Coupled Systems

Using Exothermic Processes to Drive Endothermic Processes

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Where do we get the heat?

An endothermic reaction or process requires heat from the surroundings.

Alternatively, we can use the heat liberated by an *exothermic* process to provide the heat to drive an *endothermic* process.

These are referred to as “coupled systems”.

Heat Stoichiometry

Coupled systems problems are stoichiometry problems. “Heat” is either a reactant or a product and can be determined the same way you would determine the amount of any other reactant or product.

An Example:

Propane gas can be burned in oxygen to produce carbon dioxide gas, gaseous water, and heat. That heat can be used to warm water for a cup of tea.

How many grams of propane gas must you burn (in excess oxygen) to heat 200.0mL of water from 14.03°C to 96.72°C?

(Assume 100% efficient heat transfer from the reaction to the liquid water.)

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What’s happening?

First, identify the two parts of the coupled system and summarize what you know.

Part 1 – Burning propane. This is a chemical reaction, so we’ll need to write out a balanced equation and figure out the enthalpy change.

Part 2 – This is just heating liquid water, there’s no chemical reaction happening here. That’s a specific heat problem.

Start with the “known” part

In this problem, we have more information about the water half. We are heating 200.0mL of water from 14.03°C to 96.72°C. Look up the specific heat of water; it’s 4.184 $^{J}/_{g*°C}$. Now we can figure out how much heat is required for this part:

$$\left(\frac{4.184\,J}{g*°C}\right)\left(200.0mL*\frac{1g}{1mL}\right)(96.72°C-14.03°C)=69194.992J$$

1st half notes

A couple notes about that 69194.992J

1. That is how much heat we need to get from the other part of this coupled system.
2. What about those sig figs? IF this were the final answer, we would round to 4 sig figs (69.19kJ), BUT since this is the middle of a problem, let's keep that entire number in our calculator for the next step.

The other half…

Let's figure out how much propane to burn.
Start with a balanced chemical equation:

$$C_3H_8(g) + 5\ O_2(g) \rightarrow 3\ CO_2(g) + 4\ H_2O(g)$$

Now we need the enthalpy:

$$\Delta H_{rxn} = (+103.8\,{}^{J}/_{mol}) + 5(0\,{}^{J}/_{mol}) + 3(-393.51\,{}^{J}/_{mol}) + 4(-241.82\,{}^{J}/_{mol})$$

$$\Delta H_{rxn} = -2044.01\,{}^{J}/_{mol} = -2.04401\,{}^{kJ}/_{mol}$$

→ Again, if this were our final answer, we would evaluate sig figs and round to the tenths place (defined by the "103.8" value, but since we're in the middle of a calculation, keep all the digits for now.

Putting them together

Now we can put the halves together. We need 69.194992kJ of energy, and each mole of propane we burn gives us 2.04401kJ of energy

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$$(69.194992kJ)\left(\frac{1\ mol\ C_3H_8(g)}{2.04401kJ}\right)\left(\frac{44.194g\ C_3H_8(g)}{mol\ C_3H_8(g)}\right) = 1496.0805 \text{ grams}$$

Now we're at the "end", so we need to evaluate sig figs. The formula mass should have had 4 sig figs, the ΔH should have had 5, and the specific heat result should have had 4. These were all multiplied together, so the result should be rounded to 4 sig figs → 1496g propane, or 1.496kg propane.

The Approach

None of the individual steps of a coupled systems problem are especially complex, so the “trick” is usually just figuring out how to put them together. Work through them one bite at a time and they are a little easier to “chew”.

Signs, Signs, Everywhere're Signs

You may have noticed that I dropped some negative signs in this problem. If you want to be very mathematical about it, then you need to keep all the signs in and keep all the signs straight. Signs tell us what direction the energy is moving, but the *magnitude* of that energy stays the same. Words like "liberate", "consume", "absorb", and "release" are all ways of describing the direction the heat is moving without using an explicit sign.