Balancing chemical equations

Teacher instructions and answers

Chemical processes in water treatment

Instructions

1. Print student worksheets and equation cards (single-sided makes it easier to arrange).
2. Hand worksheets and the three sets of equation cards to each group of students.
3. Ask the students to work as a group to balance equations. Students will need to draw on their understanding of chemical principles to answer the reasons why they were added.

Answers

1. **Write the relevant chemical equations of the dissolution of calcium hydroxide (lime). Then, identify the reasons why calcium hydroxide was added to the water after screening.**

$$Ca(OH)_{2\,(aq)} \leftrightarrow Ca^{2+}{}_{(aq)} + 2OH^{-}{}_{(aq)}$$

OH^- increases pH to reach the targeted pH range ($6.3 < pH < 10.3$) for a water filtration plant. At Orchard Hills this is $pH \sim 9$ to:

- boost pH to buffer the acidic reactions that occur from subsequent additions of chemicals eg, coagulants like ferric chloride.
- create an optimal environment for stable floc formation. Floc occurs better in a mildly alkaline environment. Ferric hydroxide floc is insoluble in the pH range.
- add calcium ions to harden the water, reducing calcium being leached from pipes, increasing the life of the pipes.

2. **In all surface waters, carbon dioxide dissolves into water, forming the carbon dioxide/carbonic acid buffer system.**

   Write the relevant chemical equations to explain how the water is naturally buffered. How does calcium hydroxide from above manipulate this buffer? When would an unwanted by-product be produced?

$$CO_{2(g)} \leftrightarrow CO_{2(aq)}$$
$$CO_{2(aq)} + H_2O_{(l)} \leftrightarrow H_2CO_{3(aq)}$$
$$H_2CO_{3(aq)} \leftrightarrow H^+{}_{(aq)} + HCO_{3^-}{}_{(aq)}$$
$$HCO_{3^-}{}_{(aq)} \leftrightarrow H^+ + CO_3{}^{2-}{}_{(aq)}$$

This is an equilibrium buffer system; different species will dominate at different pH ranges.

When pH is:

- < 6.3, H_2CO_3 dominates

- 6.3 < pH < 10.3, HCO3$^-$ dominates
- > 10.3, CO_3^{2-} dominates.

When HCO_3^- dominates, there's enough buffering capacity for the acidic coagulation reaction compared to lower pH range. HCO_3^- is a proton donor and proton acceptor. This means that when chemicals are added to the system that may produce acidic (or alkaline) conditions, it can resist changes in pH.

When $Ca(OH)_2$ is added to the water at 6.3<pH<10.3, the following will occur:

$$Ca^{2+} + HCO_{3^-} \rightarrow Ca(HCO_3)_{2\,(aq)}$$

This ensures the calcium remains in solution and decreases the amount of solid to be filtered.

When $Ca(OH)_2$ is added to the water at this pH > 10.3, the following will occur:

$$Ca^{2+} + {CO_3}^{2-} \rightarrow CaCO_{3\,(s)}$$

This solid is not harmful but will require a great deal more work to remove it and cause aesthetic and functional issues (scaling). It is not efficient and costly.

3. **Balance and write the chemical equations to identify the difference between chlorination using chlorine gas compared to sodium hypochlorite. Hint: HOCl (hypochlorous acid) is the disinfectant. List some pros and cons of each chlorination technique.**

$$Cl_{2\,(g)} + H_2O_{(l)} \leftrightarrow HOCl_{(aq)} + H^+{}_{(aq)} + Cl^-{}_{(aq)}$$

$$NaOCl_{(aq)} + H_2O_{(l)} \leftrightarrow HOCl_{(aq)} + NaOH_{(aq)}$$

Chlorine gas is a better oxidising agent (more effective) than sodium hypochlorite. Only part of the chlorine in the hypochlorite solution is freely available to react with water. Sodium hypochlorite is a liquid which easier and safer to use than gas.

Being a weak acid, HOCl partially dissociated to hypchlorite ion (OCl^-).

$$HOCl_{(aq)} \leftrightarrow H^+{}_{(aq)} + OCl_{(aq)}$$

The degree of dissociation varies with temperature and pH. An increase in pH will shift the equilibrium to the right. Sodium hypochlorite does this very slightly.

