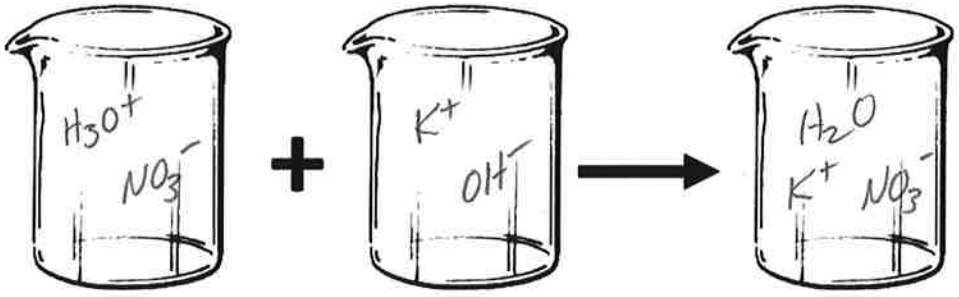
Chapter 4 Day 3 (Sections 4.10, 4.11, 4.4, 4.13) (Unit 3) 29 September 2017

1. In the first beaker sketch $HNO_3(aq)$ – remember, like HCl, it dissociates 100% into ions – one of which is H_3O^+. In the second beaker sketch KOH(aq). In the third beaker sketch what happens when the two solutions are mixed. Write the overall reaction and the net ionic reaction below the beakers.

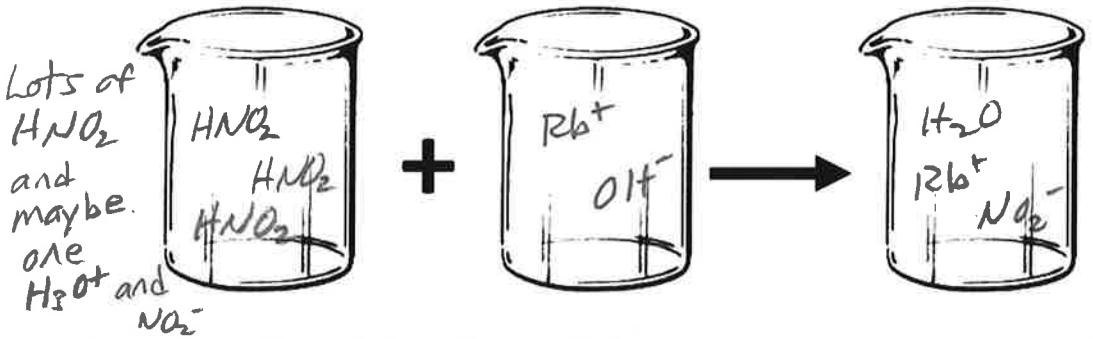


Balanced overall chemical equation: $HNO_3(aq) + KOH(aq) \rightarrow H_2O(l) + KNO_3(aq)$

Balanced ionic equation: $H_3O^+(aq) + NO_3^-(aq) + K^+(aq) + OH^-(aq) + 2\,H_2O(l) + K^+(aq) + NO_3^-(aq)$

Balanced net ionic equation: $H_3O^+(aq) + OH^-(aq) \rightarrow 2\,H_2O(l)$

2. In the first beaker sketch $HNO_2(aq)$ – a weak acid that does not dissociate 100% into ions. In the second beaker sketch RbOH(aq). In the third beaker sketch what happens when the two solutions are mixed. Write the overall reaction and the net ionic reaction below the beakers.



Balanced overall chemical equation: $HNO_2(aq) + RbOH(aq) \rightarrow H_2O(l) + RbNO_2(aq)$

Balanced ionic equation: $HNO_2(aq) + Rb^+(aq) + OH^-(aq) \rightarrow H_2O(l) + Rb^+(aq) + NO_2^-(aq)$

Balanced net ionic equation: $HNO_2(aq) + OH^-(aq) \rightarrow H_2O(l) + NO_2^-(aq)$

3. What volume of 0.0788 M LiOH(aq) is required to react with 0.00970 mol H_2SO_4 according to the reaction:

$H_2SO_4(aq) + 2\,LiOH(aq) \rightarrow Li_2SO_4(aq) + 2\,H_2O(aq)$

$n_{LiOH} = 0.00970\,mol\,H_2SO_4 \,/\, 2\,mol\,LiOH \,/\, 1\,mol\,H_2SO_4$

$= 0.0194\,mol\,LiOH$

$n = MV$

$V = \frac{n}{M} = \frac{0.0194\,mol \,/\, L}{0.0788\,mol}$

$= 0.246\,L$

4. A 25.00 mL sample of acetic acid was titrated with 0.1199 M NaOH(aq). If it took 30.44 mL of the NaOH solution to titrate the solution, what is the molarity of the original acetic acid solution?

$n_{NaOH} = M \cdot V = 0.1199\,\frac{mol}{L} * 0.03044\,L$

$= 0.00365\,mol\,NaOH$

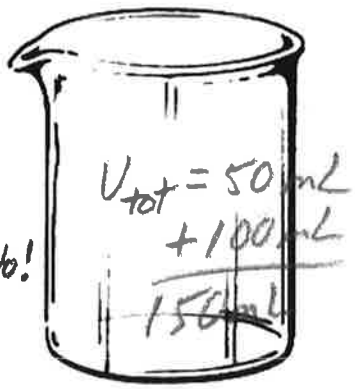
$M_{acetic\,acid} = \frac{0.00365\,mol\,HC_2H_3O_2}{0.02500\,L}$

$= 0.146\,mol/L$

5. What mass of nitric acid is in 22.2 mL of 0.2027 M HNO_3?

$n = MV$
$= 0.2027 \text{ mol} \times 0.0222 \text{ L}$
$= 0.00450 \text{ mol}$
$m = 0.00450 \text{ mol} \times 63 \text{ g/mol} = 0.283 \text{ g}$

6. Flash-back from our first day on Chapter 4: Suppose 50.0 mL of 0.200 M NaCl(aq) is mixed with 100.0 mL 0.300 M Na_2SO_4(aq). Is a precipitate formed? Sketch what things are present in the beaker. What is the molar concentration (molarity) of each ion?



No!

from NaCl: $n_{Na^+} = MV = 0.200 \times 0.050 = 0.010 \text{ mol } Na^+$
$n_{Cl^-} = MV = 0.010 \text{ mol } Cl^-$

from Na_2SO_4: $n_{Na^+} = MV = 0.300 \times 0.100 \times 2 = 0.060 \text{ mol } Na^+$
$n_{SO_4^{2-}} = MV = 0.300 \times 0.100 = 0.03 \text{ mol } SO_4^{2-}$
$M_{Na^+} = 0.07 \text{ mol} / 0.150 \text{ L} = 0.47 \text{ mol/L}$ etc

$[Cl^-] = 0.067 M$

$[SO_4^{2-}] = 0.20 M$

7. When an aqueous solution of sodium chromate is added to an aqueous solution of silver nitrate. Write the balanced equation, the ionic equation and the net ionic equation for this reaction.

$Na_2CrO_4(aq) + 2 AgNO_3(aq) \rightarrow NaNO_3(aq) + Ag_2CrO_4(s)$
(ionic not shown)
$2 Ag^+(aq) + CrO_4^{2-}(aq) \rightarrow Ag_2CrO_4(s)$

8. We saw aluminum metal placed in a solution of copper(II) chloride. The solution turned from blue to colorless and chunks of aluminum metal "dissapeared." Write and balance the oxidation-reduction reaction that occurred. Identify what was oxidized and what was reduced. Identify the oxidizing agent and the reducing agent.

net ionic: $2 Al(s) + 3 Cu^{2+}(aq) \rightarrow 2 Al^{3+}(aq) + 3 Cu(s)$
Al was oxidized Cu^{2+} was ox agent
Cu^{2+} was reduced Al was reducing

$2 Al(s) + 3 CuCl_2(aq) \rightarrow 2 AlCl_3(aq) + 3 Cu(s)$
($CuCl_2$ was ox agent; Al reducing agent)

9. What is the molarity of bromide in a solution if 25.00 mL of the solution was titrated with 17.40 mL of 0.00525 M permanganate with excess acid?

$2 MnO_4^-(aq) + 10 Br^-(aq) + 16 H^+(aq) \rightarrow$
$2 Mn^{2+}(aq) + 5 Br_2(aq) + 8 H_2O(aq)$

$n_{MnO_4^-} = MV = 0.00525 \times 0.0174$
$= 9.135 \times 10^{-5} \text{ mol}$
$n_{Br^-} = 9.135 \times 10^{-5} \text{ mol } MnO_4^- \times \frac{10 \text{ mol } Br^-}{2 \text{ mol } MnO_4^-}$
$= 4.568 \times 10^{-4} \text{ mol}$
$[Br^-] = 4.568 \times 10^{-4} \text{ mol} / 0.025 \text{ L} = 0.01827 M$

Questions in final exam format (multiple choice):

10. What is the oxidation number change for the iron atom in the following reaction?

$2 Fe_2O_3(s) + 3 C(s) \rightarrow 4 Fe(s) + 3 CO_2(g)$

A. -6 (B) -3
C. +3 D. +6

11. Which species functions as the oxidizing agent in the following reduction-oxidation reaction?

$5 Fe^{2+}(aq) + MnO_4^-(aq) + 8 H^+(aq) \rightarrow$
$Mn^{2+}(aq) + 5 Fe^{3+}(aq) + 4 H_2O(l)$

A. $Fe^{2+}(aq)$ B. $H^+(aq)$
C. $Mn^{2+}(aq)$ (D) $MnO_4^-(aq)$

12. Balance the chemical equation given below, and determine the number of milliliters of 0.0600 M phosphoric acid required to neutralize 25.00 mL of 0.0150 M calcium hydroxide.

$3 Ca(OH)_2(aq) + 2 H_3PO_4(aq) \rightarrow$
$1 Ca_3(PO_4)_2(s) + 6 H_2O(l)$

A. 1.69 mL
(B) 4.17 mL
C. 6.25 mL
D. 12.50 mL

Now try these problems from the book:
Section 4.10. (Redox) Problems 12, 104, 106, 108 and 133
Section 4.11. (Redox) Problems 13, 110 and 111
Section 4.4. (Types of Chemical Reactions – much of this was done in lab) Problems 68, 69
Section 4.13. (Solution stoichiometry) Problems 15, 44, 94, 116, 118, 120 and 124