Answers to Exercise 11.4
Nonmetal Oxides as Acids

1.

(a) i. $m_S = 70\ Tg$ $M_S = 32.066\frac{g}{mol}$ $M_{SO_2} = 64.065\frac{g}{mol}$

$M = \frac{m}{n}$ therefore $n = \frac{m}{M}$

$n_S = n_{SO_2}$ therefore $\frac{m_S}{M_S} = \frac{m_{SO_2}}{M_{SO_2}}$

$$m_{SO_2} = \frac{m_S \cdot M_{SO_2}}{M_S} = \frac{(70Tg)\left(64.065\frac{g}{mol}\right)}{\left(32.066\frac{g}{mol}\right)} = 140\ Tg = 1.4 \times 10^2\ Tg$$

ii. $m_S = 70\ Tg$ $M_S = 32.066\frac{g}{mol}$ $M_{SO_3} = 80.064\frac{g}{mol}$

$M = \frac{m}{n}$ therefore $n = \frac{m}{M}$

$n_S = n_{SO_3}$ therefore $\frac{m_S}{M_S} = \frac{m_{SO_3}}{M_{SO_3}}$

$$m_{SO_3} = \frac{m_S \cdot M_{SO_3}}{M_S} = \frac{(70Tg)\left(80.064\frac{g}{mol}\right)}{\left(32.066\frac{g}{mol}\right)} = 175\ Tg = 1.8 \times 10^2\ Tg$$

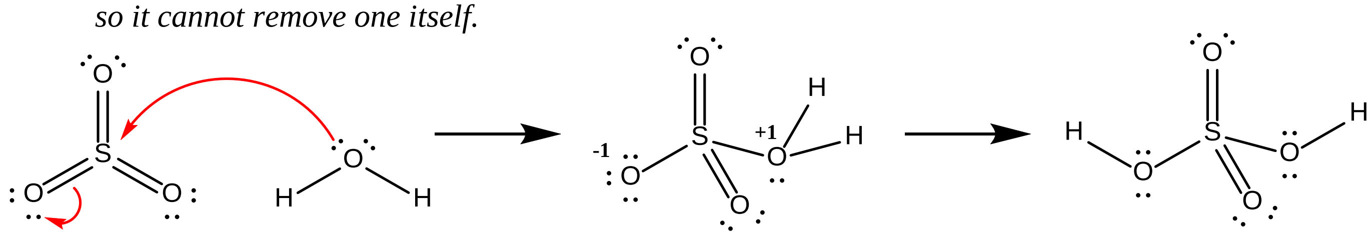
iii. $m_S = 70\ Tg$ $M_S = 32.066\frac{g}{mol}$ $M_{H_2SO_4} = 98.079\frac{g}{mol}$

$M = \frac{m}{n}$ therefore $n = \frac{m}{M}$

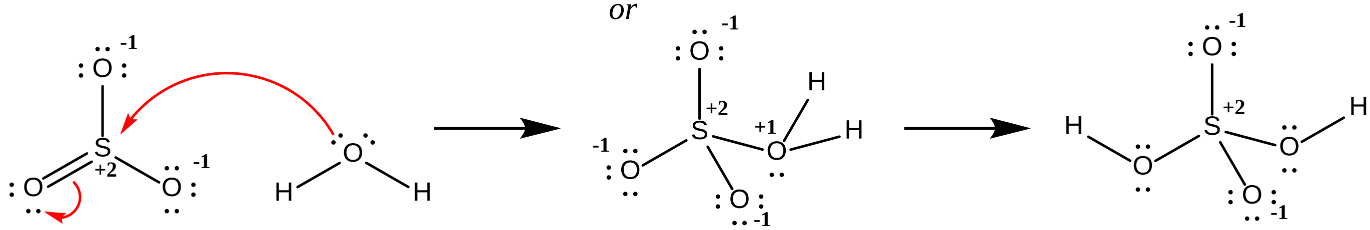
$n_S = n_{H_2SO_4}$ therefore $\frac{m_S}{M_S} = \frac{m_{H_2SO_4}}{M_{H_2SO_4}}$

$$m_{H_2SO_4} = \frac{m_S \cdot M_{H_2SO_4}}{M_S} = \frac{(70Tg)\left(98.079\frac{g}{mol}\right)}{\left(32.066\frac{g}{mol}\right)} = 214\ Tg = 2.1 \times 10^2\ Tg$$

(b) *The second step is shown to show the correct Lewis diagram for H_2SO_4. You are only expected to show electron movement using curly arrows for the first step. In the second step, two things must happen: a water molecule must remove one H^+ bonded to the positively charged O, and the negatively charged O must acquire H^+ (most easily shown as removing H^+ from an H_3O^+ ion). The O^- cannot reach either of the H attached to O^+ so it cannot remove one itself.*



or



(c) Lewis acid

(d) oxidation state of S in SO_3 is +6

oxidation state of S in H_2SO_4 is +6

When a nonmetal oxide reacts with water, the central atom of the oxoacid produced has the same oxidation state as it did in the nonmetal oxide!

(e) Like all nonmetal oxides, when carbon dioxide dissolves in water, it reacts with it to form an oxoacid:

$$CO_2(g) + H_2O(l) \rightleftharpoons H_2CO_3(aq)$$

This makes the ocean water more acidic:

$$H_2CO_3(aq) \rightleftharpoons H^+(aq) + HCO_3^-(aq)$$

The oceans now contain ~30% more H^+ than they did 250 years ago.

Carbonates like $CaCO_3$ react with acid to give soluble bicarbonates (e.g. $Ca(HCO_3)_2$) and, if there is enough acid, water and carbon dioxide:

$$CaCO_3(s) + H^+(aq) \rightleftharpoons Ca^{2+}(aq) + HCO_3^-(aq)$$

$$CaCO_3(s) + 2H^+(aq) \rightleftharpoons Ca^{2+}(aq) + H_2O(l) + CO_2(g)$$

Therefore, if the concentration of H^+ is too high, shells made of $CaCO_3$. Even at lower concentrations of H^+, new generations of aquatic creatures will not be able to make shells because the $CaCO_3$ will not precipitate. Because this is an equilibrium system, the presence of too much H^+ pushes the equilibrium away from $CaCO_3$ and toward Ca^{2+}.

This can happen at pH values above 7; it requires "enough H^+" not "more H^+ than OH^-".

Fortunately, carbonic acid (H_2CO_3) is a weaker acid than sulfuric acid (H_2SO_4); however, that does not make ocean acidification any less real. It just means that the oceans would already be "dead" if carbonic acid was as strong as sulfuric acid. (Humans "woke up" to the dangers of acid rain when "dead lakes" started becoming more common. Hopefully, we will not get to the point of having any "dead oceans".)