6. Acid Rain and the Greenhouse Effect

Several of our major environmental problems arise from interactions between water and the atmosphere. One problem is acid rain, which affects large areas of the United States, Canada, and Europe. **Acid rain** forms when certain air pollutants dissolve in raindrops or in lakes and streams. The acid in the rain damages many plants, sometimes entire forests. It alters the quality of streams and lakes, harming aquatic organisms. Runoff from acid rain pollutes estuaries and seeps into groundwater. The acid also damages artifacts, the products of human activity. Steel bridges, limestone buildings, and marble statues are being eroded by acid rain.

How does acid rain form? To answer this question we need to know some **acid-base** chemistry.

Acids and Bases

We can think of **acids** as compounds that dissolve in water and release hydrogen ions (H^+). **Bases** are compounds that dissolve in water to release hydroxide ions (OH^-). When there are more H^+ ions than OH^- ions in a water solution, it is **acidic**. **Basic** solutions have more OH^- ions than H^+ ions. Some common acids and bases and the ions they form in water are listed in Table 6–1. Note that other definitions of acids and bases exist. The ones used here are the ones that are most helpful in describing acid-base chemistry in seawater.

Pure water is neither an acid nor a base. It is **neutral,** meaning that it contains equal amounts of H^+ and OH^- ions. Since water is a covalent molecule, it does not readily dissociate into ions. But among the billions and billions of molecules in a container of water, strong polar attractions break a few apart, forming a small but equal number of H^+ ions and OH^- ions. These ions readily reform into water, so the reaction is said to be **reversible**. See Fig. 6–1.

The pH Scale

The **pH scale** shown in Fig. 6–2 gives a sense of the concentration of hydrogen ions in a solution.

- *Neutral solutions have a pH of 7.* At pH 7 there are equal numbers of H^+ ions and OH^- ions. Pure water is neutral.

- *Acid solutions have low pH readings.* Strong acids have pHs of 2 or less; weak acids range from 2 to less than 7.

- *Basic solutions have high pH readings.* Strong bases have pHs of 12 or more; weak bases range from more than 7 to less than 12.

$$H^+ + OH^- \rightleftharpoons HOH$$

H^+	OH^-	HOH
hydrogen ion	hydroxide ion	water

Fig. 6–1. Reversible water-ion reaction

Table 6–1. Some common acids and bases

Acid or base	Common name (Chemical name)	Ions formed in water		
Strong acids				
HCl	Hydrochloric acid (Hydrogen chloride)	H^+ hydrogen ion	+	Cl^- chloride ion
HNO_3	Nitric acid (Hydrogen nitrate)	H^+ hydrogen ion	+	NO_3^- nitrate ion
H_2SO_4	Sulfuric acid (Hydrogen sulfate)	H^+ hydrogen ion	+	HSO_4^- hydrogen sulfate ion
Weaker acids				
$HC_2H_3O_2$	Acetic acid, vinegar (Hydrogen acetate)	H^+ hydrogen ion	+	$C_2H_3O_2^-$ acetate ion
H_2CO_3	Carbonic acid (Hydrogen carbonate)	H^+ hydrogen ion	+	HCO^{3-} bicarbonate ion
$H_2SO_3^-$	Sulfurous acid (Hydrogen sulfite)	H^+ hydrogen ion	+	HSO_3^- hydrogen sulfite ion
Strong bases				
KOH	(Potassium hydroxide)	K^+ potassium ion	+	OH^- hydroxide ion
$NaOH$	Lye (Sodium hydroxide)	Na^+ sodium ion	+	OH^- hydroxide ion
Weak bases				
NH_4OH	Household ammonia (Ammonium hydroxide)	NH^{4-} ammonium ion	+	OH^- hydroxide ion
$Ca(OH)_2$	Slaked lime (Calcium hydroxide)	Ca^{2+} calcium ion	+	$2OH^-$ hydroxide ions

Acidic						Neutral						Basic		
0	1	2	3	4	5	6	7	8	9	10	11	12	13	14

Fig. 6–2. pH scale

An acid is **neutralized** when excess H^+ ions are chemically tied up or removed from the solution, leaving equal numbers of H^+ and OH^- ions. One way to neutralize an acid (H^+) is to add enough base (OH^-) to form water (H_2O). Another way is to add certain metals to the acid. The acid reacts with the metal, converting the H^+ ions to gaseous hydrogen (H_2), which then escapes. See Fig. 6–3.

Measuring pH

Electrical **pH meters** give precise measures of pH. A less precise but faster method uses **pH paper** strips with indicator dyes in them. Different concentrations of acid turn the paper different colors. We will use pH paper in Activity 1. The paper has a range of pH 1 to pH 14.

ACTIVITY 1

Determine the pH of common substances and list them in order on a pH scale.

MATERIALS

- substances from Table 6–2
- small beakers or watch glasses
- glass stirring rod
- pH test papers with standard color chart
- colored pencils or pens

PROCEDURE

1. Determine the pH of common substances by testing as follows:
   a. To determine the pH of a liquid substance, put a few drops in a small beaker or on a watch glass. Dip the end of a clean glass stirring rod into the liquid. Touch the moist end of the

A. Acid reacts with a base.

H^+	+	OH^-	→	HOH
hydrogen ion from an acid		hydroxide ion from a base		water (usually written as H_2O)

B. Acid reacts with metal.

Fe	+	2 H^+	→	Fe^{2+}	+	H_2
iron		hydrogen ions		ferrous ion		hydrogen gas

Fig. 6–3. Two ways to neutralize an acid

Table 6–2. Approximate pH of common substances Ⓦ

Substance	pH
Ammonia	
Apple juice	
Baking soda	
Battery acid	1.0
Bleach	
Blood	7.3–7.5
Carrots	5.0
Cola	
Cow's milk	
Crackers	7.0–8.5
Distilled water	7.0
Lemon juice	
Lye (drain opener)	13.0
Milk of magnesia	10.5
Saliva (human)	
Seawater	
Tap water	
Tomatoes	
Vinegar	

stirring rod to a small piece of pH test paper. Compare the color on the pH test strip with the standard colors. Determine the pH. Record data on Table 6–2.

b. To determine the pH of a solid substance, put a small piece of the solid in a beaker. Add a few drops of distilled water, mash the material, and stir it into the water. Test for pH as you did in step (a).

2. If time permits, test other substances, such as foods, beverages, cosmetics, household cleaning substances, and aquarium water.

3. Sketch the pH scale from Fig. 6–2 in your notebook. Write the names of the substances above the line in order of their increasing pH. Include all the substances listed in Table 6–2 and any others you tested.

4. Using colored pencils, color-code the pH scale to show acids, neutral substances, and bases. Make a key to explain your code.

QUESTIONS

1. Explain why pure water is described as neutral.

2. When is a substance said to be acidic? When is it said to be basic?

3. Judging from your observations and the information in Table 6–2, are most common foods neutral, acidic, or basic?

4. How do the pH values compare
   a. between seawater and distilled water?
   b. between distilled water and tap water?

pH of Rain

Pure water itself is neutral, but rainwater is not pure water. Rainwater is naturally slightly acidic, ranging in pH from 5.0 to 6.0. The falling drops in rain wash the air of particles and soluble gases. Carbon dioxide (CO_2) dissolves in the raindrops, forming carbonic acid (H_2CO_3), a weak acid. See Fig. 6–4.

Acid rain is any form of water precipitation (rain, fog, snow) having a pH lower than normal rainwater. It forms when certain gaseous air pollutants dissolve in atmospheric water droplets. In Europe, the eastern United States, and eastern Canada, acid rain

CO_2	+	H_2O	$\longrightarrow$	H_2CO_3
carbon dioxide		water		carbonic acid (a weak acid)

Fig. 6–4. Carbon dioxide dissolves in water, forming a weak acid.

is common in the range of pH 4.0 to pH 4.6 and occasionally as low as pH 2.

Acid rain is linked to the water cycle. It may fall from clouds near where it forms or hundreds of miles away. When it falls on land, the acid percolates into the soil or flows as runoff into streams and lakes. Some bodies of water are also acidic from the natural presence of decaying vegetation; some become acidic from farm fertilizer runoff and waste from mines and other industries.

Air Pollutants That Form Acid Rain

Two of the the major substances that form acid rain are sulfur dioxide (SO_2) and nitrogen dioxide (NO_2). Small amounts of these gases are produced naturally by soil bacteria and volcanoes. These compounds of nitrogen and sulfur are needed to sustain life, so small quantities benefit the environment. But in excessive quantities they form destructive acid rain.

Excess sulfur dioxide and nitrogen dioxide are produced by industrial power plants burning coal and gas. Oxides of nitrogen are also produced by motor vehicles. On smoggy days, reddish layers of nitrogen dioxide (NO_2) gas are sometimes visible over urban areas, particularly during heavy traffic.

The chemical reactions that produce acid rain are complex. Simplified summaries of the reactions are shown in Fig. 6–5. Both **ozone** (O_3), a gas, and water must be present in the air to produce nitric acid

Nitric acid production

N_2 nitrogen	+	$2\ O_2$ oxygen			$\longrightarrow$	$2\ NO_2$ nitrogen dioxide		
$2\ NO_2$ nitrogen dioxide	+	O_3 ozone	+	H_2O water	$\longrightarrow$	$2\ HNO_3$ nitric acid	+	O_2 oxygen

Sulfuric acid production

S sulfur	+	O_2 oxygen	$\longrightarrow$	SO_2 sulfur dioxide
$2\ SO_2$ sulfur dioxide	+	O_2 oxygen	$\longrightarrow$	$2\ SO_3$ sulfur trioxide
SO_3 sulfur trioxide	+	H_2O water	$\longrightarrow$	H_2SO_4 sulfuric acid

Fig. 6–5. Summary of reactions that form acids in the atmosphere

(HNO_3) from nitrogen dioxide (NO_2). Sulfuric acid (H_2SO_4) forms in the atmosphere when gaseous sulfur dioxide reacts with oxygen and water.

Neutralizing Acid Rain

Acid rain can be neutralized by reaction with calcium (Ca) and magnesium (Mg) compounds in soil. Limestone ($CaCO_3$), a substance common in soil, neutralizes acids when H^+ ions replace calcium-producing Ca^{2+} ions. See Fig. 6–6. Deep soils rich in metallic compounds have a greater capacity to neutralize acid rain than shallow soils or soils composed of nonmetals do. If the acid rain is not neutralized, the acid percolates into the ground or flows into lakes and streams.

Changes in the pH of water have damaging effects on aquatic life, including microorganisms, plants, and animals living in the lakes. All aquatic organisms have a pH **tolerance** or range they can live within. See Table 6–3. They soon die in water outside their tolerance. Some organisms appear to be more tolerant of acidic environments than others, and some adult organisms may be more tolerant than their eggs and embryos are. Ocean animals tolerate a pH range that is more basic than freshwater animals can tolerate. The pH at the ocean surface is 8.1 to 8.3. In the midocean column it is close to pH 7.

Table 6–3. Preferred pH tolerance for some freshwater aquatic organisms

Organism	pH tolerance
Frog & toad embryos	
American toad	5.0 to 6.5
bullfrog	5.0 to 6.5
woodfrog	4.5 to 6.5
Fish	
flathead minnow	6.5
pumpkinseed sunfish	5.5 to 6.5
smallmouth bass	6.0 to 6.5
brown trout	5.5 to 6.5
rainbow trout	6.0 to 6.5
brook trout	5.0 to 6.5
yellow perch	5.0 to 6.5
Others	
clam (some species)	6.5
crayfish	6.0 to 6.5
snail	6.5
spotted salamander embryo	5.5 to 6.5

$$CaCO_3 + H^+ \longrightarrow Ca^{2+} + HCO_3^-$$

$CaCO_3$	H^+	Ca^{2+}	HCO_3^-
calcium carbonate (limestone)	hydrogen ion (from acid)	calcium ion	bicarbonate ion

Fig. 6–6. An acid can be neutralized by reacting with limestone. (This equation shows only the reacting compounds and ions. The H^+ ion could have been contributed by any acid, for example, the air pollutant HNO_3. In this kind of equation the charges on either side of the arrow must balance.)

Other Effects of Acid Rain

Acid rain appears to contribute to the rapid dying off of forests since 1960 in some areas of Europe and the eastern United States and Canada. About 40% of the maple trees in some Canadian forests and 50% of the red spruce in mountainous areas of New England have died. Findings show that acid rain may cause nutrient deficiency by depriving roots of necessary trace metals through acid-induced leaching. It may decrease photosynthesis through direct damage to leaves and needles. Weakened plants are less resistant to disease and rot, and they are less able to withstand such natural stresses as drought and cold.

Acid rain also damages our human creations. Because acids corrode metals, acid rain weakens bridges. It discolors paint and dissolves marble and limestone buildings. Ancient Greek temples that have withstood centuries of weathering are deteriorating rapidly as acid rain dissolves the stone from which they are built.

QUESTIONS

5. What is meant by acid rain? What causes it? What problems does it cause?

6. In what way is acid rain related to
   a. the water cycle?
   b. the chemistry of water?
   c. the chemistry of soil?

7. How is acid rain neutralized in nature?

Acid Rain and the Oceans

Acid rain has little effect on seawater for several reasons. Acid rain falling on the ocean is quickly diluted, and the acids react with seawater, which is naturally basic. In addition, seawater is a **buffered** solution, meaning that its chemical reactions can keep the pH of seawater quite constant, somewhere between 8.1 and 8.3.

The ocean's major buffering system is the **carbonate buffering system**. This buffer system has two components. The first is the weak acid, carbonic acid (H_2CO_3), which forms when carbon dioxide (CO_2) from animal and plant respiration and from the atmosphere dissolves in seawater. By a **weak acid** we mean one that tends to ionize slightly and release only a little of its hydrogen component as H+ ion.

The second component is the carbonate ion (CO_3^{2-}) Carbonate ions come from carbonate components dissolved in seabed muds, from carbonates in the skeletons of coral and other animals, and from carbonate minerals washed into the sea from land.

Nitric acid (HNO_3) and sulfuric acid (H_2SO_4), the major acids in acid rain, are called **strong acids**. Strong acids readily ionize in water, releasing their hydrogen as H^+ ions. The carbonate (CO_3^{2-}) ions in the seawater tend to react with the hydrogen ions (H^+), making the weak acid carbonic acid (H_2CO_3). Since carbonic acid is a weak acid that holds its hydrogen tightly, H^+ ion is removed from the water and the pH rises. See Fig. 6–7.

This buffer system also works to neutralize excess hydroxide ions. If, for example, there is a sudden excess of hydroxide (OH^-) ions from a spill of a sodium hydroxide (NaOH) solution, the weak carbonic acid (H_2CO_3) will produce the hydrogen ions (H^+) necessary to neutralize the excess OH^- ion, reducing the pH. See Fig. 6–7.

Acid reaction								
CO_3^{2-}	+	H_2SO_4	$\longrightarrow$	H_2CO_3	+	SO_4^{2-}		
Net acid reaction								
CO_3^{2-}	+	$2\,H^+$	$\longrightarrow$	H_2CO_3				
Base reaction								
H_2CO_3	+	$2\,NaOH$	$\longrightarrow$	$2\,H_2O$	+	$2\,Na^+$	+	CO_3^{2-}
Net base reaction								
H_2CO_3	+	$2\,OH^-$	$\longrightarrow$	$2\,H_2O$	+	CO_3^{2-}		

Fig. 6–7. The carbonate buffer interacting with an acid and a base

QUESTIONS

8. Explain what is meant by a buffer solution.

9. In what ways do the following act to keep pH relatively constant?
   a. seawater
   b. calcium carbonate sediments
   c. green plants

10. What changes in pH might cause the ocean
    a. to release carbon dioxide into the atmosphere?
    b. to convert carbon dioxide into other substances?

11. Compare fresh water and seawater in terms of their buffering capacity.

ACTIVITY 2

Design an experiment that simulates acid rain.

MATERIALS

- plastic cups or beakers
- one or more weak acids (a carbonated drink can be used)
- a strong acid (pH of 2 or less)
- pH paper
- metal objects: coins, nails, bits of metals, pieces of wire
- chalk, chunks of coral, pieces of limestone or marble
- forceps or spoons
- seawater

PROCEDURE

1. Use pH paper to determine the pH of the acids you use. If you use a carbonated drink for your weak acid, test its pH immediately after opening it and again after it has gone flat, that is, when all bubbles have escaped. Record your data in your notebook.
2. Design an experiment investigating a specific problem related to pH and acid rain. Refer to Table 6–4 for suggestions. Carry out the steps listed below.
   a. With your team members, make a written research plan. State your hypothesis or research question.

b. List the materials or equipment you will need.
c. Write out your procedure and construct the data tables you will use to record your observations.
d. Review your plan with your teacher and others in the class.
e. Obtain your teacher's approval before beginning the experiment.

3. Report the results of your investigation.
   a. Organize your results in data tables and graphs.
   b. Write your conclusions.
   c. Prepare an oral report for the class. Use charts and diagrams to explain your experiment and its results.

4. Take notes on reports of experiments conducted by others in the class. Use this information to answer the questions that follow.

Table 6–4. Suggestions for investigating pH and acid rain

1. Simulate the effects of acid rain on metal.
   a. Test the effect of acid on different types of metals. Determine which metals react the most, the least, and not at all.
   b. Test the effect of acid on combinations of two or more different metals in contact with each other.
   c. Determine how the metal affects the acid, if at all. Measure the pH before and after the reactions with metals.

2. Test what acid rain does to materials made of calcium carbonate. Substances that contain calcium carbonate are chalk, coral, seashells, marble, and limestone.
   a. Devise a test that compares how acids react with solid substances containing calcium carbonate. Observe reactions after 14 to 20 minutes and after 24 hours. Record your data.
   b. Determine how the calcium carbonate affects the acid. When the acid stops reacting with the calcium carbonate, measure the pH of the final liquid.

3. Simulate acid rain being diluted in progressively larger bodies of fresh water. Start with a raindrop added to a puddle, then to a small pool, and then to a larger pool. Determine how pH changes when an acid is diluted from 100% to 10%, to 1%, to 0.1%, and to 0.01%. Measure the pH of each dilution.

4. Compare the effects of acid rain on
   a. fresh water,
   b. hard water (water containing high quantities of dissolved calcium), and
   c. seawater.

   Obtain a 100-mL sample of each liquid. Measure its pH. Add 5 mL of "acid rain," stir, then measure the pH.

5. Make a study of the pH of bodies of water in your community. Collect and test samples of rainwater and water from puddles, ponds, lakes, the ocean, and other bodies of water.

QUESTIONS

12. How does the pH of an acid change when the acid is diluted? When it is concentrated?

13. Some rain in a region may be acid. Other rain may be freshwater rain. How would light and heavy freshwater rain affect acid rain already in a rain gauge?

14. On the basis of your experiments, what statements can you make about reactions between acid rain and
    a. metallic objects?
    b. calcium carbonate objects?

15. What objects in our environment would be damaged by acid rain? What is your evidence?

Global CO_2 Problems

Another gaseous substance that can damage the environment when there's too much of it is carbon dioxide. Carbon dioxide (CO_2) is one of the naturally formed gases that make up the atmosphere. Natural processes that constantly release CO_2 into the atmosphere are shown by solid arrows in Fig. 6–8. Each year about 77 trillion tons of CO_2—about 280 parts of CO_2 per million parts of air—are released into the atmosphere by such natural processes as

- respiration by plants and animals,
- decay of organic matter,
- weathering of rocks and skeletons,
- release of dissolved CO_2 from water, and
- natural forest fires.

All plants and animals in their respiration return both CO_2 and water vapor to the

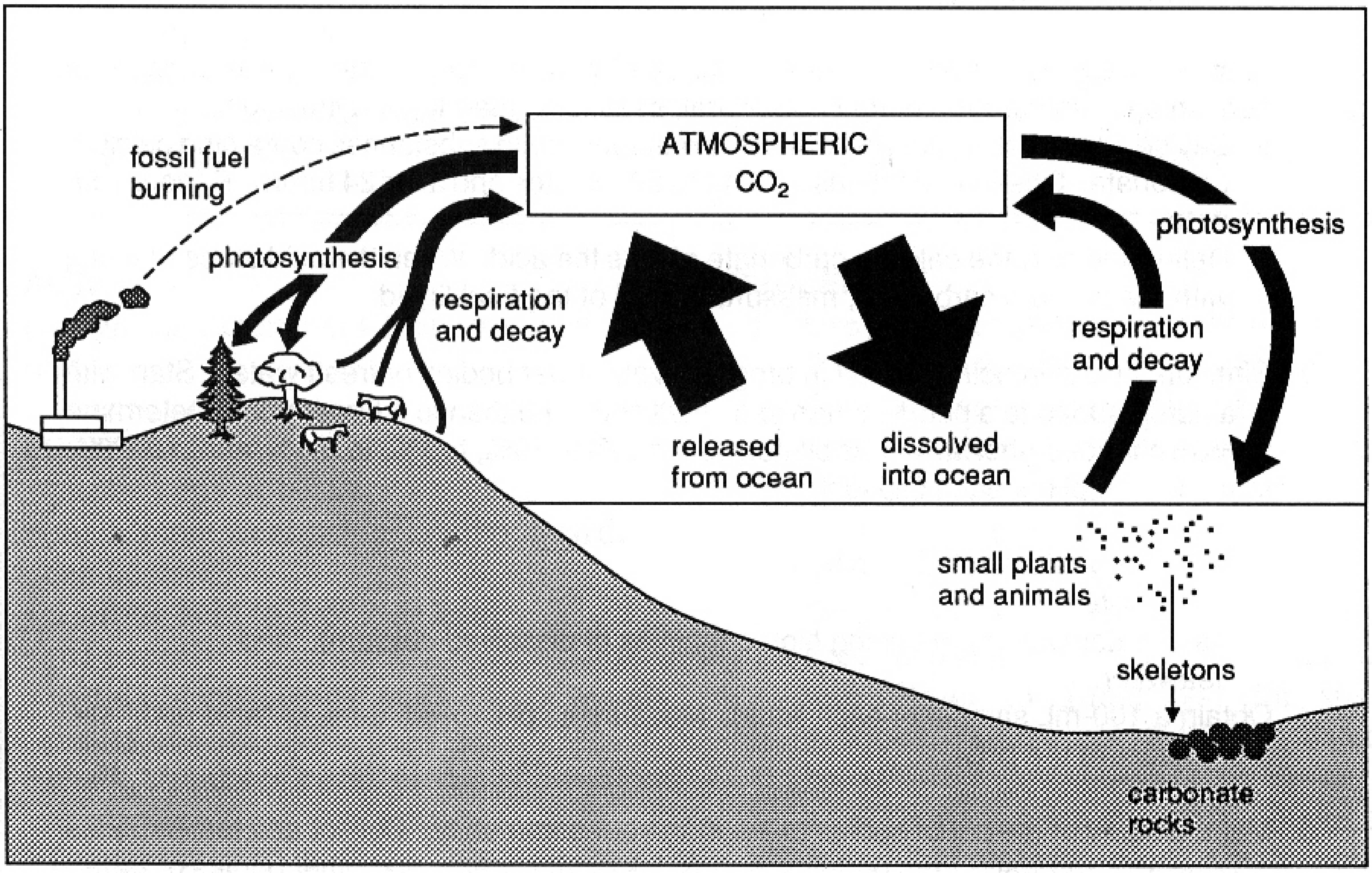

Fig. 6–8. Global CO_2 cycle

air. **Respiration** produces energy by combining carbohydrates from food with oxygen from the air according to this equation:

$$C_6H_{12}O_6 + 6\,O_2 \longrightarrow 6\,CO_2 + 6\,H_2O + \text{energy}$$

Carbohydrate + oxygen ⟶ carbon dioxide + water + energy

Processes that remove CO_2 from the atmosphere are also shown by solid arrows in Fig. 6–8. They include

- photosynthesis by green plants on land and by algae in the ocean.
- the dissolving of CO_2 molecules from the air into water.

Green plants on land such as grasses, shrubs, and trees use water from the soil and carbon dioxide from the air in the process of **photosynthesis** to manufacture new carbohydrates. In the ocean small plants called algae carry on the same process. They make more carbohydrate in photosynthesis than they consume in respiration. The excess carbohydrate produced by plants becomes the food consumed by animals. In the photosynthetic process plants convert the radiant energy of the sun into chemical bond energy to make the carbohydrates and produce oxygen as a byproduct. The oxygen is used by animals and plants to oxidize food and by industry to oxidize fuel. The equation for photosynthesis is

$$\text{Energy} + 6\,CO_2 + 6\,H_2O \longrightarrow C_6H_{12}O_6 + 6\,O_2$$

Energy + carbon dioxide + water ⟶ carbohydrate + oxygen

In addition, human activities, primarily the burning of **hydrocarbon fuels** (ones containing carbon, such as wood, coal, natural gas, gasoline, and oil) also add to the amount of CO_2 in the atmosphere. See the dashed arrow in Fig. 6–8.

Because all natural fuels—wood, oil, gasoline, coal—come from once-living things, they are made from carbon and hydrogen, which release carbon dioxide (CO_2) and water (H_2O) when they burn. For example, the gasoline component octane burns as follows:

$$2\,C_8H_{18} + 25\,O_2 \longrightarrow 16\,CO_2 + 18\,H_2O$$

Octane (gas) + oxygen ⟶ carbon dioxide + water

Recent studies of gases in the atmosphere show that the amount of CO_2 has increased steadily over the past hundred years. See Fig. 6–9. Analyses of gases trapped in ice show similar increases in CO_2 levels at both poles. See Fig. 6–10. Most of the rise comes from the tremendous increase in the burning of hydrocarbon fuels by utility companies, other industrial plants, motor vehicles, and airplanes.

To complicate matters, many plants that once removed CO_2 from the atmosphere to make food have been destroyed. As trees are cut for fuel, forests are cleared for agriculture, and vegetation is destroyed to build cities and towns, the amount of leaf surface acting to remove CO_2 from the atmosphere has dropped sharply.

Some scientists believe that seawater has reached the limit of its capacity to dissolve atmospheric CO_2. If they are right, CO_2 will continue to build up in the atmosphere.

QUESTIONS

16. What was the average atmospheric concentration of CO_2 in 1958? In 1985?

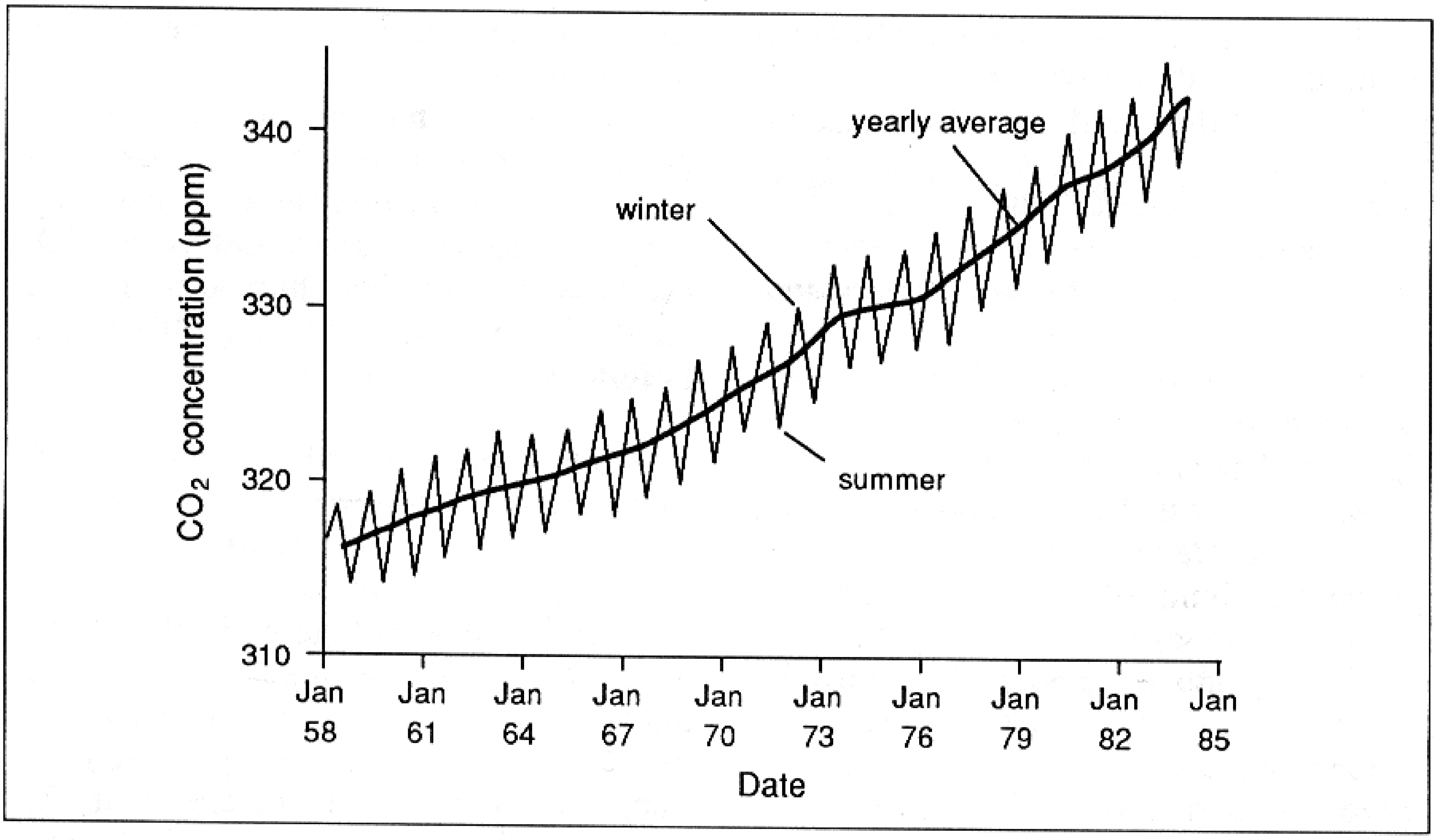

Fig. 6–9. Concentrations of atmospheric CO_2 measured at the Mauna Loa observatory in Hawaii

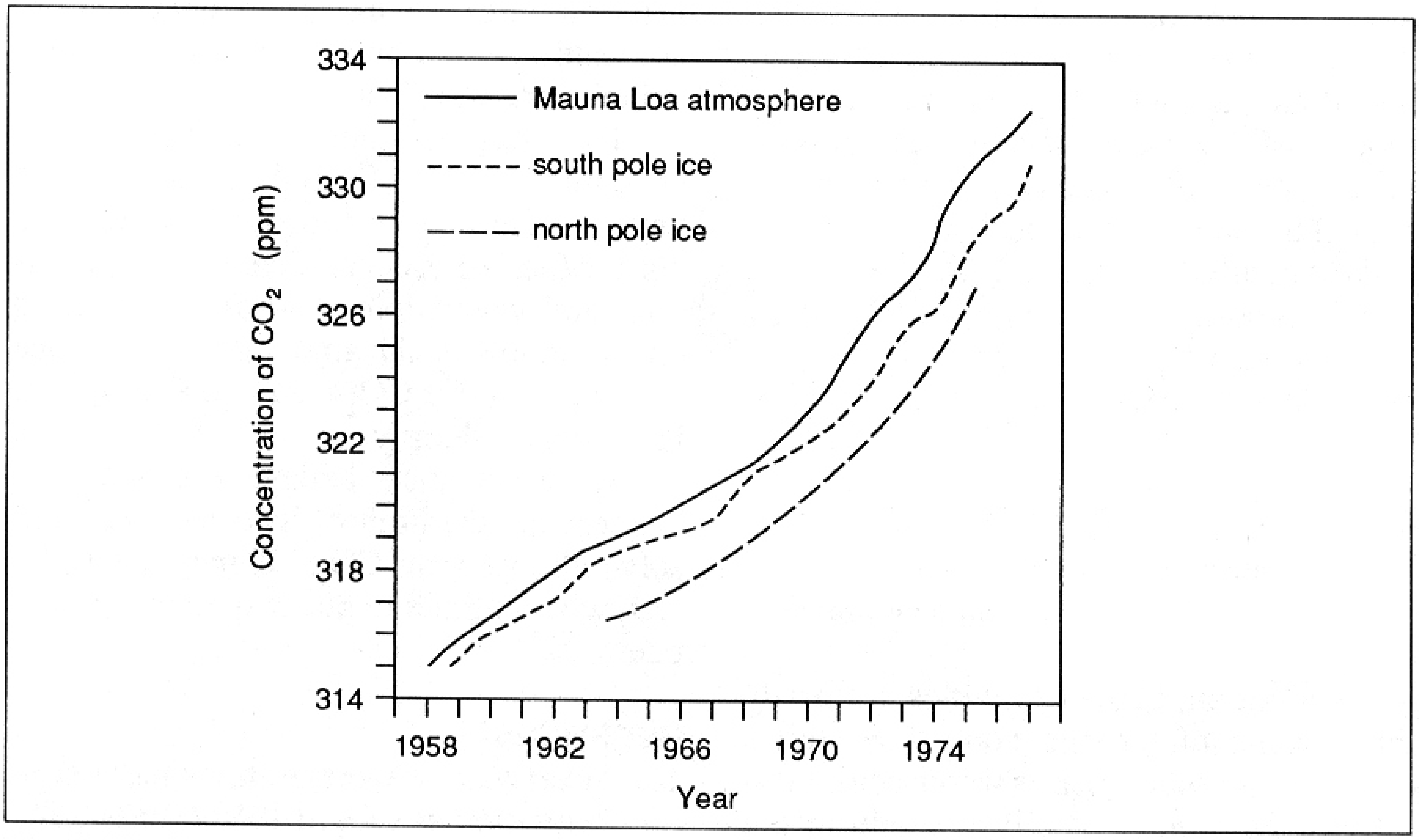

Fig. 6–10. Concentrations of CO_2 measured in gases trapped in ice

17. What is the average yearly increase in the atmospheric concentration of CO_2? Show your calculations.

18. What natural processes might account for the high concentration of CO_2 in the winter and its low concentration in the summer?

The Greenhouse Effect

When gases like CO_2 accumulate in the atmosphere, they have the effect of a greenhouse. They warm the earth the way glass holds in radiant energy, warming the interior of a greenhouse. By holding in the heat, the layer of CO_2 raises the temperature of the air, the land, and the water. The natural concentration of CO_2 in the atmosphere is enough to keep the earth at a habitable average temperature of 15°C. But human uses of hydrocarbon fuels could increase the concentration enough to cause an "enhanced," or magnified, greenhouse effect that could warm the earth enough to cause global problems. See Fig. 6–11.

QUESTIONS

19. What would be the average global temperature in °F
    a. if there were no greenhouse effect?
    b. if there were only the natural greenhouse effect?
    c. if the atmospheric carbon dioxide were to double?

20. In °F, how much warmer
    a. is the earth now because of the natural greenhouse gases?
    b. would the earth become if atmospheric carbon dioxide doubles?

21. Explain the difference between the greenhouse effect and the "enhanced" greenhouse effect.

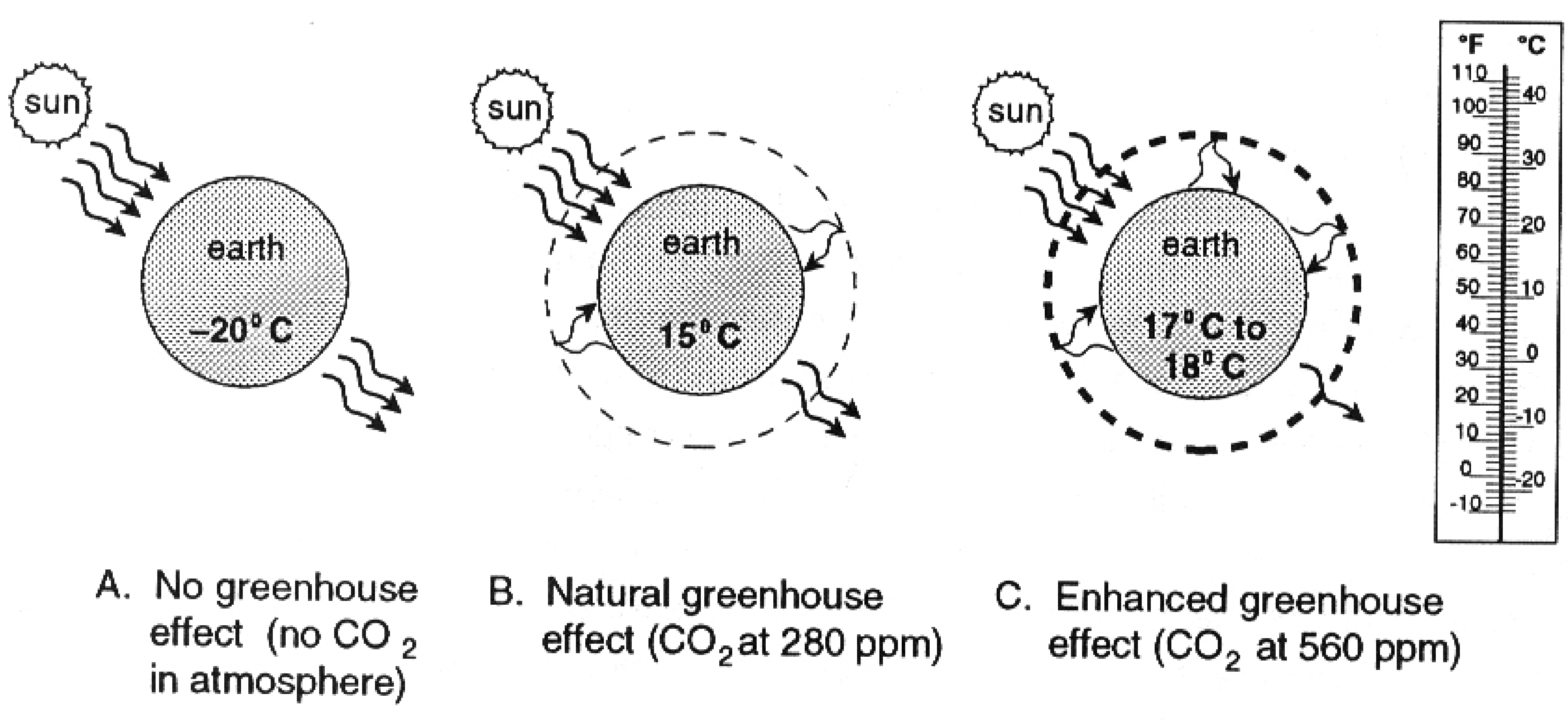

Fig. 6–11. CO_2 levels and average temperatures at the earth's surface

Other Greenhouse Gases

CO_2 is the chief culprit in the greenhouse effect, but other gases from industry and agriculture add to it. These include

- methane (CH_4), a gas produced by bacterial decomposition of organic plant and animal matter in such places as landfills, marshes, mudflats, flooded rice fields, sewage treatment plants, and the guts of cattle and termites.
- nitrogen oxides from the runoff of nitrogen-based fertilizers and as a by-product from the burning of hydrocarbons.
- ozone (O_3), a gas normally present in trace amounts in the atmosphere but also produced in industrial processes.
- chlorofluorocarbons (CFCs), synthetic gases used in cleaning solvents, refrigerants, and plastic foam.

Scientists have made computer models to predict the consequences of the greenhouse effect. They say an increase of 2°C to 3°C in the earth's temperature could cause one or more of the global changes listed in Table 6–5. Fig. 6–12 shows three predictions of rise in sea level.

Scientists warn that we do not yet have enough information to make firm predictions about the greenhouse effect. Global temperature records have been kept for only 100 years. The warmest years were in the latter part of the twentieth century. Furthermore, the average global temperature for landmasses has climbed more than 0.7°C since the 1880s. Whether we are seeing the beginning of a global warming or whether these are temporary changes in global temperature we do not yet know.

Table 6–5. Predicted global consequences of the enhanced greenhouse effect

- The volume of ocean surface water will increase.
- Sea levels will rise.
- Low-lying islands, coasts, estuaries, marshes, and deltas (including farmlands and cities) will be flooded.
- Rich, productive coastal ecosystems (including bayous, salt marshes, and mangrove swamps) will be destroyed.
- Domestic and agricultural water systems will be disrupted.
- The snowpack that supplies freshwater reservoirs will be reduced.
- Global patterns of weather and precipitation will change.
- Temperature zones in the northern hemisphere will move north.
- Agricultural areas will move north with changes in climate.
- The number of days with temperatures over 37°C will increase.
- Ocean water will evaporate faster, producing more clouds, rain, and snow in some places.
- Some deserts will get heavy rain; some fertile areas will dry up.
- Storms will be more frequent, and hurricanes will be more severe.
- Some plants and animals will become extinct.
- Photosynthesis and plant growth will accelerate.

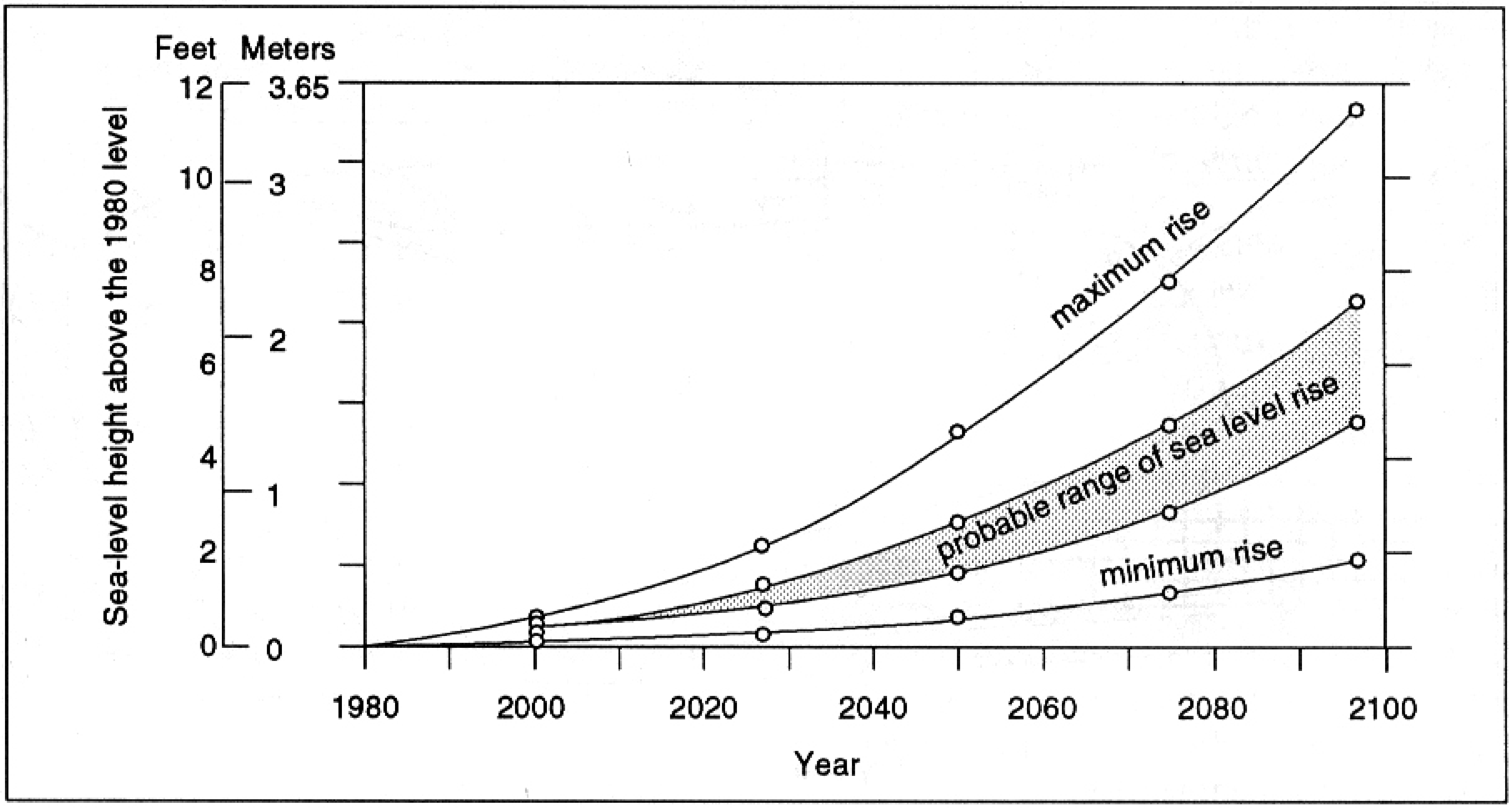

Fig. 6–12. Predictions of rises in sea level because of an enhanced greenhouse effect

Some scientists studying the data on the burning of hydrocarbons make a reverse prediction. They think that pollutants in the atmosphere could block sunlight, just as particles from forest fires and volcanic eruptions do. If that happens, the blockage could offset the warming effect of the increase in carbon dioxide.

ACTIVITY 3

Analyze a coastal map to predict shoreline changes due to the greenhouse effect.

MATERIALS

- colored pencils, pens, or crayons

PROCEDURE

1. Examine the map in Fig. 6–13 showing two imaginary coastal towns, "Seaside City" and Baytown"

2. Using colored pencils,
   a. find and draw in sea level.
   b. shade the river that carries water to the towns.
   c. find and color the major highways into and out of the towns.
   d. find the airport and color-code it.

3. Imagine that the greenhouse effect causes sea level to rise. See Fig. 6–12.
   a. Shade the area(s) that will be flooded if the most probable (average) sea-level rise occurs.
   b. Shade the area(s) that will be flooded if the maximum predicted sea-level rise occurs.
   c. Shade the contour lines for sea-level rises of 1 m, 2 m, and 4 m.

4. Attach a key to the map.

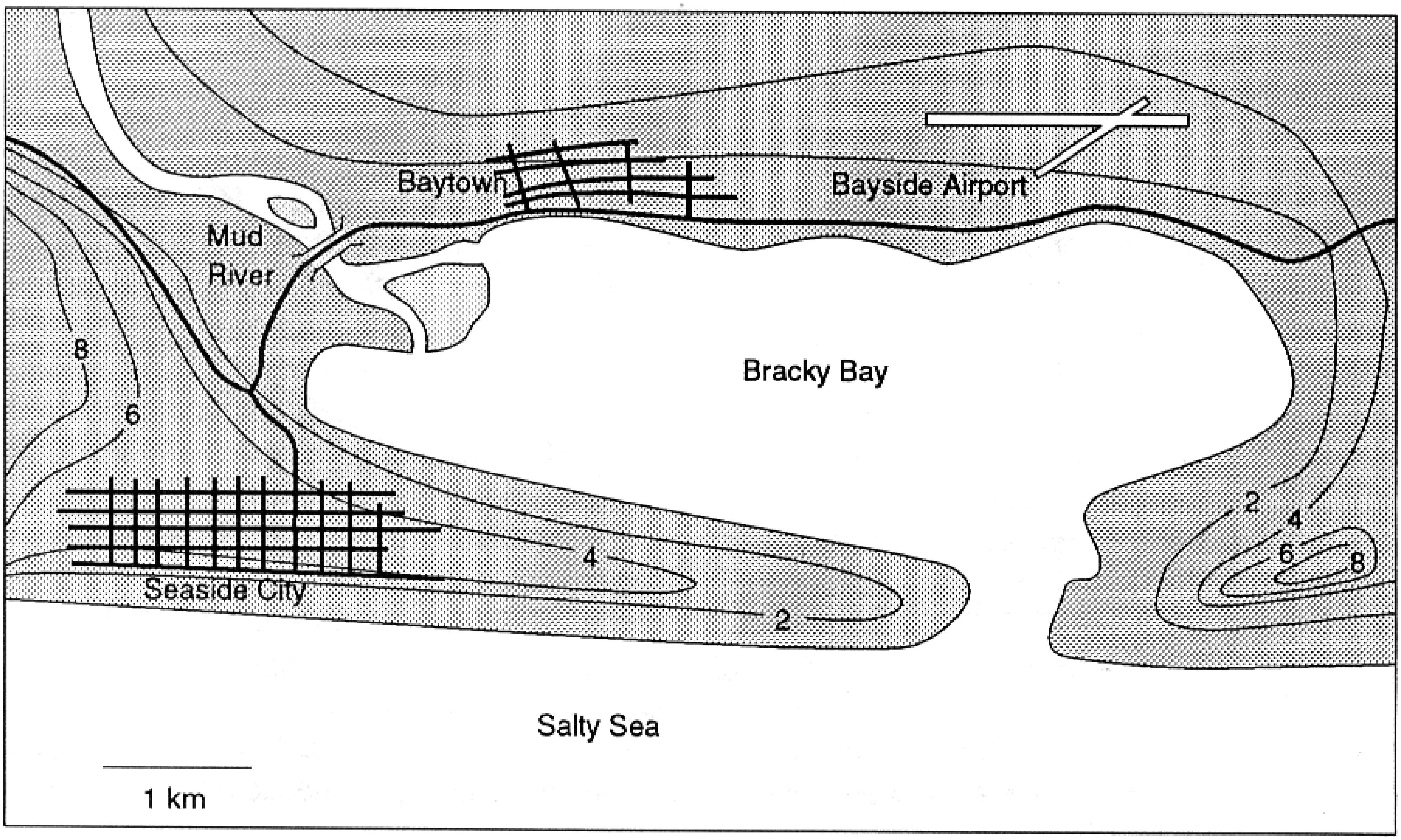

Fig. 6–13. Map of imaginary coastal towns (elevation in meters)

QUESTIONS

22. How far inland does the coastline move with
    a. a 1-m rise in sea level?
    b. a 2-m rise in seal level?
    c. a 4-m rise in sea level?

23. How would the most probable rise in sea level affect
    a. homes and business buildings in the towns?
    b. coastal roads and airports?
    c. the sources of fresh drinking water for coastal areas?
    d. coastal wildlife?

24. How would the maximum sea-level rise affect these areas?

What Can Be Done?

Many suggestions have been made for reducing the CO_2 content of the atmosphere:

- Reduce CO_2 emissions by reducing the consumption of fossil fuels.
- Switch from burning oil or coal to burning fuels (such as hydrogen) that produce less CO_2 or none.
- Increase the use of solar, wind, and geothermal energy.
- Replant and restore forests.

QUESTIONS

25. List the things that you, your family, and the people in your community could do to reduce the buildup of CO_2.

26. What are businesses, industries, and

governments in your region doing to reduce the causes of the greenhouse effect?

27. If the sea level begins to rise, what might be done to minimize
    a. damage to coastal structures?
    b. disruption of coastal wildlife?

FURTHER INVESTIGATIONS

1. Study the pH of water in your community. Include water in your home, in swimming pools, in puddles, and in lakes or streams. What evidence, if any, do you find of acid rain? How might you account for your findings?

2. Read library references on acid rain. Learn more about ways to reduce pollutants that produce acid rain. Report to the class.

3. Read library references on the greenhouse effect.
    a. Look for computer predictions of changes in precipitation, sea level, and regional, climatic, and atmospheric temperature.
    b. Read about international conferences on the greenhouse effect. Report actions that have been called for to control CO_2 emissions.

4. Ozone (O_3) is a greenhouse gas, but scientists are also concerned about its depletion in the upper atmosphere, exposing life on earth to additional harmful rays from the sun. Find out more about this and report to the class.

5. Read and report on which government agencies are trying to solve environmental problems such as acid rain and the greenhouse effect. Who monitors the environment? Where do the data go? How are decisions made, and by whom?

6. Learn more about computer models that predict environmental change. What information goes into the computer? How reliable are computer-generated maps? What are their limitations?