Chemistry 101
Natural World Assignment 3 (20 pts)
Due Wednesday, December 3rd

Background

CFCs, or chlorofluorocarbons, are compounds which contain only carbon, chlorine and fluorine. These compounds have certain properties which make them very effective as refrigerants. In particular, they are:

- very stable
- nonflammable
- non-toxic
- inexpensive
- widely available

In addition, their boiling points are in the correct range for use as refrigerants (-10 to -30°C is the ideal range, but values outside this range may also work). As a result, CFCs were widely used in refrigerators and auto, home and commercial refrigeration units. In addition, some of these properties made them useful as propellants for aerosols, and as blowing agents for plastic production (e.g. Styrofoam production).

CFCs are artificially produced and do not occur naturally. When first developed, they were hailed as a great achievement of the chemical industry. For many applications, CFCs replaced earlier refrigerants, such as the more toxic and reactive chemicals sulfur dioxide (SO_2) and ammonia (NH_3). Two common CFCs in use in the mid 1980's were trichlorofluromethane, CCl_3F (CFC-11), and dichlorodifluoromethane, CCl_2F_2 (CFC-12). Worldwide production of these two compounds in 1985 was 850,000 tons.

The stability of CFCs makes them an excellent choice for many commercial applications. However, this stability leads to special environmental problems. Because they are so stable, CFCs survive to reach the stratosphere, the region of the atmosphere 15-50 kilometers above sea level. This is the region of the atmosphere that contains the ozone layer, a blanket that protects us from ultraviolet (UV) light. In particular, it shields us from UV light with wavelengths in the range of 220 nanometers to 320 nanometers. UV light in this wavelength range is harmful to biological organisms. It is capable of ionizing and breaking bonds in DNA, for example, leading to some cancers. At sufficiently high levels, UV radiation can be fatal. Thus, the ozone layer is essential to life on earth.

High in the stratosphere, the intensity of ultraviolet light is such that CFCs can be broken down. Energy from the ultraviolet light from the sun is capable of breaking carbon-chlorine bonds (C-Cl) easily. When this happens, free chlorine atoms are formed:

$$CCl_2F_2 + UV\ light \longrightarrow CClF_2 + Cl$$

These chlorine atoms are VERY reactive. They react with ozone to form O_2 and ClO, another very reactive species:

$$Cl + O_3 \longrightarrow ClO + O_2$$

ClO then goes on to regenerate Cl atoms:

$$ClO + O \longrightarrow Cl + O_2$$

This cycle is repeated on average 100,000 times before the Cl atom finds a "sink" - something that removes it from the stratosphere. Thus, one CFC molecule can be responsible for destroying about 100,000 ozone molecules. F. Sherwood Rowland, Mario Molina and Paul Crutzen received the Nobel Prize in 1995 for their work uncovering CFCs as a major cause of ozone depletion. Their work, along with the work of many other scientists around the globe, has lead to a global phaseout of CFCs and other ozone depleting chemicals.

Finding replacement compounds for CFCs is not easy. Changing the composition of the compounds even slightly alters their chemical and physical properties, potentially making them less suitable for the commercial application. Two classes of compounds have been suggested as alternatives: HCFCs and HFCs. HCFCs, or hydrochlorofluorocarbons, contain hydrogen as well as chlorine, fluorine and carbon. HFCs, or hydrofluorocarbons, contain no chlorine, but only hydrogen, fluorine and carbon. The ozone depletion potential of HCFCs and HFCs is lower than CFCs for a number of reasons.

First, having fewer (or no) chlorines in a molecule leads to a lower ozone depletion potential for that molecule, as chlorine is the major culprit in ozone depletion. However, it should be noted that any chlorinated compound will still have some ozone depletion potential.

The second reason these compounds have lower ozone depletion potentials is because HCFCs and some HFCs are less stable than CFCs. The presence of hydrogen in the compound decreases the compound's stability, making it more likely to react, ultimately leading to a shorter atmospheric lifetime. This is generally desirable, as the molecule may not survive long enough to reach the stratosphere. Unfortunately, too many hydrogens makes some of these compound flammable.

One additional thing to consider regarding CFC replacements is the role such compounds may play in global warming. All these compounds, CFCs, HFCs and HCFCs, can potentially contribute to global warming. However, very stable compounds have long atmospheric lifetimes. Molecules with long atmospheric lifetimes will continue to contribute to global warming for many years. CFCs and their replacement exhibit a wide range of atmospheric lifetimes.

As you can see finding suitable replacements for CFCs is a complicated process.

Your Task

You work for a refrigeration company that uses CFC-12 (CCl_2F_2), but must now must switch to a different chemical. Your company uses about 250 tons of CFC-12 annually.

You have been assigned to review some basic information about five possible substitute compounds. (You can find this information on the next page.) You must choose one of these to recommend to your immediate supervisor to replace CFC-12.

Write a two page "mini-essay" that lays out your chosen replacement and your argument(s) in favor of your choice. There is no right or wrong answer to this question. Rather, you should seek to lay out a well reasoned argument for your choice.

If you would like some additional information, links to the U. S. Environmental Protection Agency's (EPA) website are provided on the course web site (see the homework and assignments link). You are free to consult our textbook, other books/articles and additional websites as well.

Cite the sources of all information. Any standard reference format is fine. (The library's "writing and citing" link off the main page may be helpful.)

The essay should be word processed in 10 or 12 point font, with 1" margins. Please put your name at the top.

Special note for this assignment: Late assignments will lose 2 points per weekday. (A certificate may be used to avoid the late penalty.)

CFCs: Here is some information about CFC-12 (the one your company is currently using, and CFC-11, another common refrigerant. (NOTE: ODP and GWP are defined at the bottom of the next page.)

Name	Formula	ODP*	atmospheric lifetime	GWP	cost per Kg (retail)	boiling point
CFC-11 trichloro-fluoromethane	CCl_3F	1.0	45 years	4680	$3.30	24°C
CFC-12 dichloro-difluoromethane	CCl_2F_2	1.0	100 years	10720	$3.45	-30°C

CFC Alternatives: Here is some information about the five possible replacement copmounds. (NOTE: ODP and GWP are defined at the bottom of the page.)

Name	Formula	ODP*	atmospheric lifetime	GWP	cost per Kg (retail)	boiling point
HCFC-22 chloro-difluoromethane	$CHClF_2$	0.05	12.0 years	1780	$3.90	40°C
HCFC-123 dichloro-trifluoroethane	$CHCl_2$-CF_3	0.02	1.3 years	76	$9.00	27°C
HCFC-141b* dichloro-fluoroethane	CH_3-CCl_2F	0.12	9.3 years	713	$3.80	32°C
HFC-134a tetrafluoro-ethane	CF_3-CH_2F	0 (no chlorines)	14 years	1320	$8.30	-26°C
HFC-125 pentafluoro-ethane	CHF_2-CF_3	0 (no chlorines)	29 years	3450	$8-10	-46°C

* denotes flammable gas.

The global warming potential (GWP) represents how much a given mass of a chemical contributes to global warming over a given time period compared to the same mass of carbon dioxide. Carbon dioxide's GWP is defined as 1.0.

The ozone depletion potential (ODP) is the ratio of the impact on ozone of a chemical compared to the impact of a similar mass of CFC-11.

References

1. Baird, Colin; *Environmental Chemistry*; W. H. Freeman: New York; 1995.
2. American Chemical Society; *Chemistry in Context: Applying Chemistry to Society*; 4th Ed.; McGraw Hill: Boston; 2003
3. Wayne, R. P. *Chemistry of Atmospheres*; 2nd Ed.; Oxford University Press: Oxford, 1991.
4. U.S. Environmental Protection Agency; *Class I Ozone-Depleting Substances*; available online: http://www.epa.gov/ozone/ods.html (accessed November 2003).
5. U.S. Environmental Protection Agency; *Class II Ozone-Depleting Substances*; available online: http://www.epa.gov/ozone/ods2.html (accessed November 2003).
6. The National Institute of Standards and Technology (NIST); *NIST Standard Reference Database*; No. 69, 2003; available online: http://webbook.nist.gov/chemistry/ (accessed November 2003).