The University of Texas at Austin
Department of Chemistry
College of Natural Sciences

Reaction of Alkali Metals & Water

Performer's Version

Safety Hazards

- Personal Protective Equipment
    - Nitrile gloves
    - Safety glasses/goggles
    - Chemical & flame retardant lab coat
    - Face shield recommended
- Safety Equipment
    - Fume hood
    - Blast shield
    - Fire extinguisher
    - Spill kit
- Physical Hazards
    - In contact with water, can release flammable gas that may spontaneously ignite and result in severe burns and eye damage
    - Can cause respiratory irritation
- Chemical Hazards
    - Flammable solid, may cause fire
    - In contact with water, produces metal hydroxides which may cause skin and eye corrosion and are harmful if swallowed
    - Phenolphthalein is a suspected mutagen and carcinogen, and may damage fertility or target organs through prolonged or repeated exposure

Safety Data Sheet(s)

- Sodium metal
- Potassium metal
- Sodium hydroxide (byproduct)
- Potassium hydroxide (byproduct)
- Phenolphthalein

Materials

- Sodium metal (approximately 0.5 – 0.75 g)
- Potassium metal (approximately 0.5 g)
- 2 large polycarbonate reaction vessels
- Hexanes in a small squirt bottle
- Phenolphthalein indicator solution
- Deionized water
- Phenolphthalein indicator solution
- Metal forceps
- 2 plastic weigh boats OR two watch glasses
- Paper towels

Procedure

WARNING: This demonstration must be done in a fume hood. If a portable fume hood is not available, the demonstration must be done via webcam or video from a laboratory equipped with one. If the performer is someone other than the Demonstrations Coordinator, please schedule an appointment for a practice demonstration. The Demonstrations Coordinator will be present for the demonstration in lecture.

1. Using the metal forceps, drop the provided piece of sodium into the water, lower the fume hood sash completely, and back away quickly.
2. The metal will begin to react violently – releasing hydrogen gas. As metal hydroxides are produced, the water will begin to turn pink, indicating the presence of a strong base. The hydrogen gas will likely ignite and/or pop/explode.
3. Repeat these steps for potassium metal.

The University of Texas at Austin
Department of Chemistry
College of Natural Sciences

<u>Pedagogy & Supplemental Information</u>

The first group of the Periodic table are known as *alkali metals*. Alkali metals are familiar and commonplace to many, but primarily in their ionic forms. Sodium and potassium ions (Na^+ and K^+) are not only present in the salt we eat, but also drive the electrophysiology of our neurons. When they're in their solid, neutral metal state, however, they are violently reactive with water.

The reaction of a pure alkali metal, like sodium or potassium, with water liberates hydrogen gas and produces metal hydroxides; metal hydroxides from Group 1 elements are considered strong bases. Bases are substances that make a solution become *alkaline* by reacting in a manner that increases the concentration of hydroxide ions in an aqueous solution as compared to hydronium ions ($[OH^-] > [H_3O^+]$). This shift in the ratio between hydroxide and hydronium results in a higher pH (or, conversely, a lower pOH). The production of these free hydroxides in solution is what earned these metals the moniker "*alkali*". The reaction of a Group 1 metal, M, with water is shown below:

$$2M\ (s) + 2H_2O\ (\ell) \rightarrow 2MOH\ (aq) + H_2\ (g)$$

$$2\ M\ (s) + 2H_2O\ (\ell) \rightarrow 2M^+\ (aq) + 2OH^-\ (aq) + H_2\ (g)$$

The liberation of hydrogen gas is where the spectacle of these reactions comes in. Hydrogen gas is highly flammable, and the reaction of alkali metals with water is also *very* exothermic – meaning that a large quantity of heat energy is released by the reaction. This results in the potential for the hydrogen gas to spontaneously ignite in an array of vibrant colors unique to each metal.

These colors serve as an exciting and explosive display of atomic emissions. Within an atom, orbitals are the regions in three-dimensional space in which there is some non-zero probability of finding an electron. Orbitals are arranged and divided into different quantized energy levels and even different spatial distributions (giving these probable regions 'shapes'), and this arrangement is unique from one element to the next.

Electrons, like most matter, will occupy the lowest energy state possible at rest – this is the configuration that offers the most stability, and is known as the *ground state*. In the context of atomic orbitals, this means that an electron will tend to be housed in the lowest energy orbital available to it. When energy is absorbed by that electron, however, it can jump up to a higher energy orbital; this energy jump is known as *excitation*. An excited electron won't stay excited forever, though – eventually, the electron will want to return to the ground state. In order to drop back down, the electron needs to release the energy it absorbed to become excited in the first place; this release, which is characterized by the magnitude of the energy gap between the excited state and ground state, is called *emission*. Electromagnetic radiation of a wide variety of energies can be emitted through this process – infrared waves, microwaves, x-rays, ultraviolet rays – depending on how much energy is absorbed by the given electron. What's most exciting for the classroom is when this emission falls within the visible light spectrum, energies of light directly relating to wavelengths we perceive as colors.