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Acid-Base Titrations & Redox Reactions (Chapter 4 Pt 2)

I. **Big Idea**

Previously, we've talked about precipitation reactions and how they're useful to us in lab. Today we're going to discuss two more types of reactions: **acid-base reactions** & **redox reactions.** Before we get there, let's define some terms.

Arrhenius definition of "Acid" -

Makes H^+ in solution

Arrhenius definition of "Base" -

Makes OH^- in solution

The general reaction between an acid and a base is:

Acid + base → water + salt   i.e   HCl (aq) + NaOH (aq) → H_2O (l) + NaCl (aq)

II. **Acid-Base Titrations**

What is a titration? - the slow addition of one solution of a known concentration (called a titrant) to a known volume of another solution of unknown concentration until the reaction reaches neutralization.

1. Determine the concentration of NaOH if 10.0 mL of the solution took 11.5 mL of 2.0 M HCl to reach the equivalence point in a titration.

Balanced eq: HCl (aq) + NaOH (aq) → H_2O (l) + NaCl (aq)

$$(11.5 \times 10^{-3})\ L\ of\ HCl \times \frac{2.0\ mol\ HCl}{1L\ of\ HCl} \times \frac{1\ mol\ NaOH}{1\ mol\ HCl} = 0.023\ mol\ NaOH$$

$$\frac{0.023\ mol\ NaOH}{10.0 \times 10^{-3} L} = 2.3\ M\ NaOH$$

2. How many milliliters of 0.610 M NaOH solution are needed to neutralize 20.0 mL of a 0.245 M H_2SO_4 solution?

Balanced eq: H_2SO_4 + 2NaOH = $2H_2O$ + Na_2SO_4

$$(20.0 \times 10^{-3})\ L\ of\ H_2SO_4 \times \frac{0.245\ mol\ H_2SO_4}{1L\ of\ H_2SO_4} \times \frac{2\ mol\ NaOH}{1\ mol\ H_2SO_4} \times \frac{1\ L\ NaOH}{0.610\ mol\ NaOH} \times \frac{1000\ mL}{1\ L} = 16.1\ mL\ NaOH$$

3. You are titrating a solution of sodium hydroxide of unknown concentration with a solution of phosphoric acid that has a concentration of 0.300 M. Starting with 29.0 mL of the sodium hydroxide solution, you use 16.8 mL of the acid to titrate the base to completion. Calculate the concentration of the sodium hydroxide solution.

Balanced eq: 3NaOH+H_3PO_4 →Na_3PO_4+$3H_2O$

$$(16.8 \times 10^{-3})\ L\ of\ H_3PO_4 \times \frac{0.300\ mol\ H_3PO_4}{1L\ of\ H_3PO_4} \times \frac{3\ mol\ NaOH}{1\ mol\ H_3PO_4} = 0.01512\ mol\ NaOH$$

$$\frac{0.01512\ mol\ NaOH}{29.0 \times 10^{-3} L} = 0.52\ M\ NaOH$$

4. A 0.350-g sample of an acid, HX, requires 25.4 mL of a 0.140 *M* NaOH solution for complete reaction. Calculate the molar mass of the acid.
   (A) 42.3 g/mol (B) 68.4 g/mol (C) 98.4 g/mol (D) 121.3 g/mol
   (E) none of these

Balanced eq: HX (aq) + NaOH (aq) → H_2O (l) + NaX (aq)

$$(25.4 \times 10^{-3})\ L\ of\ NaOH \times \frac{0.140\ mol\ NaOH}{1L\ of\ H_2SO_4} \times \frac{1\ mol\ HX}{1\ mol\ NaOH} = 0.003556\ mol\ HX$$

$$\frac{0.350\ g\ HX}{0.003556\ mol\ HX} = 98.4\ g/mol\ HX$$

III. Redox Reactions

Reduction – gains electrons

Oxidation – loses electrons

Oxidation Agent – the species that is itself reduced

Reducing Agent – the species that is itself oxidized

5. Which of the following reactions does not involve oxidation-reduction?
   (A) $MnO_2 + 4HI \rightarrow I_2 + 2H_2O + MnI_2$
   (B) $LiOH + HCl \rightarrow H_2O + LiCl$
   (C) $2Na + 2H_2O \rightarrow 2NaOH + H_2$
   (D) $Mg + 2HI \rightarrow MgI_2 + H_2$
   (E) All the above are oxidation-reduction reactions

6. In the following reaction, which species is the reducing agent?
   $3Cu + 6H+ + 2HNO_3 \rightarrow 3Cu^{2+} + 2NO + 4H_2O$
   a. Cu^{2+} b. HNO_3 c. Cu d. NO e. H^+
   The reducing agent is the species that gets oxidized. Cu loses electrons to become Cu^{2+}.

7. Identify the oxidation numbers of all of the elements in the following equation as well as indicate which element was oxidized and which was reduced.

   $2Al(s) + 3CuSO_4(aq) \rightarrow Al_2(SO_4)_3(aq) + 3Cu(s)$

O:$Al^{(0)} \rightarrow Al^{(+3)}{}_2(S^{(+6)}O^{(-2)}{}_4)_3$
R:$Cu^{(+2)}(S^{(+6)}O^{(-2)})_4 \rightarrow Cu^{(0)}$
Al was oxidized, Cu was reduced.

IV. Balancing Redox Reactions in Acidic or Basic Solution

8. Balance the following equation in acidic solution. When the following reaction is balanced in acidic solution, what is the coefficient of water?
   $Zn(s) + NO_3^-(aq) \rightarrow Zn^{2+}(aq) + NH_4^+(aq)$

   Oxi: $4 \times (Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-)$

   Red: $NO_3^-(aq) + 10H^+ + 8e^- \rightarrow NH_4^+(aq) + 3H_2O$

   Add them together to get:

$4Zn(s) + NO_3^-(aq) + 10H^+ \rightarrow 4Zn^{2+}(aq) + NH_4^+(aq) + 3H_2O$

Coefficient of water = 3 on the right.

First, we write the skeleton equation then separate it into half-reactions. We start with balancing all the atoms except hydrogen and oxygen (zinc and nitrogen look good here). Then we add water to balance the oxygens and hydrogen ions to balance the hydrogens. We then balance the charge on each half reaction by adding electrons to either side. We then scale the reaction so that the oxidation half reaction loses the same number of electrons that the reduction reaction will gain. Add the reactions together and simplify to get above.

9. Balance the following reaction occurs in aqueous acid solution. What is the coefficient of NO_3^-?
$NO_3^- + I^- \rightarrow IO_3^- + NO_2$

Oxi: $I^- + 3H_2O \rightarrow IO_3^- + 6H^+ + 6e^-$
Red: $6 \times (NO_3^- + 2H^+ + e^- \rightarrow NO_2 + H_2O)$
Add them together to get:

$I^- + 3H_2O + 6NO_3^- + 12H^+ \rightarrow IO_3^- + 6H^+ + 6NO_2 + 6H_2O$

$I^- + 6NO_3^- + 6H^+ \rightarrow IO_3^- + 6NO_2 + 3H_2O$

First, we write the skeleton equation then separate it into half-reactions. We start with balancing all the atoms except hydrogen and oxygen (nitrogen and iodine looks good here). Then we add water to balance the oxygens and hydrogen ions to balance the hydrogens. We then balance the charge on each half reaction by adding electrons to either side. We then scale the reaction so that the oxidation half reaction loses the same number of electrons that the reduction reaction will gain. Add the reactions together and simplify to get above.

10. Balance the following equation in basic solution and determine the coefficient of OH- and its location (right side or left side) in the equation. Also identify the oxidizing and reducing agents in the reaction.

$$Ce(s) + PO_4^{3-}(aq) \rightarrow HPO_3^{2-}(aq) + Ce(OH)_3(s)$$

Oxi: $2 \times (Ce + 3H_2O + 3OH^- \rightarrow Ce(OH)_3 + 3H_2O + 3e^-)$
Red $3 \times (PO_4^{3-} + 3H_2O + 2e^- \rightarrow HPO_3^{2-} + H_2O + 3OH^-)$
Add them together to get:

$2Ce + 3PO_4^{3-} + 15H_2O + 6OH^- + 6e^- \rightarrow 2Ce(OH)_3 + 3HPO_3^{2-} + 9H_2O + 6e^- + 9OH^-$
$2Ce + 3PO_4^{3-} + 6H_2O \rightarrow 2Ce(OH)_3 + 3HPO_3^{2-} + 3OH^-$

First, we write the skeleton equation then separate it into half-reactions. We start with balancing all the atoms except hydrogen and oxygen (cesium and phosphorus looks good here). Then we add water to balance the oxygens and hydroxides to balance the hydrogens. (Alternatively, you could add hydrogen ions and balance it in acidic medium, then switch it later). We then balance the charge on each half reaction by adding electrons to either side. We then scale the reaction so that the oxidation half reaction loses the same number of electrons that the reduction reaction will gain. Add the reactions together and simplify to get above.