text{ (keep extra sig dig)}$$

Now calculate pH

$$\begin{aligned} pH &= -\log [H_3O^+] \\ &= -\log (1.54 \times 10^{-12}) \\ &= 11.81 \end{aligned}$$

or we could calculate this another way:

Find the pOH, since we know the $[OH^-]$

$$\begin{aligned} pOH &= -\log [OH^-] \\ &= -\log (0.0065) \\ &= 2.187 \text{ (keep an extra sig dig)} \end{aligned}$$

then use $pH + pOH = 14$

$pH = 14 - 2.187$

$pH = 11.81$ (two sig dig)

Either method is correct - you have options.

Example #2 Calculate the pOH of a solution of 0.59 M nitric acid.

Either method you chose, you first need to work with some numbers and concentrations, so write the dissociation equation:

HNO_3 (aq) → H^+(aq) + NO_3^- (aq)
0.59 (x 1/1) 0.59

$\therefore [H^+] = 0.59$

$pH = -\log [H^+]$

$= -\log (0.59)$
$= 0.229$

$\therefore$ pOH $= 14 - 0.229$
$= 13.77$

or we could have done it this way, using K_w :

Use your dissociation equation to find the concentration of $[H^+] = 0.59$

$$[OH^-] = \frac{K_w}{[H^+]}$$
$$= \frac{1.0 \times 10^{-14}}{0.59}$$
$$= 1.695 \times 10^{-14}$$

then find pOH

pOH $= -\log [OH^-]$
$= -\log (1.695 \times 10^{-14})$
$= 13.77$

Calculation of Acid/Base Concentrations Given the pH / pOH

*****This can only be done for strong acids and strong bases because they are the only ones who completely dissociate** . (Weak acids and weak bases will be calculated later using K_a and K_b)

We must consider the mole ratio when given the formulas for strong acids and strong bases because it is possible for them to contain more than one hydronium or hydroxide ion.

Example: Calculate the concentration of a solution of aluminum hydroxide given it has a pOH of 2.49.

Find the hydroxide ion concentration using the pOH

$pOH = - \log [OH^-]$

$[OH^-] = 10^{-pOH}$

$[OH^-] = 10^{-2.49}$

$[OH^-] = 0.00324$

Use your mole ratio to find out the $[Al(OH)_3]$

$Al(OH)_3 \rightarrow Al^{3+} + 3OH^-$

$[Al(OH)_3] = 1/3\ [OH^-]$

$= 1/3 \times 0.00324$

$= 1.1 \times 10^{-3}$ M

We can extend these types of problems to include other topics covered in 2202, like calculating concentration using mass and volume (n= m/M and C=n/v), and using dilutions ($C_i v_i = C_f v_f$)

Example: Calculate the pH of a solution formed when 15 g of calcium hydroxide is dissolved in water to make a 350.0 mL solution.

This involves a lot of steps!

1. Balanced equation

$Ca(OH)_2 \rightarrow Ca^{2+} + 2\ OH^-$

2. In order to find the pH, you need $[H^+]$ or $[OH^-]$, of these the only one we can find is $[OH^-]$.

In order to find the $[OH^-]$ we need the concentration of the solution. This requires the use of C= n/v, but we don't know n, so this is our first calculation:

$n = \frac{m}{M} = \frac{15g}{74.10\ g/mol}$ M = 74.10 g/mol

$= 0.202$ mol

3. then, C = n/v (350.0mL = 0.3500L)

$= \frac{0.202\ mol}{0.3500L}$

$= 0.577$ mol/L

4. Now that we have a concentration of the solution we can use mole ratio to find the concentration of the hydroxide ions.

$Ca(OH)_2 \rightarrow Ca^{2+} + 2\ OH^-$

0.577 mol/L x 2/1 ratio→ 1.15 mol/L of OH^-

5. Now we can find the pOH and subtract from 14 to get the pH **<u>or</u>** use the K_w and $[OH^-]$ to get $[H^+]$

and then calculate the pH directly.

pOH = -log $[OH^-]$
= -log (1.15)
= $^-$0.0607

pH = 14 - $^-$0.0607
= 14.06

You try this one:

Find the pH of a solution formed when 0.250g of hydrogen chloride gas is added to 1.5 L of water.

(Hint: this is not as long as the example)

Remember with dilutions we took our C= n/v and built it into $C_iv_i = C_fv_f$; we can do this because $n_i = n_f$

We can apply pH, pOH, $[H^+]$, $[OH^-]$ to dilution problems as well

Example: 25.8 mL of a 0.10 M HCl solution was diluted to a volume of 326 mL. Calculate the pH of the dilute solution.

Remember before we can get pH we need $[H^+]$, and before we can get $[H^+]$, we need the final concentration of the HCl.

$C_i = 0.10$ M
$v_i = 25.8$ mL
$C_f = ?$
$v_f = 326$ mL

$$C_iv_i = C_fv_f$$

so $$C_f = \frac{C_iv_i}{v_f}$$

$$= \frac{(0.10M)\,(25.8\text{ mL})}{326\text{ mL}}$$
$$= 0.00791\text{ M}$$

Use your dissociation equation and mole ratio to determine there is a 1:1 ratio between the [HCl] and the $[H^+]$

$HCl \rightarrow H^+ + Cl^-$

We determined [HCl] to be 0.00791

so $[H^+] = 0.00791$ M

$\therefore$ pH $= -\log [H^+]$
$= -\log (0.00791)$
$= 2.10$

Example #2 12.6 mL of HNO_3 (aq) has a pH of 2.4 The acid was added to water (in other words diluted) to make a solution of unknown volume whose pH was 5.4 Calculate the final volume.

Step One : Look at the dissociation equation to determine the relationship between the nitric acid and the hydrogen ions. Nitric acid is a strong acid so -

$HNO_3\text{ (aq)} \rightarrow H^+\text{ (aq)} + NO_3^-\text{ (aq)}$

There is a 1:1 so whatever the $[H^+]$, there will be the same $[HNO_3]$

Step Two: Use the pH of the initial concentration to calculate the $[H^+]$ which will also give you the $[HNO_3]$ (because of the 1:1)

pH $= -\log [H^+]$
$[H^+] = 10^{-pH}$
$[H^+] = 10^{-2.4}$
$\therefore [HNO_3]_i = 10^{-2.4}$

$[HNO_3]_i = 4.0 \times 10^{-3}$ M

Step Three: Use the pH of the final concentration to find the $[HNO_3]_f$ same way you did in Step Two.

$$pH = -\log [H^+]$$
$$[H^+] = 10^{-pH}$$
$$[H^+] = 10^{-5.4}$$
$$\therefore [HNO_3]_f = 10^{-5.4}$$
$$[HNO_3]_f = 4.0 \times 10^{-6} \text{ M}$$

Step Four: Now we have C_i, C_f, we were given v_i , so we are solving for v_f.

$$C_i v_i = C_f v_f$$
$$v_f = \frac{C_i v_i}{C_f}$$
$$= \frac{(4.0 \times 10^{-3})\ (12.6 \text{ mL})}{4.0 \times 10^{-6}}$$
$$= 1.26 \times 10^{4} \text{ mL}$$
$$= 10 \text{ mL}$$