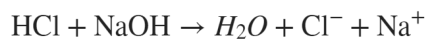


Chemistry

What volume of 0.75M Hydrochloric Acid is required to neutralize 32mL of 0.10M Sodium Hydroxide?



Both are strong acid/base. From Stoichiometry,

$$0.75 \frac{\text{mol}}{\text{L}} * X = 0.032\text{L} * 0.10 \frac{\text{mol}}{\text{L}}$$

$$u = \text{symunit};$$

$$X = 0.032 * 0.10 / 0.75 * u.1$$

$$X =$$

$$\frac{8}{1875} \text{ l}$$

If 50.0mL of 0.535M Sulfuric Acid is added to 100.0mL of 0.250M Sodium Hydroxide, what is the pH of the resulting solution?

From stoichiometry, $\text{H}_2\text{SO}_4 + 2\text{NaOH} \rightarrow 2\text{H}_2\text{O} + \text{SO}_4^{2-} + 2\text{Na}^+$, we see that the ratio is 2.

$$\text{pH} \equiv -\log([H^+]) = -\log\left(\frac{(0.050 * 0.535 * 2 - 0.1 * 0.250)\text{mol}}{0.150\text{L}}\right)$$

$$\text{mol} = 0.050 * 0.535 * 2 - 0.1 * 0.250$$

$$\text{mol} = 0.0285$$

$$L = 0.050 + 0.100$$

$$L = 0.1500$$

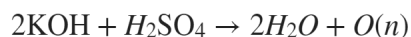
$$\text{mol/L}$$

$$\text{ans} = 0.1900$$

$$\text{pH} = -\log_{10}(\text{mol/L})$$

$$\text{pH} = 0.7212$$

How many milliliters of 0.750M Potassium Hydroxide are required to neutralize 215mL of 0.375M Sulfuric Acid?



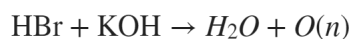
$$0.750 * x = 0.215 * 0.375 * 2$$

$$x = 0.215 * 0.375 * 2 / 0.750 \text{ u.l}$$

$$x =$$

$$\frac{43}{200} \text{ l}$$

A 60.0mL volume of 0.250M Hydrobromic Acid is titrated with 0.550M Potassium Hydroxide. Calculate the pOH after the addition of 17.0mL of Potassium Hydroxide?



$$\bullet 0.060 * 0.250 = x * 0.550 \therefore x = 0.027 \text{ L}$$

$$x = 0.060 * 0.250 / 0.550$$

$$x =$$

$$0.027272727272727$$

- After titration, 0.017L of KOH is added.
- $L = 0.060 + 0.027 + 0.017$ and there are $\text{OH} = 0.017 * 0.550 \text{ mol}$

$$\text{mol} = 0.017 * 0.550$$

$$\text{mol} =$$

$$0.0093500000000000$$

$$L = 0.060 + x + 0.017$$

$$L =$$

$$0.104272727272727$$

$$\text{pOH} = -\log_{10}(\text{mol}/L)$$

$$\text{pOH} =$$

$$1.047359121870525$$

What is the pH in the titration of 49.0mL of 0.100M Hydrochloric Acid after the addition of 58.0mL of 0.100M Sodium Hydroxide?

$$\text{mol} = 0.049 \text{ L} * 0.100 \frac{\text{mol}}{\text{L}} - 0.058 \text{ L} * 0.100 \frac{\text{mol}}{\text{L}}$$

$$L = 0.049 + 0.058$$

$$\text{mol} = 0.049 * 0.1 - 0.058 * 0.1$$

$$\text{mol} = -8.999999999999998\text{e-}04$$

$$L = 0.049 + 0.058$$

$$L = 0.1070000000000000$$

$$\text{pH} = -\log(\text{mol}/L)$$

$$\text{pH} = 4.778189350119733 - 3.141592653589793i$$

The problem here is that based on our conventional calculation, the concentration of H^+ is zero. But in the regime of very low H^+ concentration, one must use equilibrium constant to calculate **Naturally dissolving H^+ ions** Given as $\text{H}_2\text{O} \rightleftharpoons \text{H}^+ + \text{OH}^-$ $K = 10^{-14} = [\text{H}^+][\text{OH}^-]$

$$[\text{OH}] = -\text{mol}$$

$$\text{OH} = -\text{mol}$$

$$\text{OH} = 8.999999999999998\text{e-}04$$

$$L = L$$

$$L = 0.1070000000000000$$

$$\text{OH}_{\text{con}} = \text{OH} / L$$

$$\text{OH}_{\text{con}} = 0.008411214953271$$

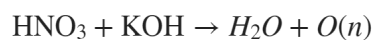
$$\text{H}_{\text{inferred}} = 10^{-14} / \text{OH}_{\text{con}}$$

$$\text{H}_{\text{inferred}} = 1.18888888888889\text{e-}12$$

$$\text{pH} = -\log_{10}(\text{H}_{\text{inferred}})$$

$$\text{pH} = 11.924858731754115$$

How many mL of 2.98M Nitric Acid would be necessary to neutralize 130.0mL of 2.35M Potassium Hydroxide?



$$x * 2.98 = 130 * 2.35$$

$$x = 130 * 2.35 / 2.98 \text{ *u.ml}$$

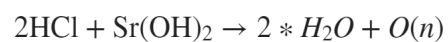
x =

$$\frac{15275}{149} \text{ ml}$$

vpa(x)

ans = 102.51677852348993288590604026846 ml

What is the concentration in terms of Molarity of a Hydrochloric Acid solution if 137.4mL of 2.21M Strontium Hydroxide is required to neutralize a 190.0mL sample?



$$2 * 137.4 * 2.21 = 190 * x$$

$$x = 2 * 137.4 * 2.21 / 190$$

x =

3.196357894736842