

Smith et al. (1969) did not focus on sulfate formation kinetics; instead, they illustrated through a novel experiment the ability of solid particles to adsorb  $\text{SO}_2$  and to release  $\text{SO}_2$  during passage through a tube with a wall that adsorbes  $\text{SO}_2$ . They measured the number of  $\text{SO}_2$  monolayers absorbed on suspended  $\text{Fe}_3\text{O}_4$  as function of  $\text{SO}_2$  partial pressure. (The monolayer coverage data reported in their Table I are in error by a factor of 100 too large; e.g., the number of monolayers at 1.13 ppm should be  $0.38 \times 10^{-2}$ .)

Chun and Quon (1973) measured the reactivity of ferric oxide to  $\text{SO}_2$ , using a flow system involving a filter containing suspended particles. They determined a removal rate constant of  $9.4 \times 10^{-3} \text{ ppm}^{-1} \text{ min}^{-1} [-d(\ln \rho)/dt]$ , where  $\rho$  is the fraction of surface sites available for reaction. Extrapolating this to an atmospheric particle concentration of  $100 \text{ ug}/\text{M}^3$  with an equivalent reactivity and an  $\text{SO}_2$  concentration of 0.1 ppm, the data project an atmospheric removal rate of 0.1 percent per hour.

Stevens et. al. (1978) report total iron concentrations in six U. S. cities ranging between 0.5 and  $1.3 \text{ ug}/\text{m}^3$ . Other species such as manganese, copper, or vanadium had total concentrations usually below  $0.1 \text{ ug}/\text{m}^3$ . Thus actual ambient air concentrations are a factor of approximately 50 times less than that assumed by the authors in the above papers. A reactive particle concentration of  $2 \text{ ug}/\text{m}^3$  would yield a predicted  $\text{SO}_2$  removal rate of no more than 0.002 percent