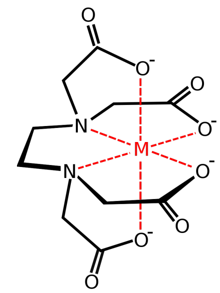
CHEM 1032 – Spring 2023 - Challenge Question 5 - key

How can metals be removed from water?



Removing metal ions from water is an important tool in health and in the environment. One way this can be achieved is by binding the metal ion with a molecule called a chelator and then removing the metal+chelator from solution. EDTA (ethylenediaminetetraacetic acid) is an example of one of these chelators and binds in 1:1 stoichiometries with many metal ions to make (M(EDTA)) complexes (as shown at right).

You may use "EDTA" as the chemical formula below in your equations.

A. Write out the chemical reaction and equilibrium expression for the binding of EDTA with Mg^{2+} and, separately, with Li^{+}.

EDTA + $Mg^{2+} \leftrightarrows (Mg^{2+}(EDTA))$

$$K = \frac{[(Mg^{2+}(EDTA))]}{[EDTA][Mg^{2+}]}$$

EDTA + $Li^{+} \leftrightarrows (Li^{+}(EDTA))$

$$K = \frac{[(Li^{+}(EDTA))]}{[EDTA][Li^{+}]}$$

B. EDTA binds more strongly with Mg^{2+} than with Li^{+}. With this information, assign the K values, $K_1 = 617$ and $K_2 = 4.63 \times 10^{8}$, to the correct metal-EDTA equilibrium expression. Explain your answer.

Since the question states that Mg^{2+} binds more strongly EDTA compared to Li^{+} it suggests that the $(Mg^{2+}(EDTA))$ is more stable than the $(Li^{+}(EDTA))$. If it is more stable, then more product will be present in the Mg^{2+} case than the Li^{+} case. If more product is present than the value of K should be larger comparatively, thus $K_2 = 4.63 \times 10^{8}$ should go with Mg^{2+} and $K_1 = 617$ should go with Li^{+}.

C. If a reaction were run where 100.0 mL of 0.100 M EDTA mixed with 100.0 mL of 0.100 M free metal, what would be the equilibrium concentration of free metal? Solve and show your work for one metal ion with EDTA and then repeat with the other metal ion.

100.0 mL 0.100 M EDTA
+
100.0 mL 0.100 M Mg^{2+}

total V 200.0 mL

M_2 EDTA = (100.0 mL)(0.100 M) = (200.0 mL)(x M)
= 0.050 M EDTA

M_2 Mg^{2+} = math same
= 0.050 M Mg^{2+}

	EDTA +	Mg^{2+} ⇌	Mg^{2+}(EDTA)
I	0.050	0.050	0
C	−x	−x	+x
E	0.050−x	0.050−x	x

All same for Li^+, except use 617 for K

$$4.63\times10^{8} = \frac{x}{(0.050-x)(0.050-x)}$$

$$4.63\times10^{8} = \frac{x}{0.0025 - 0.1x + x^2}$$

$$1.16\times10^{6} - 4.63\times10^{7}x + 4.63\times10^{8}x^2 = x$$

$$1.16\times10^{6} - 4.63\times10^{7}x + 4.63\times10^{8}x^2 = x$$

$$\frac{+4.63\times10^{7} \pm \sqrt{(-4.63\times10^{7})^2 - 4(4.63\times10^{8})(1.16\times10^{6})}}{2(4.63\times10^{8})}$$

$$\frac{4.63\times10^{7} \pm \sqrt{2.15\times10^{15} - 2.15\times10^{15}}}{9.26\times10^{8}}$$

$$\frac{4.63\times10^{7} \pm 0}{9.26\times10^{8}}$$

$x = 0.05$

Free Metal = 0.05 − 0.05 = 0 Mg^{2+}

$$617 = \frac{x}{(0.050-x)(0.050-x)}$$

$$617 = \frac{x}{0.0025 - 0.1x + x^2}$$

$$1.5425 - 61.7x + 61.7x^2 = x$$

$$1.5425 - 60.7x + 61.7x^2$$

$$\frac{+60.7 \pm \sqrt{(-60.7)^2 - 4(61.7)(1.5425)}}{2(61.7)}$$

$$\frac{60.7 \pm \sqrt{3684.5 - 380.7}}{123.4}$$

$$\frac{60.7 \pm 57.5}{123.4}$$

(cannot be value because final conc would be (−))

$x = 0.958$ or 0.026

Free Metal = 0.05 − 0.026 = 0.024 M Li^+

D. If you were to add more free metal to each of equilibrium solutions in Part C, what would you expect to happen? Explain your answer with Le Chatelier's principle and discuss the values of the reaction quotients.

If free metal ions were added to either reaction this would increase the concentration of the reactants and so the system would shift forwards generating more (M(EDTA)). For both metals, the increase in reactant concentration will result in reaction quotients which are smaller than the K values. When Q is less than K the forward reaction will occur until equilibrium is reestablished. So both would shift forward! However it's important to notice that in the Mg^{2+} the calculation in C suggests there is almost 0 unbound EDTA…thus there is not really any way for the forward reaction to happen any more.

E. If Mg^{2+} free metal were added to the equilibrium solution of $Li(EDTA)^{+}$ what would you expect to happen? Consider the equilibrium constants in your answer.

Building off Part B, where the question states that Mg^{2+} binds more strongly EDTA compared to Li^{+} it suggests that the (Mg^{2+}(EDTA)) is more stable than the (Li^{+}(EDTA)). We assigned that $K_2 = 4.63 \times 10^8$ should go with Mg^{2+} and $K_1 = 617$ should go with Li^{+}. Additionally, Part C shows there is a sizeable amount of unbound EDTA present in solution in the Li^{+} case. So if Mg^{2+} were added it would bind with the unbound EDTA. Thus would decrease the unbound EDTA in the Li+ equilibrium, forcing some of (Li^{+}(EDTA)) to go backwards and separate, due to Le Chatelier's principle.