ESS C113/213: Homework 2
Due April 18, 2013

1. Tropospheric Ozone
In the urbanized LA Basin, ozone forms mainly from the reaction of NO_2 with sunlight, and its concentration reflects a steady-state homogenous equilibrium between NO, NO_2 and O_2:

$$NO_2 + O_2 \leftrightarrow NO + O_3$$

Seinfeld (1989) proposed a relationship to estimate the equilibrium abundance of O_3 in such environments:

$$O_3 \text{ (ppmv)} = 0.021 \, [NO_2]/[NO]$$

Use this relationship to estimate the abundance of ozone when 18 ppbv NO_2 and 15 ppbv NO are present (ppbv = parts per billion by volume; ppmv = parts per million by volume).

Convert this abundance to the Air Quality Index (AQI) scale (http://daq.state.nc.us/Ozone/codecalc.shtml), assuming a constant abundance over an 8-hour period. How much ozone would be enough to trigger a "red" or "unhealthy" air quality? Find the current ground-level ozone AQI for Los Angeles, and convert it to ppm Ozone. You can find this data online at http://www.airnow.gov, our at our local air quality management district's site (http://www.aqmd.gov/smog/Reading.html). Please note the date of the observation, and its location (i.e., Santa Monica, LA inland, San Bernardino, etc.).

According to Schlesinger, figure 3.10, what (roughly) would the ozone concentration be in a forested area during the growing season, if NO_2 = 18 ppbv, NO = 15 ppbv, and the total of all other reactive nitrogen species is 4 ppbv? Why is this different from urban air?

2. CFC's and Stratospheric Ozone
We learned in class that enough air moves from the troposphere to the stratosphere each year to replace about 75% of the stratosphere's mass, indicating a mean residence time of 4/3 years for air in the stratosphere (this is an oversimplification, but let's keep things simple for now). Given that the troposphere has 4 times the mass of the stratosphere, what is the mean residence time of air in the troposphere?

The widespread use and release of chlorofluorocarbons into the atmosphere has led to the rapid seasonal loss of ozone in the stratosphere. This follows, at least in part, from the inertness of CFC molecules in the troposphere. Freon-11 ($CFCl_3$) has a mean atmospheric lifetime of about 52 years. This means that:

$$dF/dt \approx -(1/52) \bullet F \text{ /year}$$

where "F" is the molecular abundance of $CFCl_3$. Show that the expression:

$$F(t) = F_{initial} \bullet \exp(-t/52)$$

is a solution to this differential equation.

If your (hypothetical, obsolete) refrigerator springs a tiny leak and releases 1000 molecules of Freon-11, how many will survive 5 years to enter the stratosphere at least once?

Production of CFC's was essentially ended in the 1990's by the Montreal Protocol (http://en.wikipedia.org/wiki/Montreal_Protocol you don't have to read it, this link is just for kicks). Assume that production and release stopped completely by 1995 (see Fig. 1).

On Fig. 1b, plot the expected trend in CFC-11 concentration from 1995 to 2009, if production stopped completely in 1995. Does the trend match observed concentrations?

Given the mean lifetime of CFC-11 in the atmosphere, and assuming production ceased in 1995, in what year will the concentration take fall below its average 1977 abundance?

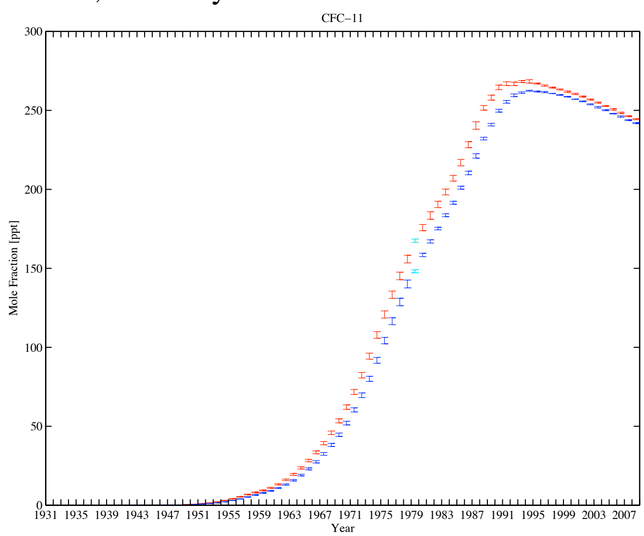

Figure 1. Concentrations of $CFCl_3$ in air. a) historical concentrations in the southern (blue) and northern (red) hemispheres, in pptv. Data from Walker et al. (2009) Reconstructed histories of the annual mean atmospheric mole fractions for the halocarbons CFC-11, CFC-12, CFC-113 and Carbon Tetrachloride. http://bluemoon.ucsd.edu/pub/cfchist/ b) recent globally averaged concentrations (1978-2013). Data set from NOAA Earth System Research Laboratory Global Monitoring Division. http://www.esrl.noaa.gov/gmd/hats/combined/CFC11.html