# Why Copper Is Reddish in Color

Comparing Copper, Silver, and Gold

## Introduction

Did you ever wonder why most common metals are silver in color except copper and gold? Why does copper have a reddish color and gold have a yellowish color? Is your curiosity piqued? Read on!

## Concepts

- Reflection/absorption spectra
- Electron configurations of metals

## Discussion

Most metals appear silver in color. Silver metals (such as silver or aluminum) do not absorb much or any of the visible light that hits it, the incident light. Most of the light is reflected and this causes very little distortion in the color of the reflected light. Notice the spectrum of silver metal and how the incident light is almost 100% reflected, as apparent by the almost complete overlap of the two curves in the light intensity spectrum (see Figure 1).

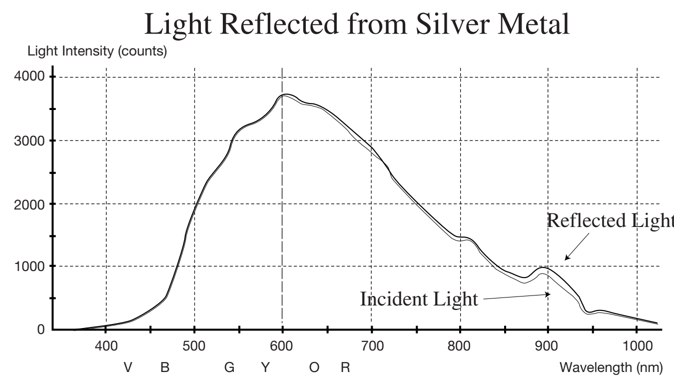


**Figure 1.**

Copper, on the other hand, has a distinctive red-orange color. The first and most obvious reason that any object is colored is that the object absorbs some wavelengths of light and reflects other wavelengths of light. Looking at the light intensity spectrum of copper, when light is shined upon copper metal, the copper atoms absorb some of the light in the blue-green region of the spectrum (see Figure 2). When an object absorbs one color of light, its complementary color (see the color wheel on page 2) is reflected back to our eyes. Since blue-green light is absorbed, its complementary color, red-orange, is reflected. Hence copper appears a red-orange color. Like copper, gold also lacks the typical "silvery" color of most metals. Instead it has a distinctive yellowish color. Looking at the light intensity spectrum of gold, less green light and a greater proportion of blue light is absorbed than with copper (see Figure 3). This results in the reflection of the more orange-yellow complementary color and, therefore, gold is a more yellow-colored metal.

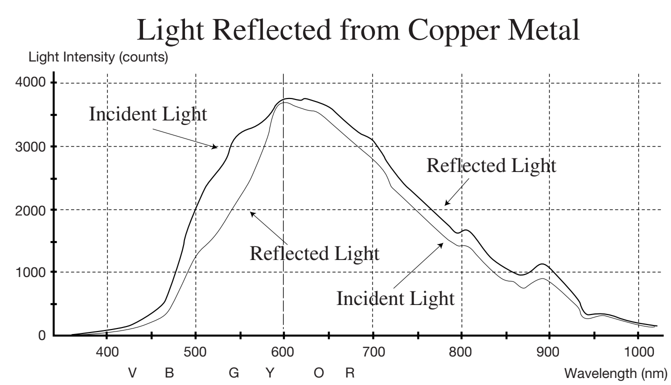


**Figure 2.**

The light intensity spectra show that copper and gold absorb some light, so it makes sense that they are colored. The question now is *why* do copper and gold behave this way when the other metals do not? Let's look at the electronic configurations of the three metals for comparison. Copper, silver, and gold—all members of the Group 12 (IB) family—have similar electron configurations. The electron configuration of copper is $[Ar]3d^{10}4s^1$, that of silver is $[Kr]4d^{10}5s^1$, and that of gold is $[Xe]5d^{10}6s^1$. The electron configurations of these three metals are exceptions to the Aufbau Principle, in which electrons fill orbitals of the least energy first, thus showing $[Ar]3d^9 4s^2$ for copper, $[Kr]4d^9 5s^2$ for silver, and $[Xe]5d^9 6s^2$ for gold.



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Notice that instead of the expected [noble gas]$d^9s^2$ electron configuration, the electron configuration becomes [noble gas] $d^{10}s^1$. This aberration in the electron configuration is due to the low energy difference between the *d* subshell and the next higher s subshell. In copper, the energy difference of the two subshells (*d* and *s*) is only 2.7 eV. In gold, the difference in the energy of the two subshells is only 1.9 eV. In silver, the difference is greater with a value of 4.8 eV.

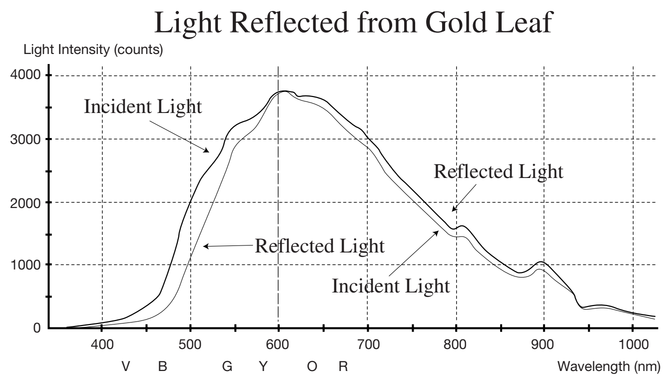
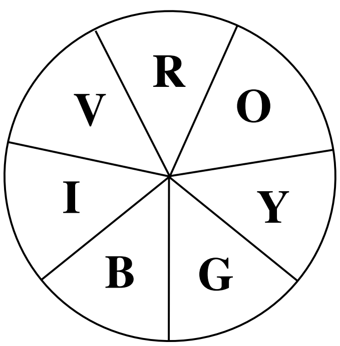



**Figure 3.**

When a ground state 3d electron in copper is promoted to the excited 4s energy level, energy is absorbed as blue-green light. As this excited electron attempts to return to the original, lower energy state, it may often lose energy in a series of smaller steps in the red-to-infrared part of the spectrum. This infrared energy is dissipated in the form of heat. In the copper spectrum (see Figure 2), notice the loss of blue-green light and an increase in the red-to-infrared light energy in the 600-nm to 1000-nm region. Gold shows a similar spectrum to copper in terms of this red-to-infrared region (see Figure 3). In silver, however, there is little difference between the spectrum of the visible incident light and the reflected light. Silver does not absorb visible light, giving it the silvery metallic luster typical of metals. Silver, like most metals, absorbs ultraviolet light instead, as explained below.




The primary reason for the difference between the spectra of copper, gold and silver is shielding. Shielding is a situation where inner electrons in an atom "shield" or "screen" the outer electrons from the nucleus. These inner d electrons in the copper-group metals repel the outer s electrons, raising their energy content. As a result, one of these s electrons drops back into the more stable d sublevel, forming the [noble gas]$d^{10}s^1$ configuration instead of the predicted [noble gas]$d^9s^2$ configuration. Silver not only has 3d electrons (as in copper) but also additional 4d electrons, which cause greater shielding and contraction and provide its 5s electron with an even greater energy. The large energy difference between these sublevels requires greater energy for its transition—an energy in the UV range. It would be expected that gold would have an even greater energy difference between its 5d and 6s electrons. However, gold has fourteen 4f electrons. These electrons cause greater shielding and contraction than in silver. These 4f electrons not only repel the 6s electrons in gold but the 5d electrons as well, lowering the energy **difference** between the two sublevels to an energy value in the blue-green portion of the visible spectrum.

*Note:* The above represents an overly simplified explanation of the actual mechanism that takes place in metals. For further information and a discussion of band theory, consult an advanced chemistry textbook.

## Acknowledgment

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