Mentoring Advanced Placement
Chemistry

Topic: Exothermic and endothermic reactions

Quick Question:

Two identical thermal reservoirs are adjacent and brought into thermal contact. One is at a higher temperature than the other, and the second law would have us assume that heat flows towards the lower temperature reservoir.

Q: What affect on entropy does this transfer have for both reservoirs individually? The system as a whole? What characteristics does a change in entropy affect?

Challenge:

Physical therapists and sports enthusiasts find frequent use for hot and cold packs. Use knowledge of exothermic/endothermic reactions to design the pack of your choice.

- How much energy flow is appropriate?
- What are possible material sources for this level of energy, and what amounts will the reaction require?
- If the user wants a relatively constant amount of heat/cold for 20 minutes, how would you alter the packs design?

QQ Solution:

A: The entropy of the hot body lowers, that of the cool body rises, and the system, as a whole, experiences an increase in entropy. The "disorder" affects movement of molecules and occupied electron energy states. Molecules move faster, spinning, rotating about bonds with more degrees of freedom, stretching, and vibrating. Electrons can move to higher levels, and the general number of material positions and system orientations allowed increases. Most specifically, it's an increase in the number of allowed quantum states.

Challenge Solution and discussion:
Go through solutions put together by students. If one of the designs looks reasonable, and materials are available, give it a shot.

1.) One possible method, useing salts and the heat energy of solvation.

Enthalpy of Solvation (H_{solv} kJ/mol) of Some Common Electrolytes,
http://www.science.uwaterloo.ca/~cchieh/cact/applychem/hydration.html

Substanc	H_{solv}	Substance	H_{solv}
$AlCl_3$(s)	-373.63	H_2SO_4(l)	-95.28
$LiNO_3$(s)	-2.51	LiCl(s)	-37.03
$NaNO_3$(s)	20.50	NaCl(s)	3.88
KNO_3(s)	34.89	KCl(s)	-17.22
NaOH(s)	-44.51	NH_4Cl(s)	14.77

Example of necessary assumptions and calculations:

-Make the hot pack 0.2L; allowing good area for heat transfer, and a thermal reservoir to hold energy given off during solvation.

-44C° sounds like a toasty goal temperature.

-Assume the water starts at 22C°.

-Water has a specific heat of 4.2J/gC°, the salt in solution will not greatly affect this.

-Use sodium hydroxide, since it's easy to access and has good thermal output.

0.2L of water weighs 200g, 44C° - 22C° = 22C° = ΔT

$mC\Delta T$ = (200g)(4.2J/gC°)(22C°) = 18.5kJ

With ΔH = -44.51kJ/mol, 18.5kJ / -44.51kJ/mol = 0.42mol

NaOH = (22.99g/mol)+(16.00g/mol)+(1.01g/mol) = 40.00g/mol

(0.42mol)(40.00g/mol) = 16.8g of NaOH required

Mix the salt and water in a suitably sized plastic bag

- be watchful of quick heating and avoid burns

Possible issues to address:

- Spike and precipitous drop of water temperature
- Inconvenient size and weight of the pack
- Durability of the bag
- Method of salt dispersion

2.) A second method uses the heat given off during oxidation of metal.

This is the process behind most store bought chemical hand warmers.

The activity is from a site by the University of Illinois Urbana-Champaign, at
http://matse1.mse.uiuc.edu/home.html

The oxidation of iron to iron-oxide is exothermic. Most iron in the earth's crust exists in oxidated form, and lines called red beds, (bands of iron oxide rust), are taken to mark the geologic oxygenation of earth's atmosphere.

Converting iron oxide ore to pure iron requires a large input of thermal energy, and this energy is given off when the iron spontaneously converts back. Accelerating the reaction produces a useable amount of heat, driving the common hand warmer.

Supplies:
25 g Iron powder
1 g sodium chloride
1 Tbs. small vermiculite (sand)
plastic baggy

Procedure:
Mass 25 g of iron powder or very fine iron filings and 1 g of sodium chloride. Place these in a small plastic bag. Shake the bag to mix.
Add about a tablespoon of vermiculite to the bag and shake well.
Add 5 ml of water and seal the bag. Shake it. The reaction will start after about a minute.

Further points for discussion:
1. Store bought hand warmers come in a plastic bag with directions stating that the bag should not be removed until the user wants heat. Why should the bag not be removed?
2. In this reaction, iron metal oxidizes to iron (III) oxide, releasing heat. Plants reducing ores to metal take a large amount of energy. Where is this energy going?

Answers:
1. The bag keeps out moisture – stalling oxidation, as catalyzed by water in the air.
2. As the reverse of the hand warmer reaction, the reaction to form elemental iron from iron (III) oxide is very endothermic.