Pick up the **e⁻**'s.	However, one cannot have a full reaction with only a reduction couple or its reverse, an oxidation couple. Similarly, you cannot have an acid–base reaction with only an acid and its conjugate base or a base and its conjugate acid. Each of these is only half of the reaction and will not occur without a complementary half. Cleverly enough, in redox chemistry, each of these is called a *half-reaction*.
Move **e⁻**'s away and toward you. Move **e⁻**'s toward and away from the student.	Two complementary half-reactions make a full reaction. The copper oxidation half-reaction and the iron reduction half-reaction need to be complementary.
Moving to the end of the row away from the student with the iron, pass the **e⁻**'s down the row.	If we separate the two half-reactions and allow the electrons to move along an external circuit we have an electrochemical cell. One problem with all of this is that we are not insisting upon electrical neutrality. Two electrons are moving from the copper to the iron and no corresponding charge is moving back. There is nothing nature abhors more than an unbalanced charge.
Give **Cl⁻**'s to two students and stand them near you.	In fact, one never has just $Cu(II)$ in solution; you have $Cu(II)$ with charge balancing counterions. Today, we are working with chloride ions as our counterions, so to balance the charge we need two Cl^- for the one $Fe(II)$. So, as the electrons move to me, the chlorides move to the $Fe(II)$.
Give two of them each a **Cl⁻** and the other two each a **K⁺**.	These chlorides do not necessarily have to move. We need ions to balance the chloride charge in the solution around the Cu and the $Fe(II)$ charge in its solution. We now need four people up in the center to make a salt bridge.
Have the two salt bridge **Cl⁻**'s move to the **Cu(II)** side.	Now let us move the two electrons from me, the $Cu(II)$, down the row to the $Fe(II)$. At my end, I now have an excess positive charge. What do I need to balance this charge: obviously, negative charges. I can get these from the chlorides in the salt bridge. At the other end, with the new electrons, we have an excess of negative charge. What is attracted by negative charges? Positive charges! What do we have that has positive charges? Potassium ions!
Have the two salt bridge **K⁺** move to the **Fe(II)** side.	Notice how the positive ions followed the negative charged electrons and the negative ions moved toward the electrons used to be. The charges (ionic plus electronic) then go full circle. In the language of electricity this is called a circuit. If the circle is not closed, you have a broken circuit and somewhere along the line, charge builds up without being compensated. As I said before, nature abhors this and does not let it happen for long and the current stops.
	Oxidation occurs at the electrode called the anode. Where there is oxidation (where electrons are lost) there will be a deficiency of negative charge. Negatively charged ions are attracted to the anode to replace the lost charge; they are called anions.
	Reduction occurs at the electrode called the cathode. Where there is reduction (where electrons are gained) there always will be a surplus of positive charge. This will attract positively charged ions; they are called cations.

Linking acid–base and redox reactions by the realization that they are transfer of protons and electrons, respec-tively, is helpful for the student to gain a rational foothold on the idea of chemical reactions as transformations.

Conclusion

The elegance of these kinetic classroom demonstrations is the simplicity of both the basic transformations and the materials necessary. Yet with all of its simplicity, one can develop a large number of interconnected concepts, as well as ignite student interest. Concepts involving transformation are inherently difficult for introductory students. These concrete examples in the classroom with instructor and student movement help with difficult concepts. They do so by developing the meaning before assigning definitions. After the meaning is developed, the terms are defined and the students are introduced to the language of transformation chemistry.

Acknowledgment

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Pictorial Analogies XI: Concentrations and Acidity of Solutions

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Students frequently confuse the concentration of a reagent in solution and the strength of the reagent itself. The concentration simply has to do with the number of molecules (moles) present per unit volume. This is analogous to the number of soldiers in an army where one army is considered equivalent to a unit of volume. See Figure 1. Note that the picture is meant simply to illustrate the different number of acid molecules (soldiers) in identical volumes (armies) and the regular spacing and orientations of the soldiers should not infer that particles in liquids have such regularities.

The strength of an acid in solution has to do with the relative ability of an acid to donate protons, where an acid is considered strong only if it is such a good proton donor that each and every acid molecule will give up at least one hydrogen ion to a water molecule (100% acid dissociation), which results in the leveling effect. Perhaps it is less confusing if reagent strength is first related to the strength of an oxidizing agent, with the relative measure of strength given in the half-cell potentials, which are more clearly distinguishable from concentrations. In either case, the strength of a reagent in solution can be distinguished from concentration by relating that strength to the muscle power of individual soldiers in an army as illustrated in Figure 2. One can have a more dilute solution of a stronger reagent, just as one can have a smaller army, but one that has a larger percentage of strong soldiers. Note that Figure 2 also can be used to illustrate that a more dilute solution

Concentration of Solution

Number of Soldiers in an Army

more concentrated | less concentrated
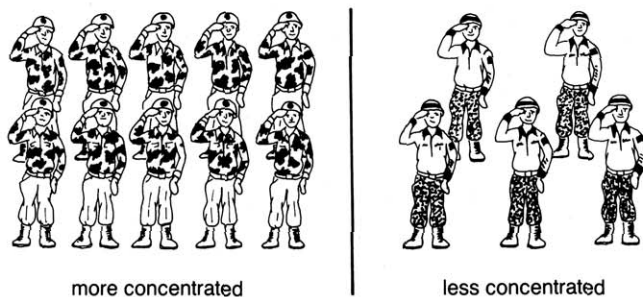
100 vs. 50

Figure 1. Analogy of solution strength to armies.

Strength of Solution

Percent of Soldiers who can lift 300 lbs.

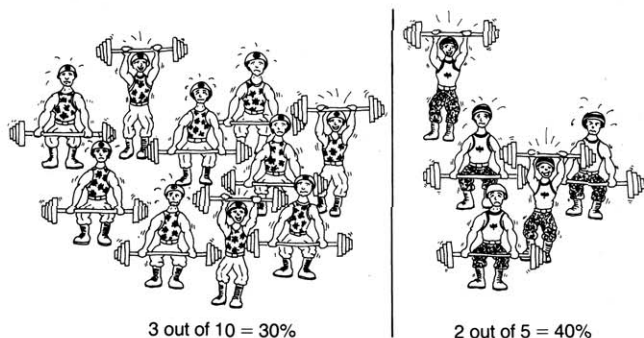3 out of 10 = 30% | 2 out of 5 = 40%

Figure 2. Analogy of acid strength to percent of strong soldiers.

Super Saturated Solution

Too many in theater for the number of seats, | but you don't realize it until all seats are filled
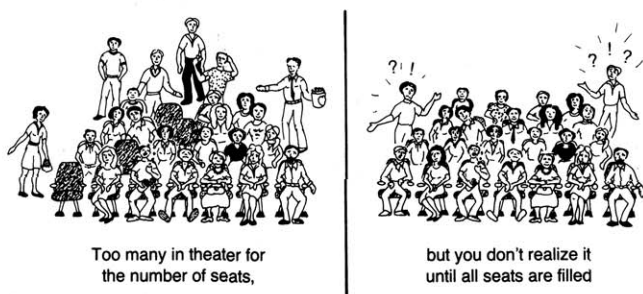
Figure 3. Analogy to a supersaturated solution.

of the same weak acid has a higher percent dissociation than a more concentrated solution of that acid.

Other published analogies to acid strengths and the leveling effect relate them to a father and his son (1), bricklayers (2), and spending money (3).

Another confusing concept concerning solutions is the supersaturated solution. How can the unstable condition
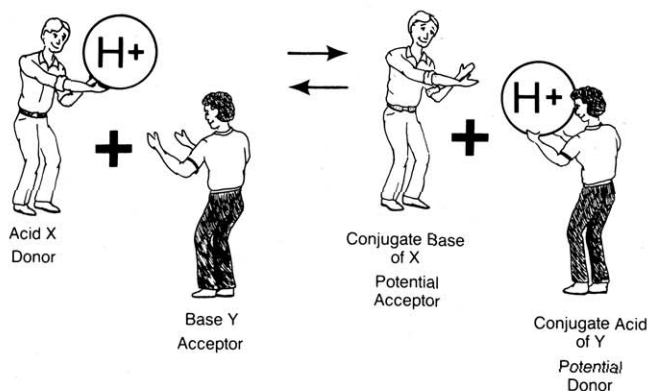
Conjugate Acid - Base Pairs

Acid X Donor

Base Y Acceptor

Conjugate Base of X Potential Acceptor

Conjugate Acid of Y Potential Donor

Figure 4. Analogy to conjugate acid–base character.
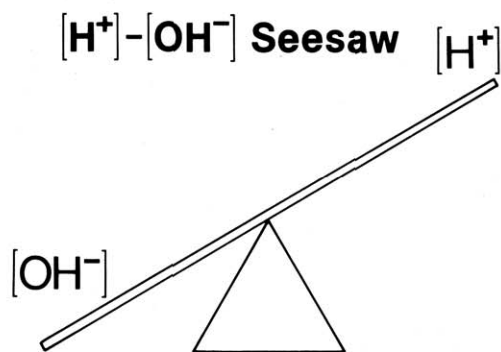
$[H^+]$ - $[OH^-]$ Seesaw

Figure 5. Analogy to the hydrogen ion concentration to hydroxide ion concentration relationship.
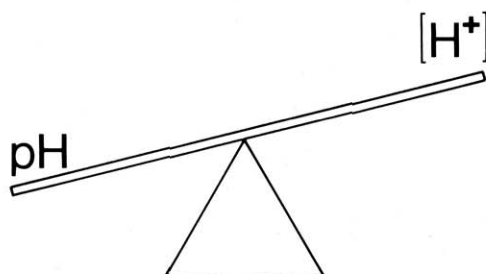
pH - $[H^+]$ Seesaw

Figure 6. Analogy to the inverse relationship between hydrogen ion concentration and pH.

exist in which an amount of solvent contains more solute than it should be able to contain at that temperature? Figure 3 relates the supersaturated condition to a theater audience who doesn't realize that there are too many people for the number of seats until everyone has to find a seat.

Conjugate acid–base pairs also give students difficulties. Figure 4 illustrates a good analogy 4, 5 in terms of two people passing a beachball back and forth. The ball represents the acidic proton. Something that gives it up has the potential to receive it back. Furthermore, a strong person

who passes the proton forcefully to a weaker person is less likely to accept it back. Finally, the reason that a person (base) can accept the ball (proton) is that he has a pair of hands (electrons), that ties the Brønsted–Lowry concept to the Lewis definitions nicely.

Most students can follow the concept that when the hydrogen ion concentration goes up the hydroxide ion concentration goes down, just like a seesaw or teetertotter. See Figure 5. Something harder for the beginner to remember is that, due to the negative logarithmic definition of pH, when the hydrogen ion concentration goes up the pH goes down. This is presented in Figure 6.

Large photocopies of these illustrations will be sent on request for those wishing to make overhead transparencies.

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