(i) (This answer assumes the disodium salt of EDTA was used in the experiment.) During the titration, $H_2Y^{2-}_{(aq)}$ chelates (binds) free $Ca^{2+}_{(aq)}$ ions in solution, shifting the following equilibrium to the right:

$Ca^{2+}_{(aq)} + H_2Y^{2-}_{(aq)} \rightleftharpoons CaY^{2-}_{(aq)} + 2\ H^{+}_{(aq)}$

Performing this titration in a basic solution shifts this equilibrium to the right, ensuring that all the calcium in the sample has been complexed with EDTA. This is particularly important because calcium–EDTA complexes are weak as compared to metal–EDTA complexes.

Erichrome black T is a dark dye used to signal the endpoint of an EDTA titration. Like acid–base indicators, Erichrome black T is a weak acid (abbrev. $HIn^{2-}_{(aq)}$) which is in equilibrium with its conjugate base, In $In^{3-}_{(aq)}$:

$In^{3-}_{(aq)} + H^{+}_{(aq)} \rightleftharpoons HIn^{2-}_{(aq)}$

Initially, in the titration, there is a large excess of $Ca^{2+}_{(aq)}$. Some calcium combines with the indicator to produce a red-coloured complex, $CaIn^{-}_{(aq)}$.

$CaIn^{-}_{(aq)} + H^{+}_{(aq)} \rightleftharpoons HIn^{2-}_{(aq)} + Ca^{2+}_{(aq)}$

(red) (blue)

At the endpoint, chelation of the last trace of free $Ca^{2+}_{(aq)}$ shifts the Erichrome black T equilibrium to the right, resulting in the endpoint colour change.

(j) The sharpness of the endpoint increases with pH. However, there is an upper limit to how high the pH can be. If the pH is too high, calcium and magnesium ions begin to precipitate out of solution as $CaCO_{3(s)}$ and $Mg(OH)_{2(s)}$. As a result, a compromise of pH 10 is used for this titration. To maintain a constant pH during the titration, an ammonia/ammonium buffer solution is used. The buffering equilibrium involved is:

$NH^{+}_{4(aq)} \rightleftharpoons H^{+}_{(aq)} + NH_{3(aq)}$

Small amounts of acid ($H^{+}_{(aq)}$) produced in the flask are consumed by the forward reaction. Small excesses of base ($OH^{-}_{(aq)}$) are consumed by the reverse reaction.

(k) EDTA can be found in a large variety of consumer and pharmaceutical products:

Soaps – EDTA is sometimes added to soap to act as a water-softening agent. EDTA softens water by chelating calcium and magnesium from the water.

Canned fruits and vegetables – Metals are sometimes introduced into canned vegetables either from the soil or from harvesting or processing machinery. Metals can degrade food by catalyzing the oxidation of fat. EDTA chelates metals, preventing them from decomposing food.

Meat products – EDTA prevents the discoloration of some meat products.

Chelation therapy – EDTA is useful in the treatment of a variety of disorders, such as

- arteriosclerosis (hardening of the arteries). EDTA removes calcium deposits, making the arteries flexible again.
- metal poisoning. EDTA chelates the metal (e.g., lead), which can then be eliminated from the body in urine.
- kidney stones (which consist primarily of calcium compounds such as calcium oxalate). EDTA chelates calcium, which helps to shrink the stones.

(l) The medical use of EDTA is not without risk. Large doses of EDTA can damage the kidneys. Also, EDTA can cause a drop in the blood sugar levels. This is a particular concern for diabetics who use zinc-based insulin.

UNIT 4 SELF-QUIZ

(Page 636)

1. False: The equilibrium concentrations depend on the value of the equilibrium constant at any given temperature.
2. True
3. True
4. False: Catalysts lower the activation energy for both the forward and reverse reactions.
5. False: Inert gases have no effect on equilibrium concentrations.
6. False: The value of the equilibrium constant will decrease.
7. False: Calcium fluoride is more soluble.
8. True
9. False: The activation energy depends on whether the reaction is endothermic or exothermic.

10. False: The spontaneity of a reaction depends on enthalpy changes as well as entropy changes.
11. False: The pH of acetic acid is greater than 1.
12. False: The hypochlorite ion is a weaker base than ammonia.
13. False: Metal oxides form basic solutions while nonmetal oxides form acidic solutions.
14. False: Potassium sulfate forms a neutral solution.
15. True
16. True
17. False: The pH at the equivalence point depends on the type of acid and base involved.
18. True
19. False: Buffering action occurs during the flat portions of the graph.
20. True
21. False: An effective acid–base buffer contains approximately equal amounts of a weak acid and its conjugate base.
22. True
23. (b)
24. (b)
25. (e)
26. (b)
27. (b)
28. (c)
29. (b)
30. (e)
31. (c)
32. (c)
33. (d)
34. (c)
35. (e)
36. (c)
37. (b)
38. (a)
39. (d)
40. (d)
41. (e)

UNIT 4 REVIEW

(Page 639)

Understanding Concepts

1. $$\frac{[SO_{2(g)}]^2[O_{2(g)}]}{[SO_{3(g)}]^2} = K$$
$$= \frac{1}{279}$$
$$K = 3.58 \times 10^{-3}$$
The equilibrium constant for the given reaction has a value of 3.58×10^{-3}.

2. $$K = \frac{[NO_{(g)}]^2}{[N_{2(g)}][O_{2(g)}]}$$
$$= \frac{[0.15]^2}{[0.63][0.21]}$$
$$K = 1.7 \times 10^{-3}$$
The equilibrium constant is 1.7×10^{-3}.

3. (a) [CO] decreases
(b) [CO] decreases
(c) [CO] increases