MINILAB [15 pts]
Intro to Acids & Bases

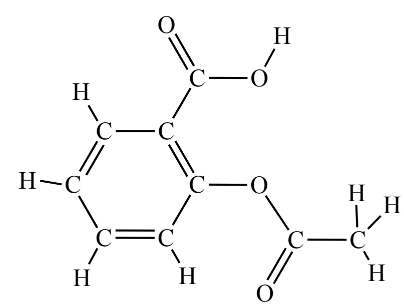
NAME ____________________________
Period ____ **Date** ___________________

Part A: pH of Familiar Acids and Bases

1) Test the pH of the following familiar acids and bases by using the pH and litmus paper provided. Please use only as much pH and litmus paper as you need—I do not have an unlimited supply. Just put a dot of each liquid on a strip of the paper, putting as many liquids on a strip as possible. For the pH paper, compare the color with the "color key" provided on the vial. Fill in pH, red litmus paper, blue litmus paper and A, B, N? (acid, base or neutral) columns.
2) Fill in the missing chemical names at any time. I have given you the difficult ones, but you should be able to name the ionic ones yourself!!

Substances [4 pts]	Chemical formula	Chemical Name	pH	Red Lit.	Blue Lit.	A, B, N?
1) Vinegar	CH_3COOH					
2) Ammonia	NH_3	ammonia				
3) Drano™	NaOH					
4) Dish soap	$NaC_{18}H_{35}O_2$	sodium stearate				
5) Milk of Magnesia™	$Mg(OH)_2$					
6) Lemon juice	$H_3C_6H_5O_7$	citric acid				
7) Baking soda (aq)	$NaHCO_3$					
8) Tums™ antacid (aq)	$CaCO_3$					
9) Regular aspirin (aq)	$HC_9H_7O_4$	acetylsalicylic acid				
10) Buffered aspirin (aq)	$HC_9H_7O_4$ + ?	xxxxx				

1) [1 pt] Aspirin is slightly acidic. Thus, there must be a hydrogen in its structure which is easy to lose when dissolved in water. **Only hydrogen atoms in polar bonds can be lost.** In the structure at right, put in all significant partial charges (δ^+, δ^-). Circle the hydrogen that is lost when dissolved in water.
2) [1 pt] Why is buffered aspirin "easier on your stomach" than regular aspirin. (Compare their pH values.)



Part B: Carbonated water (demo)

1) [1 pt] I will put some water in an Erlenmeyer flask add 2 drops of universal indicator. **initial color/pH:** ______ Obtain a pH meter and check its pH. **initial pH using with pH meter =** ______
2) [1 pt] Blow air into the flask until a color change is observed. **final color observed** ______ Check pH with pH meter. **What is your final pH with meter?** _____
3) [1 pt] Did the pH get **more acidic or more basic** when air was blown into the water?_______. Based on this demo, do your think carbonated drinks are slightly **acidic** or **basic**?
4) [1 pt] Carbonated water is made by bubbling carbon dioxide gas into water (like blowing air in). This causes some carbon dioxide gas to dissolve into the water. Some of the carbon dioxide actually reacts with the water to form carbonic acid, H_2CO_3. This reaction is shown here: CO_2 (g) + H_2O (l) → H_2CO_3 (aq). Since carbonic acid is an acid, it donates an H^+ to water. Write the products formed when it reacts with water. (only ONE H^+ is transferred.).

$$\mathbf{H_2CO_3\ (aq) + H_2O\ (\mathit{l}) \rightleftharpoons}$$

Part C: Questions about article, "An Invisible Fire"

1) [1 pt] What was the name of the acid in the rust remover? What is the chemical formula of the acid? (Hint: it sounds like hydrochloric acid, except with a different element in place of Cl)

2) [2 pts] This acid is considered relatively weak (compared to H_2SO_4 or HCl), yet it is so VERY dangerous. Explain why this acid is so dangerous by answering these questions.
   a) Why is it easy for this acid to pass through the skin?

   b) What happens to the acid when it reaches the living cells in the epidermis and dermis?

   c) What damage does the free hydrogen ion do?

   d) What damage does the free fluoride ion do to one's body?

3) [1 pt] What was the purpose of the calcium gluconate in the gel that doctors applied to William's skin?

4) [1pt] What other two procedures did the doctor use (after the gel) to try to save his arm?

FROM THE APRIL 1996 ISSUE

An Invisible Fire

By Jeremy Brown | Monday, April 01, 1996

During a scene in the movie Alien, crew members are startled to see a fist-size hole in their spaceship's ceiling, still sizzling from some substance that has just burned clear through the metal. That metal-eating stuff is, of course, the blood from the alien. I have met the closest thing there is to that alien's blood. It came in a small plastic bottle, and it was eating its way through my patient's hand.

On an otherwise ordinary evening, William Turner, a 37-year-old truck driver, noticed a paint stain on his coat. Looking for something to remove the stain, he wandered into the basement of his rented house and rummaged around. At the back of a dusty shelf stood a small bottle labeled Industrial Laundry Rust

Remover. The side of the bottle carried the warning CAUTION: DO NOT USE WITHOUT GLOVES. William didn't read that bit, however, and he removed the cap from the bottle, spilled some liquid onto a rag, and began rubbing it into his stained coat. But as the stain began to fade, his right hand, the hand he was using to apply the liquid, began to hurt. After 20 minutes the pain was so intense that he had to stop. Within 40 minutes he could no longer move his fingers. Frightened and in terrible pain, he managed to drive himself to our emergency room. Fortunately, he brought the rust remover with him.

While William writhed in agony, I took a close look at his right hand. Except for some mild swelling of the fingertips, it looked just like his left hand. But if I even gently pressed on his fingernails, he grimaced and begged for a painkiller.

When we burn our skin by touching something extremely hot, it is the high temperature that kills the cells. Chemical burns are different: cells are killed by a chemical reaction on the surface of our skin. Probably the most common type of severe chemical burn comes from the sulfuric acid in a car battery. But sulfuric is not the acid to be most feared. That distinction belongs to hydrofluoric acid, a compound commonly used in solvents and rust removers, and so powerful that it can be used to etch images on glass. Although the burn it produces initially causes no blisters or changes in skin color, it can leave behind a scarred limb.

Hydrofluoric acid can severely damage the deep tissues of the body yet leave little trace of damage on the skin surface. It can even kill. People have died after a patch of skin no bigger than the sole of the foot was exposed to the substance.

Counterintuitively, perhaps, what makes hydrofluoric acid so deadly is not that it is a strong acid. In fact, compared with hydrochloric acid or sulfuric acid, it is actually weak. Acids are formed when charged hydrogen atoms bind with nonmetal atoms, and they are judged strong or weak depending on the tenacity of that bond. The weaker the acid, the less easily the bond is broken.

A charged atom, or ion, is an atom that has gained or lost one or more electrons, and it is the attraction between a hydrogen ion and a fluoride ion that creates the chemical bond in hydrofluoric acid. The bond is relatively stable because the fluoride ion--which can hold an electron more strongly than any other ion can--wants the electron that hydrogen has to offer. Yet under the right conditions, that stable bond can be broken. Because fluoride is so electron-hungry, it will latch onto chemicals that can provide electrons. And unfortunately, the tissues of the body are swarming with chemical partners that are far more attractive to fluoride than hydrogen.

To do its deadly work, hydrofluoric acid must first pass through the skin. This is easily done because hydrofluoric acid doesn't carry a charge, and uncharged molecules can easily slip through the fatty surfaces of membranes. Hand in hand, as it were, the hydrogen ion and the fluoride ion pass down through the stratum corneum, a tough, waterproof layer of dead cells. Eventually they reach living cells in the epidermis and dermis, where they meet up with a slew of new chemical partners. And that's when the terrible damage begins. The electron-hungry fluoride ion breaks free of the hydrogen ion and binds to calcium or magnesium, two electron-rich minerals. The hydrogen ion, now free of its fluoride partner, binds to enzymes that neutralize acids and keep the pH in our blood and tissue stable. This devastating disruption of the normal chemical balance--both inside and outside cells--kills cells beneath the surface of the skin.

The worst damage occurs when fluoride grabs onto calcium and magnesium, minerals crucial to a host of electrochemical reactions. Without enough free calcium and magnesium, nerves fail and cell membranes collapse. The degree of damage depends on just how low the levels fall. A mild decline can cause numbness, cramps, or horrendous pain. A more severe decline can cause extreme muscle spasms, convulsions, an irregular heartbeat, and even death.

When William first arrived in the emergency room, the doctors flooded his injured hand with water to wash away the hydrofluoric acid. Then they applied a gel--a mixture of calcium gluconate powder and surgical jelly. The reasoning was simple. If the tissue damage occurs because hydrofluoric acid causes dangerously low levels of calcium, then providing extra calcium to the affected tissue should prevent the damage. But William was still in excruciating pain. That meant the hydrofluoric acid was still at work deep inside the tissue. Something more had to be done. Without treatment, William could lose his hand.

Hydrofluoric acid burns are rare enough to require the advice of an expert. When I called our local poison control center, the toxicologist advised that we boost William's calcium levels by injecting calcium into his bloodstream.

There were two ways we could do this. The first was to inject calcium into a vein. Technically, this is very easy to do. The problem is that veins carry blood away from tissues, so any calcium we gave would quickly be washed away from where it was most needed. The second way was to inject calcium directly into William's radial artery, the artery in the arm that carries blood to the wrist and hand. Since arteries carry blood into the tissues, calcium injected into an artery would be most likely to reach the affected area. The problem is that arteries are much more sensitive to injections than veins. An injection of calcium into an artery could cause it to spasm, cutting off the blood supply. That was the last thing William's already damaged hand needed.

I decided to inject calcium into his vein, but I would use a variant of a technique that the toxicologist told me might decrease the blood flow out of William's arm and thus increase the delivery of calcium to the injured tissue. The technique, long used for local anesthesia, is known as a Bier's block. After placing a small needle into one of the veins in his right arm, I raised William's arm up above his head and wrapped an Ace bandage around his fingers and up his arm until I reached his elbow. The pressure of the bandage squeezed blood out of his arm. I then inflated a blood pressure cuff to a pressure high enough to prevent any more blood from leaving his arm. William now had, in effect, a right arm that was almost completely detached from the rest of his circulation. Any calcium we injected into this arm would stand a good chance of staying where it was and raising the calcium levels back to normal.

The Bier's block went according to plan. As I waited to see if the treatment eased William's pain, I thought about what he would face if he lost the use of his right hand. William was the sole breadwinner for his large family and ailing parents. Saving William's hand might also save his family.

But after nearly an hour, William felt no better. His arm looked just about the same as when I first examined it--almost normal, though very pale. But the danger of hydrofluoric acid is that the burn it causes just doesn't look like a burn at all. It was time to use the last weapon we had.

I asked the nurse who was caring for William to infuse calcium directly into his radial artery. She knew this was an unusual procedure. Although patients in the intensive care unit often have a catheter placed into the radial artery, the catheter is typically used to measure the pressure within the artery. It is rarely used to give drugs. Infusing calcium through the artery made the nurse uneasy. She worried that it would constrict, cutting off the blood supply to his hand and fingers. After I explained why we needed to do the procedure, she shrugged and made sure that I signed the patient's medical chart. That way, at least, if any problem occurred, it would be clear that she was just following the doctor's written request.

A half hour later, I returned to William's bedside. He was smiling for the first time. The throbbing sensation in his arm had eased, and although his hand remained tender, he said the pain was subsiding. Within two hours he was totally free of pain, and to prove it, he extended his right hand and gently shook my hand.

William remained under close observation until the next morning. His pain had nearly vanished, and I discontinued his calcium therapy. I checked on him the following day, just before he was discharged. He was still very sensitive to pressure on his fingertips, but there was almost no trace of the burn. Our treatment had worked. William had uncorked a flesh- eating chemical monster in a dusty bottle of rust remover but would survive unscathed.
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