Chemistry

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

$\mathrm{HCl} + \mathrm{NaOH} \rightarrow H_2O + \mathrm{Cl}^- + \mathrm{Na}^+$

Both are strong acid/base. From Stoichiometry,

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

```
u = symunit;

X = 0.032*0.10/0.75*u.l
```

```
X =
```

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

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, $H_2\mathrm{SO}_4 + 2\mathrm{NaOH} \rightarrow 2H_2O + \mathrm{SO}_4^{2-} + 2\mathrm{Na}^+$, we see that the ratio is 2.

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

```
mol = 0.050*0.535*2 - 0.1*0.250
```

```
mol = 0.0285
```

```
L = 0.050 + 0.100
```

```
L = 0.1500
```

```
mol/L
```

```
ans = 0.1900
```

```
pH = -log10(mol/L)
```

```
pH = 0.7212
```

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

$2\text{KOH} + H_2\text{SO}_4 \rightarrow 2H_2O + O(n)$

$0.750 * x = 0.215 * 0.375 * 2$

```
x = 0.215*0.375*2 / 0.750 *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?

$\text{HBr} + \text{KOH} \rightarrow H_2O + O(n)$

- $0.060 * 0.250 = x * 0.550 \therefore x = 0.027L$

```
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 OH = 0.017 * 0.550 mol

```
mol = 0.017 * 0.550
```

```
mol =
   0.009350000000000
```

```
L = 0.060 + x + 0.017
```

```
L =
   0.104272727272727
```

```
pOH = -log10(mol/L)
```

```
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?

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

$$L = 0.049 + 0.058$$

```
mol = 0.049 * 0.1 - 0.058 * 0.1
```

```
mol =
    -8.999999999999998e-04
```

```
L = 0.049 + 0.058
```

```
L =
   0.107000000000000
```

```
pH = -log(mol/L)
```

```
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 $H_2O \Leftrightarrow H^+ + \mathrm{OH}^- \; K = 10^{-14} = [H^+]\{\mathrm{OH}^-\}$

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

```
OH = - mol
```

```
OH =
    8.999999999999998e-04
```

```
L = L
```

```
L =
   0.107000000000000
```

```
OH_con = OH / L
```

```
OH_con =
   0.008411214953271
```

```
H_inferred = 10^-14 / OH_con
```

```
H_inferred =
    1.188888888888889e-12
```

```
pH = -log10(H_inferred)
```

```
pH =
  11.924858731754115
```

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

$$\mathrm{HNO_3} + \mathrm{KOH} \rightarrow H_2O + O(n)$$
$$x * 2.98 = 130 * 2.35$$

```
x = 130*2.35/2.98 *u.ml
```

```
x =
```

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

```
vpa(x)
```

ans = $102.51677852348993288590604026846\ \mathrm{ml}$

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?

$2\mathrm{HCl} + \mathrm{Sr(OH)_2} \rightarrow 2 * H_2O + O(n)$

$2 * 137.4 * 2.21 = 190 * x$

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

```
x =
   3.196357894736842
```